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Teacher's Edition



Chemistry - Second Edition

Teacher's Edition

Richard Parsons, (RichardP)

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CHAPTER

1

Introduction to Chemistry C-SE-TE

Chapter Outline

- 1.1 WHAT IS CHEMISTRY?
- 1.2 THE PROCESS OF SCIENCE
- 1.3 WORKSHEETS FOR CHAPTER 1
- 1.4 EXTRA READINGS FOR CHAPTER 1
- 1.5 CHAPTER 1 ASSESSMENT
- 1.6 REFERENCES

Lessons and Number of Activities for Lessons

TABLE 1.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. What is Chemistry?	0	0	0	0
2. The Process of Science	2	2	1	5

1.1 What is Chemistry?

Student Behavioral Objectives

- The student will give a brief history of how chemistry began.

Timing, Standards, Activities

TABLE 1.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>What is Chemistry?</i>	1.5	None

Activities for What is Chemistry?

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for What is Chemistry? (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

1.2 The Process of Science

Student Behavioral Objectives

The student will:

- describe the steps involved in the scientific method.
- appreciate the value of the scientific method.
- recognize that in some cases, not all the steps in the scientific method occur, or they do not occur in any specific order.
- explain the necessity for experimental controls.
- recognize the components in an experiment that represent experimental controls.

Timing, Standards, Activities

TABLE 1.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Process of Science</i>	3.0	1c, 1d, 1f, 1g, 1j, 1k, 1n

Activities for Lesson 2

Laboratory Activities

- 1. Candle Observation
- 2. Dazoo

Demonstrations

- 1. The Seven-of-diamonds psychic game
- 2. “This” or “That” psychic Game

Worksheets

- 1. Scientific Method Worksheet

Extra Readings

- 1. Science Background Information
- 2. Rene Blondolot and N-Rays
- 3. James Randi vs the Dowsers
- 4. The “Mysterious” Bermuda Triangle
- 5. Historical Comparisons

Answers for The Process of Science (L2) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

Multimedia Resources for Chapter 1

The following web site contains an article titled ‘How the Scientific Method Works’ and also other articles such as ‘History of the Scientific Method’ and more.

- <http://science.howstuffworks.com/innovation/scientific-method6.htm>

The *learner.org* website allows users to view the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The video called “The World of Chemistry” relates chemistry to other sciences and daily life.

- <http://learner.org/resources/series61.html>

These five 10 minute videos from BBC chronicle a failed search for cold fusion.

- <http://www.youtube.com/watch?v=vrs01XsmpLI>
- <http://www.youtube.com/watch?v=WQ1oghHOjHA>
- http://www.youtube.com/watch?v=4u_-h6IGb6E
- <http://www.youtube.com/watch?v=6mBu7-y2mM>
- <http://www.youtube.com/watch?v=QqiJtv5DR-4>

The *learner.org* website allows users to view the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The website has two videos that apply to this lesson. One video called “Thinking Like Scientists” relates to the scientific method. The other video is called “Modeling the Unseen.”

- <http://learner.org/resources/series61.html>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 1-2 is on the scientific method.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson12.htm>

This video is a ChemStudy film called “High Temperature Research”. The film is somewhat dated but the information is accurate.

- <http://www.youtube.com/watch?v=Tt2JEwbOtq8>

Laboratory Activities for Chapter 1

Teacher's Resource Page for Candle Observation

Safety Issues:

If students are allowed to light their own candles, they should be instructed to strike matches on the striker pad in a direction away from the body so that any flying pieces of burning match head move away from the body. Extinguished matches should be held until cool, and then placed in solid waste containers (wastebasket). Students should be reminded that during any lab involving an open flame (candles, Bunsen burners, etc.) long hair must be restrained behind the head so that it does not fall past the face when looking down. Students should be instructed not to handle candles once they are lit. Dripping hot wax can be painful.

Observation List

The candle is cylindrical in shape¹ and has a diameter² of about 2 *cm*. The length³ of the candle was initially about 18 *cm* and changed slowly during observation⁴, decreasing about 4 *mm* in 20 *minutes*.

The candle is made of a translucent,⁵ white⁶ solid⁷ which has a slight odor⁸ and no taste.⁹ It is soft enough to be scratched with a fingernail.¹⁰ There is a wick which extends from the top to bottom of the candle along its central axis¹¹ and protrudes¹² about 10 *mm* above the top of the candle. The wick is made of three strands¹³ of string braided¹⁴ together.

The candle is lit by holding a source of flame close to the wick for a few seconds.¹⁵ Thereafter, the source of the flame can be removed and the flame sustains itself¹⁶ at the wick.

The burning candle makes no sound.¹⁷ While burning, the body of the candle remains cool to the touch¹⁸ except near the top.¹⁹ Within about 5 *mm* from the top of the candle, it is warm²⁰ but not hot, and sufficiently soft to mold²¹ easily.

The flame flickers²² in response to air currents and tends to become smoky²³ while flickering. In the absence of air currents, the flame is in the form shown in the picture²⁴ although it retains some movement²⁵ at all times.

The flame begins²⁶ about 4 *mm* above the top of the candle, and at its base, the flame has a blue tint.²⁷ Immediately around the wick in a region about 5 *mm* wide²⁸ and extending about 8 *mm* above²⁹ the top of the wick, the flame is dark.³⁰ This dark region is roughly conical³¹ in shape.

Around this dark zone and extending about 5 *mm* above the dark zone is a region which emits yellow light,³² bright³³ but not blinding. The flame has rather sharply defined sides³⁴ but a ragged³⁵ top.

The wick is white³⁶ where it emerges from the candle, but from the base of the flame to the end of wick, it is black,³⁷ appearing burnt except for the last 2 *mm* where it glows red.³⁸ The wick curls³⁹ over about 4 *mm* from its end. As the candle becomes shorter, the wick shortens⁴⁰ too, so as to extend roughly a constant distance above the top of the candle.

Heat is emitted⁴¹ by the flame, enough so that it becomes uncomfortable in a few seconds to hold ones' fingers near the flame.

The top of a quietly burning candle becomes wet⁴² with a colorless⁴³ liquid and becomes bowl-shaped.⁴⁴ Sometimes, the liquid in the bowl drains⁴⁵ down the side of the candle, cools, gradually solidifying⁴⁶ and attaching itself to the candle.⁴⁷

Under quiet conditions, a stable⁴⁸ pool of clear liquid remains in the bowl-shaped top of the candle. The liquid rises⁴⁹ slightly around the wick, wetting⁵⁰ the base of the wick as high as the base⁵¹ of the flame.

Candle Observation

Materials:



Each student or pair of students is given a candle to observe. The candles should be between $\frac{1}{2}$ and $\frac{2}{3}$ inch in diameter (so their length will significantly shorten during the period) and mounted securely in a candle holder (a jar lid will work fine). The teacher should light the candles, give appropriate instructions about hair (it burns), and other safety issues. The student should be instructed to make as many observations about the burning candle as they can in the allotted time. The teacher should encourage estimated quantitative observations.

After the observation period, the teacher can ask for observations from the class and get as many as possible on the board. Further observations can be added from the list above. It is useful for students to recognize that there are many more observations about a simple system than they may have imagined.

Teacher's Resource Page for DAZOO

TABLE 1.4: Answers for DAZOO

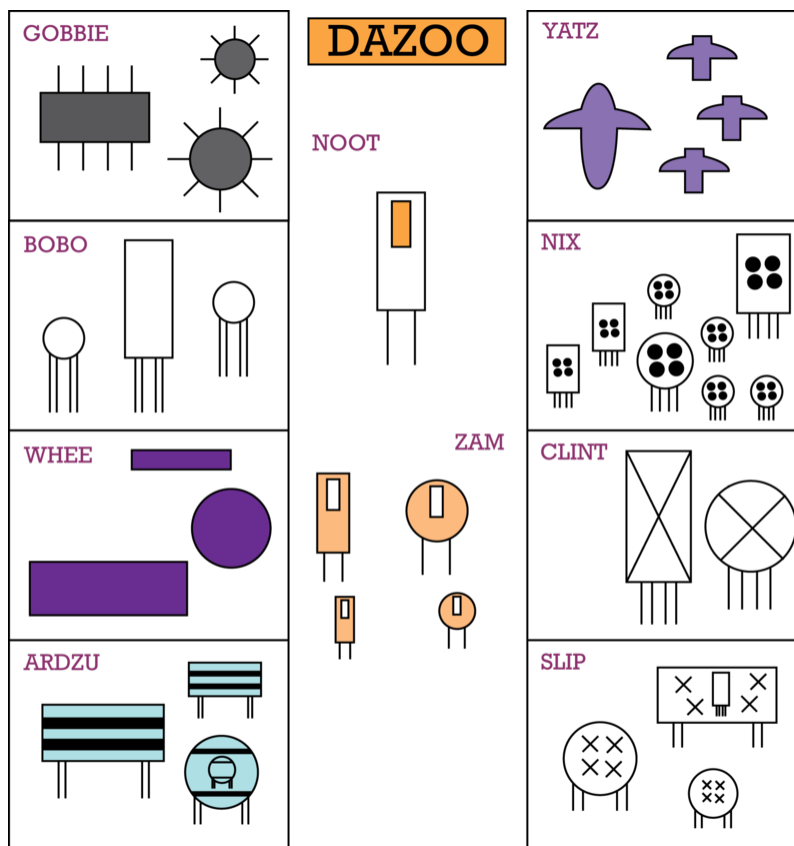
Question Number	Answer	Reason
1.	ZAM	It is a family and they are outside the cages.
2.	NOOT	It is a single male outside the cages.
3.	NIX	It is the largest number of individuals in a cage.
4.	GOBBIE	They have eight legs.
5.	BOBO	There is no large circle in the cage.
6.	YATZ	There is no large rectangle in the cage.
7.	BOBO	There are two small circles in the cage.
8.	YATZ	There are three small rectangles in the cage.
9.	CLINT	There are no small circles or rectangles in the cage.
10.	ARDZU	Circle in circle - mother is pregnant - can't be SLIP because the child is inside of a male.

TABLE 1.4: (continued)

Question Number	Answer	Reason
11.	Girl	The baby is a circle - which represents female.
12.	SLIP	There is a small BOBO rectangle inside the large SLIP rectangle.
13.	BOBO	The crocodile ate a small giraffe.
14.	YATZ	They have wings.
15.	WHEE	They have no legs.
16.	NIX	They live between YATZ and CLINT.

Observation Game: DAZOO

SET-UP: Teacher prints out a copy of the game image and a set of questions for each student. With a little thought, most students can answer the questions without teacher input.



Only one family and the zookeeper are at the zoo today. The zookeeper is a single male, but the family visiting the zoo has both males and females in the family. The family groupings of the zoo families have been given surnames and are identified by the surnames in the diagram. Try to answer the questions below.

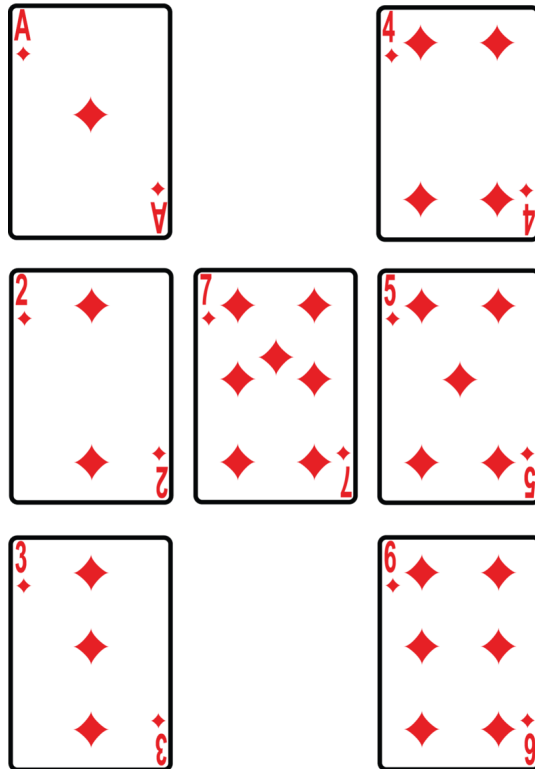
1. What is the name of the family visiting the zoo?
2. What is the name of the zookeeper?
3. Which family in the zoo has the most members?

4. What is the name of the family of spiders?
5. Which mother is away at the hospital?
6. Which family has no father?
7. Which family has two daughters?
8. Which family has three sons?
9. Mr. and Mrs. Elephant have no children. Which family are they?
10. What will be the last name of the baby tiger when it is born?
11. Will the baby tiger be a boy or a girl?
12. Mr. Crocodile has swallowed the giraffes son. Which is the family of crocodiles?
13. Which is the family of giraffes?
14. Which is the family of pelicans?
15. Which is the family of snakes?
16. The aardvarks live between the pelicans and elephants. Which is the family of aardvarks?

Demonstrations for Chapter 1

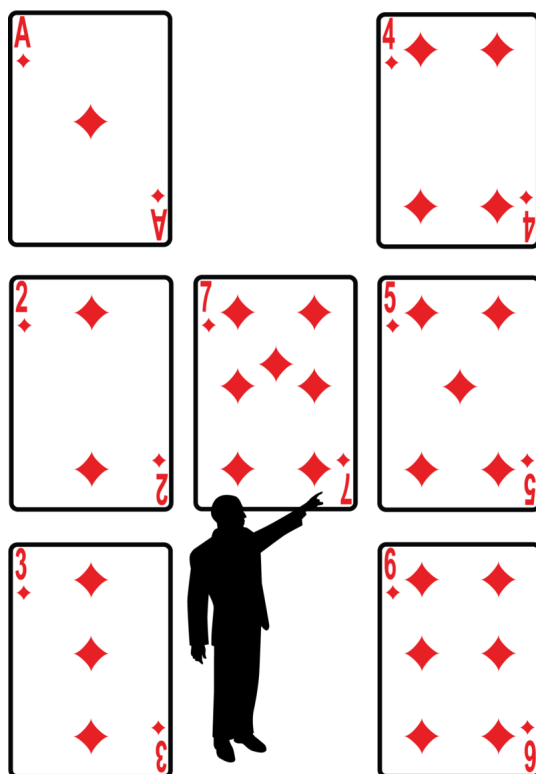
The Seven-of-Diamonds Game

This is the easiest of the observation and hypothesis games. To play the game, the teacher must select and instruct an assistant to play the role of “psychic.” The teacher draws the seven-of-diamonds set up on the board as shown below.



To begin the game, the “psychic” is sent out of the room. While the psychic is out of the room, the students select one of the cards and inform the teacher which card. Then the psychic is called back into the room and the teacher points to one of the cards and asks the psychic, *“Is this the card?”* The psychic responds either “yes” or “no” and

the process continues until the teacher points to the correct card and the psychic correctly identifies the card as the one the students had selected. The students are to observe the game and after each trial run, make hypothesis about how the trick is being done. They can do this as individuals or in groups. the game can be played over and over until at least one student or group figures out how the trick is being done. When a student or group thiks they know the trick, they can go out of the room with the psychic and then play the role of the psychic when they return. If they can correctly identify the selected card, they win the game. The game can continue until more students figure it out or the winners can explain the trick to those who didn't figure it out.



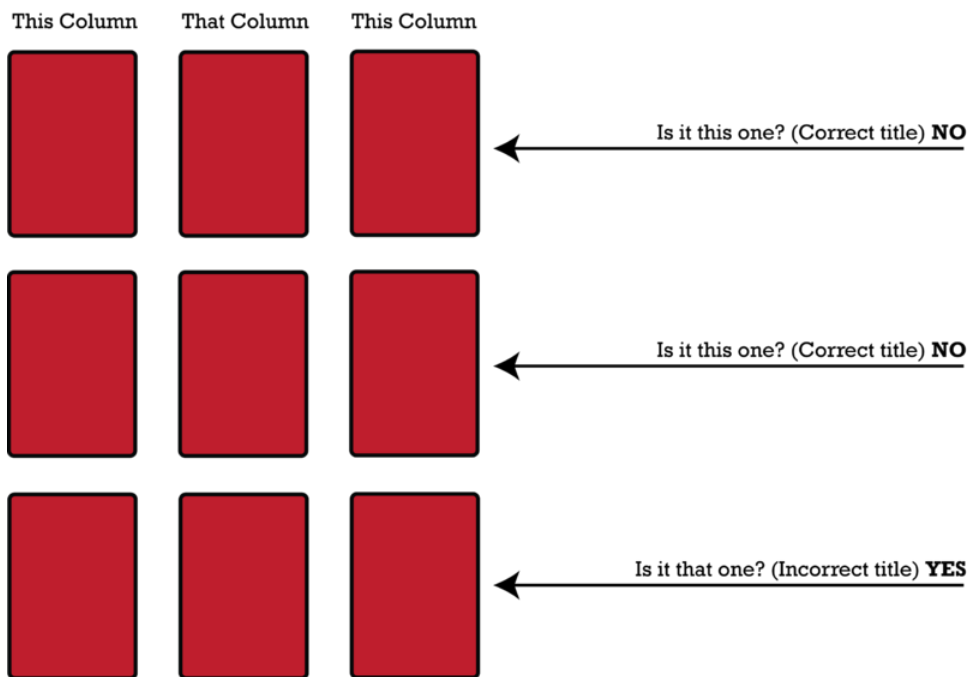
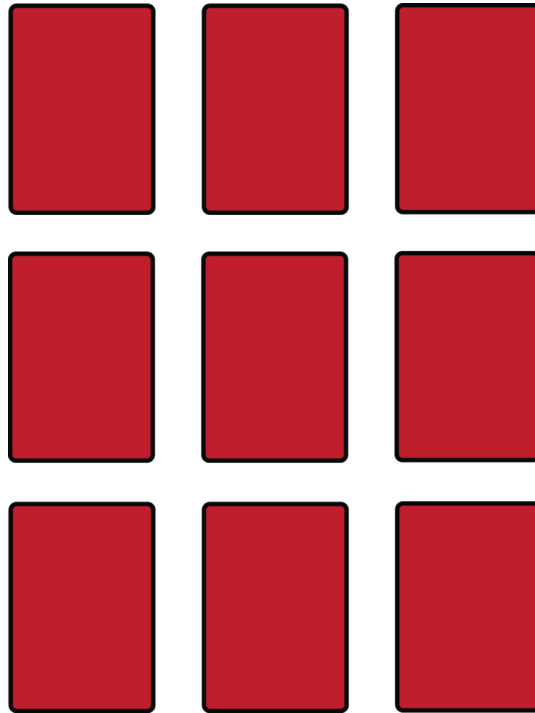
THE TRICK

The layout of the cards on the board and the spots on the seven-of-diamonds exactly correspond. The psychic will not know the correct card until the teacher points to the seven-of-diamonds and asks, “Is this the card?” When the teacher points to the seven-of-diamonds, he/she points to the spot on the card that corresponds to the card selected. In the picture above, the teacher is pointing to the spot on the seven-of-diamonds that corresponds to the position of the “6” in the layout. Therefore, the six-of-diamonds is the selected card for this trial. The psychic continues to say “no” until the teacher points to the six-of-diamonds and then says “yes.” You should note, it is not possible to have the psychic identify the correct card on the first try unless the selected card is the seven-of-diamonds. If the seven is the selected card, the teacher can point to the seven first and point to the center position. The psychic must be alert to get this one. The teacher can vary the sequence of asking so that sometimes, the selected card is pointed to on the second try or the fourth try, and so forth.

“This” or “That” Psychic Game

This is the most difficult of these observation/hypothesis games for the students to figure out. The teacher draws 3 columns of 3 squares each on the board as shown at right. Once again, the teacher needs an assistant to act as psychic. Secretly, the teacher and the “psychic” conspire and assign the two outside columns to be called “this” columns and the middle column to be called a “that” column.

As usual, the psychic leaves the room and the students select one of the squares to be “psychically” identified. The psychic is called back into the room and the teacher proceeds to point at various squares and ask the psychic, “Is it this one?”, or “Is it that one?” The code known only to the teacher and the psychic is that if the teacher uses the correct name of the column when inquiring about a square, the psychic answers “no.” If the teacher uses the incorrect name of the column when inquiring, the psychic replies “yes.”



One of the things that make this game so difficult is that the teacher can ask about the correct square on the first try. On the very first trial, the teacher can point to a square in the middle column (the “that” column) and ask, “Is it this one?” and the psychic replies “yes.”

1.3 Worksheets for Chapter 1

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Scientific Method Worksheet

1. A scientific study showed that the depth at which algae were found in a lake varied from day to day. On clear days, the algae were found as much as 6 meters below the surface of the water but were only 1 meter below the surface on cloudy days. Which hypothesis best explains these observations?
A. Nitrogen concentration affects the growth of algae. B. Rainfall affects the growth of algae. C. Light intensity affects the growth of algae. D. Wind currents affect the growth of algae.
2. Experiments may have only two variables; the “cause” variable and the “effect” variable. The “cause” variable is known as the:
A. control. B. independent variable. C. constant. D. dependent variable.
3. A scientist plants two rows of corn for experimentation. She puts fertilizer on row 1 but does not put fertilizer on row 2. Both rows receive the same amount of water and light intensity. She checks the growth of the corn over a period of 3 months. Which of the following is a constant in this experiment?
A. plant height B. corn without fertilizer C. corn with fertilizer D. amount of water
4. An experiment testing a new asthma medication was set up with two groups. Group one was given a pill containing the new drug for asthma, while group two was given a sugar pill. The sugar pill serves as a:
A. control. B. constant. C. independent variable. D. dependent variable.
5. The first step in the scientific method involves:
A. forming a hypothesis. B. making observations. C. performing an experiment. D. predicting the result of an experiment.
6. A pattern or relationship that has been established in a large amount of experimental data is a:
A. theory. B. hypothesis. C. law.
7. Ultimately, all scientific knowledge comes from:
A. experimentation. B. observation. C. textbooks. D. creative theory formation.
8. In order to be a scientific hypothesis, the hypothesis must be:
A. proven correct. B. simple. C. testable. D. popular.
9. A scientist testing the affect of a chemical on apple yield sprays an orchard with the chemical. A second orchard does not receive the chemical. In the fall, the number of apples harvested from each orchard is counted. Which of the following is the independent variable in this experiment?
A. the number of apples B. the chemical C. the treated orchard D. the untreated orchard
10. In order for the apple tree experiment in question 10 to be valid, both orchards must:
A. receive the same amount of sunlight. B. received the same amount of water. C. have the same species of apply tree. D. all of these.

11. If a theory is challenged by new evidence, which of the following could occur?
A. the theory could be altered or replaced. B. the evidence is discarded because it doesn't agree with the theory. C. a vote is taken on whether to accept the new evidence or not.
12. Eijkman fed a group of chickens exclusively on rice whose seed coat had been removed (polished rice or white rice). The chickens all developed polyneuritis (a disease of chickens) and died. He fed another group of chickens unpolished rice (rice that still had its seed coat). Not a single one of them contracted polyneuritis. He then gathered the polishings from rice (the seed coats that had been removed) and fed the polishings to other chickens that were sick with polyneuritis. In a short time, the birds all recovered. Eijkman had accurately traced the cause of polyneuritis to a faulty diet. For the first time in history, a food deficiency disease had been produced and cured experimentally.

Which of the following is a reasonable statement of Eijkman's hypothesis?

- A. Polyneuritis is a fatal disease for chickens. B. White rice carries a virus for the disease polyneuritis.
C. Unpolished rice does not carry the polyneuritis virus. D. The rice seed coat contains a nutrient that provides protection for chickens against polyneuritis. E. None of these is a reasonable statement of Eijkman's hypothesis.

Questions 13, 14, and 15 relate to the following paragraphs.

Scientist A noticed that in a certain forest area, the only animals inhabiting the region were giraffes. He also noticed that the only food available for the animals was on fairly tall trees, and as the summer progressed, the animals ate the leaves higher and higher on the trees. The scientist suggested that these animals were originally like all other animals, but generations of animals stretching their necks to reach higher up the trees for food, caused the species to grow very long necks.

Scientist B conducted experiments and observed that stretching muscles does NOT cause bones to grow longer, nor change the DNA of animals so that longer muscles would be passed on to the next generation. Scientist B, therefore, discarded Scientist A's suggested answer as to why all the animals living in the area had long necks. Scientist B suggested instead that originally many different types of animals including giraffes had lived in the region but only the giraffes could survive when the only food was high in the trees, and so all the other species had left the area.

13. Which of the following statements is an interpretation, rather than an observation?
A. The only animals living in the area were giraffes. B. The only available food was on tall trees. C. Animals which constantly stretch their necks will grow longer necks. D. A, B, and C are all interpretations. E. A, B, and C are all observations.
14. Scientist A's hypothesis was that:
A. The only animals living in the area were giraffes. B. The only available food was on tall trees. C. Animals which constantly stretch their necks will grow longer necks. D. The animals which possess the best characteristics for living in an area, will be the predominant species. E. None of the above are reasonable statements of Scientist A's hypothesis.
15. Scientist A's hypothesis being discarded is:
A. Evidence that the scientific method doesn't always work. B. A result achieved without use of the scientific method. C: An example of what happened before the scientific method was invented. D. An example of the normal functioning of the scientific method. E. An unusual case.

Answers to Worksheets

- **The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.**

1.4 Extra Readings for Chapter 1

Science Background Information

What is Mass-Energy Equivalence?

Albert Einstein is best known for his theories of relativity. There are two parts to the theory. The first part is the special theory of relativity, which was proposed in 1905. The second is the general theory of relativity, which was proposed in 1915. Einstein's special theory of relativity describes the motion of particles moving close to the speed of light. Mass-energy equivalence is a consequence of the special theory of relativity. Mass-energy equivalence is the concept that a measured quantity of energy is equivalent to a measured quantity of mass. The formula $E = mc^2$ expresses the connection between mass and energy. Here E represents energy, m represents mass, and c represents the speed of light in a vacuum. Because the speed of light is a very large number (299,792,458 m/s) and it is squared, the equation shows that very small amounts of mass can be converted into very large amounts of energy and vice versa.

Rene Blondlot and N-Rays

In 1903, Rene Blondlot was a distinguished professor of physics at the University of Nancy, France. He was a member of the French Academy of Sciences and had won several scientific awards. He had designed and carried out a brilliant experiment to measure the speed of electricity traveling through a conductor. Other scientists duplicating Blondlot's methods found that the method worked and they got the same result as Blondlot. Thus, his results were verified.

Later in his career, while trying to polarize x-rays, Blondlot claimed to have discovered a new invisible radiation similar to x-rays which he called N-rays (after the town of Nancy). Blondlot claimed that N-rays were emitted by all substances except wood and he detected them with an instrument he designed using an aluminum prism to scatter the rays and a fluorescent thread to detect the rays. Fourteen of Blondlot's friends (also scientists) confirmed the existence of the N-rays. Some other scientists tried Blondlot's experimental set up and agreed with his results. The French Academy of Science was preparing to award Blondlot their highest prize, the LaLande Prize.

Dr. Robert Wood, an American scientist, attempted to reproduce Blondlot's experiment in his own lab. Not only was Wood unable to obtain Blondlot's results but some of the observations reported by Blondlot seemed to Wood to be impossible. *Nature* magazine was skeptical of Blondlot's result because other scientists in England and Germany were also unable to duplicate Blondlot's result. The magazine sent Dr. Wood to investigate Blondlot's discovery.

Wood visited Blondlot's lab and asked for a demonstration of the experiment. Wood looked through the eyepiece of the instrument but saw no effect of N-rays on the thread. He was told by Blondlot that his eyes weren't properly sensitized. At one point, while Blondlot and his assistant were operating the instrument, Wood secretly reached into the machine and removed the prism. Both Blondlot and his assistant, however, continued to "see" the evidence of N-rays when it was impossible for the instrument to work. Wood then tried to secretly put the prism back in place, but the assistant saw him and thought that Wood was removing the prism. The next time they ran the experiment, neither Blondlot nor the assistant could see any N-rays, even though the machine was in proper working order.

Wood published the results of his visit to Blondlot's lab and the contentions of Blondlot, his assistant, and colleagues was discredited. The French Academy of Science had already published over 100 papers about N-rays. The

Academy went ahead and awarded the LaLande prize to Blondlot but it was presented as rewarding his entire career and no mention was made of N-rays. Ten years later, all mention of N-rays had been removed from French science books and French encyclopedias.

There are other similar stories - some about scientists who made serious errors in experiments and others about scientists who faked data - if you are interested in reading about some, you could search internet stories on *polywater* or *cold fusion*.

James Randi versus the Dowzers

Dowsing is the process of using wooden twigs or metal rods to locate hidden water or pieces of metal. The dowser holds the sticks or rods in his hands and when they swing together, it indicates the presence of water or metal.



Starting Position of Rods



Position of Rods when Locating

With the rods in the starting position, the dowser walks across a search area and when he/she passes over underground water or a hidden piece of metal, the rods will swing together indicating the presence of water or metal.

James Randi, a former magician, who now spends his time debunking paranormal charlatans, has made a standing offer, originally \$10,000 but now \$1,000,000, to anyone who can pass controlled tests to prove they have paranormal or supernatural powers (this includes dowsing). Mr. Randi uses what are called double blind experiments to test the claims of dowzers. A double blind test requires that neither the dowser nor the judges know the position of the dowzers search object. Over 1,000 people have attempted to claim the prize money, none have succeeded.

Consider the case of Stanley Wojcik, who claimed to be an expert dowser who could locate hidden pieces of metal in over 90% of his trials. Mr. Wojcik supplied reference letters from individuals who supported his claims. Mr. Wojcik's dowsing rods were two coat hangers straightened out to form L-shaped pieces. His procedure was to proceed forward with the rods projecting straight out in front of him until some object was "sensed" and then the rods would swing together.

The test began with Mr. Wojcik being asked to locate a small pile of coins placed on a table in plain sight. This was done to show the judges how the dowsing rods behaved when locating the hidden object and to offer proof that there was nothing in the location to inhibit the function of the dowsing rods. It is common for dowzers who fail the test to offer excuses for the failure . . . the most common excuse is that there is something in the location that interfered with the test . . . like water pipes in the floor or something of that sort.

Mr. Wojcik walked around the room with the dowsing rods extended and when he reached the coins on the table, the dowsing rods came together. Then Mr. Wojcik was asked if the rods would still work if the coins were placed in an

envelope and Wojcik replied in the affirmative. During the second test, when the coins were placed in an envelope and placed on the same table in the same place as before, the rods again came together precisely over the envelope. In the next dozen tests, nine more envelopes identical to the first but containing small lumps of paper to match the lump caused by the coins were placed around the room. Even though the odds would indicate that the dowser would correctly find the envelope containing the coins once in ten tries, Mr. Wojcik failed to find the coins even one time. When Mr. Wojcik indicated the test was flawed because of water pipes, the test was moved to another room and Mr. Wojcik still failed every time. The dowser had scored 100% on the trials where he could see the object and 0% on the blind trials.

Tests such as this have been performed with many dowsers over the years, but Mr. Randi still has the \$1,000,000.

Checking the Data on the “Mysterious” Bermuda Triangle

The “Bermuda Triangle” is a triangular expanse of ocean between the three vertices of Bermuda, Puerto Rico, and Miami. The “mystery” of the Bermuda Triangle was set forth in a series of three books written by Charles Berlitz in the 1970’s. Since the publication of the three books, dozens of other books, articles, stories, and several TV movies about the Bermuda Triangle have appeared.

Berlitz’s books contained a collection of stories of boats, airplanes, and people mysteriously lost at sea in the Bermuda Triangle, and claimed that all the stories were true and that they offered proof that there was something strange about the Triangle. Berlitz convinced millions of people that there was some unknown force in the Triangle that caused planes, boats, and people to disappear. This unknown force has variously been attributed to a sunken flying saucer, the lost city of Atlantis, or some distortion in the earth’s magnetic field.

When an unbelievable story is claimed to be true, the best place to start checking the story is to examine the data upon which the hypothesis is based.



Berlitz identified approximately 80 incidents that he claimed occurred in the Bermuda Triangle. Subsequent authors have stated that there are thousands of such incidences but do not identify any of them. Skeptics who investigated the original 80 incidents have determined that 41 of them did not occur at all. That is, there was never an airplane or boat by the name given in the story; there was never a report made to the U.S. Navy, Coast Guard; or to any police

department; the people named cannot be located by the names given in the story; and there were no flight plans or travel plans filed at the airport or harbor of origination. It is presumed, therefore, that these are fictional incidents.

Of the total number of incidents claimed, only 39 have any evidence indicating that they actually took place. Of the 39, 10 were accidents in which a ship was found abandoned. Bermuda Triangle authors indicate that the people disappeared with no explanation. For these 10 cases, however, the crews were rescued and produced quite normal explanations of what happened and why the ship was abandoned.

The other 29 incidents are indicated on the map.

Of the original 80 incidents, 41 were fictitious, 10 turned out to be quite normal, and of the 29 remaining incidents, only 4 of them actually occurred inside the Triangle. As you can see on the map, one incident occurred in the Gulf of California, over 2,000 *miles* from the Bermuda Triangle. You can also see three incidents that occurred on the European side of the Atlantic Ocean, also over 2,000 *miles* from the Triangle. These three occurred off the coast of Ireland, off the coast of Portugal, and near the Azores Islands. It is absurd to include these events in any examination of the Bermuda Triangle.



Berlitz reported incidents in his book using language and shortage of details to make the incidents seem as mysterious as possible. More complete reports often remove the mystery.

Example report by Berlitz: “Thirty-nine persons vanished north of the Triangle on a flight to Jamaica on February 2, 1953. An SOS, which ended abruptly without explanation, was sent by the *British York* transport just before it disappeared. No trace was ever found.”

A more complete report: The flight plan of the transport was to fly from the Azores (near Portugal) to Newfoundland, Canada. After an overnight stopover, the plane was to continue on to Jamaica the following day. On the flight to Canada, the plane encountered strong winds up to 75 *miles* per hour and torrential rains in the mid-Atlantic. The

crew sent an SOS which ended abruptly and no parts of the airplane were found.

This airplane did not crash in the Bermuda Triangle, in fact, it never even flew through the Bermuda Triangle. The only connection to the Triangle was a future flight plan. When metal airplanes fall into the ocean, they sink . . . and when they sink, radio messages cease abruptly. Berlitz uses words like “vanished” and “disappeared” rather than “crashed in the ocean and sank” to make the incident seem mysterious.

Another Berlitz report: Berlitz’s description of the loss of Eastern Flight 401 indicated that while flying through the Triangle, the Eastern flight “suffered a loss by disintegration”. This description would lead us to believe that the flight was somewhere in the Triangle when suddenly pieces of the airplane began to fall off for no apparent reason.

Surviving crew member’s report: The crew members reported that while over the Florida Everglades (not in the Triangle), they turned off the autopilot and while trying to fix a navigation problem, failed to notice a loss in altitude. The plane flew into the ground and “disintegrated.” End of mystery.

Another Berlitz report: This incident involves Christopher Columbus. Berlitz reports in his book (and quotes from Columbus’ logbook) that Columbus wrote about a “fireball which circled his flagship.”

Other readers of the logbook report: Columbus wrote of “a great flame which fell into the sea.” There is no indication or implication in Columbus’ logbook that the flame circled his ship. A meteor burning through the atmosphere is a spectacular sight and somewhat rare. One that flew around in a circle would indeed be a mystery.

The primary incident that Bermuda Triangle enthusiasts would point to is an incident that occurred on December 15, 1945. According to Berlitz, five fully equipped Avenger torpedo bombers took off from Fort Lauderdale Naval Air Station on a flight into the Triangle and back. At the time the planes should have returned, the flight leader reported over the radio they were lost and confused about directions. About 45 *minutes* later, the planes vanished from radar screens. A rescue plane sent to find them also disappeared. No trace was found of either flight. Berlitz also reported some strange radio transmissions by the pilots and flight leader.

The official Naval report of the incident, however, again, is a somewhat different story. The flight was a training flight for new pilots. Only the flight leader’s compass was turned on. During the flight, the flight leader’s compass failed and this failure was reported by radio. The failure was discovered after the planes were dangerously low on fuel. The pilots were lost and flew around in confusion until out of fuel and then fell into the sea. The rescue plane that was sent was known to be a dangerous plane because it leaked gas fumes inside the plane. A ship in the area observed this rescue plane explode and fall into the sea. Records and witness accounts of the radio transmissions from the flight show no strange or unusual transmissions. From the official report, there appears to be no mystery. The media, however, preferred the mysterious scenario and chose not to check the facts.

The desire to see favorable results where none exist is the source of much of the “data” presented by supporters of the paranormal. There are also examples of this failing in orthodox science. To protect ourselves from such wrong-headed thinking, we must always be skeptical and when we suspect flawed procedures, CHECK THE DATA AND HOW IT WAS COLLECTED.

Historical Comparisons

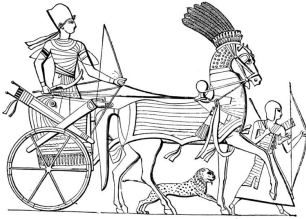
Introduction to Science

What is science? Is it a list of marvelous inventions and how they work? Or is it a list of theories about matter and energy and biological systems? Or is science a subject that you learn by carrying out activities in a laboratory? Science is all of these, but it is also something even more basic. Science is a method of thinking that allows us to discover how the world around us works.

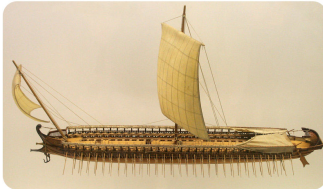
To begin this study of one form of science, we will review the last 3,000 years in the history of human transportation, communication, and medicine. The following summary lists humankind’s accomplishments in these areas during three periods in the last 3,000 years.

Transportation in 1000 B.C.

In 1000 B.C., people could transport themselves and their goods by walking, riding an animal, or by riding in a cart pulled by an animal (**Figure 1.1**). Crossing water, people could paddle a boat or have an animal walk beside the river and pull the boat (**Figure 1.2**). These methods of transportation required muscle power, either human muscles or animal muscles.

**FIGURE 1.1**

A horse-drawn Egyptian chariot. Chariots originated in Mesopotamia around 3000 B.C.

**FIGURE 1.2**

A photo of a wooden model of a Greek ship that has both sails and oars.

A few societies had designed rowboats or sailboats, which used muscle power or the force of the wind to move the boat. These early means of transportation were very limited in terms of speed and therefore, also limited in terms of distances traveled. The sail and rowboats were used on rivers and inland seas, but were not ocean-going vessels.

Transportation in 1830

By the year 1830, people were still walking and riding in carts pulled by animals. Iron ore was moved along canals by animals pulling barges. American pioneers crossed the United States in covered wagons pulled by animals (**Figure 1.3**). Large cities had streetcars pulled by horses (**Figure 1.4**). Ocean crossing was accomplished in sailing ships. The only improvement in transportation was the addition of springs and padded seats to carts and wagons to make the ride less jolting. In the period from 1000 B.C. to 1830, a span of 2,830 years (about 100 generations of people), there were *no significant changes* in the mode of human transportation.

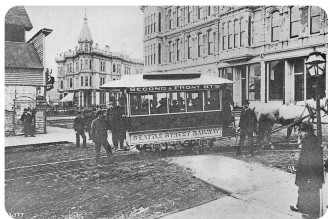
**FIGURE 1.3**

A covered wagon of the type used by pioneers to cross the US in the mid-1800s.

Transportation in 1995

By the year 1995, steam engines, gasoline engines, automobiles, propeller-driven and jet engines, locomotives, nuclear-powered ships, and inter-planetary rocket ships were invented (**Figure 1.5**). In all industrialized countries, almost anyone could own an automobile and travel great distances in very short times.

In the mid-1800s, several months were required to travel from Missouri to California by covered wagon and the trip was made at considerable risk to the traveler's life. In 1995, an average family could travel this same distance easily

**FIGURE 1.4**

The first horse-drawn street car in Seattle, Washington in 1884.

in two days and in relative safety. An ordinary person in 1995 probably traveled a greater distance in one year than an ordinary person in 1830 did in an entire lifetime. The significant changes in the means of transportation in the 165 years between 1830 and 1995 (perhaps 5 generations) were phenomenal.

**FIGURE 1.5**

A modern jetliner.

Communication in 1000 B.C.

**FIGURE 1.6**

Shopping list chiseled on a rock.

Essentially, people's only means of communicating over large distances (more than 15 miles) in 1000 B.C. was to send hand-carried messages (**Figure 1.6**). Some societies, for short distances, had developed the use of smoke signals, light signals, or drum signals, but these methods were useless for long distances. Since the means of communicating required hand-carried messages, the speed of communication was limited by the speed of transportation. Sending messages over distances of 1,000 miles could require several weeks and even then delivery was not guaranteed.

Communication in 1830

By the year 1830, people's means of communication over large distances was still the hand-carried message. While the paper and ink used to write the message had been improved, it still had to be hand-carried. In the United States, communication between New York and San Francisco required more than a month. When a new president was elected, Californians would not know who it was for a couple of months after the election.

For a short period of time, the Pony Express was set up and could deliver a letter from St. Louis, Missouri to Sacramento, California in eleven days, which was amazing at the time (**Figure 1.7**). The means of communication in 1830 was essentially the same as in 1000 B.C.

Communication in 1995

**FIGURE 1.7**

A pony express rider, circa 1861.

**FIGURE 1.8**

A modern cell phone.

By the year 1995, the telegraph, telephone, radio, television, optical fibers, and communication satellites were invented (**Figure 1.8**). People could communicate almost anywhere in the industrialized world instantaneously. Now, when a U.S. president is elected, people around the globe know the name of the new president the instant the last vote is counted. Astronauts communicate directly between the earth and the moon. An ordinary person in an industrialized country can speak with people around the world while simultaneously watching events occur in real time globally. There have been truly extraordinary changes in people's ability to communicate in the last 165 years.

Medical Treatment in 1000 B.C.

**FIGURE 1.9**

Physician letting blood from a patient.

Medical treatment in 1000 B.C. consisted of a few natural herbs and some superstitious chants and dances. The most advanced societies used both sorcerers and herbalists for medical treatment. Some of the natural herbs helped the patient and some did not. Cleaning and bandaging wounds decreased opportunity for infection while some herbs such as sesame oil demonstrated moderate antiseptic properties. Dances, chants, incense burning, and magic spells were absolutely useless in curing illnesses. At some point in time, bloodletting was added to the physician's repertoire (**Figure 1.9**). Bloodletting was accomplished by cutting the patient and allowing the blood to drip out or

by applying leeches (which doctors often carried with them). However, bloodletting was not helpful to the patient, and in many cases, it was harmful. Bloodletting was flourishing by 500 B.C. and was carried out by both surgeons and barbers. It wasn't until around 1875 that bloodletting was established as quackery.

In those times, for an ordinary person, broken bones went unset and injuries like deep cuts or stab wounds were often fatal due to infection. Infant mortality was high and it was common for at least one child in a family die before adulthood. The death of the mother in childbirth was also quite common.

In the Middle Ages, knowledge of germs, hygiene, and contagion was non-existent. People who were seriously ill might have their disease blamed on the planets going out of line (astrology) or "bad odors," or retribution for sins, or an imbalance in body fluids. Cures could involve anything from magic spells, bleeding, sweating, and vomiting to re-balance bodily fluids. Between 1340 AD and 1348 AD, the Black Death (bubonic plague) was responsible for killing in the vicinity of half the population of Europe. The bacterium causing the disease was carried by fleas, but, of course, none of this was known by the physicians of the time. Efforts to stop the plague included burning incense to eliminate "bad odors," causing loud noises to chase the plague away (the constant ringing of bells or firing of canon), and a number of people used self-flagellation to attempt to cure the disease.

Medical Treatment in 1830

Medical treatment in 1830 remained in the form of natural herbs and bloodletting. During this time, the ability to set broken bones and to amputate limbs was also developed. Amputation saved many lives from infection and gangrene. Gangrene occurs when the blood supply to tissue is interrupted and the tissue dies. The dead tissue remains part of the body, invites infection, and causes death as the poisons from the rotting tissue are carried through the body. Once gangrene afflicted an arm or leg, the poison from the limb would eventually kill the patient. During the American Civil War (1861 - 1863), a common means of treatment for wounds in field hospitals was amputation. Along with amputation was the ability to cauterize wounds to stop bleeding.

Even though bloodletting did not help patients, it continued in use through 1830. There is a tale (which may or may not be true) that George Washington was suffering from pneumonia and his doctors removed so much blood trying to cure him that they actually caused his death.

Medical Treatment in 1995



FIGURE 1.10

Receiving a vaccination.

By 1995, medical science had discovered chemical medicines, antiseptic procedures, surgery, and probably most important of all, vaccination . . . the ability to prevent disease rather than cure it after it had been contracted (**Figure 1.10**).

Diseases that had killed and crippled hundreds of thousands of people in the past are seldom heard of today (polio, smallpox, cholera, bubonic plague, etc.). These diseases have been controlled by scientific understanding of their causes and carriers and by vaccination. Average life expectancy has nearly doubled in the last 165 years. Both infant mortality and death during childbirth rates have dropped to less than 25% of what they were in 1830.

Methods of Learning About Nature: Opinion, Authority, and Superstition

Why did humans make so little progress in the 2,800 years before 1830 and then such incredible progress in the 160 years after 1830?

Socrates (469 B.C. - 399 B.C.), Plato (427 B.C. - 347 B.C.), and Aristotle (384 B.C. - 322 B.C.) are among the most famous of the Greek philosophers. Plato was a student of Socrates and Aristotle was a student of Plato. These three were probably the greatest thinkers of their time. Aristotle's views on physical science profoundly shaped medieval scholarship and his influence extended into the Renaissance (14th century). Aristotle's opinions were the authority on nature until well into the 1300s.

Unfortunately, many of Aristotle's opinions were wrong. It is not intended here to denigrate Aristotle's intelligence; he was without doubt a brilliant man. It was simply that he was using a method for determining the nature of the physical world that is inadequate for that task. The philosopher's method was logical thinking, not making observations on the natural world. This led to many errors in Aristotle's thinking on nature. Let's consider just two of Aristotle's opinions as examples.

In Aristotle's opinion, men were bigger and stronger than women, and therefore, it was logical that men would have more teeth than women. Therefore, Aristotle concluded it was a true fact that men had more teeth than women. Apparently, it never entered his mind to actually look into the mouths of both genders and count their teeth. Had he done so, he would have found that men and women have exactly the same number of teeth.

In terms of physical science, Aristotle thought about dropping two balls of exactly the same size and shape but of different masses to see which one would strike the ground first. In his mind, it was clear that the heavier ball would fall faster than the lighter one and he concluded that this was a law of nature. Once again, he did not consider doing an experiment to see which ball fell faster. It was logical to him, and in fact, it still seems logical. If someone told you that the heavier ball would fall faster, you would have no reason to disbelieve it. In fact, it is not true and the best way to prove this is to try it. Eighteen centuries later, Galileo decided to actually get two balls of different masses, but with the same size and shape, and drop them off a building (legend says the Leaning Tower of Pisa), and actually see which one hit the ground first. When Galileo actually did the experiment, he discovered, by observation, that the two balls hit the ground at exactly the same time . . . Aristotle's opinion was, once again, wrong.

In the 16th and 17th centuries, innovative thinkers were developing a new way to discover the nature of the world around them. They were developing a method that relied upon making observations of phenomena and insisting that their explanations of the nature of the phenomena corresponded to the observations they made. In order to do this, they had to overcome the opinions of the ancient Greeks, the authority of the church, and the superstitions of ordinary people.

In the opinion of the ancient Greeks, the earth was the center of the universe and did not move, while the sun, moon, planets, and stars revolved around the Earth in orbits. The astronomer Ptolemy (around 150 A.D.) observed the positions of the planets and recognized that the positions where he observed the planets did not match up with the positions predicted by the orbits of the Greeks. Ptolemy designed new orbits that had circles within circles and complicated retrograde motion (planets moving backward in their orbits at certain times). His descriptions came closer but still could not accurately predict where the heavenly orbs would be on a given night. It wasn't until Nicolaus Copernicus (1473 - 1543) suggested a heliocentric (sun-centered) system that the positions of the planets came close to matching predictions. Copernicus was hesitant to publish his ideas - some say because he feared ridicule from his peers and others say because he feared persecution by the church - but eventually, he sent his work for publication just before his death.

The publication of Copernicus' heliocentric theory didn't seem to cause much controversy for the next 50 years until the idea was supported by a scientist named Giordano Bruno who was promptly prosecuted and burned at the stake by Cardinal Bellarmine in 1600. The most famous supporter of the Copernican system was Galileo Galilei (1564 - 1642) who had developed an improved telescope (1610) and turned it toward the sky. Galileo published a small work describing what he saw with his telescope and how his observations supported the Copernican theory. The book was banned by the church in 1616 and Galileo was instructed not to write about the subject any further. In 1632, Galileo published another work, again supporting the Copernican theory and was arrested by the church, prosecuted, and punished by house arrest for the remainder of his life.

But the method of learning by experimenting, observing, and hypothesizing had been launched and many scientists would not turn back. It should be mentioned that the supporters of the methods of opinion, authority, and superstition

did not give and have not given up today. We still have “scientists” claiming that unsupported opinions are “facts” and we still have people deciding the “truth” about nature by voting on it. Nor has superstition died. You may remember from your history classes that the pilgrims of Massachusetts were still drowning and hanging women accused of being witches as late as 1693. It is easy to think that the people of those times were not very smart, and nowadays, no one would think that way. However, you should be aware that a student was suspended from school in Tulsa, Oklahoma in 1999 for “casting spells” and a substitute teacher in Florida was fired in 2008 for “wizardry” after performing a magic trick for his students.

1.5 Chapter 1 Assessment

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

Chapter 1 Quiz

1.6 References

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6. CK-12, Richard Parsons. [Shopping list chiseled on a rock..](#) CC-BY-SA
7. US National Archives and Records Administration. [A pony express rider, circa 1861..](#) Public Domain
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CHAPTER **2** Measurement in Chemistry C-SE-TE

Chapter Outline

- 2.1 MAKING OBSERVATIONS**
- 2.2 MEASUREMENT SYSTEMS**
- 2.3 THE SI SYSTEM OF MEASUREMENT**
- 2.4 SIGNIFICANT FIGURES**
- 2.5 USING ALGEBRA IN CHEMISTRY**
- 2.6 SCIENTIFIC NOTATION**
- 2.7 EVALUATING MEASUREMENTS**
- 2.8 GRAPHING**
- 2.9 WORKSHEETS FOR CHAPTER 2**
- 2.10 EXTRA READINGS FOR CHAPTER 2**
- 2.11 CHAPTER 2 ASSESSMENT**

Lessons and Number of Activities for Lessons

TABLE 2.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Making Observations	1	0	1	0
2. Measurement Systems	0	0	0	1
3. The SI System of Measurement	1	1	0	1
4. Significant Figures	0	0	1	0
5. Using Algebra in Chemistry	1	1	1	0
6. Scientific Notation	0	0	1	0
7. Evaluating Measurements	0	0	1	0
8. Graphing	0	0	1	0

2.1 Making Observations

Student Behavioral Objectives

The student will:

- define qualitative and quantitative observations.
- distinguish between qualitative and quantitative observations.
- explain the difference between qualitative and quantitative observations.
- use quantitative observations in measurements where possible.

Timing, Standards, Activities

TABLE 2.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Making Observations</i>	1.0	1a, 4e, 4f, 4g

Activities for Lesson 1

Laboratory Activities

- If the *Candle Observation* or *Dazoo* were not used in chapter 1, they could be used here.

Demonstrations

- None

Worksheets

- About Observations

Extra Readings

- None

Answers Making Observations (L1) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

2.2 Measurement Systems

Student Behavioral Objectives

The student will:

- identify the different measurement systems used in chemistry.
- state the different prefixes used in the metric system.

Timing, Standards, Activities

TABLE 2.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Measurement Systems</i>	0.5	None

Activities for Lesson 2

Laboratory Activities

- 1. None

Demonstrations

- 1. None

Worksheets

- 1. None

Extra Readings

- 1. The Metric System

Answers for Measurement Systems (L2) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

2.3 The SI System of Measurement

Student Behavioral Objectives

The student will:

- explain the difference between mass and weight.
- identify SI units of mass, distance (length), volume, temperature, and time.
- define derived unit.
- describe absolute zero.

Timing, Standards, Activities

TABLE 2.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The SI System of Measurement</i>	2.5	4e, 4f

Activities for Lesson 3

Laboratory Activities

- Thermometer Calibration

Demonstrations

- Absolute Temperature Determination

Worksheets

- None

Extra Readings

- What is Kelvin Temperature

Answers for The SI System of Measurement (L3) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

2.4 Significant Figures

Student Behavioral Objectives

The student will:

- explain the necessity for significant figures.
- determine significant figures of the equipment pieces chosen.
- identify the number of significant digits in a measurement.
- use significant figures properly in measurements and calculations.
- determine the number of significant digits in the result of a calculation.
- round calculated values to the correct number of significant digits.

Timing, Standards, Activities

TABLE 2.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Significant Figures</i>	1.0	None

Activities for Lesson 4

Laboratory Activities

- None

Demonstrations

- None

Worksheets

- Significant Figures

Extra Readings

- None

Answers for Significant Figures (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

2.5 Using Algebra in Chemistry

Student Behavioral Objectives

The student will:

- perform algebraic manipulations to solve equations.
- use the density equation to solve for the density, mass, or volume when the other two quantities in the equation are known.
- construct conversion factors from equivalent measurements.
- apply the techniques of dimensional analysis to solving problems.
- perform metric conversions using dimensional analysis and the factor label method for conversions.

Timing, Standards, Activities

TABLE 2.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Using Algebra in Chemistry</i>	2.0	None

Activities for Lesson 5

Laboratory Activities

- Density Determination

Demonstrations

- Density of Diet Soda vs Regular Soda

Worksheets

- Measurement Worksheet

Extra Readings

- None

Answers for Using Algebra in Chemistry (L5) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

2.6 Scientific Notation

Student Behavioral Objectives

The student will:

- use scientific notation in writing measurements and in calculations.

Timing, Standards, Activities

TABLE 2.7: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Scientific Notation</i>	1.0	None

Activities for Lesson 6

Laboratory Activities

- None

Demonstrations

- None

Worksheets

- Scientific Notation Worksheet

Extra Readings

- None

Answers for Scientific Notation (L6) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

2.7 Evaluating Measurements

Student Behavioral Objectives

The student will:

- define accuracy and precision.
- explain the difference between accuracy and precision.
- indicate whether a given data set is precise, accurate, both, or neither.
- calculate percent error in an experiment.

Timing, Standards, Activities

TABLE 2.8: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Evaluating Measurements</i>	1.0	None

Activities for Lesson 7

Laboratory Activities

- None

Demonstrations

- None

Worksheets

- Mass vs Weight Worksheet

Extra Readings

- None

Answers for Evaluating Measurements (L7) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

2.8 Graphing

Student Behavioral Objectives

The student will:

- correctly graph data utilizing dependent variable, independent variable, scale and units of a graph, and best fit curve.
- recognize patterns in data from a graph.
- solve for the slope of given line graphs.

Timing, Standards, Activities

TABLE 2.9: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Graphing</i>	1.0	None

Activities for Lesson 8

Laboratory Activities

- None

Demonstrations

- None

Worksheets

- Graphing Worksheet

Extra Readings

- None

Answers for Graphing (L8) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 2

This website will help to build observation skills.

- <http://www.mrsoshouse.com/pbl/observe/indexobserve.htm>

Introduction to the Metric System text.

- <http://videos.howstuffworks.com/hsw/5890-scientific-method-the-metric-system.htm>

The following website provides more information about the metric system and measurements in chemistry.

- http://www.chemistry24.com/teach_chemistry/measurement-and-math-in-chemistry.html

Metric Conversion instruction and examples.

- <http://www.tecc.tv/oldmhh/manthhomeworkhotline.com/old/metric.html>

Metric Equivalents or conversion factors.

- http://www.harcoutschool.com/activity/con_math/g03c25.html

This video is an explanation of how to convert among the Celsius, Kelvin, and Fahrenheit temperature scales and includes a sample problem.

- <http://www.youtube.com/watch?v=SASnMMGp5mo>

This video is an explanation of particle temperature, average temperature, heat flow, pressure, and volume.

- http://www.youtube.com/watch?v=tfE2y_7LqA4

Comparing and Ordering Numbers in Scientific Notation

http://www.learnalberta.ca/content/mejhm/html/video_interactive/exponents/exponentsInteractive.html

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 1-3 is on measuring matter, lesson 1-7 is on temperature conversion, lesson 2-1 is on the International System of measurements, lesson 2-3 is on significant figures, lesson 2-4 is on dimensional analysis, lesson 2-5 is on scientific notation, lesson 2-2 is on accuracy and precision.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lessindex.htm>

The *learner.org* website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The website has a video that apply to this lesson called “Measurement: The Foundation of Chemistry” that details the value of accuracy and precision.

- <http://learner.org/resources/series61.html>

Laboratory Activities for Chapter 2

Density Determination

Pre-Lab Discussion

Density is defined as the mass per unit volume of a substance. The table below lists the densities of some well known substances.

TABLE 2.10: Density of Some Common Substances

Substance	Density	Substance	Density
Water	1.0 g/cm ³	Aluminum	2.7 g/cm ³
Oxygen gas	0.0013 g/cm ³	Iron	8.9 g/cm ³
Sugar	1.6 g/cm ³	Lead	11.3 g/cm ³
Table salt	2.2 g/cm ³	Gold	19.3 g/cm ³

Density measurements allow scientists to compare the masses of equal volumes of substances. If you had a piece of lead as large as your fist and a piece of gold as large as your thumb, you would not know which substance was innately heavier because the size of the pieces are different. Determining the density of the substances would allow you to compare the masses of the same volume of each substance. The process for finding the density of a substance involves measuring the volume and the mass of a sample of the substance and then calculating density using the following formula.

$$\text{Density} = \frac{\text{mass in grams}}{\text{volume in mL}}$$

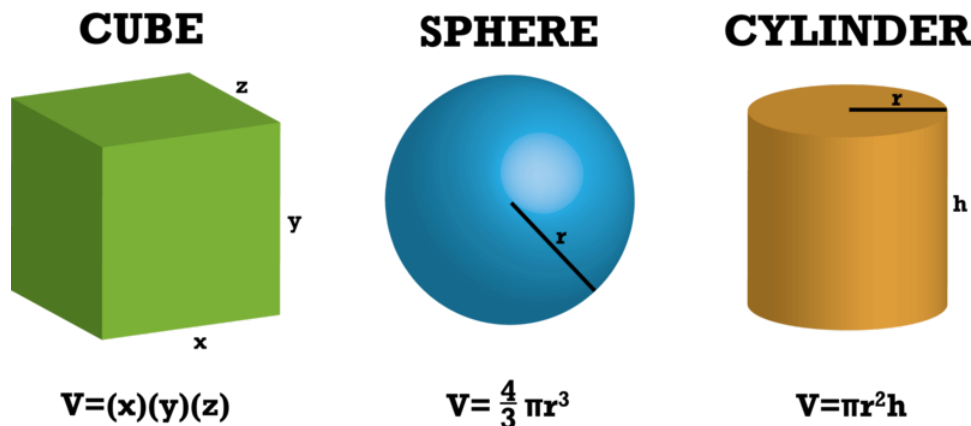
Example: Calculate the density of a piece of lead whose mass is 226 grams and whose volume is 20.0 mL. Also calculate the density of a sample of gold whose mass is 57.9 grams and whose volume is 3.00 mL.

Solution:

$$\begin{aligned} \text{Density of Lead} &= \frac{\text{mass}}{\text{volume}} = \frac{226 \text{ g}}{20.0 \text{ mL}} = 11.3 \text{ g/mL} \\ \text{Density of Gold} &= \frac{\text{mass}}{\text{volume}} = \frac{57.9 \text{ g}}{3.00 \text{ mL}} = 19.3 \text{ g/mL} \end{aligned}$$

Methods of Measuring Mass and Volume

The mass of substances is measured with a balance. In the case of a solid object that will not react with the balance pan, the object may be placed directly on the balance pan. In the case of liquids or reactive solids, the substance must be placed in a container and the container placed on the balance pan. In order to determine the mass of the substance, the mass of the container is determined before hand (empty) and then the container's mass is subtracted from the total mass to determine the mass of the substance in the container. There are several common procedures for determining the volume of a substance. The volume of a liquid is determined by pouring the liquid into a graduated cylinder and reading the bottom of the meniscus. For a regularly shaped solid, the volume can be calculated from various linear measurements.



For an irregularly shaped object, a graduated cylinder is partially filled with water and the volume measured. The object is then submerged in the water and the new volume measured. The difference between the two volumes is the volume of the submerged object.

Equipment: Specific gravity blocks, graduated cylinders (10 mL and 100 mL), thread, millimeter ruler, balance, distilled water, glycerol. (If you have an overflow can, it also works well for submersion.)

Procedure:

1. Obtain a regularly shaped object from your teacher. Measure and record its mass in grams and its dimensions in centimeters.
2. Add approximately 50 mL of tap water to a 100 – mL graduate and record its exact volume. Tie a thread to the block and carefully immerse it in the cylinder of water. Record the new volume in the cylinder.
3. Measure and record the mass of a clean, dry, 10 – mL graduated cylinder.
4. Add exactly 10.0 mL of distilled water to the cylinder. Measure and record the combined mass of the cylinder and water.
5. Repeat steps 3 and 4 with glycerol instead of water.

Data Table

Object

code name or letters _____

width _____

height _____

length _____

volume of water before block _____

volume of water after block _____

Distilled Water

volume _____

mass of empty graduate _____

combined mass _____

Glycerol

volume _____

mass of empty graduate _____

combined mass _____

Calculations

1. Find the volume of the solid object using the dimensions and appropriate formula.
2. Find the volume of the solid object using water displacement.
3. Find the density of the block using the volume from calculation 1.
4. Find the density of the block using the volume from calculation 2.
5. Find the density of the distilled water.
6. Find the density of the glycerol.
7. If your teacher gives you the accepted values for the densities in this lab, calculate the percent error for your values.

$$\% \text{ error} = \frac{(\text{experimental value}) - (\text{actual value})}{(\text{actual value})} \times 100$$

Teacher's Pages for Thermometer Calibration

Notes:

Thermometers should be stored vertically when not in use. You can stand them upright in a large beaker or in tall test tube racks. When thermometers are stored horizontally, they sometimes suffer a separation of the liquid near the top. If you have thermometers with separated liquid, you can sometimes shake them down or bounce the bulb gently on a folded towel to rejoin the liquid.

Answers to Pre-Lab Questions

1. Why is a mixture of ice and water, rather than ice alone, used in calibrating a thermometer?

You cannot be sure of the temperature of solid ice. It might be -10°C . When ice and water are both present and in equilibrium, you can be sure the temperature of the mixture is 0°C .

2. Why does the boiling point of a liquid vary with the barometric pressure?

Water boils when its vapor pressure is equal to the surrounding pressure. If the surrounding pressure is above or below normal atmospheric pressure, then the boiling point of a liquid will be above or below its normal boiling point.

3. What is the approximate boiling point of pure water at 380 Torr?

Between 80°C and 82°C .

4. What is the approximate boiling point of pure water at 800 Torr?

Near 101°C

5. Food products such as cake mixes often list special directions for cooking the products in high altitude areas. Why are special directions needed? Would a food product needing such directions require a longer or shorter time period to cook under such conditions?

At high altitudes, the atmospheric pressure is less than normal atmospheric pressure and therefore, the boiling point of water is below 100°C . Since boiling water is less than 100°C , cooking in boiling water will take longer.

TABLE 2.11: Vapor Pressure of Water at Various Temperatures

Temperature in °C	Vapor Pressure in mm of Hg	Temperature in °C	Vapor Pressure in mm of Hg
-10	2.1	52	102.1
-5	3.2	54	112.5
0	4.6	56	126.8
2	5.3	58	136.1
4	6.1	60	149.4
6	7.0	62	163.8
8	8.0	64	179.3
10	9.2	66	196.1
12	10.5	68	214.2
14	12.0	70	233.7
16	13.6	72	254.6
18	15.5	74	277.2
20	17.5	76	301.4
22	19.8	78	327.3
24	22.4	80	355.1
26	25.2	82	384.9
28	28.3	84	416.8
30	31.8	86	450.9
32	35.7	88	487.1
34	39.9	90	525.8
36	44.6	92	567.0
38	49.7	94	610.9
40	55.3	96	657.6
42	61.5	98	707.3
44	68.3	100	760.0
46	75.7	102	815.9
48	83.7	104	875.1
50	92.5	106	937.9

Thermometer Calibration

Background Information

The most common device for measuring temperature is the thermometer. The typical thermometer used in general chemistry labs has a range from -20°C to 120°C . Most laboratory thermometers are constructed of glass and therefore are very fragile. Older thermometers contain mercury as the temperature sensing liquid while newer thermometers contain a red colored fluid. The mercury thermometers are hazardous if they break because mercury vapors are poisonous over long periods of inhalation and the mercury vaporizes slowly and so when it is spilled, the lab is toxic for several months unless every drop of mercury is picked up. The red colored liquid thermometers are also hazardous if they break because the liquid is flammable and may be toxic. Great care should be exercised when handling thermometers of either kind.

The typical laboratory thermometer contains a bulb (reservoir) of temperature sensing fluid at the bottom; it is this portion of the thermometer which actually senses the temperature. The glass barrel of the thermometer above the liquid bulb contains a fine capillary opening in its center, into which the liquid rises as it expands in volume when heated. The capillary tube in the barrel is very uniform in its cross-section all along the length of the thermometer.

This insures that the fluid will rise and fall uniformly when heated or cooled.

(NOTE: laboratory thermometers look like clinical thermometers for taking people's temperatures but they are not the same. The clinical thermometer has a constriction in the tube so that after the temperature goes up and the thermometer is removed from the heat source, the liquid will not go back down. Such clinical thermometers must be shaken to lower the temperature reading before each use. Lab thermometers have no such constriction and hence the temperature reading immediately starts down when the heat source is removed. For that reason, lab thermometers must be read while the bulb is still in contact with the material whose temperature is being taken.)

Because thermometers are so fragile, it is a good idea to check them, now and then, to make sure they are still working properly. To check a thermometer, a process of calibration is used. To do this, you will determine the reading given by your thermometer in two systems whose temperature is known with certainty. If the readings of your thermometer differ by more than one degree from the true temperatures, it should be removed from use.

A mixture of ice and water which has reached equilibrium has a temperature of exactly 0°C and will be used as the first calibration point. The second calibration point will be boiling water whose exact temperature must be determined using the barometric pressure in the lab.

Pre-Lab Questions

1. Why is a mixture of ice and water, rather than ice alone, used in calibrating a thermometer?
2. Why does the boiling point of a liquid vary with the barometric pressure?
3. What is the boiling point of pure water at 380 Torr?
4. What is the boiling point of pure water at 800 Torr?
5. Food products such as cake mixes often list special directions for cooking the products in high altitude areas. Why are special directions needed? Would a food product needing such directions require a longer or shorter time period to cook under such conditions?

Purpose

In this experiment, you will check a thermometer for errors by determining the temperature of two stable equilibrium systems.

Apparatus and Materials

- Thermometer
- 400 mL beaker
- 250 mL beaker
- distilled water
- ice
- hot plate
- stirring rod
- boiling chips.

Safety Issues

Mercury thermometers are hazardous if they break because mercury vapors are poisonous over long periods of inhalation and the mercury vaporizes slowly and so when it is spilled, the lab is toxic for several months unless every drop of mercury is picked up. The red colored liquid thermometers are also hazardous if they break because the liquid is flammable and may be toxic. Great care should be exercised when handling thermometers of either kind.

Procedure

Fill a 400 mL beaker with ice and add tap water until the ice is covered with water. Stir the mixture with a stirring rod for one minute. Dip the thermometer into the ice water mixture so that the thermometer bulb is approximately centered in the mixture (not near the bottom or sides). Leave the thermometer in the mixture for two minutes and

then read the thermometer to the nearest 0.2 *degree* while the thermometer is still in the ice water bath. Record the temperature.

Allow the thermometer to return to room temperature by resting it in a safe place on the laboratory table.

Half fill a 250 mL beaker with distilled water and place it on a hot plate. Add 2 or 3 boiling chips to the water. Heat the water to boiling. Dip the thermometer into the boiling water making sure the thermometer does not get near the bottom, sides, or top of the water. Hold it there for 2 *minutes* and record the temperature reading to the nearest 0.2 *degree*.

Ask your instructor for the current barometric pressure reading in the laboratory room, look up the actual boiling point of water at this pressure and record.

Data

Actual freezing point of water = _____

Freezing point determined by your thermometer = _____

Difference between correct and trial values = _____

Barometric pressure in the room = _____

Actual boiling point of water at this pressure = _____

Boiling point determined by your thermometer = _____

Difference between correct and trial values = _____

Post-Lab Questions

1. Calculate the percent error of your measurement of the freezing point of water.

$$\% \text{ error} = \frac{\text{actual value} - \text{trial value}}{\text{actual value}} \times 100 =$$

2. Calculate the percent error of your measurement of the boiling point of water.

$$\% \text{ error} = \frac{\text{actual value} - \text{trial value}}{\text{actual value}} \times 100 =$$

TABLE 2.12: Actual Boiling Point of Water versus Various Room Pressures

Room Pressure (mm of Hg)	Boiling Point of Water (°C)
750	99.6
751	99.7
752	99.7
753	99.8
754	99.8
755	99.8
756	99.9
757	99.9
758	99.9
759	100.0
760	100.0
761	100.1
762	100.1
763	100.1

TABLE 2.12: (continued)

Room Pressure (mm of Hg)	Boiling Point of Water ($^{\circ}\text{C}$)
764	100.2
765	100.2
766	100.2
767	100.2
768	100.3
769	100.3
770	100.3

Demonstrations for Chapter 2

Density of Diet Soda vs. Regular Soda

Brief description of demonstration

A can of diet soda and regular soda are placed into a clear container of water. The diet soda floats while the regular soda sinks.

Materials

- 12 oz. can of diet soda
- 12 oz. can of regular soda, preferably the same brand
- Clear container with enough volume so the can has room to sink totally

Procedure

Fill the clear container to within 5 *cm* of the top with water. Place the diet soda into the container. It will float. Place the regular soda into the container. It will sink.

Hazards

None.

Disposal

Pour the water down the sink.

Discussion

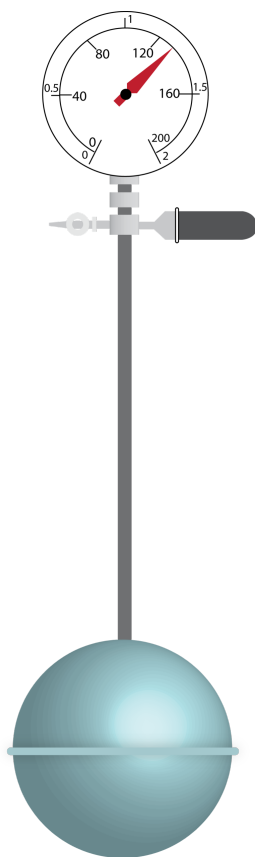
This is a good demonstration of density and for the discussion of dependent and independent variables. The only significant difference between the cans is their contents. One may want to try different sodas of both diet and regular varieties to show this.

Absolute Zero Determination Demo

Brief description of demonstration

An Absolute Zero Apparatus is placed in various liquids at different temperatures. The temperatures of each solution are known. The pressure is read from the pressure gauge on the apparatus. A graph is made with Celsius

temperature on the vertical axis and pressure on the horizontal axis. The plot is then extrapolated to zero pressure. The extrapolated line will cross the temperature axis at absolute zero.



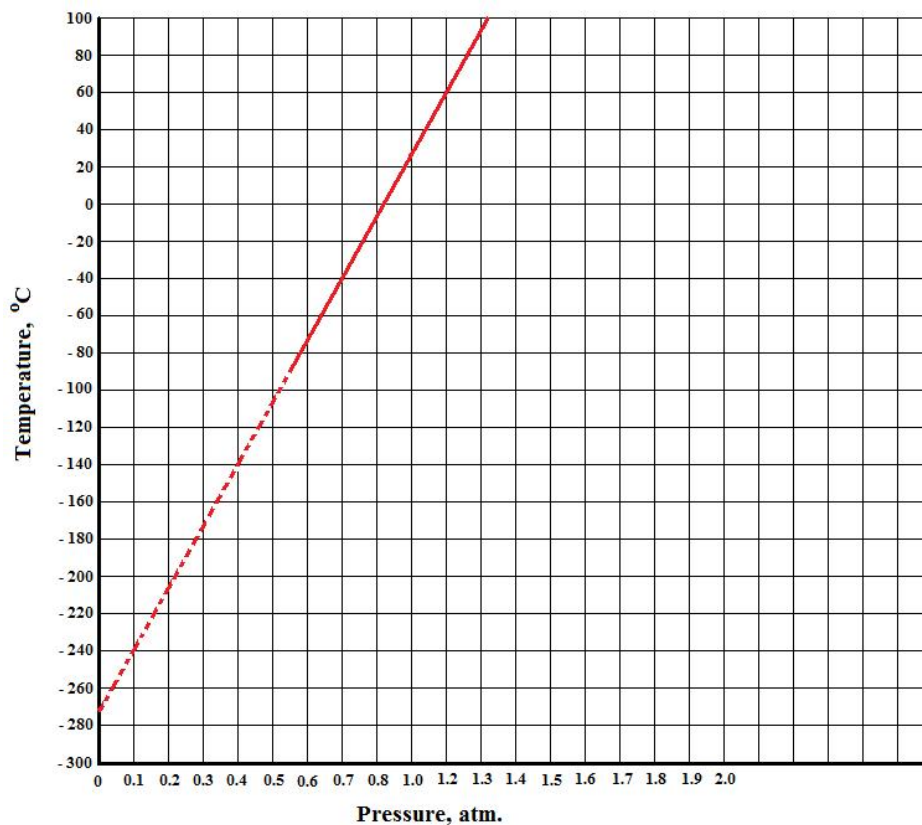
Apparatus and Materials

- Absolute zero apparatus (available from science supply companies for around \$150)
- 3 - Pyrex or Kimex 2.0 liter beakers
- Hotplate
- Ice
- Dry ice (you can find dry ice suppliers on the internet – dry ice can be stored in a Styrofoam cooler but do not put the lid on tightly)
- Ethanol – 600 mL
- If you have a mercury thermometer that covers the range -100°C to $+100^{\circ}\text{C}$, you can use it to measure the temperatures of the baths.

Procedure

1. Fill one of the Pyrex beakers half full of tap water and place it on the hotplate to boil.
2. Fill the another Pyrex beaker half full of tap water and crushed ice.
3. Fill the third beaker about one-fourth full will broken pieces of dry ice and then add ethanol slowly (lots of fog) until the beaker is about half full.
4. Place the bulb of the Absolute Zero Apparatus into the boiling water and leave it there until a constant pressure is reached. Record the temperature of the bath (taken to be 100°C) and the pressure on the gauge.
5. Place the bulb of the Absolute Zero Apparatus into the ice water and leave it there until a constant pressure is reached. Record the temperature of the bath (taken to be 0°C) and the pressure on the gauge.
6. Place the bulb of the device into the dry ice and alcohol slush and leave it there until a constant pressure is reached. Record the temperature (known to be -81°C) and the pressure on the gauge.

- Plot a graph of the temperatures and pressures recorded. Make sure that the temperature axis on your graph extends below -280°C .
- After the three points are plotted on the graph, lay a straight edge on the graph line and extend it to the zero pressure line. You will get a graph similar to the one shown below.



Hazards

Do not handle dry ice with bare hands. Pot holders or thermal protection gloves are necessary to handle the boiling water beaker and the dry ice – alcohol slush beaker.

Disposal

Once the dry ice has all melted, the solutions can be poured down the sink.

Discussion

Pressure is caused by the collisions of gas particles with each other and the walls of their container. When the temperature is lowered, the particles move more slowly, decreasing the frequency and strength of these collisions. In turn, the pressure falls.

Absolute temperature can be defined as the temperature at which molecules cease to move. Therefore absolute zero temperature corresponds to zero pressure.

Extending a graph beyond actual data points is called **extrapolation** ... a not-always acceptable procedure.

2.9 Worksheets for Chapter 2

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

About Observations Worksheet

1. Tools for making accurate measurements are essential for a scientist's work.
A. True B. False
2. A scientist can make a general conclusion after doing an experiment one time.
A. True B. False
3. If the results of an experiment do not support the hypothesis, then the experiment was a waste of time.
A. True B. False
4. Science has played a large role in improving our standard of living.
A. True B. False
5. Organized science has been a part of human existence for thousands of years.
A. True B. False
6. Scientists have already found almost all the answers to questions about nature.
A. True B. False
7. As a rule, men make better scientists than women.
A. True B. False
8. Scientists have no definite method they can follow when they set out to solve a problem.
A. True B. False

Identify each of the following observations as quantitative or qualitative.

9. The lamp is light green in color. _____
10. The liquid tastes sour. _____
11. The longest leaf is 9 *cm* long. _____
12. When you remove the cork, a loud popping sound occurs. _____
13. The mass of the computer is 1.5 *kg*. _____
14. The veins are about 3 *mm* wide. _____
15. Over time, the solution becomes darker and darker. _____

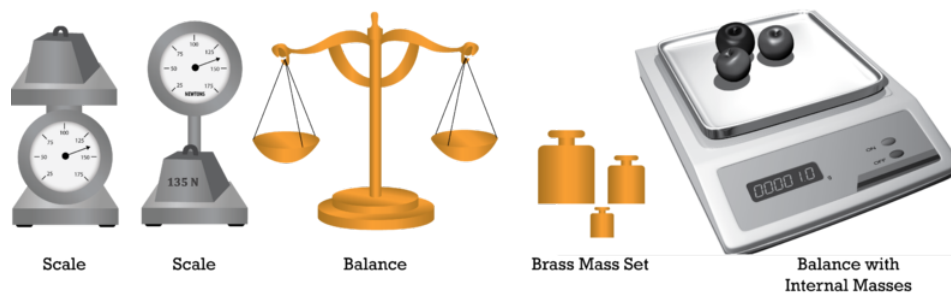
Mass Versus Weight Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

The mass of an object is a measure of the amount of matter in it. The mass (amount of matter) of an object remains the same regardless of where the object is placed. For example, moving a brick to the moon does not cause any matter in it to disappear or be removed. The weight of an object is the force of attraction between the object and the earth (or whatever large body it is resting on). We call this force of attraction, **the force of gravity**. The gravitational pull on the object varies depending on where the object is with respect to the earth or other gravity producing object. For example, a man who weighs 180 *pounds* on earth would weigh only 45 *pounds* if he were in a stationary position, 4,000 *miles* above the earth's surface. This same man would weigh only 30 *pounds* on the moon because the moon's gravity is only one-sixth that of earth. If this man were in outer space with no planet or moon nearby, **his weight would be zero**. There would be gravitational pull on him at all. The mass of this man, however, would be the same in all those situations because the amount of matter in him is constant.

We measure **weight** with a **scale**, which is a spring that compresses when a weight is placed on it. If the gravitational pull is less, the spring compresses less and the scale shows less weight. We measure **mass** with a **balance**. A balance compares the unknown mass to known masses by balancing them on a lever. If we take our balance and known masses to the moon, an object will have the same measured mass that it had on the earth. The weight, of course, would be different on the moon.



On, or near the surface of, the earth, the force of gravity is constant and so we can determine either the mass or the weight of an object if we know one of those two. On or near the surface of the earth, the conversion factor between mass and weight is: 1.00 *kg* of mass will have a weight of 9.80 *Newtons* (the standard unit of force in the SI system).

Example: What is the weight in Newtons of a 3.0 *kg* mass on the surface of the earth?

$$(\text{gravitational force} = (3.00 \text{ kg})(9.80 \text{ N/kg}) = 29.4 \text{ N})$$

Example: If an object weighs 200. *N* on the surface of the earth, what is its mass?

$$\text{mass} = (200. \text{ N}) \left(\frac{1.00 \text{ kg}}{9.80 \text{ N}} \right) = 20.4 \text{ kg}$$

Exercises

1. If an object weighs 400. *N* on the earth, how much mass does it contain?
2. What is the weight, in Newtons, of a 50 *kg* mass on the surface of the earth?

3. On the surface of the earth, how much mass is contained in a 600. N weight?
4. If an object weighs 1200 N on the earth, how much will it weigh on the moon?
5. If an object has a mass of 120 kg on the earth, what is its mass on the moon?

Measurements Worksheet

Name _____ Date _____

Measurement makes it possible to obtain more exact observations about the properties of matter such as the size, shape, mass, temperature, or composition. It allows us to make more exact quantitative observations. For example, the balance makes it possible to determine the mass of an object more accurately than we could by lifting the object and a clock gives a better measure of time than we could determine by observing the sun's position in the sky.



Measurements were originally made by comparing the object being measured to some familiar object. Length was compared to the length of one's foot. Other measures were handspans, elbow to fingertip, and so on. As people's needs increased for more consistent measurements, STANDARD systems of measurement were devised. In a standard system of measurement, some length is chosen to be the standard and copies of this object can then be used by everyone making measurements. With a standard system of measurement, two people measuring the same distance will get the same measurement.

For a time, the standard for length (one meter) was a platinum bar which was marked and stored at constant temperature in a vault. It was stored at constant temperature so that it did not expand or contract. Standard masses are also stored in airtight containers to insure no change due to oxidation. Presently, the standard *meter* is the distance light travels in a vacuum in $\frac{1}{299,792,458}$ second and the standard *second* is based on the vibrations of a cesium –133 atom.

For any system of measurements, all measurements must include a **unit** term; a word following the number that indicates the standard the measurement is based on. Systems of measurement have several standards such as length, mass, and time, and are based on physical objects such as platinum bars or vibrating atoms. Standards based on physical objects are called **undefined units**. All the other standards are expressed in terms of these object-based standards. For example, length and time are object-based standards and velocity (meters/second) and acceleration (m/s^2) are expressed in terms of length and time. Volume is expressed in terms of the length standard, volume = length \times length \times length, such as cm^3 .

There are two major systems of standards used in the United States. The one commonly used by the public (pounds, feet) and the system used for all scientific and technical work (kilograms, meters). The system used for scientific work is called the **Metric System** in its short form and is called the **International System (SI)** in its complete form. The undefined units in the SI system are the meter, gram, and second. All the sub-divisions in the SI system are in decimal form.

Conversion Factors, English to Metric

$$1.00 \text{ inch} = 2.54 \text{ centimeters}$$

$$1.00 \text{ quart} = 0.946 \text{ liter}$$

$$1.00 \text{ pound} = 4.54 \text{ Newtons} (= 454 \text{ grams on earth})$$

Units and Sub-Divisions for the SI System

Basic unit for length = meter

Basic unit for mass = gram

Basic unit for time = second

Unit for volume = liter (lee-ter)

$$1000 \text{ millimeters} = 1 \text{ meter}$$

$$100 \text{ centimeters} = 1 \text{ meter}$$

$$1000 \text{ meters} = 1 \text{ kilometer}$$

$$10 \text{ centimeters} = 1 \text{ millimeter}$$

$$1000 \text{ milligrams} = 1 \text{ gram}$$

$$1000 \text{ grams} = 1 \text{ kilogram}$$

$$1000 \text{ milliliters} = 1 \text{ liter}$$

$$1 \text{ milliliters} = 1 \text{ cubic centimeter} = 1 \text{ cm}^3$$

All the relationships between units are defined numbers and therefore, have an infinite number of significant figures. When converting units, the significant figures of the answer are based on the significant figures of the measurement, not on the conversion factors.

The unit terms for measurements are an integral part of the measurement expression and must be carried through every mathematical operation that the numbers go through. In performing mathematical operations on measurements, the unit terms as well as the numbers obey the algebraic laws of exponents and cancellation.

Examples:

TABLE 2.13: Unit Terms Follow the Rules of Algebra

Math Operations	Unit Term Operations
$6x + 2x = 8x$	$6 \text{ mL} + 2 \text{ mL} = 8 \text{ mL}$
$(5x)(3x) = 15x^2$	$(5 \text{ cm})(3 \text{ cm}) = 15 \text{ cm}^2$
$\frac{9x^3}{3x} = 3x^2$	$\frac{9 \text{ cm}^3}{3 \text{ cm}} = 3 \text{ cm}^2$
$\frac{21x}{3a} = 7\left(\frac{x}{a}\right)$	$\frac{21 \text{ grams}}{3 \text{ cm}^3} = 7 \frac{\text{grams}}{\text{cm}^3}$

Converting Units

Frequently, it is necessary to convert units measuring the same quantity from one form to another. For example, it may be necessary to convert a length measurement in meters to millimeters. This process is quite simple if you follow a standard procedure called *unit analysis*. This procedure involves creating a **conversion factor** from equivalencies between various units.

For example, we know that there are 12 *inches* in 1 *foot*. Therefore, the conversion factor between inches and feet is $12 \text{ inches} = 1 \text{ foot}$. If we have a measurement in inches and we wish to convert the measurement to feet, we would generate a conversion factor $\left(\frac{1 \text{ foot}}{12 \text{ inches}}\right)$ and multiply the measurement by this conversion factor.

Example: Convert 500. *inches* to *feet*.

$$(500. \text{ inches}) \left(\frac{1 \text{ foot}}{12 \text{ inches}} \right) = 41.7 \text{ feet}$$

We design the conversion factor specifically for this problem so that the unit term “inches” will cancel out and the final answer will have the unit “feet”. This is how we know to put the unit term “inches” in the denominator and the unit term “foot” in the numerator.

Example: Convert 6.4 *nobs* to *hics* given the conversion factor, 5 *hics* = 1 *nob*.

$$(6.4 \text{ nobs}) \left(\frac{5 \text{ hics}}{1 \text{ nob}} \right) = 32 \text{ hics}$$

Example: Convert 4.5 *whees* to *dat*s given the conversion factor, 10 *whees* = 1 *dat*.

$$(4.5 \text{ whees}) \left(\frac{1 \text{ dat}}{10 \text{ whees}} \right) = 0.45 \text{ dat}$$

Sometimes, it is necessary to insert a series of conversion factors.

Example: Convert 5.00 *wags* to *pix* given the conversion factors, 10 *wags* = 1 *hat*, and 1 *hat* = 2 *pix*.

$$(5.00 \text{ wags}) \left(\frac{1 \text{ hat}}{10 \text{ wags}} \right) \left(\frac{2 \text{ pix}}{1 \text{ hat}} \right) = 1.00 \text{ pix}$$

Solved Conversion Problems

1. Convert 1.22 *cm* to *mm*.

$$(1.22 \text{ cm}) \left(\frac{10 \text{ mm}}{1 \text{ cm}} \right) = 12.2 \text{ mm}$$

2. Convert 5.00 *inches* to *mm*.

$$(5.00 \text{ inches}) \left(\frac{2.54 \text{ cm}}{1 \text{ inch}} \right) \left(\frac{10 \text{ mm}}{1 \text{ cm}} \right) = 127 \text{ mm}$$

3. Convert 66 *lbs* to *kg*. As long as the object is at the surface of the earth, pounds (force) can be converted to grams (mass) with the conversion factor 454 *g* = 1 *lb*.

$$(66 \text{ lbs}) \left(\frac{454 \text{ g}}{1 \text{ lb}} \right) \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right) = 30. \text{ kg}$$

The mathematical answer for this conversion comes out to be 29.964 but must be rounded off to two significant figures since the original measurement has only two significant figures. When 29.964 is rounded to two significant figures, it requires a written in decimal after the zero to make the zero significant. Therefore, the final answer is 30. *kg*.

4. Convert $340. \text{ mg/cm}^3$ to lbs/ft^3 .

$$\left(\frac{340. \text{ mg}}{1 \text{ cm}^3}\right) \left(\frac{1 \text{ g}}{1000 \text{ mg}}\right) \left(\frac{1 \text{ lb}}{454 \text{ g}}\right) \left(\frac{16.39 \text{ cm}^3}{1 \text{ in}^3}\right) \left(\frac{17.28 \text{ in}^3}{1 \text{ ft}^3}\right) = 21.2 \text{ lbs/ft}^3$$

You should examine the units yourself to make sure they cancel and leave the correct units for the answer.

Exercises

- Convert 40. *cots* to togs given the conversion factor, $10 \text{ cots} = 1 \text{ tog}$.
- Convert 8.0 *cur*s to nibbles given the conversion factor, $1 \text{ cur} = 10 \text{ nibbles}$.
- Convert 100. *gags* to bobos given the conversion factor, $5 \text{ gags} = 1 \text{ bobo}$.
- Convert 1.0 *rat* to utes given the conversion factors, $10 \text{ rats} = 1 \text{ gob}$ and $10 \text{ gobs} = 1 \text{ ute}$.
- Express 3.69 *m* in *cm*.
- Express 140 *mm* in *cm*.
- Convert 15 *inches* to *mm*.
- Express 32.0 *grams* in *pounds*. (Be aware that such a conversion between weight and mass is only reasonable on the surface of the earth.)
- Express 690 *mm* in *m*.
- Convert 32.0 *lbs/qt* to *g/mL*.
- Convert 240. *mm* to *cm*.
- Convert 14,000 *mm* to *m*.

Significant Figures Worksheet

Name _____ Date _____

Working in the field of science almost always involves working with numbers. Some observations in science are qualitative and therefore, do not involve numbers, but in chemistry, most observations are quantitative and so, require numbers. You have been working with numbers for many years in your math classes thus numbers are not new to you. Unfortunately, there are some differences between the numbers you use in math and the numbers you use in science.

The numbers you use in math class are considered to be exact numbers. When you are given the number **2** in a math problem, it does not mean **1.999** rounded to **2** nor does it mean **2.000001** rounded to **2**. In math class, the number **2** means exactly **2.0000000...** with an infinite number of zeros - **a perfect 2!** Such numbers are produced only by definition, not by measurement. That is, we can define 1 *foot* to contain exactly 12 *inches*, and these two numbers are perfect numbers, but we cannot measure an object to be exactly 12 *inches* long. In the case of measurements, we can read our measuring instruments only to a limited number of subdivisions. We are limited by our ability to see smaller and smaller subdivisions, and we are limited by our ability to construct smaller and smaller subdivisions. Even using powerful microscopes to construct and read our measuring devices, we eventually reach a limit, and therefore, even though the actual measurement of an object may be a perfect number of inches, we cannot prove it to be so. **Measurements do not produce perfect numbers** and since science is greatly involved with measuring, science does not produce perfect numbers (except in defined numbers such as conversion factors).

It is very important to recognize and report the limitations of measurements along with the magnitude and unit of the measurement. Many times, the analysis of the measurements made in a science experiment is simply the search for regularity in the observations. If the numbers reported show the limits of the measurements, the regularity, or lack there of, becomes visible.

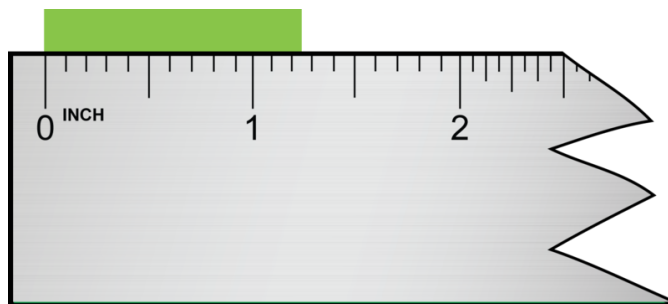
TABLE 2.14: Two Sets of Observations

Observations List A	Observations List B
22.41359 <i>m</i>	22.4 <i>m</i>
22.37899 <i>m</i>	22.4 <i>m</i>
22.42333 <i>m</i>	22.4 <i>m</i>
22.39414 <i>m</i>	22.4 <i>m</i>

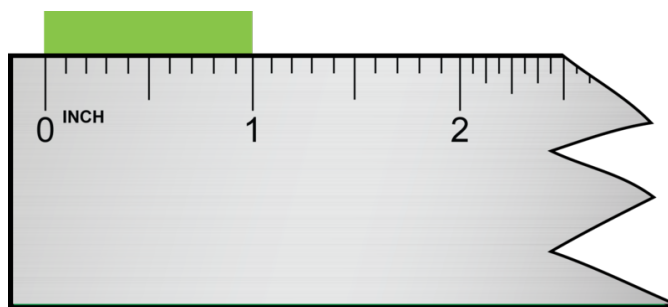
In the lists of observations above, it is difficult to perceive a regularity in List A, but when the numbers are reported showing the limits of the measurements as in List B, the regularity becomes apparent.

One of the methods used to keep track of the limit of a measurement is called **Significant Figures**. In this system, when you record a measurement, the written number must indicate the limit of the measurement, and when you perform mathematical operations on measurements, the final answer must also indicate the limit of the original measurements.

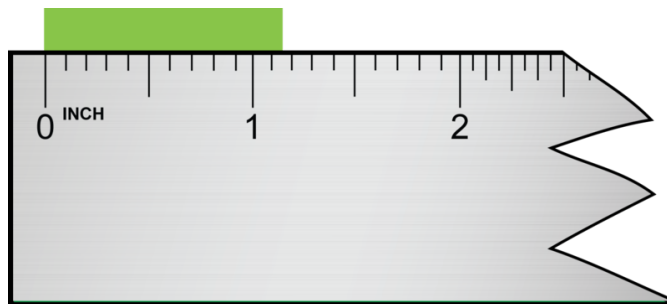
To record a measurement, you must write down all the digits actually measured, including measurements of zero and you must NOT write down any digit not measured. The only real problem that occurs with this system is that zeros are sometimes used as measured numbers and are sometimes used simply to locate the decimal point and ARE NOT measured numbers.



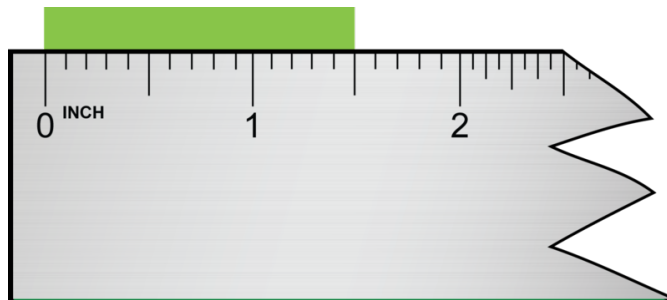
In the case shown above, the correct measurement is greater than 1.2 *inches* but less than 1.3 *inches*. **It is proper to estimate one place beyond the calibrations** of the measuring instrument. Therefore, this measurement should be reported as either 1.23, 1.24, 1.25, 1.26, or 1.27 *inches*.



In this second case, it is apparent that the object is, as nearly as we can read, exactly at 1 *inch*. Since we know the tenths place is zero and can estimate the hundredths place to be zero, the measurement should be reported as 1.00 *inch*. It is vital that you include the zeros in your measurement report because these are measured places.



This is read as 1.13, 1.14, 1.15, or 1.16 inches.



This is read 1.50 inches.

These readings indicate that the measuring instrument had subdivisions down to the tenths place and the hundredths place is estimated. There is some uncertainty about the last and only the last digit.

In our system of writing significant figures, we must distinguish between measured zeros and place-holding zeros. Here are the rules for determining the number of significant figures in a measurement.

RULES FOR DETERMINING THE NUMBER OF SIGNIFICANT FIGURES

1. All non-zero digits are significant.
2. All zeros between non-zero digits are significant.
3. All beginning zeros are NOT significant.
4. Ending zeros are significant if the decimal point is actually written in but not significant if the decimal point is an understood decimal.

Examples of the Rules

1. All non-zero digits are significant.

543 has 3 significant figures.

22.437 has 5 significant figures.

1.321754 has 7 significant figures.

2. All zeros between non-zero digits are significant.

7,004 has 4 significant figures.

10.3002 has 6 significant figures.

103.0406 has 7 significant figures.

3. All beginning zeros are NOT significant.

00013.25 has 4 significant figures.

0.0000075 has 2 significant figures.

0.000002 has 1 significant figure.

4. Ending zeros are significant if the decimal point is actually written in but not significant if the decimal point is an understood decimal.

37.300 has 5 significant figures.

33.00000 has 7 significant figures.

1.70 has 3 significant figures.

1,000,000 has 1 significant figure.

302,000 has 3 significant figures.

1,050 has 3 significant figures.

1,000,000. has 7 significant figures.

302,000. has 6 significant figures.

1,050. has 4 significant figures.

Exercises

How many significant figures are given in each of the following measurements?

1. 454 g _____
2. 2.2 lbs _____
3. 2.205 lbs _____
4. 0.3937 L _____
5. 0.0353 L _____
6. 1.00800 g _____
7. 500 g _____
8. 480 ft _____
9. 0.0350 kg _____
10. 100. cm _____
11. 1,000 m _____
12. 0.625 L _____
13. 63.4540 mm _____
14. 3,060 m _____
15. 500. g _____
16. 14.0 mL _____
17. 1030 g _____
18. 9,700 g _____
19. 125,000 m _____
20. 12,030.7210 g _____
21. 0.0000000030 cm _____
22. 0.002 m _____
23. 0.0300 cm _____
24. 1.00 L _____
25. 0.025 m/s _____
26. 0.100 kg _____
27. 0.00300 km _____
28. 303.0 g _____
29. 250 g _____
30. 1,000. m _____

Maintaining Significant Figures Through Mathematical Operations

In addition to using significant figures to report measurements, we also use them to report the results of computations made with measurements. The results of mathematical operations with measurements must include an indication of the number of significant figures in the original measurements. There are two rules for determining the number of significant figures after a mathematical operation. One rule is for addition and subtraction, and the other rule is for multiplication and division. (Most of the errors that occur in this area result from using the wrong rule, so always double check that you are using the correct rule for the mathematical operation involved.)

Significant Figure Rule for Addition and Subtraction

The answer for an addition or subtraction problem must have digits no further to the right than the shortest addend.

Example:

$$\begin{array}{r} 13.3843 \text{ cm} \\ 1.012 \text{ cm} \\ + 3.22 \text{ cm} \\ \hline 17.6163 \text{ cm} = 17.62 \text{ cm} \end{array}$$

Note that the vertical column farthest to the right has a **3** in the top number but that this column has blank spaces in the next two numbers in the column. In elementary math classes, you were taught that these blank spaces can be filled in with zeros, and in such a case, the answer would be 17.6163 cm . In science, however, these blank spaces are NOT zeros but are unknown numbers. Since they are unknown numbers, you cannot substitute any numbers into the blank spaces and you cannot claim to know, for sure, the result of adding that column. You can know the sum of adding (or subtracting) any column of numbers that contains an unknown number. Therefore, when you add these three columns of numbers, the only columns for which you are sure of the sum are the columns that have a known number in each space in the column. When you have finished adding these three numbers in the normal mathematical process, you must round off all those columns that contain an unknown number (a blank space). Therefore, the correct answer for this addition is 17.62 cm and has four significant figures.

Example:

$$\begin{array}{r} 12 \quad m \\ + 0.00045 \text{ m} \\ \hline 12.00045 \text{ m} = 12 \text{ m} \end{array}$$

In this case, the **12** has no numbers beyond the decimal and therefore, all those columns must be rounded off and we have the seemingly odd result that after adding a number to **12**, the answer is still **12**. This is a common occurrence in science and is absolutely correct.

Example:

$$\begin{array}{r} 56.8885 \text{ cm} \\ 8.30 \text{ cm} \\ + 47.0 \text{ cm} \\ \hline 112.1885 \text{ cm} = 112.2 \text{ cm} \end{array}$$

This answer must be rounded back to the tenths place because that is the last place where all the added numbers have a recorded digit.

Significant Figure Rule for Multiplication and Division

The answer for a multiplication or division operation must have the same number of significant figures as the factor with the least number of significant figures.

Example: $(3.556 \text{ cm})(2.4 \text{ cm}) = 8.5344 \text{ cm}^2 = 8.5 \text{ cm}^2$

In this case, the factor 2.4 has two significant figures and therefore, the answer must have two significant figures. The mathematical answer is rounded back to two significant figures.

Example: $(20.0 \text{ cm})(5.0000 \text{ cm}) = 100 \text{ cm}^2 = 100. \text{ cm}^2$

In this example, the factor 20.0 cm has three significant figures and therefore, the answer must have three significant figures. In order for this answer to have three significant figures, we place an actual decimal after the second zero to indicate three significant figures.

Example: $(5.444 \text{ cm})(22 \text{ cm}) = 119.768 \text{ cm}^2 = 120 \text{ cm}^2$

In this example, the factor 22 cm has two significant figures and therefore, the answer must have two significant figures. The mathematical answer is rounded back to two significant figures. In order to keep the decimal in the correct position, a non-significant zero is used.

Exercises

Add, subtract, multiply, or divide as indicated and report your answer with the proper number of significant figures.

1.

$$\begin{array}{r} 703 \text{ g} \\ 7 \text{ g} \\ + 0.66 \text{ g} \\ \hline \end{array}$$

2.

$$\begin{array}{r} 5.624 \text{ ft} \\ 0.24 \text{ ft} \\ + 16.8 \text{ ft} \\ \hline \end{array}$$

3.

$$\begin{array}{r} 34 \text{ kg} \\ - 0.2 \text{ kg} \\ \hline \end{array}$$

4.

$$\begin{array}{r} 18.7 \text{ m} \\ + 0.009 \text{ m} \\ \hline \end{array}$$

5. Add 65.23 cm, 2.666 cm, and 10 cm.

6. Multiply 2.21 cm and 0.3 cm.

7. Multiply: $(2.002 \text{ cm})(84 \text{ cm})$
8. Multiply: $(107.888 \text{ cm})(0.060 \text{ cm})$
9. Divide 72.4 cm by 0.0000082 cm .
10. Multiply 0.32 cm by 600 cm and then divide the product by 8.21 cm .

Exponential Notation Worksheet

Name _____ Date _____

Work in science frequently involves very large and very small numbers. The speed of light, for example, is $300,000,000 \text{ meters/second}$; the mass of the earth is $6,000,000,000,000,000,000,000,000 \text{ kg}$; and the mass of an electron is $0.0000000000000000000000000000009 \text{ kg}$. It is very inconvenient to write such numbers and even more inconvenient to attempt to carry out mathematical operations with them. Imagine trying to divide the mass of the earth by the mass of an electron! Scientists and mathematicians have designed an easier method for dealing with such numbers. This more convenient system is called **Exponential Notation** by mathematicians and **Scientific Notation** by scientists.

In scientific notation, very large and very small numbers are expressed as the product of a number between 1 and 10 and some power of 10. The number $9,000,000$, for example, can be written as the product of 9 times 1,000,000 and $1,000,000$ can be written as 10^6 . Therefore, $9,000,000$ can be written as 9×10^6 . In a similar manner, 0.00000004 can be written as 4 times $\frac{1}{10^8}$ or 4×10^{-8} .

This is called the "coefficient." → 6.5×10^4 ← This is called the "exponent."

TABLE 2.15: Examples

<i>Decimal Notation</i>	<i>Scientific Notation</i>
95,672	9.5672×10^4
8,340	8.34×10^3
100	1×10^2
7.21	7.21×10^0
0.014	1.4×10^{-2}
0.0000000080	8.0×10^{-9}
0.00000000000975	9.75×10^{-12}

As you can see from the examples above, to convert a number from decimal to exponential form, you count the spaces that you need to move the decimal and that number becomes the exponent of 10. If you are moving the decimal to the left, the exponent is positive, and if you are moving the decimal to the right, the exponent is negative. One and only one non-zero digit exists to the left of the decimal and ALL significant figures are maintained. The value of using exponential notation occurs when there are many non-significant zeros.

Exercises

Express the following decimal numbers in exponential form. The exponential form should have exactly one non-zero digit to the left of the decimal and you must carry all significant figures.

1. 1000

2. 150,000
3. 243
4. 9.3
5. 435,000,000,000
6. 0.0035
7. 0.012567
8. 0.0000000000100
9. 0.000000000000467
10. 0.000200
11. 186,000
12. 9,000,000,000,000
13. 105
14. 77,000
15. 502,000

Carrying Out Mathematical Operations with Exponential Numbers

When numbers in exponential notation are added or subtracted, the exponents *must* be the same. If the exponents are the same, the coefficients are added and the exponent remains the same.

Consider the following example.

$$4.3 \times 10^4 + 1.5 \times 10^4 = (4.3 + 1.5) \times 10^4 = 5.8 \times 10^4 \quad (43,000 + 15,000 = 58,000)$$

$$8.6 \times 10^7 - 5.3 \times 10^7 = (8.6 - 5.3) \times 10^7 = 3.3 \times 10^7 \quad (86,000,000 - 53,000,000 = 33,000,000)$$

If the exponents of the numbers to be added or subtracted are not the same, then one of the numbers must be changed so that the two numbers have the same exponent.

Examples

The two numbers given below, in their present form, cannot be added because they do not have the same exponent. We will change one of the numbers so that it has the same exponent as the other number. In this case, we choose to change 3.0×10^4 to 0.30×10^5 . This change is made by moving the decimal one place to the left and increasing the exponent by 1. The two numbers can now be added.

$$8.6 \times 10^5 + 3.0 \times 10^4 = 8.6 \times 10^5 + 0.30 \times 10^5 = 8.9 \times 10^5$$

We also could have chosen to alter the other number. Instead of changing the second number to a higher exponent, we could have changed the first number to a lower exponent.

$$8.6 \times 10^5 \rightarrow 86 \times 10^4$$

Now, we can add the numbers, $86 \times 10^4 + 3.0 \times 10^4 = 89 \times 10^4$

The answer, in this case, is not in proper exponential form because it has two non-zero digits to the left of the decimal. When we convert the answer to proper exponential form, it is exactly the same answer as before, $89 \times 10^4 \rightarrow 8.9 \times 10^5$.

Exercises

Add or subtract the following exponential numbers as indicated.

- $(8.34 \times 10^5) + (1.22 \times 10^5) =$
- $(4.88 \times 10^3) - (1.22 \times 10^3) =$
- $(5.6 \times 10^{-4}) + (1.2 \times 10^{-4}) =$
- $(6.38 \times 10^5) + (1.2 \times 10^4) =$
- $(8.34 \times 10^5) - (1.2 \times 10^4) =$
- $(8.34 \times 10^{-5}) + (1.2 \times 10^{-6}) =$
- $(4.93 \times 10^{-1}) - (1.2 \times 10^{-2}) =$
- $(1.66 \times 10^{-5}) + (6.4 \times 10^{-6}) =$
- $(6.34 \times 10^{15}) + (1.2 \times 10^{16}) =$
- $(6.34 \times 10^{15}) - (1.2 \times 10^1) =$

Multiplying or Dividing with Numbers in Exponential Form

When multiplying or dividing numbers in scientific notation, the numbers do not have to have the same exponents. To multiply exponential numbers, **multiply the coefficients** and **add the exponents**. To divide exponential numbers, **divide the coefficients** and **subtract the exponents**.

Examples of Multiplying Exponential Numbers

$$\text{Multiply: } (4.2 \times 10^4)(2.2 \times 10^2) = (4.2 \times 2.2)(10^{4+2}) = 9.2 \times 10^6$$

The coefficient of the answer comes out to be 9.24 but since we can only carry two significant figures in the answer, it has been rounded to 9.2.

$$\text{Multiply: } (2 \times 10^9)(4 \times 10^{14}) = (2 \times 4)(10^{9+14}) = 8 \times 10^{23}$$

$$\text{Multiply: } (2 \times 10^{-9})(4 \times 10^4) = (2 \times 4)(10^{-9+4}) = 8 \times 10^{-5}$$

$$\text{Multiply: } (2 \times 10^{-5})(4 \times 10^{-4}) = (2 \times 4)(10^{-5-4}) = 8 \times 10^{-9}$$

$$\text{Multiply: } (8.2 \times 10^{-9})(8.2 \times 10^{-4}) = (8.2 \times 8.2)(10^{(-9)+(-4)}) = 32.8 \times 10^{-13}$$

The product in the last example has too many significant figures and is not in proper exponential form, so we must round to two significant figures, 33×10^{-13} , and then move the decimal and correct the exponent, 3.3×10^{-12} .

Examples of Dividing Exponential Numbers

$$\text{Divide: } \frac{8 \times 10^7}{2 \times 10^4} = \left(\frac{8}{2}\right)(10^{7-4}) = 4 \times 10^3$$

$$\text{Divide: } \frac{8 \times 10^{-7}}{2 \times 10^{-4}} = \left(\frac{8}{2}\right)(10^{(-7)-(-4)}) = 4 \times 10^{-3}$$

$$\text{Divide: } \frac{4.6 \times 10^3}{2.3 \times 10^{-4}} = \left(\frac{4.6}{2.3}\right)(10^{(3)-(-4)}) = 2.0 \times 10^7$$

In the final example, since the original coefficients had two significant figures, the answer must have two significant figures and therefore, the zero in the tenths place is carried.

Exercises

- Multiply: $(2.0 \times 10^7)(2.0 \times 10^7) =$
- Multiply: $(5.0 \times 10^7)(4.0 \times 10^7) =$
- Multiply: $(4.0 \times 10^{-3})(1.2 \times 10^{-2}) =$
- Multiply: $(4 \times 10^{-11})(5 \times 10^2) =$
- Multiply: $(1.53 \times 10^3)(4.200 \times 10^5) =$
- Multiply: $(2 \times 10^{-13})(3.00 \times 10^{-22}) =$
- Divide: $\frac{4.0 \times 10^5}{2.0 \times 10^5} =$
- Divide: $\frac{6.2 \times 10^{15}}{2.0 \times 10^5} =$
- Divide: $\frac{8.6 \times 10^{-5}}{3.1 \times 10^3} =$
- Divide: $\frac{8.6 \times 10^{-5}}{3.1 \times 10^{-11}} =$

Graphing Worksheet

You will need a piece of graph paper for this activity.

Experimental data consists of information about the variables in an experiment from specific measurements. Graphs can be prepared from the data. A straight line or curve is drawn using the data points as a guide. The data points are not connected in a “dot-to-dot” manner. Rather, the line is drawn that “best fits” the data.

Readers of the graph may need to know the value of a variable at a point that was not measured. **Interpolation** is a method used to approximate values that are between data points on a graph. **Extrapolation** is a method for approximating values that are beyond the range of the data. Data must be extrapolated when needed values are not in the range of the measurements obtained. Unless you have a reason to believe that the graph line continues smoothly beyond the ends of the data points, extrapolation is risky . . . that is, it may produce wildly incorrect data.

The data in the following table were obtained from an experiment conducted to find out how the volume of a gas changes when its temperature is changed. Use this data to construct a graph.

Data Table: Volume vs. Temperature	
Temperature (K)	Volume (mL)
0	
100	71
140	
210	155
273	
280	195
360	257
400	
600	

Procedure

1. Put a title on the graph paper.
2. Mark the x-axis and y-axis lines on the graph paper. Use the x-axis for the independent variable (temperature) and use the y-axis for the dependent variable (volume). Write the names of the variables along the axis including the units of measure for each variable.
3. Write in the temperature measures along the temperature axis. The temperature axis should include all temperatures from 0 K to 600 K.
4. Write in the volume measures along the volume axis. The volume axis should include all volumes from 0 mL to 500 mL.
5. Plot a point for set of data where you have a measurement for both temperature and pressure.
6. Draw a line that best fits the data points.

Questions

1. Use your graph to predict values for the volume of the gas at 0 K, 140 K, 273 K, 400 K, and 600 K. Place the values of these volumes in the data table.

2. What technique did you use to estimate the volume measurements at 140 K and 273 K?
3. What technique did you use to estimate the volume measurements at 0 K and 600 K?
4. Write a sentence that describes the relationship between the temperature and the volume of a gas.

Answers to Worksheets

- **The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.**

2.10 Extra Readings for Chapter 2

The Metric System

In the late 18th century, Louis XVI of France charged a group of scientists to reform the French system of weights and measures. It was widely recognized at the time that it was an inconsistent and disorganized collection of measurements that varied with location and often on obscure bases. Providing a scientifically observable system with decimally based divisions was the charge assigned to a group from the French Academy of Sciences, which included Pierre Simon LaPlace and J.J. Lagrange. They sought to create bases of measurement linked to the scientifically verifiable values such as the Earth's circumference.

The unit of length, defined as a meter, was introduced in 1791 after careful measurement of the Earth's radius and the recognition that the planet was not perfectly spherical but instead possessed an oblate spheroid shape. The meter was designated as one ten-millionth of the length of the Earth's meridian through the city of Paris from the North Pole to the Equator.

For the measurement of volume, the SI unit devised in 1795 was the cubic meter, which was based on the volume of a cube with sides of one meter each. The large size of this unit has largely resulted in the more common use of the smaller metric unit of the liter, defined as 0.001 cubic meters.

The kilogram was settled upon in 1799 as the mass standard, based on the value of a platinum bar. Now the contemporary standard for the kilogram is stored at the Bureau International des Poids et Mesures (BIPM) in Sevres, France as a Platinum-iridium alloy.

The original definition of the principal time unit, the second was considered to be $\frac{1}{86,400}$ th of the mean solar day. Due to inconsistencies in the rate of the Earth's rotation, the modern definition is linked to the radiation correlating to the orbital transitions of the cesium -133 isotope.

Since the 1960s, the International System of Units has been internationally agreed upon as the standard metric system.

What is the Kelvin Temperature Scale?

There are three different temperature scales in use in the world today. Mainly the United States utilizes the Fahrenheit scale, which was introduced by Daniel Gabriel Fahrenheit in 1724. The non-intuitive reference points on the Fahrenheit system ($212^{\circ}F$ and $32^{\circ}F$) for the boiling and freezing points of water, respectively) are replaced in the more universally accepted Celsius, or Centigrade system, devised by Anders Celsius in 1742, by $100^{\circ}C$ and $0^{\circ}C$ for scientific applications, however, both scales are inconveniently constructed in that a substantial portion of the scale consists of negative values for temperature. For many physical considerations, the use of a Celsius or Fahrenheit temperature that is a negative number produces an impossible result, such as in the Ideal Gas Law, ($pV = nRT$).

In 1848, William Thomson Kelvin, a British physicist proposed the scale that is now named in his honor. In the design of this system, there are no negative values for temperature with the lowest value on the scale known as absolute zero. Substances at this theoretical point would display a complete absence of kinetic energy, thus atoms at absolute zero would cease all motion.

The Kelvin and Celsius scales are routinely used in chemical measurements and are conveniently constructed in

that temperature change between any two points are exactly the same. Most laboratory thermometers available today are graduated in the Celsius system yet transition to the accepted SI Kelvin units is straightforward; since $0K = -273.15^{\circ}C$, adding 273.15 *degrees* to the Celsius temperature will yield the correct Kelvin value. Note that because the Kelvin system is an absolute scale, the degree symbol ($^{\circ}$) is omitted in reporting the Kelvin temperature.

2.11 Chapter 2 Assessment

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

Chapter 2 Quiz

CHAPTER 3 Matter and Energy C-SE-TE

Chapter Outline

- 3.1 WHAT IS MATTER?
- 3.2 PROPERTIES AND CHANGES OF MATTER
- 3.3 ENERGY
- 3.4 WORKSHEETS FOR CHAPTER 3
- 3.5 EXTRA READINGS FOR CHAPTER 3
- 3.6 CHAPTER 3 ASSESSMENT

Lessons and Number of Activities for Lessons

TABLE 3.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. What is Matter?	0	0	0	1
2. Properties and Changes of Matter	1	1	1	0
3. Energy	1	0	0	1

3.1 What is Matter?

Student Behavioral Objectives

- The student will:
 - define matter and explain how it is composed of building blocks known as atoms.
 - explain the differences between substances and mixtures.
 - classify mixtures as homogeneous or heterogeneous.
 - identify the chemical symbols of common elements.
 - explain the difference between an element and a compound by their symbols or formulas.
 - demonstrate the proper use of parentheses and subscripts in writing chemical formulas.
 - determine the number of atoms and name of each element in a compound.

Timing, Standards, Activities

TABLE 3.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>What is Matter?</i>	2.0	None

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. What is Matter Worksheet

Extra Readings

1. The Scientific Method and the Socratic Method

Answers for What is Matter (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

3.2 Properties and Changes of Matter

Student Behavioral Objectives

The student will:

- explain the difference between physical and chemical properties of matter.
- list examples of physical properties.
- list examples of chemical properties.
- classify properties as chemical properties or physical properties.
- explain the difference between physical and chemical changes in matter.
- list examples of physical changes.
- list examples of chemical changes.
- classify changes as physical changes or chemical changes.

Timing, Standards, Activities

TABLE 3.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Properties and Changes of Matter</i>	1.5	None

Activities for Lesson 2

Laboratory Activities

1. Chemical and Physical Changes

Demonstrations

1. Separating Mixtures: Extracting Iron from Breakfast Cereal

Worksheets

1. Elements, Compounds, and Mixtures Worksheet

Extra Readings

1. None

Answers for Properties and Changes of Matter (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

3.3 Energy

Student Behavioral Objectives

The student will:

- explain the difference between kinetic and potential energy.
- state the law of conservation of matter and energy.
- define heat.
- define work.

Timing, Standards, Activities

TABLE 3.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Energy</i>	1	None

Activities for Lesson 3

Laboratory Activities

1. Energy Lab - Recognizing Potential Energy

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. The Nature of Energy

Answers for Energy (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 3

The following web site has a video that defines matter and energy.

- <http://videos.howstuffworks.com/hsw/20431-matter-and-energy-defining-energy-video.htm>

You may listen to Tom Lehrer's humorous song "The Elements" with animation at this website.

- <http://www.privatehand.com/flash/elements.html>

This website provides a review about matter and the categories of matter.

- <http://www.thetech.org/exhibits/online/topics/50a.html>

This website provides some free PowerPoint presentations. The presentation on "Matter and Energy" provides a review of some properties of matter, as well as provide examples of the topics covered in this lesson.

- <http://science.pppst.com/energy.html>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 1-5 is on physical and chemical properties, as well as physical and chemical changes.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson15.htm>

Summary of concepts of matter and energy and benchmark review.

- <http://broncho2.uco.edu/funeral/Bill%20Lewis/BoardReview/ChemLessons/Lesson%201.pdf>

Classroom videos about energy.

- <http://www.energyclassroom.com/>

Laboratory Activities for Chapter 3

Teacher's Pages for Chemical and Physical Changes

Lab Notes

This lab requires extensive setup. It is best to set up the equipment for the students so they can come in and not waste time. Allow two periods or a double period to complete this lab.

It is best to prepare the copper turnings for the students. To prepare a ball of turnings, cut some of the turnings off of the mass of turnings and place it in the palm of your hand. Roll the turnings between the palms of your hands until the ball becomes small enough to fit wholly within a crucible. When rolling between the palms, do not apply too much pressure. This can cause the sharp edges of the copper turnings to abrade your hands. Do NOT pull the turnings apart with your fingers. The edges of the turnings are sharp enough to cause deep cuts to you fingers and hands if placed under even moderate tension, and these cuts are painful and slow to heal.

The copper(II) nitrate is hygroscopic and is often clumped together. If needed, chip the material apart with a scoopula, and if necessary, break the larger chunks up with a mortar and pestle. Do NOT grind nitrates too vigorously, and do not mix them with an easily oxidizable substance. Store the copper(II) nitrate in an airtight re-sealable container to prevent caking.

The destructive distillation of wood produces methanol vapor, which sometimes ignites. It is a good tip to remind the students to sometimes expect the unexpected. It is a good way to keep them alert and not tip them off to an interesting observation. It may not even occur, but often does. After this experiment, the test tubes are so fouled with tar and other decomposition products that it is actually time effective to discard them. When the students are done heating the tubes have them set it aside to a place where the likelihood of it being touched is low. They stay hot for a long time.

For the *NaCl* procedure, use a watch glass. Using a glass slide almost always leads to the slide breaking due to thermal stress.

The destructive distillation of wood produces quite a bit of smoke. Do this lab in a well ventilated room, and inform the administration that this lab is being performed so they do not think that there is a wood fire in the building by the odor.

Solutions to Prepare

Per Lab Group

- 1 mL of approximately 0.10 M *NaCl* solution (0.6 g/100 mL)
- 1 mL of approximately 0.10 M *AgNO₃* solution (1.7 g/100 mL)
- 1 mL of approximately 0.10 M *FeCl₃* solution (1.6 g/100 mL)
- 1 mL of approximately 0.10 M *KSCN* solution (1 g/100 mL)
- 1 mL of approximately 0.10 M *HNO₃* (dilute 6.25 mL conc nitric acid to 1.0 liter)

Lab - Chemical and Physical Changes

Background Information

Matter is characterized in different ways using physical and chemical properties. Some physical properties are: color, odor, density, hardness, magnetism, solubility, melting point, and boiling point. Chemical properties are determined by the reaction of a substance with other substances. Examples of chemical properties are the following: reacts with acids, reacts with oxygen in air, decomposes on heating, and is acidic or basic. Chemical and physical properties must not be confused with chemical and physical changes. Physical changes refer to a transformation of a state. The chemical composition of the substance remains the same. Chemical change refers to the production of new substances that may have physical and chemical properties different from the original substance. Chemical changes are occurring when there is a drastic change in color, a gas or light or sound is produced, a mass change occurs, or a solid is formed where there was none before.

In this experiment, you will investigate the difference between chemical and physical changes in matter. You are to perform several short exercises and observe the changes that occur during each. You will then determine if your procedure resulted in chemical or physical changes.

Purpose

The purpose of this activity is to observe and document chemical and physical changes in matter.

Apparatus and Materials

- Binocular microscope
- crucible

- test tube rack
- test tube holder
- forceps
- disposable pipettes
- scoopula
- 2 test tubes ($13 \times 10 \text{ mm}$)
- laboratory burner
- 3 test tubes ($18 \times 150 \text{ mm}$)
- ring stand and ring
- clay triangle
- watch glass
- wood splint
- matches
- crucible tongs
- solid NaCl
- solid copper turnings
- solid $\text{Cu}(\text{NO}_3)_2$
- 1 mL of approximately 0.10 M NaCl solution
- 1 mL of approximately 0.10 M AgNO_3 solution
- 1 mL of approximately 0.10 M FeCl_3 solution
- 1 mL of approximately 0.10 M KSCN solution
- 1 mL of approximately 0.10 M HNO_3
- solid calcium carbide, Ca_2C
- distilled water

Safety Issues

Hot glassware; handle with forceps or crucible tongs. The solutions used can be corrosive and or poisonous. Heating wood can cause copious fumes. Perform this lab in a well ventilated room

Procedure for Part I

1. Heating a wood splint. Obtain a wood splint and break it into small pieces. Place several pieces in a large test tube. Using a test tube holder, heat the test tube strongly for several minutes. CAUTION: Hot glass. Record your observations in the data table.
2. NaCl and water. Place a spatula of sodium chloride in a watch glass. Add 3 – 5 mL of water and stir the salt and water to dissolve. Using the crucible tongs, hold the watch glass over a low flame of a Bunsen burner until the water has evaporated. After the watch glass and the solid have cooled, examine the residue with a binocular microscope. Compare the residue to a fresh sample of NaCl .
3. Heating copper. Clean and dry your crucible. To dry the crucible, heat strongly for 2 to 3 minutes and let cool. CAUTION: Hot crucibles should be handled only with tongs or forceps. While the crucible is cooling, obtain a small amount of copper turnings from the supply table. If not done already, roll them into a ball about 2 cm in diameter and place the ball in the crucible. Measure the mass of the crucible and copper to the nearest 0.01 g. Heat the metal in the crucible over a hot flame for five minutes, cool and measure the mass.
4. Heating $\text{Cu}(\text{NO}_3)_2$ CAUTION: Proper ventilation must be used. Do not inhale the fumes; avoid skin contact. Using a spatula, obtain a few crystals of copper(II) nitrate, $\text{Cu}(\text{NO}_3)_2$, and place them in a large test tube. Heat slowly over a low flame until a change is observed. Then, heat strongly until the reaction is complete. Let the mixture cool, and then add 10 drops of dilute nitric acid, HNO_3 . Heat gently.
5. Combining solutions. Obtain 3 clean small test tubes and mix the following solutions:
 - a. To 5 drops of sodium chloride, NaCl , solution add 5 drops of silver nitrate, AgNO_3 , solution. CAUTION: Silver nitrate can cause burns. Avoid skin contact.
 - b. To 5 drops of iron (III) chloride, FeCl_3 , solution add 1 to 3 drops of Potassium Thiocyanate, KSCN , solution. Observe and record your observations.

- c. To 5 drops of $FeCl_3$ solution add 1 to 2 drops of silver nitrate, $AgNO_3$, solution. CAUTION: $AgNO_3$ causes burns, avoid skin contact. Observe and record your observations.
6. Calcium carbide and water. Add 5 mL of water to a watch glass, and place a chunk about the size of a pea of calcium carbide into it. Record your observations. Light the emitted gas with a wooden splint or match.

TABLE 3.5: Data and Observations

Experiment	Observations	Chemical or Physical Change
Heating a Wooden Splint		
$NaCl$ and Water		
Heating Copper		
Heating $Cu(NO_3)_2$		
Solutions of $NaCl$ and $AgNO_3$		
Solutions of KSCN and $FeCl_3$		
Calcium Carbide and Water		

Energy Lab – Recognizing Potential Energy

Most of us intuitively know that a moving object has energy. As a result, we do not have trouble spotting examples of kinetic energy in the world around us. Recognizing potential energy, though, can be more challenging. In this lab, you will look at several examples of potential energy. In each case, you will prove that potential energy exists by *converting* the potential energy into kinetic energy. You must be careful to prove *two* points in each experiment:

1. You must prove that the potential energy is stored in the object itself. To do this you will have to be careful not to add any of your own energy to the object. You are not allowed to throw the object, or push the object, or pull the object. You are only allowed to release the object by removing your hands from the object.
2. You must prove that the object has *potential* energy by *storing* that energy for 10 *seconds*. To do this, you will hold the object still for a count of 10. Only *after* you have counted to 10 are you allowed to release the object.

Mini-lab One: Gravitational Potential Energy

When you hold an object above the Earth's surface, the object has potential energy due to its position relative to the ground below.

Materials:

A marble sized chunk of modeling clay

Method:

Using two fingers, hold the modeling clay several feet above your desk, and count to 10. By counting to 10, you are proving that the energy can be stored and thus, that it is *potential* energy. When you have finished counting to 10, release the modeling clay by removing your fingers. The modeling clay will drop to the desk. Notice how the potential energy that the modeling clay had when you held it above the ground is converted into kinetic energy as soon as you remove your fingers.

Mini-lab Two: Magnetic Potential Energy

When you hold two magnets close to each other, they have potential energy due to their relative positions

Materials:

Two fairly large bar magnets with opposite poles painted red and blue

Method:

Part One: Hold the two magnets so that the red end of one magnet is almost, (but not quite) touching the blue end of the other magnet. Count to 10. By counting to 10, you are proving that the energy can be stored and thus that it is *potential* energy. When you have finished counting to 10, release the magnets by removing your hands. The magnets will attract each other and thus, will slide together. Notice how the potential energy that the two magnets had when they were held near each other is converted into kinetic energy as soon as you remove your hands.

Part Two: Hold the two magnets so that the red end of one magnet is almost, (but not quite) touching the red end of the other magnet. Count to 10. By counting to 10, you are proving that the energy can be stored and thus, that it is *potential* energy. When you have finished counting to 10, release the magnets by removing your hands. The magnets will repel each other, and will slide apart. Notice how the potential energy, the two magnets had when they were held near each other, is converted into kinetic energy as soon as you remove your hands.

Mini-lab Three: Elastic Potential Energy

When you stretch or compress a spring, it has potential energy due to the positions of the coils relative to each other.

Materials:

A large spring

Method:

Part One: Set the spring on the table in front of you. Using your thumb and forefinger, compress the spring as much as you can. When the spring is fully compressed, count to 10. By counting to 10 you are proving that the energy can be stored and thus that it is *potential* energy. When you have counted to 10, remove your fingers from the spring. The spring will bounce back to its normal length. Notice how the potential energy in the compressed spring is converted into kinetic energy as soon as you remove your fingers.

Part Two: Set the spring on the table in front of you. Using both hands, stretch the spring as much as you can. When the spring is fully stretched, count to 10. By counting to 10, you are proving that the energy can be stored and thus, that it is *potential* energy. When you have counted to 10, remove your hands from the spring. The spring will bounce back to its normal length. Notice how the potential energy in the stretched spring is converted into kinetic energy as soon as you remove your hands.

Demonstrations for Chapter 3

Separating Mixtures: Extracting Iron from Breakfast Cereal

Brief description of demonstration

A 150 mL volume of breakfast cereal is ground in a Ziploc bag containing a magnet. The metallic iron present in the cereal adheres to the magnet.

Materials

- Breakfast cereal (Cheerios or Total works well)
- 1 quart Ziploc bag or larger
- Magnetic stir bar
- 250 mL beaker

Procedure

Measure about 150 mL of dry breakfast cereal into the 250 mL beaker. Add the cereal to the bag and add the magnet. A magnetic stir bar is preferred, because the iron is easily visible against the white plastic of the stir bar. Expel as

much air as you can from the Ziploc bag before sealing it. Grind the cereal and the stir bar together thoroughly with your hands, and against a desktop if needed until the cereal is pulverized. Remove the magnet from the bag. Metallic iron can be seen adhering to the magnet.

Hazards

None

Disposal

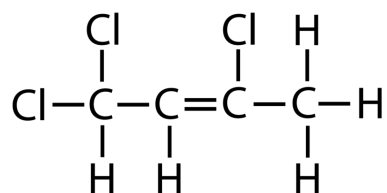
Throw the cereal and bag away in the trash.

3.4 Worksheets for Chapter 3

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Elements, Compounds, and Mixtures Worksheet

- Identify each of the following as an element, a compound, a homogeneous mixture, or a heterogeneous mixture.
 - water _____
 - pure air _____
 - salad dressing _____
 - chromium _____
 - gasoline _____
 - brass _____
 - salt water _____
 - iron _____
 - protein _____
 - mud _____
 - sand and gravel _____
 - pepsi cola _____
- Answer the following questions about one molecule of sulfuric acid, H_2SO_4 .
 - How many different elements are in the molecule? _____
 - Name the elements present. _____
 - How many atoms of oxygen are in the molecule? _____
 - What is the total number of atoms in the molecule? _____
- Identify each of the following as an element or not an element and for those that are elements, further identify them as a metal or a non-metal.
 - granite
 - carbon
 - francium
 - bronze
- The structural formula shown below is 1,1,3-trichloro-2-butene.



- What elements are present in this compound? _____
- How many hydrogen atoms are present in one molecule of this compound? _____
- What is the molecular formula for this compound? _____

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

3.5 Extra Readings for Chapter 3

The Scientific Method and the Socratic Method

The development of the scientific method was the result of centuries of cultural and societal evolution. Ranging from the philosophers of the Golden Age of Greece, through the applications of the Islamic scientists and into the ultimate flowering of the Scientific Revolution. The main premise of the scientific method is the synthesis of a hypothesis and the collection of evidence, and the persistent application of experimentation designed to support or disprove that hypothesis.

Among the first practitioners of what developed into the scientific method was Al Hazen (965 – 1039), an Islamic mathematician renowned for his extensive studies in the fields of optics, physics and psychology. In particular, Al Hazen may have been among the very first to collect experimental evidence and to assemble his observations. For example, he conducted a series of tests on observing the light of external lanterns from an inner room to lead to the conclusion that the light emanated from the lanterns, not from the long held idea that light instead was the result of particles emerging from the eyes.

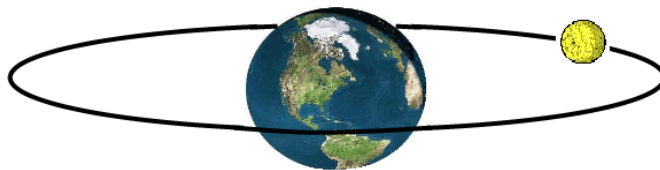
An alternative approach, called the Socratic method, consists of a method of inquiry in some ways following a parallel approach to the scientific method. The dialogues of Socrates, as collected by his student, Plato, consisted of framing a question, often about a philosophical dilemma, and addressing this issue with a logical answer. The strategy was pursued with series of questions intended to support or undermine the problem at hand. The goal of the Socratic method was to arrive at a conclusion via this sequence, mainly by uncovering any inconsistencies in their logic. This type of reasoning, utilizing only logic and the “thought experiment,” lead to early misconceptions about the nature of physical realities. At times, the lack of simple experimentation produced erroneous conclusions that remained entrenched in many cultures for many years. The eminent philosopher Aristotle wrote about objects moving with “natural motion,” that is, moving according to their composition and their speed, a result of their weight. More than a thousand years elapsed before the experiments conducted by Galileo rolling different objects down a ramp removed the role of weight in free fall acceleration.

The Nature of Energy

The Four Fundamental Forces

There are four fundamental forces within all atoms, that dictate interactions between individual particles, and the large-scale behavior of all matter throughout the universe. They are the strong nuclear force, the weak nuclear force, the electromagnetic force, and the gravitational force.

Gravity is a force of attraction that acts between each and every particle in the universe. It is always attractive, never repulsive. It pulls matter together. It is gravity that keeps the planets in their orbits around the sun, the moon in its orbit around the earth, binds galaxies together in clusters, causes apples to fall from trees, and keeps you standing on the earth.



The electromagnetic force determines the ways in which electrically charged particles interact with each other and also with magnetic fields. This force can be attractive or repulsive. Like charges (two positive or two negative charges) repel each other; unlike charges attract. The electromagnetic force binds electrons in electron clouds around the positively charged nucleus of an atom and also governs the emission and absorption of light and other forms of electromagnetic radiation. Since the outside of atoms is an electron cloud, the electromagnetic force controls the interaction of materials when they touch each other and thus is the cause of the existence of liquids and solids and allows you to talk, move, breathe, and so on. All of the interactions between objects that you see every day is controlled by the electromagnetic force.

The strong nuclear force is the force that binds the atomic nucleus together. You may not have thought about it at the time the atomic nucleus was introduced to you but the atomic nucleus contains a number of positively charged protons held tightly together in a tiny space. From what we know about the repulsive force between like charges, the atomic nucleus should not stay together. The positive protons should repel each other strongly and fly apart. The fact that the protons and neutrons stay together in an atomic nucleus is because they are held there by an extremely strong force – namely, the strong nuclear force. Both the strong and weak nuclear forces operate only when the particles being attracted are extremely close together.

At this level, the weak nuclear force will be skipped over with little consideration. We will just note that protons and neutrons are also composed of smaller particles (quarks, etc.) and these particles have a force which holds them together to form protons and neutrons. This force is the weak nuclear force.

Energy

Energy, like matter, is an important factor in our universe. Without energy, all matter – living and non-living – would be at a standstill; nothing would move, nothing would live. Energy is considered to be the “mover of matter”. The idea of energy is one that unites all the sciences. Energy does not have mass and does not take up space so it is not matter. Energy is defined as the ability to do work. An example of doing work (in the physics sense) is when you lift an object from the ground onto a table. The amount of work done depends on the force you had to apply to lift the object (its weight – or – the force of gravity on it) and the height you lifted the object. The greater the weight of the object or the higher it is lifted, the greater the amount of work done.

Energy comes in many forms. Besides mechanical energy, there is heat, light, sound, electricity, magnetism, chemical, and nuclear energy. Almost any form of energy can be converted into any other form. Our chief source of energy is the sun. It provides us with light, which can then be converted into other forms of energy. Light can be absorbed by matter and converted into heat. Light can also be absorbed by plants in the process of photosynthesis and be converted into chemical energy.

Kinetic and Potential Energy (Mechanical Energy)

Energy can be classified as either kinetic energy or potential energy. The original definition we gave for energy was the ability to do work. The ability to do work could also be stated as the “ability to make matter move”. Anything that can make a piece of matter move has energy. It should be obvious that a moving object has the ability to make another piece of matter move simply by colliding with it. Therefore, all moving objects have energy. This type of energy (the energy of moving objects) is called kinetic energy.

There are also non-moving objects that have the ability to make matter move. These objects have the ability to make matter move because of their position. For example, a rock held up in the air has the ability to make matter move – all that is required is that whatever is holding the rock up must release it. A stretched rubber band has the ability to make matter move – all that is required is that whatever is holding the rubber band must release it. This type of energy is stored energy or potential energy.

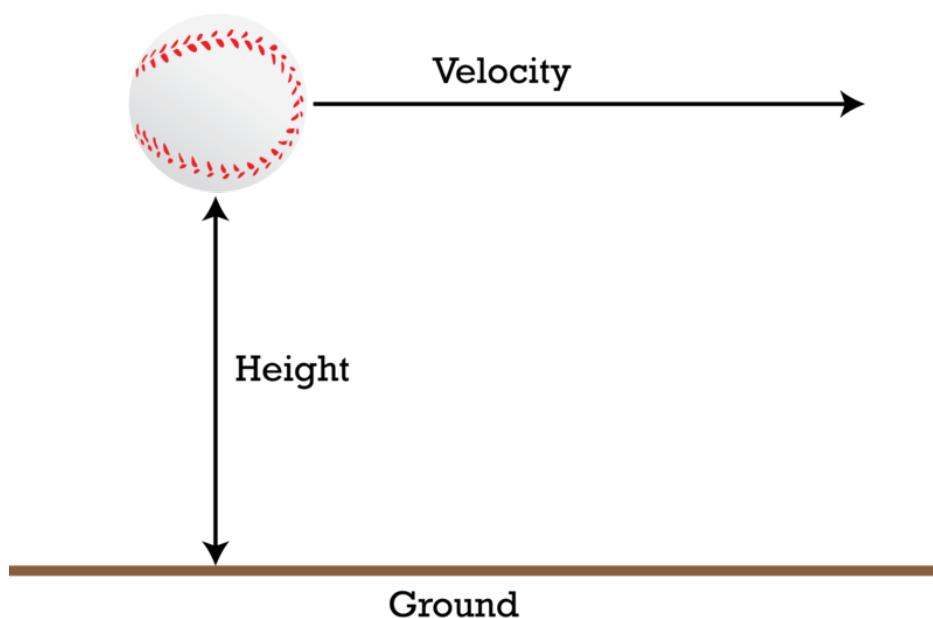
This baseball flying through the air has both kinetic and potential energy. The kinetic energy is due to its motion and the potential energy is due to the balls' height above the ground.

The kinetic energy of an object can be calculated by multiplying one-half of its mass times its velocity squared.

$$KE = \frac{1}{2}mv^2$$

The gravitational potential energy of an object can be calculated by multiplying the mass of the object times the acceleration due to gravity times the height the object can fall.

$$PE = mgh$$



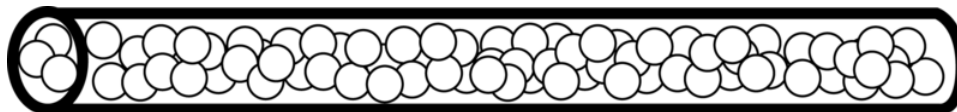
Potential energy is always present when two objects are attracted or repelled and are held in position. The most obvious case is an object that is held above the earth. The object is attracted to the earth by gravity but is kept from falling (gravitational potential energy). This same type of energy is present in bent sticks, compressed or stretched springs, stretched rubber bands, like or unlike electrical charges, and like or unlike magnetic poles. In all these cases, the potential energy can be calculated by multiplying the force of attraction (or repulsion) by the distance one object will move.

Energy Transmission

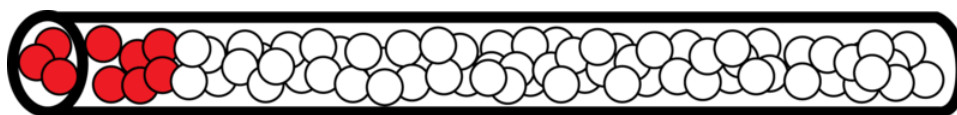
Scientists use three words to indicate the different methods by which energy moves from place to place. These three words are conduction, convection, and radiation.

Conduction

We are all familiar with the concept of molecules in constant random motion. This molecular motion increases when we heat the molecules and decreases when we cool the molecules. The energy of these moving molecules is kinetic energy. Kinetic energy is transferred between molecules when molecules at different temperatures collide with each other. When molecules at different temperatures collide with each other, energy is transferred from the "hotter" molecules to the "colder" molecules. Consider an object such as an iron bar – we can view the bar as a long chain of molecules crowded very close together. Remember that molecules or atoms in a solid are in a tightly packed pattern.



If this bar lies on a table for a few minutes, all the particles will be at about the same temperature. This is because each molecule is constantly bumping into its neighbors and these collisions transfer kinetic energy from a faster moving particle to a slower moving particle. If one end of this bar is placed into a flame and heated, the bar particles that are in the flame will get very hot.



When the hot particles bump into cold ones, the cold particles gain kinetic energy from the hot ones and thus the cold ones also become hot. Those particles then bump into their neighbors down the bar and eventually, all the molecules in the bar will be hot. This process of passing heat (kinetic energy) from particle to particle by collision is known as conduction.

Any time two objects at different temperatures touch each other, heat will be conducted from the hot one to the cold one by this process.

Molecules that make up living systems (like you) are called organic molecules. Most of these molecules are much more fragile than non-organic molecules. Non-organic molecules can usually reach quite a high temperature before the molecules are damaged. Organic molecules, however, are frequently long chains of carbon atoms and are easily to break if they are jerked around. If a hot object conducts heat to your hand, like all other conduction, the increased temperature causes the molecules of your hand to move around more rapidly. At temperatures at or below 40°C , your nervous system reacts in such a way that the sensation is not unpleasant. But at higher temperatures, your molecules begin moving around so rapidly that some of them break apart. When this happens, your nervous system sends a signal to your brain that you are in PAIN so that you will remove your hand from the heat as fast as possible. If many molecules are broken, the tissue is permanently damaged and must be replaced (healed) by the body.



Convection

Another way to move heat (energy) from one place to another is to heat up some substance, like air or water for example, and then to move the heated substance to another place. Essentially, the matter holds the energy in the form of heat, and when you move the matter to another place, you are also moving the energy it contains. In most homes, we use a furnace to heat air and use a fan to blow the hot air through ducts to various places in the house to warm it. In some places, water is heated and then pumped through pipes and into radiators to transport the heat from where it is produced to other areas. The process of moving matter that contains heat to other places is called convection.

Nature has its own convection system of heating up air in one place and then wind blows the air to another location. There are also convection currents in lakes and oceans, where cold water sinks and warm water rises causing water flow (and thus heat transfer) from place to place.



Radiation

When you build a campfire, the heat produced from the flames mostly goes upward. This is because hot air is less dense than cold air and so the hot air behaves like a helium balloon and goes straight up. If you stand a few feet to the side of a campfire, however, you will also feel heat coming from the fire. This heat does not get there by conduction or convection. This heat arrives at your position by radiation. Radiation is a type of energy transfer that can occur even through a vacuum – it needs no air or any other matter to carry it. This is quite different from conduction and convection which both require molecules of matter to transfer the energy.



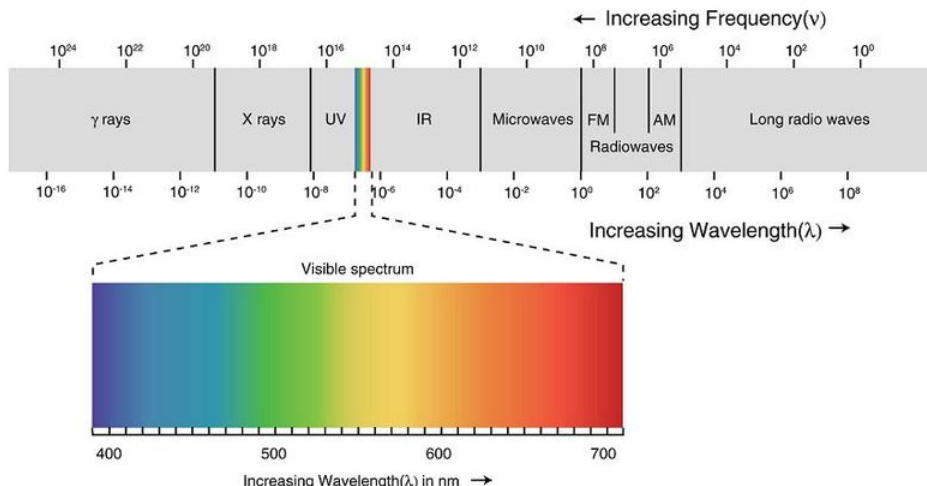
This is the way that light from the sun travels through the vacuum of outer space and arrives at the earth. This type of energy is called electromagnetic radiation. There are various levels of energy for electromagnetic radiation. One of the lower energy forms of EMR are radio waves. As the energy of EMR increases, we encounter infra-red light, visible light, ultra-violet light, and x-rays. The highest energy form of EMR is gamma radiation. The radio and television signals we use for communication are electromagnetic radiation. Astronauts in outer space can still communicate with people on earth because the radio and TV signals do not need matter to travel through – they can travel through a vacuum.

Infra-red light is frequently used in remote control devices like your TV remote. The energy that cooks your food in a microwave oven is EMR. Radio signals are used to operate automatic garage door openers and infra-red “eyes” are used to stop the garage door if something is in the way of the door. Doctors and dentists use x-rays to make pictures of bones and teeth. X-rays are such a powerful form of EMR that the energy passes through skin, flesh, and many other substances. Electromagnetic radiation at the level of x-rays and gamma rays are so powerful that they can be dangerous to human beings.



Ultraviolet light is another form of EMR. UV light is the part of the sunlight responsible for tanning skin, burning skin, and in some cases, causing skin cancer.

For humans, the most common form of EMR is visible light, which our eyes use for vision.



A Little More on Gravity

Every particle of matter attracts every other particle of matter with a force which Isaac Newton named the force of gravity. We know this force exists, we can calculate the size and direction of the force, but we cannot yet explain how or why the force works. This force between particles of matter exists everywhere in the universe and it attracts all matter everywhere in the same way.

The size of the force of gravity is dependent on the masses of the two attracting objects and also on the distance between the centers of the two objects. The amount of matter in an object is called its mass and whenever the mass of one or both of the objects is increases, the force of gravity between the objects also increases. If the objects are brought closer together, the force of gravity increases and if the objects are moved farther apart, the force of gravity decreases.

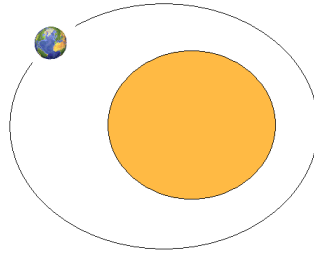
The force of gravity between two small objects, such as two people, who are standing one meter apart is so small that we cannot even measure it. When one of the objects is very large, such as the earth, the force of gravity becomes large. The force of gravity pulling on a 50 *kilogram* person standing on the surface of the earth would be about 500 *Newtons*. In this English system, this force would correspond to about 110 *pounds*. The weight of a person is, in fact, the force of gravity acting on that person. If you weigh 120 *pounds*, it means that the earth is pulling on you with a force of 120 *pounds*.



If two objects the size of the earth were right next to each other, the force of gravity between them would be about 10,000,000,000,000,000,000,000 *Newtons*. As you can see, small objects have very small attractions to due gravity but very large objects can produce a gigantic force of gravity.

If we hold two 1,000,000 *kg* objects (like a battleship) one meter apart, the force of gravity between them would be about 70 *Newtons*. If we move the objects to a 10 *meter* separation, the force would become about 0.7 *N* and if we separate the objects by 100 *meters*, the force becomes 0.007 *Newton*. The force of gravity weakens rapidly as the distance between objects increases.

With two really large objects, like the earth and the sun, there is a certain distance where the force of gravity is just enough to keep the objects circling each other. The smaller object does most of the moving in this circle (actually, more like an oval).



3.6 Chapter 3 Assessment

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 4 The Atomic Theory C-SE-TE

Chapter Outline

- 4.1 THE ATOMIC THEORY
- 4.2 FURTHER UNDERSTANDING OF THE ATOM
- 4.3 ATOMIC STRUCTURE
- 4.4 WORKSHEETS FOR CHAPTER 4
- 4.5 EXTRA READINGS FOR CHAPTER 4
- 4.6 CHAPTER 4 ASSESSMENT

Lessons and Number of Activities for Lessons

TABLE 4.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. The Atomic Theory	0	0	0	0
2. Further Understanding of the Atom	1	0	0	1
3. Atomic Structure	1	0	1	1

4.1 The Atomic Theory

Student Behavioral Objectives

The student will

- give a short history of how the concept of the atom developed.
- describe the contributions of Democritus and Dalton to the atomic theory.
- summarize Dalton's atomic theory and explain its historical development.
- state the law of definite proportions.
- state the law of multiple proportions.

Timing, Standards, Activities

TABLE 4.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Atomic Theory</i>	1.0	1g, 1n

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for The Atomic Theory(L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

4.2 Further Understanding of the Atom

Student Behavioral Objectives

The student will:

- explain the observations that led to Thomson's discovery of the electron.
- describe Thomson's plum-pudding model of the atom.
- draw a diagram of Thomson's plum-pudding model of the atom and explain why it has this name.
- describe Rutherford's gold foil experiment and explain how this experiment disproved the plum-pudding model.
- draw a diagram of the Rutherford model of the atom and label the nucleus and the electron cloud.

Timing, Standards, Activities

TABLE 4.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Further Understanding of the Atom</i>	2.0	1e, 1h

Activities for Lesson 2

Laboratory Activities

1. Rutherford's Experiment on a Large Scale

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. Timeline for the Discovery of the Elements

Answers for Further Understanding of the Atom (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

4.3 Atomic Structure

Student Behavioral Objectives

The student will:

- identify the three major subatomic particles and their charges, masses, and location in the atom.
- briefly describe the discovery of the neutron.
- define atomic number.
- describe the size of the nucleus in relation to the size of the atom.
- explain what is meant by the atomic mass of an element and describe how atomic masses are related to carbon-12.
- define mass number.
- explain what isotopes are and how isotopes affect an element's atomic mass.
- determine the number of protons, neutrons, and electrons in an atom.
- calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

Timing, Standards, Activities

TABLE 4.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Atomic Structure</i>	1.5	1a, 1e

Activities for Lesson 3

Laboratory Activities

1. Early Development of a Theory Lab

Demonstrations

1. None

Worksheets

1. Basic Atomic Structure Worksheet

Extra Readings

1. Who Discovered the Neutron and How?

Answers for Atomic Structure (L3) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

Multimedia Resources for Chapter 4

Jefferson Lab question archive containing questions on The Atomic Theory.

- http://education.jlab.org/qa/history_03.html

A website with an interactive atomic theory quiz.

- http://www.visionlearning.com/library/quiz_taker.php?qid=6&mid=50

A website with a demonstration of Rutherford's gold foil experiment.

- <http://micro.magnet.fsu.edu/electromag/java/rutherford/>

Laboratory Activities for Chapter 4

Early Development of a Theory Lab

Understanding The Law of Definite Proportions and The Law of Multiple Proportions

In today's lesson, we have discussed two different laws that are often confused – *The Law of Definite Proportions* and *The Law of Multiple Proportions*. In this lab, we will try to clarify the difference, and also to illustrate how each arises as a consequence of Dalton's Atomic Theory. Materials:

Scissors

A copy (preferably in color) of the "Paper Atoms" page for each student or group of students tape

Method:

At the beginning of the experiment, you should cut each of the nitrogen, *N*, atoms and each of the oxygen, *O*, atoms out, placing the *N* atoms and *O* atoms in two separate piles. These "paper atoms" will be used in both mini-labs. Notice that the *N* atoms have "14 *u*" on them. This means that a nitrogen atom has a mass of 14 "atomic mass units." (Atomic mass units can be used to measure mass, just as hours can be used to measure time, and miles can be used to measure distance.) Similarly, the *O* atoms have "16 *u*" on them. This means that an oxygen atom has a mass of 16 "atomic mass units."

According to Dalton's Atomic Theory:

1. Matter is made of tiny particles called atoms. *Notice in this lab, we are representing atoms as small circles of paper. We will use these paper circles and assume that, when put together, they form "matter".*

- Atoms are indivisible. During a chemical reaction, atoms are rearranged, but they do not break apart, nor are they created or destroyed. *This means that you **should not cut your paper atoms into pieces**. Nor should you throw out any paper atoms, or borrow any paper atoms from your friends. You should have 5 nitrogen atoms, and 20 oxygen atoms **at all times**.*
- All atoms of a given element are identical in mass and other properties. *Notice how all of the nitrogen, N, atoms are the same size, “mass” and color. Similarly, all of the oxygen, O atoms are the same size, “mass” and color.*
- The atoms of different elements differ in mass and other properties. *Notice, however, that the nitrogen, N, atoms are not the same size, “mass” or color as the oxygen, O, atoms.*
- Atoms of one element can combine with atoms of another element to form “compounds” – new, complex particles. In a given compound, however, the different types of atoms are always present in the same relative numbers. *We’ll look at this rule in more detail in the two mini-labs.*

Mini-lab One: The Law of Definite Proportions

*In a given chemical substance, the elements are always combined in the **same proportions by mass**.*

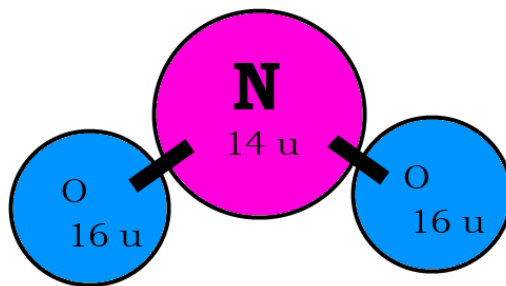
Method:

We’ll consider the Law of Definite Proportions by “making” the substance “ NO_2 ”, or “nitrogen dioxide”. Draw the following table in your notebook.

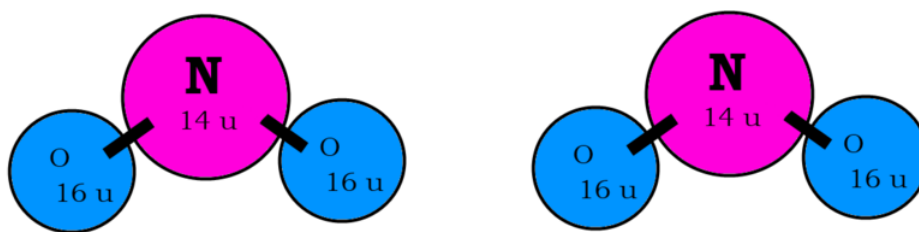
TABLE 4.5: Molecular Composition

Number of molecules	Total mass of nitrogen (u)	Total mass of oxygen (u)	Mass Ratio ($N \div O$)
1	14	32	0.4375
2	28	64	
3			
4			
5			

- Begin by constructing a single NO_2 “molecule” (“molecule” is a word given to a complex particle formed by joining several atoms together) by taping an oxygen atom to either side of one of your nitrogen atoms.



- Add up the total mass of N atoms in your one NO_2 “molecule.” This is easy, of course – since you only have one N atom, the total mass is $14 u$. Fill this number into the first row of your table under the column heading “Total mass of nitrogen (u).”
- Add up the total mass of O atoms in your one NO_2 “molecule”. This is slightly more challenging – since you have two O atoms, the total mass is $16 u + 16 u = 32 u$. Fill this number into the first row of your table under the column heading “Total mass of oxygen (u)”.
- Next, construct a *second* NO_2 molecule by taping an oxygen atom to either side of another one of your nitrogen atoms. You should now have two NO_2 molecules.



5. Add up the total mass of *N atoms* in both NO_2 “molecules.” Since you now have 2 nitrogen atoms, you have a total mass of $14\text{ u} + 14\text{ u} = 28\text{ u}$. Fill this number into the second row of your table under the column heading “Total mass of nitrogen (u).”

6. Add up the total mass of *O atoms* in both NO_2 “molecules.” Because you now have 4 oxygen atoms, you have a total mass of $16\text{ u} + 16\text{ u} + 16\text{ u} + 16\text{ u} = 64\text{ u}$. Fill this number into the second row of your table under the column heading “Total mass of oxygen (u).”

7. Continue in this manner, filling out the table as you go, until you’ve made a total of 5 molecules of NO_2 . (As you get more nitrogen atoms and more oxygen atoms, you may need to use your calculator to do the adding for you. If you’re good at math, you could also use multiplication, instead of addition. For instance, when you have 4 oxygen atoms, you can find the total mass as $4 \times 16\text{ u} = 64\text{ u}$.)

8. Once you have filled out the entire first two columns of the table, you can fill out the third column. This column is the mass ratio of nitrogen to oxygen. In other words, for each row, you need to divide the total mass of *N* (the number in the first column) by the total mass of *O* (the number in the second column). For the first row, then, this would be $14 \div 32 = 0.4375$. For the second row it would be $28 \div 64$.

What do you notice about the numbers in the third column? When nitrogen, *N*, and oxygen, *O*, react to form nitrogen dioxide, NO_2 , do they always *react in the same proportions by mass*? Does the formation of NO_2 follow the Law of Definite Proportions?

Notice that in order to prove the Law of Definite Proportions for NO_2 , we only had to make a few assumptions.

According to Dalton’s Atomic Theory:

1. Matter is made of tiny particles called atoms. *We made nitrogen dioxide by assuming that it was composed of N atoms and O atoms.*
2. Atoms are indivisible. During a chemical reaction, atoms are rearranged, but they do not break apart, nor are they created or destroyed. *In order to produce nitrogen dioxide, we connected oxygen atoms to nitrogen atoms. We did not, however, cut, throw out, or duplicate our paper atoms.*
3. All atoms of a given element are identical in mass and other properties. *Notice how all of the nitrogen, N, atoms were the same size, and color, and they all had a “mass” of 14 u. Similarly, all of the oxygen, O, atoms were the same size and color, and they all had a “mass” of 16 u.*
4. The atoms of different elements differ in mass and other properties. *The nitrogen, N, atoms were not the same size, color or “mass” as the oxygen, O, atoms.*
5. Atoms of one element can combine with atoms of another element to form “compounds” – new, complex particles. In a given compound, however, the different types of atoms are always present in the same relative numbers. *Notice that every nitrogen dioxide molecule contained one nitrogen atom and two oxygen atoms. In other words, no matter how much nitrogen dioxide we made, we always had two oxygen atoms for every nitrogen atom.*

Mini-lab Two: The Law of Multiple Proportions

“When two elements react to form more than one substance, the different masses of one element (like oxygen) that are combined with the same mass of the other element (like nitrogen) are in a ratio of *small whole numbers*.”

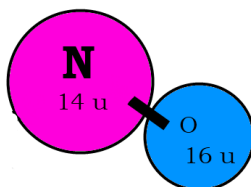
Method:

You'll need to cut the tape holding together your NO_2 molecules from Mini-lab One: The Law of Definite Proportions, since this lab will also require that you start with 5 N atoms and 20 O atoms. We'll consider the Law of Multiple Proportions by "making" the substances "NO", or "nitrogen monoxide", NO_2 , or "nitrogen dioxide", NO_3 , or "nitrogen trioxide" and NO_4 or "nitrogen tetroxide". In each case, we'll use five nitrogen atoms, meaning we'll always have the same mass of nitrogen. Draw the following table in your notebook.

TABLE 4.6: Data

Substance	Total mass of oxygen (u) in 5 molecules
NO	80
NO_2	
NO_3	
NO_4	

1. Begin by constructing **five** separate NO molecules by taping a single oxygen atom to each of your nitrogen atoms.

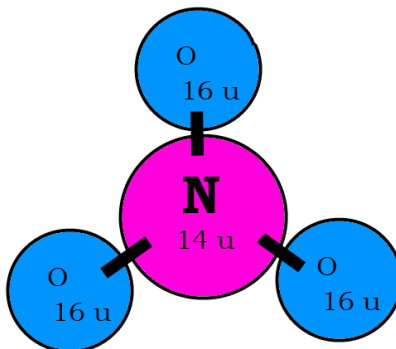


2. Add up the total mass of O atoms in *all five* NO molecules. This gives a total of $16\text{ u} + 16\text{ u} + 16\text{ u} + 16\text{ u} + 16\text{ u} = 80\text{ u}$. Fill this number into the first row of your table under the column heading "Total mass of oxygen (u)".

3. Now cut all of your NO molecules apart so that you have your original 5 N atoms, and your original 20 O atoms. Next, construct **five** separate NO_2 molecules by taping an oxygen atom to either side of each of your nitrogen atoms.

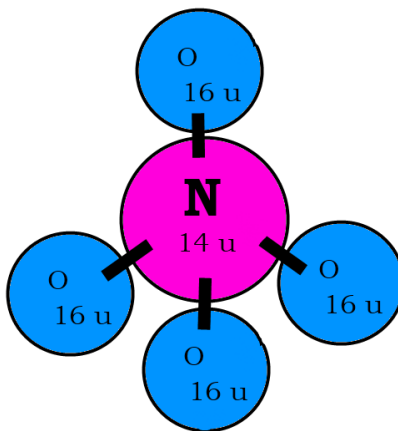
4. Add up the total mass of O atoms in *all five* NO_2 molecules. Fill this number into the second row of your table under the column heading "Total mass of oxygen (u)". (Note: If you'd like to save time at this point, you do not actually need to construct all 5 NO_2 molecules, since that's what you already did in Mini-lab One: The Law of Definite Proportions. You can use your results from Mini-lab One: The Law of Definite Proportions to find the total mass of O atoms in five NO_2 molecules.)

5. Now cut all of your NO_2 molecules apart so that you have your original 5 N atoms, and your original 20 O atoms. Next, construct **five** separate NO_3 molecules by taping three oxygen atoms to each of your nitrogen atoms.



6. Add up the total mass of O atoms in *all five* NO_3 molecules. Fill this number into the third row of your table under the column heading "Total mass of oxygen (u)".

7. Now cut all of your NO_3 molecules apart so that you have your original 5 N atoms, and your original 20 O atoms. Next, construct **five** separate NO_4 molecules by taping four oxygen atoms to each of your nitrogen atoms.



8. Add up the total mass of *O atoms* in *all five NO₄* molecules. Fill this number into the fourth row of your table under the column heading “Total mass of oxygen (u)”.

Notice that in every substance you formed, you used *5 N atoms*. As a result, you used the same mass of *N* in each case. *5 N atoms* gives a total of $14\text{ u} + 14\text{ u} + 14\text{ u} + 14\text{ u} + 14\text{ u} = 70\text{ u}$. Since you used the same mass of nitrogen in each example, but formed different substances with oxygen, The Law of Multiple Proportions suggests that the ratios of the masses of oxygen used in the different substances should be small whole numbers. Use your table to fill out the following questions and prove that this is true.

a. The ratio of the mass of oxygen in *NO₂* to the mass of oxygen in *NO* for *70 u* of *N*:

$$= (\text{total mass of oxygen in } NO_2) \div (\text{total mass of oxygen in } NO)$$

$$= (\underline{\hspace{2cm}}) \div (\underline{\hspace{2cm}})$$

$$= \underline{\hspace{2cm}}$$

b. The ratio of the mass of oxygen in *NO₃* to the mass of oxygen in *NO* for *70 u* of *N*:

$$= (\text{total mass of oxygen in } NO_3) \div (\text{total mass of oxygen in } NO)$$

$$= (\underline{\hspace{2cm}}) \div (\underline{\hspace{2cm}})$$

$$= \underline{\hspace{2cm}}$$

c. The ratio of the mass of oxygen in *NO₄* to the mass of oxygen in *NO* for *70 u* of *N*:

$$= (\text{total mass of oxygen in } NO_4) \div (\text{total mass of oxygen in } NO)$$

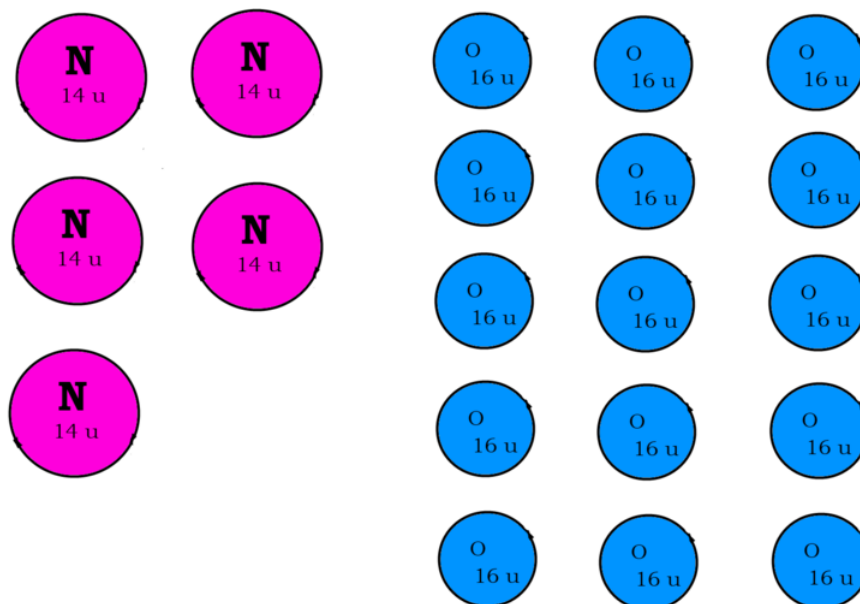
$$= (\underline{\hspace{2cm}}) \div (\underline{\hspace{2cm}})$$

$$= \underline{\hspace{2cm}}$$

What do you notice about the ratios? Are they small whole numbers? Does the Law of Multiple Proportions hold for the formation of *NO*, *NO₂*, *NO₃* and *NO₄*? Notice, again, how the Law of Multiple Proportions occurs naturally provided we assume that *NO*, *NO₂*, *NO₃* and *NO₄* are formed according to rules in Dalton’s Atomic Theory.

To the Teacher: We could do the same calculations as above, but compare the mass of oxygen in *NO₂* to the mass of oxygen in *NO₃*, or the mass of oxygen in *NO₃* to the mass of oxygen in *NO₄*. This wouldn’t give obvious whole numbers like 2, 3 or 4, but it would give fractions that could easily be converted into small whole number ratios like 2:3 or 3:4. Since this adds a level of complication, in this lab we only compare the higher nitrogen oxides with *NO*.

PAPER ATOMS



Rutherford's Experiment on a Large Scale

This lab requires a pre-lab preparation, either by the teacher, or by the students

We have discussed Rutherford's Gold Foil Experiment. In this lab, we'll try to recreate the principle that Rutherford used in his Gold Foil experiment, but we'll do it with objects that are big enough for us to hold and see. This lab can be done at the front of the classroom by a few volunteers, or it can be done in groups (provided that there are enough supplies and that enough experimental arenas have been constructed).

Materials:

small rubber ball (the kind that can be purchased from coin machines at shopping malls)

blindfold

cardboard

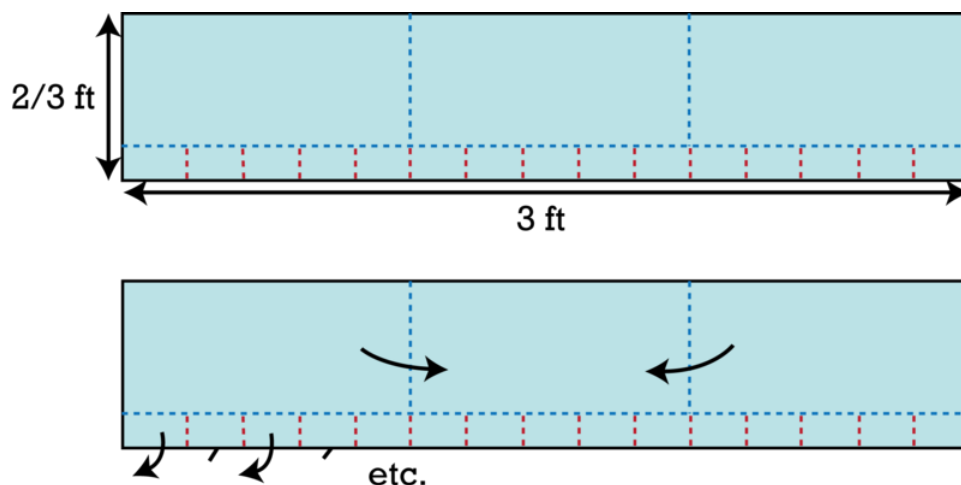
glue gun

spice jar (approximately 6 inches tall and 2 inches wide) filled with sand or some other heavy substance

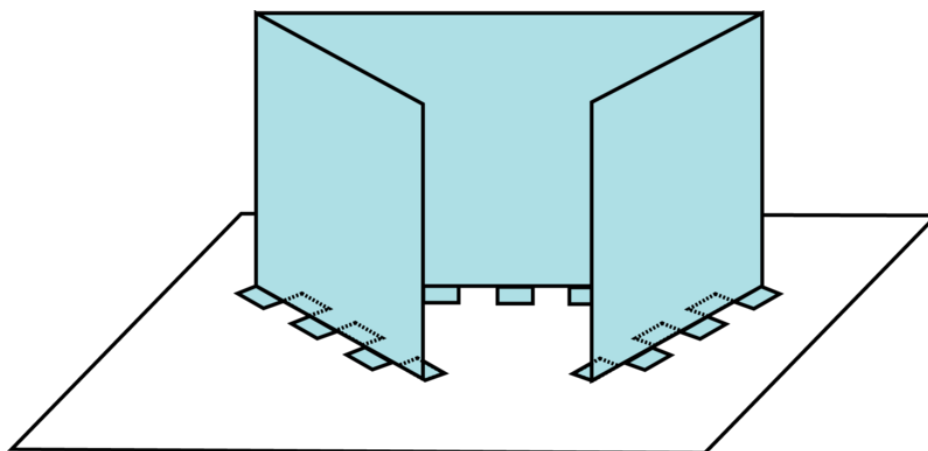
Method:

Pre-lab preparation

Cut a strip of cardboard according to the diagram below:



The top figure shows the dimensions, while the bottom figure shows the directions in which to make all folds. In both diagrams the red lines are to be cut, while the blue lines are folded. Once you have the strip cut and folded, assemble it into the experimental arena as shown below, and use a glue gun to fasten the experimental arena to a lower cardboard sheet.



You have now completed the pre-lab preparation.

Main Lab

1. Blindfold the student performing the experiment. Then choose to either leave the experimental arena empty, or to place the spice jar in the middle of the arena (the blindfolded student should not know which choice you have made).
2. Give the blindfolded student a small rubber ball, and allow her to *gently* roll the rubber ball into the opening in the experimental arena (since the student is blindfolded, she may require some guidance in order to get the rubber ball into the arena. Help to make sure that she is rolling the rubber ball in the right direction).
3. Make a mark where the rubber ball *first* hits the walls of the arena. (If the rubber ball bounces back out of the mouth of the arena, make a mark right on the edge of the arena mouth).
4. Allow the blindfolded student to roll the ball about 14 more times. Each time record a mark on the wall where the rubber ball *first* hits. (Do not count trials where the blindfolded student does not manage to get the ball into the arena at all).
5. After the student has rolled the ball about 15 times, remove the spice jar (if it was in the experimental arena; if it wasn't, there won't be anything to remove), and then take the blindfold off of the student who was performing the experiment.

6. Show the previously blindfolded student where each of her rubber balls hit the walls of the arena. See if she can guess whether or not the spice jar was in the arena based on the marks.
7. Repeat the experiment taking turns being blindfolded. Make sure you try both with and without the spice jar in the arena (although don't make your choices predictable!). If you are trying the experiment several different times, you may need to use different colored marking pens, or different symbols to mark on the walls each time – otherwise you may get the results from the different experiments confused.

Notice how the rubber ball bounced differently and struck the wall in different places when the spice jar was in the arena. By determining whether or not the rubber ball was deflected by the spice jar, the student should have been able to tell whether or not the spice jar was in the arena, even though she was blindfolded, and so had no way of seeing it.

This is the same principle that Rutherford used to prove that atoms had hard nuclei. If there hadn't been any hard nuclei in the gold atoms, the alpha particles would have sailed right through, and made a mark on the wall behind the gold foil. Any time that an alpha particle hit a gold nucleus, though, the alpha particle bounced off of the nucleus, and made a mark on the wall in another location.

Demonstrations for Chapter 4

4.4 Worksheets for Chapter 4

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Basic Atomic Structure Worksheet

- List the three basic sub-atomic particles and their respective charges
(a) _____ (b) _____ (c) _____
- The number of protons in one atom of an element determines the atom's _____.
- The atomic number tells you the number of _____ in one atom of an element. It also tells you the number of _____ in a neutral atom of that element.
- The _____ of an element is the average mass of an element's naturally occurring atoms.
- The _____ of an element is the total number of protons and neutrons in the nucleus of the atom.
- In order to calculate the number of neutrons in a nucleus, you must subtract the _____ from the _____.
- Give the symbol of and the number of protons in one atom of:
 - Lithium _____
 - Iron _____
 - Oxygen _____
 - Krypton _____
 - Uranium _____
 - Boron _____
 - Chlorine _____
- Give the symbol of and the number of neutrons in one atom of the most common isotope of:
 - Barium _____
 - Carbon _____
 - Fluorine _____
 - Europium _____
 - Xenon _____
 - Hydrogen _____
- Name the element which has the following numbers of particles:
 - (a) 26 electrons, 29 neutrons, 26 protons _____
 - (b) 53 protons, 74 neutrons _____
 - (c) 2 electrons (neutral atoms) _____
 - (d) 20 protons _____

10. If you know ONLY the following information can you ALWAYS determine what the element is? (Yes/No)

- (a) Number of protons _____ (b) Number of neutrons _____ -
_____ (c) Number of electrons in a neutral atom _____ (d) Number of electrons

Answers to Worksheets

- **The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.**

4.5 Extra Readings for Chapter 4

Timeline for the Discovery of the Elements

TABLE 4.7: Discovery of the Elements by Year

Method of Discovery	Year of Discovery	Element Symbols
Found Free in Nature or Simple Metallurgy	Before 1 A.D.	C, S, Hg, Sn, Pb, Fe, Cu, Ag, Au
Simple Metallurgy	Alchemists to 1735	Zn, P, Bi, Sb, As
Simple Metallurgy	1735 – 1745	Pt, Co
Simple Metallurgy	1745 – 1755	Ni
	1755 – 1765	
Electrochemistry	1765 – 1775	F, Mn, Cl, O, N, H
Electrochemistry	1775 – 1785	Te, W, Mo
Electrochemistry	1785 – 1795	Y, Ti, Sr, U
Electrochemistry	1795 – 1805	Ir, Os, Rh, Pd, Ce, Ta, Nb, Be, Fr, V
Electrochemistry	1805 – 1815	I, B, Mg, Ca, Ba, K, Na
Electrochemistry	1815 – 1825	Zr, Si, Se, Cd, Li
Electrochemistry	1825 – 1835	Th, Br, Al
Electrochemistry	1835 – 1845	Ru, Er, Tb, La
	1845 – 1855	
Electrochemistry	1855 – 1865	In, Tl, Rb, Cs
	1865 – 1875	
Separation Techniques	1875 – 1885	Tm, Ho, Sc, Sm, Yb, Ga
Separation Techniques	1885 – 1895	Ar, Ge, Dy, Gd, Nd, Pr
Separation Techniques	1895 – 1905	Rn, Ac, Ra, Po, Xe, Ne, Kr, Eu, He
Separation Techniques	1905 – 1915	Lu
Separation Techniques	1915 – 1925	Pa, Hf
Separation Techniques	1925 – 1935	Re
Nuclear Synthesis	1935 – 1945	Cm, Pu, Np, At, Fr, Tc
Nuclear Synthesis	1945 – 1955	Cf, Bk, Pm, Am, Es, Fm, Md
Nuclear Synthesis	1955 – 1965	No, Lr, Rf, Db, Sg, Bh, Hs, Mt
Nuclear Synthesis	1965 – Present	Ds, Rg, Cp, 114, 116, 118

Who Discovered the Neutron and How?

The construction of the modern atomic model consisting of the central nucleus and orbiting electrons was the result of years of experimentation and the dedication and insight of countless scientists. Yet, well into the twentieth century, the picture remained incomplete and inconsistencies and questions remained to be elucidated. By 1930, the fundamental positive and negative particles, the proton and the electrons had been identified and their dispositions relative to each other characterized. Ernest Rutherford, the discoverer of the proton, suggested the existence of what he termed a “proton-electron” pair, a heavy, yet neutral particle found in the nucleus, mainly on the basis of the

differences between the atomic number (Z) of several atoms and their atomic mass. Further contemplation of another heavy fundamental particle arose to account for the unusual radiation emitted by beryllium atoms when bombarded by a stream of alpha particles. Irene and Frederic Joliot-Curie found this unusual radiation to be capable of ejecting protons, therefore had a mass comparable to that of protons. This was a confusing result, in that most physicists were under the assumption that this radiation better corresponded with the high energy but zero mass gamma radiation.

James Chadwick, who had worked for Ernest Rutherford at Manchester University and later at Cambridge University, replicated the Joliot-Curie experiment but with the intention of searching for a new fundamental neutral particle. Chadwick found that other light atoms other than beryllium gave off these new particles upon bombardment. He found the mass of this newly proposed particle to be about 10% greater than that of the proton, by comparing the velocity of the protons emitted by striking a hydrogen target with the neutral rays. He disproved the possibility of a proton-electron dual particle by illustrating that under no circumstances did the neutral ray particles degrade into smaller entities. For his efforts and insight, James Chadwick was awarded the Nobel Prize in Physics in 1935.

Since the discovery of the neutron, the advent of particle accelerators has produced evidence for hundreds of subatomic particles. Protons and neutrons are defined as baryons or heavy particles, whereas the list of leptons (light particles) is extensive.

4.6 Chapter 4 Assessment

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 5 The Bohr Model of the Atom C-SE-TE

Chapter Outline

- 5.1 THE NATURE OF LIGHT
- 5.2 ATOMS AND ELECTROMAGNETIC SPECTRA
- 5.3 THE BOHR MODEL OF THE ATOM
- 5.4 WORKSHEETS FOR CHAPTER 5
- 5.5 EXTRA READINGS FOR CHAPTER 5
- 5.6 ASSESSMENT FOR CHAPTER 5

Lessons and Number of Activities for Lessons

TABLE 5.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. The Nature of Light	0	0	0	1
2. Atoms and Electromagnetic Spectra	1	1	0	0
3. The Bohr Model of the Atom	0	0	1	0

5.1 The Nature of Light

Student Behavioral Objectives

The student will:

- perform calculations involving the relationship between the wavelength and frequency of electromagnetic radiation, $v = \lambda f$.
- perform calculations involving the relationship between the energy and the frequency of electromagnetic radiation, $E = hf$.
- state the velocity of electromagnetic radiation in a vacuum.
- name at least three different areas of the electromagnetic spectrum.
- when given two comparative colors or areas in the electromagnetic spectrum, identify which area has the higher wavelength, the higher frequency, and the higher energy.

Timing, Standards, Activities

TABLE 5.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Nature of Light</i>	1.5	1h, 1j

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. What is the Electromagnetic Spectrum?

Answers for The Nature of Light (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

5.2 Atoms and Electromagnetic Spectra

Student Behavioral Objectives

The student will:

- describe the appearance of an atomic emission spectrum.
- explain why an element can be identified by its emission spectrum.

Timing, Standards, Activities

TABLE 5.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Atoms and Electromagnetic Spectra</i>	1.0	None

Activities for Lesson 2

Laboratory Activities

1. Light and the Atomic Spectra Lab

Demonstrations

1. Atomic Spectra Viewed Through a Diffraction Grating Demo

Worksheets

1. None

Extra Readings

1. None

Answers for Atoms and Electromagnetic Spectra (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

5.3 The Bohr Model of the Atom

Student Behavioral Objectives

The student will:

- describe an electron cloud containing Bohr's energy levels.
- describe how the Bohr model of the atom explains the existence of atomic spectra.
- explain the limitations of the Bohr model and why it had to be replaced.

Timing, Standards, Activities

TABLE 5.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Bohr Model of the Atom</i>	2.0	1i, 1j

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Electromagnetic Radiation and the Bohr Atom Worksheet

Extra Readings

1. None

Answers for The Bohr Model of the Atom (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 5

This website has an interactive lesson on light and color.

- <http://www.learner.org/teacherslab/science/light/>

This website has a video on the electromagnetic spectrum.

- <http://www.teachersdomain.org/resource/phy03.sci.phys.energy.nasaspectrum/>

This website provides more information about the properties of electromagnetic waves and includes an animation showing the relationship between wavelength and color.

- <http://micro.magnet.fsu.edu/primer/java/wavebasics/index.html>

This website “Spectral Lines” has a short discussion of atomic spectra. It also shows the emission spectra of several elements.

- <http://www.colorado.edu/physics/2000/quantumzone/index.html>

This video provides a summary of the Bohr atomic model and how the Bohr model improved upon Rutherford’s model.

- <http://www.youtube.com/watch?v=bDUxygs7Za8>

This video describes the important contributions of many scientists to the modern model of the atom. It also explains Rutherford’s gold foil experiment.

- <http://www.youtube.com/watch?v=6773jO6fMnM>

The following video shows the production of the hydrogen spectrum.

- <http://www.youtube.com/watch?v=QI50GBUJ48s&feature=related>

The following video shows the story of the Rutherford Experiment.

- <http://www.youtube.com/watch?v=hpKhjKrBn9s>

The following video relates quantum mechanics to atomic structure.

- <http://www.youtube.com/watch?v=-YYBCNqNqYNM&feature=related>

The following video demonstrates the emission of light from an atom.

- <http://www.youtube.com/watch?v=XVZSD23mD1Q&feature=related>

Laboratory Activities for Chapter 5

Light and the Atomic Spectra Lab

White light is really a mixture of all the possible wavelengths of light in the visible spectrum.

We have discussed the difference between a continuous spectrum and a discontinuous spectrum, and you've learned that when all of the light waves in the visible portion of the continuous electromagnetic spectrum are mixed together, your eye sees 'white'. In this lab, we will show how white light can be spread out into the rainbow of colored light that it is composed of using a prism. We will also show how colors can be 'stirred' back together to form white using a color wheel.

Mini-lab One: The Rainbow of Colors in White Light

Materials:

Slide projector

Glass prism

Screen

Method:

Turn off the lights in the classroom. Then allow white light from a slide projector to pass through the glass prism and project onto the screen behind. You should see a rainbow of colors appear where the light hits the screen. This is a continuous spectrum, and should convince you that 'white' really isn't a 'color', but rather 'all colors' mixed together.

Mini-lab Two: Mixing a Rainbow into White

Materials:

A color wheel (color wheels can be purchased from a school supplier, for instance at <http://www.teachersource.com/LightAndColor/PersistenceofVision/MagicColorWheel.aspx>). It is possible to make your own. However, if you choose to do this, be sure to get the colors in the correct proportions. If you choose to make your own, it can be spun by fastening it to the end of a drill, and powering up the drill.

Method:

Start with the color wheel at rest, noticing how it contains all of the colors in the rainbow. Begin spinning the color wheel, slowly at first, and then, with increasing speed. Notice how, once it is spinning fast enough, the colored wheel turns white. Again, this should convince you that 'white' really isn't a 'color,' but rather, all colors at once.

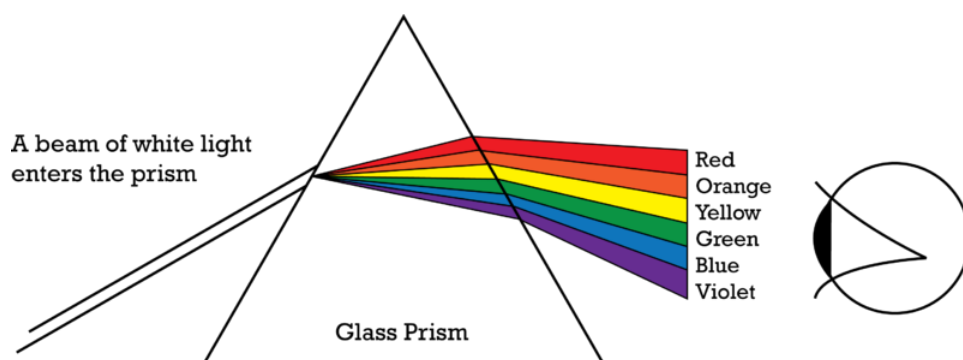
Demonstrations for Chapter 5

Atomic Spectra Viewed Through a Diffraction Grating Demo

Background:

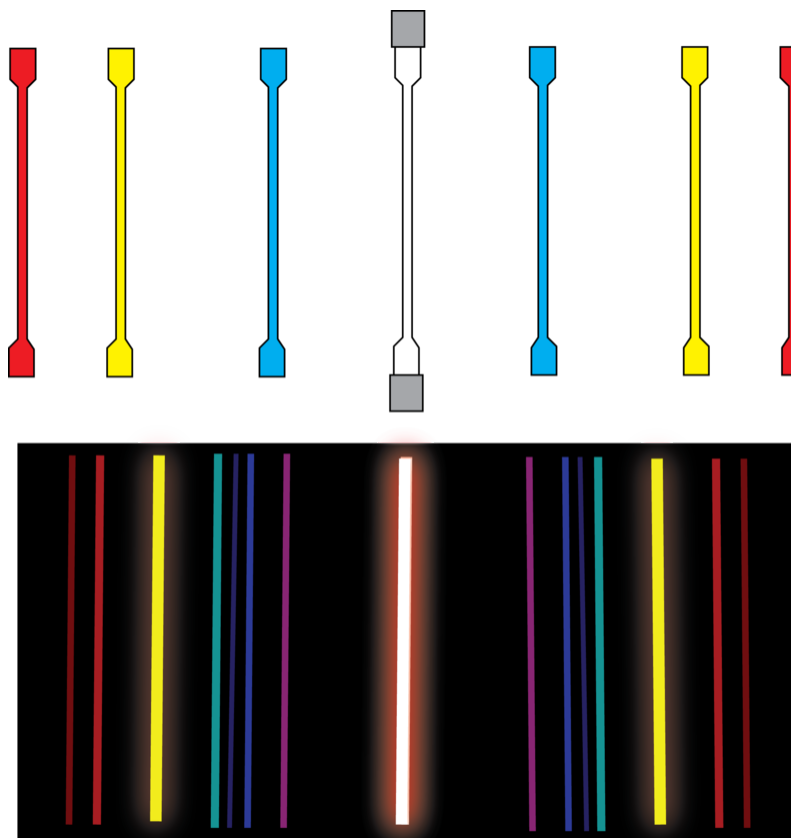
The speed of light in a vacuum is 3.0×10^8 meters/second. All frequencies of light in a vacuum travel at the same speed. When light passes through media other than a vacuum, such as glass or water, however, all frequencies do not travel at the same speed. In general, light travels at a speed of about 2×10^8 m/s in glass and not all frequencies travel

at the same speed. Long wavelengths of light, such as red, are slowed down less in glass than short wavelengths like blue. When light traveling in air passes into glass at an angle, its path is bent (diffracted) due to one side of the wave front slowing down sooner than the other side. Since each frequency of light is slowed by a slightly different amount, each frequency of light will bend by a slightly different amount. This allows us to use a prism (a triangular piece of glass) to separate the frequencies of light from each other.



When a beam of white light enters a prism, the red light is slowed and bends slightly as it passes through the boundary between air and glass. The light travels through the prism and is bent again as it leaves the prism. Blue light, which is slowed more when it enters the glass, would also bend upon entering the glass and it bends at a slightly greater angle than red light. The blue light would bend again as it leaves the prism and again would bend slightly more than red light. When the light comes out the other side of the prism, the different frequencies that made up the white light would be traveling at slightly different directions. If the light is projected onto a surface, the observer will see a rainbow of colors as shown above. The rainbow can also be observed with the naked eye looking back into the prism.

Light composed of a mixture of frequencies can also be separated into individual frequencies with a diffraction grating. A diffraction grating is a thin piece of glass or plastic which has thousands of vertical scratches per centimeter on its surface. The method by which a diffraction grating functions is more complex than a prism; it isn't necessary to examine the method here. When you look through a diffraction grating at a thin beam of light, the light source is still visible, but images of the light source in each frequency present in the original light source will also appear several times beside the original light source.



In the sketches above, the original light source appears in the center and this light source will be the composite color – i.e. the color due to the mixture of the frequencies. To the sides of the original source will be images of the source. There will be an image for each frequency present in the original source. When the images are produced by a diffraction grating, the higher frequencies will appear closest to the original source.

If you look at a source of truly white light, such as a tungsten filament, through a prism or diffraction grating, you will see a complete rainbow of the electromagnetic spectrum. This is because it is heated. If you look at an atomic spectrum (the light emitted by excited atoms of a particular element), you will see only a few lines because the atoms of a particular element emit only a few frequencies of light.

Apparatus and Materials:

- Incandescent light bulb (15 watt, tungsten filament) with a plug in socket.
- Spectrum tube power supply
- Several spectrum tubes
- Diffraction gratings (one for each student or pair of students)

Procedure:

Important note: The spectrum tube power supply produces a very high voltage. Never attempt to replace spectrum tubes unless the power supply is turned off. Never touch or even point toward the spectrum tubes unless the power supply has been turned off. It is advisable, that when changing the spectrum tube, you should not only turn the power supply off but also unplug it. This avoids the circumstance of accidentally bumping the on/off switch with your arm while fingers are near the spectrum tubes. The spectrum tube power supply is quite dangerous if not handled properly.

1. Pass out diffraction gratings to students and warn them that oily fingerprints on the surface of the grating degrade their function considerably. Effort must be made to handle the gratings only by the frame or mount.

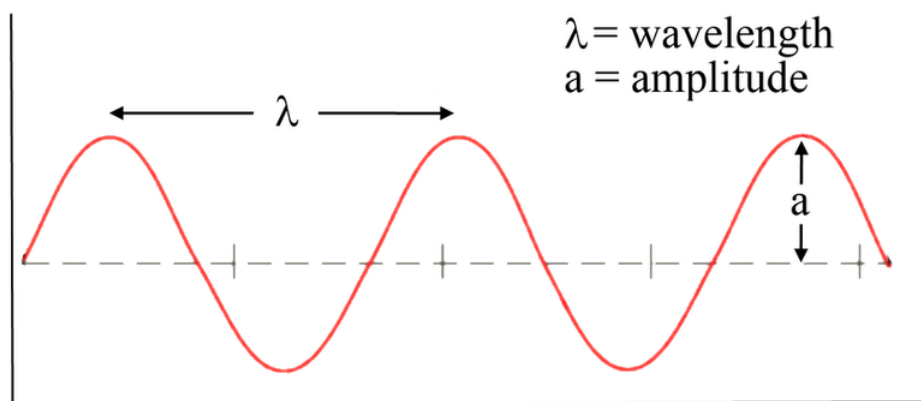
2. Use the incandescent bulb first so the students can see the entire white light spectrum. It is best to view the spectra in a darkened room. You can also use this opportunity for the students to properly orient their diffraction gratings. The gratings should be held so that the scratches are vertical.
3. Remove the white light and replace it with the spectrum tube power supply. Begin with a simpler spectrum such as hydrogen, helium, or sodium. Insert the spectrum tube into the power supply, keep your hands away from the tube, and turn the power supply on. The spectrum tubes are intended to be turned on for only about 30 *seconds* at a time. You may turn them on for 30 *seconds* and then off for 30 *seconds* and then back on if the students need more time for viewing.
4. If there is sufficient light in the room, have the students sketch the spectrum of some of the elements as they view them.
5. When you change spectrum tubes, turn off and unplug the power supply. Use hand protection (such as a pot holder) to remove the tube because it will be HOT! After replacing the tube, plug the power supply in and turn it on.
6. When you have gone through your supply of spectrum tubes, you may use one of the tubes you used before and see if the students can identify the element from their sketches of the various spectra.
7. The demo should be followed by a discussion of how the atoms produce these light spectra.

5.4 Worksheets for Chapter 5

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Electromagnetic Radiation and the Bohr Atom Worksheet

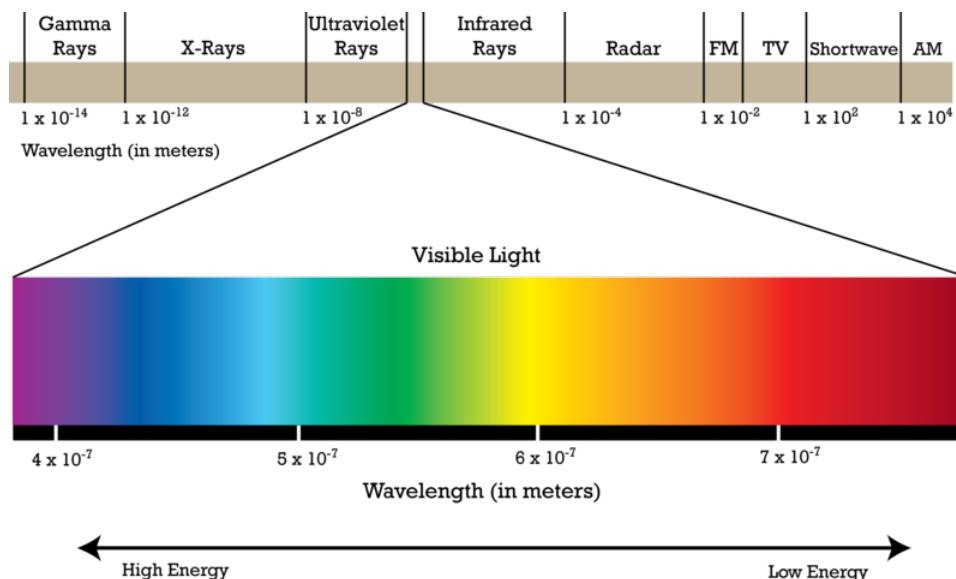
Light is known to have the wave-like properties of frequency, f , and wavelength, λ . These are illustrated below. The x-axis is a measure of time.



The distance between the peaks is called the wavelength and the number of waves that pass a certain point per unit time is called the frequency. The frequency is expressed in waves/second or 1/seconds or hertz (Hz). As the frequency increases, the wavelength decreases. In electromagnetic radiation, the wavelength and frequency are related by the equation

$$c = \lambda f$$

Where c = the speed of light, $2.998 \times 10^8 \text{ m/s}$, λ = wavelength in meters, and f = frequency in s^{-1} or hertz. The electromagnetic spectrum is shown below.



Planck recognized that energy is quantized and related the energy of electromagnetic radiation to its frequency.

$$E = hf$$

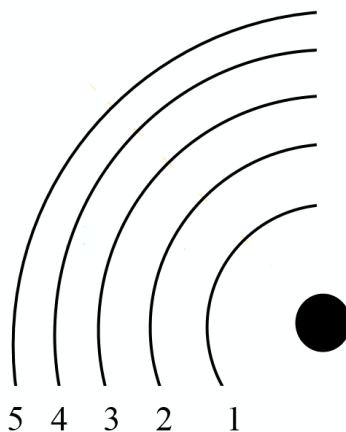
Where E = energy in Joules, h = Planck's constant, $6.626 \times 10^{-34} \text{ J} \cdot \text{s}$.

Exercises

- Which color of visible light has the shortest wavelength?
- Which type of electromagnetic radiation has the shortest wavelengths?
- Which type of electromagnetic radiation has wavelengths slightly longer than visible light?
- The wavelength of green light is about 522 nm . What is the frequency of this radiation?
- What is the wavelength of a photon that has a frequency of $2.10 \times 10^{14} \text{ Hz}$? Give your answer in nanometers and identify the type of radiation this is.
- For each of the following pairs of quantities, indicate whether they are related directly or indirectly.
 - energy and wavelength
 - wavelength and frequency
 - frequency and energy
- A classical radio station broadcasts at 93.5 megahertz, 10^6 hertz. Find the wavelength of this radiation, in meters, and the energy of one of these photons, in Joules.
- What is the energy of a photon with a wavelength of 827 nm ?

Bohr invented the concept of fixed electron energy levels for atoms and applied this concept to the line spectra of elements. He proposed that only certain energy levels are allowed for electrons within the structure of an atom. Electrons are allowed to move between these energy levels. When electrons move up in energy levels, they absorb very specific amounts of energy and when they move down in energy levels, they emit very specific amounts of energy. The energies absorbed when electrons move up correspond to the dark line atomic spectra and the energies emitted when electrons move down correspond to the bright line atomic spectra. The energies absorbed and emitted are measures of the energy difference between the two energy levels. Ionization energy is the energy required to completely remove an electron from an atom. This can be thought of as the transition between the starting energy level and the infinite energy level.

The energy level concept can be illustrated for a hydrogen atom as shown below.



An equation has been developed for calculating the energy of each energy level.

$$E = (-2.178 \times 10^{-18} \text{ J})\left(\frac{Z^2}{n^2}\right)$$

where E is the energy of the energy level, Z is the nuclear charge, and n is the energy level number. Bohr's model only worked correctly for atoms with a single electron (essentially hydrogen), and therefore, Z is almost always equal to 1.

This equation can be used to find the energy of each energy level in a hydrogen atom and can also be used to find the difference between two energy levels by calculating the energies of the two levels and subtracting.

Exercises continued

9. Calculate the energy of energy level 2 in a hydrogen atom.
10. If an electron in a hydrogen atom dropped from $n = 4$ to $n = 2$, what energy would be emitted?
11. What is the frequency of the photon emitted in question 10?
12. How much energy would be released as an electron moved from the $n = 4$ to the $n = 3$ energy levels in a hydrogen atom?
13. What frequency of light would be emitted in the electron transition mentioned in question 15?
14. What would be the wavelength of light emitted in question 15?
15. What type of light would be emitted in question 15?

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

5.5 Extra Readings for Chapter 5

What is the Electromagnetic Spectrum?

The visible light or radiant energy that illuminates our portion of the universe and enriches our existence with the appearance of different colors and hue intensities, was the first type of electromagnetic radiation evident to mankind. The remaining regions of the electromagnetic spectrum have only recently been elucidated. These varied regions can be differentiated on the basis of their wavelength (in length units of meters or millimeters), or frequency (in sec^{-1} or Hertz units).

The first region other than the section of the spectrum visible to human eyes was the infrared portion. William Herschel, also known as the discoverer of the first planet to be revealed in modern times, Uranus, was responsible for slowing rays of light with a prism, and redirecting the light rays into heat-absorbing bulbs. He found that the “caloric rays” were most intense beyond the red portion, producing the highest absorption temperatures yet the rays could be refracted and reflected like visible light.

In Germany, Johann Ritter, learning about Herschel’s discovery, attempted to identify the complementary radiation beyond the violet region of the visible spectrum by exposing silver chloride crystals to refracted sunlight. Ritter originally called this new discovery “chemical radiation” but in time, this radiation became known as ultraviolet (beyond the violet).

James Clerk Maxwell created the Electromagnetic Theory, which served to unify the initially disparate fields of electricity and magnetism utilizing Maxwell’s Equations. His work suggested that light itself was one of several types of electromagnetic waves, all traveling at the velocity of light, c .

The next portion of the electromagnetic spectrum to be identified was located in the low energy region. In 1887, German physicist Heinrich Hertz added very long wavelength radio waves to the spectrum, but his research did not pursue applications of this technology as he felt that there was no practical use for it. It was left to Nicola Tesla and Guglielmo Marconi to find ways to utilize “wireless telegraphy” for the public.

The discovery of X-rays followed soon thereafter. Wilhelm Roentgen, a Bavarian physicist, studied the passage of cathode rays from an induction coil through a glass tube that had been partially evacuated. He noticed that these rays when projected upon a fluorescent screen caused it to glow. Roentgen also found that these rays penetrated skin and could cast an image of the bones within upon on photographic plate. The first X-ray image published was that of Frau Roentgen’s hand.

Interest in uncovering new elements and new phenomena such as X-rays was all consuming as the end of the nineteenth century approached. Henri Becquerel, in Paris, discovered that uranium salts were the source of radioactivity. Another Parisian researcher, Paul Villard, also studied radioactive sources and in 1900, established that certain radioactive materials emitted what become known as gamma rays, high-energy radiation with even shorter wavelengths than X-rays.

By the early twentieth century, most of the regions of the electromagnetic spectrum had been explored and applications of the different manifestations such as radio waves and X-rays had been explored. One region, however remained largely unexamined until the 1940s. This type of electromagnetic radiation, initially known as ultrashort radio waves, consisting of wavelengths in the 1 *meter* – one millimeter range, was used to send radar signals to establish distance. Use of this application expanded during World War II. Engineers at the Raytheon corporation building the vacuum tubes for military uses noticed that the heat emitted by the tubes could be used to warm their hands in the winter months. The idea of incorporating this technology to construct microwave ovens was

implemented by Raytheon engineers John Spencer and Marvin Bock. The ubiquitous modern cell phones also utilize microwave radiation to send signals, at an intensity level too low to result in thermal heating.

New applications are continually being added to the complement of uses for the different ranges of wavelengths and frequencies encompassed by the electromagnetic radiation, shedding “light” on previously unexplored areas of potential technology.

5.6 Assessment for Chapter 5

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER

6

The Quantum Mechanical Model of the Atom C-SE-TE

Chapter Outline

- 6.1 THE DUAL NATURE OF LIGHT
- 6.2 CHARACTERISTICS OF MATTER
- 6.3 QUANTUM NUMBERS, ORBITALS, AND PROBABILITY PATTERNS
- 6.4 WORKSHEETS FOR CHAPTER 6
- 6.5 EXTRA READINGS FOR CHAPTER 6
- 6.6 ASSESSMENT FOR CHAPTER 6

Lessons and Number of Activities for Lessons

TABLE 6.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. The Dual Nature of Light	0	0	0	0
2. Characteristics of Matter	0	0	0	0
3. Quantum Numbers, Orbitals, and Probability Patterns	0	0	1	1

6.1 The Dual Nature of Light

Student Behavioral Objectives

The student will:

- name the model that replaced the Bohr model of the atom.
- explain the concept of wave-particle duality.
- solve problems involving the relationship between the frequency and the energy of a photon.

Timing, Standards, Activities

TABLE 6.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Dual Nature of Light</i>	1.0	1g, 1n

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for The Dual Nature of Light (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

6.2 Characteristics of Matter

Student Behavioral Objectives

The student will:

- describe a standing wave.
- state the Heisenberg uncertainty principle.

Timing, Standards, Activities

TABLE 6.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Characteristics of Matter</i>	2.0	None

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Characteristics of Matter (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

6.3 Quantum Numbers, Orbitals, and Probability Patterns

Student Behavioral Objectives

The student will:

- state the relationship between the principal quantum number (n), the number of orbitals, and the maximum number of electrons in a principal energy level.

Timing, Standards, Activities

TABLE 6.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Quantum Numbers, Orbitals, and Probability Patterns</i>	2.5	None

Activities for Lesson 3

Laboratory Activities

- None

Demonstrations

- None

Worksheets

- Quantum Numbers and Orbital Shapes Worksheet

Extra Readings

- Spin State and Nuclear Magnetic Resonance Imaging

Answers for Quantum Numbers, Orbitals, and Probability Patterns (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 6

This website provides an interactive tour of the atom.

- <http://ParticleAdventure.org/>

This website describes the double-slit experience and provides a simulation of the double-slit experiment.

- <http://www.colorado.edu/physics/2000/schroedinger/two-slit2.html>

A question and answer session on electrons behaving as waves.

- <http://www.colorado.edu/physics/2000/quantumzone/debroglie.html>

This website provides a video on quantum mechanics.

- <http://www.teachersdomain.org/resource/phy03.sci.phys.fund.quantum/>

A video discussing the relationship between spectral lines and electron transitions.

- <http://www.youtube.com/watch?v=fKYso97eJs4>

A short animation of s and p orbitals.

- <http://www.youtube.com/watch?v=VfBcfYR1VQo>

Another example of s, p and d electron orbitals.

- <http://www.youtube.com/watch?v=K-jNgq16jEY>

The following is a video on the quantum mechanical model of the atom.

- http://www.youtube.com/watch?v=IsA_oIXdF_8&feature=related

This video is a ChemStudy film called “Hydrogen Atom and Quantum Mechanics.” The film is somewhat dated but the information is accurate. The video also contains some data supporting quantum theory.

- <http://www.youtube.com/watch?v=80ZPe80fM9U>

Laboratory Activities for Chapter 6

Demonstrations for Chapter 6

6.4 Worksheets for Chapter 6

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Quantum Numbers and Orbital Shapes Worksheet

CK-12 Foundation Chemistry . . . Name _____ Date _____

TABLE 6.5:

Energy Level	Sub-Level	Number of Sub-Level Orbitals	Maximum Number of Electrons in Sub-Energy Level	Quantum Numbers for this Sub-Level
1	s	1	2	$n = 1, l = 0$
2	s	1	2	$n = 2, l = 0$
2	p	3	6	$n = 2, l = 1$
3	s	1	2	$n = 3, l = 0$
3	p	3	6	$n = 2, l = 1$
3	d	5	10	$n = 1, l = 2$
4	s	1	2	$n = 4, l = 0$
4	p	3	6	$n = 4, l = 1$
4	d	5	10	$n = 1, l = 2$
4	f	7	14	$n = 4, l = 3$
5	s	1	2	$n = 5, l = 0$
5	p	3	6	$n = 5, l = 1$
5	d	5	10	$n = 5, l = 2$
5	f	7	14	$n = 5, l = 3$
5	g	9	18	$n = 5, l = 4$
6	s	1	2	$n = 6, l = 0$
6	p	3	6	$n = 6, l = 1$
6	d	5	10	$n = 6, l = 2$
6	f	7	14	$n = 6, l = 3$
6	g	9	18	$n = 6, l = 4$
6	h	11	22	$n = 6, l = 5$

Mathematically, from Schrodinger's Equation, energy level 5 would have a fifth sub-level named *g*. It would have 9 orbitals and hold a maximum of 18 electrons. Similarly, energy level 6 would have this *g* sub-level and another sub-level named *h*. Sub-level *h* would have 11 orbitals and would hold a maximum of 22 electrons. This pattern would continue through all the larger energy levels. In terms of usefulness, however, we have no atoms that contain

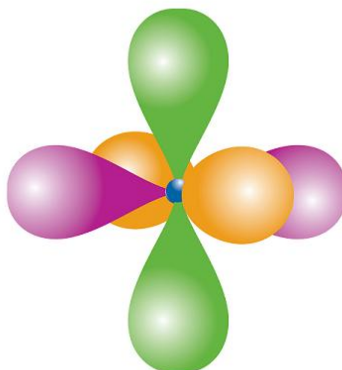
enough electrons to use the $5g, 6g, 6h, 7g, 7h$ sub-levels. The known atoms never use any energy sub-levels beyond $5f, 6f$, and $7f$. Therefore, in most listings of energy levels and sub-levels, energy levels 5, 6, and 7 will look exactly like energy level 4, with only s, p, d , and f sub-levels listed.

The probability patterns for these sub-levels are shown below.

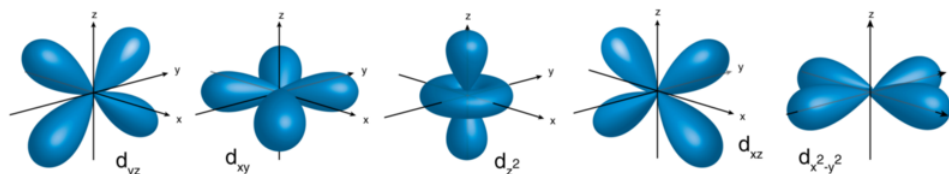
The s orbitals in every energy level are spherical.



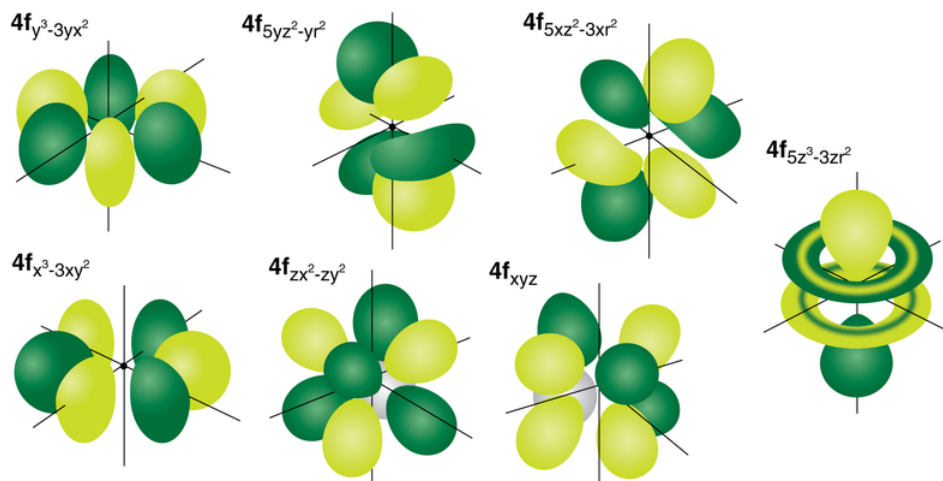
The three p orbitals in energy levels 2 – 7 are dumbbell shaped.



The five d orbitals in energy levels 3 – 7 are sometimes referred to as butterfly shaped.



The seven f orbitals in energy levels 4 – 7 are too complex to describe.



Exercises

True/False

- All sub-energy levels with $\ell = 1$, regardless of the principal energy level quantum number will have dumbbell shape.
 - True
 - False
- Theoretically, it is possible for a principal energy level to have n^2 sub-energy levels.
 - True
 - False
- It is impossible for an electron in an atom to have the quantum numbers $n = 3, \ell = 2, m_\ell = 3, m_s = +\frac{1}{2}$.
 - True
 - False

Multiple Choice

- How many sub-energy levels may be present if the principal quantum number is 3?
 - 1
 - 2
 - 3
 - 4
 - None of these.
- How many possible orbitals are there when $n = 3$?
 - 1
 - 3
 - 4
 - 5
 - 9
- How many electrons can be accommodated in the energy level for which $n = 3$?
 - 1
 - 3
 - 4
 - 5
 - 9

A. 2

B. 6

C. 8

D. 10

E. 18

7. How many atomic orbitals are present in the subshell for which $n = 3$ and $\ell = 2$?

A. 1

B. 3

C. 5

D. 7

E. 9

8. How many orbitals are present in the subshell for which $n = 5$ and $\ell = 4$?

A. 1

B. 3

C. 5

D. 7

E. 9

9. What is the shape of an orbital in the subshell for which $n = 3$ and $\ell = 0$?

A. spherical

B. dumbbell

C. butterfly or clover shaped

D. Could be any of these.

E. None of these.

10. What is the shape of an orbital in the subshell for which $n = 7$ and $\ell = 0$?

A. spherical

B. dumbbell

C. butterfly or clover shaped

D. Could be any of these.

E. None of these.

11. Which type of orbital is described by the quantum numbers $n = 2$, $\ell = 1$?

A. $2s$

B. $2p$

C. $2d$

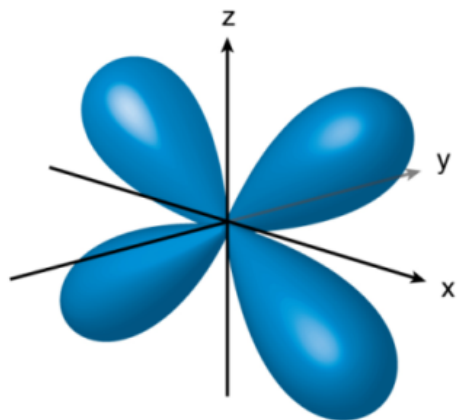
D. $2f$

E. None of these.

12. If the principal quantum number of an atomic orbital is 4, what are the possible values of ℓ ?

- A. 0, 1, 2, 3, 4
- B. 1, 2, 3, 4
- C. 0, 1, 2, 3
- D. 0, 1, 2
- E. None of these.

Use the image below to answer questions 13, 14, and 15.



13. Identify the image above as an s -orbital, p -orbital, d -orbital, f -orbital or none of these.

- A. s
- B. p
- C. d
- D. f
- E. None of these.

14. What is the ℓ value for the type of orbital pictured above?

- A. 0
- B. 1
- C. 2
- D. 3
- E. 4

15. Will an orbital of the shape pictured above be found in the $n = 2$ energy level?

- A. Yes
- B. No

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

6.5 Extra Readings for Chapter 6

Spin State and Nuclear Magnetic Resonance Imaging

While the concepts in the study of quantum mechanics may seem elusive to some students, many may be more familiar with one application based on the spin states of the hydrogen atom: Magnetic Resonance Imaging. MRI technology is based on the differing spin states of the hydrogen atom, usually those associated with biological water and fat molecules, and their interaction with strong magnetic fields and radiofrequency waves.

Atoms with an odd number of protons or neutrons in their nucleus possess an intrinsic spin that is quantized. There are two magnetic spin states for the hydrogen nucleus, which can be described as the opposite types of physical spinning: clockwise ($+\frac{1}{2}$) and counterclockwise spinning ($-\frac{1}{2}$). Under normal circumstances, a collection of hydrogen nuclei would display random alignment and both spin states would be equal in energy (degenerate). In the presence of an external magnetic field, however, the hydrogen nuclei can orient with or against the magnetic field, with more nuclei lining up with the magnetic field at a lower energy value. Those nuclei with spins opposing the magnetic field would then be higher in energy by a value of E . If that precise amount of energy is added to the system, in the form of radio waves applied at right angles to the magnetic field, it can cause the lower energy nuclei to absorb and perform a “spin flip” to the higher energy configuration. The radio frequency must match or be in resonance with the nucleus’ natural spin. As the “flipped” nuclei gradually relax and realign with the magnetic field, they resume their lower energy spin states and release the absorbed energy. The relaxation rates are a function of the interaction of the nucleus and its physical environment. Incorporation of adjustable magnetic fields can generate a map of very slight difference in resonance frequencies, and thus produce the magnetic resonance image, allowing practitioners to construct images of body tissues with a three-dimensional quality.

The introduction of Magnetic Resonance Imaging has provided a diagnostic revolution in the medical world. Unlike X-rays and CT (computer tomography) scans, no radiation is utilized to produce the image. The key limitations to MRI include eliminating any interference of the magnetic field with metals, thus MRI patients may not have pacemakers, insulin pumps or prosthetic implants. Enhancement of the distinction between normal and diseased tissue is often needed as well, and provided by the introduction of contrast agents. These are usually molecules containing paramagnetic ions, such as Gd^{2+} , which has seven unpaired electrons. These agents are administered intravenously and they serve to highlight visualization of tissue by shortening the relaxation time of the nuclei. Other paramagnetic agents, such as iron oxide and manganese agents are also used for certain applications.

MRI examinations are performed by placing the patient inside the bore of a very large magnet. Other obstacles to this diagnostic tool is that some patients experience claustrophobic anxiety inside the magnet, while others have found the time needed and noise incurred during the data acquisition to be uncomfortable. Despite these hindrances, Magnetic Resonance Imaging has emerged to become an impressive tool for the practice of modern medicine.

6.6 Assessment for Chapter 6

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 7 The Electron Configuration of Atoms C-SE-TE

Chapter Outline

- 7.1 ELECTRON ARRANGEMENT
- 7.2 VALENCE ELECTRONS
- 7.3 WORKSHEETS FOR CHAPTER 7
- 7.4 EXTRA READINGS FOR CHAPTER 7
- 7.5 ASSESSMENT FOR CHAPTER 7

Lessons and Number of Activities for Lessons

TABLE 7.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Electron Arrangement	0	0	2	0
2. Valence Electrons	1	0	2	1

7.1 Electron Arrangement

Student Behavioral Objectives

The student will:

- draw the orbital representation for selected atoms.
- write the electron configuration code for selected atoms.
- identify the principal, angular momentum, magnetic, and spin quantum numbers.
- identify the four quantum numbers for indicated electrons.
- identify the energy level, sub-energy level, orbital, and spin for an electron given the four quantum numbers for the electron

Timing, Standards, Activities

TABLE 7.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Electron Arrangement</i>	3.0	1g

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Orbital Configuration Worksheet
2. Quantum Numbers Worksheet

Extra Readings

1. None

Answers for Electron Arrangement (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

7.2 Valence Electrons

Student Behavioral Objectives

The student will:

- define valence electrons.
- indicate the number of valence electrons for selected atoms.
- draw Lewis dot diagrams for selected atoms.

Timing, Standards, Activities

TABLE 7.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Valence Electrons</i>	1.0	2e

Activities for Lesson 2

Laboratory Activities

1. Diamagnetic Levitation

Demonstrations

1. None

Worksheets

1. Electron Configuration Worksheet
2. Electron Dot Formulas Worksheet

Extra Readings

1. Orbital Filling Order Exceptions

Answers for Valence Electrons (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 7

This website has a lesson on atomic structure.

- <http://www.chemtutor.com/struct.htm>

This website gives the apartment analogy of electron configuration.

- <http://kaffee.netfirms.com/Science/activities/Chem/Activity.Electron.Configuration.html>

This video discusses the physical interpretation of quantum numbers.

- <http://www.youtube.com/watch?v=e9N2h8c6dE4>

This video explains the four quantum numbers which give the "address" for an electron in an atom.

- <http://www.youtube.com/watch?v=63u7A2NiyU>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 3-6 is on electron configuration.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson36.htm>

These two websites examine the discovery of the spin quantum number.

- <http://www.lorentz.leidenuniv.nl/history/spin/goudsmit.html>
- http://www.ethbib.ethz.ch/exhibit/pauli/elektronenspin_e.html

This video provides an introduction to the electron configuration of atoms.

- <http://www.youtube.com/watch?v=fv-YeI4hcQ4>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 3-8 is on Lewis dot diagrams.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson38.htm>

Laboratory Activities for Chapter 7

The Electron Spin Quantum Number Lab – Diamagnetic Levitation

In this chapter, we discussed the differences between diamagnetic and paramagnetic material. In this lab, we suggest demonstrating how a diamagnetic material is slightly repelled by a magnetic field.

Materials:

A thin sheet of pyrolytic graphite

A set of magnets

A complete diamagnetic levitation kit can be bought from certain school supply websites. A reasonably priced kit can be found at <http://www.teachersource.com/ElectricityAndMagnetism/ElectricityAndMagnetism/DiamagneticLevitationKit.aspx>

Method:

Place the graphite above the magnets and watch it float!

Demonstrations for Chapter 7

7.3 Worksheets for Chapter 7

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Orbital Configuration Worksheet

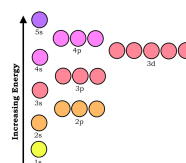
CK-12 Foundation Chemistry

Name _____ Date _____

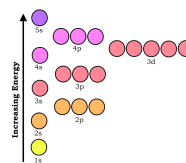
TABLE 7.4: Draw the Orbital Configuration for these Atoms

Symbol
Mg

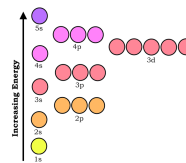
Orbital Diagram



P



Ge



Kr

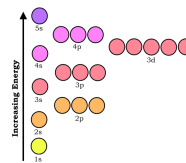
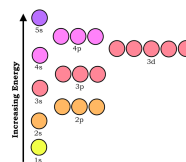


TABLE 7.4: (continued)

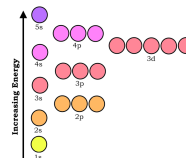
Symbol

Orbital Diagram

O



F



Pb

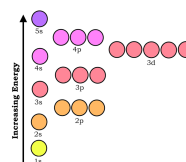


TABLE 7.5: Write the Electron Configuration Code for these Atoms

Atom

Electron Configuration Code

V

 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

Mg

P

Ge

Kr

O

F

Pb

Quantum Numbers Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

1. Which quantum number indicates the electron's energy level?
2. Which quantum number indicates the electron's sub-energy level?
3. Which quantum number indicates the electron's orbital within the sub-energy level?
4. Which quantum number indicates the electron's spin?
5. What is the lowest energy level that has a d sub-level?
6. What is the total number of electrons that can exist in the 3^{rd} energy level?

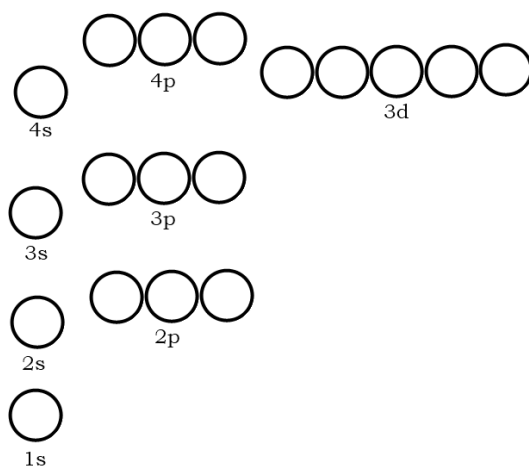
7. Which sub-energy level is indicated by $\ell = 1$?
8. which sub-energy level is indicated by $\ell = 2$?
9. What is the maximum number of electrons that can be held in an f sub-energy level?
10. What does it mean for an electron to be “excited”?
11. What are the n and ℓ quantum numbers for the last electron in bromine?
12. What are the n and ℓ quantum numbers for the last electron in iron?
13. What are the n and ℓ quantum numbers for the electron in hydrogen?
14. The three electrons in the $2p$ sub-energy level of nitrogen have the n and ℓ quantum numbers. What are the m_ℓ quantum numbers for each of these three electrons?
15. What is the basic tenet of the quantum theory?
16. Why are the quantum numbers $n = 2, \ell = 2, m_\ell = 2, s = \frac{1}{2}$, not an acceptable set of quantum numbers for an electron?
17. Sketch a picture of the $2s$ sub-energy level showing any nodes present.
18. Give the full set of quantum numbers for each of the electrons in a helium atom.
19. What maximum number of electrons in an atom can have the quantum numbers $n = 2, \ell = 1$?
20. What maximum number of electrons in an atom can have the quantum numbers $n = 3, \ell = 3$?

Electron Configuration Worksheet

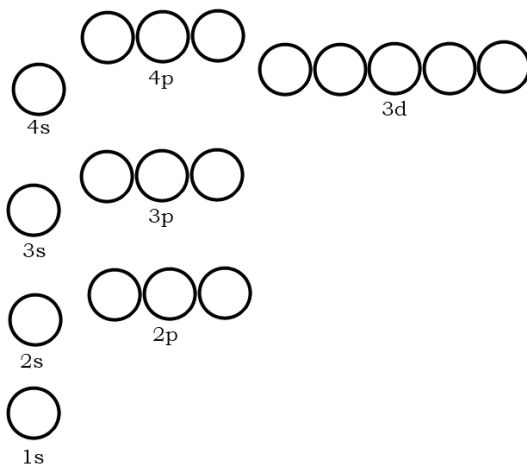
CK-12 Foundation Chemistry

Name _____ Date _____

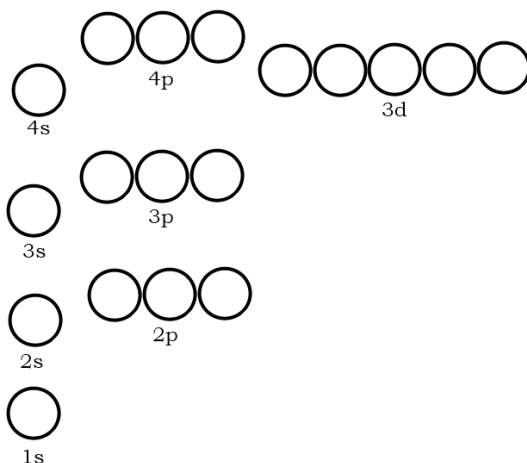
1. Fill in the orbital electron representation for phosphorus.



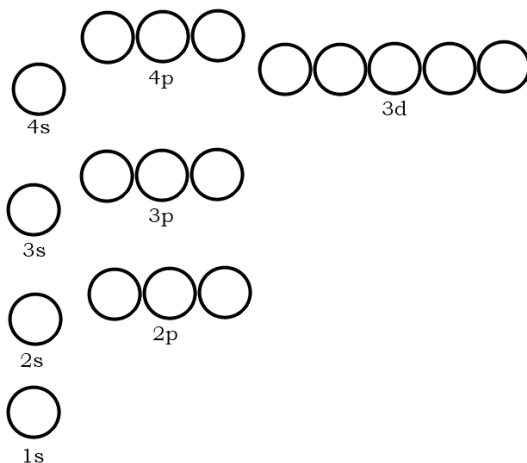
2. Fill in the electron orbital configuration for cobalt.



3. Fill in the electron orbital configuration for bromine.



4. Draw the orbital representation for the electron configuration of calcium.

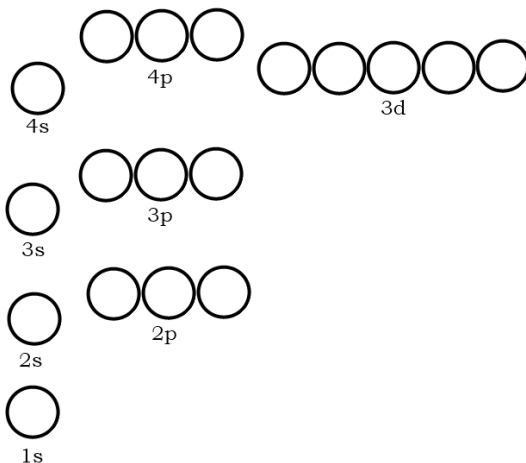


5. Write the electron configuration code for phosphorus.

6. How many valence electrons does phosphorus have?

7. Write the electron configuration code for cobalt.

8. How many valence electrons does cobalt have?
9. Write the electron configuration code for bromine.
10. How many valence electrons does bromine have?
11. Write the electron configuration code for calcium.
12. How many valence electrons does calcium have?
13. How many valence electrons does tellurium have?
14. What will be the outer energy level electron configuration for element #118?
15. Draw the orbital representation of the electron configuration for silicon.



16. How many valence electrons does silicon have?

Electron Dot Formulas Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

1. Draw the electron-dot formula for phosphorus.
2. Draw the electron-dot formula for bromine.
3. Draw the electron-dot formula for cobalt.
4. Draw the electron-dot formula for calcium.
5. Draw the electron-dot formula for silicon.
6. Draw the electron-dot formula for element #118.

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

7.4 Extra Readings for Chapter 7

Orbital Filling Order Exceptions

In assembling the electron configurations for many-electron atoms, one tool that students find valuable is the diagonal rule. This rule provides a guideline that is readily remembered and easily followed to produce accurate electron configurations for even complicated d - and f -block atoms. One confusing consequence of the diagonal rule is the order of filling the $4s$ and $3d$ subshells.

When these orbitals are filled, they are very close in energy. Though as the electrons begin to occupy the empty orbitals, the $4s$ level is slightly lower in energy than the $3d$, thus it is filled first. On the other hand, when both are occupied with electrons, the $4d$ orbital becomes higher in energy. Thus, in the case that both of these filled levels are composed of valence electrons, the $4s$ level loses its valence electrons before the $3d$ level.

The preferential filling of the $4s$ orbital can also be explained by means of the electron penetration effect. Due to the spherical shape of the s orbital probability density distribution, the likelihood that an electron is found closer to the nucleus is greater than the multi-lobed $3d$ orbitals.

The similarity in the energy levels of the $4s$ and $3d$ orbitals also leads to another interesting consequence. In the electron configuration of the neutral Chromium atom with 24 electrons, the diagonal rule suggests an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$. The actual electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$, where due to the similarity in energy between the $4s$ and $3d$ orbitals, one electron transfers from the $4s$ to the $3d$ orbital. The net effect of this exchange yields half-filled $4s$ and $3d$ orbitals, and therefore can be justified in terms of generating additional stability. This is also the case for neutral copper atoms, with 29 electrons and a putative electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$. Again as in the example of chromium, an electron transfer occurs, shifting one electron from the $4s$ orbital to the $3d$ orbital. For copper, the $4s$ orbital is now half-filled but added stability is attained by completing the $3d$ subshell.

The stability afforded to half-filled orbitals is also noted among the f -block elements. For example, the electron configuration for Europium (atomic number 63) is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^7$ whereas the next atom, Gadolinium, with atomic number 64, has the additional electron added to the $5d$ orbital in order to maintain the half-filled stability of the $4f^7$ configuration. The electron configuration for Gadolinium is therefore $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^7 5d^1$.

The unusual stability of half-filled orbitals can be explained in terms of the disruption afforded by the addition of another electron to this configuration. After the orbital is half-filled, the next additional electron must pair up with another electron, increasing the spin-spin interaction energy and destabilizing the configuration.

7.5 Assessment for Chapter 7

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER

8

Electron Configuration and the Periodic Table C-SE-TE

Chapter Outline

- 8.1 MENDELEEV'S PERIODIC TABLE
- 8.2 FAMILIES AND PERIODS OF THE PERIODIC TABLE
- 8.3 WORKSHEETS FOR CHAPTER 8
- 8.4 EXTRA READING FOR CHAPTER 8
- 8.5 ASSESSMENT FOR CHAPTER 8

Lessons and Number of Activities for Lessons

TABLE 8.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Mendeleev's Periodic Table	0	0	0	0
2. Families and Periods of the Periodic Table	0	0	1	1

8.1 Mendeleev's Periodic Table

Student Behavioral Objectives

The student will:

- identify the person credited for organizing the periodic table.
- state the basis for the organization of Mendeleev's periodic table.

Timing, Standards, Activities

TABLE 8.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Mendeleev's Periodic Table</i>	1.0	None

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Mendeleev's Periodic Table (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

8.2 Families and Periods of the Periodic Table

Student Behavioral Objectives

The student will:

- identify groups in the periodic table.
- state the number of valence electrons for each A group in the periodic table.
- explain the relationship between the chemical behavior of families in the periodic table and their electron configuration.
- identify periods in the periodic table.
- describe the similarities among elements in the same period in the periodic table.

Timing, Standards, Activities

TABLE 8.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Families and Periods of the Periodic Table</i>	1.5	1b, 1c, 1f

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. The Periodic Table and Electron Configuration Worksheet

Extra Readings

1. The Upper Limit of the Periodic Table

Answers for Families and Periods of the Periodic Table (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 8

The following websites explore the history behind the periodic table.

- <http://www.wou.edu/las/physci/ch412/perhist.htm>
- <http://www.aip.org/history/curie/periodic.htm>

The following video is an introduction to the electronic organization of the periodic table.

- <http://www.youtube.com/watch?v=5MMWpeJ5dn4>

This website provides an interactive periodic table with basic information about each element.

- <http://www.chemicalelements.com/http://www.chemicalelements.com/> <http://www.chemicalelements.com/>

This video describes the organization and family properties within the periodic table.

- <http://www.youtube.com/watch?v=rD0maRGQhU8>

This website provides a dynamic periodic table that also provides information about the orbitals and electron configurations of the elements.

- <http://www.dayah.com/periodic>

This website reviews the different groups in the periodic table.

- <http://www.chemtutor.com/perich.htm>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 3-4 is on the periodic table.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson34.htm>

This video is a ChemStudy film called “Transuranium Elements.” The film is somewhat dated but the information is accurate.

- <http://www.youtube.com/watch?v=SpQSEN1n4HE>

This video is a ChemStudy film called “Inert Gas Compounds.” The film is somewhat dated but the information is accurate.

- <http://www.youtube.com/watch?v=QASDy4bLQxs>

Laboratory Activities for Chapter 8

Demonstrations for Chapter 8

8.3 Worksheets for Chapter 8

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

The Periodic Table and Electron Configuration Worksheet

When Mendeleev organized the periodic table, he placed the elements in vertical columns according to their chemical behavior. That is, elements were placed in the same vertical columns because they behaved similarly in chemical reactions. All the alkali metals (*Li, Na, K, Rb, Cs*) react with water to produce heat, hydrogen gas, and the metal hydroxide in solution. Essentially, the only difference in the reactions is that the larger alkali metals react faster than the smaller ones. The vertical columns of elements are frequently referred to chemical “families” because of their similar chemical characteristics.

When quantum theory generated electron configurations which demonstrated that the elements in the same family have the same outer energy level electron configuration, the reason these elements behaved similarly became clear. Since chemical behavior is determined by outer energy level electron configuration, it was clear that elements that behaved similarly should have similar electron configuration.

TABLE 8.4: The Electron Configuration of Family 1A Elements

Element	Electron Configuration
Li	$1s^2 2s^1$
Na	$1s^2 2s^2 2p^6 3s^1$
K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$
Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$

TABLE 8.5: The Electron Configuration of Family 7A Elements

Element	Electron Configuration
F	$1s^2 2s^2 2p^5$
Cl	$1s^2 2s^2 2p^6 3s^2 3p^5$
Br	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
I	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$

Exercises

- If the outermost energy level electron configuration of an atom is $ns^2 np^1$,
 - to which family does it belong? _____
 - is the atom a metal, metalloid, non-metal, or a noble gas? _____
 - how many valence electrons does it have? _____
- If the outermost energy level electron configuration of an atom is $ns^2 np^4$,
 - to which family does it belong? _____
 - is the atom a metal, metalloid, non-metal, or a noble gas? _____

- c. how many valence electrons does it have? _____
3. If the outermost energy level electron configuration of an atom is ns^2np^6 ,
- to which family does it belong? _____
 - is the atom a metal, metalloid, non-metal, or a noble gas? _____
 - how many valence electrons does it have? _____
4. The electron configuration of an element is $[Ar]4s^23d^3$.
- What is the identity of the element? _____
 - In what period does the element belong? _____
 - In what group does the element belong? _____
 - Is the element a main group element, a transition element, a lanthanide, or an actinide? _____
5. Write the electron configuration of only the outermost energy level for an element that is in family 5A of the fifth period of the periodic table. _____
6. Write the electron configuration of only the outermost energy level for an element that is in family 8A of the third period of the periodic table. _____

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.**

8.4 Extra Reading for Chapter 8

The Upper Limit of the Periodic Table

The Periodic Table has been acknowledged as one of the most influential keys to understanding modern chemistry. A wealth of information is organized into a readily interpretable array of essential atomic data. Since the days of Dmitri Mendeleev, who is credited with arranging our modern periodic table on the basis of physical similarities, atomic physicists have drastically extended the number of elements by the preparation of artificial elements. These are atoms not found naturally on Earth due to radioactive decay instability but have been created synthetically by atomic bombardment and collisions.

The very first synthetic element was the result of many years of searching for the elusive missing element to be inserted between molybdenum and ruthenium, an omission noted and a space left open by Mendeleev. Many efforts claiming to have identified element 43 were made but not substantiated. Conclusive evidence for the production of a new element was made by Emilio Segré and Carlo Perrier in 1937 after they collided molybdenum atoms with the heavy isotope of hydrogen known as deuterium. Later trace amounts of technetium were identified among the decay products of uranium fission. The name technetium was chosen from the Greek word for artificial.

The next synthetic element, 61, promethium, was produced by a similar method. Jakob Marinsky and Larry Glendenin at MIT bombarded neodymium atoms with neutrons obtained as byproducts of uranium decay. Their 1946 announcement named the new element after the mythological Prometheus, who, according to legend was responsible for bringing fire to mankind.

The decade of the 1940's also marked the creation of the first trans-uranium element. Neptunium was the result of Berkeley scientists Edwin McMillan and Philip Abelson colliding uranium with neutrons as was the concurrent production of element 94, named plutonium in the sequence correlating with the modern group of solar system planets. One name suggested for element 94 was "extremium" offering the proposition that this artificially produced element was the upper limit or heaviest possible atom.

Since that time, the quest for producing super-heavy elements has continued with the question of where and when that upper limit, if it exists, will be reached. Currently, (2009) the as-yet unnamed Element 118, a member of the noble gas family, maintains its status as the heaviest element. Three atoms of element 118 were reportedly created by fusing californium atoms with calcium atoms in 2006 at Lawrence Livermore Laboratory. In the last year, claims suggesting the existence of Element 122 have also been reported but as yet, experimental replications have failed to reproduce this evidence.

Is there an upper limit to the periodic table? The intrinsic instability with respect to nuclear decay appears to limit the production of elements with atomic numbers greater than that of uranium. Most of the trans-uranium elements have extremely short half-lives and very limited production quantities. Attempting to load the tiny atomic nucleus with 100+ protons appears to provide a barrier that may have reached its synthetic limit.

8.5 Assessment for Chapter 8

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 9

Chemical Periodicity C-SE-TE

Chapter Outline

- 9.1 THE PERIODIC TABLE
- 9.2 PERIODIC TRENDS IN ATOMIC SIZE
- 9.3 PERIODIC TRENDS IN IONIC SIZE
- 9.4 PERIODIC TRENDS IN IONIZATION ENERGY
- 9.5 PERIODIC TRENDS IN ELECTRONEGATIVITY
- 9.6 PERIODIC TRENDS IN ELECTRON AFFINITY
- 9.7 WORKSHEETS FOR CHAPTER 9
- 9.8 EXTRA READING FOR CHAPTER 9
- 9.9 ASSESSMENT FOR CHAPTER 9

Lessons and Number of Activities for Lessons

TABLE 9.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. The Periodic Table	0	0	0	0
2. Periodic Trends in Atomic Size	0	0	0	0
3. Periodic Trends in Ionic Size	0	0	0	0
4. Periodic Trends in Ionization Energy	0	0	0	0
5. Periodic Trends in Electronegativity	1	0	0	0
6. Periodic Trends in Electron Affinity	0	0	1	1

9.1 The Periodic Table

Student Behavioral Objectives

The student will:

- explain the Periodic Law.
- explain how the periodic table is arranged in terms of the Periodic Law.
- describe the differences among metals, nonmetals, and metalloids.
- draw a rough sketch of the modern periodic table, and indicate the portion of the table occupied by metals, nonmetals, and metalloids.
- identify the stair-step line that separates the metallic elements from the nonmetallic ones.
- compare the properties of metals, nonmetals, and metalloids.

Timing, Standards, Activities

TABLE 9.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Periodic Table</i>	0.5	1c

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for The Periodic Table (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

9.2 Periodic Trends in Atomic Size

Student Behavioral Objectives

The student will:

- define atomic radius.
- state the boundary issue with atomic size.
- describe measurement methods for atomic size.
- define the shielding effect.
- describe the factors that determine the trend of atomic size.
- describe the general trend in atomic size for groups and for periods.
- describe the trend of atomic radii in the rows in the periodic table.
- describe how the trend of atomic radii works for transition metals.
- use the general trends to predict the relative sizes of atoms.

Timing, Standards, Activities

TABLE 9.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Periodic Trends in Atomic Size</i>	1.0	1c

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Periodic Trends in Atomic Size (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org

to request sample answers.

9.3 Periodic Trends in Ionic Size

Student Behavioral Objectives

The student will:

- describe the factors that determine the trend of ionic size.
- describe the general trend in ionic size for groups and for periods.

Timing, Standards, Activities

TABLE 9.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Periodic Trends in Ionic Size</i>	1.0	1c

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Periodic Trends in Atomic Size (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

9.4 Periodic Trends in Ionization Energy

Student Behavioral Objectives

The student will:

- define ionization energy.
- describe the trend that exists in the periodic table for ionization energy.
- describe the ionic size trend that exists when elements lose one electron.

Timing, Standards, Activities

TABLE 9.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Periodic Trends in Ionization Energy</i>	1.0	1c

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Periodic Trends in Ionization Energy (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

9.5 Periodic Trends in Electronegativity

Student Behavioral Objectives

The student will:

- define electronegativity.
- describe the trend for electronegativity on the periodic table.

Timing, Standards, Activities

TABLE 9.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Periodic Trends in Electronegativity</i>	0.5	1c

Activities for Lesson 5

Laboratory Activities

1. Paramagnetic Behavior of Manganese Compounds

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Periodic Trends in Electronegativity (L5) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

9.6 Periodic Trends in Electron Affinity

Student Behavioral Objectives

The student will:

- define electron affinity.
- describe the trend for electron affinity on the periodic table.

Timing, Standards, Activities

TABLE 9.7: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Periodic Trends in Electron Affinity</i>	1.0	1g

Activities for Lesson 6

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Trends in the Periodic Table Worksheet

Extra Readings

1. The Chemistry of Glass

Answers for Periodic Trends in Electron Affinity (L6) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 9

Key information about the chemical elements.

- <http://www.webelements.com/>

This video is on the periodic table.

- <http://www.youtube.com/watch?v=1geccHiylcU&feature=fvw>

A video introducing the electronic organization of the periodic table.

- <http://www.youtube.com/watch?v=35cWAXtHUGw>

This video discusses the ionization energy trends in the periodic table.

- <http://www.youtube.com/watch?v=xE9YOBXdTSo>

This video shows the relationships between atomic size, ionization energy, and electron affinity.

- <http://www.youtube.com/watch?v=iCwYjpl8eeY&feature=channel>

This pdf document reviews the causes and relationships of the trends in atomic size, ionization energy, electronegativity, and electron affinity.

- <http://www.oakland.k12.mi.us/Portals/0/Learning/PeriodicTable.pdf>

Laboratory Activities for Chapter 9

Teacher's Pages for Paramagnetism Lab

Lab Notes Capsules for this lab can be pre-loaded with the manganese salts. This will save time with a large class, and is safer. MnO_2 can stain the skin and is a mild oxidizing agent. MnO_4^- is a strong oxidizing agent and can cause burns.

If you have trouble with the procedure involving the balance due to balance sensitivity, you may try placing the gel capsule onto a smooth surface and bring a strong magnet towards it. The capsule may exhibit an attraction, repulsion, or no reaction. The paramagnetism of the material may be gauged ordinally.

Answers to Pre-Lab Questions

1. Paramagnetism is a weak attraction of a substance for a magnetic field due to the presence of unpaired electrons. The unpaired electrons each exert a weak magnetic field due to spin, and the spins are additive when unpaired.
2. Paramagnetism is a weak attraction for a magnetic field, while diamagnetism is a weak repulsion to a magnetic field due to opposing electron spins within an orbital.
3. The presence of d sublevel in transition metals gives nine opportunities for the sublevel to have an unpaired electron, not accounting for d^4 or d^9 promotion. This is more than the p sublevel or the s sublevel.

Lab - Paramagnetic Behavior of Manganese Compounds

Background Information

When we construct electron configurations for the atoms and ions in the periodic table, one of the cardinal rules that chemists observe is the pairing of electrons in orbitals whenever possible. Substances with paired electrons display diamagnetism, a phenomenon that causes that material to be slightly repelled by a magnetic field. Atoms with unpaired electrons, known as paramagnetic substances, are weakly attracted to magnetic fields. One interesting example of a molecule that displays paramagnetism is the oxygen diatomic molecule. While a Lewis dot structure can be constructed pairing all twelve valence electrons into an oxygen – oxygen double bond with two pairs of non-bonded electrons on each oxygen atom, the molecule displays paramagnetism when poured between the poles of a strong magnet, attracted to the magnetic field.

Pre-Lab Questions

1. Explain what is meant by the term paramagnetic.
2. Explain how paramagnetism differs from diamagnetism.
3. Why might paramagnetic transition metal compounds be more prevalent than those of main group elements?

Purpose

The purpose of this experiment is to demonstrate the effects of a magnetic field on a paramagnetic material.

Apparatus and Materials

- MnO_2
- Mn_2O_3
- $KMnO_4$
- Styrofoam cups
- Magnets
- Empty gel capsules

Safety Issues

Manganese oxide is a strong oxidant and can irritate tissues. Potassium permanganate is also a powerful oxidizing agent and can irritate skin. The use of eye protection is strongly recommended.

Procedure

Place about 1.0 gram of each of the materials to be tested in empty gel capsules. The capsules are then placed on the base of an inverted Styrofoam cup. A second Styrofoam cup is then placed beneath the first cup. The entire assembly is then placed on a top loading balance. Obtain the mass of the cup and gel capsule assembly. Now bring the magnet to the cup/capsule assembly and reassess the mass. Note the decrease in mass as the magnet attracts the capsule contents.

TABLE 9.8: Data

	With Magnet	Without Magnet
Mass of the MnO_2 /capsule/cup assembly		
Mass of the Mn_2O_3 /capsule/cup assembly		

TABLE 9.8: (continued)

Mass of the $KMnO_4$ /capsule/cup assembly	With Magnet	Without Magnet
<hr/>		

Post-Lab Questions

1. Which of the samples displayed the largest mass difference?
2. How many unpaired electrons does Manganese possess in MnO_2 ?

In Mn_2O_3 ?

In $KMnO_4$?

Demonstrations for Chapter 9

9.7 Worksheets for Chapter 9

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Trends in the Periodic Table Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

1. Which atom is larger, by volume, hydrogen or helium?
2. What is the smallest atom, by volume, in the third period?
3. Describe the relationship between atomic volume and ionization energy.
4. Which atom has the greatest electron affinity?
5. What is the most stable number of electrons for an atom's outermost energy level?
6. Which is larger in volume, oxygen or sulfur?
7. Which is chemically more reactive, potassium or cesium?
8. Which is chemically more reactive, oxygen or sulfur?
9. Which atom in period 3 has the greatest electron affinity?
10. Which atom in period 3 has the largest volume?
11. Which atom has greater ionization energy, aluminum or gallium?
12. Which atom has greater second ionization energy, potassium or calcium?
13. What is the outer energy level electron configuration of a noble gas?
14. Which atom in period 3 has the lowest ionization energy?
15. Explain why fluorine, even though it is larger than neon, has a greater electron affinity.

Answers to Worksheets

- **The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.**

9.8 Extra Reading for Chapter 9

The Chemistry of Glass

In the world of the twenty-first century, it is difficult to imagine our day-to-day existence without glass. As a transparent material with great resistance to corrosive substances, it has many uses from container duty, to its architectural impact as windows, to its role in optical devices. Its versatility places glass in an indispensable position in our array of material choices. Glass is chemically defined as a member of a group of compounds that solidify from the molten state without crystallization.

The original formulation of glass may have been based on observation of the naturally occurring opaque black glass known as obsidian, produced when volcanic lava is cooled abruptly by contact with water. Whether it was inspired or the result of a fortuitous accident, the first glasses originated in the Near East about 3000 B.C. Instructions for glassmaking have been identified on Mesopotamian cuneiform tablets. Their basic “recipe” included the same three main ingredients utilized to create modern glass formulations: formers, flux and stabilizers. A former is a material that forms the basis upon which the rest of the formulation is set; in many glasses, old and new, the former is silica, SiO_2 , or sand. The flux is the substance added in a minor quantity, mainly to lower the melting temperature of the mixture. Alkalis such as soda (sodium carbonate) and potash (potassium carbonate) have long been employed in this capacity. Lastly, the stabilizer, such as lime, CaO , calcium oxide, strengthens the glass and also adds water resistance.

Various inorganic materials have been added to this classic formulation since antiquity and their incorporation has been shown to impart novel characteristics to the glass produced. One of the earliest substitutions seen in glass formulations include the addition of lead oxide as a flux material. Lead glass, which may have first been used in Han dynasty China, was shaped into artificial gemstones, and later for lead crystal stemware. Lead glass has a lower melting temperature and a reputation for brilliance and sparkle due to its high refractive index. Lead glassware is also known for its ability to “ring” when struck, that distinguishes it from ordinary silicate glass.

The addition of small amounts of various metal oxide salts to the basic glass formula produce “stained” glass, renowned in its use in churches and cathedrals. Cobalt oxide is responsible for the vivid blue coloration, red from gold salts and copper oxide imparts a brilliant green hue.

The use of borax (boric oxide, B_2O_3) in place of soda and lime was a nineteenth century innovation attributed to Otto Schott, a German glassmaker, who originally called this new material “Duran”. Borosilicate glass has a higher melting temperature and a much greater thermal stability. Its modern commercial name “Pyrex”, is well known both to cooking enthusiasts and laboratory scientists.

The underlying chemical rationale for the differing properties of these disparate glass formulations may be due to atomic size mismatches between the various components. Since glass is not a crystalline substance, without a regular, repeating microscopic structure, it is better represented as a disordered network with defects, or “empty spaces” in the network. The presence of atoms with variously sized atomic radii in these defect regions, can then alter the macroscopic characteristics of the glass. For example, the substitution of the smaller boron atom in place of larger alkali metals may provide more efficient silica packing, and possibly account for the enhanced thermal stability of Pyrex. Lead ions, more comparably sized to the alkali ions, present a glass product that has similar melting characteristics.

What does the future hold for new glasses? Silicate fiber optics play a vital role in modern telecommunications. Heat resistant glasses are employed on the exterior of spacecraft for protection upon re-entry into Earth’s atmosphere.

Smart glass windows can control the amount of incoming solar radiation, and there is research into “self-cleaning” window glass. Chicago’s Sears Tower recently installed a glass observation deck flooring with load-bearing tempered glass. The material that caught the eye and imagination of humans many years ago may have many more surprising applications in store.

9.9 Assessment for Chapter 9

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 10 Ionic Bonds and Formulas C-SE-TE

Chapter Outline

- 10.1 IONS AND ION FORMATION
- 10.2 IONIC COMPOUNDS
- 10.3 WRITING IONIC FORMULAS
- 10.4 NAMING IONIC COMPOUNDS
- 10.5 WORKSHEETS FOR CHAPTER 10
- 10.6 EXTRA READING FOR CHAPTER 10
- 10.7 ASSESSMENT FOR CHAPTER 10

Lessons and Number of Activities for Lessons

TABLE 10.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Ions and Ion Formation	0	0	1	1
2. Ionic Compounds	0	0	0	0
3. Writing Ionic Formulas	0	0	1	1
4. Naming Ionic Compounds	0	0	1	0

10.1 Ions and Ion Formation

Student Behavioral Objectives

- The students will define an ion.
- The students will identify the atoms most likely to form positive ions, and the atoms most likely to form negative ions.
- The students will explain why atoms form ions.
- The students will predict the charge on ions from the electron affinity, ionization energies, and electron configuration of the atom.
- The students will describe how atoms form an ionic bond.
- The students will state why, in terms of energy, atoms form ionic bonds.
- The students will state the octet rule.
- Given the symbol of a representative element, the students will indicate the most likely number of electrons the atom will gain or lose.
- Given the electron configuration of a representative element, the students will indicate the most likely number of electrons the atom will gain or lose.
- Given the successive ionization energies of a metallic atom, the students will indicate the most likely number of electrons the atom will lose during ionic bond formation.

Timing, Standards, Activities

TABLE 10.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Ions and Ion Formation</i>	1.0	1d, 2c

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Ion Formation Worksheet

Extra Readings

1. The History of Chemical Symbols

Answers for Ions and Ion Formation (L1) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

10.2 Ionic Compounds

Student Behavioral Objectives

- The student will give a short, generic description of a lattice structure.
 - The student will identify distinctive properties of ionic compounds.
-

Timing, Standards, Activities

TABLE 10.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Ionic Compounds</i>	1.0	2c

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None
-

Answers for Ionic Compounds (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

10.3 Writing Ionic Formulas

Student Behavioral Objectives

- Given the elements to be combined, the student will write correct formulas for binary ionic compounds, compounds containing metals with variable oxidation numbers, and compounds containing polyatomic ions.
-

Timing, Standards, Activities

TABLE 10.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Writing Ionic Formulas</i>	1.0	None

Activities for Lesson 3

Laboratory Activities

- None

Demonstrations

- None

Worksheets

- Formula Writing Worksheet

Extra Readings

- Chemical Bonding
-

Answers for Writing Ionic Formulas (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

10.4 Naming Ionic Compounds

Student Behavioral Objectives

- Given the formula for a binary ionic compound, a compound containing metals with variable oxidation numbers, or a compound containing polyatomic ions, the students will be able to name it.
- Given the name for a binary ionic compound, a compound containing metals with variable oxidation numbers, or a compound containing polyatomic ions, the students will write the correct formula for it.

Timing, Standards, Activities

TABLE 10.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Naming Ionic Compounds</i>	1.0	None

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Inorganic Nomenclature Worksheet

Extra Readings

1. None

Answers for Naming Ionic Compounds (L4) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

Multimedia Resources for Chapter 10

This website provides printable periodic tables for handouts.

- <http://science.widener.edu/svanbram/ptable.html>

This website provides a fill-in-the-blank worksheet generator.

- <http://www.theteacherscorner.net/printable-worksheets/make-your-own/fill-in-the-blank/>

This website provides an interactive lesson on chemical bonding.

- http://www.visionlearning.com/library/module_viewer.php?mid=55

This website provides an interactive lesson on ionic bonding.

- <http://www.teachersdomain.org/resource/lsp07.sci.phys.matter.ionicbonding/>

This website provides a writing the formulas of ionic compounds flowchart.

- <http://www.phs.princeton.k12.oh.us/departments/Science/lusch/honorspdfs/namingchpt5/Flowcharts.pdf>

This website provides more information about ionic compounds.

- <http://misterguch.brinkster.net/ionic.html>

This video is a ChemStudy film called “Electric Interactions in Chemistry.” The film is somewhat dated but the information is accurate.

- <http://www.youtube.com/watch?v=o9TaQLVCFDM>

This website provides more details about ionic bonding, including a conceptual simulation of the reaction between sodium and chlorine. The website also discusses covalent bonding, the focus of the chapter “Covalent Bonds and Formulas.”

- http://visionlearning.com/library/module_viewer.php?mid=55

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 5-4 is on naming compounds.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson54.htm>

Laboratory Activities for Chapter 10

Demonstrations for Chapter 10

10.5 Worksheets for Chapter 10

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Ion Formation Worksheet

Questions 1 - 4 relate to element *X* whose first six ionization energies are shown in the table below. Element *X* is a representative element.

TABLE 10.6: The First Six Ionization Energies of Element

Number of Ionization Energy	Ionization Energy (kJ/mol)
1 st	800
2 nd	1,400
3 rd	15,000
4 th	18,000
5 th	21,000
6 th	25,000

1. Is element *X* more likely to be a metal or a non-metal?
2. Which family of elements does element *X* belong to?
3. How many electrons is element *X* most likely to gain or lose in a normal chemical reaction?
4. What is the most likely charge for an ion of element *X*?

Questions 5 - 8 relate to element *Y* whose first six ionization energies are shown in the table below. Element *Y* is a representative element.

TABLE 10.7: The First Six Ionization Energies of Element

Number of Ionization Energy	Ionization Energy (kJ/mol)
1 st	500
2 nd	4,800
3 rd	6,800
4 th	9,000
5 th	13,000
6 th	15,000

5. Is element *Y* more likely to be a metal or a non-metal?
6. Which family of elements does element *Y* belong to?
7. How many electrons is element *Y* most likely to gain or lose in a normal chemical reaction?
8. What is the most likely charge for an ion of element *Y*?

Questions 9 - 12 relate to element *M* whose first eight ionization energies are shown in the table below. Element *M* is a representative element.

TABLE 10.8: The First Eight Ionization Energies of Element

Number of Ionization Energy	Ionization Energy (kJ/mol)
1 st	1,100
2 nd	1,800
3 rd	2,800
4 th	4,000
5 th	6,000
6 th	8,000
7 th	27,000
8 th	36,000

- Is element *M* more likely to be a metal or a non-metal?
- Which family of elements does element *M* belong to?
- How many electrons is element *M* most likely to gain or lose in a normal chemical reaction?
- What is the most likely charge for an ion of element *M*?

The table below gives the electron affinities for period 3 of the periodic table.

TABLE 10.9: The Electron Affinities of Elements in Period Three

Family	1A	2A	3A	4A	5A	6A	7A	8A
Electron Affinity (kJ/mol)	52	0	42	134	72	200	349	0

The table below gives the electron affinities for period 4 of the periodic table.

TABLE 10.10: The Electron Affinities of Elements in Period Four

Family	1A	2A	3A	4A	5A	6A	7A	8A
Electron Affinity (kJ/mol)	48	2	29	119	78	195	325	0

While family 5A is somewhat anomalous, the general trend is apparent in this data.

- If a representative element has an electron affinity greater than 150 kJ/mol , would you expect it to be a metal or a non-metal?
- If all the elements in a family have an electron affinity of 0 kJ/mol , what family is it most likely to be?
- The first ionization energies (in kJ/mol) of *Li*, *Na*, *K*, *Rb*, and *Cs* in random order are 370, 520, 400, 500, and 420.
 - Which first ionization energy do you think belongs to *Li*?
 - Which first ionization energy do you think belongs to *Cs*?
 - What knowledge about chemical families did you use to make those choices?
- Given the electron configuration of the outermost energy level of an atom to be s^2p^4 :
 - is the element a metal or non-metal?
 - is it most likely to gain or lose electrons?
 - how many electrons is it most likely to gain or lose in a normal chemical reaction?

D. what is the most likely charge on an ion of this element?

Formula Writing Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Fill in the squares with the appropriate formula for the compound formed by the combination of the atoms or ions that intersect.

TABLE 10.11: Formula Writing Practice

	bromine	acetate	sulfate	phosphate	hydroxide	sulfur
potassium						
calcium						
aluminum						
ammonium						
iron(III)						
lead(II)						

Inorganic Nomenclature Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

TABLE 10.12:

Number	Formula	Name
1.	<i>LiF</i>	_____
2.	<i>Na₃PO₄</i>	_____
3.	<i>Al(OH)₃</i>	_____
4.	<i>Cl₂O₇</i>	_____
5.	<i>PbO</i>	_____
6.	<i>Fe₂S₃</i>	_____
7.	<i>TeO₂</i>	_____
8.	<i>CuSO₄</i>	_____
9.	<i>Ca₃(PO₄)₂</i>	_____
10.	<i>HNO₃</i>	_____

TABLE 10.13:

Number	Name	Formula
1.	copper(I) sulfide	_____
2.	boron trichloride	_____
3.	potassium carbonate	_____
4.	sulfur hexafluoride	_____
5.	chlorine monofluoride	_____
6.	dinitrogen tetroxide	_____

TABLE 10.13: (continued)

Number	Name	Formula
7.	tin(IV) oxide	_____
8.	silver acetate	_____
9.	diphosphorus pentoxide	_____
10.	lithium nitrate	_____

Answers to Worksheets

- **The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.**

10.6 Extra Reading for Chapter 10

The History of Chemical Symbols

The one or two letter shorthand used to represent chemical elements is a familiar feature in modern science. The tradition of using symbols to represent elements is quite ancient. Long before those interested in studying the composition and behavior of matter were known as chemists, mystical practitioners of alchemy devised symbols often to obfuscate their experimentation, and to cloak their work in secrecy. Their coded imagery drew inspiration from astrology, as well as ancient writing systems like the hieroglyphs. Alchemists linked certain metals with celestial bodies to describe their behavior, such as the connection between the rapidly moving planet Mercury and the metallic liquid quicksilver.

As chemistry became an experimental science and new methods produced scores of newly discovered elements, the need for a shorthand technique to describe chemical changes became apparent. One of the first chemists to attempt to introduce a symbolic system for identifying the elements was John Dalton, known for his relative mass scale of the atomic weights. His symbols, introduced in 1808 in his “New System of Chemical Philosophy”, consisted mainly of circles, some with inscribed alphabetic letters and others with dots or lines within the circles. Compounds were written as combinations of circles representing the constituent atoms. His system did not lend itself to ready memorization and did not catch on with his contemporaries.

Our modern method of using one or two letter shorthand for the elements was devised in 1813 by Jöns Jakob Berzelius, citing the ease of implementation, particularly for typesetters. Due to the common employment of Latin terminology in scientific communication, Berzelius suggested using the first or first two letters of the element’s Latin name as the symbol for that atom. In the case of confusion or duplication of the letters, exceptions includes the use of Hg (hydrargyrum for Mercury and plumbum for lead). Some modern modifications have been introduced for new elements, especially those named in honor of famous scientists. Berzelius is also responsible for the use of subscripts in a chemical formula to designate the ratio of atoms.

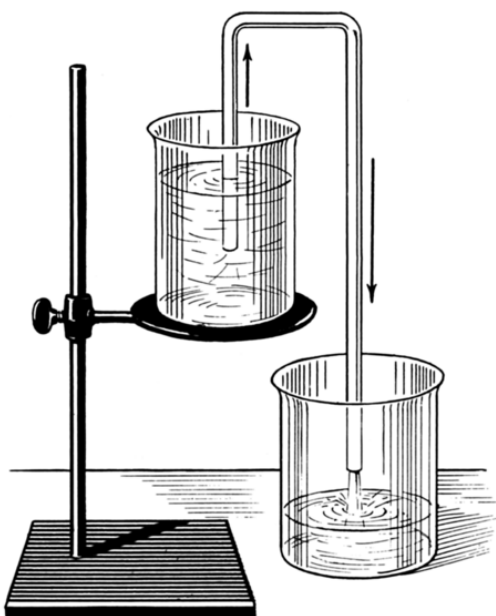
Chemical Bonding

One of the primary causes of change in physical systems is the tendency toward minimum potential energy. Objects roll downhill, objects above the Earth fall, stretched rubber bands contract, objects with like charges separate, and objects with unlike charges move together. All of these changes involve a decrease in potential energy.

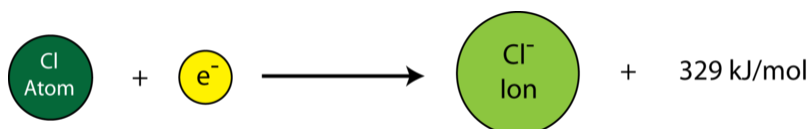
In many situations, the potential energy of a system increases somewhat, at first, in order to achieve a position from which the potential energy will significantly decrease. A siphon is an example of this. In a siphon, water will run up the hose as long as the final position of the water is lower in potential energy than the original position.

Ionic Bonding

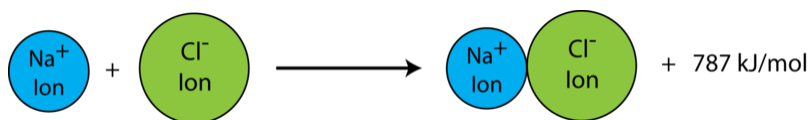
In ionic bonding, the metallic atom must lose one or more electrons in order to form a bond. This loss of electrons by metallic atoms requires an input in energy. The necessary ionization energy (energy to remove an electron) must be provided to form a cation (positive ion). Suppose we use a sodium atom as an example.



An input of energy ionizes the sodium atom to a sodium ion, Na^+ . In the presence of chlorine atoms, the electron can then add to a chlorine atom to form a negative chloride ion. In this process, energy is given off. The electron affinity of chlorine is -329 kJ/mol . That means that adding an electron to a chlorine atom is a reduction in potential energy.



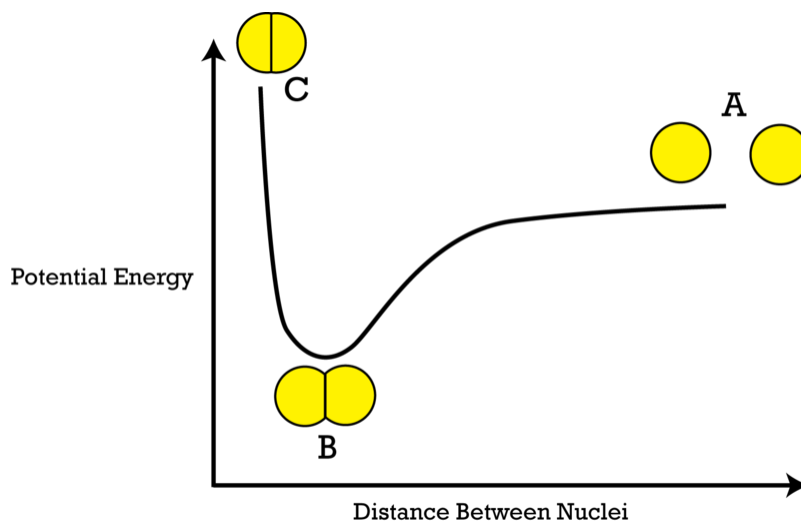
The sodium ions and the chloride ions have opposite charges, and therefore, are attracted to each other. When the ions move closer together, potential energy is again lowered. As the oppositely charged ions move together to form a crystal lattice, energy is given off. For each mole of sodium ions and chloride ions that move together to form a lattice, 787 kJ of energy are given off.



The overall process of removing an electron from a sodium atom (energy input = 496 kJ/mol), adding the electron to a chlorine atom (energy output = 329 kJ/mol), and the ions moving together in a lattice structure (energy output = 787 kJ/mol) has a net energy output of 620 kJ/mol ; therefore, like a siphon, the process occurs because it has a net lowering of potential energy.

Covalent Bonding

In the case of covalent bonding, there is no electron losing, gaining, or transferring. In covalent bonds, the bonding electrons are shared. The potential energy lowering that occurs in covalent bonding can be represented by showing the relationship between the potential energy of the system and the distance between the nuclei of the bonding atoms.



When individual non-metallic atoms are completely separated, there is an attraction (electron affinity) between the electrons of each atom and the protons in the nucleus of the other atom. This attraction and the distance between the atoms cause potential energy to exist. As the atoms move closer together, the potential energy of the system decreases because the distance between the attracting objects is becoming less.

As the distance between the two nuclei decreases, the potential energy becomes less and less. Since these are non-metallic atoms, their electron outer energy levels are not full and this allows the electron clouds of the two atoms to overlap, and the atoms may continue to move closer together. When the atoms reach position B in the diagram, the potential energy of the system is at its lowest possible for these two atoms. If the nuclei continue to move closer together, the potential energy increases dramatically due to the repulsion of the positively charged nuclei. In position A, the nuclei are too close together so potential energy is high. In position C, the atoms are too far apart so the potential energy is high. In position B, the atoms are the proper distance apart for lowest potential energy. Since these two atoms reach lowest potential energy in a position where their electron clouds are overlapped, they will be bonded. The distance between the nuclei in position B is known as the bond length. If the atoms attempt to move either closer together or farther apart, the potential energy increases. The tendency toward minimum potential energy causes these atoms to remain in the bonded position.

10.7 Assessment for Chapter 10

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 11

Covalent Bonds and Formulas C-SE-TE

Chapter Outline

- 11.1 THE COVALENT BOND
- 11.2 WRITING FORMULAS AND NAMING COVALENT COMPOUNDS
- 11.3 ELECTRONIC AND MOLECULAR GEOMETRY
- 11.4 THE GEOMETRICAL ARRANGEMENT OF ELECTRONS AND MOLECULAR SHAPE
- 11.5 WORKSHEETS FOR CHAPTER 11
- 11.6 EXTRA READING FOR CHAPTER 11
- 11.7 ASSESSMENT FOR CHAPTER 11

Lessons and Number of Activities for Lessons

TABLE 11.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. The Covalent Bond	0	0	0	0
2. Writing Formulas and Naming Covalent Compounds	0	0	1	0
3. Electronic and Molecular Geometry	1	0	0	1
4. The Geometrical Arrangement of Electrons and Molecular Shape	0	0	1	1

11.1 The Covalent Bond

Student Behavioral Objectives

The student will:

- describe how covalent bonds form.
- explain the difference between ionic and covalent bond formation and structure.
- state the relationship between molecular stability and bond strength.
- identify pairs of atoms that will form covalent bonds.
- draw Lewis structures for simple covalent molecules.
- identify sigma and pi bonds in a Lewis structure.
- describe (chemistry) resonance.
- explain the equivalent bond strengths in a resonance situation.

Timing, Standards, Activities

TABLE 11.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Covalent Bond</i>	2.0	2a, 2b, 2e

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for The Covalent Bond (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

11.2 Writing Formulas and Naming Covalent Compounds

Student Behavioral Objectives

- The students will name binary covalent compounds using the IUPAC nomenclature system.
 - The students will provide formulas for binary covalent compounds given the IUPAC name.
-

Timing, Standards, Activities

TABLE 11.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Writing Formulas and Naming Covalent Compounds</i>	1.0	2b

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Completion of Worksheet from Chapter 10

Extra Readings

1. None
-

Answers for Writing Formulas and Naming Covalent Compounds (L2) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

11.3 Electronic and Molecular Geometry

Student Behavioral Objectives

- Given binary formulas and an electronegativity chart, students will identify the most likely bonding type (ionic, covalent, or polar covalent) for each compound.
- The students will describe a polar covalent bond and explain why it forms.
- Given binary formulas and an electronegativity chart, students will identify the most likely bonding type (ionic, covalent, or polar covalent) for each compound.
- The students will draw Lewis structures for simple molecules that violate the octet rule.
- Given a list of binary compounds, the students will identify those that require electron promotion in the explanation of their bonding.
- The students will identify the type of hybridization in various molecules.
- The students will explain the necessity for the concept of hybridized orbitals.
- The students will state the main postulate in VSEPR theory.
- The students will identify both the electronic and the molecular geometry for simple binary compounds.

Timing, Standards, Activities

TABLE 11.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Electronic and Molecular Geometry</i>	2.0	2a, 2b, 2g

Activities for Lesson 3

Laboratory Activities

1. Molecular Models and Shapes Lab

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. Organic Conductors

Answers for Electronic and Molecular Geometry (L3) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

11.4 The Geometrical Arrangement of Electrons and Molecular Shape

Student Behavioral Objectives

- The students will determine whether bonds are polar or non-polar.
- The students will determine whether simple molecules are polar or non-polar.

Timing, Standards, Activities

TABLE 11.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Geometrical Arrangement of Electrons and Molecular Shape</i>	1.0	2f

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Molecular Geometry Worksheet

Extra Readings

1. Chelates

Answers for The Geometrical Arrangement of Electrons and Molecular Shape (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 11

This website provides a lesson on Chemical Bonding.

- http://www.visionlearning.com/library/module_viewer.php?mid=55

This website provides a Polar Bears and Penguins Bonding Activity.

- <http://www.keypress.com/Documents/chemistry/SampleLessons/SmellsTG.pdf>

This website provides examples of polar and non-polar molecules.

- <http://www.ausetute.com.au/molpolar.html>

This website provides a tutorial on drawing resonance structures.

- http://www.chem.ucla.edu/harding/tutorials/resonance/draw_res_str.html

This website provides an animation showing ionic and covalent bonding.

- <http://www.youtube.com/watch?v=QqjCvzWwww>

This website provides an introduction to drawing Lewis electron dot symbols.

- <http://www.youtube.com/watch?v=y6QZRBIO0-o>

The learner.org website allows users to view the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The video called “Chemical Bonds” explains the differences between ionic and covalent bonds using models and examples from nature.

- http://www.learner.org/vod/vod_window.html?pid=800

This animation explores the differences between ionic and covalent bonding.

- http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/bom1s2_11.swf

This website provides more information about the different types of chemical bonding.

- http://www.visionlearning.com/library/module_viewer.php?mid=55

The website below provides a guide to drawing Lewis structures.

- <http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch8/lewis.html#step>

This website reviews the rules for naming binary covalent compounds and includes a quiz to test your understanding of these rules.

- http://www.chem.purdue.edu/gchelp/nomenclature/covalent_2009.htm

For an introduction to polar bonds see these two videos. The first video defines polar bonds. The second one consists of lecture and animation about polar bonds and polar molecules.

- <http://www.youtube.com/watch?v=rRoLSBeOjME>

- <http://www.youtube.com/watch?v=LKAjTE7B2x0>

This animation reviews the differences in ionic, covalent, and polar covalent bonds.

- <http://www.youtube.com/watch?v=Kj3o0XvhVqQ>

The following websites provide more information on electron promotion and orbital hybridization.

- <http://www.mikeblaber.org/oldwine/chm1045/notes/Geometry/Hybrid/Geom05.htm>
- <http://pages.towson.edu/ladon/carbon.html>

This video is a ChemStudy film called “Shapes and Polarities of Molecules.”

- <http://www.youtube.com/watch?v=BnU2idxQ3Xc>

An animation showing the molecular shapes that are generated by sharing various numbers of electron pairs around the central atom (includes shapes when some pairs of electrons are non-shared pairs).

- http://www.classzone.com/cz/books/woc_07/resources/htmls/ani_chem/chem_flash/popup.html?layer=act&src=qtifw_act065.1.xml

The *learner.org* website allows users to view the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The video called “Molecular Architecture” is related to this lesson.

- <http://learner.org/resources/series61.html>

This website reviews how to predict molecular structure by using the VSEPR theory.

- http://www.up.ac.za/academic/chem/mol_geom/mol_geometry.htm

Laboratory Activities for Chapter 11

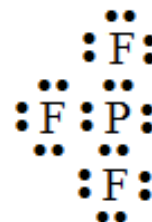
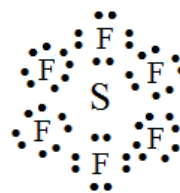
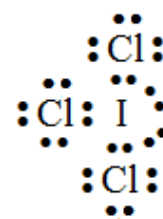
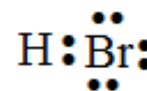
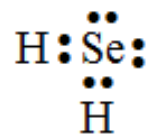
Teacher’s Resource Page for Molecular Models and Shapes

Lab Notes:

The molecular model kits must contain the 5-hole ball (trigonal bipyramidal) and the 6-hole ball (octahedral). Model kits containing these are available from Arbor Scientific.

You will need 1 kit for each student or pair of students.

If you wish to have complete sets for many years, you will need to supervise students closely. Small parts are easily lost. You may wish to check the kits in and out or assign the task to a lab assistant.

TABLE 11.6: Answers to Pre-Lab Questions**Formula****Electron-Dot Formula***PF₃**SF₆**ICl₃**HBr**H₂Se*

Molecular Models and Shapes Laboratory

Background:

The shapes exhibited by molecules are often very difficult for beginning chemistry students to visualize, especially since most students' training in geometry is limited to plane geometry. To understand the geometric shapes exhibited

by molecules, a course in solid geometry would be more useful. In this experiment, you will encounter some unfamiliar geometrical arrangements that will help you appreciate the complexity of molecular geometry.

The basic derivation and explanation of molecular shapes arises from the Valence Shell Electron Pair Repulsion theory, usually known by its abbreviation, VSEPR. This theory considers the environment of the most central atom in a molecule, and imagines first how the valence electron pairs of that central atom must be arranged in three-dimensional space around the atom to minimize repulsion among the electron pairs. The general principle is: for a given number of pairs of electrons, the pairs will be oriented in three-dimensional space to be as far away from each other as possible. For example, if a central atom were to have only two pairs of electrons around it, the electron pairs would be expected to be 180° from each other.

The VSEPR theory then also considers which pairs of electrons are bonding pairs (have atoms attached) and which are non-bonding pairs (lone pairs). The overall shape of the molecule as a whole is determined by how many pairs of electrons are around the central atom (electronic geometry), and how many of these pairs are bonding pairs (molecular geometry). A simple example that clearly makes this distinction concerns the case in which the central atom has four valence electron pairs. Consider the Lewis structures of the following four molecules: hydrogen chloride, HCl ; water, H_2O ; ammonia, NH_3 ; and methane, CH_4 . The central atom of each of these molecules is surrounded by four pairs of valence electrons. According to the VSEPR theory, these four pairs of electrons will be oriented in three-dimensional space to be as far away from each other as possible. The four pairs of electrons point to the corners of the geometrical figure known as a tetrahedron. The four pairs of electrons are said to be tetrahedrally oriented, and are separated by angles of approximately 109.5° . In all four cases, the electronic geometry is tetrahedral but the overall molecular geometry is tetrahedral for only one of the four molecules.

In the case of HCl , even though there are four pairs of electrons around the chlorine atom, three of them are not shared - there is no atom bonded by them. These spaces are empty. Since there are only two atoms joined by a bond (a pair of electrons), the molecular geometry will be linear. For H_2O , there are two shared pairs of electrons and two unshared pairs. The molecular geometry of this molecule will be angular (or bent). In NH_3 , three of the electrons are shared and one is unshared. This molecular geometry will be pyramidal. In the final molecule, CH_4 , all four pairs of electrons are shared and so the molecular geometry matches the electronic geometry and is tetrahedral.

Pre-Lab Questions

Draw Lewis structures and predict bond angles for the following molecules.



Purpose:

Models built according to the predictions of the VSEPR theory illustrate the regular patterns of molecular shapes.

Apparatus and Materials:

- 1 molecular model kit for each lab group
- 1 protractor for each lab group

Procedure:

Build each of the 19 models indicated in the following table of geometries. Measure all bond angles in your models. Sketch each of your models including bonding and non-bonding electron pairs and bond angles. See if you can find real molecules that would be expected to have each shape. (There are 5 models for which no real molecule is known.) You may need to sort through your textbook to find the examples.

TABLE 11.7: Table of Electronic Geometries

	Pairs on Central Atom	Electronic Geometry	Bonding Pairs	Molecular Geometry	Formula Type
1.	2	linear	2	linear	AB_2
2.	3	trigonal planar	1	linear	AB^*
3.	3	trigonal planar	2	bent	AB_2^*
4.	3	trigonal planar	3	trigonal planar	AB_3
5.	4	tetrahedral	1	linear	AB
6.	4	tetrahedral	2	bent	AB_2
7.	4	tetrahedral	3	pyramidal	AB_3
8.	4	tetrahedral	4	tetrahedral	AB_4
9.	5	trigonal bipyramidal	1	linear	AB
10.	5	trigonal bipyramidal	2	linear	AB_2
11.	5	trigonal bipyramidal	3	T-shape	AB_3
12.	5	trigonal bipyramidal	4	distorted tetrahedron	AB_4
13.	5	trigonal bipyramidal	5	trigonal bipyramidal	AB_5
14.	6	octahedral	1	linear	AB^*
15.	6	octahedral	2	linear	AB_2^*
16.	6	octahedral	3	T-shape	AB_3^*
17.	6	octahedral	4	square planar	AB_4
18.	6	octahedral	5	square pyramid	AB_5
19.	6	octahedral	6	octahedral	AB_6

*Indicates no molecules known.

Demonstrations for Chapter 11

11.5 Worksheets for Chapter 11

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Molecular Geometry Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Lewis structures only show how many bonding pairs of electrons, and unshared pairs of electrons, surround a given atom on a flat page. The molecules are actually three dimensional which is not shown by Lewis structures. To convey a sense of three dimensionality, we use “ball and stick” models.

There is a correlation between the number of electron pairs, (sigma bonds plus non-shared pairs) around the central atom of a molecule, and the electronic geometry of that molecule.

The idea that allows us to predict the electronic geometry is that each pair of electrons (shared or unshared) repels all the other electron pairs. The electron pairs move as far apart as possible, but since they are all tied to the central atom, they can only orient themselves in such a way that they make the angles between them as large as possible. This is the essence of the *Valence Shell Electron Pair Repulsion (VSEPR) Theory* for predicting molecular shapes.

To use VSEPR theory, we must first be able to determine the number of valence shell electron pairs around the central atom. These pairs consist of all sigma bond pairs and all unshared pairs of electrons. Pi bond electrons are excluded because the electrons are not placed between bonding atoms and therefore, do not contribute to electronic geometry.

TABLE 11.8: Visualizing Electron Pairs

Electron Pairs

To visualize the electron pairs that contribute to electronic geometry, imagine them situated on the surface of a sphere with the central atom at the center.

If there are only two pairs of electrons in the valence shell of the central atom, the two pairs can avoid each other best if they are 180° apart. This means that the two pairs and the central atom are in a straight line; the arrangement is **linear**.

Image

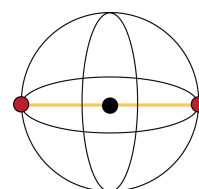
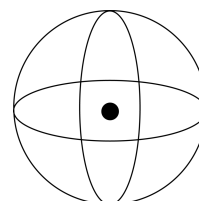
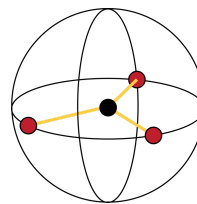


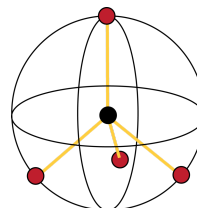
TABLE 11.8: (continued)

Electron Pairs

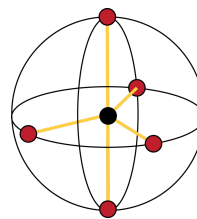
If a third pair of electrons is added, the three pairs push around to the shape shown at right. The angles between electron pairs would be 120° and we call the shape **trigonal planar**. The three pairs of electrons and the central atom are all in a single plane.

Image

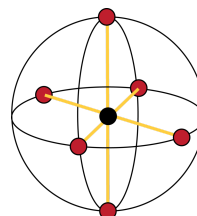
A fourth pair of electrons causes the electrons to push around into the shape shown at right, the **tetrahedron**. The angles in this shape are 109.5° .



A fifth pair of electrons produces the shape known as **trigonal bipyramidal**. The angles between the three pairs of electrons around the center is 120° and the angles between the pairs around the center and the pairs on the ends is 90° .



Finally, the sixth pair of electrons produces the **octahedral** shape shown at right. All angles in this shape are 90° .



Once the number of electron pairs surrounding the central atom is determined, the **electronic geometry** is known.

TABLE 11.9: The Relationship Between Number of Electron Pairs and Electronic Geometry

Electron Pairs Around the Central Atom	Electronic Geometry
1 pair	Linear
2 pairs	Linear
3 pairs	Trigonal Planar
4 pairs	Tetrahedral
5 pairs	Trigonal Bipyramidal
6 pairs	Octahedral

The **molecular** geometry may be different from the **electronic** geometry because many times, not all the electron pairs are shared. An unshared electron pair will not have an atom in that position of the electronic geometry. In order

to determine molecular geometry, we must recognize which pairs of electrons have an atom attached and which are lone pairs. The overall shape of the molecule is determined by how many pairs of electrons are around the central atom, and how many of these have atoms attached.

It is sometimes difficult for students to recognize the difference between the orientation of electron pairs (called electronic geometry) and the overall shape of the molecule (called molecular geometry). We will look at an example that shows the difference between electronic and molecular geometry. Consider the following four molecules: hydrogen chloride, HCl ; water, H_2O ; ammonia, NH_3 ; and methane, CH_4 .

TABLE 11.10: The Relationship Between Shared Pairs and Molecular Geometry

Shared Pairs

The central atom of each of these molecules is surrounded by four pairs of electrons. According to VSEPR theory, these four pairs will be oriented in three-dimensional space to be as far away from each other as possible. The four pairs will point to the corners of the geometrical shape known as a tetrahedron. The angles between the electron pairs will be approximately 109.5° . In all four cases, the **electronic geometry** is **tetrahedral** but only one of the molecules will have tetrahedral molecular geometry.

In the case of HCl , even though there are four pairs of electrons around the chlorine atom, three of them are not shared. There is no atom attached to them. These spaces are empty. Since there are only two atoms joined by a bond, the molecular geometry will be **linear**.

In the water molecule, two electron pairs are shared and two are unshared. So while the electronic geometry is tetrahedral, the molecular geometry is **bent** (aka angular, aka V-shaped).

In the ammonia molecule, one pair of electrons is unshared and the other three are shared. This results in a molecular shape called **pyramidal**.

Molecular Geometry

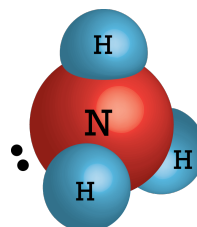
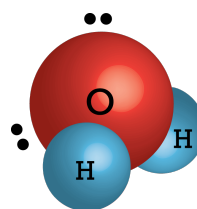
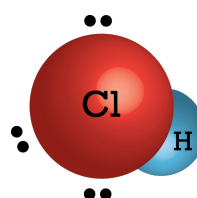
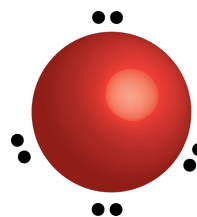


TABLE 11.10: (continued)

Shared Pairs

In the methane molecule, all four pairs of electrons are shared, and so not only is the electronic geometry tetrahedral but the molecular geometry is also **tetrahedral**.

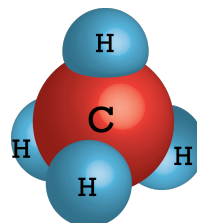
Molecular Geometry

TABLE 11.11:

Central Atom	Electron Pairs	Electronic Geometry	Bonding Pairs	Molecular Geometry	Sketch
	2	Linear	2	Linear	
	3	Trigonal Planar	1	Linear	
	3	Trigonal Planar	2	Bent	
	3	Trigonal Planar	3	Trigonal Planar	
	4	Tetrahedral	1	Linear	
	4	Tetrahedral	2	Bent	

TABLE 11.11: (continued)

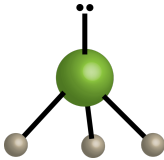
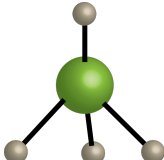
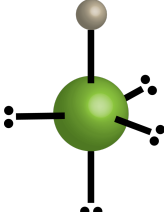
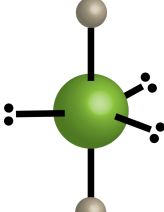
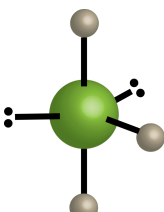
Central Electron Pairs	Atom	Electronic Geometry	Bonding Pairs	Molecular Geometry	Sketch
4		Tetrahedral	3	Pyramidal	
4		Tetrahedral	4	Tetrahedral	
5		Trigonal Bipyramidal	1	Linear	
5		Trigonal Bipyramidal	2	Linear	
5		Trigonal Bipyramidal	3	T-shape	

TABLE 11.11: (continued)

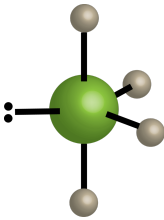
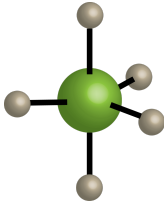
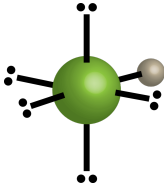
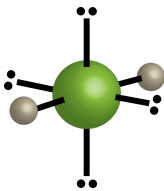
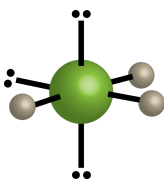
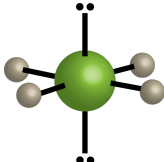
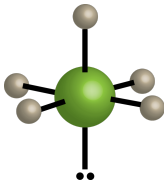
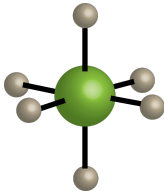
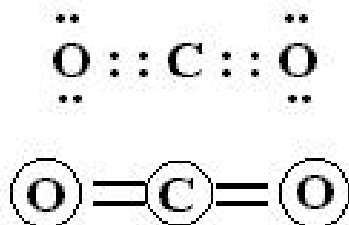
Central Electron Pairs	Atom	Electronic Geometry	Geome-try	Bonding Pairs	Molecular Geome-try	Sketch
5		Trigonal Bipyramidal	Bipyrami-dal	4	Distorted Tetrahedron	
5		Trigonal Bipyramidal	Bipyrami-dal	5	Trigonal Bipyramidal	
6		Octahedral		1	Linear	
6		Octahedral		2	Linear	
6		Octahedral		3	T-shape	
6		Octahedral		4	Square Planar	

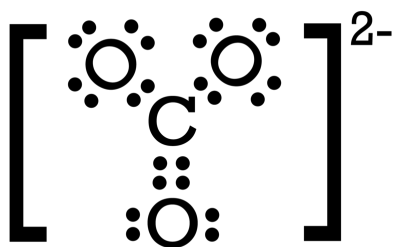
TABLE 11.11: (continued)

Central Atom	Electronic Geometry	Bonding Pairs	Molecular Geometry	Sketch
6	Octahedral	5	Square Pyramidal	
6	Octahedral	6	Octahedral	

In order to choose the correct molecular geometry, you must keep in mind that only electron pairs involved in sigma bonds and unshared pairs contribute to electronic geometry. Pi bonds are not directed bonds, and those electron pairs do not contribute to electronic geometry. In the Lewis structure for the carbon dioxide molecule (shown at right), it is clear that the central atom is carbon, and the carbon atom is surrounded by 4 pairs of electrons. But these four pairs of electrons are involved in two sigma bonds and two pi bonds. Therefore, the electronic geometry of carbon dioxide is based on **two** pairs of electrons around the central atom, and will be linear. Since both pairs of electrons are shared, the molecular geometry will also be linear.



The Lewis structure for the carbonate ion, shown at right, shows the central atom is carbon and it is surrounded by 4 electron pairs. One of those pairs, however, is a pi bond, and therefore the electronic geometry of the carbonate ion is based on 3 pairs of electrons around the central atom. Thus, the electronic geometry is trigonal planar and since all three pairs are shared, the molecular geometry is also trigonal planar.



Polarity

Bonds between atoms that are of the same element are non-polar bonds. Molecules composed of all the same atom such as Cl_2 , O_2 , H_2 , S_8 , P_4 , have no polar bonds and therefore do not have dipoles. That is, the molecules will be non-polar. A molecule that does have polar bonds can still be non-polar. If the polar bonds are symmetrically distributed, the bond dipoles cancel and do not produce a molecular dipole.

TABLE 11.12: Symmetrical and Non-Symmetrical Molecular Shapes

Molecular Shape	Symmetry
Linear	Symmetrical
Bent	Non-Symmetrical
Trigonal Planar	Symmetrical
Pyramidal	Non-Symmetrical
Tetrahedral	Symmetrical
T-shaped	Non-Symmetrical
Distorted Tetrahedron	Non-Symmetrical
Trigonal Bipyramidal	Symmetrical
Square Planar	Symmetrical
Square Pyramidal	Non-Symmetrical
Octahedral	Symmetrical

Exercises

Fill in the table with electronic geometry, molecular geometry, and indicate whether the molecular will be polar or non-polar.

TABLE 11.13: Polarity Table

Formula	Electronic Geometry	Molecular Geometry	Polarity
AsH_3			
BCl_3			
IF_3			
$SiBr_4$			
SeH_4			
XeI_4			
OF_2			
KrF_2			
ICl_5			
CCl_2F_2			

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

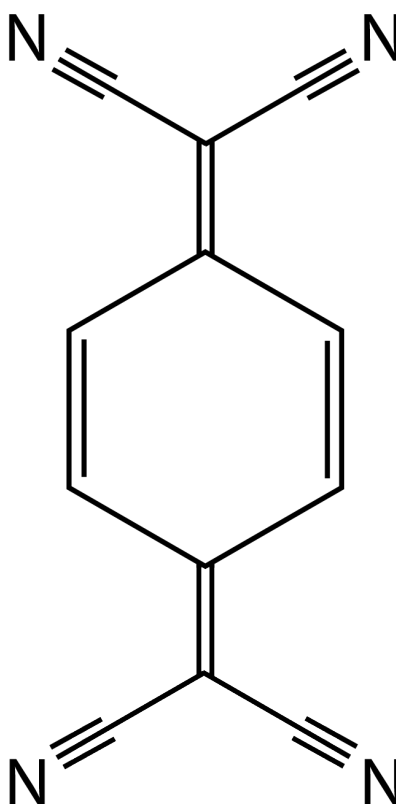
11.6 Extra Reading for Chapter 11

Organic Conductors

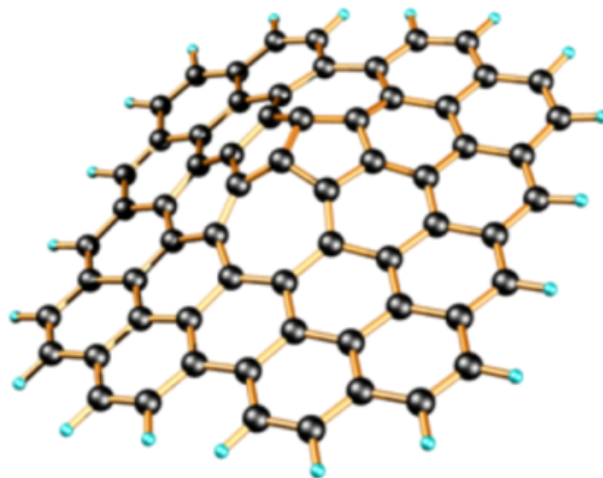
Metallurgy was one of the first applied sciences to be mastered by humankind, and successive generations have garnered increasing expertise in processing metal ores into many different versatile and valuable substances. Despite the overwhelming dependence of modern technology on metals and their applications, new supplies of many different common metals have become increasingly more difficult to locate and procure. Many metals once considered plentiful are now deemed semi-precious, and the possibility exists that we may be restricted to the reserves on hand in the future.

The properties that make metals so valuable, such as their electrical and heat conductivity, malleability, hardness, and density are difficult to replicate in other materials. One attempt to retain the conductive characteristics of metals in more readily available materials, is the development of organic conductors. Although most organic molecules are considered to be insulators, organic materials have been developed to produce semiconductors as well as truly conductive systems.

The first organic conductors were constructed as charge transfer complexes; these systems consisted of two molecules with one acting as an electron donor and the other an electron acceptor. For example, tetracyanoquinodimethane (TCNQ) was first identified in 1962. As its structural formula indicates, TCNQ contains alternating single, double and triple bonds, and this structure readily accepts electrons while



resulting in reallocation of the pi bonding electrons into new bonding arrangements. Several TCNQ complexes with a variety of electron donors, with high conductivities even into temperature ranges when the salt complexes melted. Organic conductors are compelling research targets due to the vast availability of the raw materials used to prepare them, and new research suggests the possibility of producing conductive biomaterials for medical applications. The graphene molecule has already been demonstrated to form attachments with nerve cells which display electrical conductance.



Chelates

The expected shapes of molecules containing non-metal atoms can be predicted from Valence Shell Electron Pair Repulsion (VSEPR) Theory. The basis of this theory dictates that the optimal shape of the molecule maximizes the spatial distance between groups situated around a central atom.

Metals also, may have groups oriented around them utilizing the same premises for assigning their shape. In the case of metal ions, the attached groups are usually referred to as ligands. When one ligand is attached to more than one site in the coordination sphere of the central metal, this is an example of a group known as a chelate. The term chelate comes from the Greek word “chele” meaning the claw, such as that of a crab or lobster. The ready attachment of these multidentate groups has been employed to extract the metal ion in certain situations, such as in what is known as chelation therapy. This technique is used to remove certain undesirable or toxic metal ions, such as lead or mercury ions, from the body in cases of heavy metal poisoning.

The first use of chelating agents was between the world wars as an antidote to the arsenic-based poisonous gas, Lewisite, used on the battlefields of World War I. With what became known as British anti-Lewisite (BAL), a sulfur-based chelation agent was successfully applied to treat the gassing victims. In addition, the application of a chelate can be used to sequester metal ions such as radioactive thorium or plutonium for waste stream remediation.

Chelates have also been used to stabilize metal ions and in some cases, improve their solubility as well. Gadolinium ions are desirable paramagnetic agents for use as contrast agents in Magnetic Resonance Imaging, although the metal ions themselves have considerable toxicity. The use of DTPA (Diethylene triamine pentaacetic acid), has proved to be an effective agent for the enhanced solubility, improved biodistribution, but most importantly, superior stability in vivo. Gd-DTPA contrast agents were approved for use in human MRI scans in 1988.

Chelates have also been used in metalworking applications to control the availability of the metal ion. In many cases, chelates are used in place of other more toxic ligands, such as the cyanide ion.

Metal chelates are also employed in agricultural applications to provide improved interaction of metal ions with soil

components. Also for better migration of the metal ion and therefore better distribution, particularly for those metal ions with important roles as macronutrients and micronutrients.

One possible application of the use of chelates in medical treatment may be their use in the arteriosclerosis therapy. Research in progress utilizes chelates to sequester the calcium ions in arterial plaques. As calcium ions may serve as the binders that keep these plaques intact, the exploitation of the chelate effect may prove to be a key breakthrough in improving the longterm health of cardiac patients.

11.7 Assessment for Chapter 11

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 12 The Mole Concept C-SE-TE

Chapter Outline

- 12.1 DETERMINING FORMULA AND MOLECULAR MASS**
- 12.2 THE MOLE**
- 12.3 PERCENT COMPOSITION**
- 12.4 EMPIRICAL AND MOLECULAR FORMULAS**
- 12.5 WORKSHEETS FOR CHAPTER 12**
- 12.6 EXTRA READING FOR CHAPTER 12**
- 12.7 ASSESSMENT FOR CHAPTER 12**

Lessons and Number of Activities for Lessons

TABLE 12.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Determining Formula and Molecular Mass	0	0	1	0
2. The Mole	0	0	1	1
3. Percent Composition	1	0	1	0
4. Empirical and Molecular Formulas	1	0	2	0

12.1 Determining Formula and Molecular Mass

Student Behavioral Objectives

- When given the formula or name of a compound and a periodic table, the student will be able to calculate the formula mass.

Timing, Standards, Activities

TABLE 12.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Determining Formula and Molecular Mass</i>	2.0	None

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Calculating Molar Masses Worksheet

Extra Readings

1. None

Answers for Determining Formula and Molecular Mass (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

12.2 The Mole

Student Behavioral Objectives

- Given the number of particles of a substance, the student will use Avogadro's number to convert to moles and vice versa.
- Given the number of moles of a substance, the student will use the molar mass to convert to grams and vice versa.

Timing, Standards, Activities

TABLE 12.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Mole</i>	2.0	3b, 3c, 3d

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Moles Worksheet

Extra Readings

1. Avogadro's Number

Answers for The Mole (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

12.3 Percent Composition

Student Behavioral Objectives

- Given either masses of elements in a compound, the student will calculate the percent composition by mass.
 - Given the formula or name of a compound, the student will calculate the percent composition by mass.
-

Timing, Standards, Activities

TABLE 12.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Percent Composition</i>	1.5	3d

Activities for Lesson 3

Laboratory Activities

1. Water of Hydration Lab

Demonstrations

1. None

Worksheets

1. Percent Composition Worksheet

Extra Readings

1. None
-

Answers for Percent Composition (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

12.4 Empirical and Molecular Formulas

Student Behavioral Objectives

- The student will reduce molecular formulas to empirical formulas.
- Given either masses or percent composition of a compound, the student will determine the empirical formula.
- Given either masses or percent composition of a compound and the molar mass, the student will determine the molecular formula.

Timing, Standards, Activities

TABLE 12.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Empirical and Molecular Formulas</i>	2.0	None

Activities for Lesson 4

Laboratory Activities

1. Empirical Formula of Magnesium Oxide

Demonstrations

1. None

Worksheets

1. Empirical Formulas Worksheet
2. Molecular Formulas Worksheet

Extra Readings

1. None

Answers for Empirical and Molecular Formulas (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 12

This website provides an interactive quiz on moles.

- <http://glencoe.com/qe/science.php?qi=978>

This website provides a Mole Conversion practice problem generator.

- <http://science.widener.edu/svb/tutorial/massmoles.html>

This website provides “Chemical Composition” flashcards.

- http://college.cengage.com:80/chemistry/general/zumdahl/world_of_chem/1e/students/flashcards/ch06/index.html

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 5-6 is on molecular and formula mass.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson56.htm>

This video introduces the concept of the mole.

- <http://www.youtube.com/watch?v=AsqEkF7hcII>

The *learner.org* website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The website has a video that apply to this lesson called “The Mole.”

- <http://learner.org/resources/series61.html>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 9-1 is on the mole.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson91.htm>

The website below reviews how to calculate the molar mass of chemical compounds.

- <http://misterguch.brinkster.net/molarmass.html>

This website has solved example problems for a number of topics covered in this lesson, including the calculation of percent composition by mass.

- <http://www.sciencejoywagon.com/chemzone/05chemical-reactions/>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 5-8 is on percent composition.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson58.htm>

The website below reviews how to calculate percent composition.

- <http://www.ausetute.com.au/percentc.html>

This website has solved example problems for a number of topics covered in this lesson, including the determination of empirical and molecular formulas.

- <http://www.sciencejoywagon.com/chemzone/05chemical-reactions/>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 5-5 is on empirical formulas.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson55.htm>

The website below reviews how to calculate empirical formulas.

- <http://www.chem.lsu.edu/lucid/tutorials/empiricalform.html>

Laboratory Activities for Chapter 12

Teacher's Pages for Empirical Formula of Magnesium Oxide

Lab Notes

It is not crucial that the magnesium ribbon be exactly 35 *cm*, but it should be clean. A good way to clean the ribbon is to dip it in 0.1 *MHCl* for a couple of seconds, then rinse it in distilled water, and dry it in alcohol or acetone.

The more finely divided the ribbon is, the faster it will react. 0.5 *cm* to 1.0 *cm* pieces seems to work best. The crucibles often react with the magnesium during this process. This can cause greenish-black discoloration to the crucibles. This does not really affect chemical behavior for later reactions, as long as they are cleaned. However, the crucibles will often crack during this procedure. If a crucible cracks, discard it. Stress to the students that the crucibles get extremely hot. A hot crucible can cause a very serious burn. Show the students how to handle a crucible properly by using crucible tongs. A common source of error in this experiment is not to react the nitride. Make sure the students do this portion of the lab – it is not washing.

Answers to Pre-Lab Questions

1. Assume you have 100 *grams* of the compound. This then changes the percent composition to grams of each element.

Find the number of moles of each element present. (Divide the grams of each element by its molar mass.)

Divide each of these answers by the smallest answer. This will give the empirical formula.

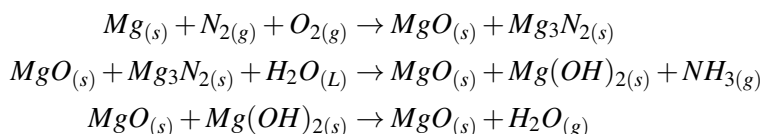
2. The molecular weight of the compound is needed.
3. Yes. The charge of magnesium increases from 0 to +2, and the charge of the oxygen is reduced from 0 to –2.
4. 3.93 g for MgO , 3.28 g for Mg_3N_2

Lab - Empirical Formula of Magnesium Oxide

Background Information

In this lab, magnesium metal (an element) is oxidized by oxygen gas to magnesium oxide (a compound). Magnesium reacts vigorously when heated in the presence of air. The $Mg - O_2$ reaction is energetic enough to allow some Mg to react with gaseous N_2 . Although there is a higher percentage of N_2 gas in the atmosphere than O_2 , O_2 is more reactive, and the magnesium oxide forms in a greater amount than the nitride. The small amount of nitride that forms can be removed with the addition of water, which converts the nitride to magnesium hydroxide and ammonia gas. Heating the product again causes the loss of water and conversion of the hydroxide to the oxide.

The unbalanced equations are:



Pre-Lab Questions

1. If the mass percent of each element in a compound is known, what steps are taken to determine the compound's empirical formula?
2. If the empirical formula of a compound is known, what additional information is required to determine the molecular formula of the compound?
3. Is the reaction of magnesium metal and oxygen gas an oxidation-reduction reaction? If so, what is the change in oxidation number of each type of atom?
4. What is the theoretical yield in grams of MgO if 2.37 g Mg metal reacts with excess O_2 ? What is the theoretical yield of Mg_3N_2 if the same amount of Mg reacts with excess N_2 ?

Purpose

To determine the empirical formula of magnesium oxide, and to reinforce the concepts of the law of mass conservation and the law of multiple proportions.

Apparatus and Materials

- Safety goggles
- Magnesium ribbon, Mg
- Balance (to 0.01 g or better)
- Ring stand
- Bunsen burner
- Ring support with clay triangle
- Crucible with lid
- Crucible Tongs
- Heat resistant tile or pad

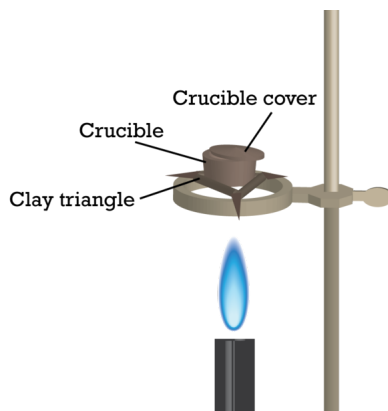
Safety Issues

The crucible and all of the apparatus gets very hot. The ammonia emitted during the secondary reaction is irritating. Open flames can be dangerous. Do not place a hot crucible on an electronic balance. It can damage the electronics. In addition, a hot crucible causes the air above it to become buoyant. If placed on a balance, the buoyant air will cause a mass reading, which is less than the actual mass.

Procedure

1. Heat the empty crucible and lid for about 3 *minutes* to remove water, oils, or other contaminants and to make sure there are no cracks. The bottom of the crucible should glow red-hot for about 20 *seconds*. Remove the flame and cool the crucible with lid.

- Record the mass of crucible and lid once it has cooled. Handle the crucible with tongs.
- Obtain about 0.3 g (35 cm) magnesium ribbon (do not handle the ribbon with your hands). Cut the magnesium into 0.5 – 1.0 cm pieces with scissors.
- Record the mass of the magnesium ribbon, lid and crucible.
- Place the crucible securely on the clay triangle. Set the lid slightly off-center on the crucible to allow air to enter but to prevent the magnesium oxide from escaping.



- Place the Bunsen burner under the crucible, light it, and heat the bottom of the crucible with a gentle flame for about 1 *minute*; then, place the burner under the crucible and heat strongly.
- Heat until all the magnesium turns into gray-white powder (around 10 *minutes*).
- Stop heating and allow the crucible, lid and contents to cool.
- Add about 1 ml (approx. 10 drops) of distilled water directly to the solid powder. Carefully waft some of the gas that is generated toward your nose, but be very careful. Record any odor.
- Heat the crucible and contents, with the lid slightly ajar, gently for about 2 *minutes* and then strongly for about another 3 to 5 *minutes*.
- Allow the crucible to cool and then record the mass of the crucible, lid and contents.
- Follow instructions for oxide disposal given by your teacher. Clean all equipment thoroughly.

Data

Mass of crucible and lid = _____ g

Mass of the crucible, crucible lid, and the magnesium = _____ g

Mass of the crucible, crucible lid, and magnesium oxide = _____ g

Post-Lab Questions

- Determine the mass of magnesium ribbon used in the experiment by subtracting the mass of the crucible and lid from the mass of the crucible, lid, and magnesium.

Mass of magnesium = _____ g

- Determine the number of moles of magnesium used. Remember: $\frac{\text{mass}}{\text{atomic weight}} = \text{number of moles}$. The atomic weight of magnesium is 24.3 g/mol.

Number of moles of magnesium = _____ mole

- Determine the mass of magnesium oxide that was formed by subtracting the mass of the mass of the crucible and lid from the mass of the crucible, lid, and magnesium oxide.

Mass of magnesium oxide formed = _____ g

4. Determine the mass of oxygen that combined with the magnesium.

Mass of oxygen = mass of magnesium oxide - mass of magnesium

Mass of oxygen that combined with the magnesium = _____ g

5. Determine the number of moles of oxygen atoms that were used. This is elemental oxygen so use 16.0 g/mol for the atomic weight.

Number of moles of oxygen atoms that were used = _____ mole

6. Calculate the ratio between moles of magnesium atoms used and moles of oxygen atoms used. Remember, this is simple division. Divide the number of moles of magnesium by the number of moles of oxygen. Round your answer to the nearest whole number, as we do not use part of an atom. This represents the moles (and also atoms) of magnesium. The moles (and also atoms) of oxygen, are represented by 1, because it was on the bottom of the division.

Moles of Magnesium : Moles of Oxygen

_____:

Teacher's Pages for Water of Hydration Lab

Lab Notes

This lab is fairly straightforward and simple in execution. The biggest hazard is with the crucibles. They become very, very hot – enough to weld your skin on to it. Make sure to handle hot crucibles with crucible tongs, and never to weigh a hot crucible. Teach your students how to gauge the temperature of a piece of glassware by having them approach the hot item with the back of their hand. If they can bring their hand to within a centimeter of the piece of glassware and it is hot, it is too hot to handle! If it is too hot for the back of your hand, it is certainly too hot for the front. Instruct the students on the proper procedure for using a desiccator.

Water of Hydration Lab

Background Information

When ionic crystals form, they often incorporate water molecules within their structure. This water within the crystal is called the water of hydration, and the compounds themselves are called hydrates. This water of hydration can often be removed by simply heating the hydrate, because the water molecules are only weakly attracted to the ions present. When this is done, the resulting leftover substance is called the anhydrous form of the crystal.

The amount of water present is often in a whole-number stoichiometric amount relative to the anhydrous form. Examples include barium chloride-2-hydrate, $BaCl_2 \cdot 2H_2O$, and cobalt nitrate-6-hydrate, $Co(NO_3)_2 \cdot 6H_2O$. Hydration numbers are most often integers, but in calcium compounds they are often fractional. Portland cement is an example of an important hydrate. When water is introduced to the anhydrous form, the water incorporates into the structure, and the cement hardens. Since the removal of water requires an input of heat, it should not be surprising that adding water to a hydrate gives off heat, and it gives off as much heat as was put into the system to remove the water in the first place. This is a problem for civil engineers who pour large amounts of cement: the heat given off by the hardening cement can be so great as to break the cement that has already hardened, due to heat stress. Steps must be taken to remove this heat. The Hoover Dam is such a large piece of concrete (cement + aggregate) that the dam is still cooling – and the last of the cement was poured in 1935.

Purpose

To determine the hydration number and empirical formula of copper(II) sulfate hydrate.

Apparatus and Materials

- Ring Stand and Ring
- Crucible
- Clay Triangle
- Bunsen burner
- Wash bottle
- Matches
- Electronic Balance
- Copper(II)sulfate hydrate (approximately 3 g per lab group)
- desiccator
- watch glass

Safety Issues

Always handle crucibles with crucible tongs. Never place a hot crucible on a balance. It can damage the electronics and give a measurement, which is less than the actual mass.

Procedure

1. Clean a porcelain crucible with soap and water. Rinse and dry the crucible by placing the crucible and cover on a clay triangle over a laboratory burner and heating until red-hot.
2. Carefully remove the crucible and cover with crucible tongs and let it cool. Handle the crucible and cover with tongs for the remainder of the experiment.
3. Measure the mass of the empty crucible and cover to the nearest 0.01 *gram*.
4. Add about 3 g of $CuSO_4$ hydrate crystals to the crucible, replace the cover, and measure the mass to the nearest 0.01 g.
5. Begin heating slowly. Increase the heat until you have heated the crucible strongly for about 10 *minutes*.
6. Remove the crucible from the triangle support, let it cool in a desiccator, and measure the mass.
7. Reheat with a hot flame for a few minutes, cool, and measure the mass again. If the mass is different from that recorded in Step 6, continue to heat and measure until the masses agree.
8. Remove the button of anhydrous copper sulfate by tipping it into a watch glass. Add a few drops of water to the anhydrous copper sulfate, and record your observations below.

TABLE 12.6: Data

Number	Object	Mass (grams)
1.	Mass of Crucible + Cover	_____ grams
2.	Mass of Crucible + Cover + Copper Sulfate Hydrate	_____ grams
3.	Mass of Crucible + Cover + Anhydrous Copper Sulfate	_____ grams
4.	Mass of Water (2 – 3)	_____ grams
5.	Mass of Anhydrous Copper Sulfate (3 – 1)	_____ grams
6.	Moles of Anhydrous Copper Sulfate	_____ moles
7.	Moles of water driven off	_____ moles
8.	Ratio of moles of water to moles of anhydrous $CuSO_4$ ($\frac{7}{6}$)	

9. Describe the behavior of anhydrous copper sulfate when water is added.

Post-Lab Questions

1. Compare the number of moles of anhydrous $CuSO_4$ to the number of moles of water in the hydrate. Use the ratio of these two values to predict a formula for the hydrated $CuSO_4$.
2. Why is it necessary to let the crucible cool before measuring mass? Why should the mass of the crucible be measured immediately after the crucible cools, and not later?
3. In this experiment, you cooled your crucible in a desiccator. What is a desiccator? How does a desiccator work?
4. How would your experimental results be affected if you did not use a desiccator when cooling the crucible and contents?
5. How can you account for the behavior of the anhydrous form of the copper sulfate when the water was added? What do you think the new substance is?

Demonstrations for Chapter 12

12.5 Worksheets for Chapter 12

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Calculating Molar Masses Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

The relative masses of atoms, in units called Daltons, are listed in the periodic table. The relative masses of molecules, in the same units, can be determined by adding up the masses of all the atoms that make up the molecule. For example, the periodic table lists the relative mass of a hydrogen atom as 1.01 *Dalton* and relative mass of the an oxygen atom to be 16.00 *Daltons*. Therefore, on this same scale, the relative mass of a water molecule, H_2O , would be the sum of two hydrogen atoms and one oxygen atom, $1.01 + 1.01 + 16.00 = 18.02$ *Daltons*.

When an Avogadro's number, 6.02×10^{23} , of atoms or molecules are taken, the mass of the group will be the same number as the relative mass, but the units will be grams. That is, the mass in grams, of 6.02×10^{23} water molecules is 18.02 *grams*. An Avogadro's number of particles is called one **mole** and the mass of that group of particles is called the **molar mass** (or mass of one mole) of that substance.

Example: Find the molar mass of calcium phosphate, $Ca_3(PO_4)_2$.

TABLE 12.7: Adding Up a Molar Mass

Atoms of Element	Atoms \times Atomic Mass	Total
3 <i>Ca atoms</i> =	3×40.1	= 120.3
2 <i>P atoms</i> =	2×31.0	= 62.0
8 <i>O atoms</i> =	8×16.0	= 128.0
		310.3

Therefore, the molar mass of calcium phosphate is 310.3 *grams/mole*.

Exercises

Find the molar masses of the following compounds. (Do not fail to include units in your answers.)

1. $NaOH$
2. $NaBr$
3. $PbSO_4$
4. $Ca(OH)_2$
5. AgF
6. $C_6H_{12}O_6$
7. $Ba(C_2H_3O_2)_2$
8. $ZnCl_2$
9. $(NH_4)_2SO_4$
10. $(NH_4)_3PO_4$

Moles Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

An Avogadro's number of particles of a substance is called *one mole* of that substance. When an Avogadro's number, 6.02×10^{23} , of atoms or molecules are taken, the mass of the group will be the same number as the relative molecular mass, but the units will be grams. The mass of one mole of a substance (6.02×10^{23} particles) is the relative molecular mass **in grams**.

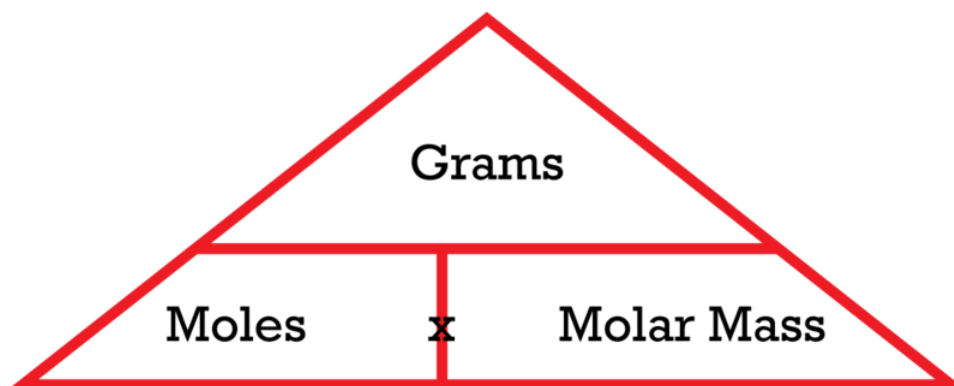
The relationship between the moles and mass of a substance is given by:

$$\text{grams} = (\text{moles})(\text{molar mass})$$

This relationship can be solved for any one of the three variables in the expression.

$$\text{grams} = (\text{moles})(\text{molar mass}) \qquad \text{moles} = \frac{\text{grams}}{\text{molar mass}} \qquad \text{molar mass} = \frac{\text{grams}}{\text{moles}}$$

Some students find the triangle below to be a useful crutch. You put your thumb over the quantity you are solving for and the part of the triangle not covered shows the correct formula.



Example 1: How many moles are present in 10.0 *grams* of sodium hydroxide, NaOH ?

Solution: The molar mass of NaOH is 40.0 *g/mol*. (“mol” is the abbreviation of mole.)

$$\text{moles} = \frac{\text{grams}}{\text{molar mass}} = \frac{10.0 \text{ g}}{40.0 \text{ g/mol}} = 0.250 \text{ moles}$$

Example 2: What is the mass, in grams, of 4.20 *moles* of $\text{Ca}(\text{NO}_3)_2$?

Solution: The molar mass of $\text{Ca}(\text{NO}_3)_2$ is 164.1 *g/mol*.

$$\text{grams} = (\text{moles})(\text{molar mass}) = (4.20 \text{ mol})(164.1 \text{ g/mol}) = 689 \text{ grams}$$

Example 3: What is the molar mass of an unknown substance if 0.250 *moles* of the substance has a mass of 52.6 *grams*?

Solution:

$$\text{molar mass} = \frac{\text{grams}}{\text{moles}} = \frac{56.2 \text{ g}}{0.250 \text{ mol}} = 225 \text{ g/mol}$$

Example 4: What is the mass of 3.01×10^{23} molecules of ammonia, NH_3 ?

Solution: This problem involves converting the number of molecules to moles (divide by Avogadro's number), and then multiplying the moles by the molar mass.

$$\text{mass} = (3.01 \times 10^{23} \text{ molecules}) \left(\frac{1.00 \text{ mol}}{6.02 \times 10^{23} \text{ molecules}} \right) \left(\frac{17.0 \text{ g}}{1.00 \text{ mol}} \right) = 8.50 \text{ grams}$$

Exercises

1. How many moles are present in 5.00 grams of NaOH ?
2. How many grams are present in 2.50 moles of NH_3 ?
3. How many moles are present in 100. g of $\text{Ca}(\text{NO}_3)_2$?
4. What is the mass of 0.468 moles of $\text{C}_6\text{H}_{12}\text{O}_6$?
5. How many moles are present in 1.00×10^{24} molecules of water?
6. What is the mass, in grams, of one molecule of water?
7. What is the molar mass of a substance if 0.336 moles of it has a mass of 70.0 grams?
8. Convert 4.00 grams of CH_4 to moles.
9. Convert 4.00 moles of CH_4 to grams.
10. How many molecules are present in 1.00 g of $\text{Al}(\text{C}_2\text{H}_3\text{O}_2)_2$

Percent Composition Worksheet**CK-12 Foundation Chemistry**

Name _____ Date _____

The percent composition (or percentage composition) of a compound is a measure of the percentage of each different element present in the compound. To calculate the percent composition, the mass of each individual element is divided by the total mass of the compound and then multiplied by 100 (to get its percentage). The percent composition of a compound can be calculated either from the known masses of the elements in the compound (determined in the lab) or from the formula of the compound.

Example: The composition of a compound is determined in the laboratory to be 5.748 grams of sodium and 8.862 grams of chlorine. What is the percent composition of the compound?

Solution: The total mass of this sample of the compound is 14.61 grams.

$$\begin{aligned} \% \text{ sodium} &= \frac{5.748 \text{ g}}{14.61 \text{ g}} \times 100 = 39.34\% \\ \% \text{ chlorine} &= \frac{8.862 \text{ g}}{14.61 \text{ g}} \times 100 = 60.66\% \end{aligned}$$

When you add up all the percentages of elements, you should get 100%, although on many occasions, rounding may cause the last digit of the total to be off by 1. That is, on occasion, you get a total of 99.9% or 100.1% due to several individual percentages all being rounded up or all being rounded down.

Example: Calculate the percent composition of all the elements in $(NH_4)_3PO_4$.

Solution:

$$3 \text{ N atoms} = 3 \times 14.01 = 42.03$$

$$12 \text{ H atoms} = 12 \times 1.01 = 12.12$$

$$1 \text{ P atom} = 1 \times 30.97 = 30.97$$

$$4 \text{ O atoms} = 4 \times 16.00 = 64.00$$

Formula weight for $(NH_4)_3PO_4 = 149.12$

$$\begin{aligned} \% \text{ N} &= \frac{42.03}{149.12} \times 100 = 28.19\% & \% \text{ P} &= \frac{30.97}{149.12} \times 100 = 20.77\% \\ \% \text{ H} &= \frac{12.12}{149.12} \times 100 = 8.13\% & \% \text{ O} &= \frac{64.00}{149.12} \times 100 = 42.92\% \end{aligned}$$

When the four percentages are added in this case, the total is 100.01%. The extra 0.01% is due to the fact that all four of these percentages were rounded up.

Exercises

1. Determine the percent composition of Na_2SO_4 .
2. Determine the percent composition of $NaOH$.
3. Determine the percent composition of $AlCl_3$.
4. Determine the percent composition of $Ca(C_2H_3O_2)_2$.
5. Determine the percent composition of $C_6H_{12}O_6$.

Empirical Formulas Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Empirical formulas represent the simplest whole number ratio of the atoms that make up a compound. In some cases, such as CO_2 , the empirical formula is exactly the same as the actual molecular formula. In other cases such as benzene, C_6H_6 , whose empirical formula is CH , the molecular formula is some multiple of the empirical formula.

Empirical formulas can be determined either from the masses of the elements making up the compound or from the percent composition.

Example 1: What is the empirical formula of a compound that contains 0.0134 grams of iron, 0.00769 grams of sulfur, and 0.0115 grams of oxygen?

Step 1: Convert each of the masses into moles of atoms of that element. This is accomplished by dividing the grams of each element by the atomic mass of the element.

$$\text{moles Fe} = \frac{0.0134 \text{ g}}{55.8 \text{ g/mol}} = 0.000240 \text{ mol}$$

$$\text{moles S} = \frac{0.00769 \text{ g}}{32.1 \text{ g/mol}} = 0.000240 \text{ mol}$$

$$\text{moles O} = \frac{0.0115 \text{ g}}{16.0 \text{ g/mol}} = 0.000719 \text{ mol}$$

It is important to note that we are determining the number of moles of each atom that exists in the compound and therefore, for the diatomic gases, we use the atomic mass of a single atom of the element (not the diatomic molar mass).

Step 2: The ratio of moles that we determined in step 1 is the correct ratio for the compound. We are not allowed, however, to write a formula in the form, $Fe_{0.000230}S_{0.000240}O_{0.000719}$. Before we can write the formula, we must get the ratio into a simplest whole number ratio. This is often accomplished by dividing each of the moles by the smallest of them.

$$\text{moles Fe} = \frac{0.000240}{0.000240} = 1.00$$

$$\text{moles S} = \frac{0.000240}{0.000240} = 1.00$$

$$\text{moles O} = \frac{0.000719}{0.000240} = 3.00$$

Therefore, the empirical formula for this compound is $FeSO_3$.

Example 2: Find the empirical formula of a compound that contains 48.78% carbon, 2.439% hydrogen, 26.02% oxygen, and 22.77% nitrogen.

Solution: When the empirical formula is to be determined from percent composition, it is easiest to assume a 100. gram sample, take each percentage of the 100. grams to get grams for each element, and then proceed as in **Example 1**. Using this technique, each of the percentages in the problem becomes the mass of the element in grams.

$$\text{grams C} = 48.78 \text{ g, grams H} = 2.439 \text{ g, grams O} = 26.02 \text{ g, grams N} = 22.77 \text{ g}$$

Step 1:

$$\text{moles C} = \frac{48.78 \text{ g}}{12.01 \text{ g/mol}} = 4.062 \text{ mols}$$

$$\text{moles H} = \frac{2.439 \text{ g}}{1.01 \text{ g/mol}} = 2.415 \text{ mols}$$

$$\text{moles O} = \frac{26.02 \text{ g}}{16.00 \text{ g/mol}} = 1.626 \text{ mols}$$

$$\text{moles N} = \frac{22.77 \text{ g}}{14.01 \text{ g/mol}} = 1.625 \text{ mols}$$

Step 2: Divide each of the moles by the smallest.

$$\begin{aligned} \text{moles C} &= \frac{4.062}{1.625} = 2.50 \\ \text{moles H} &= \frac{2.415}{1.625} = 1.49 \\ \text{moles O} &= \frac{1.626}{1.625} = 1.00 \\ \text{moles N} &= \frac{1.625}{1.625} = 1.00 \end{aligned}$$

Step 3: In a case, such as this one, where step 2 does NOT produce a simple whole number ratio, we then choose a multiplier with which to multiply each of the final numbers such that we do get a simple whole number ratio. This is usually an integer between 2 and 5 but could possibly be a larger integer. In this case, the multiplier is 2.

$$\text{moles C} = 5, \text{moles H} = 3, \text{moles O} = 2, \text{moles N} = 2$$

Therefore, the empirical formula for this compound is $C_5H_3O_2N_2$.

Exercises

- Find the empirical formula for a compound that is 75.0% carbon and 25.0% hydrogen.
- Find the empirical formula for a compound that is 32.8% chromium and 67.2% chlorine.
- Find the empirical formula for a compound that is 67.1% zinc and the rest oxygen.
- A sample of a compound was found to contain 0.62069 g of carbon, 0.10345 g of hydrogen, and 0.27586 g of oxygen. What is the empirical formula?
- A sample of a compound was found to contain 48.65% carbon, 8.11% hydrogen, and 43.24% oxygen. What is its empirical formula?

Molecular Formulas Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Empirical formulas show the simplest whole number ratio of the atoms of the elements that make up a compound. Molecular formulas show the actual number of atoms of each element that make up the compound. The molecular formula for benzene is C_6H_6 but the empirical formula for benzene would be the simplest whole number ratio for these atoms, which would be CH . The empirical formula can be determined from either the masses of the elements in a compound or from percent composition. In order to determine the molecular formula, we also need the molar mass of the compound. The molecular formula will always be some whole number multiple of the empirical formula. In the case of benzene, the multiplier is 6.

The molecules C_2H_4 , C_3H_6 , C_4H_8 , and C_5H_{10} all have the same empirical formula, namely CH_2 . If we have the empirical formula CH_2 and the molar mass of 56 g/mol for a compound, we can determine the molecular formula by dividing the formula mass of CH_2 into the molar mass to find the multiplier. The formula mass of CH_2 is 14 g/mol. When we divide the formula mass, 14 g/mol, into the molar mass, 56 g/mol, we get the multiplier 4. Therefore, the molecular formula for this compound is 4 times the empirical formula. $CH_2 \times 4 = C_4H_8$.

Example: What is the molecular formula for a compound with the empirical formula HCO_2 and a molar mass of 90. g/mol?

Solution: The formula mass of HCO_2 is 45 g/mol . Dividing 45 g/mol into $90. \text{ g/mol}$ yields a multiplier of 2. Therefore, the molecular formula for this compound is $2 \times CHO_2 = H_2C_2O_4$.

Exercises

1. A compound has the empirical formula C_2OH_4 and a molar mass of 88 g/mol . What is its molecular formula?
2. A compound has the empirical formula C_4H_4O and a molar mass of 136 g/mol . What is its molecular formula?
3. A compound has the empirical formula $CBrO$ and a molar mass of 254.7 g/mol . What is its molecular formula?
4. A compound is 7.692% hydrogen and 93.308% carbon. Its molar mass is 104 g/mol . What is its molecular formula?
5. A compound is 47.0% potassium, 14.5% carbon, and 38.5% oxygen. Its molar mass is 166.2 g/mol . What is its molecular formula?

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

12.6 Extra Reading for Chapter 12

Avogadro's Number

1811 was the year that Lorenzo Romano Amedeo Carlo Avogadro de Quaregna e di Cerreto - better known as Amedeo Avogadro, published his now famous hypothesis. It stated that equal volumes of gases contain the same number of particles. The nature of those particles was still a topic of considerable debate. Avogadro produced his theory based on the results of Joseph-Louis Gay-Lussac, who showed that when different gases combine, they do so in simple whole number ratios. His contemporary, John Dalton, responsible for the similar sounding Law of Multiple Proportions, reacted critically to Gay-Lussac's work. Dalton suggested that the atoms in gases were not capable of attaching; he argued that they would repel each other. Avogadro recognized that the viewpoints of both Dalton and Gay-Lussac could both be operable if, in his words, the same volume of gas contained the same number of molecules. It must be understood that the distinction between atoms and molecules did not exist in 1811 and the two words were used interchangeably.

Avogadro's principle did not gain adherents until the concept of the atom became more solidly established. Likewise, the actual determination of what has become known as Avogadro's number, was not accomplished until after Avogadro's death in 1856. Johann Josef Loschmidt, an Austrian chemist, developed a method for the first estimate of the actual number. His technique entailed measuring the difference in volume between a given liquid substance, and the volume of that material upon evaporation into the gas phase. He reasoned that in the liquid phase, all of the liquid molecules touched their adjacent molecules and that there was no empty space. Thus the total volume of the liquid was equivalent to the volume of all of the liquid molecules added together. Comparing the volumes of the liquid and gas phases, he estimated that there were about 5×10^{22} molecules in a volume of gas. By defining the number of molecules in a cubic meter of gas at standard temperature and pressure, he derived what is now known as "Loschmidt's number" or 2.686×10^{25} .

The establishment of a more carefully calculated value for the number of particles in one mole of any substance was made by Albert Einstein in the early twentieth century. Rather than using a gas model, Einstein's method was based on evaluating the number of sugar molecules in a sample of sugar water. His calculation was based on the average velocity that the individual molecules diffused through a membrane and was initially published as 2.1×10^{23} . Several years later, with new data from more accurate measurements, he redefined the value as 6.1×10^{23} .

The value now used in chemistry texts, 6.022×10^{23} , was arrived at by a different technique. The current value has been calculated by using x-ray diffraction in crystal lattices of silicon atoms.

12.7 Assessment for Chapter 12

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 13

Chemical Reactions C-SE-TE

Chapter Outline

- 13.1 CHEMICAL EQUATIONS
- 13.2 BALANCING EQUATIONS
- 13.3 TYPES OF REACTIONS
- 13.4 WORKSHEETS FOR CHAPTER 13
- 13.5 EXTRA READING FOR CHAPTER 13
- 13.6 ASSESSMENT FOR CHAPTER 13

Lessons and Number of Activities for Lessons

TABLE 13.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Chemical Equations	0	1	0	0
2. Balancing Equations	1	0	1	0
3. Types of Reactions	2	1	1	1

13.1 Chemical Equations

Student Behavioral Objectives

The student will:

- read chemical equations, and provide requested information contained in the equation including information about substances, reactants, products, and physical states.
- convert symbolic equations into word equations and vice versa.
- use the common symbols, +, (s), (L), (g), (aq), and \rightarrow appropriately.
- describe the roles of subscripts and coefficients in chemical equations.

Timing, Standards, Activities

TABLE 13.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Chemical Equations</i>	0.5	3a

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. Briggs-Rauscher Oscillating Reaction

Worksheets

- 1.

Extra Readings

1. None

Answers for Chemical Equations (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

13.2 Balancing Equations

Student Behavioral Objectives

The student will:

- demonstrate the Law of Conservation of Matter in a chemical equation.
- explain the roles of coefficients and subscripts in a chemical equation.
- balance equations using the simplest whole number coefficients.

Timing, Standards, Activities

TABLE 13.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Balancing Equations</i>	2.0	3a

Activities for Lesson 2

Laboratory Activities

1. Chemical Reactions in Microscale

Demonstrations

1. None

Worksheets

1. Balancing Equations Worksheet

Extra Readings

1. None

Answers for Balancing Equations (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

13.3 Types of Reactions

Student Behavioral Objectives

The student will:

- identify the types of reactions.
- predict the products in different types of reactions.
- distinguish between the different types of reactions.
- write balanced chemical equations and identify the reaction type given only the reactants.

Timing, Standards, Activities

TABLE 13.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Types of Reactions</i>	2.0	None

Activities for Lesson 3

Laboratory Activities

1. Types of Chemical Reactions
2. Chemical Reactions Using Probeware

Demonstrations

1. Explosive Mixtures of Ethyne and Air

Worksheets

1. Types of Chemical Reactions Worksheet

Extra Readings

1. Fireworks

Answers for Types of Reactions (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 13

This website provides an equation balancing activity.

- <http://www.middleschoolscience.com/balance.html>

This website provides a balancing chemical equations game.

- <http://funbasedlearning.com/chemistry/chemBalancer/default.htm>

This video shows ten amazing chemical reactions that are fun to watch but dangerous to carry out.

- <http://listverse.com/2008/03/04/top-10-amazing-chemical-reactions/>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 8-1 is on balancing equations.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson81.htm>

This video contains classroom demonstrations of several reaction types and then shows the balancing process for the reaction equations.

- <http://www.youtube.com/watch?v=4B8PFqbMNIw>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 8-2 is on the classification of chemical reactions.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson82.htm>

The website below also reviews the different types of chemical reactions.

- <http://library.thinkquest.org/2923/react.html>

Laboratory Activities for Chapter 13

Chemical Reactions in Microscale

Purpose

To write balanced chemical reactions, identifying the type of reaction, and the physical characteristics that indicate a reaction has taken place.

Background

Within the five main types of chemical reactions studied in the *Chemical Reactions* chapter, four of these (synthesis, decomposition, single replacement, and double replacement), have subgroups of reactions that can be classified. These subgroups are known as precipitation reactions, neutralization reactions, combustion reactions, and the like. Physical changes such as: the formation of a precipitate (hence the precipitation reaction subclass), change in color,

gas formation, change in temperature, tell us that a reaction has taken place. As well, within each of these reactions, whether an observable physical change has occurred or not, the Law of Conservation of mass is always maintained.

In this lab you are given seven nitrate solutions with which you are going to react seven sodium and one ammonium solutions. Planning is everything! Added to this you are only working with 10 drops of solution in total.

Materials

- 0.1 mol/L $Cu(NO_3)_2$
- 0.1 mol/L Na_2CO_3
- 0.1 mol/L $Pb(NO_3)_2$
- 0.1 mol/L Na_2SO_4
- 0.1 mol/L $Ni(NO_3)_2$
- 0.1 mol/L $NaCl$
- 0.1 mol/L $Co(NO_3)_3$
- 0.1 mol/L NaI
- 0.1 mol/L $Fe(NO_3)_3$
- 0.1 mol/L Na_2CrO_4
- 0.1 mol/L $AgNO_3$
- 0.1 mol/L $Na_2Cr_2O_7$
- 0.1 mol/L HNO_3
- 0.1 mol/L $NaOH$
- 0.1 mol/L NH_4OH
- H_2O
- 24-well micro plate
- toothpicks
- beral pipettes
- 24-well micro plate (x 4)

TABLE 13.5: Data Table

	1	2	3	4	5	6	7	8
A								
B								
C								
D								
E								
F								
G								

Safety

Avoid contact with the solutions. If solutions get on your skin, rinse the area thoroughly with running water.

Procedure

Part 1: Place 4 micro plates in a grid so that you have at least 8 wells in a line and 7 lines down.

Part 2: In wells A1 through to A8, add 5 drops of 0.1 mol/L $Cu(NO_3)_2$.

Part 3: In wells B1 though B8, add 5 drops of 0.1 mol/L $Pb(NO_3)_2$.

Part 4: In wells C1 though C8, add 5 drops of 0.1 mol/L $Ni(NO_3)_2$.

Part 5: In wells D1 though D8, add 5 drops of 0.1 mol/L $Co(NO_3)_3$.

Part 6: In wells E1 though E8, add 5 drops of 0.1 mol/L $Fe(NO_3)_3$.

Part 7: In wells F1 through F8, add 5 drops of 0.1 mol/L $AgNO_3$.

Part 8: In wells G1 through G8, add 5 drops of 0.1 mol/L HNO_3 .

Part 9: In wells A1 through G1, add 5 drops of 0.1 mol/L Na_2CO_3 .

Part 10: In wells A2 through G2, add 5 drops of 0.1 mol/L Na_2SO_4 .

Part 11: In wells A3 through G3, add 5 drops of 0.1 mol/L $NaCl$.

Part 12: In wells A4 through G4, add 5 drops of 0.1 mol/L NaI .

Part 13: In wells A5 through G5, add 5 drops of 0.1 mol/L Na_2CrO_4 .

Part 14: In wells A6 through G6, add 5 drops of 0.1 mol/L $Na_2Cr_2O_7$.

Part 15: In wells A7 through G7, add 5 drops of 0.1 mol/L $NaOH$.

Part 16: In wells A8 through G8, add 5 drops of 0.1 mol/L NH_4OH .

Part 17: Record all of your observations into your data table.

Part 18: Clean Up. Empty the contents of the micro plate into the sink and rinse the plate and the sink with plenty of water. Wash your hands and the container thoroughly.

Analysis

1. Which reactions resulted in the formation of a precipitate?
2. Write balanced chemical equations for the reactions found in question 1. Can you determine based on your observations in this lab what the precipitate is likely to be? If so, indicate that in your chemical reaction.
3. What other physical changes were observed?
4. Write the balanced chemical equations for these reactions.
5. What three questions do you have as a result of doing this experiment?

Conclusion

What conclusions did you make as a result of doing this experiment?

Chemical Reactions using Probeware

An alternate method of performing an experiment to demonstrate chemical reactions would be to use technology such as that found with the DataStudio, the GLX XPlorer, or with Vernier probeware. All three of these technology alternatives come complete with labs manuals or you can go online to <http://www.pasco.com/experiments/chemistry/home.html> to find other labs.

More specifically to reactions and the Law of Conservation of Mass, Pasco provides a lab such as the one entitled Classical Exothermic Reaction: Steel Wool & Vinegar at http://www.pasco.com/experiments/chemistry/october_2002/home.html

In this lab handout, students pour vinegar over steel wool and measure the change in mass and temperature as the steel wool reacts with the vinegar. Instructions for both the DataStudio 500 Interface and the XPlorerGLX are given.

Types of Chemical Reactions

Pre-lab Questions

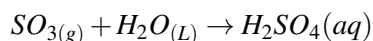
1. How does a decomposition reaction differ from the other types of chemical reactions?

2. In a combination reaction, what other products form in addition to any new compound?
3. Which two ions are bioaccumulative and should be used in very small quantities?
4. During which steps (give numbers) of the procedure would you expect to produce a gas?
5. What safety precaution applies when heating a test tube?

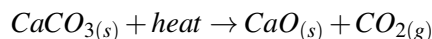
Introduction

In this experiment, you will learn to differentiate among five general types of chemical reactions. You will carry out certain representative reactions yourself, while others will be demonstrated by your teacher. From your observations you will attempt to identify the products of each reaction, and to determine the type of reaction that has taken place. The types of reactions that you will consider are the following: synthesis reactions, decomposition reactions, single replacement reactions, double replacement reactions, and combustion reactions. The majority of common chemical reactions can be classified as belonging to one of these categories. A brief description of each reaction type is provided below.

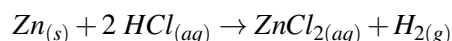
(a) Synthesis reactions are reactions in which two or more substances combine to form a single product. The reactants may be elements or compounds, but the product is always a single compound. An example of a combination reaction is a reaction of sulfur trioxide and water to form sulfuric acid.



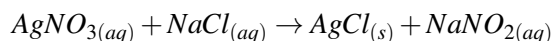
(b) Decomposition reactions are reactions in which a single substance breaks down into two or more simpler substances. There is always just a single reactant in a decomposition reaction. An example of a decomposition reaction is the breakdown of calcium carbonate upon heating.



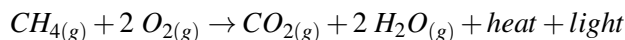
(c) Single replacement reactions are reactions in which an element within a compound is displaced to become a separate element. This type of reaction always has two reactants, one of which is always an element. An example of a single replacement reaction is the reaction of zinc metal with hydrochloric acid.



(d) Double replacement reactions are reactions in which a positive ion from one ionic compound exchanges with the positive ion from another ionic compound. These reactions typically occur in aqueous solution and result in either the formation of a precipitate, the production of a gas, or the formation of a molecular compound such as water. An example of a double-replacement reaction is the reaction that occurs between aqueous silver nitrate and aqueous sodium chloride. A precipitate of solid silver chloride is formed in the reaction.



(e) Combustion reactions are reactions in which an element or compound reacts rapidly with oxygen gas to liberate heat and light energy. Commonly, the compounds combining with oxygen in these reactions are hydrocarbons, compounds containing hydrogen and carbon. The well-known combustible fuels kerosene and gasoline, for instance, are hydrocarbon mixtures. The complete combustion of a hydrocarbon yields carbon dioxide and water as well as reaction products. If insufficient oxygen is available, combustion will not be complete and carbon monoxide and elemental carbon may be obtained as additional products in the reaction. An example of a combustion reaction is the burning of methane gas to give water (in the form of steam), carbon dioxide, heat, and light.

**Objectives:**

- To observe chemical reactions in order to determine the reaction type.
- To write balanced chemical equations for each reaction.

Apparatus and Materials

- Iron filings
- safety glasses
- Copper(II) sulfate pentahydrate
- 2 × small test tubes
- Magnesium, turnings
- 2 × medium test tubes
- 0.1 mol/L copper(II) sulfate
- 1 × large test tube
- 0.2 mol/L lead(II) nitrate
- 1 × test tube holder
- 0.2 mol/L potassium iodide
- gas burner
- 3% hydrogen peroxide
- ring stand & clamp
- 5 mol/L hydrochloric acid
- dropper pipette
- 3% sulphuric acid (teacher demo)
- crucible tongs
- sodium bicarbonate (teacher demo)
- electrolysis apparatus (teacher demo)
- limewater (teacher demo)
- 1-holed rubber stopper (teacher demo)
- toothpicks
- 1 glass tube, 25 cm long bent at 90° in
- matches

Procedure

1. Draw a table similar to the one below to use for collecting your observations.

TABLE 13.6: Observations

Reaction	Observations	Reaction Type
<i>Fe</i> and <i>CuSO</i> ₄		
<i>Pb(NO</i> ₃ <i>)</i> ₂ and <i>KI</i>		
<i>CuSO</i> ₄ · 5 <i>H</i> ₂ <i>O</i> and heat		
<i>Mg</i> and <i>HCl</i>		
<i>H</i> ₂ <i>O</i> ₂ and heat		
Electrolysis of <i>H</i> ₂ <i>O</i>		
<i>NaClO</i> ₂ and heat		

2. **Iron metal and copper(II) sulfate solution.** Fill a small test tube halfway with copper(II) sulfate solution. Add 2 g (about $\frac{1}{4}$ of a small test tube) of iron filings to the solution. Observe the reaction after 5 *minutes*. Record your observations in the observation table. Discard the solid contents of the test tube into the waste container provided. The liquid portion can be poured down the sink.

3. **Lead(II) nitrate and potassium iodide solutions.** Put 2 mL of lead(II) nitrate solution in the test tube. Add 5 to 10 drops of potassium iodide solution. Record your observations. Discard the contents of the test tube into the waste container and rinse the tube with water.

4. **Action of heat on copper(II) sulfate solution.** Put two or three pea-sized crystals of copper(II) sulfate pentahydrate into a large, dry test tube. Fasten a utility clamp to the upper end of the test tube. Hold the tube by the clamp so that it is almost parallel with the surface of the lab bench. **CAUTION: Do not point the open mouth of the tube at yourself or anyone else.** Heat the crystals gently at the bottom of the tube (where the crystals are located) in a burner flame for approximately 30 *seconds* recording your observations. When the test tube has cooled, discard its contents into the waste container provided.

5. **Magnesium metal and hydrochloric acid.** Fill one medium-size test tube halfway with 6 mol/L hydrochloric. **CAUTION: Hydrochloric acid is corrosive.** Place the test tube in the test tube rack. Put several pieces of magnesium turnings into the acid solution. If you observe a gas forming, test for its identity by holding a burning wood splint at the mouth of the test tube. Do not put the splint into the solution. Record your observations. Decant the liquid portion of the test tube contents into the sink; discard the solid into the waste container provided.

6. **Action of heat on hydrogen peroxide.** Add 2 mL of the 3% hydrogen peroxide solution to a medium test tube. Use a utility clamp to secure the tube to a ring stand. **CAUTION: Make sure that the mouth of the tube is pointed away from you and away from everyone else.** Heat the solution very gently. If you observe a gas forming, test for its identity by inserting a glowing wood splint at the mouth of the test tube. Do not put the splint into the solution. Record your observations. Rinse the contents of the test tube into the sink.

TEACHER DEMONSTRATIONS

7. **Action of electricity on water (Electrolysis).** Water can be broken down to its component elements by passing electricity through it. This process is called electrolysis. The apparatus used for this demonstration will be explained by your teacher. Make your observations of the reaction at several intervals during a period of 5 – 10 *minutes*.

8. **Action of heat on sodium bicarbonate.** Solid sodium bicarbonate will be heated strongly in a test tube for 2 *minutes*. The gas that is given off will be tested by exposing it to a burning splint, and by bubbling it through limewater (a saturated solution of calcium hydroxide, $Ca(OH)_2$). Record your observations of these tests.

Data Analysis

1. Decide which type of reaction is represented by each reaction observed in this experiment. Record your answers in your observation table.
2. Write a balanced chemical equation for each chemical reaction observed.
3. Although you did not work with any synthesis reactions in this experiment, can you describe one or give an example of one that you might have seen before or read about. Write a balanced equation for this reaction.

Results and Conclusions

1. Describe in your own words the five types of chemical reactions that were discussed in the introduction to this experiment. Explain how each type of reaction can be identified.
2. List the tests that were used in this experiment to identify gases.

Extension

1. Make a list of the reactions observed in previous experiments. Identify the types of reaction in as many cases as possible.

Demonstrations for Chapter 13

Briggs-Rauscher Oscillating Reaction Demonstration

Demonstration Notes

The chemistry in this reaction is too complicated to explain to students. In fact, it is too complicated for most teachers as well. It is best to use the demonstration as a, ‘Wow, look at that’ demonstration. In that sense, the demonstration could be used in the *Chemical Reactions* chapter, or the *Reaction Kinetics* chapter, or as an introduction to the acid-base indicators lesson.

Chloride contamination will cause this reaction to fail. Make sure your glassware is clean – alconox with a distilled water rinse should be sufficient.

Brief description of demonstration

Equal volumes (100 mL) of clear and colorless solutions (3) are added to a large beaker and stirred with a magnetic stirrer. The solutions first turn yellow, then very dark blue-purple, then goes to colorless again. This pattern repeats itself for several minutes: the first cycle takes about twenty seconds, then increases over a period of several minutes until the dark blue color remains.

Materials

- 16 g Malonic Acid, $CH_2(COOH)_2$
- 3.4 g Manganese(II) Sulfate 1-hydrate, $MnSO_4 \cdot H_2O$
- 0.3 g corn starch
- 20 g Sodium Thiosulfate 5-Hydrate, $Na_2S_2O_3 \cdot 5H_2O$
- 410 mL 30% Hydrogen Peroxide Solution, H_2O_2
- 43 g Potassium Iodate, KIO_3
- 4.3 mL 18 M sulfuric acid, H_2SO_4
- Magnetic Stirrer and Stir Bar
- Hot Plate
- Plastic or rubber gloves
- (3) 2 L Beakers
- Normally available lab glassware

Procedure

The procedure requires the preparation of three different solutions, which you can label solutions **a**, **b**, and **c**. Equal quantities of the prepared solutions are then mixed and stirred using the magnetic stirrer.

To prepare solution **a**, add 500 mL of distilled water into a 2 L beaker. Wearing gloves, add 410 mL 30% H_2O_2 to the water, and dilute to 1 L.

To prepare solution **b**, add 900 mL of distilled water to a 2 L beaker. Dissolve 43 g of KIO_3 into the water, and then add 4.3 mL of 18 M H_2SO_4 . Dilute the resulting solution to 1 L.

To prepare solution **c**, add 500 mL of water to a 2 L beaker. Dissolve 16 g of malonic acid and 3.4 g $MnSO_4 \cdot H_2O$ into this solution. In a separate 150 mL beaker, boil 50 mL of distilled water and add 0.3 g of corn starch. Stir the resulting mixture until it is suspended. It will be cloudy. Let the starch suspension cool until you can pick up the beaker safely, then add this suspension to solution **c** and stir. Dilute solution **c** to 1 L.

This procedure can be used for up to 6 classes, depending on the volume you use for each class. 100 mL of each solution for each class will yield 5 presentations.

Place a 400 mL beaker on a magnetic stirrer. Add 100 mL of solution a and 100 mL of solution **b** into the beaker, and stir until a small vortex is seen. Then pour 100 mL of solution **c** into the swirling mixture. The yellow to blue oscillation should begin almost immediately.

Hazards

KIO_3 is a strong oxidizing agent. 30% H_2O_2 and concentrated H_2SO_4 is a very strong oxidation agent and is corrosive. Molecular iodine is a reaction product, and is a strong irritant. Wear gloves, apron, and goggles when handling these materials. Malonic acid is a precursor to illegal drug synthesis, so place it in a locked space.

Disposal

Add 2 g of $Na_2S_2O_3$ to the reaction products to reduce the molecular iodine to iodide ion, and flush the resulting solution down the drain with more than 100-fold excess of water. This reaction produces heat, so handle the resulting solution carefully.

Explosive Mixtures of Ethyne and Air

Brief description of demonstration

Ethyne gas (acetylene) is mixed with different proportions of air and ignited. One mixture produces a very faint flame, one produces a more visible flame and copious amounts of soot, and one produces a blue flame and a very loud whistling noise.

Materials

- Pneumatic trough or washbasin
- Calcium carbide, CaC_2
- Watch glass
- Forceps
- 3 test tubes, 15 mm × 150 mm
- 3 rubber stoppers to fit test tubes
- Matches

Procedure

Fill the pneumatic trough or tray with water. Fill one of the test tubes with water, stopper it with your thumb, and invert the test tube. Put the mouth of the test tube below the surface of the water in the pneumatic trough. Take your thumb away from the mouth of the tube, and stand it in the corner. Add a small chunk of calcium carbide (about the size of a pea) to the water. It will begin to fizz. Place the mouth of the test tube over the fizzing calcium carbide. The tube will begin to fill with ethyne gas. Fill the tube completely. Stopper the tube while it is still underwater. Minimize the amount of water stoppered into the tube. Set the tube aside. Take the second test tube, and fill it halfway with water, leaving the other half filled with air. Stopper this tube with your thumb, and invert it into the trough as before. Collect ethyne gas into this tube as before, and stopper it as before. Set this tube aside. Repeat this procedure for the third and the fourth tubes, this time only filling the tubes $\frac{1}{4}$ of the way with water and $\frac{1}{8}$ to $\frac{1}{16}$ of the way with water, respectively. At the end of this procedure you should have 4 tubes : the first contains 100% ethyne, 0% air : the second, 50% ethyne, 50% air, the third, 25% ethyne, 75% air, and the fourth, 12.5% ethyne, 87.5% air.

For presentation, take the first tube containing 100% ethyne. Pointing the mouth of the tube away from spectators, light a match, remove the stopper, and quickly place the burning match over the mouth of the test tube. The gas will ignite at the surface of the ethyne/ air interface, but the rest of the gas will not burn. Put this fire out by placing it into the trough or basin: if it is allowed to fill with water slowly, the burning gas will escape from the tube as it fills with water.

Take the second tube, and shake it thoroughly. This will allow the air and ethyne to mix. Repeat the ignition

procedure. The mixture will ignite, producing a layer of soot on the surface of the test tube. Some of this soot often escapes into the air due to the turbulence produced by the burn and floats in the air. Repeat this procedure for the third tube. This burn is rapid, and sometimes produces a little soot, but more often does not. For the fourth tube, make sure the tube is shaken thoroughly, and ignite it. The ethyne ignites explosively with a blue flame and a loud whistle. If it does not, try collecting more ethyne, but add less water to the tube, thus more air.

Hazards

The ethyne produced is an explosive. Do not produce large amounts of it. A solution of calcium hydroxide is produced during the reaction. Do not let students handle it, and wash your hands after the demonstration.

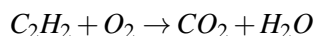
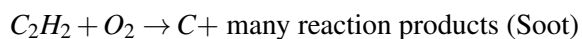
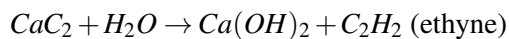
Disposal

Make sure all of the calcium carbide in the basin has reacted, then rinse the contents of the basin or trough down the sink with a 100 fold excess of water.

Discussion

The demonstration shows the effects of mixing different amount of oxygen with the ethyne produced. The production of soot in the second reaction clearly shows incomplete combustion, while the fourth reaction is clearly complete combustion, with no soot produced and water appearing in the test tube after the reaction. The garlic odors produced when reacting the calcium carbide results from sulfur compounds present during the synthesis of the carbide using coke, limestone, and an electric furnace.

The reactions producing ethyne and the incomplete and complete combustion are:



13.4 Worksheets for Chapter 13

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Balancing Equations Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Balance the following equations by inserting the smallest whole number coefficients.

1. $\text{___ CuCl} + \text{___ H}_2\text{S} \rightarrow \text{___ Cu}_2\text{S} + \text{___ HCl}$
2. $\text{___ Na} + \text{___ H}_2\text{O} \rightarrow \text{___ NaOH} + \text{___ H}_2$
3. $\text{___ Mg} + \text{___ O}_2 \rightarrow \text{___ MgO}$
4. $\text{___ Fe} + \text{___ O}_2 \rightarrow \text{___ Fe}_2\text{O}_3$
5. $\text{___ H}_2\text{O} + \text{___ N}_2\text{O}_3 \rightarrow \text{___ HNO}_2$
6. $\text{___ Fe} + \text{___ H}_2\text{O} \rightarrow \text{___ Fe}_3\text{O}_4 + \text{___ H}_2$
7. $\text{___ Al} + \text{___ Pb(NO}_3)_2 \rightarrow \text{___ Al(NO}_3)_3 + \text{___ Pb}$
8. $\text{___ KOH} + \text{___ H}_3\text{PO}_4 \rightarrow \text{___ K}_3\text{PO}_4 + \text{___ H}_2\text{O}$
9. $\text{___ C}_2\text{H}_6 + \text{___ O}_2 \rightarrow \text{___ CO}_2 + \text{___ H}_2\text{O}$
10. $\text{___ C}_2\text{H}_5\text{OH} + \text{___ O}_2 \rightarrow \text{___ CO}_2 + \text{___ H}_2\text{O}$
11. $\text{___ N}_2 + \text{___ H}_2 \rightarrow \text{___ NH}_3$
12. $\text{___ Al(OH)}_3 + \text{___ H}_2\text{SO}_4 \rightarrow \text{___ Al}_2(\text{SO}_4)_3 + \text{___ H}_2\text{O}$
13. $\text{___ SbCl}_3 + \text{___ H}_2\text{S} \rightarrow \text{___ Al}_2\text{S}_3 + \text{___ HCl}$
14. $\text{___ C}_5\text{H}_{12} + \text{___ O}_2 \rightarrow \text{___ CO}_2 + \text{___ H}_2\text{O}$
15. $\text{___ NH}_4\text{Cl} + \text{___ Ca(OH)}_2 \rightarrow \text{___ CaCl}_2 + \text{___ NH}_3 + \text{___ H}_2\text{O}$

Convert the following word equations into formula equations and then balance them.

16. Iron + oxygen yields iron (III) oxide.
17. Antimony + chlorine yields antimony (III) chloride.
18. Sodium chlorate (NaClO_3) yields sodium chloride + oxygen.
19. Lead (II) nitrate + hydrogen sulfide yields lead (II) sulfide + nitric acid (HNO_3).
20. Aluminum + sulfuric acid (H_2SO_4) yields aluminum sulfate + hydrogen gas.

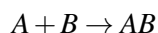
Types of Chemical Reactions Worksheet

There are millions of different compounds and therefore, there must be millions of different chemical reactions to form these compounds. When chemists are confronted with an overwhelming number of things, they tend to classify them into groups in order to make them easier to study and discuss. One popular system of classification for chemical reactions places them in five major categories. Some of the categories have different names in different books and you should become familiar with all the names.

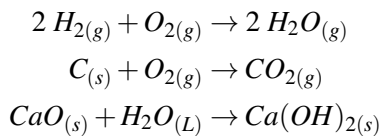
Types of Chemical Reactions

1. *Synthesis (also called Direct Combination)*

A synthesis reaction occurs when two or more substances combine to make a single, more complex substance. The reactants may be elements or compounds but the product will always be a compound. The general formula for this type of reaction can be shown as:



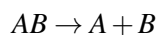
Some examples of synthesis reactions are shown below.



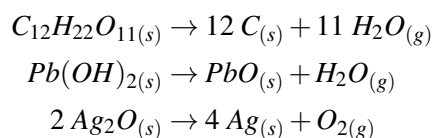
You should note in each case above, there are two or more substances in the reactants and only one substance as the product.

2. *Decomposition (also called Analysis)*

A decomposition reaction occurs when one substance is broken down into two or more simpler substances. This type of reaction is the opposite of a synthesis reaction, as shown by the general formula below:

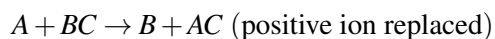


Some examples of decomposition reactions are shown below.

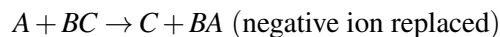


3. *Single Displacement (also called Single Replacement)*

In this type of reaction, a neutral element becomes an ion as it replaces another ion in a compound. The general form of this equation can be written as:

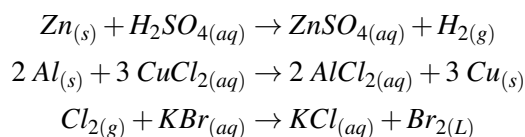


Or



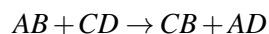
In either case, the equation is element + compound \rightarrow element + compound.

Some examples of single displacement reactions are shown below.



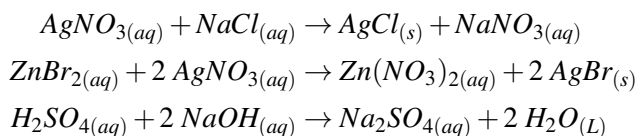
4. Double Displacement (also called Double Replacement and Metathesis)

In this reaction type, pairs of ionic compounds exchange partners. The basic form for this type of reaction is shown below.



The reaction is Compound + Compound \rightarrow Compound + Compound

Some examples of double displacement reactions are shown below.

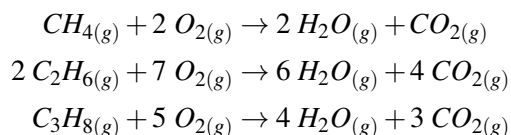


5. Combustion

When organic compounds are burned, they react with oxygen in the air to form carbon dioxide and water. The basic form of the combustion reaction is shown below.

hydrocarbon + oxygen \rightarrow carbon dioxide + water

Some examples of combustion reactions are shown below.



Exercises

Fill in the reaction type on the line following the balanced equation.

- $3 NaBr + H_3PO_4 \rightarrow Na_3PO_4 + 3 HBr$ _____
- $3 Ca(OH)_2 + Al_2(SO_4)_3 \rightarrow 3 CaSO_4 + 2 Al(OH)_3$ _____
- $3 Mg + Fe_2O_3 \rightarrow 2 Fe + 3 MgO$ _____

4. $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$ _____
5. $2PbSO_4 \rightarrow 2PbSO_3 + O_2$ _____
6. $2NH_3 + 3I_2 \rightarrow N_2I_6 + 3H_2$ _____
7. $H_2O + SO_3 \rightarrow H_2SO_4$ _____
8. $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$ _____
9. $4C_5H_9O + 27O_2 \rightarrow 20CO_2 + 18H_2O$ _____
10. $Li_3N + 3NH_4NO_3 \rightarrow 3LiNO_3 + (NH_4)_3N$ _____

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

13.5 Extra Reading for Chapter 13

Fireworks

If you enjoyed a Fourth of July evening pyrotechnic display, or perhaps witnessed a New Years' Eve event, you've witnessed the results of over a thousand years' worth of research and development into the art of fireworks. The first efforts were produced in China initially by accident as they observed that when saltpeter (potassium nitrate, KNO_3) was dropped into a charcoal fire, the mixture "popped" and produced an interesting flame color. Later, as a means to surprise their enemies in battle, the earliest "Shock and Awe" campaigns featured a mixture of saltpeter, charcoal and sulfur. The mechanism by which fireworks operate involves heating the proper ratio of these materials (75% KNO_3 , 15% carbon and 10% sulfur), and generating a chemical reaction to produce nitrogen and carbon dioxide gases. These initial "gung pow" were mainly explosive devices directed into the air, but later new additions brought whistling sound effects and a spectrum of colors to dazzle their opponents. The energy needed to propel the shell and to excite the composite atoms is still provided by a gunpowder formula.



The brilliant colors that produce the *oohs* and *aaahs* of today's displays are mainly due to elements like magnesium, which results in a blinding white effect. On an atomic level, the energy imparted by the explosion causes the atom's electrons to be promoted to a higher energy level. When the atoms relax back to the ground state, a specific amount of energy is released and the color of visible light reveals the frequency of light corresponding to that energy value. The red coloration is due to the presence of lithium or strontium salts such as lithium or strontium carbonate. Sodium salts (usually nitrate) generate a yellow hue and calcium chloride or sulfate result in orange coloration. Barium chloride supplies a green color. The all-American red, white and blue display is difficult to construct due to the complexity of finding a blue colored explosive. Usually copper chloride in a blue-violet hue is substituted. This copper salt's instability at the high temperatures of the exploding device has caused modern day pyrotechnical researchers to continue the search for a reliable source of blue color.

The shape of the image produced when the shell explodes in the air is a function of how the components are arranged in the shell. When the pyrotechnic device explodes as the resultant gases are produced, the arrangement of the salts in the mortar shell will mirror the pattern produced by the explosion. Dividing the materials into different compartments can also produce a "time – delay" effect, where the display effects occur sequentially.

The sound effects employed as a counterpoint to the visual display are also the result of chemical reactions. Adding bismuth trioxide to the mixture generates "popping" noises, whereas copper salicylate yields a "whistling" sound.

On the next occasion when a fireworks display rises to the sky, you'll not only enjoy the beautiful visual effects, but have an appreciation for the science that went into the presentation.

13.6 Assessment for Chapter 13

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER

14**Stoichiometry C-SE-TE****Chapter Outline**

- 14.1 THE MOLE CONCEPT AND EQUATIONS**
- 14.2 MASS-MASS CALCULATIONS**
- 14.3 LIMITING REACTANT**
- 14.4 PERCENT YIELD**
- 14.5 WORKSHEETS FOR CHAPTER 14**
- 14.6 EXTRA READING FOR CHAPTER 14**
- 14.7 ASSESSMENT FOR CHAPTER 14**

Lessons and Number of Activities for Lessons**TABLE 14.1: Lessons and Activities for Lessons**

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. The Mole Concept and Equations	0	0	0	0
2. Mass-Mass Calculations	1	0	1	0
3. Limiting Reactant	0	0	1	1
4. Percent Yield	0	0	1	0
5. Energy Calculations	0	2	0	0

14.1 The Mole Concept and Equations

Student Behavioral Objectives

The student will:

- express chemical equations in terms of molecules, formula units, and moles.
- determine mole ratios in chemical equations.
- explain the importance of balancing equations before determining mole ratios.
- use mole ratios in balanced chemical equations.

Timing, Standards, Activities

TABLE 14.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Mole Concept and Equations</i>	1.0	3c,3e

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for The Mole Concept and Equations (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

14.2 Mass-Mass Calculations

Student Behavioral Objectives

- The student will define stoichiometry.
- Given the mass of one reactant or product, the student will calculate the mass of any other reactant or product.
- The student will use the factor-label method in mass-mass calculations.

Timing, Standards, Activities

TABLE 14.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Mass-Mass Calculations</i>	2.0	3e

Activities for Lesson 2

Laboratory Activities

1. Mass-Mass Relationships in a Chemical Change

Demonstrations

1. None

Worksheets

1. Stoichiometry Worksheet

Extra Readings

1. None

Answers for Mass-Mass Calculations (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

14.3 Limiting Reactant

Student Behavioral Objectives

The student will:

- identify the limiting reactant in a chemical reaction.
 - identify excess reactants in chemical reactions.
 - calculate the limiting reactant using the mole-mole ratios.
 - calculate the products using the limiting reactant and the mass-mass ratios.
-

Timing, Standards, Activities

TABLE 14.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Limiting Reactant</i>	1.5	3e

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Limiting Reactant Worksheet

Extra Readings

1. Excess and Limiting Reactants
-

Answers for Limiting Reactant (L3) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

14.4 Percent Yield

Student Behavioral Objectives

The student will:

- define theoretical and actual yield.
- explain the difference between theoretical and actual yield.
- calculate percent yield (reaction efficiency).

Timing, Standards, Activities

TABLE 14.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Percent Yield</i>	1.5	3f

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Percent Yield Worksheet

Extra Readings

1. None

Answers for Percent Yield (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 14

This website provides a stoichiometry game.

- <http://www.chemcollective.org/mr/>

This website contains various resources, including PowerPoint lectures, on many topics in chemistry, including one on stoichiometry.

- <http://www.chalkbored.com/lessons/chemistry-11.htm>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 9-3 is on solving mass-mass problems.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson93.htm>

This website has several video lessons on performing stoichiometry calculations, including problems with a limiting reactant.

- <http://www.kentchemistry.com/moviesfiles/chemguy/advanced/ChemguyStoich.htm>

This video is an electronic blackboard presentation of the calculation of a percent yield.

- http://www.youtube.com/watch?v=1L12_TRSql8

Laboratory Activities for Chapter 14

Teacher's Pages for Mass-Mass Relationships in a Chemical Change

Lab Notes

The setup for this lab is easy. Prepare the 6 *M* *HCl* by diluting the stock 12 *M* acid with a $\frac{50}{50}$ (v:v) volume of water. Prepare enough so that each lab team has access to about 10 *mL*.

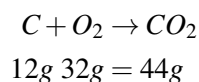
The most common error with this lab is to add too little or too much acid. The students will often leave unreacted *NaHCO*₃ on the bottom of the evaporating dish because it is difficult to see the difference in color between the dish and the *NaHCO*₃. Have students swirl the acid/*NaHCO*₃ mixture a little and look closely, the powder will appear as a sediment.

The 6 *M* acid often has a way of reminding us that chemical contamination can be dangerous. If a student reports itching, wash the affected area with copious amounts of water and send the student to the nurse if warranted.

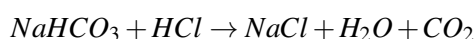
Lab – Mass-Mass Relationships in a Chemical Change

Background Information

When chemical reactions occur, the amount of product created can be theoretically predicted if you know the amounts of reactant used. This phenomenon is based on the law of mass-energy conservation. Since you cannot create or destroy matter in a chemical change, the sum of the masses of reactants and the sum of the masses of products must be equal. For example:



Since this equation is balanced, the coefficients in front of each reactant are 1. One mole of C must react with 1 mole of O_2 to produce 1 mole of CO_2 . The object of this experiment is to prove this in the laboratory. The reaction you will be performing is as follows:



As you can see, the coefficients in this reaction are also ones. One mole of $NaHCO_3$ reacted with one mole of HCl will produce one mole of $NaCl$, one mole of H_2O , and one mole of CO_2 . You will see that the number of moles of $NaHCO_3$ will be equal to the number of moles of $NaCl$ produced. The CO_2 is removed by letting it escape into atmosphere, and the H_2O is removed by evaporating it in a drying oven. The mass of salt can then be determined.

Purpose

The purpose of this activity is to study the relationships between the numbers of moles of reactant used, number of moles of products obtained, and coefficients used to balance a chemical reaction.

Apparatus and Materials

- Evaporating dish
- 10 mL graduate
- Watch glass
- Disposable pipette
- Electronic balance
- Goggles and apron
- Electric Oven
- Scoopula

Safety Issues

6 M HCl is VERY strong acid! It will cause a nasty chemical burn if you get it on your skin, and will put holes in your clothing. Wear goggles and apron, and if you get it on your skin, you will begin to itch immediately. Wash it off with plenty of water.

Procedure

1. Weigh a clean, dry evaporating dish + watch glass on the electronic balance. Record the mass.
2. Weigh out $2.0 \text{ g} \pm 0.10 \text{ g}$ of $NaHCO_3$ into the evaporating dish, and record the mass.
3. Measure 5.0 mL of 6 M HCl into a 10 mL graduate.
4. Place the watch glass on top of the evaporating dish, curve side down. Then add the HCl through the remaining opening by transferring it with a disposable pipette. Add all of it. Swirl the dish slowly. Record what you see. The reaction is complete when no solids remain. Add more acid drop wise if the reaction is not complete.
5. Place the dish, the liquid, and the watch glass into an 110°C oven for 24 hours.
6. Record the mass of the dish, the watch glass, and the remaining residue. Record this mass.

Data

Mass of empty dish + watch glass _____g

Mass of $NaHCO_3$ _____gMass of dish + watch glass + $NaCl$ _____gMoles of $NaHCO_3$ _____gMoles of $NaCl$ _____g**Post-Lab Questions**

1. From your balanced equation, what is the mole ratio between the $NaHCO_3$ and $NaCl$? How does it compare with the mole ratio from your experiment?
2. Suppose you had started with 20 moles of $NaHCO_3$, how many moles of $NaCl$ would you expect to be formed? Explain.
3. Do the results of your experiment support the law of conservation of mass-energy? Explain.

14.5 Worksheets for Chapter 14

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Stoichiometry Worksheet

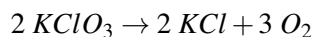
CK-12 Foundation Chemistry

Name _____ Date _____

1. How many moles are present in 58.6 *grams* of lead (II) oxide?

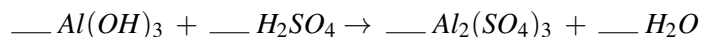
- A. 0.113 *moles*
- B. 0.158 *moles*
- C. 0.263 *moles*
- D. 0.300 *moles*
- E. None of these.

2. According to the following balanced equation, how many moles of oxygen can be produced by the complete reaction of 10.0 *moles* of potassium chlorate, $KClO_3$?



- A. 10.0 *moles*
- B. 6.67 *moles*
- C. 15.0 *moles*
- D. 4.00 *moles*
- E. None of these.

3. Balance the following equation and determine how many moles of water will be produced by the complete reaction of 0.600 *moles* of aluminum hydroxide?

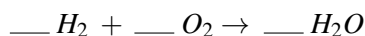


- A. 1.80 *moles*
- B. 0.200 *moles*
- C. 20.0 *moles*
- D. 0.600 *moles*
- E. None of these.

4. Using the balanced equation, $2 KClO_3 \rightarrow 2 KCl + 3 O_2$, how many moles of O_2 can be produced by the complete reaction of 100. *grams* of $KClO_3$?

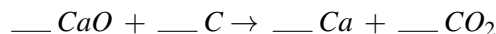
- A. 0.326 moles
- B. 0.544 moles
- C. 0.816 moles
- D. 1.22 moles
- E. None of these.

5. If hydrogen is completely reacted with oxygen and produces 180. grams of water, how many grams of hydrogen was consumed? The following equation for the reaction is not yet balanced.



- A. 2.02 g
- B. 20.2 g
- C. 10.1 g
- D. 90.0 g
- E. 180. g

6. How many grams of calcium can be produced by the complete reaction of 9.35 grams of calcium oxide, according to following, as yet unbalanced, equation?



- A. 6.70 g
- B. 3.34 g
- C. 12.4 g
- D. 7.19 g
- E. None of these.

7. In a particular reaction, iron (III) oxide and carbon solid reacted to produce iron metal and carbon monoxide. How many grams of iron (III) oxide are required to produce 150. grams of carbon monoxide?

- A. 160. g
- B. 222 g
- C. 286 g
- D. 480. g
- E. None of these.

8. How many grams of octane, C_8H_{18} , when burned in oxygen gas are required to produce 272 grams of carbon dioxide? The other product is water.

- A. 136 g
- B. 121 g
- C. 100. g
- D. 94.6 g
- E. 88.2 g

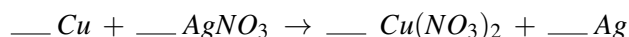
9. How many grams of bromine gas would be liberated when 25.0 *grams* of gallium bromide were heated and decomposed to form gallium metal and bromine gas?
- A. 16.4 *g*
B. 19.4 *g*
C. 21.8 *g*
D. 27.1 *g*
E. None of these.
10. 2000. *g* of potassium carbonate react completely with barium phosphate to produce potassium phosphate and barium carbonate. How many grams of barium carbonate will be formed?
- A. 1240 *g*
B. 1680 *g*
C. 2220 *g*
D. 2860 *g*
E. None of these.

Limiting Reactant Worksheet

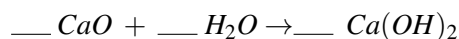
CK-12 Foundation Chemistry

Name _____ Date _____

1. If 2.5 *moles* of copper and 5.5 *moles* of silver nitrate are available to react in the following equation, what is the limiting reactant? (The equation is not yet balanced.)



- A. copper
B. silver nitrate
C. copper (II) nitrate
D. silver
E. None of these.
2. How many grams of calcium hydroxide will be formed in the following reaction when 4.44 *g* of calcium oxide and 7.77 *g* of water are available to react? (The equation is not yet balanced.)



- A. 12.2 *g*
B. 7.77 *g*
C. 5.86 *g*
D. 4.11 *g*

E. None of these.

3. Magnesium undergoes a single replacement reaction with nitric acid, HNO_3 . Write the balance equation for the reaction and determine how many grams of hydrogen gas will be formed from the reaction of 3.00 grams of magnesium with 18.00 grams of nitric acid.

A. 0.695 g

B. 0.572 g

C. 0.540 g

D. 0.492 g

E. None of these.

4. Sulfur reacts with oxygen gas to produce sulfur trioxide. Write the balanced equation for the reaction and determine how many grams of sulfur trioxide will be produced when 6.30 g of S and 10.0 g of O_2 are available for reaction.

A. 16.3 g

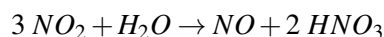
B. 15.7 g

C. 13.2 g

D. 11.9 g

E. None of these.

5. Some of the acid in acid rain is produced from the following reaction:



A falling raindrop with a mass of 0.0500 gram comes into contact with 0.200 gram of NO_2 . What mass of HNO_3 can be produced?

A. 0.183 g

B. 0.250 g

C. 0.350 g

D. 0.146 g

E. None of these.

6. In problem #5, how many grams of the excess reactant remains after the reaction?

A. 0.0415 g

B. 0.0388 g

C. 0.0264 g

D. 0.0239 g

E. None of these.

7. Consider the following reaction: $2 Al + 6 HBr \rightarrow 2 AlBr_3 + 3 H_2$. When 87.0 g of Al is combined with 401 g of HBr, how many grams of H_2 are formed?

A. 3.89 g

B. 5.01 g

C. 7.11 g

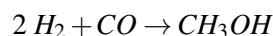
- D. 12.4 g
E. None of these.

Percent Yield Worksheet

CK-12 Foundation Chemistry

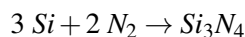
Name _____ Date _____

1. Methanol, CH_3OH can be produced by the following reaction.

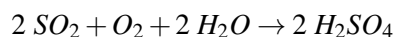


Assume CO is the limiting reactant and 2.00 *mols* of CO are used in the reaction. If 0.780 *mols* of CH_3OH are produced by the reaction, what is the percent yield?

2. Consider the following reaction.



- A. What is the theoretical yield of Si_3N_4 from this reaction when 21.45 *mols* of Si are reacted with excess N_2 ?
B. If 5.92 *mols* of Si_3N_4 are actually produced, what is the percent yield?
3. Part of the SO_2 that is introduced into the atmosphere by the combustion of sulfur containing compounds ends up being converted to sulfuric acid, H_2SO_4 by the following reaction.



- A. What is the theoretical yield of H_2SO_4 if 100. g of SO_2 is completely consumed?
B. If the actual yield from the reaction in A is 100. g of H_2SO_4 , what is the percent yield?
4. Consider the reaction: $4 FeS_2 + 11 O_2 \rightarrow 2 Fe_2O_3 + 8 SO_2$
A. If 20.0 *moles* of FeS_2 react with 60.0 *moles* of O_2 , what is the limiting reactant?
B. How many moles of SO_2 are formed?
C. How many moles of the reactant in excess will be left over at the end of the reaction?
D. If the actual yield of SO_2 is 25.0 *moles*, what is the percent yield?

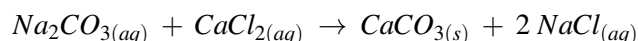
Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

14.6 Extra Reading for Chapter 14

Excess and Limiting Reactants

The chemical name for chalk is calcium carbonate. The reaction between sodium carbonate and calcium chloride produces calcium carbonate.



Stoichiometry allows us to compare the amounts of various species involved in a reaction. In order to determine which of the reactants is the limiting reactant, we must take into account both the amounts present and how they relate stoichiometrically in the balanced equation. Why do chemists use limiting reactants? The reason lies in the fact that not all reactions go to 100% completion; in fact the majority of the really interesting ones do not. However, scientists can use an equilibrium “trick” to get the stubborn reactions to go to completion. They start with an excess of one of the reactants to “push” the reaction to make more product. This essentially makes the other reactant the limiting reactant.

14.7 Assessment for Chapter 14

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER

15

The Behavior of Gases
C-SE-TE

Chapter Outline

- 15.1 THE THREE STATES OF MATTER
- 15.2 GASES
- 15.3 GASES AND PRESSURE
- 15.4 GAS LAWS
- 15.5 UNIVERSAL GAS LAW
- 15.6 MOLAR VOLUME
- 15.7 STOICHIOMETRY INVOLVING GASES
- 15.8 WORKSHEETS FOR CHAPTER 15
- 15.9 EXTRA READING FOR CHAPTER 15
- 15.10 ASSESSMENT FOR CHAPTER 15

Lessons and Number of Activities for Lessons

TABLE 15.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. The Three States of Matter	0	0	0	0
2. Gases	0	2	0	0
3. Gases and Pressure	0	1	1	1
4. Gas Laws	1	0	1	0
5. Universal Gas Law	1	0	1	0
6. Molar Volume	0	1	1	0
7. Stoichiometry Involving Gases	0	0	1	0

15.1 The Three States of Matter

Student Behavioral Objectives

- The students will describe molecular arrangement differences among solids, liquids, and gases.
- The students will describe the basic characteristic differences among solids, liquids, and gases.

Timing, Standards, Activities

TABLE 15.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Three States of Matter</i>	0.5	4a

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for The Three States of Matter (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

15.2 Gases

Student Behavioral Objectives

- The students will describe the relationship between molecular motion and Kelvin temperature.
- The students will describe random motion of gas molecules and explain how their collisions with surfaces cause pressure on the surface.
- The students will state that zero kinetic energy of molecules corresponds to 0 K and that there is no lower temperature.

Timing, Standards, Activities

TABLE 15.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Gases</i>	0.5	4a, 4f, 4g, 7a

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. Brownian Motion Demonstration
2. Molecular Motion - Kinetic Energy Demonstration

Worksheets

1. None

Extra Readings

1. None

Answers for Gases (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

15.3 Gases and Pressure

Student Behavioral Objectives

- The student will define pressure.
- The student will convert requested pressure units.
- The student will read barometers and both open-end and closed-end manometers.
- The student will apply the gas laws to relationships between the pressure, temperature, and volume of a gas.
- The student will state standard conditions for gases.

Timing, Standards, Activities

TABLE 15.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Gases and Pressure</i>	1.0	4a, 4f, 4g

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. Magdeburg Hemispheres Demonstration

Worksheets

1. Kinetic Molecular Theory Worksheet

Extra Readings

1. Elastic versus Inelastic Collisions

Answers for Gases and Pressure (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

15.4 Gas Laws

Student Behavioral Objectives

- The students will state Boyle's Law, Charles' Law, and Gay-Lussac's Law.
- The students will solve problems using Boyle's Law, Charles' Law, and Gay-Lussac's Law.
- The students will state the combined gas law.
- Using the combined gas law, and given any five of the six variables, the students will solve for the sixth variable.

Timing, Standards, Activities

TABLE 15.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Gas Laws</i>	2.0	4c, 4d

Activities for Lesson 4

Laboratory Activities

1. Charles' Law with a Balloon and Bunsen Burner

Demonstrations

1. None

Worksheets

1. Gas Laws Worksheet

Extra Readings

1. None

Answers for Gas Laws (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

15.5 Universal Gas Law

Student Behavioral Objectives

- The students will solve problems using the Universal Gas Law, $PV = nRT$.
- The students will state Avogadro's Law of equal molecules in equal volumes under the same conditions of temperature and pressure.
- The students will calculate molar mass from $mm = \frac{gRT}{PV}$, given mass, temperature, pressure, and volume.

Timing, Standards, Activities

TABLE 15.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Universal Gas Law</i>	2.0	4c, 4h

Activities for Lesson 5

Laboratory Activities

1. Finding the Molar Mass of a Gas Experimentally

Demonstrations

1. None

Worksheets

1. Universal Gas Law Worksheet

Extra Readings

1. None

Answers for Universal Gas Law (L5) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

15.6 Molar Volume

Student Behavioral Objectives

- The students will apply the relationship 1.00 *mole* of any gas at standard conditions will occupy 22.4 L.
- The students will convert gas volume at STP to moles and to molecules and vice versa.
- The students will apply Dalton's Law of Partial Pressures to describe the composition of a mixture of gases.

Timing, Standards, Activities

TABLE 15.7: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Molar Volume</i>	0.5	4b, 4d, 4i

Activities for Lesson 6

Laboratory Activities

1. None

Demonstrations

1. Rate of Diffusion at Various Temperatures Demonstration

Worksheets

1. Molar Volume and Partial Pressure Worksheet

Extra Readings

1. None

Answers for Molar Volume (L6) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

15.7 Stoichiometry Involving Gases

Student Behavioral Objectives

- The students will solve stoichiometry problems involving converting gas volume at STP to moles and vice versa.
- The students will solve stoichiometry problems involving gas volume to gas volume under any conditions of temperature and pressure.

Timing, Standards, Activities

TABLE 15.8: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Stoichiometry Involving Gases</i>	1.5	None

Activities for Lesson 7

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Stoichiometry Involving Gases Worksheet

Extra Readings

1. None

Answers for Stoichiometry Involving Gases (L7) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 15

This website provides a lesson on temperature and absolute zero.

- <http://www.colorado.edu/UCB/AcademicAffairs/ArtsSciences/physics/PhysicsInitiative/Physics2000/bec/temperature.html>

This website provides a particle motion computer simulation.

- <http://intro.chem.okstate.edu/1314F00/Laboratory/GLP.htm>

This website provides an animation of Boyle's Law.

- <http://www.grc.nasa.gov/WWW/K-12/airplane/aboyle.html>

The *learner.org* website allows users to view the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The video called "A Matter of State" examines matter in its three principle states - gases, liquids, and solids - and relates the visible world to the submicroscopic.

- <http://www.learner.org/resources/series61.html?pop=yes&pid=793#>

This website provides more details about the kinetic molecular theory.

- <http://www.chm.davidson.edu/vce/kineticmoleculartheory/basicconcepts.html>

These two videos contain a discussion of the relationship between absolute zero and kinetic energy.

- <http://www.youtube.com/watch?v=K4sOfGKEaxs>
- <http://www.youtube.com/watch?v=Mgyp94TZdqQ>

Flaws in videos:

- statement that temperature is related to average molecular velocity rather than average molecular kinetic energy.
- statement that the carbon dioxide molecule has an angular shape rather than linear shape.

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 7-1 is on the kinetic theory of gases.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson71.htm>

This video provides an introduction to gases and gas pressure.

- <http://video.google.com/videoplay?docid=8173099266343829057#>

This video is a laboratory demonstration of Boyle's Law.

- http://www.youtube.com/watch?v=J_I8Y-i4Axc

This video is a laboratory demonstration of Charles's Law.

- <http://www.youtube.com/watch?v=IkRIKGN3i0k>

This video is a black board presentation of some ideal gas law calculations and it includes the definition of standard temperature and pressure.

- http://www.youtube.com/watch?v=GwoX_BemwHs

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 7-6 is on the combined gas law.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson76.htm>

This video reviews how to perform ideal gas law (universal gas law) calculations.

- <http://video.google.com/videoplay?docid=-2010945058854921425#>

The video demonstrates diffusion in gases and liquids. It clearly shows the differences in rate of diffusion due to molar mass as well as higher temperatures.

- <http://www.youtube.com/watch?v=H7QsDs8ZRMI>

Laboratory Activities for Chapter 15

Teacher's Pages for Finding the Molar Mass of a Gas Experimentally Lab

Lab Notes:

The use of the butane lighters as directed in the lab produces few hazards. The use of the butane lighters in ways other than directed in the lab will produce many hazards. Alert supervision is required. The room should be well ventilated and no open flames allowed.

The most critical part of the lab, in terms of getting accurate values, is exactly filling the flask with butane to the 200. mL line while the line on the flask is held exactly at water level. Careful attention to detail is necessary.

Timing:

The calculations in this lab require use of the combined gas law, the universal gas law, and the law of partial pressures so the lab should be done after all topics have been considered in the course.

Answers to Pre-Lab Questions:

1. $P_{H_2} = P_{TOTAL} - P_{H_2O} = 755.0 \text{ mm of Hg} - 23.7 \text{ mm of Hg} = 731.3 \text{ mm of Hg}$
2. $V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{(730. \text{ mm of Hg})(137 \text{ mL})(273 \text{ K})}{(760. \text{ mm of Hg})(298 \text{ K})} = 121 \text{ mL}$
3. $\frac{0.51 \text{ g}}{0.250 \text{ L}} = \frac{x \text{ g}}{22.4 \text{ L}} \quad x = 45.7 \text{ g}$

Answers to Post-Lab Questions:

1. Since the pressure inside the flask is determined by measuring the atmospheric pressure in the room, it is necessary to make sure the pressures are the same. If the water levels inside and outside the flask are not exactly the same, the pressures will not be the same.

$$2. V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{(775 \text{ mm of Hg})(498 \text{ mL})(273 \text{ K})}{(760. \text{ mm of Hg})(373 \text{ K})} = 372 \text{ mL}$$

$$\frac{2.31 \text{ g}}{0.372 \text{ L}} = \frac{x \text{ g}}{22.4 \text{ L}} \quad x = 139 \text{ g}$$

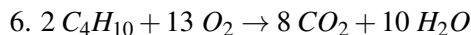
Therefore, the molar mass is 139 g/mol .

3. The molar mass of C_4H_{10} is 58.0 g/mol .

$$\text{moles } C_4H_{10} = \frac{3.60 \text{ g}}{58.0 \text{ g/mol}} = 0.0621 \text{ moles}$$

$$4. \text{ of molecules} = (\text{moles})(\text{molecules/mole}) = (0.0621)(6.02 \times 10^{23}) = 3.74 \times 10^{22} \text{ molecules}$$

$$5. V = \frac{nRT}{P} = \frac{(0.0621 \text{ mol})(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(273 \text{ K})}{(1.00 \text{ atm})} = 1.39 \text{ liters}$$



$$\frac{2 \text{ mols } C_4H_{10}}{13 \text{ mols } O_2} = \frac{0.0621 \text{ mol } C_4H_{10}}{x \text{ mol } O_2} \quad x = 0.404 \text{ mol } O_2$$

$$\text{grams } O_2 = (0.404 \text{ mol})(32.0 \text{ g/mol}) = 12.9 \text{ grams}$$

Finding the Molar Mass of a Gas Experimentally Lab

Background:

A mole of any pure gas at STP has a volume of 22.4 liters . The mass of that 22.4 L is the molar mass of the gas.

If the volume, mass, temperature and pressure of a gas is known, then, using the combined gas law, the volume can be mathematically converted to a volume at STP. The mass of 22.4 liters can be determined. This mass is the molar mass of the gas and has been determined experimentally.

When gas is collected by water displacement, you must look up the vapor pressure of water at that temperature and use Dalton's law of partial pressures to factor out the pressure of the water vapor.

Pre-Lab Questions

1. $200. \text{ mL}$ of hydrogen gas was collected over water at a pressure of 755.0 mm of Hg and a temperature of 25°C . If the vapor pressure of water at 25°C is 23.7 mm of Hg , what was the partial pressure of hydrogen in the container?
2. If a volume of gas occupies 137 mL at $730. \text{ mm of Hg}$ and 25°C , what volume will it occupy under standard conditions?
3. A $250. \text{ mL}$ sample of gas under standard conditions has a mass of 0.51 grams . What would be the mass of 22.4 L of this gas?

Purpose:

To experimentally determine the molecular mass of butane.

Safety Issues:

There should be no open flames in the room during the lab and the room should be well ventilated. Students should be closely supervised to make sure they are using the lighters only as directed.

Apparatus and Materials:

- butane lighter
- balance
- $250 - 300 \text{ ml}$ flask
- china marker or water proof marker
- 100 mL graduated cylinder
- Water trough (bucket, dishpan)
- thermometer
- plastic wrap or glass plate

Procedure:

1. Use a graduated cylinder to place exactly 200 ml of water in the flask. Using the marker, draw a highly visible line at the 200 ml water line.
2. Determine and record the exact mass of the butane lighter.
3. Fill the trough or other large container with water.
4. Fill the flask completely with water and using plastic wrap, a glass plate or your hand, invert the flask into the trough without permitting any air bubbles in it.
5. Make sure that the lighter is turned to its highest gas flow. Hold the butane lighter under the mouth of the flask and press the release lever. Be careful that all of the gas flows into the flask. Hold the flask so that the 200 ml mark is exactly even with the water level in the trough. Fill the flask exactly to the 200 ml mark that you made on the flask.
6. Determine and record the temperature of the water in the trough and the barometric pressure in the lab.
7. Thoroughly dry the butane lighter and determine and record its exact mass.

Data

1. Initial mass of lighter = _____ g
2. Final mass of lighter = _____ g
3. Mass of butane collected = _____ g
4. Volume of butane collected = 200 mL
5. Temperature = _____ °C
6. Atmospheric pressure in the lab = _____ mm of Hg
7. Vapor pressure of water at this temperature = _____ mm of Hg

TABLE 15.9: Table of water vapor pressures at normal lab temperatures.

Temperature, °C	Pressure, mm of Hg
17	14.5
18	15.5
19	16.5
20	17.5
21	18.6
22	19.8
23	21.0
24	22.4
25	23.7
26	25.2
27	26.7

Calculations:

1. Using Dalton's law of partial pressures, enter the pressure of water vapor that corresponds to the lab temperature in the data table. Then subtract this from the lab pressure to find the partial pressure of butane gas.
2. Convert the lab temperature from Celsius to Kelvin.
3. Using the combined gas law formula, calculate the volume of butane to the volume it would occupy at STP.
4. Knowing the mass of butane that occupied the volume found in calculation 3, use a proportion to determine the mass of this gas that would occupy 22.4 liters. This is the experimentally determined molar mass of butane.
5. Butane has the formula C_4H_{10} . Determine its molar mass from this formula.
6. Calculate your percent error.

$$\% \text{ error} = \frac{\text{actual value} - \text{experimental value}}{\text{actual value}} \times 100 =$$

Post-Lab Questions

1. In procedure step 5, it was required that the flask be held in the water trough so that the 200. mL line marked on the flask was exactly level with the water outside the flask. Why was this done?
2. If 2.31 grams of dry gas (no water vapor) occupied 498 mL at 100.°C and 775 mm of Hg, what is the molar mass of the gas?
3. A disposable butane lighter contains approximately 3.60 grams of butane. How many moles of butane is this?
4. How many molecules of butane would be present in 3.60 grams?
5. What volume would 3.60 grams of butane occupy under standard conditions?
6. How many grams of oxygen would be required to burn 3.60 grams of butane?

Demonstrations for Chapter 15

Charles Law with a Balloon and a Bunsen Burner

Brief description of demonstration

A Bunsen burner is lit and adjusted to a cool flame. Students then hold a partially inflated balloon over the flame. The balloon inflates rapidly, which they feel.

Materials

- Bunsen burner
- Balloons, several
- Matches

Procedure

Light the Bunsen burner, and adjust to a cool flame. Inflate the balloons by blowing into them until they reach a diameter of 20 to 25 cm. Tie the end off. Have the students hold the inflated balloon about 30 cm over the flame. The balloon will rapidly inflate.

Hazards

The Bunsen burner can cause burns or ignite clothing if the student gets too close. Demonstrate how to hold the balloon before allowing the students to do it. The balloon inflates so rapidly and is such a strange tactile feeling that the students often think it's alive, which can be startling to some. Be aware of over-reaction, and clear away any breakable items that may come into harms' way.

Disposal

Throw the used balloons away.

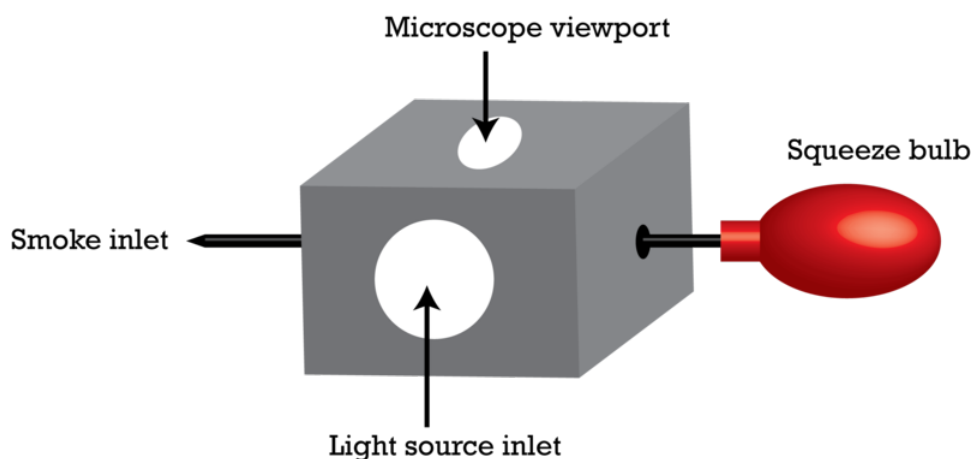
Discussion

For a balloon that is 20 cm in diameter, the circumference of the balloon will change about 7 mm for every applied 10°C temperature change. This is a relatively small change, but it does happen almost instantaneously, which surprises students.

Brownian Motion Demonstration

Brief description of demonstration

The Brownian Motion Apparatus consists of a metal chamber with a glass viewing window on top and a lens on one side to allow a light source to shine in the chamber. Smoke from a smoldering rope or an extinguished match is drawn into the chamber through an inlet tube by squeezing the rubber bulb. The chamber is illuminated by light shining through the side lens (laser light is best, if you have a laser pointer). The smoke cell sits on a microscope stage and by focusing through the top viewing window, smoke particles are visible (like shiny stars) against a dark background. The smoke particles jiggle about as they are bombarded by air molecules. There is also some sideways drift due to convection currents. If you have a projecting microscope, the image can be projected on a screen so everyone can see at the same time. Sometimes the smoke particles seem to explode as they rise or sink out of focus. The smoke activity in the cell will die down after a while so a new puff of smoke must be drawn in now and then.



Materials

- Microscope or projecting microscope
- Source of smoke (smoldering rope or an extinguished match)
- Brownian Motion Apparatus

Procedure

1. Place the Brownian Motion Apparatus on the microscope stage.
2. Adjust the side light into the lens.
3. Squeeze the rubber bulb and hold it empty.
4. Light and blow out a match. Hold smoking match near the inlet aperture and release the squeeze bulb.
5. Have the students take turn viewing unless you have a projecting microscope or a TV camera.

Hazards

None.

Disposal

Dispose of cooled matches in waste basket.

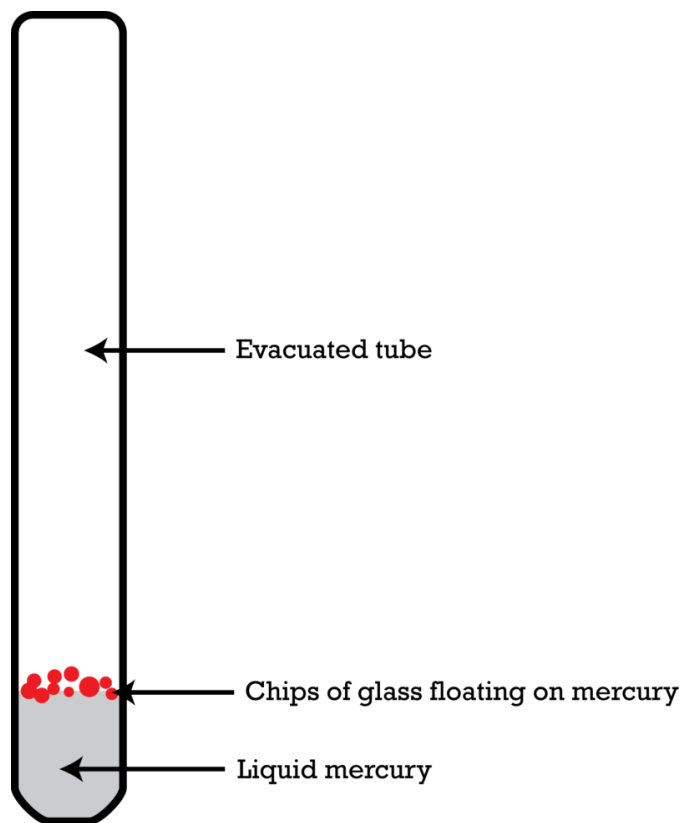
Discussion

Make sure the students understand that the smoke particles are much too large to exhibit molecular motion. The movement of the smoke particles is due to collisions from air molecules that do have molecular motion.

Molecular Motion/Kinetic Energy Demo

Brief description of demonstration

The apparatus is a glass tube with a small amount of mercury in the bottom and some glass chips floating on the mercury. (The apparatus is sometimes referred to as Stoekle's Molecular Apparatus) At room temperature, the molecular motion of the mercury molecules is not forceful enough to move the glass chips.



When the tube is held over a Bunsen burner, however, the molecular motion of the mercury molecules increases, and the collisions of the mercury molecules with the bottom of the glass chips sends them flying to the top of the tube.

Materials

- Molecular motion demonstration tube
- Utility clamp
- Bunsen burner
- Pot holder (something to lay the hot tube on while cooling)
- Matches

Procedure

Clamp the demonstration tube in a utility clamp so you can hold it without burning your fingers. Show the tube to the students at room temperature. Hold it over a Bunsen burner flame until the glass chips begin flying to the top of the tube. Allow the tube to cool before returning to storage.

Hazards

The tube gets very hot so be careful not to burn fingers. Mercury vapors are hazardous so it is important to be very careful not to break the tube. If you see any cracks in the tube, do not use it. It should be discarded according to disposal rules for mercury. It's best to store the tube from year to year in bubble wrap.

Discussion

As the mercury molecules leave the liquid due to heating, they collide with the glass chips and knock them high into the tube. When the tube is removed from the heat, tiny droplets of mercury can be seen to condense on the walls of the tube.

Rate of Diffusion at Various Temperatures Demo

Brief description of demonstration

Drops of food color are added to water at three different temperatures and the time required for complete diffusion is observed.

Materials

- Hot Plate
- Ice
- 3 – 250 mL graduated cylinders
- 3 – dropper pipets
- Thermometer
- Wall Clock with second hand

Procedure

1. Fill one graduated cylinder with tap water at room temperature, (this water needs to sit in the room for several hours to reach room temperature. (Fill a clean milk carton with tap water the day before the demonstration.
2. Heat tap water to boiling and fill the second graduated cylinder.
3. Make ice water and stir until no more ice melts. Fill the third graduated cylinder with ice water (but no ice cubes).
4. Record the temperature of the water in each cylinder.
5. Allow the water to settle for a couple of minutes.
6. Add two drops of good coloring to cylinder at the same time. Note the time.
7. Record how long it takes for the food coloring to be completely dispersed in each cylinder.

Hazards

Boiling water can be a burn hazard.

Disposal

All solutions can be disposed of down the sink.

Discussion

In which cylinder did the dye spread out the fastest.

In which cylinder did the dye spread out the slowest.

Give reasons for the different times of diffusion of the three cylinders of water.

The same molecules at different temperatures are moving at different speeds. At higher temperatures, the speed of the molecules is greater. Molecular collisions occur with greater frequency and with greater force.

If you wish, you could relate what the student sees in this demo to increased gas pressure at higher temperature and increased effusion rates (escape through a pinhole) at higher temperature.

Magdeburg Hemispheres Demonstration

Brief description of demonstration

Two five inch diameter hemispheres are placed together and the air is removed from inside the resulting sphere with a vacuum pump. A strong student is asked to pull the two hemispheres apart (unsuccessfully).

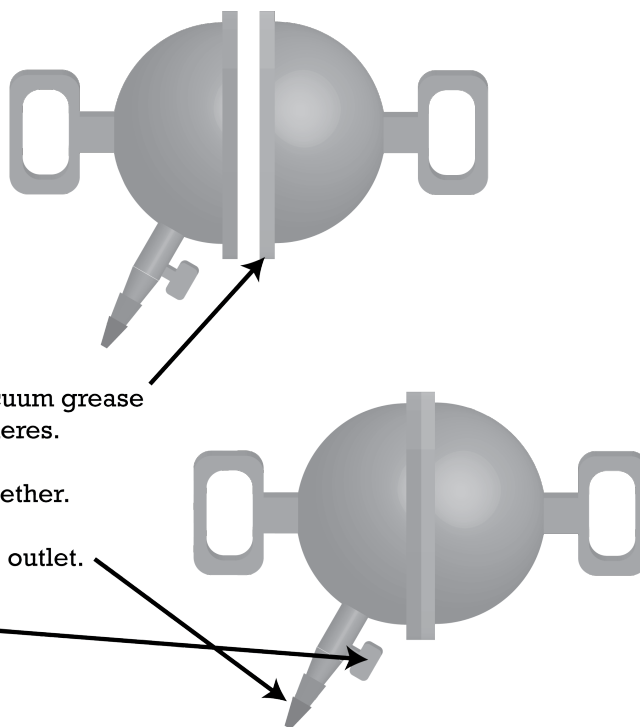
Purpose

To demonstrate the powerful force of air pressure.

Materials

- Magdeburg Hemispheres
- Vacuum grease
- Vacuum pump (a hand pump will do)

Procedure



1. Put a thin coating of vacuum grease on the lips of the hemispheres.

2. Fit the hemispheres together.

3. Attach vacuum pump to outlet.

4. Open valve.

5. Activate vacuum pump.

6. After a minute or two, close the valve and turn off the vacuum pump.

7. Have a strong student try to separate the hemispheres.

When the valve is opened and air allowed in, the hemispheres come apart easily.

If you have a bell jar large enough to enclose the hemispheres, you can put the hemispheres inside the bell jar and remove the air from the bell jar (surrounding the hemispheres) and the hemispheres will fall apart.

Hazards

An air leak may allow the hemispheres to come apart suddenly.

Discussion

The actual Magdeburg hemispheres were around 20 *inches* in diameter and were designed to demonstrate a vacuum pump that von Guericke had invented. When the air was removed from inside the hemispheres, and the valve closed, the hemispheres were held together by the air pressure of the surrounding atmosphere.

It is not known how good a vacuum von Guericke's pump was able to produce. If all the air were removed from the inside, the hemispheres would have been held together with a force of around 20,000 Newtons, equivalent to lifting a car or a small elephant.

Von Guericke's demonstration was performed on May 8, 1654. Thirty horses, in two teams of 15, could not separate the hemispheres until the vacuum was released.

15.8 Worksheets for Chapter 15

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Kinetic Molecular Theory Worksheet

CK12 Foundation Chemistry

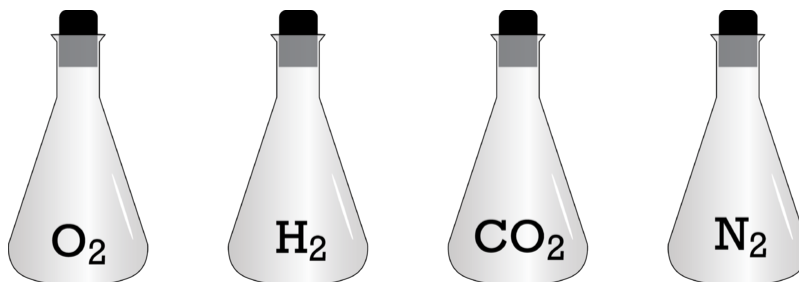
Name _____ Date _____

True or False

- _____ 1. The mass of a gas is the sum of the masses of the individual molecules.
- _____ 2. The volume of a gas is the sum of the volumes of the individual molecules.
- _____ 3. Molecules of different substances move at different velocities when they are the same temperature.
- _____ 4. Molecules of the same substance move at the same velocity when they are at the same temperature.
- _____ 5. Molecules are in motion at all temperatures above absolute zero.
- _____ 6. Gases are more compressible than solids and liquids because they have more space between the molecules.
- _____ 7. Molecules of liquid water and molecules of solid water (ice) at the same temperature have the same velocity.
- _____ 8. A liquid has its own shape and volume regardless of the container.
- _____ 9. All molecules at the same temperature have the same velocity.
- _____ 10. All molecules at the temperature has the same average kinetic energy.
- _____ 11. Molecules of different substances (different mass), at the same temperature, exert different amounts of force when they collide with the walls of their container.
- _____ 12. Smaller mass molecules move with greater velocity than larger mass molecules at the same temperature.

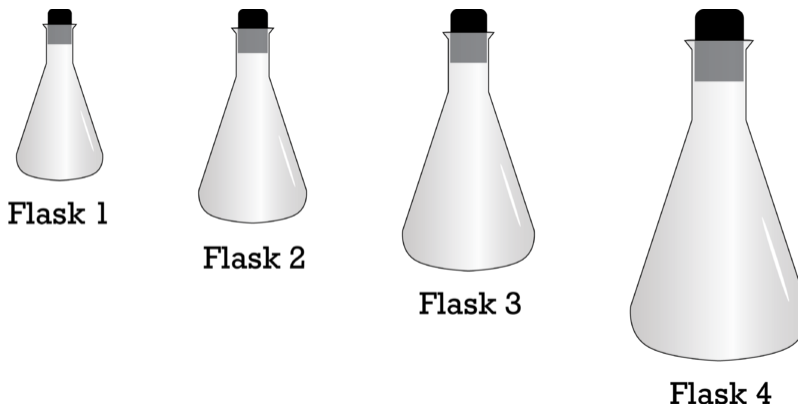
Multiple Choice

13. Four different flasks of the same size contain four different gases, all at the same temperature. In which of the following gas samples will the molecules be moving the fastest?



- A. O_2 B. H_2 C. CO_2 D. N_2 E. They will all have the same velocity.

14. For the same situation as in problem 13, in which of the gas samples will the molecules be moving the slowest?
A. O_2 B. H_2 C. CO_2 D. N_2 E. They will all have the same velocity.
15. Each of the flasks shown below contain the same number of molecules at the same temperature.



In which flask will the pressure be the highest?

- A. Flask 1 B. Flask 2 C. Flask 3 D. Flask 4 E. The pressure will be the same in all the flasks.

Combined Gas Laws Worksheet

CK12 Foundation Chemistry

Name _____ Date _____

- An ideal gas is contained in a volume V at temperature T . If the volume is doubled at constant pressure, what change will occur in the temperature?
A. the temperature will be unchanged B. the temperature will be halved C. the temperature will be doubled D. unable to determine without more information
- An ideal gas is sealed in a container at constant volume. If the temperature is increased to $4T$, what would happen to the pressure?
A. unchanged B. increased to $4P$ C. decreased to $\frac{1}{4}P$ D. unable to determine without more information
- If the temperature of an ideal gas is increased from 300 K to 1200 K , what change occurs in the kinetic energy of the gas?
A. it remains constant B. it increases by a factor of 2 C. it increases by a factor of 4 D. it increases by a factor of 8 E. it increases by a factor of 16
- A sample of gas is held at constant volume. When the temperature of the gas is $100.\text{ K}$, the pressure is 1.00 atm . What must the temperature become in order for the pressure to become 3.00 atm ?
A. 27 K B. $100.\text{ K}$ C. $300.\text{ K}$ D. None of these. E. cannot be determined from this data
- A sample of gas occupies $100.\text{ mL}$ at 1520 Torr and 323 K . What volume will this sample occupy under standard conditions?
A. $100.\text{ mL}$ B. 116 mL C. 232 mL D. 169 mL E. None of these.
- 10.0 liters of oxygen gas is held at $3800.\text{ mm of Hg}$ pressure and 27.0°C . What volume will this gas occupy if it is at -23.0°C and $380.\text{ mm of Hg}$ pressure?
A. 8.33 L B. 83.3 L C. 833 L D. 50.0 L E. None of these.

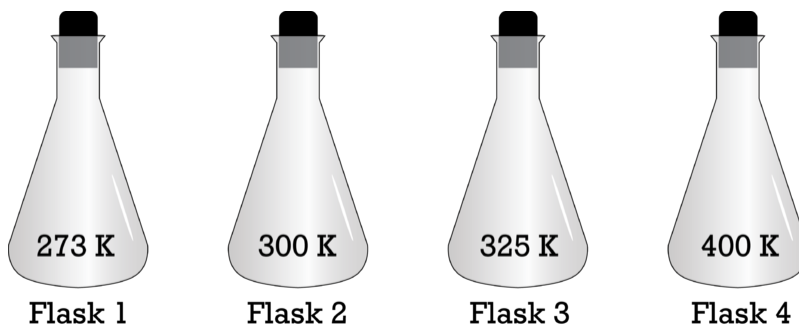
7. A gas sample occupies 3.25 liters at 24.5°C and 1825 mm of Hg . Determine the temperature at which the gas will occupy 4.25 liters at 1140 mm of Hg .
8. 2.00 L of hydrogen, originally at 25.0°C and $750.\text{ torr}$ are heated until a volume of 20.0 L and a pressure of 2660 torr is reached. What is the new temperature?
9. 73.0 mL of nitrogen gas at STP is heated to 80.0°C and the volume increases to 4.53 L . What is the new pressure?
10. The pressure of a gas is reduced from 1.50 atm to 1.00 atm as the volume of the container is increased from 85.0 mL to $350.\text{ mL}$. If the original temperature was 363 K , what is the final temperature?

Universal Gas Law Worksheet

CK12 Foundation Chemistry

Name _____ Date _____

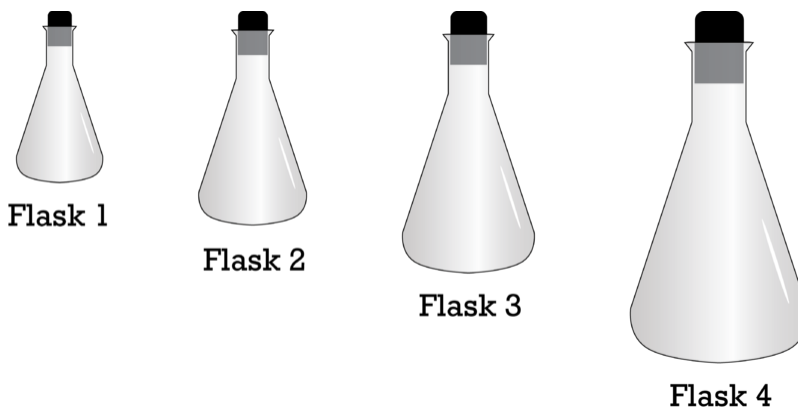
1. 10.0 atm of pressure is applied to 0.250 mol of methane gas. What must the temperature be if the volume is to be $1400.\text{ mL}$?
A. 409 K B. 682 K C. 955 K D. 0 K E. None of these.
2. Given a sample of gas at 1.0 atm pressure, what would the pressure become if the amount of gas is doubled, the volume decreased to half, and the absolute temperature quadrupled?
A. 1.0 atm B. 2.0 atm C. 4.0 atm D. 8.0 atm E. 16 atm
3. How many moles of gas are required to fill a 1.0 liter container to 5.00 atm at 27.0°C ?
A. 0.13 mols B. 0.20 mols C. 0.29 mols D. 0.38 mols E. None of these.
4. What is the mass of $100.\text{ L}$ of Br_2 gas under standard conditions?
A. 22.4 g B. 357 g C. $560.\text{ g}$ D. 714 g E. insufficient data to determine
5. What is the molar mass of a gas if 0.500 g of it occupies 0.250 L at 1.00 atm and $100.^{\circ}\text{C}$?
A. 32.0 g/mol B. 44.0 g/mol C. 61.2 g/mol D. 77.2 g/mol E. 104 g/mol
6. 10.0 liters of a gas at 27.0°C and 0.15 atm has a mass of 10.0 g . What is the molar mass of this gas?
A. $40.\text{ g/mol}$ B. $80.\text{ g/mol}$ C. $100.\text{ g/mol}$ D. $120.\text{ g/mol}$ E. 164 g/mol
7. A cylinder with 2.0 moles of an ideal gas is held at constant volume and constant temperature. If 2.0 more moles of gas are added to the cylinder, what must happen to the temperature?
A. the temperature will be unchanged B. the temperature must be doubled C. the temperature must halved D. insufficient data to determine
- 8.



Each of the flasks shown above contains the same number of gas molecules. In which flask would the pressure be lowest?

A. Flask 1 B. Flask 2 C. Flask 3 D. Flask 4 E. All four flasks would have the same pressure.

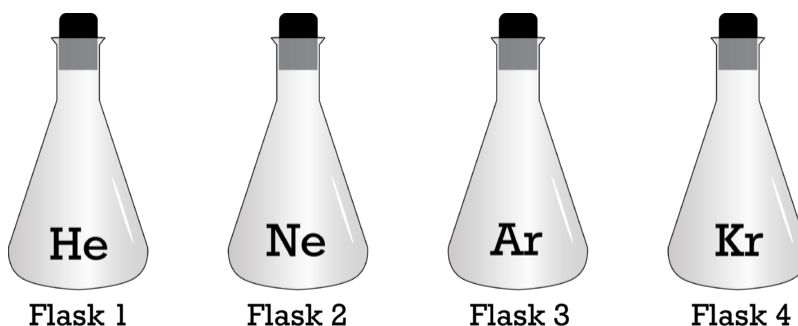
9.



Each of the flasks shown above contains the same number of gas molecules at the same temperature. In which flask would the pressure be lowest?

A. Flask 1 B. Flask 2 C. Flask 3 D. Flask 4 E. All four flasks would have the same pressure.

10.



If all the flasks above are the same size, at the same temperature, and contain the same number of molecules, in which flask will the molecules be moving slowest?

A. Flask 1 B. Flask 2 C. Flask 3 D. Flask 4 E. The molecular velocity would be the same in all flasks.

Molar Volume and Partial Pressure Worksheet

CK12 Foundation Chemistry

Name _____ Date _____

Molar Volume

- Calculate the number of moles contained in 6.50 L of sulfur dioxide at STP.
- A cylinder of O_2 gas contains a volume of 60.50 L at STP. How many moles of oxygen gas are in the cylinder?
- A cylinder contains 15.85 moles of N_2 gas at STP. What is the volume of the cylinder in liters?

- 12.35 *mol*s of Cl_2 are compressed in a cylinder at 5.00 *atm*. If the gas were released from the cylinder at STP, what volume would the gas occupy?
- What volume would a 200. *gram* sample of hydrogen sulfide gas occupy at STP?

Partial Pressure

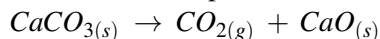
- A flask contains only 0.10 *mol* of nitrogen gas and 0.40 *mol* of oxygen gas. The total pressure in the flask is 1.00 *atm*. What is the partial pressure of the nitrogen gas?
- 1.00 *g* of H_2 gas is placed in a flask with 1.00 *g* of He gas. The total pressure in the flask is 900. *torr*. What is the partial pressure of the H_2 gas?
- A mixture of neon and argon gases exert a total pressure of 2.39 *atm*. The partial pressure of the neon alone is 1.84 *atm*. What is the partial pressure of the argon?
- A 500. *mL* sample of hydrogen gas is collected over water at 27°C. The pressure of the hydrogen and water vapor mixture is 78.5 *kPa* and the vapor pressure of water at 27°C is 3.6 *kPa*. What is the partial pressure of the dry hydrogen gas?
- A flask contains 320. *g* of O_2 , 320. *g* of N_2 , and 88.0 *g* of CO_2 . The total pressure in the flask is 500. *kPa*. What is the partial pressure of each gas in the flask?

Stoichiometry Involving Gases Worksheet

CK12 Foundation Chemistry

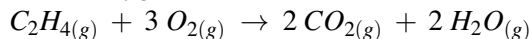
Name _____ Date _____

- Calcium carbonate decomposes at 1.00 *atm* pressure and 0°C to form carbon dioxide and calcium oxide.



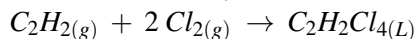
How many grams of calcium carbonate is needed to produce 3.50 liters of carbon dioxide?

- Ethene burns in oxygen at 1.00 *atm* and 500. *K* to form carbon dioxide and water.



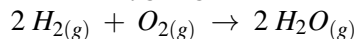
How many liters of water vapor can be formed if 1.25 *L* of ethane are consumed?

- When chlorine is added ethyne, 1,1,2,2-tetrachloroethane is formed.

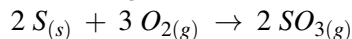


If this reaction occurs at 2.00 *atm* pressure and 100.°C, how many liters of chlorine are needed to produce 75.0 *g* of $C_2H_2Cl_4$?

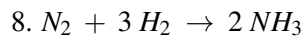
- What volume of oxygen gas would react with 35.0 *mL* of hydrogen gas at STP, according to the equation below?



- What volume of oxygen is needed to react with solid sulfur to form 2.5 *L* of sulfur trioxide, according to the equation below, if all gases are measured under the same conditions of temperature and pressure.



- How many liters of propane gas, C_3H_8 , will undergo complete combustion with 34.0 *L* of oxygen gas, forming carbon dioxide and water, if both gases are measured under the same conditions of temperature and pressure?
- When 5.00 *L* of propane gas is completely combusted to form carbon dioxide and water vapor at a temperature of 600. *K* and a pressure of 1.20 *atm*, what mass of water vapor will be formed?



Ammonia is formed by the synthesis of hydrogen and nitrogen as shown in the equation above. If all gases are measured under the same conditions, how many liters of ammonia can be formed from 15.0 L of hydrogen gas?

Answers to Worksheets

- **The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.**

15.9 Extra Reading for Chapter 15

Elastic versus Inelastic Collisions

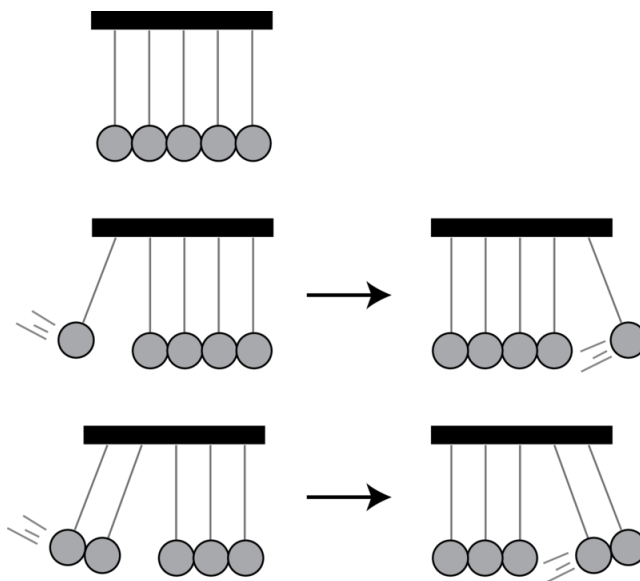
The momentum, (p), of an object is defined as the mass of the object multiplied by its velocity, mv . The velocity of an object and the momentum of the object are vectors. That is a statement of either the velocity or the momentum of an object includes the direction that the object is traveling. The direction is an integral part of the measurement. If we assign the direction north to be positive direction, then a 5.0 kg object traveling north at 7.0 meters/second will have a momentum of $+35\text{ kg}\cdot\text{m/s}$. In this same system, a 5.0 kg object traveling south at 7.0 m/s will have a momentum of $-35\text{ kg}\cdot\text{m/s}$. During collisions between objects, momentum is always conserved. In this system (consisting of these two objects), the total momentum of the system is $0\text{ kg}\cdot\text{m/s}$ because $(+35\text{ kg}\cdot\text{m/s}) + (-35\text{ kg}\cdot\text{m/s}) = 0$. If these two objects collide and bounce directly backwards with velocity exactly opposite to their original velocities, the object that had a momentum of $+35\text{ kg}\cdot\text{m/s}$ will now have a momentum of $-35\text{ kg}\cdot\text{m/s}$ and the object whose original momentum was $-35\text{ kg}\cdot\text{m/s}$ will now have a momentum of $+35\text{ kg}\cdot\text{m/s}$. The total momentum of the system is still $0\text{ kg}\cdot\text{m/s}$ and momentum has been conserved (as it always is). If these two objects collide and stick together (like two balls of Play Doh), both velocities become zero. In such a case, the momentum of each object is zero, the total momentum of the system is zero, and once again, momentum is conserved.

The kinetic energy, KE , of an object is defined as one-half the mass of an object multiplied by its velocity squared, $KE = \frac{1}{2}mv^2$. The kinetic energy of an object is NOT a vector. An 5.0 kg object traveling north at 7.0 m/s will have a $KE = \frac{1}{2}(5.0\text{ kg})(7.0\text{ m/s})^2 = 120\text{ kg}\cdot\text{m}^2/\text{s}^2 = 120\text{ Joules}$. In this same system, a 5.0 kg object traveling south at 7.0 m/s will also have a kinetic energy of 120 Joules . . . there is no direction associated with kinetic energy. Therefore, the total kinetic energy in this system is 240 Joules . . . the opposite directions of the ball's motions do not cause cancellation when dealing with kinetic energy. If these two objects collide and bounce directly backwards with velocities exactly opposite to their original velocities, each object will have the same kinetic energy it had before the collision and the total kinetic energy of the system will still be 240 Joules . . . kinetic energy has been conserved. If these two objects collide and stick together (like two balls of Play Doh), both velocities become zero. In such a case, the kinetic energy of each object is zero, the total kinetic energy of the system is zero, and kinetic energy is not conserved. Since energy, all forms considered, is conserved in all interactions except nuclear, the kinetic energy that was lost in the collision must be found in some other form, usually heat and sound. Make sure you understand that energy is conserved in non-nuclear interactions, but the form of the energy is not necessarily conserved. Specifically, KE is not conserved in the collision but energy in all forms is conserved. Mechanical energy may become electrical or electrical energy may become light, but when all forms of energy are added up, energy is conserved.

Considering collisions of all sorts, collisions between automobiles, collisions between tennis balls and walls, collisions between billiard balls, momentum is always conserved and kinetic energy is almost never conserved. When automobiles collide, metal parts are bent, causing parts to rub against each other, and friction turns kinetic energy into heat and sound. Even a tennis ball bouncing on the ground slowly loses energy of motion as it bounces. When the tennis ball strikes the ground, it is deformed and this deformation stores energy in the ball and as the ball regains its shape, the ball bounces back up in the air and the stored energy again becomes energy of motion. But, in the process, the deformation of the ball causes internal friction which is converted to heat and the ball will not bounce as high after each bounce. The tennis ball will bounce lower and lower until all the energy has been converted to heat. You probably cannot detect the temperature increase in a tennis ball but the same thing occurs when a hammer pounds on a nail and if you touch the nail after several strikes, you will feel the higher temperature.

Only a few collisions in nature come close to conserving kinetic energy. The collisions between billiard balls or between polished steel balls come quite close to conserving kinetic energy. A popular demonstration of conservation

of momentum and conservation of kinetic energy features several polished steel balls hung in a straight line in contact with each other.



If one ball is pulled back and allowed to fall and strike the line of balls, exactly one ball will fly out the other side. The other balls, including the one which was dropped will remain motionless.

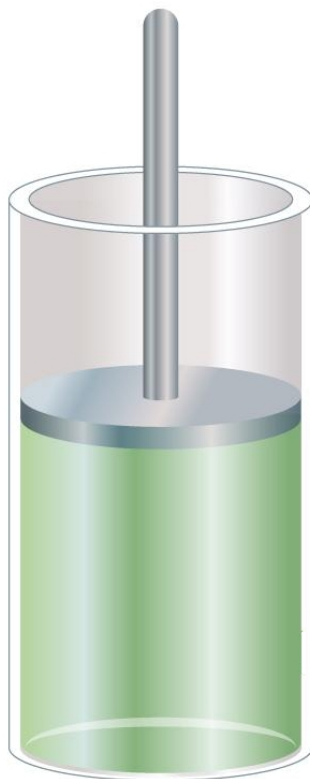
If two balls are pulled back and allowed to fall and strike the line of balls, exactly two balls will fly out the other side. The other balls, including the two that were dropped will remain motionless.

In the extreme case, if four balls are pulled back and allowed to fall, to strike the single motionless ball, four balls will fly out the other side, leaving one motionless ball.

The reason this strange phenomena occurs is that both momentum and kinetic energy are conserved in these collisions. Momentum would be conserved if one ball dropped at velocity X and two balls flew out the other side with velocity $\frac{1}{2}X$ but this would not conserve kinetic energy. In order for kinetic energy to be conserved, the same number of balls must fly out with the same velocity as the balls that were dropped.

When kinetic energy is conserved in a collision, physicists refer to the collision as a perfectly elastic collision. Why do we offer all this physics information to a chemistry teacher? The answer is that collisions between ideal gas particles are perfectly elastic collisions, that is, kinetic energy is conserved in collisions between gas particles. That's why when gas particles are bouncing around inside a container and exerting pressure, they do not gradually lose kinetic energy resulting in a lower pressure (as they would do if they were tennis balls).

There are cases, however, when gas particles do gain or lose kinetic energy without heat being added or removed from an external source. Consider a gas trapped in a closed cylinder fitted with a piston.



Consider the situation when the piston is held in position by hand. If the pressure inside the cylinder is greater than the external pressure, then releasing the piston will allow the gas inside the cylinder to push the piston higher. Moving the piston higher requires energy. Since the piston is being pushed higher by the gas inside the cylinder, the energy must come from the gas. The molecules of gas that strike the piston and push it upward are doing work ($\text{force} \times \text{distance}$) and will lose some kinetic energy. Therefore, those molecules slow down. The average kinetic energy of the molecules in the cylinder becomes less and therefore, the temperature will be lower (temperature is proportional to the average kinetic energy of the molecules). Thus, the expansion of the gas against a force (outside pressure, gravity, etc.) causes the gas to cool slightly. Conversely, if you push the piston down, thus compressing the gas, your hand is doing work on the molecules the piston strikes. Those molecules will gain kinetic energy and so the average KE of the molecules increases. The temperature of the gas will increase slightly. Thus, the compression of the gas raises its temperature slightly.

Suppose you have two boxes, one containing a gas and one containing a vacuum, and you open a valve between the boxes so the gas can expand into the vacuum. In this case, the gas is not pushing against anything, so it is not doing work and there will be no temperature change.

15.10 Assessment for Chapter 15

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 16 Condensed Phases: Solids and Liquids C-SE-TE

Chapter Outline

- 16.1 PROPERTIES OF SOLIDS AND LIQUIDS
- 16.2 INTERMOLECULAR FORCES OF ATTRACTION
- 16.3 IONIC, METALLIC, AND NETWORK CONDENSED PHASES
- 16.4 VAPOR PRESSURE AND BOILING
- 16.5 HEAT AND CHANGES OF STATE
- 16.6 PHASE DIAGRAMS
- 16.7 WORKSHEETS FOR CHAPTER 16
- 16.8 EXTRA READING FOR CHAPTER 16
- 16.9 ASSESSMENT FOR CHAPTER 16

Lessons and Number of Activities for Lessons

TABLE 16.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Properties of Solids and Liquids	0	0	0	0
2. Intermolecular Forces of Attraction	0	0	0	0
3. Ionic, Metallic, and Network Condensed Phases	0	0	0	1
4. Vapor Pressure and Boiling	0	1	0	0
5. Heat and Changes of State	1	0	3	1
6. Phase Changes	0	0	1	0

16.1 Properties of Solids and Liquids

Student Behavioral Objectives

- The student will explain the basic behavior and characteristics of solids and liquids using the molecular arrangement present in liquids.
- The students will describe the molecular arrangement in solids.
- The students will use the molecular arrangement in solids to explain the incompressibility of solids.
- The students will use the molecular arrangement in solids to explain the low rate of diffusion in solids.
- The students will use the molecular arrangement in solids to explain the ability of solids to maintain their shape and volume.

Timing, Standards, Activities

TABLE 16.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Properties of Solids and Liquids</i>	1.0	2d

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Properties of Solids and Liquids (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

16.2 Intermolecular Forces of Attraction

Student Behavioral Objectives

- The student will identify liquids whose intermolecular forces of attraction are due to London dispersion forces, polar attractions, and hydrogen bonding.
- The student will describe some of the unique properties of water that are due to hydrogen bonding.
- The student will select from comparative compounds, the ones most likely to form hydrogen bonding.
- The student will select from comparative compounds whose intermolecular forces are London dispersion forces, the one most likely to have the strongest intermolecular forces.
- The students will describe the metallic bond, and explain some of the solid characteristics that are due to metallic bonding.
- Given characteristics of a solid such as conductivity of solid and liquid phase, solubility in water, malleability, and so on, the students will identify the type of solid, i.e. the attractive forces holding the solid in solid form.

Timing, Standards, Activities

TABLE 16.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Intermolecular Forces of Attraction</i>	1.0	2d, 2h

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Intermolecular Forces of Attraction (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org

to request sample answers.

16.3 Ionic, Metallic, and Network Condensed Phases

Student Behavioral Objectives

- The students will describe the metallic bond, and explain some of the solid characteristics that are due to metallic bonding.
- Given characteristics of a solid such as conductivity of solid and liquid phase, solubility in water, malleability, and so on, the students will identify the type of solid, i.e. the attractive forces holding the solid in solid form.

Timing, Standards, Activities

TABLE 16.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Ionic, Metallic, and Network Condensed Phases</i>	1.0	2a, 7c

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. Liquid Crystals

Answers for Ionic, Metallic, and Network Condensed Phases (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

16.4 Vapor Pressure and Boiling

Student Behavioral Objectives

- The students will describe the processes of evaporation and condensation.
- The students state the factors that control the rates of evaporation and condensation.
- The students will describe vapor pressure equilibrium.
- The students will state the relationship between boiling point, vapor pressure, and ambient pressure.
- Given a vapor pressure table for water, and the ambient pressure, the students will determine the boiling point of water for specified conditions.

Timing, Standards, Activities

TABLE 16.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Vapor Pressure and Boiling</i>	1.0	7c

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. Boiling Water in a Paper Cup

Worksheets

1. None

Extra Readings

1. None

Answers for Vapor Pressure and Boiling (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

16.5 Heat and Changes of State

Student Behavioral Objectives

- The student will calculate energy changes during phase changes.
- The student will explain the slopes of various parts of heating and cooling curves.
- The students will explain why it is necessary for a solid to absorb heat during melting even though no temperature change is occurring.
- Given appropriate thermodynamic data, the students will calculate the heat required to raise temperatures of a given substance with no phase change.
- Given appropriate thermodynamic data, the students will calculate the heat required to melt specific samples of solids with no temperature change.
- Given appropriate thermodynamic data, the students will calculate the heat required to produce both a phase change and a temperature change, for a given sample of solid.

Timing, Standards, Activities

TABLE 16.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Heat and Changes of State</i>	2.0	7c, 7d

Activities for Lesson 5

Laboratory Activities

1. Race to $110^{\circ} C$.

Demonstrations

1. None

Worksheets

1. Intermolecular Forces of Attraction Worksheet
2. Heat Transfer Worksheet
3. Calorimetry Worksheet.

Extra Readings

1. Refrigeration

Answers for Heat and Changes of State (L5) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

16.6 Phase Diagrams

Student Behavioral Objectives

- The students will read specific requested information from a phase diagram.
 - The students will state the primary difference between a generic phase diagram, and a phase diagram for water.
-

Timing, Standards, Activities

TABLE 16.7: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Phase Diagrams</i>	1.0	7c, 7d

Activities for Lesson 6

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Phase Diagram Worksheet

Extra Readings

1. None
-

Answers for Phase Diagrams (L6) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.
-

Multimedia Resources for Chapter 16

This website provides a fill-in-the-blank worksheet generator.

- <http://www.theteacherscorner.net/printable-worksheets/make-your-own/fill-in-the-blank/>

This website provides a lesson on Characteristics of Solids, Liquids, and Gases.

- <http://www.chem.purdue.edu/gchelp/liquids/character.html>

This website provides a Phase Diagram Learning Activity.

- http://www.wisc-online.com/objects/index_tj.asp?objID=GCH6304

This website provides a brief animation showing the differences in molecular motion and relative position for gases, liquids, and solids.

- <http://www.youtube.com/watch?v=s-KvoVzukHo>

This website provides a video demonstration of a physical model of molecular motion and states of matter.

- <http://www.youtube.com/watch?v=ynUso6rJ0rE>

An interactive animated video showing the motion and arrangement of molecules in the three states of matter.

- http://www.harcourtschool.com/activity/states_of_matter/

This video shows demonstrations that demonstrate the states of matter.

- <http://www.youtube.com/watch?v=j2KZmRIKea8&feature=related>

This video shows and narrates how hydrogen bonds form and shows variations in boiling point of a homologous series due to hydrogen bonding.

- <http://www.youtube.com/watch?v=LGwyBeuVjhU>

This website provides an animation showing the form of a metallic bond.

- <http://www.youtube.com/watch?v=ijw8OBt4btM>

The following web site has data and explanatory reasons for the trends in melting and boiling points of some period 3 elements.

- <http://www.creative-chemistry.org.uk/alevel/module1/trends8.htm>

The *learner.org* website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The website has a video that apply to this lesson called “Metals” that details the value of accuracy and precision. In the video, malleability, ductility, and conductivity are examined, along with the methods for extracting metals from ores and blending alloys.

- http://learner.org/vod/vod_window.html?pid=811

The following sites provide more information about crystals, glasses, and amorphous materials.

- <http://dwb4.unl.edu/Chem/CHEM869A/CHEM869ALinks/www.ualberta.ca/bderksen/florin.html>
- <http://www.sciencedaily.com/releases/2008/07/080704153507.htm>
- <http://www.bestcrystals.com/crystals2.html>

The following website provides more information about vapor pressure.

- <http://www.chem.purdue.edu/gchelp/liquids/vpress.html>

This video is a ChemStudy film called “Gas Pressure and Molecular Collisions.” The film is somewhat dated but the information is accurate.

- <http://www.youtube.com/watch?v=fK6LfN6CD0w>

This video serves a blackboard lecture on the mathematics of heat involved in temperature change and phase change.

- http://www.youtube.com/watch?v=zz4KbvF_X-0

This video explores how matter changes state depending on the temperature.

- <http://www.videopediaworld.com/video/29170/Exploring-Heat-Change-of-State>

For more information about heating curves, visit the website:

- <http://www.kentchemistry.com/links/matter/heatingcurve.htm>

To look at more phase diagrams, visit the website:

- <http://www.kentchemistry.com/links/Matter/Phasediagram.htm>

Laboratory Activities for Chapter 16

Teacher’s Pages for Race to 110 Degrees Celsius

Lab Notes

As we know, the students will not be able to reach 110°C in this experiment, because the lab is carried out at atmospheric pressure. Nevertheless, they will try, and be stubborn and creative about it. Do not allow them to tamper with other lab equipment, or make makeshift apparatus to allow them to take the pressure within their container above 1 *atm*. Show the students how to mount the thermometer into the clamp correctly, how to lower it into the water, and to what depth to lower it. Some thermometers are total immersion: avoid these if possible. They will give inaccurate readings. The alcohol or hexane within the thermometers will be colored red or blue-green: this is NOT mercury. Most mercury thermometers have been eliminated from high schools because of the perceived toxicity of metallic mercury and of potential liability. If you do come across a mercury thermometer, do not allow students to use it, and give it to your department head for disposal. If you are the department head, report it to your district science supervisor and arrange for disposal.

The temperature readings yielded by student-grade thermometers are notoriously inaccurate. The readings are often off by $\pm 5^{\circ}\text{C}$. Let the students discover this on their own, and in post-lab discussion (or during the lab, for that matter) address this issue with them.

If you had your class do the thermometer calibration lab earlier in the year, they can refer back to what they learned about lab-grade thermometers in that laboratory activity.

Answers to Pre-Lab Questions

1. melting
2. boiling* (or vaporization)
3. sublimation
4. deposition
5. condensation
6. freezing

- Boiling is a common term, but somewhat inaccurate. The boiling point is defined as the temperature at which the vapor pressure of a liquid is equal to the surrounding or ambient pressure. In most cases, this is at 1 *atm*. It is often worthwhile to use the term “normal boiling point” when referring to the boiling point at a pressure of 1 *atm*.

Race to 110 Degrees Celsius Lab

Background Information

The physical properties of water, as all pure substances, has definite and predictable characteristics. Of these, melting and boiling points are most often measured and the ones that most people are familiar with. Melting occurs when a solid changes to a liquid, and boiling occurs when a liquid changes to a gas. The substance itself does not change in chemical composition. Ice and liquid water and steam all have the same chemical formula: H_2O . However, their physical properties are vastly different. The temperatures at which these changes take place are called their melting points and boiling points, respectively. The changes themselves are called phase changes, with the solid, liquid, and gas forms called phases.

Pre-Lab Questions

Give the name of the following changes of phase.

1. solid to liquid
2. liquid to gas
3. solid to gas
4. gas to solid
5. gas to liquid
6. liquid to solid

Purpose

The purpose of the lab is to investigate some of the physical behavior of water during temperature and phase changes, to learn how to use a thermometer correctly, and to construct a heating curve.

Apparatus and Materials

- Beaker, 250 *mL* or 400 *mL*
- Water, H_2O : liquid and solid
- Hotplate
- Ring stand and Thermometer clamp
- Thermometer, -20 to $110^\circ C$
- Beaker Tongs

Safety Issues

Boiling water and the hotplates can cause severe burns if touched, so do not touch them. Use normal and proper laboratory techniques to prevent injury.

Procedure

1. Fill a 250 mL or 400 mL beaker about halfway with liquid water. Add a few ice cubes to it, so the beaker is about $\frac{3}{4}$ full.
2. Take the temperature of the ice water until it stays steady, which should be around 0°C. Record this as your starting temperature.
3. Place the beaker on the hot plate, and adjust the thermometer on the clamp so the bulb is immersed (short ones) or to the line (long ones). Record the temperature again.
4. Turn your hot plate on. Record the temperature every minute until the temperature reaches 110°C. The first group to do it wins!

TABLE 16.8: Data

Time, in Minutes	Temperature, in °C	Time, in Minutes	Temperature, in °C
Start		13	
1		14	
2		15	
3		16	
4		17	
5		18	
6		19	
7		20	
8		21	
9		22	
10		23	
11		24	
12		25	

Graph: Construct a graph of the temperature vs. time. Use either a sheet of graph paper or Excel or some other graphing program.

Post-Lab Questions

1. Why did the graph level out eventually? Do you think it is possible to get a higher temperature than you got? If so, how?
2. Did your thermometer record the boiling point and the melting points of water accurately? If it did not, could you suggest a reason why it did not? How could you correct this error?

Demonstrations for Chapter 16

Boiling Water in a Paper Cup

Brief description of demonstration

Water is placed in a non – waxed paper cup. A Bunsen burner is placed underneath the cup with a medium flame. The water in the cup will boil without the paper igniting.

Materials

- Non – waxed paper cup (the conical ones used with water coolers work very well)
- Bunsen burner
- Ring Stand and Ring
- Water

Procedure(s)

Fill the cup to within 1cm of the top with tap water. Place the cup so that it is held firmly onto the iron ring: the closer the ring is to the top of the cup the better. Small diameter ring stands (about 6 to 7 *cm*) work well for this. Light the Bunsen burner and adjust the flame so that it is relatively cool (with the air vents shut). Position the tip of the flame so that it directly underneath and just touching the lower tip of the cup. Depending on the intensity of the flame the water will start boiling within 2 – 3 *minutes*, without the cup igniting.

Hazards

Boiling hot water can cause severe burns. The cup may ignite if the flame touches the portion of the cup not in direct contact with the water. The cup may scorch slightly.

Disposal

Let the water cool, and dispose of the water down the sink, and the paper in the trash can.

Discussion

The heat capacity of water is enormous in comparison to other materials $4.18 \text{ J/g}\cdot^{\circ}\text{C}$ vs. less than $1 \text{ J/g}\cdot^{\circ}\text{C}$ for other materials. Thus it can absorb a lot of heat energy before changing its temperature. Once it reaches its boiling point of 100°C , it will stay there until all of the liquid water is gone, and the paper in contact with the water will stay at that temperature too. The temperature at which paper catches fire is 233°C , which is far above the temperature of the water, so the paper cannot catch on fire. The instant the water is no longer in contact with the paper, however, the temperature of the paper will rise very quickly, so watch it carefully so the paper does not ignite and spill its hot contents.

16.7 Worksheets for Chapter 16

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Intermolecular Forces of Attraction Worksheet

Intermolecular forces of attraction are forces that exist between molecules. For example, two hydrogen atoms and one oxygen atom can be held together in the form of a water molecule by covalent chemical bonds. Many water molecules can then be held together to form a glass of water or a block of ice by intermolecular forces of attraction.

We can divide the six intermolecular forces of attraction into two groups of three. One group of three are referred to as strong intermolecular forces, namely network intermolecular forces (covalent bonds), ionic intermolecular forces (ionic attractions) and metallic intermolecular forces (metallic bonds). The other group, referred to as weak intermolecular forces consist of London Dispersion Forces, Dipole-Dipole forces, and hydrogen bonding.

The phase in which a substance exists (solid, liquid, gas) is a result of the competition between intermolecular forces of attraction pulling the molecules together and molecular motion (temperature) pushing them apart. If the intermolecular forces are much greater than molecular motion, the molecules will be pulled together tightly to form a solid. If the molecular motion is much greater than the intermolecular forces of attraction, the molecules will fly away from each other to form a gas. If the intermolecular forces of attraction are approximately equal to the molecular motion, the molecules will be pulled together but not tightly enough to produce a solid . . . only tightly enough to produce a loosely held group of molecules in the liquid phase.

As a substance is heated in the solid state, the molecular motion increases and the substance can pass into the liquid phase and eventually the gaseous phase as it continues to be heated.

As the strength of the intermolecular forces of attraction increase, so do the boiling points, melting points, heats of melting, and heats of vaporization. As the strength of the intermolecular forces of attraction increase, the vapor pressure of the substance decreases.

Network (Covalent) Bonds are the strongest of the intermolecular forces of attraction.

- In a network solid, atoms are covalently bonded into a continuous piece of solid material.
- Network solids are very hard and their strong bonds result in very high melting and boiling points.
- Network solids have localized electrons which are fixed in positions in the covalent bonds. This makes network solids very poor conductors of electricity and heat. (An exception is the highly pi bonded graphite.)
- Examples of network solids are diamond, graphite, asbestos, mica, silicon (IV) oxide, silicon carbide, and many rocks such as granite.
- Melting points of network solids are usually above 2000 K.
- Network solids are not soluble in water.

Ionic Bonds (within ionic crystal lattices) are the second strongest of the intermolecular forces of attraction.

- In an ionic solid, adjacent ions have electrostatic attractions and are arranged in a lattice structure.
- Ionic solids have high melting and boiling points.
- Ionic solids have localized electrons which are in fixed positions around the atoms. This makes ionic solids poor conductors of heat of electricity. Ionic liquids, however, (such as molten *NaCl*) are good conductors of electricity because the ions are free to move about.

- The strength of ionic intermolecular forces of attraction are related to the charge on the ions and the size of the ions. (Coulombs Law says that the force of attraction is directly proportional to the product of the charges and inversely proportional to the distance between the centers of the ions.) Therefore, the ionic attractions in MgO (charges $2+$ and $2-$) are stronger than those in NaF (charges $+1$ and -1). The ionic attractions in LiF are stronger than those in KF because Li is much smaller than K .
- Melting point of ionic solids are usually 900 – 1500 K .
- Ionic solids are soluble in water.

Metallic Bonds (bonds between metal atoms) are the weakest of the strong intermolecular forces of attraction.

- Metallic solids are often described as “a group of nuclei surrounded by a sea of mobile electrons”. The electrons in metallic solids are delocalized and are free to move about, making metals good conductors of heat of electricity.
- Because there are many mobile electrons in the valence shells of metals, many electron jumps are available and thus metals can absorb and emit many frequencies of light by electron transition . . . this frequently causes metals to be white and shiny.
- The non-localized electrons and therefore non-directional bonds in metals allow the atoms to be moved within the solid without breaking the solid. This makes metals malleable and ductile.
- Examples of metallic solids are Fe , Mg , Na .
- Metallic solids usually melt 500 – 1000 K . (There are many exceptions.)
- Metals do not dissolve in water.

Hydrogen Bonding is the strongest of the weak intermolecular forces of attraction.

- In hydrogen bonds, the partial positive charge on the hydrogen end of one molecule is attracted to the partial negative charge on the end of another molecule . . . that end must be an extremely electronegative element (fluorine, oxygen, or nitrogen – FON)
- HF , H_2O , NH_3 , and alcohols (along with other organic molecules with $-OH$ groups can form hydrogen bonds.
- Hydrogen bonding explains why water in the solid phase is less dense than it is in the liquid phase . . . contrary to most other substances. The hydrogen bonds in solid water form a crystal structure keeping the molecules further apart than they are in the liquid phase.
- Hydrogen bonds are an extremely strong form of dipole attractions.
- Substances that form hydrogen bonds have uncharacteristically higher melting and boiling points that do similar substances that don't form hydrogen bonds.
- Hydrogen bonded substances are soluble in water.

Dipole-Dipole forces are the second strongest of the weak intermolecular forces of attraction.

- Dipole-dipole force exist between neutral, polar molecules where the positive end of one molecule is attracted to the negative end of another molecule.
- The greater the polarity (difference in electronegativity of the atoms in the molecule), the stronger the dipole attraction.
- Dipole-dipole attractions are very weak and substances held together by these forces have low melting and boiling points.
- Only the strongest dipole-dipole attractions are liquids at room temperature . . . most of them are gases at room temperature.

Dipole-dipole solids are not conductive.

- Dipole-dipole solids are soluble in water.

London dispersion forces are the weakest of the weak intermolecular forces of attraction.

- LDF forces occur between neutral, nonpolar molecules. LDF occur due to the random motion of electrons.
- At any moment, one atom may have an unbalanced organization of its electrons resulting in an instantaneous polarity on the atom. During that instant, the unbalanced atom will act as a very weak dipole.
- Since LDF are dependent on the random motion of electrons, the more electrons an atom or molecule has, the stronger the London dispersion forces.
- LDF are extremely weak so substances held together by these force have extremely low melting and boiling temperatures.
- LDF solids are almost all gases and room temperature and many of them are gases well below room temperature.
- All substances have LDF but LDF are so weak, they are only considered when there are no other intermolecular forces of attraction.
- LDF solids are not conductive.
- LDF solids are not very soluble in water.

Exercises

1. Given the molar mass and the dipole moment (a measure of polarity) for the following substances, which should have the highest boiling point?
A. propane, mm = 44 g/mol, dipole moment = 0.1 B. dimethylether, mm = 46 g/mol, dipole moment = 1.3 C. methyl chloride, mm = 50 g/mol, dipole moment = 1.9 D. acetaldehyde, mm = 44 g/mol, dipole moment = 2.7 E. acetonitrile, mm = 41, dipole moment = 3.9
2. Which of the following substances has London dispersion forces as its only intermolecular force of attraction?
A. CH_3OH B. HCl C. NH_3 D. H_2S E. CH_4
3. Which of the following should have the lowest boiling point?
A. PH_3 B. HCl C. H_2O D. H_2S E. SiF_4
4. Which of the following substances is likely to have the largest heat of vaporization?
A. I_2 B. Br_2 C. Cl_2 D. F_2 E. O_2
5. Which of the following is an exothermic process?
A. melting B. subliming C. freezing D. boiling E. All of the above are exothermic.
6. If you three substances, one in solid phase, one in liquid phase, and one in gaseous phase, and all substances are at the same temperature, the strongest interparticle attractions will exist in the
A. solid. B. liquid. C. gas. D. insufficient information
7. For most substances, as the solid melts, the molecules become
A. closer together. B. farther apart. C. no change in distance between molecules.
8. An exception to the general rule implied in question 7 is
A. sodium chloride. B. diamond. C. calcium metal. D. ice. E. There are no exceptions to the rule.
9. Which of the following will form a solid held together by dipole-dipole attractions?
A. AsH_3 B. BCl_3 C. Cl_2 D. CO_2 E. XeF_4
10. Which of the following compounds will have the highest boiling point?
A. CCl_4 B. CF_4 C. CBr_4 D. Cl_4 E. CH_4

11. Hydrogen bonding is a special case of
A. London dispersion force B. dipole-dipole attraction C. ionic bonding D. metallic bonding E. None of the above.
12. Large intermolecular forces in a substance are manifested by:
A. low vapor pressure. B. high boiling point. C. high melting point. D. high heat of vaporization. E. All of the above.
13. No temperature change occurs during melting and boiling because the heat added during these phase changes is used to
A. increase kinetic energy. B. re-arrange atoms within molecules. C. overcome intermolecular forces of attraction. D. increase molecular velocity. E. increase density.
14. The predominant intermolecular force in $CaBr_2$ is:
A. dipole-dipole forces. B. hydrogen bonding. C. London dispersion forces. D. ionic bonding. E. covalent bonding.
15. Crystalline solids differ from amorphous solids in that crystalline solids have
A. no orderly structure. B. a repeating pattern of atoms, molecules, or ions. C. much larger atoms, molecules, or ions. D. stronger intermolecular forces of attraction.
16. A solid that has a very high melting point, is a poor electrical conductor, and is not soluble in water is most likely a(n)
A. ionic solid. B. metallic solid. C. network solid. D. dipole-dipole solid. E. LDF solid.
17. Which one of the following statements is incorrect?
A. Dispersion (London) forces are the weakest type of intermolecular interactions. B. The strong intermolecular attractions in H_2O result from hydrogen bonding. C. The boiling point of H_2S is lower than H_2O . D. The boiling point of non-polar substances tends to decrease with increasing molecular weight.
18. Which of the species below does not exhibit hydrogen bonding?
A. C_2H_6 B. NH_3 C. HF D. H_2O E. C_2H_5OH
19. Based on intermolecular interactions, which of the following should have the highest boiling point?
A. CH_4 B. $CHCl_3$ C. H_2S D. CH_3OH
20. Which of the following will have the highest melting point?
A. C_2H_6 B. NH_3 C. HF D. Kr E. Fe

Use the heating curve shown below to answer questions 21, 22, and 23.



The heating curve was generated by heating a 1.00 g sample of a solid substance beginning at -30°C and ending at 142°C .

21. What is the boiling point of this substance according to this heating curve?
A. 0°C B. 100°C C. -30°C D. 142°C E. None of these.
22. The specific heat of the liquid form of this substance could be determined by finding the slope of which segment?
A. A B. B C. C D. D E. E
23. The heat of melting for this solid could be calculated by adding up all the delivered during which segment?
A. A B. B C. C D. D E. E
24. The shape of a liquid's meniscus in a container is determined by:
A. the shape of the container. B. the volume of the liquid. C. the relative magnitudes of cohesive forces in the liquid and adhesive forces between the liquid and the container. D. the viscosity of the liquid. E. the amount of hydrogen bonding in the liquid.
25. What type of forces are being overcome when liquid oxygen boils at 90 K?
A. ionic bonds B. dipole-dipole forces C. covalent bonds D. London dispersion forces E. hydrogen bonding
26. A white substance melts at 1000 K. As a solid, it is a non-conductor of electricity but it dissolves in water to form a conducting solution. The white substance is:
A. a network solid. B. an ionic solid. C. a metallic solid. D. a dipole-dipole solid. E. an LDF solid.
27. A white solid melts at 1000 K. As a solid, it is a good conductor of electricity and it does not dissolve in water. The white solid is most likely:
A. a network solid. B. an ionic solid. C. a metallic solid. D. a dipole-dipole solid. E. an LDF solid.
28. Which of the following statements about liquids is NOT true?
A. liquids assume the shape of their container B. liquids assume the volume of their container C. liquids are virtually incompressible D. diffusion of liquids is slow E. liquids flow readily
29. What happens to molecules in a liquid when the liquid is heated and begins forming vapor?
A. the intramolecular forces holding the molecules together are being disrupted B. the intermolecular forces between the molecules are being disrupted C. the kinetic energy of the molecules is decreasing D. LDF are becoming stronger E. surface tension is increasing

30. Which of the following statements best explains why diamond has a higher melting point than dry ice?
 A. carbon atoms in diamond are held together by a network of covalent bonds
 B. dry ice molecules are held together by dipole-dipole forces
 C. diamond has greater internal ionic charges than dry ice
 D. the LDF in diamond are greater than in dry ice.
 E. All of the above are true.

Heat Transfer Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Temperature is defined as the **average** kinetic energy of all the molecules in a body, while **heat** is defined as the **total** kinetic energy of all the molecules in a body. A sample of matter will contain kinetic energy due to the motion of its molecules, and it also contains potential energy due to its phase (solid, liquid, gas). When two objects come into contact with each other, heat always flows from the one with higher temperature to the one with lower temperature. This transfer of KE is accomplished by the collision of molecules and continues until the two objects are at the same temperature.

Every chemical change and many physical changes involve the gain or loss of energy. In most cases, this energy gain or loss occurs in the form of heat, but light and electricity are also possible. Heat gains and losses are measured in units called **Joules**. It requires 4.18 *Joules* of energy to raise the temperature of 1.00 *gram* of water by 1.00°C. When heat energy is added to a substance, it produces one or both of the following effects: 1. it may increase the temperature of the object, which means it increases the average kinetic energy of the molecules or, 2. it may cause a phase change in that substance, which means it increases the potential energy of the substance.

When heat is absorbed by a substance as kinetic energy, the temperature of the substance increases because temperature is a measure of the average kinetic energy of the molecules of the substance. Different substances have a different amount of increase in temperature when they absorb the same amount of energy. The quantity of heat 1.00 *gram* of the substance must absorb to raise its temperature by 1.00°C is called the **specific heat** of the substance. The symbol, *C*, is often used for specific heat. The specific heat of water is 4.18 $J/g \cdot ^\circ C$. This means that 1.00 *gram* of water requires 4.18 *J* of heat to raise its temperature by 1.00°C. The specific heats of most substances are considerably less than that of water.

TABLE 16.9: Specific Heat of Various Substances

Substance	Specific Heat
Aluminum, Al	0.900 $J/g \cdot ^\circ C$
Copper, Cu	0.386 $J/g \cdot ^\circ C$
Gold, Au	0.126 $J/g \cdot ^\circ C$
Silver, Ag	0.235 $J/g \cdot ^\circ C$
Ethanol, C_2H_5OH	2.40 $J/g \cdot ^\circ C$
Butane, C_4H_{10}	2.34 $J/g \cdot ^\circ C$
Water, H_2O	4.18 $J/g \cdot ^\circ C$

Energy is also absorbed or given off by substances when they undergo a phase change. The energy gained or lost during a phase change is potential energy. This energy gain or loss does not change the temperature of the substance. When substances undergo a phase change, the average distance between the molecules changes and this requires an input or output of potential energy. When a substance changes from solid to liquid, the energy that must be absorbed is called **heat of melting**. The reverse process, changing from liquid to solid, gives off exactly the same amount of energy but for this phase change, the amount of energy is known as the **heat of fusion**. The phase change from liquid to gas requires an input of the **heat of vaporization**. The reverse process, gas condensing to liquid, gives off the same amount of potential energy but it is called the **heat of condensation**. Like specific heat, each substance has

its own heat of melting and heat of vaporization.

TABLE 16.10: Thermodynamic Data of Various Substances

Substance	Specific Heat	Heat of Fusion, ΔH_{fusion}	Heat of Vaporization, ΔH_{vap}
Aluminum, <i>Al</i>	0.900 J/g·°C	400. J/g	10,900 J/g
Copper, <i>Cu</i>	0.386 J/g·°C	205 J/g	5,069 J/g
Gold, <i>Au</i>	0.126 J/g·°C	64.5 J/g	1,578 J/g
Silver, <i>Ag</i>	0.235 J/g·°C	111 J/g	2,320 J/g
Ethanol, C_2H_5OH	2.40 J/g·°C	109 J/g	841 J/g
Butane, C_4H_{10}	2.34 J/g·°C	80.1 J/g	385 J/g
Water, H_2O	4.18 J/g·°C	334 J/g	2,260 J/g

The energy absorbed or given off by a substance during a temperature change (with no phase change) can be calculated with the equation, $Q = mC\Delta t$, where Q is the amount of heat in Joules, m is the mass in grams, C is the specific heat, and Δt is the temperature change.

Example: How many Joules are given off when 52.5 g of water cools from 67.5°C to 23.2°C?

Solution: $Q = mC\Delta t = (52.5 \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(44.3^\circ\text{C}) = 9720 \text{ J}$

The specific heat is taken from the table above and the units cancel appropriately to yield Joules.

Example: If 4490 J of heat are added to 50.0 g of solid silver at 25.0°C, what would the final temperature be?

Solution: $Q = mC\Delta t$ so $\Delta t = \frac{Q}{mC}$

$$\Delta t = \frac{Q}{mC} = \frac{4490 \text{ J}}{(50.0 \text{ g})(0.235 \text{ J/g}\cdot^\circ\text{C})} = 382^\circ\text{C}$$

Final temperature = initial temperature + $\Delta t = 25^\circ\text{C} + 382^\circ\text{C} = 407^\circ\text{C}$

The energy absorbed or given off by a substance during a phase change (with no temperature change) can be calculated with the equations, $Q = m\Delta H_{fusion}$ or $Q = m\Delta H_{vap}$, where Q is the amount of heat in Joules, m is the mass of the substance in grams, and ΔH_{fusion} or ΔH_{vap} is the heat of fusion or vaporization.

Example: How many Joules are required to melt 17.7 grams of solid aluminum at its normal melting point with no temperature change?

Solution: $Q = m\Delta H_{fusion} = (17.7 \text{ g})(400. \text{ J/g}) = 7080 \text{ J}$

When heat is added to a substance such that the substance undergoes both a temperature change and a phase change, the problem is solved separately for each process. For example, if sufficient heat is added to solid water (ice) at -20°C to raise the temperature and cause the necessary phase changes, the solid water will go through five processes; 1. the temperature of the ice will be raised to the melting point, 2. the solid water will be melted, 3. the temperature of the liquid water will be raised to the boiling point, 3. the liquid will be vaporized, and 5. the temperature of the gaseous water will be raised to the final temperature. To do calculations for this entire process, many bits of thermodynamic data will be required. We would need to know the specific heat of solid water (not the same as liquid water), the heat of fusion for water, the specific heat of liquid water, the heat of vaporization, and the specific heat of gaseous water.

Example: Calculate the heat necessary to raise 100. g of iron at 25.0°C to liquid iron at 2000.°C. The necessary thermodynamic data are: melting point of iron = 1540.°C, specific heat of solid iron = 0.450 J/g·°C, specific heat of liquid iron = 0.770 J/g·°C, heat of fusion of iron = 280. J/g.

Solution:

Step 1: Heat the solid iron from 25.0°C to its melting point at 1540.°C ($\Delta t = 1515^\circ\text{C}$).

$$Q = mC\Delta t = (100. \text{ g})(0.450 \text{ J/g}\cdot^\circ\text{C})(1515^\circ\text{C}) = 68,200 \text{ J}$$

Step 2: Melt the solid iron to liquid.

$$Q = m\Delta H_{\text{fusion}} = (100. \text{ g})(280. \text{ J/g}) = 28,000 \text{ J}$$

Step 3: Heat the liquid iron from the melting point (1540.°C) to the final temperature (2000.°C) $\Delta t = 460.^\circ\text{C}$.

$$Q = mC\Delta t = (100. \text{ g})(0.770 \text{ J/g}\cdot^\circ\text{C})(460^\circ\text{C}) = 35,400 \text{ J}$$

Step 4: Add up the heat added for each step to get the total.

$$Q_{\text{TOTAL}} = 68,200 \text{ J} + 28,000 \text{ J} + 35,400 \text{ J} = 131,600 \text{ J} = 131.6 \text{ kJ} = 132 \text{ kJ}$$

Example: Calculate the heat necessary to raise 40.00 g of ice at -50.0°C to water vapor at 180.0°C .

Necessary Thermodynamic Data

- $C_{\text{ice}} = 2.09 \text{ J/g}\cdot^\circ\text{C}$
- $C_{\text{water}} = 4.18 \text{ J/g}\cdot^\circ\text{C}$
- $C_{\text{water vapor}} = 2.01 \text{ J/g}\cdot^\circ\text{C}$
- Melting Point = 0°C
- Boiling Point = $100.^\circ\text{C}$
- $\Delta H_{\text{fusion}} = 334 \text{ J/g}$
- $\Delta H_{\text{vap}} = 2260 \text{ J/g}$

Solution:

Step 1: Raise the temperature of the ice from -50.0°C to the melting point 0°C .

$$Q = mC\Delta t = (40.00 \text{ g})(2.09 \text{ J/g}\cdot^\circ\text{C})(50.00^\circ\text{C}) = 4,180 \text{ J}$$

Step 2: Melt the ice to liquid water.

$$Q = m(\Delta H)_{\text{fusion}} = (40.00 \text{ g})(334 \text{ J/g}) = 13,360 \text{ J}$$

Step 3: Raise the temperature of the liquid water from the m.p. to the b.p. ($\Delta t = 100.^\circ\text{C}$).

$$Q = mC\Delta t = (40.00 \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(100.^\circ\text{C}) = 16,720 \text{ J}$$

Step 4: Vaporize the liquid water.

$$Q = m\Delta H_{\text{vap}} = (40.00 \text{ g})(2260 \text{ J/g}) = 90,400 \text{ J}$$

Step 5: Raise the temperature of the gaseous water from the b.p. to the final temperature ($\Delta t = 100.^\circ\text{C}$).

$$Q = mC\Delta t = (40.00 \text{ g})(2.01 \text{ J/g}\cdot^\circ\text{C})(80.^\circ\text{C}) = 6,400 \text{ J}$$

Step 6: Add up the results of each step.

$$Q_{\text{TOTAL}} = 4180 + 13360 + 16720 + 90400 + 6400 = 131,000 \text{ J} = 131 \text{ kJ}$$

Questions and Exercises

The thermodynamic data necessary for these problems can be found in the preceding pages.

1. Assuming no phase change occurs, what happens to the temperature of a substance when it absorbs heat?
2. What happens when two objects at different temperatures are brought into contact?
3. How many Joules of heat must be added to 5000. g of water to change its temperature from $20.^\circ\text{C}$ to $80.^\circ\text{C}$?
4. If 500. g of water at $25.^\circ\text{C}$ loses 10,000. J of heat, what will its final temperature be?
5. What does the temperature of an object actually measure?
6. At what temperature do molecules have zero kinetic energy?
7. Describe a situation where heat can enter a body without causing an increase in temperature?
8. How much heat is released when 44.8 g of solid gold are cooled from $80.^\circ\text{C}$ to 62°C ?
9. How much heat is needed to melt 25.0 g of silver at its normal melting point?
10. How much heat is absorbed when 24.5 g of ice at -10.0°C is warmed to liquid water at 42.5°C ?
11. Calculate the amount of heat necessary to raise 45.0 g of cesium metal from 24.0°C to 880.0°C . Use the data given below.

Necessary Thermodynamic Data

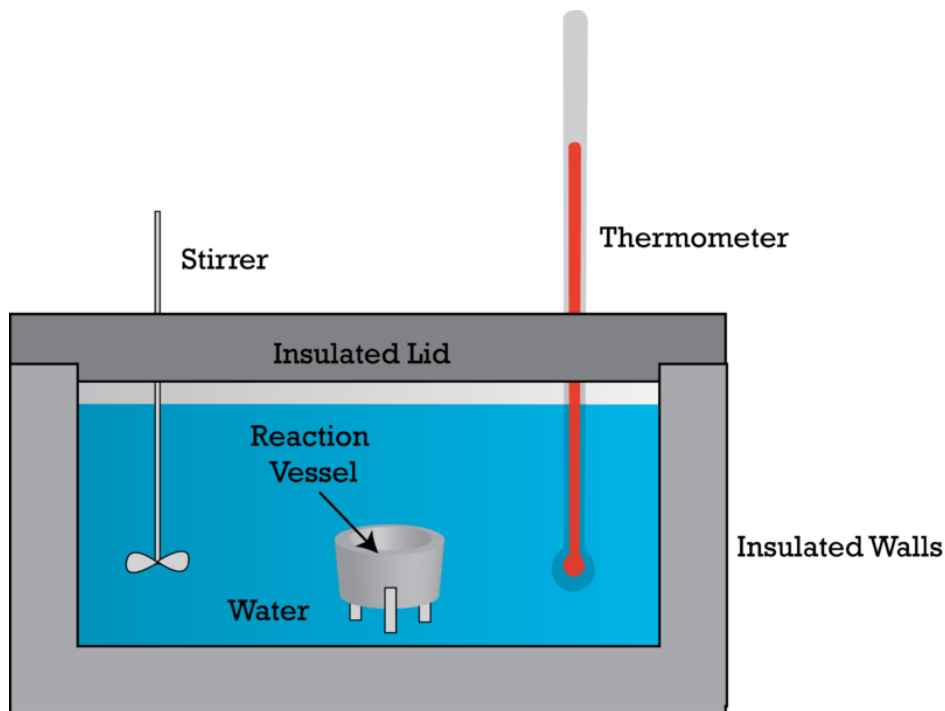
- $C_{\text{solid } Cs} = 0.251 \text{ J/g}\cdot^\circ\text{C}$
- $C_{\text{liquid } Cs} = 0.209 \text{ J/g}\cdot^\circ\text{C}$
- $C_{\text{gaseous } Cs} = 0.167 \text{ J/g}\cdot^\circ\text{C}$
- Melting Point = 29.0°C
- Boiling Point = 690.0°C
- $\Delta H_{\text{fusion}} = 16.3 \text{ J/g}$
- $\Delta H_{\text{vap}} = 669 \text{ J/g}$

Calorimetry Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

The laboratory process for measuring the amount of heat gained or during a chemical reaction or other energy exchange involves the use of an instrument called a *calorimeter*. The basic idea of a calorimeter is sketched below.



The calorimeter has an insulated container to eliminate heat exchange with the outside, a reaction vessel where the reaction to be measured will occur, a quantity of water to absorb from or give up to the heat from the reaction, a thermometer to accurately measure the temperature of the water, and a stirring rod to assure that all the water is the same temperature. Since the heat will come out of or go into the reaction vessel, it is likely that the water touching the vessel would be warmer or colder than the remainder of the water. The stirring rod is used to keep the water circulating and thus all the water will be the same temperature.

At an earlier time, the unit chemists used to measure heat was the calorie. The words *calorimeter* (the name of the instrument) and *calorimetry* (the name of the process) came from the unit, calorie. When scientists decided to use the same units in all branches of science, chemists changed their unit for heat (and all other forms of energy) from calories to **Joules**. The old unit *calorie* is equal to 4.18 *Joules*. Even though chemists don't use the calorie unit anymore, the words calorimeter and calorimetry remain with us.

Extremely accurate calorimeters are calibrated before each use. A precisely known amount of heat is added to the calorimeter and the temperature change is noted. In this way, the scientist can determine exactly how much heat is required to raise the temperature of the calorimeter by 1.00°C . This allows the scientist to measure not only the heat absorbed by the water in the calorimeter but also the heat absorbed by the reaction vessel, the stirrer, the thermometer, and the inside walls of the calorimeter. For a less precise calorimeter, the scientist assumes all the heat added to the calorimeter is absorbed by the water, ignoring the small amount of absorbed by other components.

To use a calorimeter of the less precise type, the scientist measures the amount of water inside very carefully, measures the temperature of the water before the reaction begins, and measures the maximum or minimum temperature the water reaches after the reaction. Since it is assumed that all the heat absorbed or given off by the reaction went into the water, knowing the amount of water and the temperature change of the water, the scientist can then calculate the amount of heat that the water absorbed or gave off, and that is the heat input or output by the reaction. The equation used to calculate the change in heat content of the water is the same one used before, namely $Q = mC\Delta t$.

Example: How much heat was absorbed by 1000. g of water in a calorimeter if the temperature of the water was raised from 23.5°C to 44.8°C ?

Solution: $Q = mC\Delta t = (1000. \text{g})(4.18 \text{ J/g}\cdot^{\circ}\text{C})(21.3^{\circ}\text{C}) = 89,000 \text{ J} = 89 \text{ kJ}$

Example: How much heat was absorbed by 500. g of water in a calorimeter if the water temperature changed from 25.0°C to 17.2°C ?

Solution: $Q = mC\Delta t = mC(t_2 - t_1) = (500. \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(17.2^\circ\text{C} - 25.0^\circ\text{C})$

$$Q = (500. \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(-7.8^\circ\text{C}) = -16,300 \text{ J} = -16.3 \text{ kJ}$$

The negative sign of this result indicates the water in the calorimeter lost heat to the reaction, so the reaction was endothermic.

Calorimeters are used by scientists to measure many types of heat exchanges, such as finding the specific heat of substances, the heat value of fuels, and the heat of chemical reactions. Coal mined in different areas is of different quality. When coal is purchased by users from producers, the price paid is based not only on the mass of coal purchased but also on the amount of heat produced by burning a unit quantity of the coal. When a trainload of coal is delivered, there is a scientist on hand to take samples of the coal and burn them in a calorimeter to determine the average Joules/gram of heat produced by that particular load of coal and the price is adjusted accordingly.

Physicists use calorimeters to determine the specific heat of substances. Suppose we wished to determine the specific heat of brass. We use a calorimeter containing 250. g of water at 25.0°C and into it we place a 100. g piece of brass whose temperature we have raised to 91.0°C. When the heat transfer is complete, the final temperature of the water and the piece of brass are 27.3°C. (Since they are in contact, they must eventually reach the same temperature.) The amount of heat lost by the brass will equal the amount of heat gained by the water. We can use the following equation to find the specific heat of the brass.

$$m_{\text{water}}C_{\text{water}}\Delta t_{\text{water}} = -m_{\text{brass}}C_{\text{brass}}\Delta t_{\text{brass}}$$

The negative sign on the brass side of the equation is present because the heat is being gained by the water and lost by the brass. Therefore, the Δt for the water will be positive but the Δt for the brass will be negative. The heat calculated on the two sides of the equation can only be equal if we change the sign of one of them.

Substituting from the problem yields

$$(250. \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(27.3^\circ\text{C} - 25.0^\circ\text{C}) = -(100. \text{ g})(x \text{ J/g}\cdot^\circ\text{C})(27.3^\circ\text{C} - 91.0^\circ\text{C}).$$

Solving for x yields, $x = 0.377 \text{ J/g}\cdot^\circ\text{C}$

The heat of reaction, ΔH , for a chemical reaction is commonly expressed in J/mole or kJ/mole of product. It is also standard to express the ΔH for an endothermic reaction as a positive number (the reaction is gaining energy) and the ΔH for an exothermic reaction as a negative number (the reaction is losing energy). For the reaction between hydrochloric acid and sodium hydroxide, $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$, the amount of materials necessary to produce one mole of water would be too large for the calorimeter. That is, we can't actually use molar quantities of these materials. Therefore, we use a fraction of a mole and calculate what the heat transfer would have been for an entire mole.

Example: Suppose we carry out the above reaction in a calorimeter. We use 4.00 g of NaOH with excess HCl solution. That means the NaOH will be the limiting reactant. The 4.00 g of NaOH is 0.100 mole and will produce 0.100 mole of H₂O. We use 250. g of water in the calorimeter and the temperature change during the reaction is from 22.4°C to 28.4°C. Calculate the the heat of reaction for the reaction between hydrochloric acid and sodium hydroxide.

Solution: We can calculate the heat absorbed by the water in the calorimeter in the usual way.

$$Q = (250. \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(6.0^\circ\text{C}) = 6,270 \text{ J} = 6.27 \text{ kJ}$$

We can then calculate the ΔH for the reaction by dividing the heat transferred to the water in the calorimeter by the moles of water produced during the reaction. Since the temperature of the water in the calorimeter increased, we know this is an exothermic reaction and therefore, we provide for making the ΔH a negative value . . . required by the definition of ΔH . We can use the following equation.

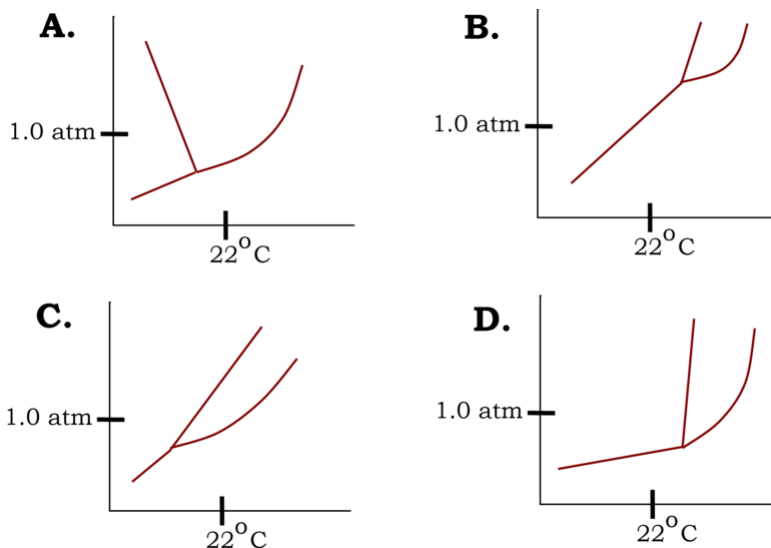
$$\Delta H = \frac{-\Delta Q}{\text{moles product}} = \frac{-6.27 \text{ kJ}}{0.100 \text{ mol}} = -62.7 \text{ kJ/mol}$$

Exercises

- How much heat is absorbed by 1.00 g of water when its temperature changes from 20.0°C to 25.0°C.
- What was the heat transfer if 800. g of water in a calorimeter underwent a temperature change from 25.0°C to 22.0°C?
- A 7.38 g sample of coal is burned in a calorimeter and raises the temperature of 1000. g of water in calorimeter from 22.0°C to 68.8°C. What is the heat content of this coal in J/g?
- A reaction that formed 10.0 g of magnesium oxide, MgO , was carried out in a calorimeter. The calorimeter contained 800. g of water and the temperature of the water increased 44.6°C. What was the ΔH for this reaction in kJ/mol?
- Using the ΔH you found in problem #4, suppose you had carried out exactly this same reaction except that you had used a calorimeter than container 250. g of water instead of 800. g of water. What would the temperature change have been? Give a reason that this reaction wouldn't be carried out with 250. g of water.

Phase Diagram Worksheet

- Which of the following phase diagrams is consistent with a solid that sublimates at 22°C and normal atmospheric pressure?



- Methanol, CH_3OH (bp : 65°C), boils nearly 230° higher than methane, CH_4 (bp : -64°C). What intermolecular forces are responsible for the higher boiling point of methanol?
 - London dispersion forces
 - hydrogen bonds
 - dipole-dipole forces
 - ionic charge-dipole attractions
 - None of these.

3. Decreasing the external pressure will:

- A. increase the vapor pressure of a liquid. B. decrease the boiling point of a liquid. C. increase the boiling point of a liquid. D. weaken the intermolecular forces of attraction in a liquid. E. None of these.

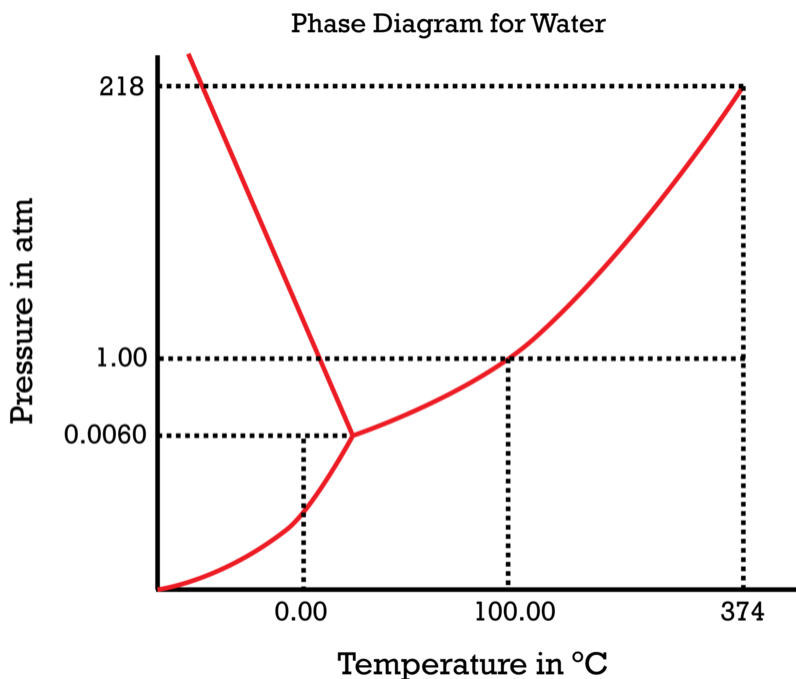
4. Which of the following statements is FALSE?

- A. Water boils at a lower temperature atop Mt. Everest than in San Jose, CA. B. Ice has a lower density than liquid water. C. The vapor pressure of a liquid is lower than the external pressure at temperatures below its boiling point. D. A phase diagram is a pressure-temperature plot which outlines the P-T regions in which a substance will exist as solid, liquid, or gas. E. All of these are true.

5. The triple point is:

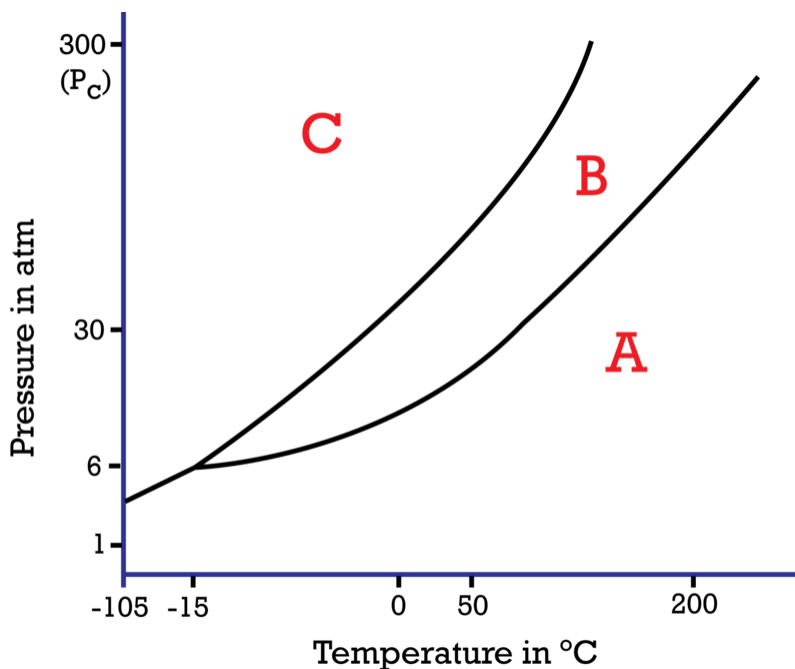
- A. the point in a phase diagram where the solid, liquid, and gas phases of a substance can coexist at equilibrium. B. three places in a phase diagram where the solid, liquid, and gas can coexist. C. the boiling point, the melting point, and the sublimation point of a substance. D. the point in a phase diagram where all molecular motion ceases. E. None of these.

6. What pressure would be required to liquefy (or condense) a sample of water vapor at 500°C ? The critical temperature of water is 374°C and its phase diagram is shown below.



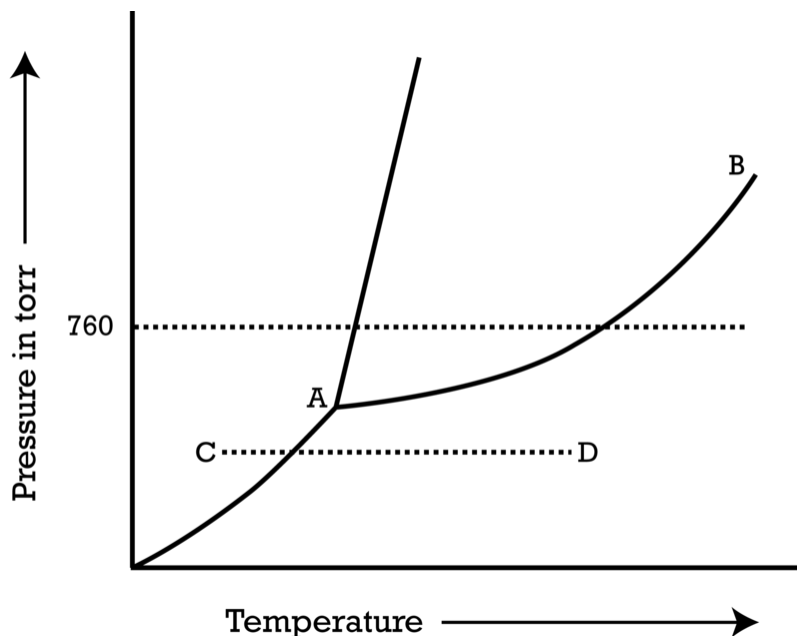
- A. 0.0060 atm B. 1.00 atm C. 218 atm D. some pressure higher than 218 atm E. no amount of pressure can condense water vapor at 500°C

Answer questions 7 – 13 using the phase diagram below.



7. If the temperature of the substance is held constant at -15°C , the phase change that would occur with a pressure increase from 1.0 atm to 30 atm would be
 A. gas to liquid B. gas to solid C. liquid to solid D. solid to liquid E. solid to gas
8. A phase change from area B in the diagram to area C in the diagram is
 A. freezing B. vaporization C. sublimation D. melting E. condensation
9. A phase change from area A in the diagram to area B in the diagram is
 A. freezing B. vaporization C. sublimation D. melting E. condensation
10. The area of the graph that represents the solid phase is
 A. A B. B C. C
11. The critical temperature for this substance is
 A. between -106°C and -16°C B. between -16°C and 0°C C. between -16°C and 200°C D. less than 200°C
 E. greater than 200°C
12. At STP, this substance would exist as a
 A. liquid B. gas C. solid
13. At 30 atm pressure, the boiling point of this substance is approximately
 A. -15°C B. 0°C C. 50°C D. 100°C E. 200°C
14. Sublimation is
 A. exothermic B. endothermic
15. Freezing (solidification) is
 A. exothermic B. endothermic
16. Vaporization is
 A. exothermic B. endothermic

Use the phase diagram for Weissium, We, below to answer questions 17 - 19.



17. Examine the phase diagram for Weissium above and select the correct statement.
- Solid Weissium has a lower density than liquid Weissium.
 - The triple point for We is higher than the normal melting point for We.
 - We changes from solid to liquid as one follows the line from C to D.
 - We changes from liquid to gas as one follows the line from C to D.
 - Point B represents the critical temperature and pressure for We.
18. In the phase diagram for Weissium, identify the feature represented by point A.
- melting point
 - critical point
 - triple point
 - sublimation point
 - boiling point
19. Consider the phase diagram for Weissium and identify the process occurring as one goes from point C to point D.
- increasing temperature with a phase change from solid to liquid
 - increasing temperature with a phase change from solid to vapor
 - increasing temperature with a phase change from liquid to vapor
 - increasing temperature with no phase change
 - increasing temperature beyond the critical point
20. Neon condenses due to
- dipole-dipole forces.
 - London dispersion forces.
 - hydrogen bonding.
 - covalent bonding.
 - metallic bonding.
21. A crystal with a moderate melting point readily conducts heat and electricity and is malleable (changes shape upon hammering without shattering). What is the most likely type of intermolecular forces for this crystal?
- network
 - ionic
 - metallic
 - dipole-dipole
 - London dispersion forces
22. Which of the following substances would produce the hardest, most brittle crystals?
- SiC
 - Xe
 - Cu
 - KNO_3
 - H_2O
23. When water boils, the temperature remains constant. Describe what is happening to the heat added during this time.
24. Explain why different substances can be in different phases at the same temperature.
25. Explain the difference between evaporation and boiling.

Answers to Worksheets

- **The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.**

16.8 Extra Reading for Chapter 16

Refrigeration

Vaporization is the phase change from liquid to gas at the boiling point of the liquid. When this phase change occurs below the boiling point of the liquid, it is called evaporation. Liquids undergo evaporation because while the average temperature of its molecules is less than the boiling point, some of the molecules have temperatures above the boiling point. These hot molecules are the ones that leave the liquid phase and enter the gaseous phase. During both vaporization and evaporation, the amount of liquid that leaves the liquid phase and enters the gaseous phase absorb the heat of vaporization. When the ambient temperature of a gaseous substance is above the boiling point of the liquid of that substance, scientists called the substance a gas. But when the ambient temperature of a gaseous substance is below the boiling point of the liquid of that substance, they call it a vapor. Hence, gaseous water at an ambient temperature of 120°C is water gas and gaseous water at an ambient temperature of 70°C is water vapor.

The process of evaporation has long been used to cool food and drink.

Canteens are frequently covered in fabric or carried in a fabric holder. The canteen user wets the fabric when filling the canteen so that as the water evaporates from the fabric, it absorbs the heat of vaporization from the canteen and cools the canteen, making the water more pleasant to drink.

Some people put butter on the dinner table with the dish holding the butter sitting inside another dish half-filled with water. As the water in the outside dish evaporates, it absorbs the heat of vaporization, cools the butter dish and keeps the butter from melting.

Before the days of the portable ice chest, people who took bottled or canned drinks on a picnic would often keep the drinks in a fabric bag that they would soak with water on arrival at the picnic spot. The evaporation of the water would keep the drinks much cooler than if they were sitting out on a table.

The function of refrigerators and air conditioners also involve the heat of vaporization of liquids.

Many gaseous substances can be compressed until they become liquids. That is, the molecules are pushed together forcefully until they touch and the gas becomes a liquid. In this process, the gas also gives up the heat of vaporization as it becomes a liquid. By compressing the coolant to a liquid outside the refrigerator, the phase change gives up the heat of vaporization outside the refrigerator. The liquid coolant is then pumped through a tube inside the refrigerator where it is allowed to vaporize back to gas, thus absorbing the heat of vaporization inside the refrigerator. Then the gas is pumped outside the refrigerator and again compressed to liquid, giving up the heat of vaporization. In this manner, heat is absorbed from inside the refrigerator and given off outside the refrigerator. The inside gets colder and the outside gets warmer. This is why you can feel heat coming from the back or from underneath a refrigerator. This is also why the compressor for an air conditioner must be outside the house. It wouldn't do much good to absorb the heat and release the heat both inside the house. So much for the idea of cooling the house by leaving the refrigerator door open.

Liquid Crystals

LCD or liquid crystal displays have become a ubiquitous part of our technology landscape. Now appearing as computer and television screens and other electronic displays, even in new automobile dashboard devices. Yet for

chemistry students with an alert ear, the seemingly contradictory term Liquid Crystal, should at very least, merit additional questions. We know that crystalline lattices are structures characteristic of the solid phase, with atoms or ions limited in their positions to vibrational motion, in place of the translational capabilities due to the decrease in intermolecular attractive forces in the liquid state.

There are currently thousands of different compounds, however that display behaviors intermediate between that of the liquid and solid state. Liquid crystals are arranged in a regular, orderly pattern yet their individual molecules can flow like liquids. The types of molecules that tend to form liquid crystals are usually cylindrically – shaped, with a polar group at one end of the molecule. This shape allows different opportunities for arrangement, such as orientation in the same directions (nematic), or alignment in layers (smectic). These arrangements are due to the presence of dipole – dipole or hydrogen bonding interactions or, at times, a combination of both forces.



Due to their unusual structural arrangements, liquid crystals exhibit interesting thermal, optical, and electronic properties. Some liquid crystal samples will react to changes in temperature. You may have used a body thermometer that display the temperature with a liquid crystal. Pressure–sensitive liquid crystals have been implemented in the design of fingerprint detection devices. More commonly, the application of an electric or magnetic field can result in the realignment of a liquid crystal sample which in turn causes a change in the visual display. Most nematic liquid crystals are transparent or translucent but with the application of an electrical field, the molecular orientation alters and the display becomes opaque.

The current popularity of devices containing liquid crystals continues to grow as the demand for light–weight, flexible display technology increases. Some future uses for this technology include the incorporation of liquid crystals into carbon nanotubes to create three–dimensional arrays. Another interesting potential application of liquid crystals includes their use in an anti–cancer drug, as well as the use of liquid crystals in cosmetics and personal care products.



16.9 Assessment for Chapter 16

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 17

Solutions and Their Behavior C-SE-TE

Chapter Outline

- 17.1 PROPERTIES OF SOLUTIONS
- 17.2 SOLUTION FORMATION
- 17.3 MEASURING CONCENTRATIONS
- 17.4 FACTORS AFFECTING SOLUBILITY
- 17.5 SOLUBILITY GRAPHS
- 17.6 COLLIGATIVE PROPERTIES
- 17.7 SEPARATING MIXTURES
- 17.8 REACTIONS BETWEEN IONS IN SOLUTIONS
- 17.9 WORKSHEETS FOR CHAPTER 17
- 17.10 EXTRA READING FOR CHAPTER 17
- 17.11 ASSESSMENT FOR CHAPTER 17

Lessons and Number of Activities for Lessons

TABLE 17.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Properties of Solutions	0	0	0	0
2. Solution Formation	0	1	0	2
3. Measuring Concentration	0	0	3	1
4. Factors Affecting Solubility	0	0	0	0
5. Solubility Graphs	0	0	0	0
6. Colligative Properties	0	0	3	1
7. Separating Mixtures	1	0	0	0
8. Reactions Between Ions in Solution	1	0	1	0

17.1 Properties of Solutions

Student Behavioral Objectives

The student will:

- define solutions.
- describe the composition of homogeneous solutions.
- describe the different types of solutions that are possible within the three states of matter.
- identify homogeneous solutions of different types.

Timing, Standards, Activities

TABLE 17.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Properties of Solutions</i>	1.0	6a

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Properties of Solutions (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

17.2 Solution Formation

Student Behavioral Objectives

The student will:

- describe why solutions occur; the “like dissolves like” generalization.
- determine if solutions will occur by studying the molecular structure.
- state the importance of water as the “universal solvent.”

Timing, Standards, Activities

TABLE 17.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Solution Formation</i>	1.0	6b

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. Conductivity of Solutions

Worksheets

1. None

Extra Readings

1. Solutions
2. Properties of Solutions

Answers for Solution Formation (L2) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

17.3 Measuring Concentrations

Student Behavioral Objectives

The student will:

- define molarity, mass percent, ppm, and molality.
- calculate molarity, mass percent, ppm, and molality.
- explain the importance of quantitative measurement in concentration.

Timing, Standards, Activities

TABLE 17.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Measuring Concentrations</i>	2.0	6a, 6d

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Concentration by Percent Mass Worksheet
2. Mole Fraction and Molality Worksheet
3. Molarity Worksheet

Extra Readings

1. Methods of Expressing Concentration

Answers for Measuring Concentrations (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

17.4 Factors Affecting Solubility

Student Behavioral Objectives

The student will:

- define solute, solvent, soluble, insoluble, miscible, immiscible, saturated, unsaturated, concentrated, and dilute.
- describe the factors that affect solid solubility.
- describe the factors that affect gas solubility.
- describe how pressure can affect solubility.

Timing, Standards, Activities

TABLE 17.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Factors Affecting Solubility</i>	1.0	6c, 6d

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Factors Affecting Solubility (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

17.5 Solubility Graphs

Student Behavioral Objectives

- Students will read and report data from solubility graphs.
 - Students will read and report saturation points from a solubility graph.
-

Timing, Standards, Activities

TABLE 17.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Solubility Graphs</i>	1.0	6c

Activities for Lesson 5

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None
-

Answers for Solubility Graphs (L5) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

17.6 Colligative Properties

Student Behavioral Objectives

The student will:

- describe vapor pressure lowering.
- define boiling point elevation and freezing point depression.
- describe what happens to the boiling points and freezing points, when a solute is added to a solvent.
- describe the importance of the Van't Hoff factor.
- calculate the boiling point elevation for electrolyte and non-electrolyte solutions.
- calculate the freezing point depression for electrolyte and non-electrolyte solutions.

Timing, Standards, Activities

TABLE 17.7: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Colligative Properties</i>	1.5	6c, 6e

Activities for Lesson 6

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Dilution Worksheet
2. Colligative Properties: Solution Vapor Pressure Worksheet
3. Colligative Properties: B.P. Elevation and M.P. Depression Worksheet

Extra Readings

1. Colligative Properties

Answers for Colligative Properties (L6) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org

to request sample answers.

17.7 Separating Mixtures

Student Behavioral Objectives

- The students will describe differences between the physical properties of pure substances and solutions.
- The students will list and describe methods of separation for mixtures.
- The students will explain the principles involved in chromatographic separation.
- The students will identify the mobile and stationary phases in a chromatography design.
- Given appropriate data, the students will calculate Rf values.

Timing, Standards, Activities

TABLE 17.8: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Separating Mixtures</i>	1.5	6e, 6f

Activities for Lesson 7

Laboratory Activities

1. Separation of a Mixture Lab

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Separating Mixtures (L7) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

17.8 Reactions Between Ions in Solutions

Student Behavioral Objectives

The student will:

- use the solubility chart and/or solubility rules to determine if substances are soluble in water.
 - use the solubility chart and/or the solubility rules to determine if precipitates will form.
 - write molecular, ionic, and net ionic equations.
 - identify spectator ions in ionic equations.
-

Timing, Standards, Activities

TABLE 17.9: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Reactions Between Ions in Solution</i>	1.5	6f

Activities for Lesson 8

Laboratory Activities

1. Qualitative Ion Testing Lab

Demonstrations

1. None

Worksheets

1. Reactions Between Ions in Solution Worksheet

Extra Readings

1. None
-

Answers for Reactions Between Ions in Solution (L8) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 17

This website provides a virtual mixtures lab activity.

- <http://www.harcourtschool.com/activity/mixture/mixture.html>

This website provides a cleaning water activity.

- http://acswebcontent.acs.org/games/clean_water.html

This website provides a lesson on the factors affecting solubility.

- <http://www.chem.lsu.edu/lucid/tutorials/solubility/Solubility.html>

Lecture slides on solution properties and their behavior are available at the following website.

- http://www.public.asu.edu/~jpbirk/CHM-113_BLB/Chpt13/sld002.htm

A student defines solute and solvent and this animation shows the hydrolysis involved in dissolving.

- <http://www.youtube.com/watch?v=hydUVGUbyvU>

This animation shows dissociation of salt in water creating sodium chloride solution.

- <http://www.youtube.com/watch?v=EBfGcTAJF4o>

This website has lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 6-4 is on molarity.

- <http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson64.htm>

This website details some of the factors that affect solubility.

- <http://www.citycollegiate.com/chapter3d.htm>

This video explains the reasoning behind and the math involved in boiling point elevation and freezing point depression.

- <http://www.youtube.com/watch?v=z9LxdqYntlU>

The following link hosts several videos on how to solve various types of chemistry solutions problem.

- <http://www.kentchemistry.com/moviesfiles/chemguy/AP/ChemguySolutions.htm>

This video presents thin layer chromatography with fluorescent materials, and column chromatography with UV active materials. There is no narration on the video so it would be advantageous to watch with a chemistry teacher.

- <http://www.youtube.com/watch?v=gzp2S0e9o8s>

An interactive video on separating mixtures is available at the link below.

- http://www.bbc.co.uk/schools/ks3bitesize/science/chemical_material_behaviour/compounds_mixtures/activity.shtml

This web site video shows a series of precipitation reactions.

- <http://www.youtube.com/watch?v=RjBjwQF276A>

This video is a ChemStudy film called “Molecular Spectroscopy.” The film is somewhat dated but the information is accurate.

- <http://www.youtube.com/watch?v=mNaipM3WEO0>

Laboratory Activities for Chapter 17

Teacher’s Pages for Separation of a Mixture Lab

Lab Notes

The time needed to prepare this lab is minimal. Each student pair will require 2 *grams* of iron filings, 0.5 *gram* of grated styrofoam, 5 *grams* of sand, and 5 *grams* of salt. The best way to prepare the styrofoam is to grate it using a household grater, cheese grater, or rasp. You can either prepare the mixture beforehand or have the students do it - there are advantages and disadvantages to both choices. If you wish to do a percent composition analysis, it is best to eliminate the Styrofoam component due to density differentiation when the students handle the mixture, yielding differing results for all teams.

Make sure the students leave the magnets within the Ziploc bags. This will ensure clean separation of the iron filings.

Answers to Pre-Lab Questions

1)

- A. boiling point
- B. intermolecular attraction
- C. particle size and solubility
- D. solubility
- E. boiling point

Separation of a Mixture

Background Information

One of the chemists primary job is to identify what materials are present in a mixture. To do this, the components of the mixture have to be separated from one another. This is done by exploiting the fact that the components of the mixture have different physical properties such as particle size, boiling points, density, solubilities, and intermolecular attraction for different substances.

This laboratory uses the differences in magnetism, density, and solubility to separate a mixture of Styrofoam, sand, salt, and iron filings.

Pre-Lab Questions

1. State the physical property used to separate mixtures with the following techniques:
 - a. distillation
 - b. chromatography
 - c. filtration
 - d. solvent extraction
 - e. fractional crystallization

Purpose

To learn different techniques of separating a mixture.

Apparatus and Materials

- Hot plate
- Iron ring stand and ring
- Filter funnel
- Filter paper
- 250 mL beaker
- Stirring rod
- Forceps
- Magnet in a Ziploc bag

Safety Issues

No particular hazards. Salt water can be irritating. Inhalation of dust can be harmful. Use normal laboratory safety techniques.

Procedure for Part I

1. Obtain separate samples of each of the four mixture components from your teacher. Use the equipment you have available to make observations of the components and determine their properties. You may need to run several tests with each substance so don't use all your sample on the first test. Look for things like whether the substance is magnetic, whether it dissolves or whether it floats. Record your observations in your data table. **DO NOT PLACE THE MAGNET IN THE SAMPLE AT ANY TIME! ALWAYS USE THE MAGNET OUTSIDE THE PLASTIC BAG!**
2. Make a plan for what you will do to separate a mixture that includes the four components from step one. Review your plan with your teacher.
3. Obtain a sample of the mixture from your teacher. Using only the equipment you have available, run the procedure you have developed. You should have each of the components (in its pure form) when you have finished.
4. Clean your lab station. Clean all equipment and return to its proper place. **DO NOT DISPOSE OF ANY SOLIDS IN THE SINK! RETURN LEFT OVER SAMPLES TO YOUR TEACHER.**

TABLE 17.10: Data

	Magnetic?	More or Less Dense than H_2O ?	Soluble in H_2O ?
Styrofoam			
Salt			
Sand			
Iron Filings			

Post-Lab Questions

1. Do some research to determine how to separate each of the following two-part mixtures:
 - a. Lead filings and iron filings
 - b. Sand and gravel
 - c. Sand and finely ground polystyrene foam
 - d. Alcohol and water

Teacher's Resource Page for Qualitative Ion Testing Lab

Lab Notes

If you have several classes to do the lab, 500 mL of each solution will be adequate. Students have fewer spills and create less mess with dropper bottles. You will need to refill dropper bottles from larger bottles after each class. If you purchase a large supply of dropper bottles, the labeled bottles can be stored and used from year to year. Making and removing labels consumes a large amount of preparation time.

This lab may require more than one day if your lab periods are 50 minutes or less per day. You can also alter the lab so that your students only test for the cations or for the anions but not both.

Solution Preparation

3.0 M NH_4OH Concentrated NH_4OH is 14.5 M so 500 mL can be prepared by diluting 103 mL of concentrated NH_4OH to 500 mL.

3.0 M HNO_3 Concentrated nitric acid is 15.6 M so 500 mL can be prepared by diluting 96 mL of concentrated HNO_3 to 500 mL.

0.10 M $Ba(NO_3)_2$ Dissolve 13.1 grams of $Ba(NO_3)_2$ in sufficient water to make 500 mL of solution. Since this solution is also used as a testing solution (see below), you can dissolve 26.2 grams in sufficient water to make 1.0 L of solution and divide the solution to avoid preparing it again.

0.10 M $KSCN$ Dissolve 4.86 grams of $KSCN$ in sufficient water to make 500 mL of solution.

0.10 M $AgNO_3$ Dissolve 8.5 grams of $AgNO_3$ in sufficient water to make 500 mL of solution. Since this solution is also used as a testing solution (see below), you can dissolve 17.0 grams in sufficient water to make 1.0 L of solution and divide the solution to avoid preparing it again.

Unknown salts for students to test, use $BaCl_2$ or Na_2SO_4 or Li_2CO_3 (recommended).

For testing solutions:

Barium, Ba^{2+} , use 0.1 M $Ba(NO_3)_2$; Dissolve 13.1 grams of $Ba(NO_3)_2$ in sufficient water to make 500 mL of solution.

Iron(III), Fe^{3+} , use 0.1 M $Fe(NO_3)_3$; Dissolve 12.1 grams of anhydrous $Fe(NO_3)_3$ or 20.7 grams of $Fe(NO_3)_3 \cdot 9H_2O$ in sufficient water to make 500 mL of solution. Lithium, Li^+ , use 0.1 M $LiNO_3$; Dissolve 3.45 grams of $LiNO_3$ in sufficient water to make 500 mL of solution. Potassium, K^+ , use 0.1 M KNO_3 ; Dissolve 5.06 grams of KNO_3 in sufficient water to make 500 mL of solution.

Silver, Ag^+ , use 0.1 M $AgNO_3$; Dissolve 8.5 grams of $AgNO_3$ in sufficient water to make 500 mL of solution.

Sodium, Na^+ , use 0.1 M $NaNO_3$; Dissolve 4.25 grams of $NaNO_3$ in sufficient water to make 500 mL of solution.

Carbonate, CO_3^{2-} , use 0.1 M Na_2CO_3 ; Dissolve 5.30 grams of Na_2CO_3 in sufficient water to make 500 mL of solution.

Chloride, Cl^- , use 0.1 M $NaCl$; Dissolve 2.93 grams of $NaCl$ in sufficient water to make 500 mL of solution.

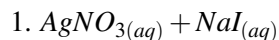
Iodide, I^- , use 0.1 *MNaI*; Dissolve 7.50 grams of *NaI* in sufficient water to make 500 mL of solution.

Sulfate, SO_4^{2-} , use 0.1 *MNa₂SO₄*; Dissolve 7.10 grams of *Na₂SO₄* in sufficient water to make 500 mL of solution.

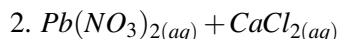
For flames tests of Li^+ , Na^+ , K^+ , and Ba^{2+} ions, place 25 mL of each of the testing solutions in separate 150 mL (labeled) beakers and stand 15 wooden splints (one for each lab group) in the beakers to soak.

Sometimes, the barrel opening of Bunsen burners are contaminated by previous spills and will produce colored flames without a testing splint in them. You should check the burners to make sure they do not produce colored flames when burning. You should also remind students of burner safety procedures . . . no loose hair that falls past your face, etc.

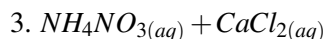
Answers to Pre-Lab Questions



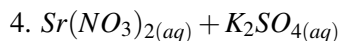
Yes, precipitate forms, *AgI*



Yes, precipitate forms, *PbCl₂*



No precipitate forms



Yes, precipitate forms, *SrSO₄*

Qualitative Ion Testing Lab

Background:

How are unknown chemicals analyzed? One method is by making comparisons to “known” chemicals. In this lab activity, ion tests will be performed and observations made for the reactions of four known anions and six known cations. Then an unknown salt will be identified by analyzing and comparing results to what is known.

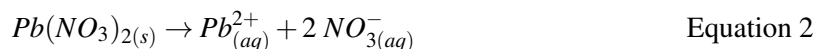
The process of determining the identities of unknown substance is called qualitative analysis. This can be contrasted to quantitative analysis, which is the process of determining how much of a given component is present in a sample. Qualitative analysis procedures use physical tests as well as chemical tests. The physical tests in this lab involve observing colors of solutions and colors produced in flame tests. The chemical tests in this lab involve chemical reactions, as evidenced by formation of a precipitate, dissolving of a precipitate to form a complex ion, a color change, or evolution of a gas.

Formation of a Precipitate

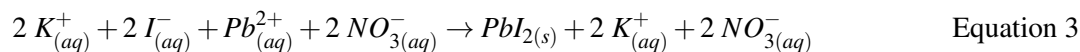
An ionic salt is a compound composed of two parts – cations (positively charged ions) and anions (negatively charged ions). When an ionic salt is dissolved in water, the salt crystal dissociates or separates into its cations and anions. For example, potassium iodide (KI) dissociates into potassium ions (K^+) and iodide ions (I^-) according to equation 1.



Similarly, the ionic salt lead(II) nitrate, $Pb(NO_3)_2$, dissociates into lead cations, Pb^{2+} , and nitrate anions, NO_3^- , according to equation 2.



When two ionic salts are mixed together in water, two new combinations of cations and anions are possible. In some cases, the cation from one salt and the anion from the other salt may combine to form an insoluble product, which is called a precipitate. For example, if solutions of potassium iodide and lead(II) nitrate are mixed together, a solid precipitate of lead(II) iodide, PbI_2 , forms, as shown in equation 3.

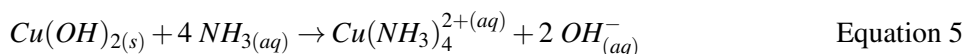


Notice that the potassium cations and the nitrate anions remain dissolved in solution. They did not change during the reaction and are therefore, referred to as spectator ions. In net ionic equations, spectator ions are omitted. A net ionic equation is one that includes only the ions participating in the reaction. Thus, equation 3 can be reduced to equation 4.



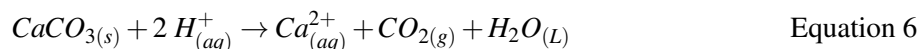
Dissolving Precipitates through Complex-Ion Formation

A complex ion is a water-soluble, charged species containing a central atom and other molecules bonded to it. The formation of a complex ion is commonly evidenced by the dissolution of a precipitate. For example, copper(II) hydroxide, $Cu(OH)_2$, is insoluble in water but will dissolve when excess ammonia is added to it, forming a soluble copper amine complex ion, $Cu(NH_3)_4^{2+}$, according to equation 5.



Evolution of a Gas

Certain anions, such as the carbonate ion, CO_3^{2-} , and sulfide ion, S^{2-} , evolve gas when treated with a dilute strong acid. For example, the reaction of calcium carbonate, $CaCO_3$, with nitric acid, HNO_3 , produces carbon dioxide gas, CO_2 , according to equation 6.



Flame Colors

Some metallic salts will display a distinctive color of light when placed in a flame. When the colored light from any one of these flames is passed through a prism or viewed through a diffraction grating, a portion of the spectrum is visible, containing only a few colors at specific wavelengths, including the colors in the original flame. A partial spectrum that contains only discrete lines is called a line spectrum. When heated in a flame, electrons in the metal absorb energy from the flame and are promoted to excited energy levels. They emit light as they relax back down to the ground state. Each line in the spectrum represents a different electronic transition. Since each element has a unique electronic configuration, an element's line spectrum, and thus its flame color, is unique and can be used for identification.

Pre-Lab Questions

On the last page of this laboratory packet is a listing of solubility rules. Use those rules to determine which of the following mixtures of solutions would produce a precipitate. Write the formula for the precipitate where one forms.

1. $AgNO_3(aq) + NaI(aq)$
2. $Pb(NO_3)_2(aq) + CaCl_2(aq)$
3. $NH_4NO_3(aq) + CaCl_2(aq)$
4. $Sr(NO_3)_2(aq) + K_2SO_4(aq)$

Purpose

In parts I and II of this lab, qualitative tests for four known anions and six known cations will be performed. Test results will be noted and recorded. In part III, the same tests will be performed on an unknown ionic salt which contains one of the six possible cations and one of the four possible anions. The cation and anion that make up the unknown salt will then be identified.

Apparatus and Materials

All solutions should be in drop control dispenser bottles.

- Ammonium hydroxide, NH_4OH , 3.0 M
- Barium nitrate solution, $Ba(NO_3)_2$, 0.10 M
- Nitric acid solution, HNO_3 , 3.0 M
- Potassium thiocyanate solution, $KSCN$, 0.10 M
- Silver nitrate solution, $AgNO_3$, 0.10 M
- Unknown salt solution
- Distilled water
- Beral-type pipets, labeled
- Bunsen burner setup
- Reaction plate, 24-well
- Sheet of notebook paper
- Wooden splints
- Beakers, 250 – mL, 15 each
- Watchglass, 15 each

Cation Testing Solutions (0.10 M solutions of the nitrates)

- Barium, Ba^{2+}
- Iron(III), Fe^{3+}
- Lithium, Li^+
- Potassium, K^+
- Silver, Ag^+
- Sodium, Na^+

Anion Testing Solutions (0.10 M solutions of the sodium or potassium compounds)

- Carbonate, CO_3^{2-}
- Chloride, Cl^-
- Iodide, I^-
- Sulfate, SO_4^{2-}

Safety Issues

All solutions are irritating to skin, eyes, and mucous membranes, particularly the 3.0 M NH_4OH and HNO_3 solutions. The silver nitrate solution will turn skin and clothes permanently black. Handle solutions with care, avoid getting the material on you, and wash your hands carefully before leaving the lab. As with all labs, do not mix any chemicals other than the ones you are directed to mix.

Procedure**Part I – Anion Testing for Cl^- , I^- , SO_4^{2-} and CO_3^{2-}** **Preparing the well plate:**

1. Obtain a 24-well reaction plate and set it on the lab bench with a piece of white paper underneath it. Label the paper as shown in Data Table 1. Notice that the 24-well reaction plate is divided into 6 columns (1-6) and 4 rows (A-D).
2. Using a pipet, add 5 drops of the Cl^- anion-testing solution to wells A, B, and C of column 1.
3. Using a pipet, add 5 drops of the I^- anion-testing solution to wells A, B, and C of column 2.
4. Using a pipet, add 5 drops of the SO_4^{2-} anion-testing solution to wells A, B, and C of column 3.
5. Using a pipet, add 5 drops of the CO_3^{2-} anion-testing solution to wells A, B, and C of column 4.

Performing the tests:**Silver Nitrate Test**

6. Add 3 drops of 0.10 $M AgNO_3$ to the first four wells across row A. Observe the formation of precipitates and/or color changes. You may need to remove the paper to see clearly. Record detailed observations in Data Table 1.
7. Add 5 drops of 3.0 $M HNO_3$ to each of the precipitates from step 6. Gently swirl the well plate to stir. Observe which precipitates dissolve and which do not. Record observations in Data Table 1.
8. Add 10-12 drops of 3.0 $M NH_4OH$ to each of the remaining precipitates from step 6. Gently swirl and observe which precipitates dissolve and which do not. Record observations in Data Table 1.

Barium Nitrate Test

9. Add 3 drops of 0.10 $M Ba(NO_3)_2$ to the first four wells across row B. Observe the formation of precipitates and/or color changes. You may need to remove the paper to see clearly. Record detailed observations in Data Table 1.
10. Add 5 drops of 3.0 $M HNO_3$ to each of the precipitates from step 9. Gently swirl the well plate to stir. Observe which precipitates dissolve and which do not. Record observations in Data Table 1.
11. Add 3 drops of 3.0 $M HNO_3$ to the first four wells across row C. Make observations, looking for the strong evolution of gas bubbles. Record observations in Data Table 1.
12. Repeat any tests for which results were unclear in Row D of the well plate. Rinse the plate with plenty of tap water and then rinse with deionized water to prepare the plate for part II.

Part II – Cation Testing for Li^+ , Na^+ , K^+ , Ag^+ , Ba^{2+} , and Fe^{3+} .**Preparing the Well Plate**

13. Obtain a 24-well reaction plate and set it on the table with a piece of notebook paper underneath it. Label the paper as shown in Data Table 2.
14. Using a pipet, add 5 drops of the Li^+ cation-testing solution to the top well (A) of column 1.
15. Using a pipet, add 5 drops of the Na^+ cation-testing solution to the top well (A) of column 2.
16. Using a pipet, add 5 drops of the K^+ cation-testing solution to the top well (A) of column 3.
17. Using a pipet, add 5 drops of the Ag^+ cation-testing solution to the top well (A) of column 4.
18. Using a pipet, add 5 drops of the Ba^{2+} cation-testing solution to the top well (A) of column 5.
19. Using a pipet, add 5 drops of the Fe^{3+} cation-testing solution to the top well (A) of column 6.

Performing the Tests:

20. Observe each solution and record the color of each solution in Data Table 2 (Row A).

Potassium Thiocyanate Test

21. Add 3 drops of 0.10 $MKSCN$ to each of the six wells across row A. Gently swirl the plate to stir. Observe the formation of precipitates and/or color changes. Record detailed observations in Data Table 2. (Row B)

Flame Tests

Note: Several of the cations may be identified using flame tests. The flame tests will be performed on the four cations that did not show a reaction in step 21.

22. Set up a Bunsen burner. Adjust the air so the flame color is blue (NO YELLOW) and a distinct inner blue cone is apparent.

23. Obtain a wooden splint which has been soaking in the Li^+ cation testing solution for at least 15 *minutes*.

24. Hold the wooden splint in the flame, flat side down. (Do not touch the top of the burner with the wooden splint.) The top end of the splint should be placed directly into the inner blue cone. A distinct color should be apparent. Record the flame color in Data Table 2 (Row C). Do not hold the splint in the flame too long or the splint will begin to burn.

25. Repeat steps 24 and 25 using the Na^+ cation-testing solution, then the K^+ cation-testing solution, and finally the Ba^{2+} cation-testing solution. Be careful not to touch the splints together when gathering them.

26. Rinse the well-plate in the sink with plenty of tap water and make a final rinse with deionized water. The splint should be discarded in the waste basket.

Part III – Identification of an Unknown Salt

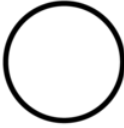
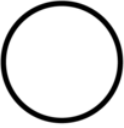
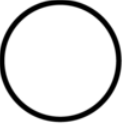
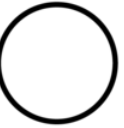
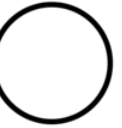
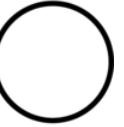
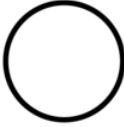
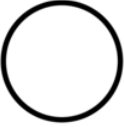
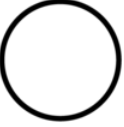
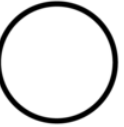
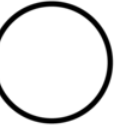
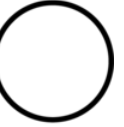
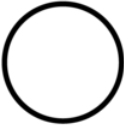
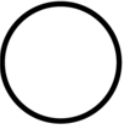
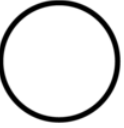
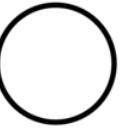
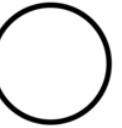
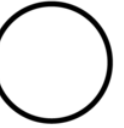
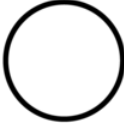
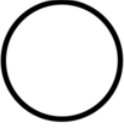
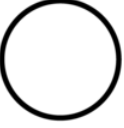
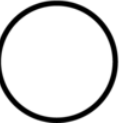
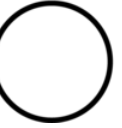
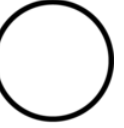
Note: the unknown salt is made up of one of the cations and one of the anions previously tested.

27. Obtain from your teacher a pipet filled with an unknown salt solution and a pre-soaked wooden splint of the same unknown salt. Be sure to record the unknown identifying letter in the Data Tables.

28. Determine the identity of the cation and the anion that make up the unknown salt. To do this, repeat the steps in Part I and Part II. Record all observations for anion testing in Data Table 1 and for cation testing in Data Table 2.

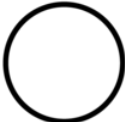
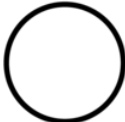
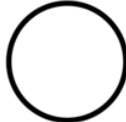
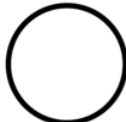
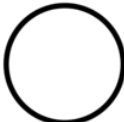
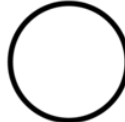
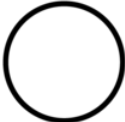
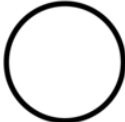
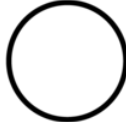
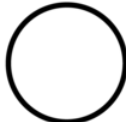
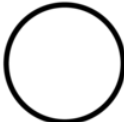
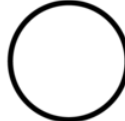
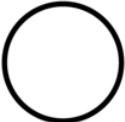
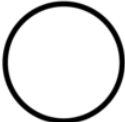
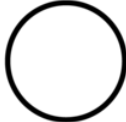
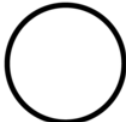
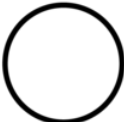
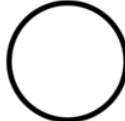
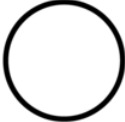
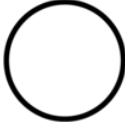
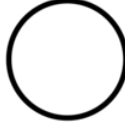
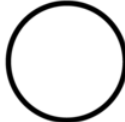
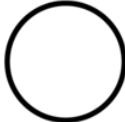
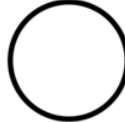
Data for Anion Testing

Record detailed observations inside the circles on the table. Record all colors that form. Record whether any gases evolve. If any solid precipitates form, use the abbreviation PPT. If no reaction occurs, use the abbreviation NR.

	Cl^- anion 1	I^- anion 2	SO_4^{2-} anion 3	CO_3^{2-} anion 4	Unknown 5	Unknown 6
A						
B						
C						
D						

Data for Cation Testing

- In row A, put the original color of the solution.
- In row B, reaction with $KSCN$.
- In row C, flame test color.
- Row D is for unknown testing. In column 1, put the original color of the solution. In column 2, put the results of reaction with $KSCN$. In column 3, put the flame test color. If you do a second unknown, use columns 4, 5, and 6.

	Cl^- anion 1	I^- anion 2	SO_4^{2-} anion 3	CO_3^{2-} anion 4	Unknown 5	Unknown 6
A						
B						
C						
D						

Post-Lab Questions

1. Write the net ionic equation for each precipitation reaction that occurred in Parts I and II. Include the well identification (A1) as shown in the example below.

Part I – Anion testing

Well A1 $Ag^+ + Cl^- \rightarrow AgCl_{(s)}$ (white ppt)

2. Unknown letter _____

(a) What cation is present in your unknown?

(b) What anion is present in your unknown?

3. Write the name and formula for your unknown salt.

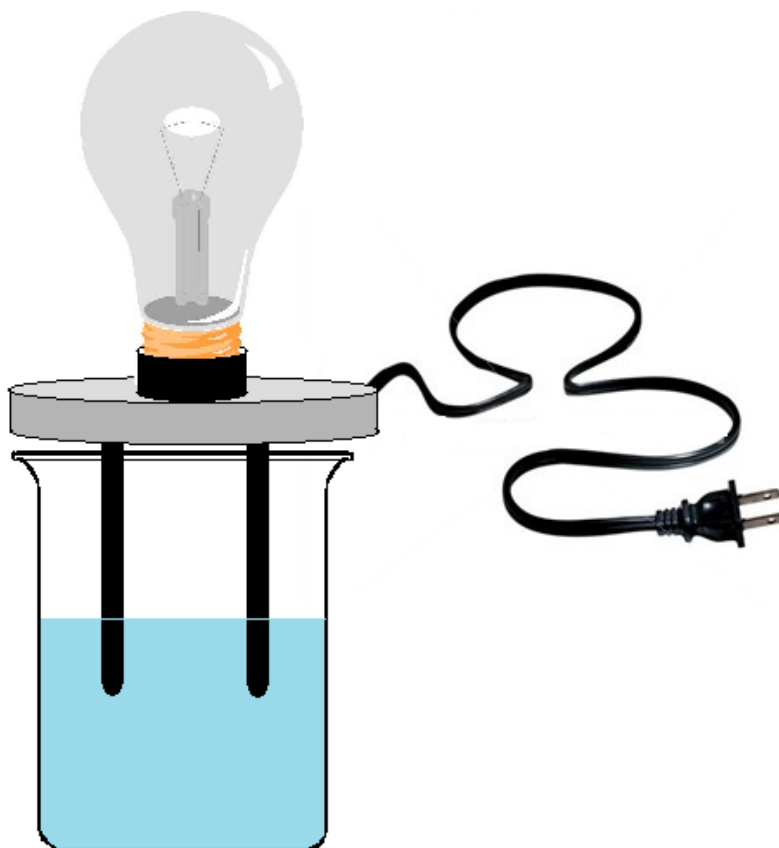
Solubility Rules

- 1. All group 1 metals and ammonium compounds are soluble.
- 2. All nitrates, chlorates, and bicarbonates are soluble.
- 3. Halides are soluble except for Ag^+ , Hg_2^{2+} , and Pb^{2+} .
- 4. Sulfates are soluble except for Ag^+ , Ba^{2+} , Ca^{2+} , Hg_2^{2+} , Sr^{2+} , and Pb^{2+} .
- 5. Carbonates, chromates, phosphates, and sulfides are insoluble except those from rule #1.
- 6. Hydroxides are insoluble except for those in rule #1, and Ba^{2+} .

Demonstrations for Chapter 17

Conductivity of Solutions

Brief description of demonstration



Electrodes of a bulb-type conductivity tester are submerged into acid solutions of increasing strength but equal concentration. The bulb glows brighter with the increasing acid strength.

Materials

- 100 mL distilled water
- 100 mL 0.1 M hydrochloric acid, HCl
- 100 mL 0.1 M acetic acid, $HC_2H_3O_2$
- 100 mL 0.1 M citric acid, $H_3C_6H_5O_7$
- 100 mL 0.1 M malonic acid, $H_2C_3H_2O_4$
- 100 mL 0.1 M ascorbic acid, $HC_6H_7O_6$
- 100 mL 0.1 M propanoic acid, $HC_3H_5O_2$
- 100 mL 0.1 M glycine, $HC_2H_4O_2N$
- 100 mL alanine, $HC_3H_6O_2N$
- 2 mL universal indicator solution, 1 – 10 pH range
- 9 250 mL beakers
- disposable pipette

- 8 stirring rods
- conductivity tester, light bulb type
- wash bottle
- 400 mL beaker

Procedure

Label each of the beakers with the appropriate acid names (or distilled water). Pour 100 mL of each acid into separate beakers. Add 4 drops of universal indicator solution to each beaker and stir. Arrange the beakers in order of the spectrum, from red to yellow. A white background behind the beakers will help color definition. Place the electrodes of the conductivity tester into the glycine solution. Plug the conductivity tester in. The bulb will glow dimly. (Darken the room if necessary to see it.) Unplug the conductivity tester, and rinse the electrodes with a wash bottle into a 400 mL beaker. Repeat this procedure with the other acid solutions, noting that the conductivity and thus the brightness of the bulb increases with increasing acid strength.

Hazards

All of the acid solutions are corrosive, *HCl* especially so because it is a strong acid. Avoid contact. The conductivity tester is a considerable electrical shock hazard, especially with the solutions of electrolytes. Make sure that it is unplugged before handling it.

Disposal

Rinse each acid solution down the sink with a 100 fold excess of water.

17.9 Worksheets for Chapter 17

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Concentration by Percent Mass Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

The definition of percent mass concentration is the ratio of the mass of solute divided by the total mass of the solution and multiplied by 100 to convert to a percentage.

$$\text{percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

Example: What is the percent concentration by mass of a solution formed by dissolving 100. *grams* of ethanol, C_2H_5OH , in 100. *grams* of water?

Solution: percent by mass = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100 = \frac{100. \text{ g}}{200. \text{ g}} \times 100 = 50.0\%$

Example: If the density of a 10.0% by mass KNO_3 solution in water is 1.19 *g/mL*, how many grams of KNO_3 are present in 100. *mL* of the solution?

Solution: We can multiply the volume times the density to the mass of the 100. *mL* of solution and then take 10.0% of the mass of the solution to get the mass of the potassium nitrate.

$$\begin{aligned} \text{grams of solution} &= (100. \text{ mL})(1.19 \text{ g/mL}) = 119 \text{ grams} \\ \text{grams of } KNO_3 &= (0.10)(119 \text{ grams}) = 11.9 \text{ grams} \end{aligned}$$

Exercises

1. If 30.0 *grams* of $AgNO_3$ are dissolved in 275 *grams* of water, what is the concentration of the silver nitrate by mass percent?
2. How many grams of MgF_2 are present in 100.0 *g* of a 20.0% MgF_2 in water solution?
3. How many grams of water are present in the solution in question #2?
4. The density of a 30.0% by mass solution of $NaOH$ in water is 1.33 *g/mL*. How many grams of $NaOH$ are required to prepare 500. *mL* of this solution?
5. The density of pure water is 1.00 *g/mL*. What is the concentration by percent mass of a solution prepared by dissolving 85.0 *grams* of $NaOH$ in 750. *mL* of water?
6. A solution is prepared by dissolving 66.0 *grams* of acetone, C_3H_6O , in 146.0 *grams* of water. The density of the solution is 0.926 *g/mL*. What is the percent concentration of acetone by mass?
7. A 35.4% solution of H_3PO_4 in water has a density of 1.20 *g/mL*. How many grams of phosphoric acid are present in 300. *mL* of this solution?

Mole Fraction and Molality Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Mole Fraction

The definition of mole fraction is the ratio of the moles of solute divided by the total moles of the solution.

$$\text{mole fraction} = \frac{\text{moles of solute}}{\text{moles of solution}}$$

Example: What is the mole fraction of ethanol in a solution prepared by dissolving 100. g of ethanol, C_2H_5OH , in 100. g of water?

Solution:

$$\text{moles ethanol} = \frac{100. \text{ g}}{46.0 \text{ g/mol}} = 2.17 \text{ moles}$$

$$\text{moles water} = \frac{100. \text{ g}}{18.0 \text{ g/mol}} = 5.56 \text{ moles}$$

$$\text{mole fraction of ethanol} = \frac{2.17 \text{ mols}}{7.73 \text{ mols}} = 0.281$$

Molality

The definition of molality is the ratio of the moles of solute divided by the kilograms of solvent.

$$\text{molality} = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

Example: What is the molality of a solution prepared by dissolving 100. g of ethanol, C_2H_5OH , in 100. g of water?

$$\text{moles ethanol} = \frac{100. \text{ g}}{46.0 \text{ g/mol}} = 2.17 \text{ moles}$$

$$\text{molality of ethanol} = \frac{2.17 \text{ mols}}{0.100 \text{ kg}} = 21.7 \text{ m}$$

Example: A 35.4% solution of H_3PO_4 in water has a density of 1.20 g/mL. What is the mole fraction of H_3PO_4 in this solution and what is the molality?

Solution: We can choose a sample volume of this solution and get the mass of it by multiplying the volume times the density. Suppose we choose a 1.00 L sample.

$$\begin{aligned} \text{mass of solution} &= (1000. \text{ mL})(1.20 \text{ g/mL}) = 1200. \text{ grams} \\ \text{mass of } H_3PO_4 \text{ in the solution} &= (0.354)(1200. \text{ grams}) = 425 \text{ grams} \\ \text{mass of } H_2O &= 1200. \text{ grams} - 425 \text{ grams} = 775 \text{ grams} \\ \text{moles } H_3PO_4 &= \frac{425 \text{ g}}{98.0 \text{ g/mol}} = 4.34 \text{ moles} \\ \text{moles } H_2O &= \frac{775 \text{ g}}{18.0 \text{ g/mol}} = 43.1 \text{ moles} \\ \text{mole fraction of } H_3PO_4 &= \frac{4.34 \text{ mol}}{47.4 \text{ mol}} = 0.0916 \\ \text{molality} &= \frac{4.34 \text{ mol}}{0.775 \text{ kg}} = 5.60 \text{ m} \end{aligned}$$

Exercises

1. What is the mole fraction of MgF_2 in a solution that has 20.0 g of MgF_2 dissolved in 80.0 grams of water?
2. What is the molality of the solution in question 1?
3. The density of a 30.0% by mass solution of $NaOH$ in water is 1.33 g/mL. What is the mole fraction of $NaOH$ in this solution?
4. What is the molality of the solution in problem 3?
5. What is the molality of a solution prepared by dissolving 4.00 g of $NaCl$ in 100. g of water?
6. How many grams of beryllium chloride would you need to add to 125 g of water to make a 0.500 m solution?
7. What would be the mole fraction of $BeCl_2$ in the solution in problem 6?
8. A solution is prepared by dissolving 66.0 g of acetone, C_3H_6O , in 146.0 g of water. The density of the solution is 0.926 g/mL. What is the molality of this solution?
9. What is the mole fraction of acetone in the solution in problem 8?

Molarity Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

The definition of molarity is the ratio of the mols of solute divided by the volume of the solution.

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Example: What is the molarity of a solution prepared by dissolving 60.0 grams of $NaOH$ in sufficient water to produce 2.00 liters of solution?

Solution:

$$\begin{aligned} \text{moles } NaOH &= \frac{60.0 \text{ g}}{40.0 \text{ g/mol}} = 1.50 \text{ moles} \\ \text{molarity} &= \frac{1.50 \text{ mol}}{2.00 \text{ L}} = 0.750 \text{ M} \end{aligned}$$

Example: What volume of 0.750 M NaOH solution will contain 10.0 gram of NaOH?

$$\begin{aligned} \text{moles NaOH} &= \frac{10.0 \text{ g}}{40.0 \text{ g/mol}} = 0.250 \text{ moles} \\ \text{volume} &= \frac{\text{mol}}{M} = \frac{0.250 \text{ mol}}{0.750 \text{ mol/L}} = 0.333 \text{ L} \end{aligned}$$

Exercises

1. What is the molarity of a solution in which 4.50 g of NaNO_3 is dissolved in 265 mL of solution?
2. How many grams of ammonia, NH_3 are present in 5.0 L of 0.100 M solution?
3. How many milliliters of 0.200 M NaOH solution is necessary to contain 6.00 grams of NaOH?
4. How many liters of 0.500 M CaF_2 solution is required to contain 78.0 g of CaF_2 ?
5. What mass of ammonium phosphate is needed to make 100. mL of 0.500 M $(\text{NH}_4)_3\text{PO}_4$ solution?
6. What is the molarity of a solution prepared by dissolving 198 g of BaBr_2 in 2.00 liters of solution?
7. How many grams of glycerine, $\text{C}_3\text{H}_8\text{O}_3$, are needed to make 100. mL of 2.60 M solution?
8. A test tube contains 10.0 mL of 3.00 M CaCO_3 solution. How many grams of calcium carbonate are in the tube?

Dilution Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

The process of dilution involves increasing the amount of solvent in a solution without changing the amount of solute. For example, you could dilute 50. mL of 0.250 M HCl solution by placing the solution in a 100. mL graduated cylinder and adding water until the solution reached the 100. mL line in the graduate. The original solution contained 0.0125 moles of HCl before it was diluted and therefore, it also contains 0.0125 moles of HCl after the dilution. In the process of dilution, the amount of solute never changes. The amount of solvent, the total volume of the solution, and the concentration change but the amount of solute remains the same.

For a solution whose concentration is expressed in molarity, the moles of solute can be calculated by multiplying the volume in liters times the molarity.

$$\text{moles solute} = (\text{molarity})(\text{liters})$$

For the moles of solute in the original solution, $\text{moles}_{\text{initial}} = \text{molarity}_{\text{initial}} \times \text{liters}_{\text{initial}}$ or $\text{mols}_i = M_i \times V_i$. After the solution has been diluted, the moles in the final solution can be calculated with $\text{mols}_f = M_f \times V_f$. Since the mols do not change during dilution,

$$\text{mols}_i = \text{mols}_f \quad \text{and} \quad M_i \times V_i = M_f \times V_f.$$

In the dilution problems you will be given, for the most part, three of the four variables or ways to find three of the four variables and you will be asked to calculate the fourth variable.

Example: How many milliliters of 6.00 M NaOH solution are necessary to prepare 300. mL of 1.20 M NaOH solution?

Solution:

$$(M_i)(V_i) = (M_f)(V_f)$$

$$V_i = \frac{(M_f)(V_f)}{(M_i)} = \frac{(1.20 M)(0.300 L)}{(6.00 M)} = 0.0600 L = 60.0 mL$$

Exercises

200. mL of 3.00 M NaCl solution is diluted to a final volume of 500. mL. What is the concentration of the final solution?
100. mL of concentrated hydrochloric acid was diluted to 1.20 liters of 1.00 M solution. What was the concentration of the original concentrated solution?
- What volume of 6.00 M NaOH is needed to prepare 250. mL of 0.600 M NaOH?
- If 25.0 mL of 16.0 M HNO₃ is diluted to 500. mL, what is the final concentration?
- To what volume must you dilute 10.0 mL of 6.00 M H₂SO₄ to produce a solution that is 1.00 M H₂SO₄?
- Solution A is 5.00 mL of 12.0 M HCl. Solution B is prepared by diluting solution A to a new volume of 100. mL. Solution C is produced by taking 5.00 mL of solution B and diluting it to 100. mL. What is the molarity of solution C?

Colligative Properties: Solution Vapor Pressure Worksheet

Colligative properties are those properties of a solution that depend on the number of particles of solute present in the solution, and not on the chemistry nor the mass of the particles. That is, the chemical behavior and the molar masses of urea, (NH₂)₂CO, and glucose, C₆H₁₂O₆, are very different, but the colligative properties of a 1.0 M solution of urea will be exactly the same as the colligative properties of a 1.0 M solution of glucose.

The colligative properties of solutions include vapor pressure lowering, boiling point elevation, freezing point depression, and changes in osmotic pressure. The changes in these properties are dependent entirely on the concentration of particles of solute in the solution. It must be noted that ionic solutes dissociate when dissolved in water and therefore, add more particles to the solution than a substance that does not dissociate in water.

Vapor Pressure Lowering

The vapor pressure of a solution can be calculated from the individual vapor pressures of the components (solute and solvent) and the mole fractions of each component. Raoult's Law is an expression of the relationship.

$$\text{Vapor Pressure}_{\text{solution}} = (X_{\text{mol fraction solvent}})(\text{Vapor Pressure}_{\text{solvent}}) + (X_{\text{mol fraction solute}})(\text{Vapor Pressure}_{\text{solute}})$$

Example: What is the vapor pressure, at 25°C, of a solution produced by dissolving 50.0 of acetone, C₃H₆O, in 50.0 grams of water? The vapor pressure of pure acetone at 25°C is 230. mm of Hg and the vapor pressure of pure water at 25°C is 23.7 mm of Hg.

Solution: 50.0 g of acetone is 0.86 moles and 50.0 g of water is 2.78 moles.

Therefore, the mole fractions in this solution are 0.236 acetone and 0.764 water.

$$\begin{aligned}
 VP_{\text{SOLUTION}} &= (0.764)(23.7 \text{ mm of Hg}) + (0.236)(230. \text{ mm of Hg}) \\
 &= 18.1 \text{ mm of Hg} + 54.3 \text{ mm of Hg} \\
 &= 72.4 \text{ mm of Hg}
 \end{aligned}$$

In this case, the vapor pressure of the solution is higher than the vapor pressure of the solvent. That is due to the fact that acetone is a volatile (weak intermolecular forces of attraction) and therefore, evaporates readily. When we refer to **vapor pressure lowering**, we are referring to solutions in which the solute is **non-volatile**. When the solute is a solid, it can be generally be assumed that the solute is non-volatile.

Suppose we are making a solution of glucose in water. Glucose is a non-volatile, solid solute whose vapor pressure at room conditions is so small that it is negligible compared to the vapor pressure of water. When we substitute the values for a glucose solution into Raoult's Law, the second term (the one for the solute) is essentially zero because the vapor pressure of the pure solute is essentially zero.

$$\begin{aligned}
 \text{Vapor Pressure}_{\text{Solution}} &= (X_{\text{Mol fraction solvent}})(\text{Vapor Pressure}_{\text{Solvent}}) \\
 &\quad + \\
 &\quad (X_{\text{Mol fraction solute}})(\text{Vapor Pressure}_{\text{Solute}})
 \end{aligned}$$

If the second term in this equation, $(X_{\text{Mol fraction solute}})(\text{Vapor Pressure}_{\text{Solute}})$, becomes zero, then for a solution with a non-volatile solute, Raoult's Law becomes:

$$\text{Mol fraction solvent}(\text{Vapor Pressure}_{\text{Solvent}}) \text{ Vapor Pressure}_{\text{Solution}} = (X$$

This is Raoult's Law for solutions whose solute is a non-volatile.

$$VP_{\text{Solution}} = (X_{\text{Solvent}})(VP_{\text{Solvent}})$$

Example: What is the vapor pressure, at 25°C, of a solution produced by dissolving 50.0 of glucose, 25°C, in 50.0 grams of water? Glucose is non-volatile and the vapor pressure of pure water at 25°C is 23.7 mm of Hg.

Solution: 50.0 g of water is 2.78 moles and 50.0 g of glucose is 0.278 moles.

Therefore, the mole fraction of water in this solution is 0.909. We do not need to calculate the mole fraction of glucose because it isn't needed in Raoult's Law for non-volatile solutes.

$$VP_{\text{Solution}} = (X_{\text{Solvent}})(VP_{\text{Solvent}} = (0.909)(23.7 \text{ mm of Hg}) = 21.5 \text{ mm of Hg}$$

In this case, and in all cases of non-volatile solutes, the vapor pressure of the solution is less than the vapor pressure of the pure solvent.

Exercises

1. If 25.0 grams of sodium chloride is added to 500. grams of water at 25°C, what will be the vapor pressure of the resulting solution in kPa? The vapor pressure of pure water at 25°C is 3.17 kPa.
2. 125 g of the non-volatile solute glucose, $C_6H_{12}O_6$, is dissolved in 125 g of water at 25°C. IF the vapor pressure of water at 25°C is 23.7 Torr, what is the vapor pressure of the solution?

- Glycerin, $C_3H_8O_3$, is a non-volatile, non-electrolyte solute. If 53.6 g of glycerin is dissolved in 133.7 g of ethanol at $40.^\circ C$, C_2H_5OH , what is the vapor pressure of the solution? The vapor pressure of pure ethanol is 113 Torr at $40.^\circ C$.
- The vapor pressure of hexane, C_6H_{14} , at $60.0^\circ C$ is 573 Torr. The vapor pressure of benzene at the same temperature is 391 Torr. What will be the vapor pressure of a solution of 58.9 g of hexane with 44.0 g of benzene?

Colligative Properties: B.P. Elevation and M.P. Depression Worksheet

When a non-volatile, solid solute is added to a solvent, the boiling point of the solution will be higher than the boiling point of the solvent, and the melting point of the solution will be lower than the melting point of the solvent. The size of the boiling point elevation and the melting point depression are colligative properties, that is, they are dependent not on the chemistry of the solute but only on the number of solute particles present in the solution.

The formula used to calculate boiling point elevation is $\Delta T_b = imK_b$, where ΔT_b is the increase in the boiling point, m is the molality of the solute, K_b is the boiling point elevation constant, and i is the van't Hoff factor.

The boiling point elevation constant, K_b , is an experimentally determined constant for the solvent. Each solvent will have its own K_b and these values are determined in the laboratory and listed in reference tables. For example, the boiling point elevation constant for water is $0.512^\circ C/m$. As the molality of the solution increases, the boiling point of the solution increases by $0.512^\circ C$ for each increase of 1.00 in the molality.

The van't Hoff factor is the ratio between the actual concentration of **particles** produced when the substance is dissolved, and the concentration of the molecules dissolved. For most non-electrolytes dissolved in water, the van't Hoff factor is essentially 1. For most ionic compounds dissolved in water, the van't Hoff factor is equal to the number of discrete ions in a formula unit of the substance. For example, a glucose solution that is 1.00 molal will have a particle concentration that is also 1.00 molal because glucose molecules do not dissociate. A 1.00 molal sodium chloride solution, on the other hand, since it dissociates into two ions will have a **particle molality** of 2.00 m. The van't Hoff factor, i , is the number of ions that the molecule will dissociate into when dissolved. Sometimes, in concentrated solutions, an ionic substance does not dissociate 100% and therefore, the value of i will not be exactly equal to the apparent number of ions produced. In such cases, the value of i must also be determined experimentally. If you are not given an actual value for i in the problem, assume that i is the number of ions apparently produced per molecule. This is true in most dilute solutions.

The formula used to calculate melting point depression is $\Delta T_f = imK_f$, where ΔT_f is the decrease in the melting point, m is the molality of the solute, K_f is the melting point depression constant, and i is the van't Hoff factor.

The melting point depression constant, K_f , is an experimentally determined constant for the solvent. Each solvent will have its own K_f and these values are determined in the laboratory and listed in reference tables. For example, the freezing point depression constant for water is $1.86^\circ C/m$. As the molality of the solution increases, the melting point of the solution decreases by $1.86^\circ C$ for each increase of 1.00 in the molality.

Example: What is the boiling point of a 5.00 m glucose solution in water? Glucose is a non-volatile, non-electrolyte solute. K_b for water = $0.512^\circ C/m$.

Solution: $\Delta T_b = imK_b = (1)(5.00 m)(0.512^\circ C/m) = 2.56^\circ C$

Since the boiling point of the pure solvent was $100.00^\circ C$, the b.p. of the solution is $100.00^\circ C + 2.56^\circ C = 102.56^\circ C$

Example: What is the melting point of a 5.00 m NaCl solution in water? Sodium chloride is a non-volatile solute that dissociates 100% in water. K_f for water = $1.86^\circ C/m$.

Solution: $\Delta T_f = imK_f = (2)(5.00 m)(1.86^\circ C/m) = 18.6^\circ C$ (Since NaCl produces two ions in solution, $i = 2$.)

Since the melting point of the pure solvent was $0.00^\circ C$, the m.p. of the solution is $0.00^\circ C - 18.6^\circ C = -18.6^\circ C$

Exercises

1. What is the melting point of a solution produced by dissolving 45.0 g of $NaCl$ in 500. g of water. K_f for water = $1.86^\circ C/m$.
2. What is the boiling point of a solution produced by dissolving 45.0 g of $NaCl$ in 500. g of water. K_b for water = $0.512^\circ C/m$.
3. Which solution will have higher boiling point: a solution containing 105 g of $C_{12}H_{22}O_{11}$ in 500. g of water or a solution containing 35.0 g of $NaCl$ in 500. g of water?
4. When 25.0 g of an unknown, non-volatile, non-electrolyte is dissolved in 130. g of water, the boiling point of the solution is $102.5^\circ C$. What is the molar mass of the unknown?
5. How many grams of $C_2H_6O_2$ (anti-freeze, a non-electrolyte) must be added to 4,000. grams of water to reduce the melting point to $-40.^\circ C$?
6. The melting point constant for benzene is $4.90^\circ C/m$. The normal melting point of benzene is $5.50^\circ C$. What is the melting point of a solution of 9.30 g of $C_{12}H_{25}OH$ (a non-electrolyte) in 250. g of benzene?
7. Assuming 100% dissociation, what is the boiling point of a solution of 200. g of AlF_3 in 500. g of water?

Reactions Between Ions in Solution Worksheet**CK-12 Foundation Chemistry**

Name _____ Date _____

For the following five reactions (all reactants are in water solution):

- Write and balance the molecular equation indicating the state of each reactant and product.
- Write the total ionic equation.
- Identify the precipitate.
- Identify the spectator ions.
- Write the net ionic equation.

1. iron (III) chloride + sodium hydroxide

Balanced molecular equation _____

Total ionic equation _____

Precipitate = _____ Spectator ions = _____

Net ionic equation _____

2. barium chloride + silver nitrate

Balanced molecular equation _____

Total ionic equation _____

Precipitate = _____ Spectator ions = _____

Net ionic equation _____

3. magnesium sulfate + potassium phosphate

Balanced molecular equation _____

Total ionic equation _____

Precipitate = _____ Spectator ions = _____

Net ionic equation _____

4. copper (II) nitrate + calcium hydroxide

Balanced molecular equation _____

Total ionic equation _____

Precipitate = _____ Spectator ions = _____

Net ionic equation _____

5. sodium chromate + strontium nitrate

Balanced molecular equation _____

Total ionic equation _____

Precipitate = _____ Spectator ions = _____

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

17.10 Extra Reading for Chapter 17

Solutions

We are all familiar with the phenomenon of a hard crystalline solid, like table salt, when placed in water, apparently disappearing quite quickly. The crystalline structure breaks up and the particles enter into the water. Why does this process occur?

More questions arise when we think about dissolving and solutions. Table salt, for example, dissolves in water, but it will not dissolve in benzene. Camphor, on the other hand, dissolves easily in benzene, but not in water. While other substances like diamonds or graphite will not dissolve in any liquid. What controls whether a solid dissolves, and in what solvent it will dissolve?

Ion-Ion Attraction vs. Ion-Solvent Attraction

Consider the example of salt dissolving in water. Recall that table salt (sodium chloride) has a simple crystal structure in which positive sodium ions and negative chloride ions are organized in a crystal lattice. The electrical interactions between the positive and negative ions causes them to be strongly held at their locations in the crystal. To break up the crystal requires a large amount of energy, or else the attraction between the ions must be replaced by some other equal or greater attraction. This is the key to understanding what happens when the ions dissolve in water. The attraction between the ions in the solid is replaced by an attraction between the ions and the water molecules (or other solvent molecules). Water is a polar liquid. The oxygen end of the molecule has a partial negative charge, while the hydrogen end has a partial positive charge. When the sodium ion enters the liquid water, the water molecules cluster around it so that the partially negative ends of the water molecules are next to the positive sodium ions. Similarly the water molecules cluster around the chloride ions so that the partially positive ends of water molecules are directed toward the negative charge of the chloride ions. It is these ion-water attractions in the solution that replace the ion-ion attractions in the solid. The ions can break away from their oppositely charged neighbors in the crystal because they have found equal or stronger attractions in the solution. High solubility requires that the attraction between the atoms, ions, or molecules in the dissolving solid be replaced by equivalent or greater attractions between these particles and the molecules of the solvent. In many cases, it still requires an input of energy for a solid to dissolve in a solvent, but if the requirement is small enough, its effect can be outweighed by that of the increased disorder of the solution. The process is then driven by the increased entropy of the solution.

Like Dissolves Like

Solids like salt, which consist of ions, dissolve in polar solvents like water in which the solvent molecules have dipoles, because the electrical attractions between the ions and the solvent replace those between ions in the solid. In crystals made of non-polar molecules like camphor, the forces are different. The molecules are held in the crystal by weak London dispersion force attractions. Similar forces exist between the molecules in a solvent like benzene. So again, the interactions between the molecules in the solid can be replaced by those between the solute molecule and solvent molecules. Hence, camphor dissolves in benzene.

Much of what we have described so far relates to water as a solvent. Water is the most widespread liquid on the surface of earth. As we have seen, water is also an excellent solvent for polar solids. Polar solids include not only those that are made of ions like sodium chloride, but also those that are composed of polar molecules, like glucose and alcohol. In some cases, the attractions between solute molecules and the solvent water are even greater due to hydrogen bonding.

Molecules or parts of molecules can be classified as hydrophilic (water loving) or hydrophobic (water hating)

depending on whether they contain polar groups. An important group of molecules of this type are soaps and detergents, which have both hydrophilic and hydrophobic ends, leading to a range of useful and remarkable properties.

Solid Solutions

We do not normally think of solids like copper or silicon as being able to dissolve because there are no common liquids in which these solids will dissolve. Sometimes, the reaction between metals and strong acids are referred to as “dissolving”, but that involves a chemical reaction. In order for a solid to dissolve, the interaction between the atoms or molecules in the solid must be replaced by comparable ones in the solution. There are no substances which are liquids at normal temperatures, in which the atoms or molecules attract metal atoms strongly enough to dissolve them (assuming no chemical reaction).

It is possible, however, for these solids to dissolve in other solids forming solid solutions. Copper will dissolve in zinc to form an alloy (a solution of one metal in another) known as brass. Like most alloys, brass is crystalline, that is, it has a regular arrangement of metal atom locations with some of the locations occupied by copper and some by zinc atoms. Alloy formation is very common; other examples are pewter (tin and zinc) and bronze (iron and copper). Dissolving a small amount of one metal in another can also have significant effects on physical and chemical properties. Stainless steel is essentially iron into which a small amount of chromium is dissolved. Stainless steel is significantly different from iron in terms of the rate at which it corrodes. Alloying iron with copper in bronze results in a much tougher, less brittle material. Solid solutions are widespread. Silicon, as we have seen, will not dissolve in any common liquid; but it will dissolve in germanium (a solid with the same crystal structure) in much the same way that copper dissolves in zinc. These solids further illustrate the point that dissolving small amounts of one solid substance in another is a vitally important way of altering the properties of materials. One that is used on an enormous scale in contemporary technology. The classic example is the semiconductor silicon. Dissolving tiny amounts (less than one part per million) of phosphorus in silicon has a significant effect on its ability to conduct electricity (making the material that is known as an ‘n-type’ semiconductor). Similar amounts of arsenic dissolved in silicon have equally large effects, but result in different electrical characteristics (the material becomes a ‘p-type’ semiconductor). Putting the two types of material together creates the famous p/n junction which allows electricity to flow only in one direction, a vital feature of some electrical circuits. Silicon with tiny quantities of deliberately introduced impurities is therefore the material basis of the technology on which the modern electronics revolution is based.

Properties of Solutions

Definitions

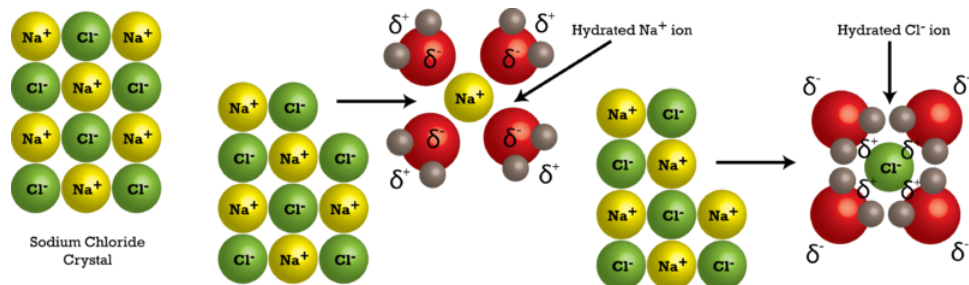
solution: homogeneous mixture of a solute dissolved in a solvent

solute: component present in smaller amount

solvent: component present in greatest amount

The Solution Process

As a solute crystal is dropped into a solvent, the solvent molecules begin to attack and pull apart the solute ions or molecules. Solvent molecules surround the solute molecules in a process called hydration, forming a solvent cage around the solute particles as the solute dissolves in the solvent.



The illustration above shows the hydration process of ions. A similar process occurs when dissolving polar molecules that do not form ions. Each polar solute molecule is attached to one or more polar solvent molecules.

Energy Changes During Solution Formation

Three types of interactions must be considered for solution formation.

- solvent-solvent attraction
- solute-solute attraction
- solvent-solute attraction

In order for a solute to dissolve, the solvent-solute attractions must be at least equal to the solvent-solvent and solute-solute attractions.

So why don't all solids dissolve in liquids?

If solvent-solute interaction can't compete with solute-solute and solvent-solvent interactions, they remain separated. When solute-solute or solvent-solvent interactions are stronger than solute-solvent interactions, solute and solvent stay separated. When solute-solvent interactions are as strong as solute-solute and solvent-solvent interactions, solute and solvent mix.

Saturated Solutions and Solubility

Definitions

solubility: Maximum amount of solute dissolved in solvent at specific temp.

unsaturated: contains less than the maximum amount of solute that a solvent can hold at specific temperature.

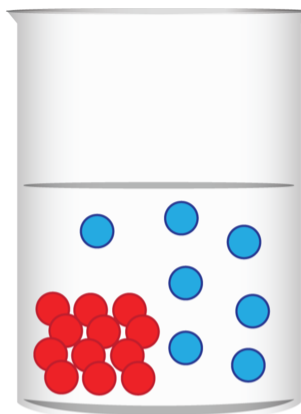
saturated: contains the maximum amount of solute that a solvent can hold at specific temperature.

supersaturated: contains more than the maximum amount of solute that a solvent can hold at specific temperature.

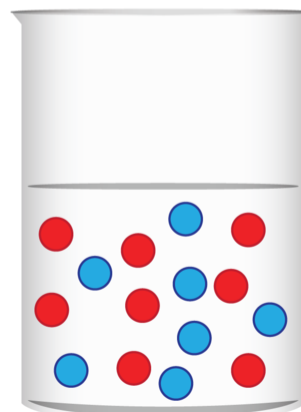
How do supersaturated solutions form?

At higher temperatures, solvents can hold more solute than at lower temperatures. If a given amount of solute is dissolved in a solvent at a higher temperature, then allowed to cool without being disturbed, the solute will remain in solution. The solution is unstable, though, and the solute will crystallize rapidly if disturbed.

Solute-solute and solvent-solvent attractions dominate = does not dissolve



Solute-solvent attractions dominate = does dissolve



Factors Affecting Solubility

A general rule referred to as “Like Dissolves Like” applies to solution formation. Ionic compounds and polar molecules dissolve in polar solvents but do not dissolve in non-polar solvents. Non-polar solutes dissolve in non-polar solvents but do not dissolve in polar solvents. The reason was referred to earlier in the section on Energy Changes During Solution Formation. When polar molecules are introduced to non-polar solvents, the solute-solute attractions are greater than the solute-solvent attractions and the solute does not dissolve. When non-polar solutes are introduced to polar solvents, the solvent-solvent attractions are greater than the solute-solvent attractions and the solute does not dissolve.

Liquid-Liquid Solutions

Polar molecules will mix (be miscible with) other polar molecules.

Non-polar molecules will mix (be miscible with) other non-polar molecules.

Polar molecules will not mix (be immiscible with) non-polar molecules.

Solid-Liquid Solutions

Ionic and polar molecular compounds dissolve in polar solvents. (This is a general rule but there are some ionic compounds that do not dissolve in polar solvents . . . for the final answer, you must check the solubility rules.)

Non-polar molecular compounds dissolve in non-polar solvent.

Molecules containing polar bonds and non-symmetrical shapes will be polar. Molecules without polar bonds or with polar bonds but in symmetrical shape will be non-polar.

Some solids do not dissolve in any solvent

- Network covalent solids (eg. graphite, quartz) do not dissolve in any solvent.
- Metals do not dissolve in any solvent. (They may react but they don’t dissolve.)

For example, of the following substances, NaCl , CCl_4 , NH_3 , and $\text{C}_{\text{diamond}}$, will be soluble in water and which will be soluble in hexane (a non-polar liquid)?

NaCl is ionic and will therefore, dissolve in water but not in hexane. CCl_4 is non-polar and will therefore, dissolve in hexane but not in water. NH_3 is polar and will therefore, dissolve in water but not in hexane. $\text{C}_{\text{diamond}}$ is a network covalent solid and will not dissolve in water or hexane.

Gas-Liquid Solutions

Gas solubility and Pressure Effects:

Henry's Law: The solubility of gas is proportional to partial pressure of gas above liquid.

The solubility of a gas increases when the partial pressure of the gas above the liquid increases and solubility decreases when the partial pressure of the gas above the liquid decreases. This is because a higher partial pressure of gas above the liquid means that a larger number of gas molecules are in contact with the liquid surface and therefore, a larger number of gas molecules will enter the liquid phase (dissolve).

Example

In a container of carbonated beverage that is sealed, the carbonation will remain indefinitely. When the container is opened, the gas above the liquid escapes, lowering the partial pressure of gas above the liquid, and the carbonation (carbon dioxide gas) will gradually decrease as the carbon dioxide gas comes out of solution.

Gas solubility and Temperature:

As the temperature increases, the solubility of a gas in a liquid decreases (in most cases). Gases are more soluble in liquids at lower temperatures.

Example

When a pan of water is heated on the stove, gas bubbles form on the bottom and sides of the pan long before the water is nearing the boiling point. The bubbles are not bubbles of water vapor but rather are bubbles of air that have come out of solution as the water warmed.

Solid solubility and Temperature:

As the temperature increases, the solubility of a solid in a liquid increases (in most cases).

Example

More sugar will dissolve in hot tea than in iced tea.

Methods of Expressing Concentration

Mass percent of solute = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$

5.0 grams of salt dissolved in 495 grams of water is a 1% solution of salt.

(ppm), parts per million = $\frac{\text{milligrams of solute}}{\text{kilograms of solvent}}$

5.0 milligrams of salt dissolved in 5.0 kg of water is 1 ppm.

(ppb), parts per billion = $\frac{\text{micrograms of solute}(\mu\text{g})}{\text{kilograms of solvent}}$

5.0 milligrams (5000 micrograms) dissolved in 5.0 kg of water is 1000 ppb.

Mole Fraction (X) (has no units since ratio of two similar quantities) = $\frac{\text{moles of solute}}{\text{total moles in solution}}$

200. grams of CaBr_2 (molar mass = 200. g/mol) dissolved in 500. grams of water (molar mass = 18.0 g/mol) = 1.00 mol of CaBr_2 dissolved in 27.8 mols of water so the mole fraction, $X_{\text{CaBr}_2} = \frac{1.00 \text{ mol}}{28.8 \text{ mols}} = 0.0347$ and $X_{\text{H}_2\text{O}} = \frac{27.8 \text{ mols}}{28.8 \text{ mols}} = 0.965$.

$$\text{Molarity (M)} = \frac{\text{mols of solute}}{\text{liters of solution}}$$

$$200. \text{ grams of } \text{CaBr}_2 \text{ dissolved in } 500. \text{ mL of solution} = \frac{1.00 \text{ mol}}{0.500 \text{ L}} = 2.00 \text{ M}$$

$$\text{Molality (m)} = \frac{\text{mols of solute}}{\text{kg of solvent}}$$

$$200. \text{ grams of } \text{CaBr}_2 \text{ dissolved in } 500. \text{ grams of water} = \frac{1.00 \text{ mol}}{0.500 \text{ kg}} = 2.00 \text{ m}$$

For molarity dilution problems, use $M_1V_1 = M_2V_2$ where M = molarity and V = volume.

All concentration expressions are independent of temperature except molarity. Since the volume of solution changes with temperature, so does molarity.

Examples of converting from one concentration unit to another.

1. Calculate the molarity of a sodium chloride solution, which is 5.0% NaCl (molar mass = 58.5 g/mol). The density of the solution is 1.03 g/mL .

Imagine a 100. mL sample of solution is taken. The mass of the total solution is 103 grams . 5.15 grams of the solution is NaCl and 97.85 grams of the solution is water. The 5.15 g of NaCl is 0.0880 mol . Therefore, the molarity will be $\frac{0.0880 \text{ mol}}{0.100 \text{ L}} = 0.880 \text{ M}$.

2. Calculate the molarity of a 0.500 m glucose (molar mass = 180. g/mol) solution if the density of the solution is 1.16 g/mL .

Imagine a sample that contains 1000. g of water. This sample will also contain 0.500 mol of glucose which has a mass of 90.0 grams . Therefore, the total mass of the sample is 1090. grams . Dividing this total mass by the density yields the volume of the sample, 940. mL . Therefore, the molarity will be $\frac{0.500 \text{ mol}}{0.940 \text{ L}} = 0.532 \text{ M}$.

Colligative Properties

colligative properties: properties that depend on the number of solute particles in solution and not on the nature of the solute particles

non-electrolytes: exist as molecules in solution (do not dissociate into ions)

electrolytes: exist as ions in solution

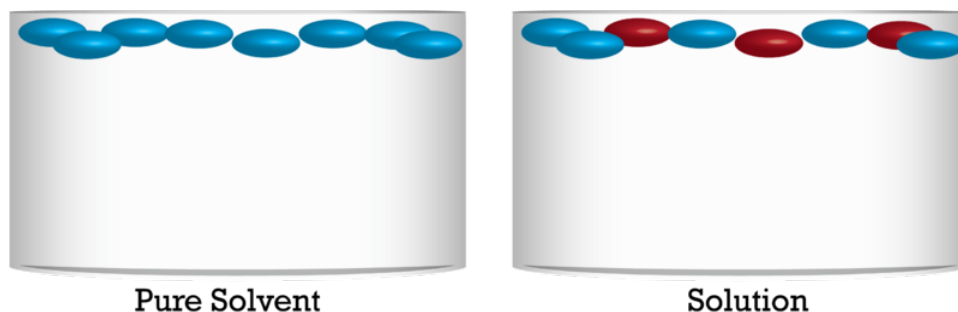
Lowering the Vapor Pressure (Non-electrolytes)

vapor pressure: pressure exerted by vapor in equilibrium with its liquid or solid

A substance that has very low vapor pressure is nonvolatile, whereas one that exhibits a vapor pressure is volatile.

Oil is considered nonvolatile while gasoline is volatile.

Adding a solute lowers the concentration of solvent molecules in liquid phase since solute particles on the surface of the solution block solvent molecules from evaporating.



Adding a solute to a solvent lowers the vapor pressure of the solvent. There are two suggested explanations for why the addition of a solute lowers the vapor pressure of a solution. Since they seem equally valid, both will be presented here. Remember that only the molecules on the surface of a liquid are able to evaporate.

In a pure solvent, all the molecules at the surface are solvent molecules. Therefore, the entire surface area is available for evaporation and the forces to be overcome are the attractive forces between the solvent molecules. One of the explanations says that in a solution, some of the surface molecules are solute molecules and since these solute molecules take up some of the surface area, less surface area is available for evaporation. Therefore, the rate of evaporation of the solvent will be lower and so the vapor pressure will be lower at the same temperature. The other explanation says that the attractive forces between the solvent molecules and the solute molecules are greater than the attractive forces between solvent molecules and therefore, the solvent molecules will not evaporate at as high a rate. Once again vapor pressure will be lowered. Both explanations start with the same premises and end with the same result so there doesn't seem to be a reason to choose between them.

Boiling-Point Elevation and Freezing Point Depression (Non-electrolytes)

Boiling-Point Elevation:

A liquid boils when its vapor pressure equals the surrounding (ambient) pressure. For example, pure water has a vapor pressure of 760 mm of Hg at 100°C. Therefore, when liquid water is raised to 100°C, it boils. If a non-volatile solute is added to water, the vapor pressure of the solution is lower than the vapor pressure of the pure solvent. Such a solution will have a vapor pressure less than 760 mm of Hg at 100°C and therefore, will not boil at this temperature. In order for the vapor pressure of the solution to exhibit a vapor pressure of 760 mm of Hg, the temperature must be raised higher than 100°C. Therefore, the boiling point of the solution is greater than the boiling point of the pure solvent.

The increased boiling point is determined as follows: $T_b = T_{bp\text{water}} + \Delta T_b$, where T_b = boiling point of solution, $T_{bp\text{water}}$ = b. p. of pure solvent, ΔT_b = change in b.p.

ΔT_b is calculated using $\Delta T_b = K_b m$, where m = molal concentration of solute and K_b = molal boiling point constant.

Example: Calculate the boiling point of a solution containing 1.25 mol of glucose in 0.250 kg of water using $K_b = 0.52^\circ\text{C}/m$.

$$\text{molality} = \frac{\text{mols solute}}{\text{kg of solvent}} = \frac{1.25 \text{ mols}}{0.250 \text{ kg}} = 5.0 \text{ m}$$

$$\Delta T_b = (5.0 \text{ m})(0.52^\circ\text{C}/m) = 2.6^\circ\text{C}$$

$$\text{B.P. of solution} = 100.^\circ\text{C} + 2.6^\circ\text{C} = 102.6^\circ\text{C}$$

Freezing-Point Depression:

Two things happen when ice and water are placed in contact: molecules on the surface of the ice escape into the water (melting), and molecules of water are captured on the surface of the ice (freezing). When the rate of freezing is the same as the rate of melting, the amount of ice and the amount of water won't change on average. The ice and water are said to be in dynamic equilibrium with each other. The balance between freezing and melting can be

maintained at 0°C , the melting point of water, unless conditions change in a way that favors one of the processes over the other.

The balance between freezing and melting processes can easily be upset. If the ice/water mixture is cooled, the molecules move slower. The slower-moving molecules are more easily captured by the ice, and freezing occurs at a greater rate than melting. Conversely, heating the mixture makes the molecules move faster on average, and melting is favored. Adding salt (or other non-volatile solute) to the system will also disrupt the equilibrium. Consider replacing some of the water molecules with molecules of some other substance. The foreign molecules dissolve in the water, but do not pack easily into the array of molecules in the solid. The total number of water molecules captured by the ice per second goes down, so the rate of freezing goes down. The rate of melting is unchanged by the presence of the foreign material, so melting occurs faster than freezing.

That's why salt melts ice. To re-establish equilibrium, you must cool the ice-saltwater mixture to below the usual melting point of water. For example, the freezing point of a 1 M NaCl solution is roughly -3.4°C . Solutions will always have such a freezing point depression. The higher the concentration of salt, the greater the freezing point depression.

The new freezing point is determined as follows: $T_{fp\text{ solution}} = T_{f\text{ solvent}} - \Delta T_f$ where $T_{fp\text{ solution}}$ = freezing point of solution, $T_{f\text{ solvent}}$ = freezing point of pure solvent, and ΔT_f = freezing point depression.

ΔT_f is calculated using $\Delta T_f = K_f m$, where m = molal concentration of solute and K_f = molal freezing point constant.

Example: Calculate the freezing point of a solution containing 1.25 mol of glucose in 0.250 kg of water using $K_f = 1.86^{\circ}\text{C}/m$.

$$\begin{aligned} \text{Molality} &= \frac{1.25\text{ mols}}{0.250\text{ kg}} = 5.00\text{ m} \\ \Delta T_f &= K_f m = (1.86^{\circ}\text{C}/m)(5.00\text{ m}) = 9.3^{\circ}\text{C} \\ T_{fp\text{ solution}} &= 0.0^{\circ}\text{C} - 9.3^{\circ}\text{C} = -9.3^{\circ}\text{C} \end{aligned}$$

Osmotic Pressure (Non-electrolytes)

semipermeable membrane: allows solvent molecules to pass through but blocks the passage of solute molecules

osmosis: net movement of solvent molecules through semipermeable membrane from pure solvent or more dilute solution to more concentrated solution

osmotic pressure (π): pressure required to stop osmosis

isotonic: when two solutions have equal osmotic pressure

hypertonic: the more concentrated of two solutions that are not isotonic

hypotonic: the less concentrated solution of two solutions that are not isotonic

We can calculate osmotic pressure (π) at a given temperature: $\pi = MRT$, where M = molarity of solute, $R = 0.0821\text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$, and T = Kelvin temperature.

Example: A 0.125 M sample of seawater was taken at 25°C . Calculate the osmotic pressure of the seawater sample.

$$\pi = MRT = (0.125\text{ mol/L})(0.0821\text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(298\text{ K}) = 3.06\text{ atm}$$

Colligative Properties of Electrolyte Solutions

For colligative properties, electrolyte solutions have another factor to consider beyond those of non-electrolyte solutions. Colligative properties are controlled by the molality of the number of particles in solution. A 1.0 m

solution of non-electrolyte will contain 1.0 mol of particles because the molecules do not dissociate in water solution. Electrolytes, on the other hand, do dissociate in water solution. In the case of an electrolyte like $NaCl$, a 1.0 m solution will contain 2.0 m solution of ions. If the electrolyte were $CaCl_2$, a 1.0 m solution would contain approximately 3.0 m solution of ions. In order to calculate colligative properties for electrolyte solutions, another factor (called the van't Hoff factor) is included in the equations.

In dilute solutions, the van't Hoff factor is equal to the number of ions that can be formed from each molecule. In concentrations solutions, sometimes the molecules do not all dissociate 100% and so the van't Hoff factor will be slightly less than the number of ions that can be formed from each molecules. Sometimes the values are referred as the theoretical and actual van't Hoff factors. In the absence of any information indicating the van't Hoff factor is less than the number of ions that can be formed from each molecule, assume full dissociation.

van't Hoff factor = i = number of moles of ions per mole of electrolyte

i = actual number of particles in solution after dissociation

Example: What is the van't Hoff factor, i for each of the following solutions?

a. Na_3PO_4 : $i = 4$

b. KOH : $i = 2$

c. $Al(NO_3)_3$: $i = 4$

d. H_2SO_4 : $i = 3$

For electrolyte solutions, the van't Hoff factor is included in the equations.

$$\Delta T_b = iK_b m$$

$$\Delta T_f = iK_f m$$

$$\pi = iMRT$$

Calculate the freezing point of a 0.100 m solution of CaS . K_f for water = $1.86^\circ C/m$.

$$\begin{aligned}\Delta T_f &= iK_f m = (2)(1.86^\circ C/m)(0.100 m) = 0.37^\circ C \\ T_{fp} \text{ for the solution} &= 0^\circ C - 0.37^\circ C = -0.37^\circ C\end{aligned}$$

Calculate the boiling point of 0.25 m solutions of $CaCl_2$. $K_b = 0.52^\circ C/m$.

$$\begin{aligned}\Delta T_b &= iK_b m = (3)(0.52^\circ C/m)(0.25 m) = 0.39^\circ C \\ T_{bp} &= 100.00^\circ C + 0.39^\circ C = 100.39^\circ C\end{aligned}$$

If the osmotic pressure of 0.010 M KI solution at $25^\circ C$ is 0.465 atm, calculate the van't Hoff factor for KI at this concentration.

$$i = \frac{\pi}{MRT} = \frac{0.465 \text{ atm}}{(0.010 \text{ mol/L})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(298 \text{ K})} = 1.9$$

Determination of Molar Mass from Freezing Point Depression

Example: Ethylene glycol (EG) is a common automobile antifreeze. It is water soluble and a non-electrolyte. Calculate the molar mass of EG if 651 g of EG was dissolved in 2505 g of water, and the freezing point for the solution is $-7.79^\circ C$. $K_f = 1.86^\circ C/m$.

$$m = \frac{\Delta T_f}{K_f} = \frac{7.79^\circ\text{C}}{1.86^\circ\text{C}/m} = 4.19 \text{ mol/kg}$$

$$\text{mols} = (4.19 \text{ mol/kg})(2.505 \text{ kg}) = 10.5 \text{ mols}$$

$$\text{molar mass} = \frac{\text{grams}}{\text{mols}} = \frac{651 \text{ g}}{10.5 \text{ mols}} = 62.0 \text{ g/mol}$$

17.11 Assessment for Chapter 17

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 18 Chemical Kinetics C-SE-TE

Chapter Outline

- 18.1 RATE OF REACTIONS
- 18.2 COLLISION THEORY
- 18.3 POTENTIAL ENERGY DIAGRAMS
- 18.4 FACTORS THAT AFFECT REACTION RATES
- 18.5 REACTION MECHANISM
- 18.6 WORKSHEETS FOR CHAPTER 18
- 18.7 EXTRA READING FOR CHAPTER 18
- 18.8 ASSESSMENT FOR CHAPTER 18

Lessons and Number of Activities for Lessons

TABLE 18.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Rate of Reactions	0	0	0	0
2. Collision Theory	0	0	0	1
3. Potential Energy Diagrams	0	0	1	0
4. Factors That Affect Reaction Rates	2	2	0	1
5. Reaction Mechanism	0	0	0	0

18.1 Rate of Reactions

Student Behavioral Objectives

The student will:

- define chemical kinetics and rates of reactions.
- write the rate expression and the units for the rate expression.
- define instantaneous rate.
- calculate instantaneous rate using a tangent line.

Timing, Standards, Activities

TABLE 18.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Rate of Reactions</i>	1.5	8a

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Rate of Reactions (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

18.2 Collision Theory

Student Behavioral Objectives

The student will:

- define the collision theory.
 - describe the conditions for successful collisions.
 - explain how the kinetic molecular theory applies to the collision theory.
 - describe the rate in terms of the conditions of successful collisions.
-

Timing, Standards, Activities

TABLE 18.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Collision Theory</i>	1.0	8b, 8d

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. Chemical Kinetics
-

Answers for Collision Theory (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

18.3 Potential Energy Diagrams

Student Behavioral Objectives

The student will:

- define enthalpy, activation energy, activated complex.
- describe and draw the difference between endothermic and exothermic potential energy diagrams.
- draw and label the parts of a potential energy diagram.

Timing, Standards, Activities

TABLE 18.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Potential Energy Diagrams</i>	1.0	7b, 8d

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Potential Energy Diagrams Worksheet

Extra Readings

1. None

Answers for Potential Energy Diagrams (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

18.4 Factors That Affect Reaction Rates

Student Behavioral Objectives

The student will:

- state how the rate of reaction changes as a function of temperature.
- explain how increased temperature increases the number of particles that can overcome the energy barrier.
- describe the effect of increasing the concentration on the rate of a reaction.
- indicate which reactants in a multi-step process can affect the rate of a reaction.
- calculate, using experimental data, the relationship between the ratio of the change in concentration of reactants, and ratio of the change in rate.
- describe the surface area to volume ratio.
- describe the effect of surface area on reaction rate.
- describe how the change in the surface area affects the collision frequency.
- describe real world examples of the effect of surface area on reaction rate.
- define a catalyst.
- identify a catalyst in a single equation.
- identify a catalyst in a multi-step process.
- describe how a catalyst affects the potential energy diagram.
- explain how a catalyst affects the rate of the reaction.
- explain how a catalyst affects our everyday lives, particularly with vitamins.

Timing, Standards, Activities

TABLE 18.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Factors That Affect Reaction Rates</i>	2.0	8b, 8c, 8d

Activities for Lesson 4

Laboratory Activities

1. Chemical Kinetics: Factors Affecting Reaction Rates
2. Chemical Kinetics: The Iodine Clock Reaction

Demonstrations

1. Catalytic Oxidation of Alcohol with Copper
2. Hydrogen Peroxide Decomposition Catalyzed by Manganese Dioxide

Worksheets

1. None

Extra Readings

1. Catalytic Converters

Answers for Factors That Affect Reaction Rates (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

18.5 Reaction Mechanism

Student Behavioral Objectives

The student will:

- define reaction mechanisms.
 - identify the rate-determining step.
 - draw a potential energy diagram for a multi-step process.
-

Timing, Standards, Activities

TABLE 18.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Reaction Mechanism</i>	1.0	8d

Activities for Lesson 5

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Reaction Mechanism (L5) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 18

Laboratory Activities for Chapter 18

Chemical Kinetics: Factors Affecting Reaction Rate

Students qualitatively compare reaction rates in regards to changes in: 1) nature of reactants - particle size, 2) temperature, 3) concentration, and 4) the presence of a catalyst. The lab material contains safety measures and a list of necessary materials.

Available at: http://www.saskschools.ca/curr_content/chem30_05/2_kinetics/labs/labs.htm

Chemical Kinetics: The Iodine Clock Reaction

This is the classic clock reaction. Students quantitatively measure reaction rates versus changes in concentration and changes in temperature. The lab material contains safety measures and a materials list.

Available at: http://www.saskschools.ca/curr_content/chem30_05/2_kinetics/labs/labs.htm

Demonstrations for Chapter 18

Catalytic Oxidation of Alcohol with Copper Demonstration

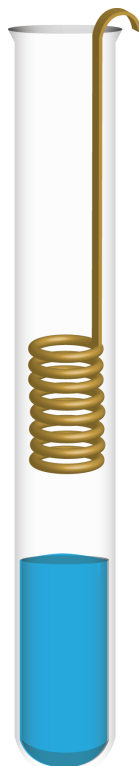
Brief description of demonstration:

A copper coil is made from #18 gauge copper wire, and heated in a Bunsen burner. The coil is then immersed in alcohol vapor. The coil immediately shows a brilliant polished copper color, and then displays a rainbow of colors ranging from blue to green to yellow to red. The reaction ceases when the wire ceases to change color. A sweet but pungent and sour odor is produced where there was none before.

Materials:

- 18 gauge copper wire, 30 *cm*
- 2 test tubes, 18 × 150 *mm*
- test tube rack
- crucible tongs
- 5 *mL* methanol, CH_3OH
- 5 *mL* ethanol, C_2H_5OH
- 10 *mL* graduated cylinder
- Bunsen burner

Procedure:



Bend the copper wire into a pigtailed spiral by bending it around a pencil. Wind between 5 and 8 turns. Pour 5 mL of water into a test tube. Insert the copper spiral into the tube so the spiral is inside of the tube. Lower it until it is about 1 cm above the surface of the water inside. Bend the remaining copper over the lip of the test tube so the copper wire stays in place when the test tube is standing. Remove the wire, and empty the contents of the test tube. Add 5 mL of methanol to the test tube, and set it aside in the stand. Grasp the copper spiral by the hooked end with crucible tongs, and lower the spiral into a hot Bunsen burner flame until it glows cherry red, but no more than that. Quickly remove the copper spiral, and lower it into the test tube. Hook the end of the copper coil around the lip of the test tube. Observe the changes in the copper wire over time. After it has stopped reacting have the students notice the odor of the vapors produced by wafting the vapors in the test tube towards their nose. Repeat the same procedure with the ethanol, using the same wire.

Hazards

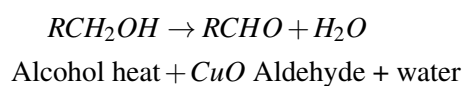
Methanol and ethanol are both poisons. Do not ingest them. Methanal and ethanal (the products of the reactions) can be irritating, so do not breathe much of the vapor. Wafting has not been a problem in the past. The copper coil gets very hot and can cause a burn.

Disposal

Pour the remaining alcohols down the sink and dilute with an excess of water. Discard the used copper wire into the trash. Make sure it has cooled first.

Discussion

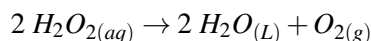
Heating copper in air causes copper(II) oxide to form on the surface of the wire. When this hot copper(II) oxide encounters alcohol vapor, the oxygen in the CuO reacts with two of the hydrogens of the alcohol, producing water and an aldehyde, which gives the strong sickly-sweet odor. The copper is regenerated, showing catalysis.



Hydrogen Peroxide Decomposition Catalyzed by Manganese Dioxide

Description of Demonstration

Hydrogen peroxide decomposes extremely slowly at room conditions to produce water and oxygen gas.



The reaction rate of this decomposition can be greatly increased by the use of a catalyst for the reaction. There are many catalysts for this reaction. Because there are many catalysts for this reaction, it is vital that the beaker used for the demonstration be clean. Any small contaminants may act as a catalyst and ruin the demo. (As a side demonstration, you may wish to show the effect on the reaction of beef liver, blood, ear wax, active yeast, etc. All of these materials will act as catalysts for this reaction. Just put a small amount of the diluted hydrogen peroxide in separate small beakers to which you can add various materials.)

30% H_2O_2 is available from most science supply companies. It comes in a dark, vented container and should be stored in oxidizer section of chemical storeroom (away from reducing agents, organics, and strong bases). If you have a chemical refrigerator, that is ideal.

You should wear protective gloves for this demonstration (latex or nitrile) and goggles. (Bursting bubbles in the solution can spray tiny droplets and will produce white speckled dots on your hands.) A solution of 10% H_2O_2 is sufficient for this demonstration – so you should dilute the 30% H_2O_2 with distilled water (2 parts water to 1 part H_2O_2).

Pour about 250 mL of 10% hydrogen peroxide into a very clean 600 mL Pyrex beaker. Show the students that no visible reaction is occurring. You should point out that the reaction is occurring but at a rate too slow to see.

Place a lump of MnO_2 (use tongs) into the solution. You can use powdered MnO_2 but the solution becomes gray and foamy, and the bubble formation is not as clear to students as it is if you use a lump of MnO_2 . MnO_2 , in the form of the mineral pyrolusite, is available cheaply from D J Mineral Company, P.O. Box 761, Butte, Montana 59703-0761, Phone: 406-782-7339, Fax: 406-494-2455 (Cost is about \$2/ lb; 5 lb is more than enough for a lifetime).

With the addition of the lump of catalyst, the formation of oxygen gas becomes rapid. If you want to go to a lot of trouble, you could determine the mass of the lump of catalyst before placing it in the solution, and then wash and dry it after you remove and re-weigh to show that the mass of the catalyst is unchanged. Most students will believe you if you simply tell them this is true.

It is important that the students understand that the MnO_2 is NOT a reactant. The reaction rate is not increasing because you have increased the concentration of a reactant but because a new reaction path has been opened that has a lower activation energy requirement. (If the lump of catalyst is removed, the rapid reaction ceases.) The net reaction is the same but the presence of the catalyst allows the reaction to occur at a much greater rate.

Light a wooden splint and allow it to burn for 20 seconds or so and then blow it out. The end of the splint should still have some part of it glowing red. Insert the glowing splint into the air above the liquid in the beaker. The splint will burst back into flame. This is the test for oxygen gas. (The test itself is based on reaction kinetics. The gas above the liquid in the beaker is somewhere near 100% oxygen gas, whereas the air in the room is about 20% oxygen gas. The 20% oxygen gas causes the splint to react fast enough to glow, and the 100% oxygen gas causes the splint to react fast enough to burst into flame. (Increasing the concentration of a reactant by a factor of 5 causes the reaction rate to increase by a factor of 5.)

You can repeat blowing out the splint and re-igniting it as long as it is entertaining.

This reaction is exothermic and the solution will heat up as you run the demo. If the solution begins to steam, it is advisable to add a little distilled water to cool it off and dilute the hydrogen peroxide so the reaction slows down.



Apparatus and Materials

- 250 mL or so of 10% H_2O_2
- 600 mL beaker (Pyrex)
- Lump of MnO_2 or MnO_2 powder
- Tongs (for adding and retrieving MnO_2 lump)
- Distilled water
- Wooden splint (perhaps several)
- Bunsen burner or candle or matches to ignite wooden splint

Hazards

10% hydrogen peroxide is caustic to the skin and eyes. Handle with care. In case of skin contact, rinse the affected area well with water. Store unused hydrogen peroxide in a freezer reserved for laboratory chemicals only.

Disposal

Retrieve the MnO_2 lump for use in following years. The solution can be poured down the sink following an excess of water.

Discussion

You should try out the demonstration before you do it in class. After observing the amount of bubbling and the ease or difficulty of getting the wooden splint to burst back into flame, you can adjust the dilution of the hydrogen peroxide to more or less than 10% to get the concentration you wish to use in class.

18.6 Worksheets for Chapter 18

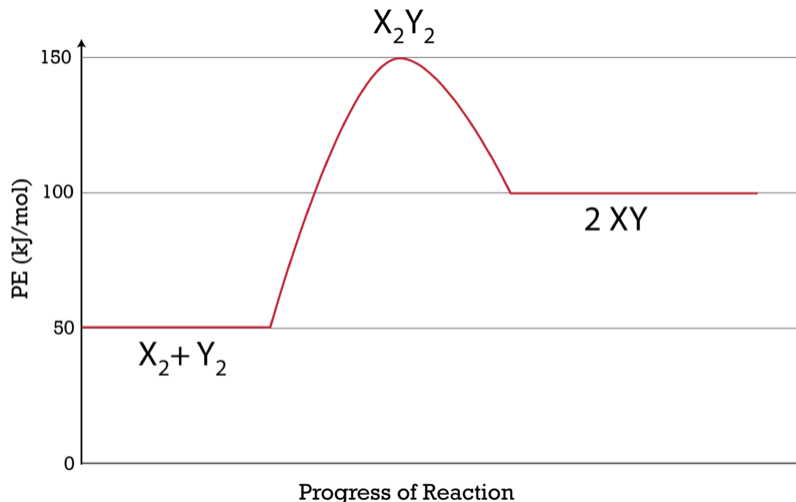
Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Potential Energy Diagrams Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Use the following Potential Energy Diagram to answer questions 1 - 12.



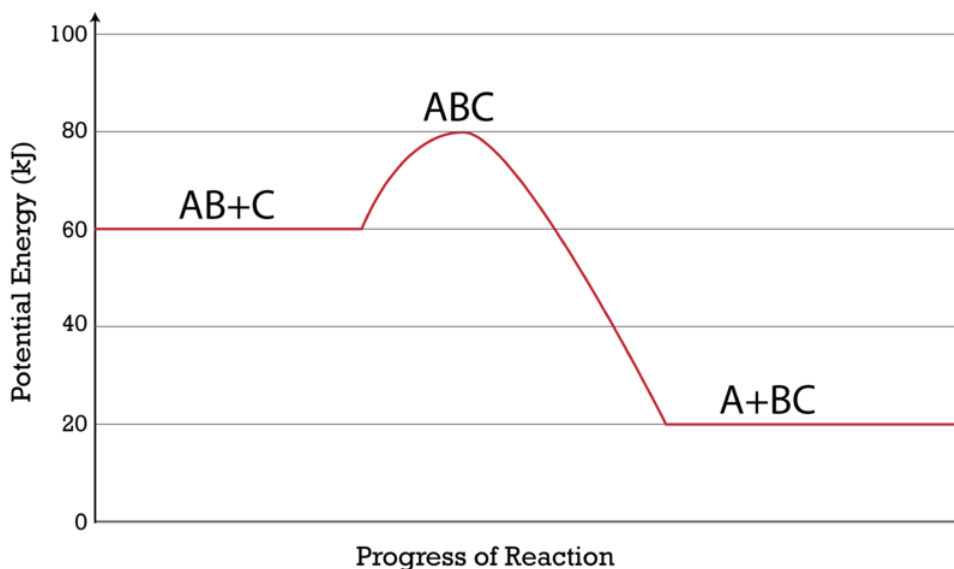
1. Is the overall reaction as shown exothermic or endothermic? _____
2. What is the activation energy for the forward reaction? _____
3. What is the activation energy for the reverse reaction? _____
4. What is the enthalpy change for (ΔH) for the forward reaction? _____
5. What is the ΔH for the reverse reaction? _____
6. Is the reverse reaction exothermic or endothermic? _____
7. Which species is the activated complex? _____
8. Which species or group of species has the *highest* potential energy? _____
9. Which species or group of species has the *weakest* bonds? _____
10. Which species or group of species has the *strongest* bonds? _____
11. Which do you think would be *faster* at that the same temperature, the forward or reverse reaction? _____ -

12. What is the threshold energy for the forward reaction? _____

13. In general, as reactant particles begin a collision, the potential energy _____ (increases, decreases, stays the same) and the kinetic energy _____ (increases, decreases, stays the same).

14. Describe what happens to two reactant particles that collide with less than the activation energy?

Use the following Potential Energy Diagram to answer questions 15 - 22.



15. What is the activation energy for the forward reaction? _____

16. What is the activation energy for the reverse reaction? _____

17. What is the ΔH for the forward reaction? _____

18. What is the ΔH for the reverse reaction? _____

19. Is the forward reaction exothermic or endothermic? _____

20. What is the threshold energy for the forward reaction? _____

21. Which bond is stronger, $A - B$ or $B - C$? _____

22. Give a reason for your answer in question 21.

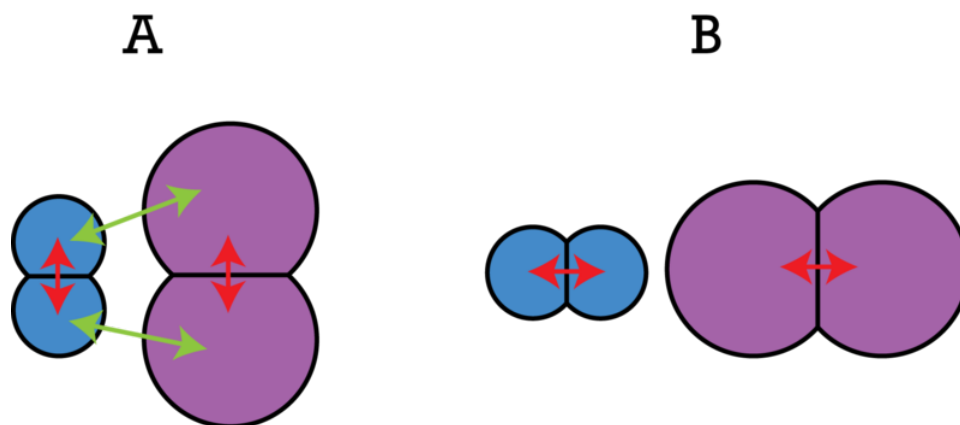
Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

18.7 Extra Reading for Chapter 18

Chemical Kinetics

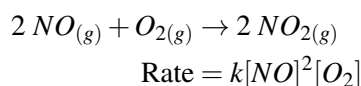
Chemical Kinetics is the study of the mechanisms and rates of chemical reactions. In order for a reaction to occur, a collision between reacting particles must occur. Assuming the reactant is a molecule, the atoms in the molecule are already bonded to at least one other atom. If the atom is to form new bonds during the reaction, the old bonds must first be broken. In some cases, the collision to break the old bonds must also have proper orientation. For example, consider the reaction between H_2 and I_2 shown below.



During the collision between H_2 and I_2 , the $H-H$ bond and the $I-I$ bond (indicated by red arrows) must be broken and $H-I$ bonds (indicated by green arrows) must be formed. In collision A, the side-to-side collision of the molecules would require the least amount of energy to break the old bonds and the atoms would be in convenient position to form HI molecules. In collision B, however, the end-to-end collision would appear to push the atoms together rather than break them apart, and the H and I atoms in the outside positions are not in convenient position to bond. If collision B were to result in a reaction, a great deal more energy would be required than that for collision A.

Reaction rates are affected by the concentrations of the reacting species, the temperature of the reaction, and whether or not a catalyst is present. Reaction rates (at constant temperature) can be expressed in a mathematical expression relating the rate of a reaction to the concentrations of the reactants. This rate law can be determined from experimental data.

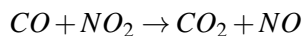
Here is an example of an overall chemical reaction and the rate law for that reaction.



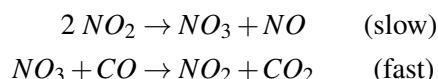
The rate expresses the rate of production of NO_2 in moles/liter/sec or M/s, and is proportional to the concentrations of the reactants where k is a proportionality constant called the reaction constant. The exponent associated with each reactant is referred to as the order of the reaction with respect to that reactant. In this case, the reaction is 2^{nd} order

with respect to NO and 1st order with respect to O_2 . The overall order of the reaction is the sum of partial orders with respect to each reactant. In this case, the overall order of the reaction is 3rd order.

The reaction mechanism is the series of collisions that describe the steps involved in the reaction. Consider the following reaction.



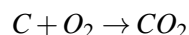
In this reaction, it has been experimentally determined that this reaction takes place according to the rate law $R = k[NO_2]^2$. Therefore, a possible mechanism by which this reaction takes place is:



The first collision in this reaction occurs between two NO_2 molecules and produces an NO_3 molecule and an NO molecule. The second collision in the mechanism occurs between the NO_3 molecule produced in step 1 and a CO molecule producing an NO_2 and a CO_2 . When all the steps in the reaction mechanism are added, the NO_3 (on both sides) cancel and NO_3 does not appear in the net reaction. The NO_2 in the product cancels one of the NO_2 molecules in the reactant and the net reaction is $CO + NO_2 \rightarrow CO_2 + NO$. The overall reaction rate for this reaction (and all reactions) will be exactly the same as the reaction rate for the slowest step. The rate law for the slowest step is $R = k[NO_2]^2$ and therefore, the rate law for the net reaction is the same, even though two NO_2 molecules do not appear in the reactants for the net reaction. That is why the rate law for the net reaction must be determined experimentally.

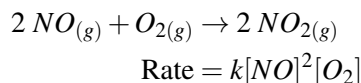
Some chemical reactions may occur with a single collision between reactant particles. The possibility of a single collision reaction is limited to reactions involving two particles or in some cases, three particles. The probability of three particles arriving at the same point at the same time for a single three-particle collision is low. Collisions involving more than three particles essentially never occur.

Suppose the reaction between carbon and oxygen to yield carbon dioxide occurred with a single collision between a carbon atom and an oxygen molecule.

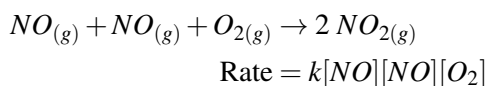


In such a case, the reaction mechanism and the net reaction are the same reaction. The net reaction represents the reaction mechanism and the slowest step in the reaction mechanism. Therefore, for this very simple reaction, the rate law may be written by looking at the net reaction; Rate = $k[C][O_2]$.

If a three particle reaction occurred in a single collision, the rate law could also be written from the net reaction.



This reaction and the rate law could also be written in the following manner,



and that's why the coefficients of the reactants become exponents in the rate law.

The great majority of reactions that involve more than two particles as reactants occur by a series of collisions (reaction mechanism) and for these reactions, the rate law must be determined experimentally.

Consider the following set of experimental data from which the rate law may be determined for the reaction between NO and O_2 .

TABLE 18.7: Rate Data Table for the Reaction Between

Trial	Initial $[NO]$	Initial $[O_2]$	Experimentally Determined Rate
1	0.10 M	0.10 M	$1.2 \times 10^{-8} M/s$
2	0.10 M	0.20 M	$2.4 \times 10^{-8} M/s$
3	0.30 M	0.10 M	$1.08 \times 10^{-7} M/s$

We pick two trials in which one of the reactant concentrations is held constant and the other reactant concentration changes. To begin, we choose trials 1 and 2 in which the concentration of NO is constant and the concentration of O_2 changes. We can determine the order of the reaction with respect to O_2 with the following mathematics.

(multiple of O_2 concentration)^x = (multiple of rate), where the exponent, x , is the order of the reaction with respect to O_2 .

The concentration of O_2 has been doubled and the rate has been doubled, so

$2^x = 2$ and therefore, $x = 1$. The order of the reaction with respect to oxygen is 1.

We now choose two trials in which the O_2 concentration is held constant and the concentration of NO varies. In trials 1 and 3, the concentration of NO has been tripled and the rate has been increased by a factor of 9.

$3^x = 9$, hence $x = 2$. The order of the reaction with respect to NO is 2. Now, we can write the rate law.

$$\text{Rate} = k[NO]^2[O_2]$$

We can determine the value of k by choosing any one of the trials and substituting the known values for the concentrations and rate. Inserting the values from trial 1 into the rate law yields

$$1.2 \times 10^{-8} M/s = k(0.10)^2(0.10), \text{ and solving for } k \text{ yields}$$

$$k = 1.2 \times 10^{-5} M^{-2}s^{-1}.$$

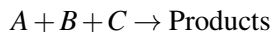
Thus, the rate law for this reaction is:

$$\text{Rate} = (1.2 \times 10^{-5} M^{-2}s^{-1})[NO]^2[O_2].$$

As long as the reaction occurs at the temperature for which this rate law was determined, the rate can be determined by plugging in the initial concentrations of the reactants. The value of k changes with temperature, so this k value is only true at the specific temperature for which the data was determined.

In a number of reactions, the order of the reaction for a particular reactant will be determined to be zero. This indicates that the reaction rate does not depend on the concentration of that reactant and the reactant will not appear in the rate law. (Anything raised to the power of 0 equals 1.)

Consider the following reaction and experimental data.

**TABLE 18.8: Experimental Rate Data Table**

Trial	Initial [A]	Initial [B]	Initial [C]	Rate
1	1.0 M	1.0 M	1.0 M	0.40 M/s
2	1.0 M	1.0 M	2.0 M	0.40 M/s
3	1.0 M	2.0 M	1.0 M	1.6 M/s
4	2.0 M	2.0 M	1.0 M	1.6 M/s

The reaction rate will be related to the equation $R = k[A]^a[B]^b[C]^c$

Comparing trials 1 and 2, we have [A] and [B] remaining constant while [C] is doubled. The rate also remains the same.

$$(\text{multiple of C concentration})^c = (\text{multiple of rate})$$

$2^c = 1$, so the exponent, c , must equal 0 anything to the power of zero equals 1.

Comparing trials 3 and 4, we have [B] and [C] remaining constant while [A] is doubled. The rate remains the same.

$$(\text{multiple of A concentration})^a = (\text{multiple of rate})$$

$2^a = 1$, so the exponent, a , must equal 0 anything to the power of zero equals 1.

Comparing trials 1 and 3, we have [A] and [C] remaining constant while [B] is doubled. The rate increases by a factor of 4.

$$(\text{multiple of B concentration})^b = (\text{multiple of rate})$$

$2^b = 4$, so the exponent, b , must equal 2.

Therefore, the rate expression will be:

$$\text{Rate} = k[A]^0[B]^2[C]^0 = k[B]^2.$$

Catalytic Converters

In 2007, there were approximately 200 million vehicles on American roads, with perhaps 450 million operating in the world. For each of these cars and trucks, their mobility stems from an internal combustion engine. In these engines, hydrocarbons combine with oxygen in a combustion process, producing carbon dioxide and water vapor gases, in addition to a host of other combustion byproducts, such as nitrogen oxides and sulfur oxides. Due to their sheer numbers and the volume of gases produced, environmental concerns about increasing levels of carbon dioxide emissions leading to global warming, and the links of nitrogen and sulfur oxides with acid rain, have led to

the development of catalytic convertors, first implemented in automobiles in 1975. These catalytic convertors are devices installed in the emission train of automobiles containing catalysts consisting of precious metal surfaces that convert harmful emission gases like NO_x and CO into N_2 and CO_2 respectively.

Catalysts are substances that can enable or accelerate chemical reactions without being consumed in the reactions themselves. In the case of catalytic convertors, originally developed by Eugene Houdry to minimize the effects of automobile exhaust, the system consists of a large surface area porous support for the actual catalyst, a precious or rare metal oxide, like those of platinum, palladium or rhodium. The amount of metals due to their extreme cost is minimal, usually consisting of a thin surface on the support.

Modern catalytic convertors are usually formulated to be three way systems: 1. reducing NO_x to N_2 ; 2. oxidizing carbon monoxide and 3. oxidizing any remaining incombustible hydrocarbons to carbon dioxide. Emission gases exiting from the engine reach the exhaust manifold at temperatures about $500^\circ C$. As the chemical reductions and oxidations occur only on the surface, only a small amount, about $\frac{1}{5}$ of an ounce of palladium metal is needed. Its enormous effectiveness in this capacity outstrips the extreme cost, currently more than \$200 per ounce leads to two additional problems: how to recover the metal when the automobile is no longer in use, and more pressingly, the issue of catalytic converter theft to recover the precious metal. Car enthusiasts also lament that the use of the catalytic converter in the exhaust system compromises the horsepower output and back pressure yet the Federal Clean Air Act mandates their inclusion in all vehicles.

The catalytic converter is a modern-day environmental success story, removing hundreds of tons of carbon monoxide, and 3050 million tons of excess hydrocarbons and nitrogen oxides. Air quality in the United States, even with a growing population and ever-increasing dependence on fossil fuel vehicles, is vastly improved in all major categories since 1975.

One problem that may exist with catalytic convertors is that in recent years, researchers at the University of California have detected significantly higher levels of another pollutant, ammonia, in the exhaust of automobiles with catalytic convertors. The researchers indicate that while catalytic convertors have played a major part in reducing air pollution caused by automobiles, these latest findings suggests that while fixing one problem, convertors may have caused an unexpected secondary problem. Further research will be necessary to determine if the catalytic convertors are producing the higher levels of ammonia in automobile exhaust.

18.8 Assessment for Chapter 18

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 19

Chemical Equilibrium C-SE-TE

Chapter Outline

- 19.1 INTRODUCTION TO EQUILIBRIUM
- 19.2 EQUILIBRIUM CONSTANT
- 19.3 THE EFFECT OF APPLYING STRESS TO REACTIONS AT EQUILIBRIUM
- 19.4 SLIGHTLY SOLUBLE SALTS
- 19.5 WORKSHEETS FOR CHAPTER 19
- 19.6 EXTRA READING FOR CHAPTER 19
- 19.7 ASSESSMENT FOR CHAPTER 19

Lessons and Number of Activities for Lessons

TABLE 19.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Introduction to Equilibrium	0	1	0	1
2. Equilibrium Constant	0	0	1	0
3. The Effect of Applying Stress to Reactions at Equilibrium	1	1	1	0
4. Slightly Soluble Salts	0	0	1	0

19.1 Introduction to Equilibrium

Student Behavioral Objectives

The student will:

- describe the three possibilities that exist when reactants come together.
- identify the three possibilities by looking at a chemical equation.
- describe what is occurring in a system at equilibrium.
- define dynamic equilibrium.
- define the conditions of dynamic equilibrium.

Timing, Standards, Activities

TABLE 19.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Introduction to Equilibrium</i>	1.0	9b

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. A Light Activated Reversible Chemical Reaction

Worksheets

1. None

Extra Readings

1. Chemical Equilibrium

Answers for Introduction to Equilibrium (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

19.2 Equilibrium Constant

Student Behavioral Objectives

The student will:

- write equilibrium constant expressions.
- use equilibrium constant expressions to solve for unknown concentrations.
- use known concentrations to solve for the equilibrium constants.
- explain what the value of K means in terms of relative concentrations of reactants and products.

Timing, Standards, Activities

TABLE 19.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Equilibrium Constant</i>	1.5	9c

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Equilibrium Worksheet

Extra Readings

1. None

Answers for Equilibrium Constant (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

19.3 The Effect of Applying Stress to Reactions at Equilibrium

Student Behavioral Objectives

The student will:

- state Le Châtelier's Principle.
- demonstrate on specified chemical reactions how Le Châtelier's Principle is applied to equilibrium systems.
- describe the effect of concentration on an equilibrium system.
- demonstrate with specific equations how Le Châtelier's Principle explains the effect of concentration.
- describe the effect of pressure as a stress on the equilibrium position.
- describe the pressure effect in Le Châtelier's Principle.
- describe the effect of temperature as a stress on an equilibrium system.
- explain how Le Châtelier's principle explains the effect of temperature.
- explain how a catalyst works in equilibrium reactions.
- explain the effect of a catalyst in equilibrium positions.

Timing, Standards, Activities

TABLE 19.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Effect of Applying Stress to Reactions at Equilibrium</i>	2.0	9a

Activities for Lesson 3

Laboratory Activities

1. Solution Equilibria Lab

Demonstrations

1. Equilibrium Between Nitrogen Dioxide and Dinitrogen Tetroxide

Worksheets

1. Le Chatelier's Principle Worksheet

Extra Readings

1. None

Answers for The Effect of Applying Stress to Reactions at Equilibrium (L3) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

19.4 Slightly Soluble Salts

Student Behavioral Objectives

The student will:

- define solubility product constants.
- write solubility product constant expressions.
- calculate solubility product constants.

Timing, Standards, Activities

TABLE 19.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Slightly Soluble Salts</i>	1.5	None

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Solubility and Solubility Product Constant Worksheet

Extra Readings

1. None

Answers for Slightly Soluble Salts (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 19

The two videos below provide an introduction to chemical equilibrium.

- <http://www.youtube.com/watch?v=vCNENmgNJYg>
- http://www.youtube.com/watch?v=KINI12Q_aJY

This website provides a video on Le Chatelier's Principle.

- <http://genchemist.wordpress.com/2007/10/05/le-chateliers-principle-the-movie/>

This website provides a video on Le Chatelier's Principle and the Haber Process.

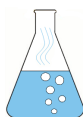
- <http://videos.howstuffworks.com/hsw/12468-chemistry-connections-le-chatelier-and-haber-bosch-video.htm>

A Khan Academy electronic lecture on Le Châtelier's Principle is available at

- <http://www.youtube.com/watch?v=4-fEvpVNTIE>

Laboratory Activities for Chapter 19

Teacher's Resource Pages for Solution Equilibria



Investigation and Experimentation Objectives

In this activity, the student will, through a series of laboratory activities, develop evidence for the existence of the equilibrium state, and answer a series of questions to develop a conclusion about the nature of the equilibrium state.

Lab Notes:

This lab involves the Fe^{3+}/SCN^- equilibria, where the $Fe[SCN]^{2+}$ complex is formed. The value of K at room temperature is about 138. The equilibrium original equilibrium position produced by the lab directions will be a light reddish orange and this position will be shifted by various activities during the lab.

Solution Preparation

To prepare 250. mL of 0.10 M $FeCl_3$ solution from solid anhydrous iron (III) chloride, dissolve 4.1 g $FeCl_3$ in a 250. mL volumetric flask and fill to the line. If you are using iron (III) chloride hexahydrate, dissolve 6.8 g $FeCl_3 \cdot 6H_2O$ in a 250. mL volumetric flask and fill to the line.

To prepare 250. mL of 0.10 M $KSCN$ solution from solid $KSCN$, dissolve 2.4 g of the solid in a 250. mL volumetric flask and fill to the line.

Answers to Pre-Lab Questions

The following questions refer to a hypothetical reversible chemical reaction in which reactant Y is a bright yellow color, reactant C is colorless, and the product B is a bright blue color.



1. At the original equilibrium position, the solution is green. Describe the color change you would expect if a large quantity of C were added to the reaction solution.

Since the addition of a large quantity of C will drive the equilibrium to the right, yellow color (Y) will decrease significantly and blue color (B) will increase . . . so the final solution will change from green to blue.

2. At the original equilibrium position, the solution is green. Describe the color change you would expect if a large quantity of C were removed from the reaction solution.

Since the removal of C from the solution would shift the equilibrium to the left, the blue color (B) would decrease and yellow color (Y) would increase . . . so the final solution would change from green to yellow.

Answers to Post-Lab Questions

1. When $FeCl_3$ was added to the equilibrium solution in test tube 2, which way did the equilibrium shift? How do you know?

Forward . . . darker red color indicated more product was formed.

2. When $KSCN$ was added to the equilibrium solution in test tube 3, which way did the equilibrium shift? How do you know?

Forward . . . darker red color indicated more product was formed.

3. When crystals of potassium phosphate were added to the equilibrium solution in test tube 4, which way did the equilibrium shift? How do you know?

Backward . . . lessening of red color indicated product was being used up.

4. If more product formed in test tube 2 when $FeCl_3$ was added, what other substance had to be present in the equilibrium solution before the $FeCl_3$ was added? How do you know?

$KSCN$. . . in order to form more product, both reactants had to be available.

5. If more product formed in test tube 3 when $KSCN$ was added, what other substance had to be present in the equilibrium solution before the $KSCN$ was added? How do you know?

$FeCl_3$. . . in order to form more product, both reactants had to be available.

6. Therefore, what substances were present in the equilibrium solution in test tube 1?

Fe^{3+} and SCN^- and $Fe[SCN]^{2+}$

7. Given the information that phosphate ion removes iron (III) ion from solution, describe what happened when phosphate ion was added to test tube 4.

The added phosphate ions removed Fe^{3+} causing the equilibrium to shift backward, thus using up $Fe[SCN]^{2+}$, resulting in a decrease in the red color.

Solution Equilibria Laboratory

Background Information

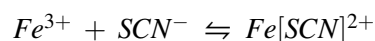
A state of equilibrium is affected by concentration of reactants, temperature, and pressure (for reactions containing gaseous substances). If a system at equilibrium is subjected to a change in one or more of these factors, a stress is placed on the equilibrium. When a stress is placed on a system at equilibrium, the equilibrium position will shift in the direction that tends to relieve the stress.

Materials and Apparatus (per lab group)

Beaker, 100 mL Graduated cylinder, 10 mL Test tubes (4), small Test tube rack Dropper pipet 0.10 M $FeCl_3$ solution, 10 mL 0.10 M $KSCN$ solution, 10 mL Solid potassium or sodium phosphate, a few crystals Distilled water

Safety Issues

Safety glasses and apron should be worn at all times while working in the chemistry laboratory.

Procedure

The iron (III) ion solution is light brown, the thiocyanate ion solution is nearly colorless and the product complex is dark red.

1. Measure 5 mL of 0.10 M $FeCl_3$ solution and place it in a 100 mL beaker. Add 5 mL of 0.10 M $KSCN$ to the same beaker. Dilute the contents of the beaker with distilled water until the solution is a light reddish-orange color. Divide the solution equally among the four numbered test tubes. Set tube 1 at one end of the rack to be used for color comparison.
2. Using a dropper pipet, add 0.10 M $FeCl_3$ solution drop by drop to the solution in test tube 2 with stirring until a significant color change occurs. Record your observations and rinse the pipet with distilled water.
3. Add 0.10 M $KSCN$ solution drop by drop to the solution in test tube 3 with stirring until a significant color change occurs. Record your observations and rinse the pipet with distilled water.
4. Drop a couple of crystals of potassium phosphate into the solution in test tube 4 with stirring. Record your observations.
5. Discard the solutions as instructed by your teacher.

Data:**TABLE 19.6: Data Table**

Test Tube Number	Color
1	
2	
3	
4	

Post-Lab Questions

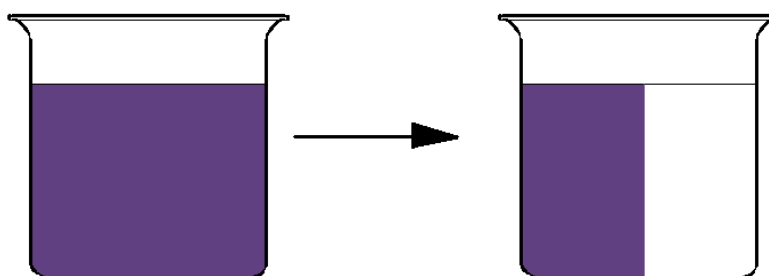
1. When $FeCl_3$ was added to the equilibrium solution in test tube 2, which way did the equilibrium shift? How do you know?
2. When $KSCN$ was added to the equilibrium solution in test tube 3, which way did the equilibrium shift? How do you know?
3. When crystals of potassium phosphate were added to the equilibrium solution in test tube 4, which way did the equilibrium shift? How do you know?
4. If more product formed in test tube 2 when $FeCl_3$ was added, what other substance had to be present in the equilibrium solution before the $FeCl_3$ was added? How do you know?
5. If more product formed in test tube 3 when $KSCN$ was added, what other substance had to be present in the equilibrium solution before the $KSCN$ was added? How do you know?
6. Therefore, what substances were present in the equilibrium solution in test tube 1?
7. Given the information that phosphate ion removes iron (III) ion from solution, describe what happened when phosphate ion was added to test tube 4.

Demonstrations for Chapter 19

A Light Activated Reversible Chemical Reaction

What the student sees.

The room is somewhat darkened. The students see a large beaker containing a purple solution sitting on an overhead projector. Half of the glass plate of the overhead project is covered with several folded layers of aluminum foil. The beaker is sitting such that half of it is on the aluminum foil and half of it is on the glass plate. When the overhead projector is turned on, the half of the solution in the beaker that is above the light turns clear . . . the other half remains purple.



The solution appears to be divided by an invisible line running vertically through the beaker. When the overhead projector is turned off, the colorless side of the solution will slowly return to purple. The reaction can be repeated and is reversible for a couple of days.

Procedure:

- Mix together in a 1-liter beaker
 - 10. mL of freshly prepared 0.001 M thionin solution. 0.001 M thionin solution can be prepared by dissolving 0.023 g of Thionin in 100. mL of distilled water. (Thionin solution has a shelf life of only a few days – so prepare it only when you are ready to use it.)
 - 100. mL of 1.0 M H_2SO_4
 - sufficient distilled water to bring the total volume to 600 mL.
 - mix thoroughly
- Turn off the room lights and add 2.0 grams of iron(II) sulfate. Stir to dissolve.

Disposal

The solution can be rinsed down the drain followed by excess water.

Discussion

The two forms of thionin differ in oxidation state and the redox reaction converting one form to the other also involves the conversion Fe^{2+} ions to Fe^{3+} ions. The activation energy for the reaction is provided by light.

Source of materials

Most chemical supply companies carry all the materials necessary for this demonstration. A kit specifically designed for this demonstration is offered by Flinn Scientific.

Equilibrium Between Nitrogen Dioxide and Dinitrogen Tetroxide Demonstration

Apparatus and Materials:

- Three sealed borosilicate glass tubes filled with nitrogen dioxide gas - demo tubes are available from most chemical supply companies (store in bubble wrap)
- Three tall form 1-liter beakers
- Hot plate
- Tap Water
- Ice

Procedure:

1. Fill one of the beakers with tap water and place it on a hot plate, heat until boiling.
2. Fill another beaker with a mixture of ice and tap water.
3. Displace the sealed tubes containing nitrogen dioxide (should be same color due to same temperature). A white background makes them much more visible.
4. Place one tube in each of the hot and cold beakers and leave the third at room temperature.
5. As the equilibrium in the tubes adjust to the new temperatures, there will be significant differences in the colors of the tubes. The hot tube becomes dark brown, the cold tube may become nearly colorless.
6. The tubes can be removed from the beakers and switched to the other beaker and the colors will change again.

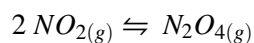


Safety Issues:

The borosilicate tubes can easily withstand the temperature extremes of these two water baths, care should be taken, however, to not damage the tubes and allow the nitrogen dioxide to escape. Nitrogen dioxide is an extremely toxic gas. It is irritating to the respiratory tract.

Discussion:

The equilibrium illustrated in this demonstration is between nitrogen dioxide, NO_2 and dinitrogen tetroxide, N_2O_4 . The chemical equation is shown below.



Nitrogen dioxide is a dark reddish brown gas and dinitrogen tetroxide is a colorless gas. When the equilibrium is shifted to the left, as written above, the amount of nitrogen dioxide increases, the amount of dinitrogen tetroxide decreases, and the color of the tube darkens. When the equilibrium is shifted to the right, as written, the amount of nitrogen dioxide decreases, the amount of dinitrogen tetroxide increases, and the color of the tube lightens. As written, the reaction releases 58 *kJ* of energy. Since this is an exothermic reaction, increasing the temperature will drive it to the left and decreasing the temperature will drive it to the right.

19.5 Worksheets for Chapter 19

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Equilibrium Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Questions 1 - 20 relate to the following reaction at equilibrium in a closed container.



1. What is the instantaneous effect on the FORWARD REACTION RATE of adding some solid phosphorus with no change in surface area?
 - A. Increase.
 - B. Decrease.
 - C. No change.
2. What is the instantaneous effect on the FORWARD REACTION RATE of adding some oxygen gas with no change in pressure?
 - A. Increase.
 - B. Decrease.
 - C. No change.
3. What is the instantaneous effect on the FORWARD REACTION RATE of adding some PO_4 gas with no change in pressure?
 - A. Increase.
 - B. Decrease.
 - C. No change.
4. What is the instantaneous effect on the FORWARD REACTION RATE of increasing the temperature?
 - A. Increase.
 - B. Decrease.
 - C. No change.
5. What is the instantaneous effect on the FORWARD REACTION RATE of increasing the pressure by reducing the volume?
 - A. Increase.
 - B. Decrease.

C. No change.

6. What is the instantaneous effect on the FORWARD REACTION RATE of adding a catalyst?

A. Increase.

B. Decrease.

C. No change.

7. What is the instantaneous effect on the REVERSE REACTION RATE of adding some solid phosphorus with no change in surface area?

A. Increase.

B. Decrease.

C. No change.

8. What is the instantaneous effect on the REVERSE REACTION RATE of adding some oxygen gas with no change in pressure?

A. Increase.

B. Decrease.

C. No change.

9. What is the instantaneous effect on the REVERSE REACTION RATE of adding some PO_4 gas with no change in pressure?

A. Increase.

B. Decrease.

C. No change.

10. What is the instantaneous effect on the REVERSE REACTION RATE of increasing the temperature?

A. Increase.

B. Decrease.

C. No change.

11. What is the instantaneous effect on the REVERSE REACTION RATE of increasing the pressure by reducing the volume?

A. Increase.

B. Decrease.

C. No change.

12. What is the instantaneous effect on the REVERSE REACTION RATE of adding a catalyst?

A. Increase.

B. Decrease.

C. No change.

13. Which direction will the equilibrium shift when solid phosphorus is added with no change in surface area?

A. Forward.

B. Reverse.

C. No shift.

14. Which direction will the equilibrium shift when oxygen gas is added with no change in pressure?

A. Forward.

B. Reverse.

C. No shift.

15. Which direction will the equilibrium shift when gaseous PO_4 is added with no change in pressure?

A. Forward.

B. Reverse.

C. No shift.

16. Which direction will the equilibrium shift when the temperature is increased?

A. Forward.

B. Reverse.

C. No shift.

17. Which direction will the equilibrium shift when the pressure is increased by reducing the volume?

A. Forward.

B. Reverse.

C. No shift.

18. Which direction will the equilibrium shift when a catalyst is added?

A. Forward.

B. Reverse.

C. No shift.

19. Which of the following changes to the system at equilibrium will change the value of the equilibrium constant?

I. Adding some solid phosphorus.

II. Adding some oxygen gas.

III. Increasing the pressure by reducing the volume.

IV. Increasing the temperature.

V. Adding a catalyst.

A. I, II, and IV.

B. III, IV, and V.

C. IV and V.

D. IV only.

E. V only.

20. If oxygen gas is added to the system at equilibrium, the equilibrium will shift forward until a new equilibrium is established. When the new equilibrium is established, how will the concentration of oxygen gas in the new equilibrium compare to the original concentration of oxygen gas before the stress was applied?

A. higher

B. lower

C. the same

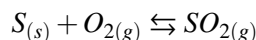
21. Here are four equations with their equilibrium constant values. Which of these reactions will have the greatest

proportion of material in the form of products?

TABLE 19.7: Equilibrium Constants for Various Equations

Choice	Equation	Equilibrium Constant
A.	$AB_{(aq)} \rightleftharpoons A_{(aq)}^+ + B_{(aq)}^-$	$K_e = 2 \times 10^{-2}$
B.	$CD_{(aq)} \rightleftharpoons C_{(aq)}^+ + D_{(aq)}^-$	$K_e = 3 \times 10^{-2}$
C.	$EF_{(aq)} \rightleftharpoons E_{(aq)}^+ + F_{(aq)}^-$	$K_e = 3 \times 10^{-3}$
D.	$GH_{(aq)} \rightleftharpoons G_{(aq)}^+ + H_{(aq)}^-$	$K_e = 6 \times 10^{-3}$

22. Solid sulfur reacts with oxygen gas to form $SO_{2(g)}$ according to the following equation.



Given that the equilibrium constant for the reaction is 5.00 and that the reaction begins with 60.0 M sulfur and 3.00 M O_2 , calculate the equilibrium concentration of SO_2 .

- A. 15.0 M
- B. 5.55 M
- C. 2.50 M
- D. 1.25 M
- E. None of these.

23. For the reaction, $N_{2(g)} + O_{2(g)} \rightleftharpoons 2 NO_{2(g)}$, the equilibrium constant is 1.0×10^{-6} . Find the equilibrium concentration of NO_2 if the beginning concentration of N_2 and O_2 are both 2.0 M?

- A. 0.0020 M
- B. 2.0×10^{-6} M
- C. 4.0×10^{-6} M
- D. 0.020 M
- E. None of these.

24. For the reaction, $H_{2(g)} + CO_{2(g)} \rightleftharpoons H_2O_{(g)} + CO_{(g)}$, the two reactants begin the reaction at 1.0 M and at equilibrium, the concentration of CO is found to be 0.80 M. What is the equilibrium constant value?

- A. 1.7
- B. 2.0
- C. 4.0
- D. 16
- E. None of these.

25. $K_e = 4.00$ for the reaction, $H_{2(g)} + CO_{2(g)} \rightleftharpoons H_2O_{(g)} + CO_{(g)}$. If all four species begin at 1.00 M, what will be the equilibrium concentration of H_2 ?

- A. 0.33 M
- B. 0.67 M
- C. 1.3 M
- D. 1.0 M

E. None of these.

Le Chatelier's Principle Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Le Chatelier's Principle is useful in predicting how a system at equilibrium will respond when certain changes are imposed. Le Chatelier's Principle does NOT explain why the system changes, and is not an acceptable explanation for the change. It merely allows you to determine quickly how the system will change when a disturbance is imposed. The explanation for why the system changes can be found in your textbook.

There are three common ways a stress may be applied to a chemical system at equilibrium:

- changing the concentration (or partial pressure) of a reactant or product.
- changing the temperature.
- changing the volume of the container (which changes partial pressure of all gases in the reaction).

You should be aware that adding a gaseous substance that is not involved in the reaction changes the total pressure in the system but does not change the partial pressure of any of the reactants or products and therefore does not affect the equilibrium.

Le Chatelier's Principle states *when a system at equilibrium is disturbed, the equilibrium shifts so as to **partially** undo (counteract) the effect of the disturbance.*

Changes in Concentration or Partial Pressure

If a system at equilibrium is disturbed by adding a reactant or removing a product, Le Chatelier's Principle predicts that the equilibrium will shift forward, thus using up some of the added reactant or producing more of the removed product. In this way, the equilibrium shift partially counteracts the disturbance. Similarly, if the disturbance is the removal of a reactant or the addition of a product, the equilibrium will shift backward, thus producing more of the removed reactant or using up some of the added product. Once again, the shift tends to "undo" the disturbance. It should be noted that when the disturbance is an increase or decrease of concentration of reactant or product, the equilibrium shift tends to **partially** return the concentration to its former value but **it never gets all the way back to the former value.**

The equilibrium constant value, K_c is not changed by the addition or removal of reactants or products. Since the concentration of solids are constant, they do not appear in the equilibrium constant expression and their concentrations do not change when disturbances cause equilibrium shifts, however, the **amount** of the solid present most certainly does change. The amount of solid can increase or decrease but the *concentration* does not change.

Changes in Temperature

Increasing the temperature of a system at equilibrium increases both forward and reverse reaction rate, but it increases the endothermic reaction more than the exothermic. Therefore, in an exothermic reaction, the reverse reaction is endothermic and so increasing the temperature will increase the reverse reaction more than the forward reaction, and the equilibrium will shift backwards. Since the forward reaction produces heat and the reverse reaction consumes heat, Le Chatelier's Principle predicts that when heat is added, the equilibrium will shift backward, consuming heat, and thus partially countering the disturbance. Cooling an exothermic reaction slows both reactions but it slows the reverse more than the forward, hence the equilibrium will shift forward producing more heat, thus partially undoing the stress.

For an endothermic reaction, all the same logic is involved except that the forward reaction is endothermic and the reverse reaction is exothermic. Therefore, heating an endothermic reaction causes the equilibrium to shift forward, and cooling an endothermic reaction causes the equilibrium to shift backward.

When an equilibrium shifts due to a temperature change all the substances on one side of the equation move in the same direction, that is, they all increase or they all decrease. Therefore, the equilibrium constant value will also change when the temperature is changed.

TABLE 19.8: Summary of

Reaction Type	Increase Temperature	Decrease Temperature
Endothermic ($\Delta H > 0$)	K increases	K decreases
Exothermic ($\Delta H < 0$)	K decreases	K increases

Changes in Volume

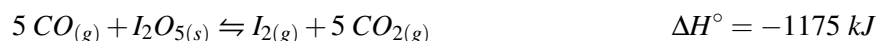
When the volume of a reaction vessel is decreased, the partial pressure (and concentration) of all gases in the container increase. The total pressure in the vessel will also increase. Le Chatelier's Principle predicts that the equilibrium will shift in a direction that tends to counteract the disturbance. Therefore, the equilibrium will shift to produce fewer moles of gaseous substances so that the pressure will decrease. Thus, decreasing the volume will cause the equilibrium to shift toward the side with fewer moles of gaseous substances. The reverse is true if the volume of the vessel is increased. The partial pressure of all gases will decrease, and the total pressure will decrease, so the equilibrium shift will be toward the side that contains more moles of gas, thus increasing pressure and partially counteracting the change.

The Addition of a Catalyst

The addition of a catalyst will increase both forward and reverse reaction rates. In the case of a catalyst, both reaction rates are increased by the same amount and therefore there will be no equilibrium shift.

Exercises

Consider the following reaction.



- If some $\text{CO}_{2(g)}$ is added to this system at equilibrium, which way will the equilibrium shift?
 - Toward the products.
 - Toward the reactants.
 - No shift.
- When equilibrium is re-established after the $\text{CO}_{2(g)}$ is added, how will the concentration of $\text{I}_{2(g)}$ compare to the original concentration?
 - Increased.
 - Decreased.
 - No change.
- When equilibrium is re-established after the $\text{CO}_{2(g)}$ is added, how will the concentration of I_2O_5 compare to the original concentration?
 - Increased.
 - Decreased.
 - No change.
- When equilibrium is re-established after the $\text{CO}_{2(g)}$ is added, how will the amount of I_2O_5 compare to the original amount?

- A. Increased.
- B. Decreased.
- C. No change.
5. When equilibrium is re-established after the $CO_{2(g)}$ is added, how will the value of K compare to the original value of K ?
- A. Higher.
- B. Lower.
- C. No change.
6. If some $I_{2(g)}$ is removed from this system at equilibrium, which way will the equilibrium shift?
- A. Toward the products.
- B. Toward the reactants.
- C. No shift.
7. When equilibrium is re-established after the $I_{2(g)}$ is removed, how will the concentration of $CO_{2(g)}$ compare to the original concentration?
- A. Increased.
- B. Decreased.
- C. No change.
8. When equilibrium is re-established after the $I_{2(g)}$ is removed, how will the concentration of $I_{2(g)}$ compare to the original concentration?
- A. Increased.
- B. Decreased.
- C. No change.
9. 5. When equilibrium is re-established after the $I_{2(g)}$ is removed, how will the value of K compare to the original value of K ?
- A. Higher.
- B. Lower.
- C. No change.
10. If the temperature of this system at equilibrium is lowered, which way will the equilibrium shift?
- A. Toward the products.
- B. Toward the reactants.
- C. No shift.
11. When equilibrium is re-established after the temperature was lowered, how will the concentration of $CO_{(g)}$ compare to its original concentration?
- A. Increased.
- B. Decreased.
- C. No change.
12. When equilibrium is re-established after the temperature was lowered, how will the value of K compare to the original value of K ?

A. Higher.

B. Lower.

C. No change.

13. If the volume of the reaction vessel for this system at equilibrium is decreased, which way will the equilibrium shift?

A. Toward the products.

B. Toward the reactants.

C. No shift.

14. When equilibrium is re-established after the volume was decreased, how will the concentration of $CO_{(g)}$ compare to its original concentration?

A. Higher.

B. Lower.

C. No change.

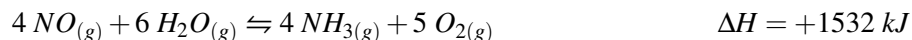
15. When equilibrium is re-established after the volume was decreased, how will the value of K compare to the original value of K ?

A. Higher.

B. Lower.

C. No change.

Consider the following reaction.



16. If some $NO_{(g)}$ is added to this system at equilibrium, which way will the equilibrium shift?

A. Toward the products.

B. Toward the reactants.

C. No shift.

17. When equilibrium is re-established after the $NO_{(g)}$ is added, how will the concentration of $NH_{3(g)}$ compare to the original concentration?

A. Increased.

B. Decreased.

C. No change.

18. If the temperature of this system at equilibrium is raised, which way will the equilibrium shift?

A. Toward the products.

B. Toward the reactants.

C. No shift.

19. When equilibrium is re-established after the temperature was raised, how will the concentration of $NO_{(g)}$ compare to its original concentration?

A. Increased.

B. Decreased.

C. No change.

20. When equilibrium is re-established after the temperature was raised, how will the value of K compare to the original value of K ?

A. Higher.

B. Lower.

C. No change.

Solubility and Solubility Product Constant Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

1. When excess solid $SrCrO_4$ is shaken with water at $25^\circ C$, it is found that 6.00×10^{-3} moles dissolve per liter of solution. Use this information to calculate the K_{sp} for $SrCrO_4$.
2. The solubility of $PbCl_2$ is 1.6×10^{-2} mol/L. What is the K_{sp} for $PbCl_2$?
3. The solubility of $AgC_2H_3O_2$ is 11.11 g/L at $25^\circ C$. What is the K_{sp} for silver acetate at this temperature?
4. The solubility of $Ag_2Cr_2O_7$ is 0.083 g/L at $25^\circ C$. What is the K_{sp} for silver dichromate at this temperature?
5. What is the solubility of AgI in grams/liter given the $K_{sp} = 8.3 \times 10^{-17}$?
6. What is the solubility of $Ca(OH)_2$ in grams/liter given the $K_{sp} = 6.0 \times 10^{-6}$?
7. Write balanced net ionic equations for the precipitation reactions that occur when the following pairs of solutions are mixed. If no reaction occurs, write "no reaction". Use the solubility table in your textbook if you need it.
 - a. Lead nitrate and hydrochloric acid.
 - b. Silver nitrate and lithium hydroxide.
 - c. Ammonium sulfide and cobalt (II) bromide.
 - d. Copper (II) sulfate and potassium carbonate.
 - e. Barium nitrate and copper (II) sulfate.
8. Lead (II) chloride has a K_{sp} value of 1.7×10^{-5} . Will a precipitate form when 140.0 mL of 0.0100 M $Pb_3(PO_4)_2$ is mixed with 550.0 mL of 0.0550 M $NaCl$?
9. A solution contains 1.0×10^{-4} M Pb^{2+} ions and 2.0×10^{-3} M Sr^{2+} ions. If a source of SO_4^{2-} ions is very slowly added to this solution, will $PbSO_4$, ($K_{sp} = 1.8 \times 10^{-8}$) or $SrSO_4$, ($K_{sp} = 3.4 \times 10^{-7}$) precipitate first? Calculate the concentration of SO_4^{2-} ions that will begin to precipitate each cation.

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

19.6 Extra Reading for Chapter 19

Chemical Equilibrium

In principle, any reaction that can be represented by a balanced chemical equation can take place. There are, however, two situations which may inhibit the reaction from occurring.

- The thermodynamic tendency (the combination of entropy and enthalpy) for the reaction to occur may be so small that the quantity of products is very low, or even negligible. This type of chemical reaction is said to be *thermodynamically inhibited*.
- The rate at which the reaction proceeds may be so slow that many years are required to detect any product at all, in which case we say the reaction is *kinetically inhibited*.

As a reaction proceeds, the quantities of the components on one side of the reaction equation will decrease and those on the other side will increase. As the concentrations of the components on one side of the equation decrease, that reaction rate slows down. As the concentrations of the components on the other side of the equation increase, that reaction rate speeds up. Eventually the two reaction rates become equal and the composition of the system stops changing. At this point, the reaction is in its *equilibrium state* and no further change in composition will occur, as long as the system is left undisturbed.

In many reactions, the equilibrium state occurs when significant amounts of both reactants and products are present. Such a reaction is said to be *reversible*. The equilibrium composition is independent of the direction from which it is approached. The labeling of substances as *reactants* and *products* is entirely a matter of convenience.

The *law of mass action* states that any chemical change is a competition between a forward reaction (left-to-right) and a reverse reaction (right-to-left). The rates of these two reactions are governed by the concentrations of the substances reacting, and the temperature. As the reaction proceeds, these two reaction rates approach each other in magnitude and at equilibrium, they become equal.

Since the reactions continue at equilibrium (at equal rates), equilibrium is referred to as *dynamic* equilibrium. At equilibrium, *microscopic changes* (the forward and reverse reactions) continue but *macroscopic changes* (changes in quantities of substances) cease.

When a chemical system is at equilibrium, any disturbance of the system, such as a change in temperature, or the addition or removal of a reactant or product, will cause the equilibrium to shift to a new equilibrium state (different quantities of reaction components). The disturbance in the system causes changes in the reaction rates and quantities of components change until the reaction rates again become identical, and a new equilibrium position is established.

19.7 Assessment for Chapter 19

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER

20**Acid-Bases C-SE-TE****Chapter Outline**

- 20.1 PROPERTIES OF ACIDS AND BASES**
- 20.2 ARRHENIUS ACIDS AND BASES**
- 20.3 THE PH CONCEPT**
- 20.4 STRENGTH OF ACIDS AND BASES**
- 20.5 BRONSTED-LOWRY ACIDS AND BASES**
- 20.6 LEWIS ACIDS AND BASES**
- 20.7 WORKSHEETS FOR CHAPTER 20**
- 20.8 EXTRA READING FOR CHAPTER 20**
- 20.9 ASSESSMENT FOR CHAPTER 20**

Lessons and Number of Activities for Lessons**TABLE 20.1: Lessons and Activities for Lessons**

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Properties of Acids and Bases	0	0	0	0
2. Arrhenius Acids and Bases	0	0	0	0
3. The pH Concept	1	0	1	1
4. Strength of Acids and Bases	0	0	2	0
5. Bronsted-Lowry Acids and Bases	0	0	1	0
6. Lewis Acids and Bases	0	0	0	0

20.1 Properties of Acids and Bases

Student Behavioral Objectives

The student will:

- list the properties of acids.
- list the properties of bases.
- name an acid or base when given the formula.
- write the formula for an acid or base given the name.

Timing, Standards, Activities

TABLE 20.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Properties of Acids and Bases</i>	1.0	5a, 5b, 5c

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Properties of Acids and Bases (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

20.2 Arrhenius Acids and Bases

Student Behavioral Objectives

The student will:

- define an Arrhenius acid and list some substances that qualify as acids under this definition.
- define an Arrhenius base and list some substances that qualify as bases under this definition.

Timing, Standards, Activities

TABLE 20.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Arrhenius Acids and Bases</i>	1.0	5a, 5b, 5e

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Arrhenius Acids and Bases (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

20.3 The pH Concept

Student Behavioral Objectives

The student will:

- calculate $[H^+]$ and $[OH^-]$ for a solution of acid or base.
- define autoionization.
- state the $[H^+]$, $[OH^-]$, and K_w values for the autoionization of water.
- define pH and describe the pH scale.
- write the formulas for pH and pOH and express their values in a neutral solution at 25°C.
- explain the relationships among pH, pOH, and K_w .
- calculate $[H^+]$, $[OH^-]$, pH, and pOH given the value of any one of the other values.
- explain the relationship between the acidity or basicity of a solution and the hydronium ion concentration, $[H_3O^+]$, and the hydroxide ion concentration, $[OH^-]$, of the solution.
- predict whether an aqueous solution is acidic, basic, or neutral from $[H_3O^+]$, $[OH^-]$, or the pH.

Timing, Standards, Activities

TABLE 20.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The pH Concept</i>	1.0	5a, 5d, 5f

Activities for Lesson 3

Laboratory Activities

1. Measuring pH with pH Paper

Demonstrations

1. None

Worksheets

1. pH Worksheet

Extra Readings

1. The pH at Which an Indicator Changes Color

Answers for The pH Concept (L3) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

20.4 Strength of Acids and Bases

Student Behavioral Objectives

The student will:

- describe the difference between a strong and weak acid.
- identify specific acids as strong or weak.
- define weak acids and bases.
- use K_a or K_b to find $[H^+]$ and vice versa.
- use K_a or K_b to find pH.

Timing, Standards, Activities

TABLE 20.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Strength of Acids and Bases</i>	2.5	1e, 5c, 5d, 5e, 5f

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Strong Acids and Bases Worksheet
2. Weak Acids and Bases Worksheet

Extra Readings

1. None

Answers for Strength of Acids and Bases (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

20.5 Bronsted-Lowry Acids and Bases

Student Behavioral Objectives

The student will:

- define Brønsted-Lowry acids and bases.
- identify Brønsted-Lowry acids and bases in chemical equations.
- define conjugate acids and bases.
- write the formula for the conjugate acid of any base and for the conjugate base of any acid.
- identify conjugate acids and bases in equations.
- given the strengths of acids and bases, identify the strength the conjugate acids and bases.

Timing, Standards, Activities

TABLE 20.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Bronsted-Lowry Acids and Bases</i>	2.0	1e, 5c, 5e

Activities for Lesson 5

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Conjugate Acids-Bases Worksheet

Extra Readings

1. None

Answers for Bronsted-Lowry Acids and Bases (L5) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

20.6 Lewis Acids and Bases

Student Behavioral Objectives

The student will:

- define a Lewis acid and a Lewis base.
- identify Lewis acids and bases in equations.

Timing, Standards, Activities

TABLE 20.7: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Lewis Acids and Bases</i>	1.0	5e

Activities for Lesson 6

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Lewis Acids and Bases (L6) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 20

This website provides a pH scale activity.

- http://www.quia.com/rd/1975.html?AP_rand=201033057

This website provides a titration demonstration.

- http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/stoichiometry/acid_base.html

This video discusses the properties of acids and bases.

- <http://www.youtube.com/watch?v=mm7Hcff5b6g>

A discussion of the difference between strong and weak acids is available at:

- <http://www.youtube.com/watch?v=XTdkWGImtSc>

Visit this website to learn more about examples and properties of acids and bases.

- <http://qldscienceteachers.tripod.com/junior/chem/acid.html>

The following website reviews the definitions of Arrhenius acids and bases.

- http://www.nyu.edu/classes/tuckerman/honors.chem/lectures/lecture_21/node2.html

The two websites below have more information about pH.

- <http://www.johnkyrk.com/pH.html>
- <http://purchon.com/chemistry/ph.htm>

This video shows an example of how to plug values into the K_a for a weak acid to find $[H^+]$ and pH.

- <http://www.youtube.com/watch?v=MOv7Z16FMK0>

The *learner.org* website allows users to view the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The video called “The Proton in Chemistry” explains pH and how it is measured, as well as the important role of acids and bases.

- http://www.learner.org/vod/vod_window.html?pid=808

The following link provides a review of Lewis acids and bases.

- http://www.nyu.edu/classes/tuckerman/honors.chem/lectures/lecture_21/node4.html

Laboratory Activities for Chapter 20

Measuring pH with pH Paper

Background

pH (Potential of Hydrogen) is a scale of acidity from 0 to 14. It tells how acidic or basic a substance is. The lower the pH, the more acidic is the solution. The higher the pH, the more basic the solution. Substances which are not acidic or basic (neutral) have a pH of 7. Acids have a pH less than 7. Bases have a pH greater than 7.

The formula for calculating pH is:

$$pH = -\log [H^+] \text{ where } [H^+] = \text{the concentration of } H^+ \text{ ions expressed in molarity.}$$

In this experiment you will be testing the pH of 5 different substances and determining whether they are acids, bases, or neither.

Pre-Lab Questions

1. If $[H^+]$ of a certain chemical is 1×10^{-14} , what is the chemical's pH?
2. If the pH of a certain chemical is 4, what is the $[H^+]$?
3. What is the pH of distilled water?

Purpose

- To gain understanding of pH and what it tells us about chemicals
- To learn how to measure pH with pH paper

Materials

- 5 Beakers/Plastic Cups
- 200 mL Distilled Water
- 200 mL Lemon Juice
- 200 mL Coca-Cola
- 200 mL Windex
- 200 mL Milk
- 5 Pieces of pH paper and a pH color comparison scale

Procedure

1. Make an estimate of the pH of each substance based on your prior knowledge and record this in your data table.
2. Take a strip of pH paper and place it about $\frac{1}{4}$ of the way into the first substance for about 2-3 seconds.
3. Take the strip and compare it to your pH color scale, recording the measured pH level in your data table.
4. Determine whether your substance is an acid or a base.
5. Repeat steps 1-4 with the other 4 substances.

Data Table

	Estimated pH	Measured pH	Acid or Base?
Water			
Lemon Juice			
Coca-Cola			
Windex			
Milk			

Post-Lab Calculations and Questions

1. Calculate the $[H^+]$ of each of these substances based on your measured pH.
2. Which of these substances was the most acidic? Which was the most basic?

Demonstrations for Chapter 20

20.7 Worksheets for Chapter 20

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

pH Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

1. Calculate the pH of a solution with $[H^+] = 7.0 \times 10^{-5} M$.
2. Calculate the pH of a solution that is $0.050 M NaOH$.
3. Calculate the pH of a solution that is $7.0 M 10^{-5} \times Mg(OH)_2$.
4. What is the $[H^+]$ in a solution with $pH = 4.4$?
5. What is the $[OH^-]$ in a solution with $pH = 3.0$?
6. $10.0 g$ of KOH is added to enough water to make $400. mL$ of solution. What is the pH ?
7. A $1.0 liter$ solution has a $pH = 2$. How many liters of water must be added to change the pH to 3?
8. If you do the regular calculations to determine the pH of a $1.0 \times 10^{-12} M HBr$ solution, you will get the $pH = 12$. You should have a feeling that something is wrong with this situation because this indicates that a solution of acid has a basic pH . What do you think is wrong with this calculation?

Complete the following table.

TABLE 20.8: Acid, Base, or Neutral

	pH	$[H^+]$	$[OH^-]$	A, B, or N
9.		$6.2 \times 10^{-4} M$		
10.			$8.5 \times 10^{-10} M$	
11.	10.75			
12.			$4.0 \times 10^{-2} M$	

Strong Acids and Bases Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

1. If the hydrogen ion concentration in a solution is $1.00 \times 10^{-4} M$, what is the hydroxide ion concentration?
2. What is the hydroxide ion concentration in a solution whose pH is 11?
3. What is the hydrogen ion concentration in a solution prepared by dissolving $0.400 grams$ of $NaOH$ in enough water to make $2.00 liters$ of solution?
4. How many mL of $0.100 M$ potassium hydroxide are required to neutralize $75.0 mL$ of $0.500 M HNO_3$?
5. If $50.0 mL$ of H_2SO_4 are neutralized by $100. mL$ of $0.200 M LiOH$, what is the molarity of the H_2SO_4 ?
6. What volume of $6.00 M HCl$ would be necessary to neutralize $400. mL$ of $3.00 M Ba(OH)_2$?

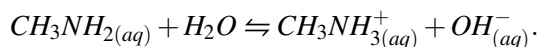
200. mL of 0.0150 M NaOH is mixed with 300. mL of 0.00100 M HCl. What is the final $[H^+]$ and $[OH^-]$?
- What is the pH of the final solution in problem 7?
700. mL of 1.00×10^{-4} M H_2SO_4 is mixed with 300. mL of 1.00×10^{-3} M $Ba(OH)_2$. What is the final $[H^+]$ and $[OH^-]$?
- What is the pH of the final solution in problem 9?
- 25.0 mL of 0.0100 M HCl is mixed with 35.0 mL of 0.0300 M NaOH. What is the final $[H^+]$ and $[OH^-]$?
- What is the pH of the final solution in problem 11?
- What is the final $[H^+]$ and $[OH^-]$ in a solution made by adding 100. mL of 0.000200 M HNO_3 to 100. mL of 0.0000990 M $Ba(OH)_2$?
- What is the pH of the final solution in problem 13?
- What is the molar mass of a solid monoprotic acid if 0.300 grams of the acid requires 30.0 mL of 0.200 M NaOH to neutralize it?

Weak Acids and Bases Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

- Explain the difference between the designations “strong” acid and “weak” acid.
- The K_a of acid A is 6.4×10^{-4} and the K_a of acid B is 1.7×10^{-5} . Which acid is the stronger acid?
- Explain what happens to the pH of a solution of acetic acid when a solution of sodium acetate is added to it.
- Explain why a solution of sodium acetate will be basic.
- What is the pH of a 0.0100 M solution of a weak acid, HX, if the K_a for HX is 8.1×10^{-7} .
- The pH of a 0.100 M solution of a weak acid, HQ, is 4.0. What is the K_a of this acid?
- What is the pH of a 0.150 M solution of NH_4OH ? The K_b for NH_4OH is 1.80×10^{-5} .
- The pH of a 1.00 M solution of the weak base methylamine is 12.3. The equation for the reaction of methylamine in water is



What is the K_b for methylamine?

- Will a 1.00 M solution of potassium acetate be acidic, basic, or neutral?
- Will a 1.00 M solution of NH_4NO_2 be acidic, basic, or neutral? Use 1.8×10^{-5} as the K_b for NH_4OH and 7.1×10^{-4} as the K_a for HNO_2 .

Conjugate Acids-Bases Worksheet

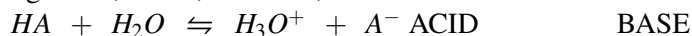
CK12 Foundation Chemistry

Name _____ Date _____

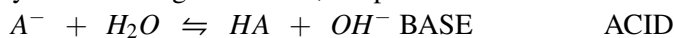
Conjugate Acid-Base Pairs

Acids and bases exist as conjugate acid-base pairs. The term conjugate come from the Latin meaning “joining together” and refers to things that are joined, particularly in pairs, such as Bronsted acids and bases.

Every time a Bronsted acid acts as a hydrogen ion donor, it forms a conjugate base. Imagine a generic acid, HA. When this acid donate a hydrogen to water, one product of the reaction is the A⁻ ion, which is then capable of accepting a hydrogen ion, and is, therefore, a Bronsted base.



Conversely, every time a base gains an H⁺, the product is a Bronsted acid.



Acids and bases in the Bronsted model, therefore, exist as conjugate pairs whose formulas are related by the gain or loss of a hydrogen ion.

The use of the symbols HA and A⁻ for the conjugate acid-base pair does not mean that all acids are neutral molecules or that all bases are negative ions. It signified only that the acid contains a hydrogen ion that is not present in the conjugate base. Bronsted acids or bases can be neutral molecules, positive ions, or negative ions.

Exercises

- Write the formula for the conjugate base of each of the following acids.
 A. HCN . . . conjugate base = _____ B. HSO₄⁻ . . . conjugate base = _____ C. HF . . . conjugate base = _____
 D. HNO₂ . . . conjugate base = _____
- Write the formula for the conjugate acid of each of the following bases.
 A. NH₃ . . . conjugate acid = _____ B. HCO₃⁻ . . . conjugate acid = _____ C. HS⁻ . . . conjugate acid = _____
 D. Br⁻ . . . conjugate acid = _____
- For each given formula, indicate whether its conjugate partner is an acid or base and write its formula.
 A. SO₄²⁻ . . . conjugate _____ = _____ B. HI . . . conjugate _____ = _____ C. S²⁻ . . . conjugate _____ = _____
 D. HNO₃ . . . conjugate _____ = _____
- In each of the following acid-base reactions, identify the acid and base on the left and their conjugate partners on the right.
 A. CHOOH(aq) + H₂O(L) ⇌ HCOO⁻(aq) + H₃O⁺(aq) B. H₂S(aq) + NH₃(aq) ⇌ NH₄⁺(aq) + HS⁻(aq)

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

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Answers for pH Worksheet

- Calculate the pH of a solution with $[H^+] = 7.0 \times 10^{-5} M$.

$$pH = -\log [H^+] = -\log (7.0 \times 10^{-5}) = 4.15$$

- Calculate the pH of a solution that is 0.050 M NaOH.

$$\text{Since } [OH^-] = 0.050 \text{ M then, } [H^+] = \frac{1.0 \times 10^{-14}}{0.050} = 2.0 \times 10^{-13} \text{ M}$$

$$pH = -\log [H^+] = -\log (2.0 \times 10^{-13}) = 12.7$$

3. Calculate the pH of a solution that is $7.0 \times 10^{-5} \text{ M Mg(OH)}_2$.

$$\text{If } [Mg(OH)_2] = 7.0 \times 10^{-5} \text{ M, then } [OH^-] = (2)(7.0 \times 10^{-5} \text{ M}) = 1.4 \times 10^{-4} \text{ M}$$

H^+

$$= \frac{1.0 \times 10^{-14}}{1.4 \times 10^{-4}} = 7.1 \times 10^{-11} \text{ M}$$

$$pH = -\log [H^+] = -\log (7.1 \times 10^{-11}) = 10.1$$

4. What is the $[H^+]$ in a solution with $pH = 4.4$?

$$[H^+] = 10^{-pH} = 10^{-4.4} = 3.98 \times 10^{-5} \text{ M}$$

5. What is the $[OH^-]$ in a solution with $pH = 3.0$?

$$\text{If } pH = 3.0, \text{ then } [H^+] = 1.0 \times 10^{-3} \text{ M and } [OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-3}} = 1.0 \times 10^{-11} \text{ M}$$

6. 10.0 g of KOH is added to enough water to make 400. mL of solution. What is the pH ?

$$\text{mols } KOH = \frac{10.0 \text{ g}}{56.1 \text{ g/mol}} = 0.178 \text{ mol}$$

KOH

$$= \frac{0.178 \text{ mol}}{0.400 \text{ L}} = 0.446 \text{ M}$$

H^+

$$= \frac{1.0 \times 10^{-14}}{0.446} = 2.2 \times 10^{-14} \text{ M}$$

$$pH = -\log [H^+] = -\log (2.2 \times 10^{-14}) = 13.6$$

7. A 1.0 liter solution has a $pH = 2$. How many liters of water must be added to change the pH to 3?

$$\text{At } pH = 2, [H^+] = 1.0 \times 10^{-2} \text{ M}$$

If $[H^+] = 1.0 \times 10^{-2} \text{ M}$, then 1.0 liter of solution contains 0.010 mol $[H^+]$.

$$\text{At } pH = 3, [H^+] = 1.0 \times 10^{-3} \text{ M}$$

$$\text{Total liters} = \frac{\text{mols}}{M} = \frac{0.010 \text{ mol}}{0.0010 \text{ mol/L}} = 10.0 \text{ liters}$$

Beginning with 1.0 liter and ending with 10.0 liters would require an additional 9.0 liters.

8. If you do the regular calculations to determine the pH of a $1.0 \times 10^{-12} \text{ M HBr}$ solution, you will get the $pH = 12$. You should have a feeling that something is wrong with this situation because this indicates that a solution of acid has a basic pH . What do you think is wrong with this calculation?

This is such a dilute solution of HBr that the contribution of $[H^+]$ from the dissociation of water is the predominant source of $[H^+]$. The pH of this solution will be nearly 7 because the contribution of hydrogen ion from the HBr is not significant.

Complete the following table.

TABLE 20.9: short caption

	pH	$[H^+]$	$[OH^-]$	A, B, or N
9.	3.2	$6.2 \times 10^{-4} M$	$1.6 \times 10^{-11} M$	A
10.	4.9	$1.2 \times 10^{-5} M$	$8.5 \times 10^{-10} M$	A
11.	10.75	$5.6 \times 10^{-10} M$	$1.8 \times 10^{-5} M$	B
12.	12.6	$2.5 \times 10^{-13} M$	$4.0 \times 10^{-2} M$	B

Answers for Strong Acids and Bases Worksheet

1. If the hydrogen ion concentration in a solution is $1.00 \times 10^{-4} M$, what is the hydroxide ion concentration?

$$[H^+][OH^-] = 1.00 \times 10^{-14}$$

OH^-

$$= \frac{1.00 \times 10^{-14}}{[H^+]} = \frac{1.00 \times 10^{-14}}{1.00 \times 10^{-4}} = 1.00 \times 10^{-10} M$$

2. What is the hydroxide ion concentration in a solution whose pH is 11?

$$[H^+] = 10^{-pH} = 1.00 \times 10^{-11} M$$

OH^-

$$= \frac{1.00 \times 10^{-14}}{[H^+]} = \frac{1.00 \times 10^{-14}}{1.00 \times 10^{-11}} = 1.00 \times 10^{-3} M$$

3. What is the hydrogen ion concentration in a solution prepared by dissolving 0.400 grams of $NaOH$ in enough water to make 2.00 liters of solution?

$$\text{moles } NaOH = \frac{0.400 \text{ grams}}{40.0 \text{ grams/mole}} = 0.0100 \text{ mole}$$

OH^-

$$= \frac{0.0100 \text{ mole}}{2.00 \text{ liter}} = 0.00500 M$$

H^+

$$= \frac{1.00 \times 10^{-14}}{0.00500 M} = 2.00 \times 10^{-12} M$$

4. How many mL of 0.100 M potassium hydroxide are required to neutralize 75.0 mL of 0.500 M HNO_3 ?

At neutrality, moles of hydrogen ion are equal to moles of hydroxide ion.

$$\begin{aligned} \text{moles hydrogen ion} &= (M)(L) = (0.500 M)(0.0750 L) = 0.0375 \text{ moles} \\ \text{moles hydroxide ion} &= (0.100 M)(x L) = 0.0375 \text{ moles} \\ x &= 0.375 \text{ liters} = 375 \text{ mL} \end{aligned}$$

5. If 50.0 mL of H_2SO_4 are neutralized by 100. mL of 0.200 M $LiOH$, what is the molarity of the H_2SO_4 ?

$$\begin{aligned} \text{moles } OH^- &= (0.200 M)(0.100 L) = 0.0200 \text{ moles} \\ \text{moles } H^+ &= (x M)(0.0500 L)(2) = (0.100)(x) \\ \text{moles } H^+ &= \text{moles } OH^- \\ (0.100)(x) &= 0.0200 \\ x &= 0.200 M \end{aligned}$$

6. What volume of 6.00 M HCl would be necessary to neutralize 400. mL of 3.00 M $Ba(OH)_2$?

$$\begin{aligned} \text{moles hydrogen ion} &= \text{moles hydroxide ion} \\ (6.00 M)(x L) &= (3.00 M)(0.400 L)(2) \\ x &= 0.400 L \end{aligned}$$

7. 200. mL of 0.0150 M $NaOH$ is mixed with 300. mL of 0.00100 M HCl . What is the final $[H^+]$ and $[OH^-]$?

$$\begin{aligned} \text{moles } H^+ &= (0.00100 M)(0.300 L) = 0.000300 \text{ moles} \\ \text{moles } OH^- &= (0.0150 M)(0.200 L) = 0.00300 \text{ moles} \\ \text{excess moles of } OH^- &= 0.00300 \text{ moles} - 0.000300 \text{ moles} = 0.00270 \text{ moles} \end{aligned}$$

OH^-

$$= \frac{0.00270 \text{ moles}}{0.500 \text{ liters}} = 0.00540 M$$

H^+

$$= \frac{1.00 \times 10^{-14}}{0.00540} = 1.85 \times 10^{-10} M$$

8. What is the pH of the final solution in problem 7?

$$pH = -\log [H^+] = -\log (1.85 \times 10^{-10} M) = 11.7$$

9. 700. mL of $1.00 \times 10^{-4} M H_2SO_4$ is mixed with 300. mL of $1.00 \times 10^{-3} M Ba(OH)_2$. What is the final $[H^+]$ and $[OH^-]$?

$$\begin{aligned} \text{moles } H^+ &= (2)(1.00 \times 10^{-4} M)(0.700 L) = 1.40 \times 10^{-4} \text{ moles} \\ \text{moles } OH^- &= (2)(1.00 \times 10^{-3} M)(0.300 L) = 6.00 \times 10^{-4} \text{ moles} \\ \text{excess moles of } OH^- &= 6.00 \times 10^{-4} M - 1.40 \times 10^{-4} M = 4.60 \times 10^{-4} \text{ moles} \end{aligned}$$

OH^-

$$= \frac{4.60 \times 10^{-4} \text{ moles}}{1.00 \text{ liter}} = 4.60 \times 10^{-4} \text{ M}$$

H⁺

$$= \frac{1.00 \times 10^{-14}}{4.60 \times 10^{-4} \text{ M}} = 2.17 \times 10^{-11} \text{ M}$$

10. What is the *pH* of the final solution in problem 9?

$$pH = -\log [H^+] = -\log (2.17 \times 10^{-11} \text{ M}) = 10.7$$

11. 25.0 mL of 0.0100 M *HCl* is mixed with 35.0 mL of 0.0300 M *NaOH*. What is the final $[H^+]$ and $[OH^-]$?

$$\text{moles } H^+ = (0.0100 \text{ M})(0.0250 \text{ L}) = 0.000250 \text{ moles}$$

$$\text{moles } OH^- = (0.0300 \text{ M})(0.0350 \text{ L}) = 0.00105 \text{ moles}$$

$$\text{excess moles of } OH^- = 0.000800 \text{ moles}$$

OH⁻

$$= \frac{0.000800 \text{ moles}}{0.0600 \text{ liter}} = 0.133 \text{ M}$$

H⁺

$$= \frac{1.00 \times 10^{-14}}{0.133 \text{ M}} = 7.52 \times 10^{-14} \text{ M}$$

12. What is the *pH* of the final solution in problem 11?

$$pH = -\log [H^+] = -\log (7.52 \times 10^{-14} \text{ M}) = 13.1$$

13. What is the final $[H^+]$ and $[OH^-]$ in a solution made by adding 100. mL of 0.000200 M *HNO₃* to 100. mL of 0.0000990 M *Ba(OH)₂*?

$$\text{moles } H^+ = (0.000200 \text{ M})(0.100 \text{ L}) = 0.0000200 \text{ moles}$$

$$\text{moles } OH^- = (2)(0.0000990 \text{ M})(0.100 \text{ L}) = 0.0000198 \text{ moles}$$

$$\text{excess moles of } H^+ = 2.00 \times 10^{-7} \text{ moles}$$

H⁺

$$= \frac{2.00 \times 10^{-7} \text{ moles}}{0.200 \text{ liter}} = 1.00 \times 10^{-6} \text{ M}$$

OH⁻

$$= \frac{1.00 \times 10^{-14}}{1.00 \times 10^{-6} \text{ M}} = 1.00 \times 10^{-8} \text{ M}$$

14. What is the *pH* of the final solution in problem 13?

$$pH = -\log [H^+] = -\log (1.00 \times 10^{-6} \text{ M}) = 6.00$$

15. What is the molar mass of a solid monoprotic acid if 0.300 grams of the acid requires 30.0 mL of 0.200 M $NaOH$ to neutralize it?

$$\text{moles } OH^- = (0.200 \text{ M})(0.0300 \text{ L}) = 0.00600 \text{ moles}$$

Therefore, $\text{moles } H^+ = 0.00600 \text{ moles}$

$$\text{molar mass} = \frac{\text{grams}}{\text{moles}} = \frac{0.300 \text{ grams}}{0.00600 \text{ moles}} = 50.0 \text{ grams/mole}$$

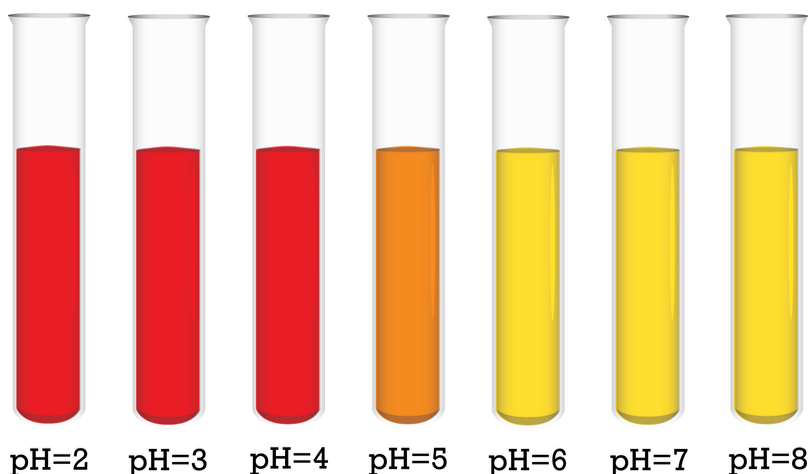
Answers for Weak Acids and Bases Worksheet

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

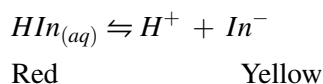
20.8 Extra Reading for Chapter 20

The pH at which an Indicator Changes Color

Many acid-base indicators exhibit exactly three colors. There is the color of the indicator when it is predominantly in its molecular form, the color of the indicator when it is predominantly in its ionic form, and there is the color of the indicator when it is close to 50% in each form. Consider a fictitious indicator, HIn , whose K_a is 1.0×10^{-5} . At pH values below 5, this indicator is distinctly red, at pH values above 5, it is distinctly yellow, and exactly at a pH of 5, the indicator is orange.



The dissociation equation for the indicator, HIn , is



When the hydrogen ion concentration is high, the equilibrium is shifted toward the reactants, most of the indicator particles are in the form of undissociated molecules, and the solution is red. When the hydrogen ion concentration is low, the equilibrium is shifted toward the products, most of the indicator particles are in the form of anions, and the solution is yellow. At some exact pH , the equilibrium will be adjusted so that exactly 50% of the indicator particles are in the form of undissociated molecules, and 50% in the form of anions. In this case, the solution will be a mixture of equal numbers of red molecules and yellow ions, hence will be orange.

Here is the equilibrium constant expression for the indicator.

$$K_a = \frac{[H^+][In^-]}{[HIn]} = 1.0 \times 10^{-5}$$

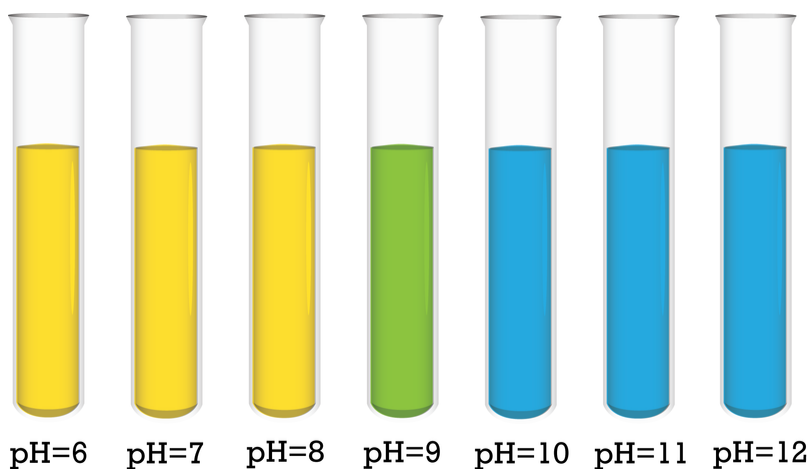
For the pH at which the color changes, we are seeking the point where half of the indicator particles are in each form; in other words, $[In^-] = [HIn]$. When these two values are exactly equal, they will cancel from the expression.

$$K_a = \frac{[H^+][In^-]}{[HIn]} = 1.0 \times 10^{-5}$$

As you can see, mathematically, the $[H^+]$ for this exact point will be equal to the K_a value and $pH = -\log(1.0 \times 10^{-5}) = 5$, which is in agreement with the pictures of the indicator colors at various pH s.

Consider the indicator thymol blue. The undissociated molecules of thymol blue are yellow and the anions are blue. The K_a for thymol blue is 1.0×10^{-9} . When this indicator is 50% in the form of undissociated molecules and 50% anions, the 50 – 50 mixture of yellow and blue would result in a green color. Calculations of the same type as shown for the previous example indicate that the green color will be present when the $[H^+]$ is equal to the value of the K_a , 1.0×10^{-9} . Therefore, the color change pH for thymol blue is $pH = 9$. When the pH value is less than 9, the indicator will be yellow, at exactly 9, it will be green, and above 9, it will be blue.

It should be clear that putting a few drops of thymol blue in a solution and getting a resultant yellow color does not tell you the pH of the solution. It only tells you that the pH is less than 9. Similarly, a resulting blue solution of thymol blue only tells you that the pH is greater than 9. There is only one color of a thymol blue solution that tells you the pH and that is green.



20.9 Assessment for Chapter 20

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER

21**Neutralization C-SE-TE****Chapter Outline**

- 21.1 NEUTRALIZATION**
- 21.2 TITRATION**
- 21.3 BUFFERS**
- 21.4 WORKSHEETS FOR CHAPTER 21**
- 21.5 EXTRA READING FOR CHAPTER 21**
- 21.6 ASSESSMENT FOR CHAPTER 21**

Lessons and Number of Activities for Lessons**TABLE 21.1: Lessons and Activities for Lessons**

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Neutralization	1	0	1	1
2. Titration	2	0	1	0
3. Buffers	0	0	1	0

21.1 Neutralization

Student Behavioral Objectives

The student will:

- explain what is meant by a neutralization reaction and give an example of one.
- write a balanced equation for the reaction that occurs when an acid reacts with a base.
- describe the formation of a salt in terms of the Arrhenius definitions of acids and bases.
- predict the salt that will be produced from the neutralization reaction between a given acid and base.
- identify acidic, basic, and neutral salts from a neutralization reaction.

Timing, Standards, Activities

TABLE 21.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Neutralization</i>	1.0	5a, 5d

Activities for Lesson 1

Laboratory Activities

1. Hydrolysis of Salts

Demonstrations

1. None

Worksheets

1. Hydrolysis of Salts

Extra Readings

1. Acid Rain

Answers for Neutralization (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

21.2 Titration

Student Behavioral Objectives

The student will:

- explain what an acid/base indicator is.
- explain how acid-base indicators work.
- explain the difference between natural and synthetic indicators.
- explain how indicators are used in the lab.
- explain what a titration is.
- describe how titrations can be used to determine the concentration of an acid or a base in solution.
- explain the difference between the equivalence point and the end point.
- define a standard solution in terms of acid-base titrations.
- calculate the concentration of an acid or base solution using a standard solution.
- calculate the concentration of unknown acid or base when given the concentration of the other and the volume needed to reach the equivalence point in a titration.

Timing, Standards, Activities

TABLE 21.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Titration</i>	2.0	1b, 5d

Activities for Lesson 2

Laboratory Activities

1. pH Measurements Using Indicators
2. Acid-Base Titration

Demonstrations

1. None

Worksheets

1. Titration Worksheet

Extra Readings

1. None

Answers for Titration (L2) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

21.3 Buffers

Student Behavioral Objectives

The student will:

- define and give an example of a buffer.
- explain the effect of a strong acid or base on a buffer system.
- explain the mechanism by which a buffer solution resists changes in pH.
- given appropriate information, calculate the pH of a buffer.
- describe how to make a buffer solution.

Timing, Standards, Activities

TABLE 21.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Buffers</i>	2.0	5g

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Buffers Worksheet

Extra Readings

1. None

Answers for Buffers (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 21

This website provides a lesson on buffer solutions.

- <http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/buffer12.swf>

This website provides a list of household acid-base indicators.

- <http://antoine.frostburg.edu/chem/senese/101/acidbase/faq/household-indicators.shtml>

This website provides a pH calculation problem generator.

- <http://science.widener.edu/svb/tutorial/phcalcs.html>

This website provides an animation of the autoionization of water.

- <http://chemmovies.unl.edu/ChemAnime/AUTOWD/AUTOWD.html>

The following link shows a video of a neutralization reaction.

- http://www.youtube.com/watch?v=_P5hGzA6Vb0

This video shows the technique for performing a titration using an indicator.

- <http://www.youtube.com/watch?v=9DkB82xLvNE>

The following link is to a video about acid-base neutralization and titration.

- <http://link.brightcove.com/services/player/bcpid9113583001?bctid=1405713919>

The video at the link below shows the lab techniques needed for titration.

- <http://chem-ilp.net/labTechniques/TitrationVideo.htm>

This video is a ChemStudy film called “Acid Base Indicators.” The film is somewhat dated but the information is accurate.

- <http://www.youtube.com/watch?v=yi8QrjmV6Sw>

The following link is to a video lecture about acid-base buffers.

- http://www.youtube.com/watch?v=O_QlZe4fv4g

To see a short animated video showing concentration changes as strong acid or base is added to a buffer, follow the link below.

- <http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/buffer12.swf>

Laboratory Activities for Chapter 21

Teacher's Pages for Hydrolysis of Salts

Lab Notes

Preparation of Solutions:

200 mL of each solution should be more than enough to complete this lab. To prepare each solution, mass the specified amount of reagent, dissolve it in 150 mL of water, and dilute the resulting solution to 200 mL:

$CuSO_4 \cdot 5H_2O$: 5.0 grams

$Ca(NO_3)_2$: 3.3 grams

K_3PO_4 : 4.2 grams

KCl : 1.5 grams

$NaBr$: 2.1 grams

Na_2S : 1.6 grams

$(NH_4)_2CO_3$: 1.9 grams

Na_2CrO_4 : 3.2 grams

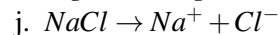
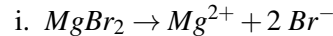
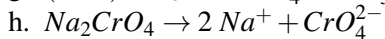
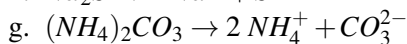
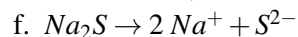
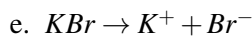
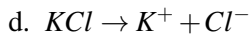
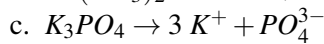
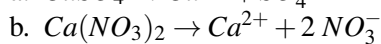
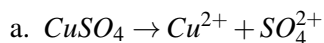
$MgBr_2$: 3.7 grams

$NaCl$: 1.2 grams

Use care when opening the container of $(NH_4)_2CO_3$. It undergoes decomposition over time, and outgases NH_3 , which collects in the container. Upon opening the NH_3 diffuses out rapidly, and is very irritating to eyes and skin. Live and learn.

Answers to Pre-Lab Questions

1. Write dissociation equations for the following salts:



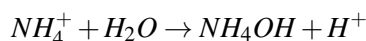
2. $NaC_6H_5CO_2 + H_2O \rightarrow HC_6H_5CO_2 + Na^+ + OH^-$ solution will be basic

Lab – Hydrolysis of Salts

Background Information

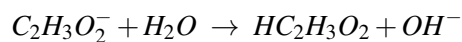
A salt is an ionic compound containing positive ions other than hydrogen and negative ions other than hydroxide. Most salts will dissociate to some degree when placed in water. In many cases, ions from the salt will react with water molecules to produce hydrogen ions, H^+ , or hydroxide ions, OH^- . Any chemical reaction in which water is one of the reactants is called a hydrolysis reaction. Salts are usually formed from the neutralization reaction between an acid and a base. A salt formed from a strong acid and a strong base will not undergo hydrolysis. The resulting solution is neutral. An example of such a salt is KBr , formed from a strong acid, HBr , and a strong base, KOH .

Salts formed from the reaction of a strong acid and a weak base hydrolyzes to form a solution that is slightly acidic. In this kind of hydrolysis, the water molecules actually react with the cation from the weak base. For example, when ammonium chloride, NH_4Cl , hydrolyzes, water molecules react with the NH_4^+ ion:



The formation of the H^+ ion from this reaction makes the solution acidic.

Salts formed from the reaction of a weak acid and a strong base hydrolyze to form a solution that is slightly basic. In this kind of hydrolysis, it is the anion from the weak acid that actually reacts with the water. For example, when sodium acetate, $NaC_2H_3O_2$, hydrolyzes, water molecules react with the acetate ion:



The formation of the OH^- ion from this reaction makes the solution basic. Salts formed from a weak acid and weak base produce solutions that may be slightly acidic, slightly basic, or neutral, depending on how strongly the ions of the salt are hydrolyzed.

In this experiment you will test several different salt solutions with pH paper and phenolphthalein solution to determine their acidity or basicity.

Purpose

To determine the relative acidity or basicity of various salt solutions, and thus predict whether hydrolysis occurred, and if so, what the reaction products are.

Pre-Lab Questions

- Write dissociation equations for the following salts:
 - Copper(II) sulfate
 - Calcium nitrate
 - Potassium phosphate
 - Potassium chloride
 - Potassium bromide
 - Sodium sulfide
 - Ammonium carbonate
 - Sodium chromate
 - Magnesium bromide
 - Sodium chloride
- Sodium benzoate is the salt formed in the neutralization of benzoic acid with sodium hydroxide. Benzoic acid is a weak acid. Write the hydrolysis reaction for the dissolution of solid sodium benzoate, $NaC_6H_5CO_2$, in water. Will sodium benzoate solution be acidic, basic, or neutral?

Apparatus and Materials

- 10 small or medium sized test tubes, or a micro reaction plate
- Test tube rack
- 0.1 *M* solutions of cupric sulfate, calcium nitrate, potassium phosphate, potassium chloride, sodium bromide, sodium sulfide, ammonium carbonate, sodium chromate, magnesium bromide and sodium chloride
- Universal *pH* indicator paper, range 0-14
- Phenolphthalein indicator solution
- 10 *mL* graduate
- Stirring rod

Safety Issues

The solutions used may be slightly acidic or basic, and as a result can be corrosive or caustic. Use proper laboratory safety equipment and techniques.

Procedure

1. Obtain a clean, dry micro reaction plate, or 10 test tubes
2. To test tubes 1 through 10 or the reaction plate, add eight to ten drops of the following solutions:
 - Tube or Well 1: Cupric sulfate
 - Tube or Well 2: Calcium nitrate
 - Tube or Well 3: Potassium phosphate
 - Tube or Well 4: Potassium chloride
 - Tube or Well 5: Sodium bromide
 - Tube or Well 6: Sodium sulfide
 - Tube or Well 7: Ammonium carbonate
 - Tube or Well 8: Sodium chromate
 - Tube or Well 9: Magnesium bromide
 - Tube or Well 10: Sodium chloride
3. Add two drops of phenolphthalein solution to each of the occupied wells of the microplate or test tube. Record your observations in the data table.
4. Test each solution with *pH* paper and record your results.

Data

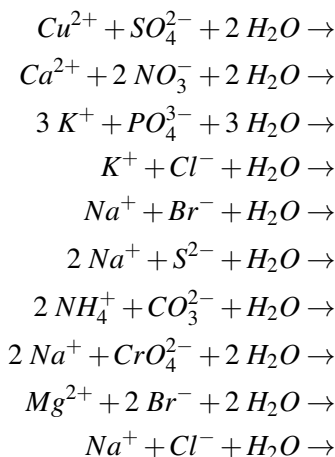
TABLE 21.5: Data Table

Well #	Salt	Effect on <i>pH</i> Indicator	Original Acid	Strong or Weak Acid	Original Base	Strong or Weak Base
1						
2						
3						
4						
5						
6						
7						
8						
9						
10						

Post-Lab Questions

1. Where a hydrolysis is likely to occur in each of the following, write a net ionic hydrolysis equation. If no

hydrolysis is likely, write NR.



- How do your observations and *pH* readings compare with the expected results based on the equations for the hydrolysis reactions?
- What is a spectator ion? Name the spectator ions present in each hydrolysis reaction in this experiment.
- A salt formed from a strong acid and a strong base produces a neutral solution. A salt of a weak acid and a weak base may or may not produce a neutral solution. Explain why.
- Bases make effective cleaning agents, because they can convert grease and oils to a water soluble substance. Trisodium phosphate (TSP) is a common commercially available cleaner. Give the reaction TSP undergoes to create a basic solution.

Teacher's Pages for pH Measurements Using Indicators

Lab Notes

Buffered solutions of various *pH* values can be purchased in dropper bottles, as can dropper bottles of indicator solutions. If you have many chemistry classes and perform the experiment for several years, it may be more economical to prepare the solutions yourself.

The solutions used in this lab can be prepared as follows.

- pH* = 1 solution: dilute 8.3 mL of concentrated HCl(12 M) to 1.00 liter
- pH* = 3 solution: dilute 10. mL of *pH* = 1 solution (above) to 1.00 liter
- pH* = 5 solution: dilute 10. mL of *pH* = 3 solution (above) to 1.00 liter
- pH* = 7 solution: distilled water
- pH* = 13 solution: dissolve 4.00 g of NaOH in sufficient water to produce 1.00 liter of solution
- pH* = 11 solution: dilute 10. mL of *pH* = 13 solution (above) to 1.00 liter
- pH* = 9 solution: dilute 10. mL of *pH* = 11 solution (above) to 1.00 liter
- methyl orange indicator: dissolve 0.1 g of methyl orange powder in 100 mL of water and filter
- bromthymol blue indicator: dissolve .01 g of bromthymol blue in 100 mL of 50% water and 50% ethanol solution and filter
- phenolphthalein solution: dissolve 1.0 g of phenolphthalein powder in 100 mL of ethanol

To prepare the unknown solutions for Part IV, select three of the known *pH* solutions used in the lab and label them as unknown. Be sure to keep a record of which *pH* values were selected as unknowns.

Answers to Pre-Lab Questions

1. How is universal indicator made?

Several indicators are mixed.

2. What distinguishes weak organic acids that are useful as acid-base indicators from weak organic acids that will not function as acid-base indicators?

The undissociated molecule of the acid and the anion of the dissociated acid must be different colors.

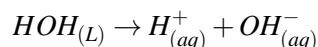
pH Measurements Using Indicators

Background Information

The nature of acids and bases have been known to man for quite sometime. Chemically speaking, acids are interesting compounds because a large number of common household substances are acids or acidic solutions. For example, vinegar contains ethanoic acid, also called acetic acid, $HC_2H_3O_2$ and citrus fruit contain citric acid. Acids cause foods to have a sour taste and turn litmus red. (Note: You should never taste substances in the laboratory.) Also, many common household substances are bases. Milk of magnesia contains the base magnesium hydroxide, $Mg(OH)_2$ and household ammonia is a common cleaning agent. Bases have a slick feel to the fingers and turn litmus blue. (Note: You should never feel chemicals in the laboratory.)

Indicator dyes, of which litmus is one, turn various colors according to the strength of the acid or base applied to it.

Pure water, which is neutral in terms of acid-base, exists mostly as H_2O molecules but does, to a very slight extent dissociate into hydrogen and hydroxide ions.



The extent of this dissociation is 1.0×10^{-7} moles/liter (at $25^\circ C$). Therefore, in all neutral water (and neutral water solutions), the concentration of hydrogen ions is $1.0 \times 10^{-7} M$ and the concentration of hydroxide ions is $1.0 \times 10^{-7} M$. The dissociation constant for this process is $K_w = [H^+][OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7} = 1.0 \times 10^{-14})$.

In 1909, a Danish chemist (Soren Sorenson), developed a mathematical system for referring to the degree of acidity of a solution. He used the term pH for “power of hydrogen” and established the equation, $pH = -\log [H^+]$.

In a neutral solution, the hydrogen ion concentration is $1.0 \times 10^{-7} M$ and therefore, the pH is 7. If the concentration of hydrogen ions is $1.0 \times 10^{-5} M$, then the pH is 5. A solution is neutral when the pH equals 7, it is acid if the pH is less than 7, and it is basic if the pH is more than 7. In commonly used solutions, pH values usually range from 1 to 14.

Living matter (protoplasm) contains a mixture of variously dissociated acids, bases, and salts and usually has a pH very near neutral. The pH of human blood is generally 7.3 and humans cannot survive if the blood becomes more basic than pH 7.8 or more acidic than pH 7.0. Life of any kind exists only between pH 3 and pH 8.5. Buffer solutions regulate the pH of the body by neutralizing excess acid or base. The chief buffers of the body are proteins, carbonates, phosphates, and hemoglobin. The kidneys play a role by eliminating excess electrolytes.

TABLE 21.6: Some Typical

pH	Substance	Acidity/Basicity
0	Sulfuric Acid (Battery Acid)	Very Highly Acidic
1	0.10 M Hydrochloric Acid	Highly Acidic
2	Stomach Acid	Acidic

TABLE 21.6: (continued)

pH	Substance	Acidity/Basicity
3	Vinegar	Acidic ($\frac{1}{100}^{th}$ as strong as pH 1)
4	Tomato Juice	
5	Black Coffee and Vitamin C	Weakly Acidic
6	Cow's Milk	Very Weakly Acidic
7	Distilled Water	Neutral
8	Sea Water	Very Weakly Basic
9	Baking Soda, $NaHCO_3$	Weakly Acidic
10	Detergents	
11		Basic
12	Household Cleaning Ammonia	
13	0.10 $MNaOH$	Strongly Basic
14	1.0 $MNaOH$ (Lye)	Very Strongly Basic

Some acids dissociate completely into ions when dissolved in water. Such acids are called strong acids (HCl , HI , HBr , HNO_3 , H_2SO_4 , $HClO_4$). Some bases dissociate completely when dissolved in water. Such bases are called strong bases ($NaOH$, $LiOH$, KOH , $RbOH$). There are other acids and bases that dissociate only slightly (although completely soluble) when dissolved in water. Such acids and bases are called weak acids or weak bases and some examples are HF , $HC_2H_3O_2$, NH_4OH .

An important method of determining pH values in the lab involves the use of substances called “acid-base indicators”. These are certain organic substances (almost always weak organic acids) that have the property of changing color in solutions of varying hydrogen ion concentration. In order for a weak organic acid to be useful as an acid-base indicator, it is necessary that the undissociated molecule and the indicator anion be different colors. For example, phenolphthalein is a colorless substance in any aqueous solution in which the hydrogen ion concentration is greater than $1 \times 10^{-9} M$ ($pH < 9$) but changes to a red or pink color when the hydrogen ion concentration is less than $1 \times 10^{-9} M$ ($pH > 9$). Such substances can be used for determining the approximate pH of solutions. Electrical measurements can determine the pH even more precisely. This lab will use three acid base indicators and what is called a “universal indicator”.

TABLE 21.7: Some Indicator Color Changes

Indicator	pH Color Change Range	Color Change
Methyl Orange	3.1 – 3.4	Red to Yellow
Bromthymol Blue	6.0 – 7.6	Yellow to Blue
Phenolphthalein	8.3 – 10.0	Colorless to Red

The universal indicator (one type is called Bogen's Universal Indicator), is made by mixing a number of indicators that all change color at different pH s. As you slowly change the pH of the indicator from 1 to 14, it goes through a series of subtle color changes. The indicator is provided with a photographic chart that shows the color of the indicator at every different pH and the pH is identified by matching the indicator color to the chart.

Pre-Lab Questions

1. How is universal indicator made?
2. What distinguishes weak organic acids that are useful as acid-base indicators from weak organic acids that will not function as acid-base indicators?

Purpose

The purpose of this lab is to have the student experience the color changes involved with acid-base indicators and to identify the approximate pH of an unknown solution using acid-base indicators.

Apparatus and Materials

- Well Plates, at least 12 wells (1 per lab group)
- drop controlled bottles of $pH = 1$
- drop controlled bottles of $pH = 3$
- drop controlled bottles of $pH = 5$
- drop controlled bottles of $pH = 7$
- drop controlled bottles of $pH = 9$
- drop controlled bottles of $pH = 11$
- drop controlled bottles of $pH = 13$
- drop controlled bottles of methyl orange
- drop controlled bottles of bromthymol blue
- drop controlled bottles of phenolphthalein
- drop controlled bottles of universal indicator
- drop controlled bottle of unknown #1
- drop controlled bottle of unknown #2
- drop controlled bottle of unknown #3

Safety Issues

All solutions are irritating to skin, eyes, and mucous membranes. Handle solutions with care, avoid getting the material on you, and wash your hands carefully before leaving the lab.

Procedure for Part I: Determining the effect of pH on indicator dyes.

1. Place the Chemplate on a sheet of white paper.
2. Place one drop of methyl orange into cavities #1 and #2.
3. Place one drop of bromthymol blue in cavities #5 and #6.
4. Place one drop of phenolphthalein in cavities #9 and #10.
5. Carefully add one drop of pH 1 to cavities #1, #5, and #9.
6. Carefully add one drop of pH 13 to cavities #2, #6, and #10.

Data for Part I

1. What color is the original methyl orange solution? _____
2. What color is methyl orange in a strong acid? _____
3. What color is methyl orange in a strong base? _____
4. What color is the original bromthymol blue solution? _____
5. What color is bromthymol blue in a strong acid? _____
6. What color is bromthymol blue in a strong base? _____
7. What color is the original phenolphthalein solution? _____
8. What color is phenolphthalein in a strong acid? _____
9. What color is phenolphthalein in a strong base? _____

Rinse the Chemplate in tap water and dry with a paper towel.

Procedure for Part II: Determining the pH color change range of indicator dyes.

1. Place one drop of methyl orange in each cavity numbered 1 – 7.
2. Carefully add one drop of pH 1 to cavity #1, pH 3 to cavity #2, pH 5 to cavity #3, pH 7 to #4, pH 9 to #5, pH 11 to #6, and pH 13 to #7.
3. Repeat the rinsing, drying, and steps 1 and 2 except using bromthymol blue and then repeat the entire process again using phenolphthalein.

Data for Part II

1. Describe the color changes and the pH s around the color change pH for methyl orange.
2. Describe the color changes and the pH s around the color change pH for bromthymol blue.
3. Describe the color changes and the pH s around the color change pH for phenolphthalein.

Rinse the Chemplate in tap water and dry with a paper towel.

Procedure for Part III: Determining a color standard for universal indicator.

1. Place one drop of universal indicator in each cavity numbered 1 – 7.
2. Carefully add one drop of pH 1 to cavity #1, pH 3 to cavity #2, pH 5 to cavity #3, pH 7 to #4, pH 9 to #5, pH 11 to #6, and pH 13 to #7.

Keep these solutions for Part IV.**Data for Part III**

1. Describe the color of the universal indicator at each pH used.

Cavity #1 ($pH = 1$), color = _____

Cavity #2 ($pH = 3$), color = _____

Cavity #3 ($pH = 5$), color = _____

Cavity #4 ($pH = 7$), color = _____

Cavity #5 ($pH = 9$), color = _____

Cavity #6 ($pH = 11$), color = _____

Cavity #7 ($pH = 13$), color = _____

Procedure for Part IV: Determining the pH of some unknown solutions.

1. Place one drop of universal indicator in cavities #10, #11, and #12.
2. Place one drop of unknown #1 in cavity #10.
3. Place one drop of unknown #2 in cavity #11.
4. Place one drop of unknown #2 in cavity #12.
5. Compare the color in each cavity with the colors in cavities # 1 – 7 that you made in activity 3.

Data for Part IV

Color of unknown #1 in universal indicator _____

Color of unknown #2 in universal indicator _____

Color of unknown #3 in universal indicator _____

Post-Lab Questions

1. What is the pH of unknown #1? _____
2. What is the pH of unknown #2? _____
3. What is the pH of unknown #3? _____

Teacher's Pages for Acid-Base Titration**Lab Notes**

Students will need two days to do both parts of the lab.

Preparation of Solutions, KHP, and unknown acids

6 M NaOH solution: Boil 600 mL of distilled water to drive off any dissolved CO₂. (The CO₂ produces carbonic acid, which drives down the concentration of NaOH.) Add 120. g of NaOH to a 500 mL volumetric flask, and add some of the freshly boiled distilled H₂O to the flask. Swirl to dissolve. Cool the resultant solution in a cold water or ice water bath, let the solution and flask return to room temperature, and dilute the resulting solution to 500 mL. Store this solution in a tightly capped bottle, preferably Nalgene or other base-resistant bottle. This will provide enough solution for (75) two-student teams with a 20% excess for spills and endpoint over-runs.

Phenolphthalein indicator solution: Dissolve 0.1 g of phenolphthalein in 50 mL of 95% ethanol, and dilute to 100 mL by adding distilled water. This will provide enough solution for (75) two student teams.

KHP – Potassium Hydrogen Phthalate – Dry 200 g of KHP in a laboratory oven for at least one hour prior to titration. Store the KHP in a dessicator. KHP is slightly hygroscopic. This is sufficient KHP for (75) two – student teams.

Unknown Acids

The following acids are suggestions for use. Their number of ionizable hydrogens vary, are stable chemically, and many schools have them on hand. Twenty grams of each acid are required for (75) two – student teams:

TABLE 21.8: short caption

Acid	Number of Ionizable Hydrogens	Molar Mass (g/mol)
Lactic	1	90.1
Malonic	2	104.1
Maleic	2	116.1
Succinic	2	118.1
Benzoic	1	122.1
Salicylic	1	138.1
Tartaric	2	150.1

NOTE: The molecular weights listed are for the anhydrous acids. If hydrates are used, use the molecular weights as shown on the reagent bottle. The number of ionizable hydrogens does not change.

Answers to Pre-Lab Questions

- Pre-rinsing the buret will remove any water or residual NaOH solution within the buret. If there were water present, it would dilute the NaOH added. If there were residual NaOH, the concentration would increase upon addition of solution due to the crystalline or concentrated solution of NaOH present.
- The KHP and the NaOH react with each other in a 1:1 stoichiometric ratio. Thus, the number of moles of NaOH required to react with the KHP will be equal to the number of moles of KHP originally present.

$$\text{Moles of NaOH added} = M \times L = (0.1 \text{ M})(0.040 \text{ L}) = 0.004 \text{ moles NaOH added: thus}$$

$$\text{Moles KHP needed} = 0.004 \text{ moles}$$

$$\text{Grams of KHP needed} = (\text{moles})(\text{g/mole}) = (0.004 \text{ mol})(204.23 \text{ g/mol}) = 0.8 \text{ grams KHP}$$

$$3. \text{ Molarity of KHP} = \frac{\text{moles KHP}}{\text{liters of solution}}$$

$$\text{Moles of KHP} = \frac{\text{grams KHP}}{\text{g/mol of KHP}} = \frac{0.759 \text{ g}}{204.2 \text{ g/mol}} = 0.00372 \text{ mol KHP}$$

$$\text{Molarity} = \frac{0.00372 \text{ mol KHP}}{0.0500 \text{ L}} = 0.0743 \text{ M KHP}$$

4. Again, the amount of water needed to dissolve the *KHP* is not needed to solve this problem. All you need is the weight of the *KHP* and its molecular weight :

$$\text{Moles of } KHP = \frac{0.521 \text{ g}}{204.2 \text{ g/mol}} = 0.00255 \text{ mols } KHP$$

Since this will be equal to the number of moles of *NaOH* reacted, the number of liters of titrant you need to add can be calculated by re-arranging the molarity formula:

$$M = \frac{\text{moles}}{L}$$

$$\text{so } L = \frac{\text{moles}}{M} = \frac{0.00255 \text{ mol}}{0.102 \text{ mol/L}} = 0.0213 \text{ L} = 21.3 \text{ mL}$$

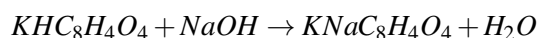
Acid-Base Titration Lab

Background Information

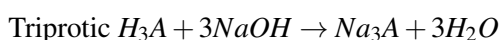
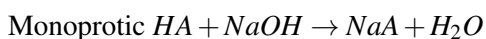
Standardization of the sodium hydroxide solution through titration is necessary because it is not possible to directly prepare a known molarity solution of sodium hydroxide with high accuracy. Solid sodium hydroxide readily absorbs moisture and carbon dioxide from the atmosphere and thus it is difficult to obtain a precise amount of the pure substance. A sodium hydroxide solution will be made close to 0.1 *M* and then the actual molarity of the solution will be determined by titration of a primary standard. A primary standard is a substance of very high purity that is also stable in air. Because the substance remains pure, it is possible to mass a sample of the substance with a high degree of accuracy. The primary standard used in this experimental procedure is potassium hydrogen phthalate, *KHC₈H₄O₄* (molar mass 204.2 *g/mole*).

This primary standard, *KHP*, is used to standardize the secondary standard, sodium hydroxide. The standardized sodium hydroxide solution can then be used to determine the molar mass of an unknown acid through titration.

In both steps a titration is performed in which a buret is used to dispense measured increments of the sodium hydroxide solution into a second solution containing a known mass of *KHP* (*NaOH* standardization). For the second reaction, a mass of acid whose molecular weight is unknown is then titrated with the solution of *NaOH* whose concentration was determined by the standardization with *KHP*. The stoichiometry of the reaction depends on the number of ionizable hydrogens within the acid. *KHP* is a weak monoprotic acid that will react with sodium hydroxide in a 1:1 mole ratio:



The unknown acid may be monoprotic, diprotic, or triprotic dependent on the number of acidic hydrogens present in the molecule. A monoprotic acid, *HA*, has one acidic hydrogen, a diprotic acid, *H₂A*, two acidic hydrogens, and a triprotic acid, *H₃A*, three acidic hydrogens. The stoichiometries of the reactions are shown below. You will be told whether your unknown acid is monoprotic, diprotic, or triprotic.



The indicator phenolphthalein is used as a signal of the equivalence point. Phenolphthalein is a weak organic acid that will change from colorless to pink near the equivalence point of the titrations. The actual point at which the indicator changes color is the end point. The endpoint and the equivalence point are not the same. The difference between the two is the titration error. Obviously, for a titration to be of value, care must be taken to select an indicator for which the difference between the equivalence point and the endpoint is small. With this particular titration, it is very small.

Pre-Lab Questions

1. Why is it necessary to rinse out the buret with the $NaOH$ solution?
2. Calculate the approximate weight of KHP required so that about 40 mL of 0.1 M sodium hydroxide is used in a titration. (MW of $KHP = 204.23 \text{ g/mol}$)
3. Calculate the molarity of a KHP solution when 0.759 g of KHP is dissolved in 50.0 mL of water.
4. 0.521 g of KHP is dissolved in 40 mL of water, and titrated with a 0.102 M $NaOH$ solution. Calculate the number of mL of the $NaOH$ solution added.

Purpose

The purpose of this experiment is to determine the concentration of a titrating solution, $NaOH$, using a stable compound, KHP . Once the concentration of the $NaOH$ solution is known, it can then be used to determine the molecular weight of an acid whose formula is unknown.

Apparatus and Materials

- 50 mL buret
- buret stand
- buret clamp
- 125 mL Erlenmeyer flask
- phenolphthalein
- $NaOH$ solution
- KHP , solid
- unknown acid, solid
- 10 mL graduate
- 400 mL beaker
- 100 mL graduate

Safety Issues

$NaOH$ is a caustic solution and will cause severe burns, especially to eye tissue. Wear goggles and aprons at all times. The solid acids cause considerable irritation if exposed to skin or mucous membranes. Avoid exposure.

Procedure for Part I

Part 1. Standardization of Sodium Hydroxide Solution

Obtain the primary acid standard from your instructor. Record the name of the acid, its molecular formula, and number of acidic hydrogens per molecule. Prepare about 300 mL of approximately 0.1 M sodium hydroxide by diluting 6 M $NaOH$ with distilled water. (Calculate, ahead of time, how much water and how much 6 M $NaOH$ will be needed.)

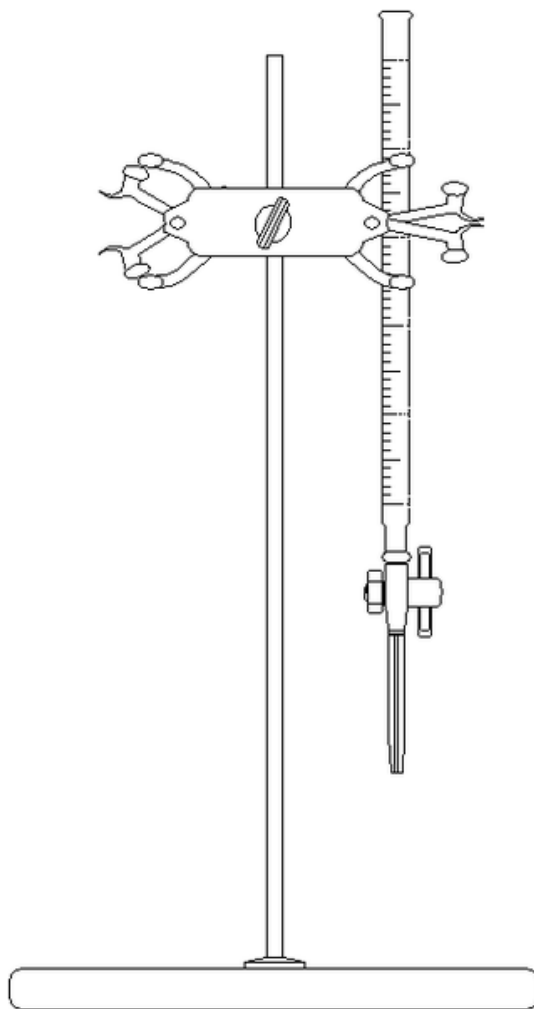
WARNING: Concentrated sodium hydroxide is corrosive and causes severe burns. Handle with care. Dilute and wash up spills with plenty of water. Wash affected skin with water until it no longer feels slippery, but feels “squeaky” clean.

Store the solution in your plastic bottle, label it “0.1 M $NaOH$ ”, and keep it tightly capped. You will determine the exact molarity of this $NaOH$ solution by standardization.

Calculate the mass of the primary acid standard that would react with about 20 mL of 0.1 M NaOH. Weigh approximately this amount into a clean 125 – mL Erlenmeyer flask by taring the balance with the flask on the pan, and then adding the acid to the flask. Record the mass of the primary acid standard to the highest precision allowed by the electronic balance. Add 30 to 40 mL of your purified water to the flask, and swirl to dissolve the primary acid standard. Add three or four drops of phenolphthalein indicator solution to the flask, and swirl to mix well. Label this flask and keep it tightly capped until ready for use.

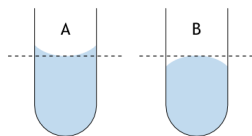
Rinse the inside of a CLEAN buret three times with small quantities of your 0.1 M sodium hydroxide solution (called “rinsing in” with the solution to be used in the buret). Drain the rinses through the stopcock and tip. Do not forget to rinse liquid through the tip, to replace water there. Fill the buret above the 0.0 mL mark with 0.1 M NaOH, and then drain it until the meniscus is slightly below

NOTE: Do not waste the time it takes to set the starting level to exactly 0.0 mL. It is more efficient and more accurate to set the level between 1 and 2 mL and read the starting level precisely.



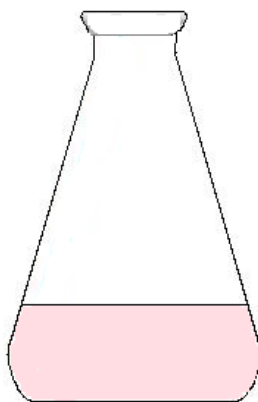
Tips on Technique

- To read the buret accurately, hold a white card with a black stripe behind the buret, with the black stripe below the meniscus, and the meniscus itself in front of the white region above the black stripe (see illustration). The meniscus will appear black against the white card. Keeping your eye level with the meniscus, read the buret.
- Remember to estimate one more digit than those marked on the scale.
- Remember that the buret scale reads increasing volume downward, not upward.



Tips on Technique

- Record the starting level to 0.01 *mL* precision (as always, estimate one more digit than marks indicate).
- Titrate the solution of primary acid standard with the 0.1 *M NaOH* until faint (see figure) phenolphthalein color appears and persists for 30 *seconds*. (Why might the color slowly disappear even after all acid is titrated?) Record the final buret reading to 0.01 *mL* precision.
- Mix the solution in the titration flask thoroughly after each addition of titrant, to ensure complete reaction before adding more.
- As you near the endpoint, wash the sides of the flask with distilled water to make sure that all delivered titrant is in solution.
- When you see that you are within a drop or two of the endpoint, split drops to avoid overshooting the endpoint.



Lab Procedure

Perform three titrations. For each, calculate the molarity of your *NaOH* solution. (From the mass of acid, and its molecular weight, you can calculate the number of moles of acid, which is equal to the number of moles of base you delivered. The molarity is found from the number of moles and the volume.)

When you have three values for the molarity of your *NaOH* solution, determine the average value.

TABLE 21.9: Data Table for Part I

	Trial 1	Trial 2	Trial 3
Initial Reading, <i>NaOH</i> buret	mL	mL	mL
Final Reading, <i>NaOH</i> buret	mL	mL	mL
Volume of <i>NaOH</i> added	mL	mL	mL
Grams of acid standard	g	g	g
Moles of acid standard	mol	mol	mol
Molarity of <i>NaOH</i>	M	M	M

Average molarity of *NaOH* = _____ M

Procedure for Part II**Part 2. Finding the Molar Mass of an Unknown Acid**

- Obtain a sample of a solid unknown from your instructor. Record its ID code in your report.
- Also, record the approximate amount of unknown to use in each titration, and the number of acid hydrogens per molecule. Your instructor will provide this information.
- Weigh the suggested amount into a clean 125 – mL erlenmeyer flask, by taring the balance with the flask on the pan, and then adding the acid to the flask. Record the mass of the sample to the precision allowed by the balance. Add 30 to 40 mL of distilled water and swirl to dissolve your sample. Add three to four drops of phenolphthalein indicator solution and swirl to mix well.
- Titrate your sample with your standardized NaOH solution until faint phenolphthalein color persists for 30 seconds.
*If your titration requires 10 to 25 mL of NaOH solution, carry out a second titration with an unknown sample of about the same mass. Otherwise, adjust the sample mass to bring the expected end-point volume to between 10 and 25 mL and do two more titrations.
- For each titration, compute the molar mass of the unknown acid, to the precision allowed by your data. Do three titrations and report the average molar mass for the solid acid.

Data for Part II

ID code of acid _____

Number of Acidic Hydrogens in Acid _____

Approximate mass of acid to be used _____ g

TABLE 21.10: short caption

	Trial 1	Trial 2	Trial 3
Mass of unknown acid sample	g	g	g
Volume of NaOH used	mL	mL	mL
Mols of NaOH used	moL	moL	moL
Moles of acid present ¹	mol	mol	mol
Molar mass of acid ²	g/mol	g/mol	g/mol
Molarity of NaOH	M	M	M

$$^1 \text{ Moles of acid present} = \frac{\text{moles NaOH}}{\text{number of } H^+ \text{ ions per acid molecule}}$$

$$^2 \text{ Molar mass of acid} = \frac{\text{grams acid}}{\text{moles acid}}$$

$$\text{Average Molar Mass} = \text{_____ g/mol}$$

Demonstrations for Chapter 21

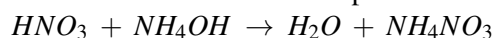
21.4 Worksheets for Chapter 21

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

The Mathematics of Salt Hydrolysis Worksheet

Part I

Reactions between acids and bases produce water and a salt.



In the reaction above, nitric reacts with ammonium hydroxide to produce water and ammonium nitrate. Ammonium nitrate is the salt in this reaction.

When salts are dissolved in water, the resulting solution could be neutral but it is also possible for the solution to be acidic or basic. Recall that the conjugate base of a strong acid, the anion in the acid, has virtually no affinity for hydrogen ions in water. That is why strong acids completely dissociate in aqueous solutions. Similarly, the cations of strong bases have no affinity for hydroxide ions in water.

The anions of weak acids, on the other hand, have significant affinity for hydrogen ions. Consider what happens when the salt sodium acetate is dissolved in water. The water, before the salt is added, has equal molarity of hydrogen and hydroxide ions. When the salt dissolves, the sodium ions (cation of strong base) will not attach to hydroxide ions but some of the acetate ion (anion of weak acid) will attach to hydrogen ions in the solution. When some of the hydrogen ions are thus removed from the solution, the resulting solution will have a higher molarity of hydroxide ions than hydrogen ions and the solution will be basic.

In the same way, when salts containing cations of weak bases and anions of strong acids are dissolved in water, the some of the cations of the weak base will attach to hydroxide ions. When hydroxide ions are removed from the solution, the solution will have excess hydrogen ions and the solution will be acidic.

The table below shows the strong acids and strong bases. All other acids and bases are weak.

TABLE 21.11: Strong Acids and Bases

Strong Acid	Formula	Strong Base	Formula
Hydrochloric Acid	HCl	Lithium Hydroxide	LiOH
Hydrobromic Acid	HBr	Sodium Hydroxide	NaOH
Hydroiodic Acid	HI	Potassium Hydroxide	KOH
Nitric Acid	HNO_3	Rubidium Hydroxide	RbOH
Perchloric Acid	HClO_4	Cesium Hydroxide	CsOH
Sulfuric Acid	H_2SO_4	Calcium Hydroxide	Ca(OH)_2
		Strontium Hydroxide	Sr(OH)_2
		Barium Hydroxide	Ba(OH)_2

The table below shows the acid or base properties of various salt solutions.

TABLE 21.12: Acidity of Salt Solutions

Parent Acid Strength	Parent Base Strength	Acidity/Basicity of Salt Solution
Strong	Strong	Neutral
Strong	Weak	Acidic
Weak	Strong	Basic
Weak	Weak	Depends on K_a and K_b

The following table shows the K_a and K_b values for various weak acids and bases.

TABLE 21.13: Acid-Base Dissociation Constants

Weak Acid	K_a	Weak Base	K_b
$HC_2H_3O_2$	1.8×10^{-5}	CH_3NH_3OH	4.4×10^{-4}
HCN	6.2×10^{-10}	NH_4OH	1.8×10^{-5}

Consider a water solution of the salt ammonium cyanide, NH_4CN . This is a salt of a weak acid and a weak base. The acidity-basicity of the salt solution will depend on the relative strength of the parent acid and base. The K_a for HCN is 6.2×10^{-10} . The K_b for NH_4OH is 1.8×10^{-5} . Since the K_a is smaller than the K_b , the cyanide ion has greater attraction for hydrogen ions than the ammonium ion has for hydroxide ions. Therefore, more hydrogen ions will be removed from the solution than hydroxide ions and the hydroxide ions will be in excess. This solution will be basic.

Part I Exercises

For the following salts, identify the parent acid and strength, the parent base and strength, and indicate whether the salt solution will be acidic or basic. Complete the table.

Salt	Parent Acid	Strong or Weak	Parent Base	Strong or Weak	Solution Acid/Base
Na_2SO_4	H_2SO_4	Strong	$NaOH$	Strong	Neutral
$NaHCO_3$	H_2CO_3	Weak	$NaOH$	Strong	Basic
NH_4Cl	HCl	Strong	NH_4OH	Weak	Acidic
K_2SO_4					
KNO_2					
$CH_3NH_3NO_3$					
$BaCO_3$					
KCN					

Part II

If we know the molarity of the salt solution and the necessary K_a and/or K_b values, we can calculate the pH of the salt solution.

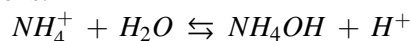
Example 1

What is the pH of a 0.100 M solution of ammonium nitrate, NH_4NO_3 ?

Solution

NH_4NO_3 is the salt product of the reaction between a strong acid, HNO_3 , and a weak base, NH_4OH . Therefore, this salt solution will be acidic. The actual pH can be calculated as shown below.

The ammonium ion in the solution acts as an acid by removing hydroxide ions from solution and thus produces hydrogen ions.



Since the ammonium ion is acting as an acid in this reaction, the equilibrium constant for the reaction is a K_a .

$$K_a = \frac{[NH_4OH][H^+]}{[NH_4^+]}$$

The K_a for ammonium ion acting as an acid is calculated by dividing the K_w by the K_b for ammonium hydroxide.

$$K_a \text{ for } NH_4^+ = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

Therefore,

$$K_a = \frac{[NH_4OH][H^+]}{[NH_4^+]} = 5.6 \times 10^{-10}$$

Since the ammonium nitrate dissociates 100%, the initial concentration of the of the ammonium ion is 0.100 M. As usual, we let x represent the molarity of the ammonium hydroxide formed and therefore, x will also represent the molarity of the hydrogen ion formed. The molarity of the ammonium ion at equilibrium will 0.100 M $- x$ but the x in this expression is much smaller than 0.100 M and can be neglected.

Therefore,

$$5.6 \times 10^{-10} = \frac{(x)(x)}{(0.100)}$$

And

$$[H^+] = x = 7.5 \times 10^{-5} \text{ and } pH = 4.1$$

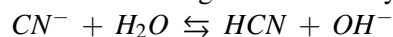
You should always check to make sure that your original determination of whether the solution will be acidic or basic is in agreement with the pH you calculate. In this case, we decided that the solution would be acidic and the pH of 4.1 is in agreement with the solution being acidic.

Example 2

What is the pH of a 1.00 M solution of sodium cyanide, $NaCN$?

Solution

Sodium cyanide is the salt of a weak acid and strong base. The cyanide ion in solution will act as a base, attaching to hydrogen ions and leaving an excess of hydroxide ions in the solution. The solution will be basic.



The K_b for cyanide ion can be calculated by dividing the K_w by the K_a for hydrocyanic acid.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-10}} = 1.6 \times 10^{-5}$$

$$K_b = 1.6 \times 10^{-5} = \frac{[HCN][OH^-]}{[CN^-]} = \frac{(x)(x)}{1.00}$$

$$[OH^-] = x = 4.0 \times 10^{-3} M$$

$$pOH = 2.4 \text{ and therefore } pH = 11.6$$

We originally decided this solution would be basic and the calculated pH verifies that it is basic.

Exercises Part II

1. What is the pH of an 0.200 M solution of potassium acetate, $KC_2H_3O_2$?
2. What is the pH of a 0.0100 M solution of ammonium bromide, NH_4Br ?

Titration Worksheet

CK12 Chemistry

Name _____ Date _____

Show ALL your work in solving these problems.

1. What is the concentration of hydroiodic acid if 59.69 mL of it is neutralized by 40.39 mL of 0.2968 M lithium hydroxide solution?
2. It requires 65.95 mL of potassium hydroxide solution to neutralize 89.29 mL of 0.2118 M nitric acid. What is the concentration of the potassium hydroxide solution?
3. What is the molarity of a CuOH solution if 50.50 mL of the solution is titrated to the endpoint with 51.99 mL of 0.3574 M carbonic acid, H_2CO_3 ?
4. What volume of 0.8351 M $\text{Fe}(\text{OH})_2$ is needed to neutralize 98.35 mL of 0.5417 M H_3PO_4 ?
5. 5.000 g of an unknown monoprotic solid acid was titrated with 163.0 mL of 0.1500 M NaOH solution. What is the molar mass of the acid?
6. A solution of NaOH was standardized against oxalic acid dehydrate, $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$ (molar mass = 126.066 g/mol). The volume of NaOH solution required to neutralize 1.500 g of oxalic acid was 45.00 mL. Oxalic acid is a diprotic acid. The standardized NaOH solution is then used to titrate 3.500 g of an unknown solid monoprotic acid, HA . The titration required 42.00 mL of the NaOH solution.
 - A. Calculate the molarity of the NaOH solution.
 - B. Calculate the moles of HA in the 3.500 g sample.
 - C. Calculate the molar mass of the unknown acid.

Buffers Worksheet

A buffer solution is one which maintains an approximately constant pH when small amounts of either a strong acid or base are added. Solutions containing either weak acids and their conjugate bases, or weak bases and their conjugate acids, can be buffering solutions. A common method of producing a buffer is to make solution of a weak acid and a salt of that weak acid.

An example of a buffer solution is a solution containing both acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, and sodium acetate, $\text{NaC}_2\text{H}_3\text{O}_2$. The weak acid is acetic acid and the conjugate base is the acetate ion from sodium acetate.

- The number of moles of the conjugate pairs must be large compared to the moles of added strong acids or bases.
- The ratio $\frac{[\text{weak acid}]}{[\text{conjugate base}]}$ or $\frac{[\text{weak base}]}{[\text{conjugate acid}]}$ should lie between 0.1 and 10, with optimal buffering of either strong acid or strong base at a 1:1 ratio.
- The pH of a 1:1 ratio buffer is equal to the pK_a of the weak acid or the pK_b of the weak base. In practical buffers, the ratio of [acid] or [base] to [conjugate] is close to 1:1 and therefore the pH of the buffer is close to the pK_a or pK_b .

Possible Buffers

Weak acid = $\text{HC}_2\text{H}_3\text{O}_2$ Conjugate base = $\text{C}_2\text{H}_3\text{O}_2^-$ $K_a = 1.8 \times 10^{-5}$

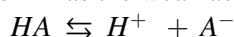
Weak base = NH_4OH Conjugate acid = NH_4^+ $K_b = 1.8 \times 10^{-5}$

Weak acid = HCO_3^- Conjugate base = CO_3^{2-} $K_a = 5.6 \times 10^{-11}$

Weak acid = $H_2PO_4^-$ Conjugate base = HPO_4^{2-} $K_a = 6.2 \times 10^{-8}$

Let's do some general calculations with a theoretical weak acid.

We will use HA as the weak acid and write the dissolving equation and K_a expression for this weak acid.



$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Let's assume the K_a for this weak acid is 1.00×10^{-4} and assume we are creating this buffer solution by pouring together 100. mL of 0.200 M HA solution and 100. mL of 0.200 M NaA solution. One of the concepts you must keep track of is the fact that as soon as the two solutions are poured together, the concentrations of the substances are cut in half due to dilution.

The initial concentrations in the solution (before any dissolving or reaction takes place) will be $[HA] = 0.100 M$, $[H^+] = 0 M$, and $[A^-] = 0.100 M$.

Some amount of the $[HA]$ will dissolve and produce $[H^+]$ and $[A^-]$. If we assign the variable x to represent the molarity of $[HA]$ that dissolves, then we can represent the concentrations as follows; $[HA] = 0.100 - x$, $[H^+] = x$, and $[A^-] = 0.100 + x$.

We can then substitute these quantities into the K_a expression.

$$K_a = 1.00 \times 10^{-4} = \frac{[H^+][A^-]}{[HA]} = \frac{(x)(0.100 + x)}{0.100 - x}$$

At this point we can make a very simplifying assumption. Weak acids do not dissociate much, perhaps one in one thousand or one in ten thousand, or even one in one hundred thousand. Therefore, the size of the value of x is extremely small compared to 0.100. The x when added to or subtracted from a much larger number can be neglected. Therefore, the expression above becomes (when the added and subtracted x 's are dropped):

$$K_a = 1.00 \times 10^{-4} = \frac{[H^+][A^-]}{[HA]} = \frac{(x)(0.100)}{0.100}$$

When we solve this equation for $[H^+]$, we get

$$[H^+] = \frac{(1.00 \times 10^{-4})(0.100)}{0.100}$$

You can easily see that the $[H^+] = 1.00 \times 10^{-4}$ and the $pH = 4$.

You can also see in this calculation why the mathematics is greatly simplified when the weak acid concentration and conjugate base concentration are equal.

In the equation solved for $[H^+]$ above, the 1.00×10^{-4} is the K_a , the 0.100 in the numerator is $[HA]$ and the 0.100 in the denominator is the $[A^-]$.

This provides you with an equation, called the Henderson-Hasselbalch equation, which allows you to calculate the pH of a buffer from the K_a and the concentrations of the weak acid and conjugate base.

$$H^+ = (K_a) \left(\frac{[HA]}{[A^-]} \right)$$

Taking the $-\log$ of both sides of the equation produces the Henderson-Hasselbalch equation. $pH = pK_a + \log \frac{[A^-]}{[HA]}$

Example Exercises

1. Use the Henderson-Hasselbalch equation to calculate the pH of a buffer produced by mixing 300. mL of 1.00 M $H_2C_2H_3O_2$ and 100. mL of 1.00 M $NaC_2H_3O_2$. $K_a = 1.8 \times 10^{-5}$

Solution:

Using the dilution equation to calculate the new molarities of the weak acid and conjugate base:

$$M_N \times V_N = M_O \times V_O$$

$$M_N = \frac{M_o \times V_o}{V_N}$$

$$[HC_2H_3O_2]_N = \frac{(1.00 M)(300. mL)}{(400. mL)} = 0.75 M$$

$$[NaC_2H_3O_2]_N = \frac{(1.00 M)(100. mL)}{(400. mL)} = 0.25 M$$

Plugging the K_a and the two concentrations into the Henderson-Hasselbalch equation:

$$pH = -\log(1.8 \times 10^{-5} + \log \frac{0.75}{0.25}) = 4.74 = 0.47 = 5.21$$

2. On page 1 of this worksheet, there is a list of four possible buffers. Which of the possible buffers listed above should be chosen to prepare a buffer of pH = 7?

Solution:

Since it is easiest to prepare a buffer where the pH of the buffer is close to the pK_a or pK_b , a buffer with a pH of 7 should be prepared from with $\frac{[H_2PO_4^-]}{[HPO_4^{2-}]}$.

3. What is the pH of a buffer formed from 50.0 mL of 15.0 M NH_4OH and 53.5 g of NH_4Cl in enough water to make 500. mL of solution?

Solution:

Step 1: Determine the moles of the NH_4^+ ion.

$$mols NH_4^+ = \frac{g}{mm} = \frac{53.5 g}{53.5 g/mol} = 1.00 mol$$

Step 2: Determine the molarity of the NH_4^+ ion.

$$[NH_4^+] = \frac{mol}{L} = \frac{1.00 mol}{0.500 L} = 2.00 M$$

Step 2: Determine the molarity of the NH_4OH (using dilution formula).

$$[NH_4OH]_N = \frac{(15.0 M)(0.050 L)}{0.500 L} = 1.5 M$$

Step 3: Find the K_b for $[NH_4OH]$.

$$K_b \text{ for } [NH_4OH] = 1.8 \times 10^{-5}$$

Step 4: Plug the data in the Henderson-Hasselbalch equation (for base).

$$pOH = pK_b - \log \frac{[NH_4OH]}{[NH_4^+]} \quad pOH = -\log(1.8 \times 10^{-5}) + -\log\left(\frac{1.5 M}{2.00 M}\right) = 4.74 + 0.12 = 4.86$$

$$pH = 14.00 - 4.86 = 9.14$$

Problems

- Calculate the pH of 0.100 M NH_4OH .
- Calculate the pH of 40.0 mL of 0.100 M NH_4OH that has been diluted with 20.0 mL of water.
- Calculate the pH of 40.0 mL of 0.100 M NH_4OH that has been mixed with 20.0 mL of 0.200 M NH_4Cl .
- A buffer solution contains 0.384 M HCO_3^- and 0.239 M CO_3^{2-} . What is the pH of this buffer? $K_a = 5.6 \times 10^{-11}$
- Calculate the pH of a solution that is 0.120 M in acetic acid and 0.210 M in sodium acetate. Use 1.76×10^{-5} as the K_a for acetic acid.

Answers to Worksheets

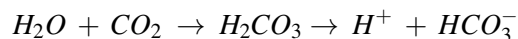
- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

21.5 Extra Reading for Chapter 21

Acid Rain

The exceptional characteristics of the substance known as water have been recognized and appreciated for millennia. In particular, its ability as a solvent provides for many of the vital processes enabling life, such as acting as the medium in which red blood cells transport oxygen in our bodies. Yet water's propensity to dissolve ions, other liquids and even gases may not always produce physical or biochemical advantages.

The water supply on Earth is continuously transported and concomitantly purified by a mechanism known as the hydrological cycle. As solar radiation heats the Earth's surface, water molecules evaporate and then condense into cloud formations as they reach higher elevations and cooler atmospheric levels. Large-scale weather patterns transport the water in these cloud formations around the globe, and return the water to the surface as precipitation. Despite the purification of the substance by this process, rainwater is found to have a *pH* that is not neutral as one might expect, but mildly acidic. During its passage through the atmosphere, water's extraordinary capacity as a solvent absorbs carbon dioxide in the air, and small quantities of carbonic acid is generated as shown:



In our modern industrialized world, there are other gases present in the atmosphere that, like CO_2 , can dissolve in atmospheric moisture. In particular, the presence of NO_x and SO_x , byproducts of fossil fuel combustion, is a specific concern.

NO_x , formed by the reaction of nitrogenous contaminants in fuels with oxygen, can react with water in the atmosphere to generate nitric acid, HNO_3 . In its concentrated form, nitric acid is a corrosive material that can dissolve some metals. Likewise, sulfur oxide contaminants react with moisture yielding sulfuric acid, the viscous acid found in lead-acid car batteries. As these acids are produced and dispersed in the atmosphere, they constitute an environmental issue that transcends borders and physical boundaries.

Acid rain is then precipitation that possesses acidity greater than that of normally slightly acidic rainwater. Its effects can be noted on both biological systems and physical structures. Trees in many areas of the world bear the evidence of acid rain damage in the form of brittle, browned leaves, but the principal destruction to plants is to the root system. Increased acidity limits access of beneficial ions such as Ca^{2+} and Mg^{2+} but encourages the solubility of damaging ions such as Al^{3+} in the soil. Aquatic organisms such as fish experience skeletal growth problems due to limited access to calcium ions. Physical structures also bear witness to the destructive nature of acid rain; marble statues erode due to long-term exposure to acidic rainfall.



21.6 Assessment for Chapter 21

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 22 Thermochemistry C-SE-TE

Chapter Outline

- 22.1 ENERGY CHANGES IN REACTIONS**
- 22.2 ENTHALPY**
- 22.3 SPONTANEOUS PROCESSES**
- 22.4 ENTROPY**
- 22.5 GIBB'S FREE ENERGY**
- 22.6 WORKSHEETS FOR CHAPTER 22**
- 22.7 EXTRA READING FOR CHAPTER 22**
- 22.8 ASSESSMENT FOR CHAPTER 22**

Lessons and Number of Activities for Lessons

TABLE 22.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Energy Change in Reactions	0	0	0	0
2. Enthalpy	1	0	2	0
3. Spontaneous Processes	0	0	0	0
4. Entropy	0	0	1	0
5. Gibb's Free Energy	0	0	1	0

22.1 Energy Changes in Reactions

Student Behavioral Objectives

The student will:

- define energy, potential energy, kinetic energy.
- define endothermic and exothermic reactions.
- describe how heat is transferred in endothermic and exothermic reactions.

Timing, Standards, Activities

TABLE 22.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Energy Changes in Reactions</i>	1.0	7a

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Energy Changes in Reactions (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

22.2 Enthalpy

Student Behavioral Objectives

The student will:

- define and understand enthalpy of reaction.
- calculate the enthalpy of reaction using $\Delta H_{rxn} = \Delta H_{products} - \Delta H_{reactants}$.
- describe, interpret, and draw potential energy diagrams.
- define and understand ΔH_f .
- define Hess's Law.
- calculate ΔH .

Timing, Standards, Activities

TABLE 22.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Enthalpy</i>	2.0	7b, 7e

Activities for Lesson 2

Laboratory Activities

1. Hess's Law Lab

Demonstrations

1. None

Worksheets

1. Enthalpy Worksheet
2. Hess's Law Worksheet

Extra Readings

1. None

Answers for Enthalpy (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

22.3 Spontaneous Processes

Student Behavioral Objectives

The student will:

- define a spontaneous and non-spontaneous reaction.
 - identify processes as either spontaneous or non-spontaneous.
 - describe how endothermic and exothermic reactions can be spontaneous or non-spontaneous.
 - explain the lack of correlation between spontaneity and speed of reaction.
-

Timing, Standards, Activities

TABLE 22.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Spontaneous Processes</i>	1.0	7b

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Spontaneous Processes (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

22.4 Entropy

Student Behavioral Objectives

The student will:

- define entropy.
- calculate change in entropy from standard entropies of formation.
- relate entropy to the tendency toward spontaneity.
- describe the factors that affect the increase or decrease in disorder.

Timing, Standards, Activities

TABLE 22.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Entropy</i>	1.0	None

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Entropy Worksheet

Extra Readings

1. None

Answers for Entropy (L4) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

22.5 Gibb's Free Energy

Student Behavioral Objectives

The student will:

- define Gibbs Free Energy.
- calculate Gibbs Free Energy given the enthalpy and entropy.
- use Gibbs Free Energy to predict spontaneity.

Timing, Standards, Activities

TABLE 22.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Gibb's Free Energy</i>	1.5	7f

Activities for Lesson 5

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Enthalpy, Entropy, and Free Energy Worksheet

Extra Readings

1. None

Answers for Gibb's Free Energy (L5) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 22

From the Annenberg videos . . . Exothermic and Endothermic reactions.

- http://www.learner.org/vod/vod_window.html?pid=805

This video provides a blackboard lecture showing the concepts involved with an example of using Hess's Law.

- <http://www.youtube.com/watch?v=j4-UrAaAy3M>

The *learner.org* website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos but there is no charge. The website has one video that relates to this lesson called "The Driving Forces."

- <http://learner.org/resources/series61.html>

This video shows an example of how to plug values into Gibbs Free Energy equation.

- <http://www.youtube.com/watch?v=ECjH1ErqzRU>

A blackboard example of using Gibbs free energy equation to determine whether a reaction is spontaneous.

- <http://www.youtube.com/watch?v=sG1ZAdYi13A>

Laboratory Activities for Chapter 22

Teacher's Resource Page for Heat of Reaction - Hess's Law

Lab Notes

Solid sodium hydroxide is highly hygroscopic . . . it absorbs water from the air. The solution produced by this absorption is extremely corrosive . . . it will damage clothes almost immediately even though the wearer may not notice for several hours. If the students leave the lid off the reagent bottle after removing *NaOH* pellets, the *NaOH* will absorb enough water to ruin the bottle of reagent. The *NaOH* pellets may appear to only be slightly damp but later they will become one solid rock and unusable. Also, any pellets spilled on the table or floor will become a serious hazard due to slippery floors and corrosive puddles for hands and sleeves to be dipped into. Spills must be cleaned up immediately and jar lids must be closed immediately. Direct and constant supervision of this step is desirable.

The 200. mL of water used by the students in step 1 may be tap water. But when tap water comes out of the faucet, it is considerably colder than room temperature. Using very cold water will encourage heat absorption from the air during the reaction and will affect the Δt . For this reason, the water should be taken from the tap at least one day ahead of the lab and left in the room to adjust to room temperature. It can be stored in used (but clean) milk containers. For 15 pairs of students, you will need about 4 liters of water.

Solution Preparation

For 15 pairs of students, you will need about 2 liters of 0.50 *MNaOH* solution. This can be prepared by dissolving 20.00 grams of solid *NaOH* per liter of solution. It is advised to fill the volumetric flask half full of distilled water,

add the solid $NaOH$, swirl to dissolve the solid, allow some time for the solution to cool (the dissolving of $NaOH$ is highly exothermic), then fill the volumetric flask to the 1.00 liter mark.

For 15 pairs of students, you will need about 2 liters of 0.50 $MHCl$ solution. This can be prepared by dissolving 42 mL of concentrated $HCl(12 M)$ per liter of solution. Always add acid to water not water to acid. Again, fill the volumetric flask half full of distilled water, add the concentrated HCl , swirl to mix, and fill to the 1.00 liter mark.

For 15 pairs of students, you will need about 4 liters of 0.25 $MHCl$. When you prepare the 0.50 $M HCl$ (see above) make 2 extra liters and then dilute these by half to make 0.25 $MHCl$.

Disposal

The solutions produced in this lab can be disposed of down the sink with excess water.

Solutions to Pre-Lab Questions

- $q = mC\Delta t = (500. g)(4.18 J/g \cdot ^\circ C)(3.5^\circ C) = 7315 J = 7.3 kJ$
- Diluting a solution does not change the moles of solute contained in it. If the moles of $NaCl$ in the original 100. mL is 0.0100 moles, then the moles of $NaCl$ in the 200. mL of diluted solution will also be 0.0100 moles.

Solutions to Post-Lab Questions

- (a) Since the grams of $NaOH$ is doubled, the heat given off will also double.
(b) Since the grams of $NaOH$ is doubled, the moles would also double.
(c) If the heat is doubled and the moles are doubled, heat per mole would remain exactly the same.

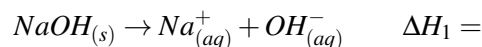
Heat of Reaction - Hess's Law

Background:

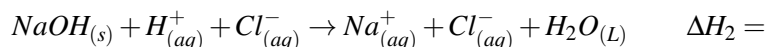
Hess's Law; *The enthalpy change for any reaction depends on the products and the reactants and is independent of the pathway or the number of steps between the reactant and the product.*

In this experiment, you will measure and compare the quantity of heat involved in three reactions. These heats of reaction will be measured using a styrofoam calorimeter. The three reactions are shown below.

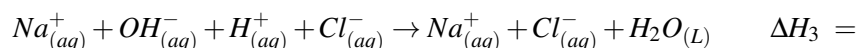
Reaction 1: The dissolving of solid sodium hydroxide in water.



Reaction 2: The reaction of solid sodium hydroxide with dilute hydrochloric acid.



Reaction 3: The reaction of sodium hydroxide solution with dilute hydrochloric acid.



The equation for calculating the heat absorbed by the water in a calorimeter is $q = mC\Delta t$, where q is the heat absorbed in Joules, m is the mass of the water in the calorimeter, C is the specific heat of water ($4.18 J/g \cdot ^\circ C$), and Δt is the temperature change of the water due to the reaction.

Pre-Lab Questions (to be done before lab day):

1. If the temperature of 500. g of water is raised by 3.5°C , how many Joules of heat were absorbed?
2. 100. mL of 0.100 M NaCl contains 0.0100 moles of NaCl (moles = $M \times L$). If 100. mL of 0.100 M NaCl solution is diluted to 200. mL, how many moles of NaCl will it contain?

Purpose:

- To physically measure and compare the amount of heat involved in three separate but related reactions.
- To provide experimental verification of Hess's Law.

Apparatus and Materials:

- 2 - large styrofoam cups (nested to provide better insulation)
- 100 mL graduated cylinder
- thermometer
- balance
- 150 mL beaker
- sodium hydroxide, NaOH (solid)
- 0.50 M NaOH solution (prepared the day before so it has time to cool)
- 0.50 M HCl solution
- 0.25 M HCl solution
- tap water

Safety Issues:

Hydrochloric acid and sodium hydroxide are corrosive. Avoid contact with skin. If any touches your skin, wash it off immediately. Solid sodium hydroxide is especially dangerous because it absorbs moisture rapidly from the air, forming an extremely concentrated solution. Avoid spilling this solid, and if a spill occurs, it must be cleaned up immediately. Ask your teacher for help. Be sure to close the lids of bottles of NaOH securely immediately after using. A lab coat or apron and goggles are required for this lab.

Procedure:**Part One: The dissolving of solid NaOH in water.**

1. Put 200. mL of room temperature water into your nested styrofoam calorimeter. Stir carefully with a thermometer until a constant temperature is reached. Record this temperature.
2. Acquire about 2 grams of solid NaOH, determine its exact mass and record the mass in your data table. It is necessary to perform this operation as quickly as possible to avoid too much water absorption by the NaOH. (This water absorption not only alters its mass but begins the reaction before you are ready.)
3. Place the solid NaOH into the water in your calorimeter. Stir gently with the thermometer until the solid is completely dissolved and record the highest temperature reached.
4. Discard the solution down the sink with excess water and rinse the cup and thermometer with water.

Part Two: The reaction of solid NaOH with dilute HCl.

5. Repeat steps 1 - 3 but replace the 200 mL of water with 200. mL of 0.25 M HCl.
6. Again discard the solution and rinse your equipment with water.

Part Three: The reaction of NaOH solution with hydrochloric acid solution.

7. Accurately measure 100. mL of 0.50 M HCl solution and pour it into your calorimeter. Accurately measure 100. mL of 0.50 M NaOH solution and place it in a 150 mL beaker. Measure and record the temperature of each solution. Be sure to rinse and dry your thermometer when switching solutions.
8. Add the sodium hydroxide solution to acid solution in your calorimeter. Stir the solution and record the highest temperature reached.

9. Discard solution and rinse equipment.

Data Table

Part One

Assuming the density of water to be 1.0 g/mL , mass of water = _____ g

Mass of added solid NaOH = _____ g

Total mass of solution = _____ g

Molar mass of NaOH = _____ g

Initial temperature of water = _____ $^{\circ}\text{C}$

Final temperature of solution = _____ $^{\circ}\text{C}$

Δt = _____ $^{\circ}\text{C}$

Part Two

Assuming the density of the HCl solution to be 1.0 g/mL , mass of acid solution = _____ g

Mass of NaOH added = _____ g

Total mass of solution = _____ g

Initial temperature of HCl solution = _____ $^{\circ}\text{C}$

Final temperature of solution = _____ $^{\circ}\text{C}$

Δt = _____ $^{\circ}\text{C}$

Part Three

Assuming the density of both the HCl and NaOH solutions to be 1.0 g/mL ,

mass of HCl solution = _____ g

mass of NaOH solution = _____ g

total mass of mixture = _____ g

Initial temperature of HCl solution = _____ $^{\circ}\text{C}$

Initial temperature of NaOH solution = _____ $^{\circ}\text{C}$

Average initial temperature = _____ $^{\circ}\text{C}$

Final temperature of mixture = _____ $^{\circ}\text{C}$

Δt = _____ $^{\circ}\text{C}$

Calculations

Note: While the chemical properties of very dilute solutions can be quite different from the chemical properties of water, the physical properties of very dilute solutions are nearly the same as water. In this lab, the density and specific heat of the dilute solutions can be assumed to be the same as water.

1. Calculate the heat absorbed by the solution in the calorimeter in part one.
2. Calculate the moles of NaOH dissolved in part one.
3. Calculate the heat released per mole of NaOH .
4. Determine ΔH_1 . (Keep in mind that exothermic reactions are assigned negative ΔH values.)
5. Calculate the heat absorbed by the solution in the calorimeter in part two.
6. Calculate the moles of NaOH reacted in part two.

7. Calculate ΔH_2 .
8. Calculate the heat absorbed by the solution in part three.
9. Calculate the number of moles of $NaOH$ reacted in part three.
10. Calculate ΔH_3 .
11. Reactions 1 and 3 will add together to produce reaction 2. Show this is true. Therefore, according to Hess's Law, $\Delta H_1 + \Delta H_3$ should equal ΔH_2 . Assuming ΔH_2 to be the true value, calculate your percent error.

$$\% \text{ error} = \frac{\text{true value} - \text{experimental value}}{\text{true value}} \times 100 =$$

Post-Lab Questions

1. Suppose you had used 4 *grams* of $NaOH$ instead of 2 *grams* in part one.
 - a. How would the answer to calculation 1 have changed?
 - b. How would the answer to calculation 2 have changed?
 - c. How would the answer to calculation 3 have changed?

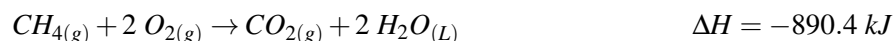
Demonstrations for Chapter 22

22.6 Worksheets for Chapter 22

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Enthalpy Worksheet

1. The combustion of methane, CH_4 , releases 890.4 kJ/mol of heat. That is, when one mole of methane is burned, 890.4 kJ are given off to the surroundings. This means that the products have 890.4 kJ less energy stored in the bonds than the reactants. Thus, ΔH for the reaction = -890.4 kJ . A negative symbol for ΔH indicates an exothermic reaction.



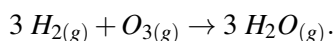
- A. How much energy is given off when 2.00 mol of CH_4 are burned?
- B. How much energy is released when 22.4 g of CH_4 are burned?
- C. If you were to attempt to make 45.0 g of methane from CO_2 and H_2O (with O_2 also being produced), how much heat would be absorbed during the reaction?

Use the following heat of formation table in questions 2 – 6.

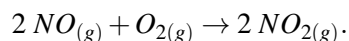
TABLE 22.7: The Standard Enthalpy and Entropy of Various Substances

Substance	ΔH_f° (kJ/mol)	S° (J/K·mol)
$C_4H_{10}(g)$	-126	310
$CaC_2(s)$	-63	70.
$Ca(OH)_2(s)$	-987	83
$C_2H_2(g)$	227	201
$CO_2(g)$	-394	214
$H_2(g)$	0	131
$H_2O(g)$	-242	189
$H_2O(L)$	-286	70.
$NH_3(g)$	-46	193
$NO(g)$	90.	211
$NO_2(g)$	34	240.
$N_2O(g)$	82	220.
$O_2(g)$	0	205
$O_3(g)$	143	239

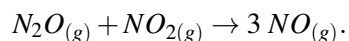
2. Using data from the heat of formation table above, calculate the enthalpy of reaction for



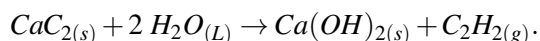
3. Using data from the heat of formation table above, calculate the heat of reaction for



4. Using data from the heat of formation table above, calculate the heat of reaction for



5. Using data from the heat of formation table above, calculate the heat of reaction for



6. Many cigarette lighters contain liquid butane, C_4H_{10} . Using the heat of formation table above, calculate the quantity of heat produced when 1.0 g of gaseous butane is completely combusted in air.

Hess's Law Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

Example Problem

Find the ΔH for the reaction below, using the following reactions and their ΔH values.

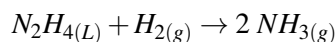


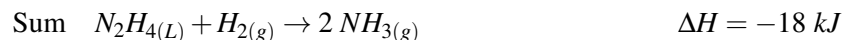
TABLE 22.8: Given Equations and

Equation	ΔH Value
$\text{N}_2\text{H}_4(L) + \text{CH}_4\text{O}_{(L)} \rightarrow \text{CH}_2\text{O}_{(g)} + \text{N}_2(g) + 3 \text{H}_2(g)$	$\Delta H = -37 \text{kJ}$
$\text{N}_2(g) + 3 \text{H}_2(g) \rightarrow 2 \text{NH}_3(g)$	$\Delta H = -46 \text{kJ}$
$\text{CH}_4\text{O}_{(L)} \rightarrow \text{CH}_2\text{O}_{(g)} + \text{H}_2(g)$	$\Delta H = -65 \text{kJ}$

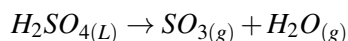
Solution

TABLE 22.9: Solution Arrangement

Changes	Equation	ΔH Value
Keep Same	$\text{N}_2\text{H}_4(L) + \text{CH}_4\text{O}_{(L)} \rightarrow \text{CH}_2\text{O}_{(g)} + \text{N}_2(g) + 3 \text{H}_2(g)$	$\Delta H = -37 \text{kJ}$
Keep Same	$\text{N}_2(g) + 3 \text{H}_2(g) \rightarrow 2 \text{NH}_3(g)$	$\Delta H = -46 \text{kJ}$
Reverse	$\text{CH}_2\text{O}_{(g)} + \text{H}_2(g) \rightarrow \text{CH}_4\text{O}_{(L)}$	$\Delta H = +65 \text{kJ}$

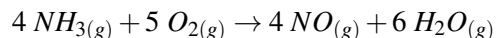
**Exercises**

1. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

**TABLE 22.10: Given Equations and**

Equation	ΔH Value
$H_2S(g) + 2 O_2(g) \rightarrow H_2SO_4(L)$	$\Delta H = -235 \text{ kJ}$
$H_2S(g) + 2 O_2(g) \rightarrow SO_3(g) + H_2O(L)$	$\Delta H = -207 \text{ kJ}$
$H_2O(L) \rightarrow H_2O(g)$	$\Delta H = +44 \text{ kJ}$

2. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

**TABLE 22.11: Given Equations and**

Equation	ΔH Value
$N_2(g) + O_2(g) \rightarrow 2 NO(g)$	$\Delta H = -180.5 \text{ kJ}$
$N_2(g) + 3 H_2(g) \rightarrow 2 NH_3(g)$	$\Delta H = -91.8 \text{ kJ}$
$2 H_2(g) + O_2(g) \rightarrow 2 H_2O(g)$	$\Delta H = -483.6 \text{ kJ}$

3. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

**TABLE 22.12: short caption**

Equation	ΔH Value
$P_4(s) + 6 Cl_2(g) \rightarrow 4 PCl_3(g)$	$\Delta H = -2439 \text{ kJ}$
$4 PCl_5(g) \rightarrow P_4(s) + 10 Cl_2(g)$	$\Delta H = +3438 \text{ kJ}$

4. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

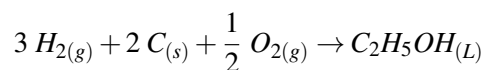
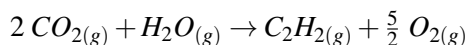


TABLE 22.13: Given Equations and

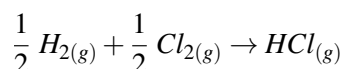
Equation	ΔH Value
$C_2H_5OH(L) + 3 O_{2(g)} \rightarrow 2 CO_{2(g)} + 3 H_2O(L)$	$\Delta H = -875.0 \text{ kJ}$
$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$	$\Delta H = -394.5 \text{ kJ}$
$H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O(L)$	$\Delta H = -285.8 \text{ kJ}$

5. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

**TABLE 22.14: Given Equations and**

Equation	ΔH Value
$C_2H_2(g) + 2 H_{2(g)} \rightarrow C_2H_6(g)$	$\Delta H = -94.5 \text{ kJ}$
$H_2O(g) \rightarrow H_{2(g)} + \frac{1}{2} O_{2(g)}$	$\Delta H = +71.2 \text{ kJ}$
$C_2H_6(g) + \frac{7}{2} O_{2(g)} \rightarrow 2 CO_{2(g)} + 3 H_2O(g)$	$\Delta H = -283.0 \text{ kJ}$

6. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

**TABLE 22.15: Given Equations and**

Equation	ΔH Value
$COCl_{2(g)} + H_2O(L) \rightarrow CH_2Cl_{2(L)} + O_{2(g)}$	$\Delta H = +48 \text{ kJ}$
$2 HCl(g) + \frac{1}{2} O_{2(g)} \rightarrow H_2O(L) + Cl_{2(g)}$	$\Delta H = +105 \text{ kJ}$
$CH_2Cl_{2(L)} + H_{2(g)} + \frac{3}{2} O_{2(g)} \rightarrow COCl_{2(g)} + 2 H_2O(L)$	$\Delta H = -403 \text{ kJ}$

Entropy Worksheet

Use the following entropy of formation table in questions 1 – 5.

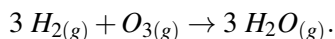
TABLE 22.16: The Standard Enthalpy and Entropy of Various Substances

Substance	ΔH_f° (kJ/mol)	S° (J/K · mol)
$C_4H_{10(g)}$	-126	310
$CaC_{2(s)}$	-63	70.
$Ca(OH)_{2(s)}$	-987	83
$C_2H_2(g)$	227	201
$CO_{2(g)}$	-394	214
$H_{2(g)}$	0	131
$H_2O(g)$	-242	189
$H_2O(L)$	-286	70.
$NH_3(g)$	-46	193
$NO(g)$	90.	211
$NO_{2(g)}$	34	240.
$N_2O(g)$	82	220.

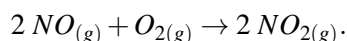
TABLE 22.16: (continued)

Substance	ΔH_f° (kJ/mol)	S° (J/K·mol)
$O_{2(g)}$	0	205
$O_{3(g)}$	143	239

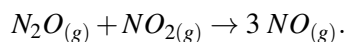
1. Using data from the entropy of formation table above, calculate the entropy of reaction for



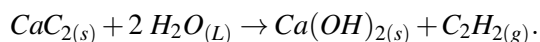
2. Using data from the entropy of formation table above, calculate the change in entropy for



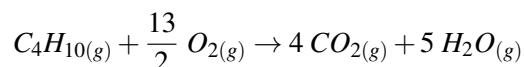
3. Using data from the heat of formation table above, calculate the ΔS° for



4. Using data from the entropy of formation table above, calculate the heat of reaction for



5. Using the entropy of formation table above, calculate the change in entropy for the following reaction.



Enthalpy, Entropy, and Free Energy Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

1. As the amount of energy required to decompose a compound increases, the thermodynamic stability of the compound _____.

- A. increases
- B. decreases
- C. remains constant
- D. varies randomly

2. The enthalpy of formation for a free element is

- A. 0 kJ/mol.
- B. 1 kJ/mol.

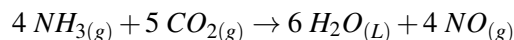
- C. 10 kJ/mol .
- D. -100 kJ/mol .
- E. variable.

Questions 3 and 4 relate to the following equation and ΔH_R value.



3. Which of the following can definitely be concluded from the equation and heat of reaction above?
- A. The reaction is spontaneous.
 - B. The reaction is non-spontaneous.
 - C. The reaction is endothermic.
 - D. The reaction is exothermic.
 - E. None of these.
4. From the equation and heat of reaction above, what is the ΔH_f of HgO ?
- A. 181.7 kJ/mol
 - B. -181.7 kJ/mol
 - C. 0 kJ/mol
 - D. 90.9 kJ/mol
 - E. -90.9 kJ/mol
5. Which of the following four substances is the most thermodynamically stable? Use the data in the Thermodynamic Data Table at the bottom of the worksheet.
- A. $\text{NH}_{3(g)}$
 - B. $\text{CO}_{2(g)}$
 - C. $\text{H}_2\text{O}_{(L)}$
 - D. $\text{NO}_{(g)}$
6. The free energy of a reaction is the combination of _____ and _____.
- A. heat and work
 - B. pressure and volume
 - C. enthalpy and entropy
 - D. internal energy and PV
 - E. None of these.
7. All reactions that occur spontaneously must have a negative _____.
- A. $T\Delta S$
 - B. ΔG
 - C. ΔH
 - D. ΔS
 - E. All of these.

Questions 8, 9, 10, and 11, relate to the equation shown below.



8. Use the data in the Thermodynamic Data Table at the bottom of this worksheet to find the ΔH_R for the reaction above?

- A. +92.8 kJ
- B. -92.8 kJ
- C. -806.3 kJ
- D. +806.3 kJ
- E. None of these.

9. Use the data in the Thermodynamic Data Table at the bottom of this worksheet to find the ΔG_R for the reaction above?

- A. -981.6 kJ
- B. +981.6 kJ
- C. -269.0 kJ
- D. +269.0 kJ
- E. None of these.

10. Use the data in the Thermodynamic Data Table at the bottom of this worksheet to find the ΔS_R for the reaction above?

- A. -575.9 J/°
- B. +575.9 J/°
- C. -1419.1 J/°
- D. +1419.1 J/°
- E. None of these.

11. Use the ΔH_R you found in question 6 and the ΔS_R you found in question 8 to calculate ΔG_R for this reaction.

- A. 634.7 kJ
- B. -634.7 kJ
- C. 977.9 kJ
- D. -977.9 kJ
- E. None of these.

12. Find ΔS for the reaction, $2 \text{NO}(g) + \text{O}_2(g) \rightarrow 2 \text{NO}_2(g)$.

- A. -146.5 J/K
- B. +146.5 J/K
- C. -16.5 J/K
- D. +16.5 J/K
- E. None of these.

13. Find ΔG_R for the reaction, $2 \text{H}_2\text{O}(g) + 2 \text{F}_2(g) \rightarrow \text{O}_2(g) + 4 \text{HF}(g)$.

- A. -1550.0 kJ

B. +1550.0 kJ

C. -635.6 kJ

D. +635.6 kJ

E. None of these.

14. What is the change in enthalpy for $4 Al_{(s)} + 3 O_{2(g)} \rightarrow 2 Al_2O_{3(s)}$?

A. 0 kJ

B. -1657.7 kJ

C. +1657.7 kJ

D. +3351.4 kJ

E. -3351.4 kJ

15. What is the change in entropy for $4 Al_{(s)} + 3 O_{2(g)} \rightarrow 2 Al_2O_{3(s)}$?

A. 0 J/K

B. -626.7 J/K

C. +626.7 J/K

D. -500.0 J/K

E. +500.0 J/K

16. Use the results from questions 14 and 15 to determine under what conditions this reaction will be spontaneous.

A. This reaction will be spontaneous at all temperatures.

B. This reaction will never be spontaneous at any temperature.

C. This reaction will be spontaneous at high temperatures.

D. This reaction will be spontaneous at low temperatures.

TABLE 22.17: Thermodynamic Properties of Some Substances (at

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol·K)
$Al_{(s)}$	0	0	+28.3
$Al_2O_{3(s)}$	-1675.7	-1582.3	+50.9
$CO_{(g)}$	-110.5	-137.2	+197.7
$CO_{2(g)}$	-393.5	-394.4	+213.7
$F_{2(g)}$	0	0	+202.8
$HF_{(g)}$	-271.1	-273.2	+173.8
$H_2O_{(L)}$	-285.8	-237.1	+69.9
$H_2O_{(g)}$	-241.8	-228.6	+188.8
$NH_{3(g)}$	-46.1	-16.5	+192.5
$NO_{(g)}$	+90.3	+86.6	+210.8
$NO_{2(g)}$	+33.2	+51.3	+240.1
$O_{2(g)}$	0	0	+205.1

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org

to request the worksheet answer keys.

22.7 Extra Reading for Chapter 22

22.8 Assessment for Chapter 22

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER **23** Electrochemistry C-SE-TE

Chapter Outline

- 23.1** ORIGIN OF THE TERM OXIDATION
- 23.2** OXIDATION-REDUCTION
- 23.3** BALANCING REDOX EQUATIONS
- 23.4** ELECTROLYSIS
- 23.5** GALVANIC CELLS
- 23.6** WORKSHEETS FOR CHAPTER 23
- 23.7** EXTRA READING FOR CHAPTER 23
- 23.8** ASSESSMENT FOR CHAPTER 23

Lessons and Number of Activities for Lessons

TABLE 23.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Origin of the Term Oxidation	0	0	0	0
2. Oxidation-Reduction	0	0	0	0
3. Balancing Redox Equations Using the Oxidation Number Method	0	0	1	0
4. Electrolysis	0	1	0	0
5. Galvanic Cells	1	0	1	1

23.1 Origin of the Term Oxidation

Student Behavioral Objectives

- The students will define the term “oxidation.”
-

Timing, Standards, Activities

TABLE 23.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Origin of the Term Oxidation</i>	0.5	None

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None
-

Answers for Origin of the Term Oxidation (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

23.2 Oxidation-Reduction

Student Behavioral Objectives

- The students will assign the correct oxidation number to any element in a compound or ion.
- In an oxidation-reduction equation, the students will identify the substance being oxidized, the substance being reduced, the oxidizing agent, and the reducing agent

Timing, Standards, Activities

TABLE 23.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Oxidation-Reduction</i>	2.0	3g

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Oxidation-Reduction (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

23.3 Balancing Redox Equations

Student Behavioral Objectives

- Given a redox reaction, the students will determine which substances are changing their oxidation state.
 - Given a redox reaction, the students will balance the equation using the oxidation number method.
-

Timing, Standards, Activities

TABLE 23.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Balancing Redox Equations</i>	2.0	3g

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Balancing Redox Equations Worksheet

Extra Readings

1. None
-

Answers for Balancing Redox Equations (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

23.4 Electrolysis

Student Behavioral Objectives

- Given a diagram of an electrolysis apparatus including the compound being electrolyzed, the students will identify the anode and the cathode.
- Given a diagram of an electrolysis apparatus including the compound being electrolyzed, the students will write the oxidation and reduction half-reactions.

Timing, Standards, Activities

TABLE 23.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Electrolysis</i>	2.0	None

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. The Electrolysis of Water

Worksheets

1. None

Extra Readings

1. None

Answers for Electrolysis (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

23.5 Galvanic Cells

Student Behavioral Objectives

- The students will describe the conditions necessary for a cell to be standard cell.
- Given a table of standard reduction potentials and a diagram or description of a Galvanic cell, the students will balance the redox equation, calculate the standard cell potential, and determine the direction of electron flow in the external circuit.

Timing, Standards, Activities

TABLE 23.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Galvanic Cells</i>	2.0	None

Activities for Lesson 5

Laboratory Activities

1. An Activity Series Lab

Demonstrations

1. None

Worksheets

1. Electrochemical Cells Worksheet

Extra Readings

1. Dependence of Cell Potential on Concentration

Answers for Galvanic Cells (L5) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 23

Gummy Bear Terminator Demo.

- [/gummi_bear_terminator.htm](http://quiz2.chem.arizona.edu/preproom/Demo%20Files/gummi_bear_terminator.htm) http://quiz2.chem.arizona.edu/preproom/Demo%20Files/gummi_bear_terminator.htm

Oxidation-reduction demos.

- <http://sites.google.com/site/chemistrydemos/7-chemical-reaction/oxidation-reduction>

Electrochemical Cell animation.

- <http://chemmovies.unl.edu/ChemAnime/ECZCELLD/ECZCELLD.html>

Electroplating lesson.

- <http://www.finishing.com/faqs/howworks.html>

This web site offers text material on the demise of the phlogiston theory and some history about some early chemists.

- <http://web.fccj.org/ethall/phlogist/phlogist.htm>

This website offers a short animated video on electron transfer during oxidation-reduction.

- <http://www.youtube.com/watch?v=e6Xxz-VBE6s>

This web site offers more information on oxidation-reduction and other aspects of electrochemistry.

- <http://www.chemtutor.com/redox.htm>

This video shows the process in balancing a redox equation.

- <http://www.youtube.com/watch?v=TBmwhTzc41o>

The following website provides more information in balancing redox reactions.

- http://preparatorychemistry.com/Bishop_Balancing_Redox.htm

This website provides an interactive animation of electrolysis.

- <http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/electroChem/electrolysis10.html>

This animation shows the function of a galvanic cell.

- <http://www.youtube.com/watch?v=A0VUsoeT9aM>

The *learner.org* website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos but there is no charge. The website has one video that relates to this lesson called “The Busy Electron” that explains the principles of electrochemical cell design through batteries, sensors, and a solar-powered car.

- <http://learner.org/resources/series61.html>

The following provides text and animations about electrochemical cells and batteries.

- <http://www.splung.com/content/sid/3/page/batteries>

Laboratory Activities for Chapter 23

Teacher's Pages for An Activity Series Lab

Lab Notes:

Metal sheets can be purchased from chemical supply companies. For this lab, only the thinnest sheets are necessary. Metal foils can be cut with sturdy scissors. Thicker sheets may require tin snips. Any surface corrosion must be cleaned off for this lab. That can be accomplished with emery cloth. It's much easier to sand the sheets before cutting them into 6 mm squares. After sanding the metal sheets, rinse and dry them.

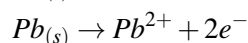
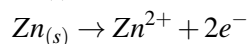
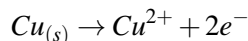
Nitrate compounds are often purchased as hydrates. Make sure you note if your $Cu(NO_3)_2$, $Mg(NO_3)_2$, $Zn(NO_3)_2$, $Pb(NO_3)_2$ are hydrates when calculating masses for making 0.10 M solutions.

You can, of course, make Cl_2 water, Br_2 water, and I_2 water, but it is much more convenient to purchase the amount you will need previous to the lab. These solutions do not store well. Purchase the amount you need year by year to guarantee strength.

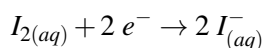
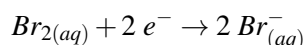
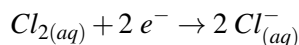
If you think your students will not have enough time to complete both parts of the lab, you can have them do part I on one day, and part II on the next day. Another way to shorten the time for the lab is to make the reference solutions for part II yourself and have them on display. Just follow the first four steps in the Procedure for Part II.

Answers to Pre-Lab Questions

1. Write oxidation half-reactions for the four metals, Cu , Zn , Mg , Pb .

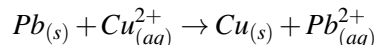
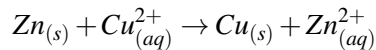
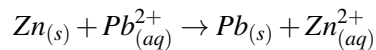
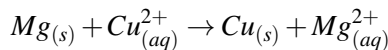
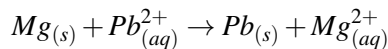
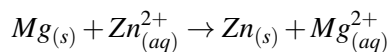


2. Write reduction half-reactions for the three halogens, Cl_2 , Br_2 , and I_2 .

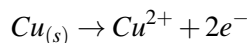
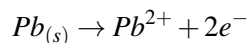
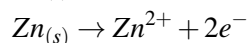
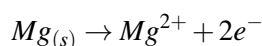


Answers to Post-Lab Questions

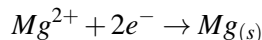
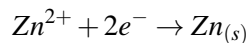
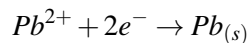
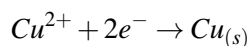
1. Write a balanced net ionic equation for all the reactions that occurred with the metals.



2. List the four oxidation half-reactions written in the pre-lab question 1 in order of decreasing ease of oxidation. That is, the one easiest to oxidize should be first and the most difficult to oxidize should be last.



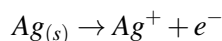
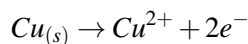
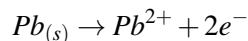
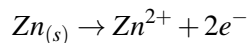
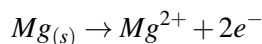
3. Write reduction half-reactions for the four ions in the experiment and put them in order of easiest to reduce first and most difficult ion to reduce last. This list should be in reverse order of your activity series in question 2.



3. Find a list of standard reduction potentials (in your textbook or elsewhere). How does your list in question 3 compare to the standard reduction potential list?

The metal ions are in the same order of ease of reduction.

4. Suppose a piece of shiny silver metal had been placed into separate solutions of each of the metallic ions in the lab and did not react with any of them. Write an oxidation half-reaction for silver and place in its proper place in your metals activity series.

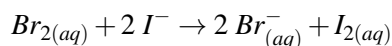
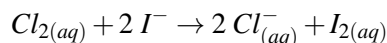
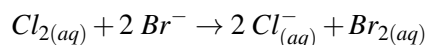


Post-Lab Questions for Part II

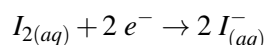
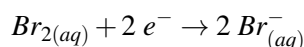
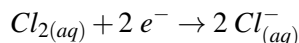
1. Why is it assumed that the halide ions will not dissolve in mineral oil?

Halide ions are charged and will dissolve readily in water but not in a non-polar solvent like mineral oil.

2. Write net ionic equations for all the reactions that occurred with the halogens.



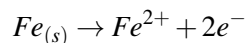
3. List the reduction half-reactions of the halogens in order of decreasing activity.



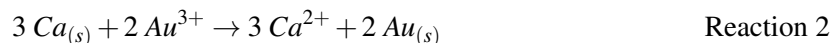
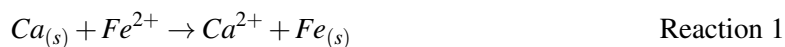
An Activity Series Lab

Background:

A ranking of elements according to their reactivity is called an activity series. For example, an activity series containing the elements calcium, gold, and iron would put the most reactive element of the three, calcium, at the top of the series; iron in the middle; and the least reactive, gold, at the bottom. The more reactive metal is the one that is most easily oxidized, so half-reactions for an activity series of metals would be written in terms of oxidation. The activity series for the three metals mentioned earlier would look like this.



This activity series could be determined experimentally by placing each of the metals in solutions containing the ions of the other two substances and observing which ones react. For example, we could place pieces of calcium metal into separate solutions containing Fe^{2+} ions and Au^{3+} ions. If the reactions occur, the equations would be:



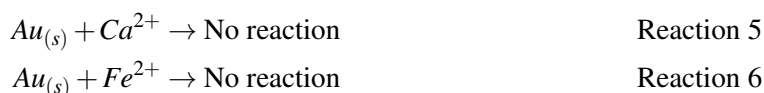
In this case, both reactions occur. The reactions would be detected by the observer because the shiny surface of the calcium metal would darken. In reaction 1, there exists a competition to give up electrons between the calcium atoms in the reactants and the iron atoms in the products. Since this reaction does occur, we know the calcium atoms are stronger at giving up electrons than iron atoms and the reaction runs forward. If the iron atoms were stronger at giving up electrons, the reaction would not occur. The calcium atoms could not transfer electrons to iron ions if the iron atoms so formed were stronger at giving up electrons. The argument for reaction 2 is the same. Since the reaction runs forward, the calcium atoms are stronger at giving up electrons than gold atoms. The fact that calcium atoms gives electrons to both these ions is the reason that calcium is placed on the top of the activity series.

Suppose we now place shiny pieces of iron metal into different solutions containing calcium ions and gold ions. If the reactions occur, the equations would be:



In this case, reaction 3 does not occur. That is, the shiny piece of iron metal remains shiny. No change occurs in this mixture. Iron atoms cannot give electrons to calcium ions, because calcium atoms are stronger at giving electrons than iron atoms. Reaction 4, on the other hand, does occur. The shiny surface of the iron metal darkens and a reaction is apparent. Iron atoms can give electrons to gold ions because iron is stronger at giving electrons than gold, iron is MORE REACTIVE. Therefore, in our activity series, iron must go below calcium but above gold.

The third situation would be to place shiny pieces of gold metal into different solutions of Ca^{2+} ions and Fe^{2+} ions.



In this case, neither reaction occurs. This is because both calcium atoms and iron atoms are stronger at donating electrons and therefore, gold cannot give electrons to either of these metals. That is the reason that gold is placed on the bottom of our activity series.

Pre-Lab Questions

1. Write oxidation half-reactions for the four metals, *Cu*, *Zn*, *Mg*, *Pb*.
2. Write reduction half-reaction for the three halogen, *Cl₂*, *Br₂*, and *I₂*.

Purpose

The purpose of this experiment is to determine an activity series for four metals. If your teacher has you carry out Part II, you will also be determining an activity series for three halogens. Part II makes use of a solvent extraction technique that will be described at the beginning of Part II.

The series of metals to be studied are copper, zinc, magnesium, and lead. Solutions of metal nitrates for each of these metals are placed in reaction wells. A piece of each metal is then placed in the other metals' nitrate solutions and observed to see if any reaction occurs. If a metal reacts with another metal nitrate, then the solid metal has reduced the other metal ion and is, therefore, the more reactive metal of the two. By comparing the results of 16 different reactions, the five metals are ranked from most reactive to least reactive.

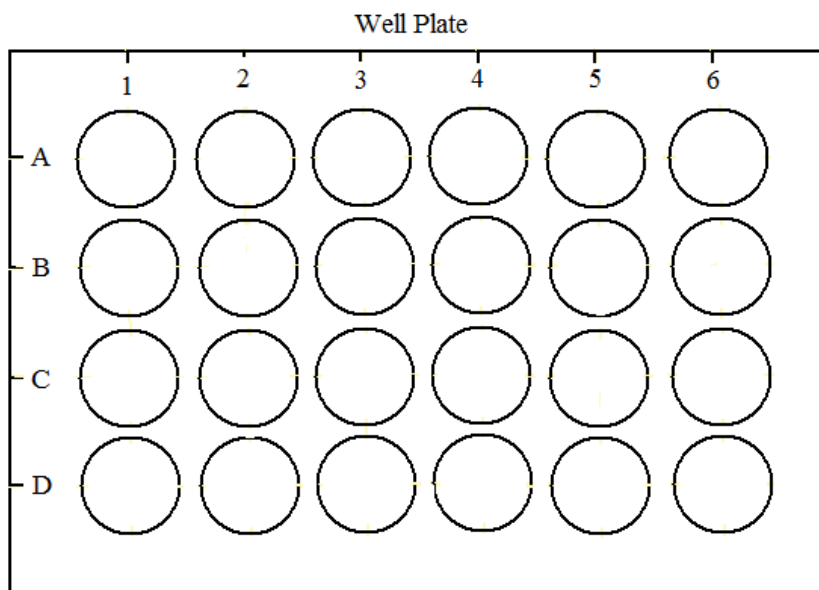
Materials and Materials (per lab group) for Part I

- 4 - copper foil, 6 × 6 mm squares
- 4 - zinc foil, 6 × 6 mm squares
- 4 - magnesium foil, 6 × 6 mm squares
- 4 - lead foil, 6 × 6 mm squares
- 4 mL of copper(II) nitrate solution, 0.10 M
- 4 mL of zinc nitrate solution, 0.10 M
- 4 mL of magnesium nitrate solution, 0.10 M
- 4 mL of lead(II) nitrate solution, 0.10 M
- 24-well reaction plate
- 5 – beral-type pipets
- Chemical forceps

- Stirring rod
- emery cloth

Safety Issues

All solutions are irritating to skin, eyes, and mucous membranes. Handle solutions with care, avoid getting the material on you, and wash your hands carefully before leaving the lab.



Procedure for Part I

1. Place the 24-well plate on top of a sheet of white paper (easier to see changes). Set the plate so there are 6 wells across and 4 wells down. Refer to the figure above to see how the wells are arranged and marked. Note that each well is identified by a unique combination of letter and number, where the letter refers to a horizontal row and the number to a vertical column. Make sure your well plate matches this one.
2. Put one dropper full (about 15 drops) of copper(II) nitrate in wells B1, C1, and D1.
3. Put one dropper full of magnesium nitrate in wells A2, C2, and D2.
4. Put one dropper full of lead(II) nitrate in wells A3, B3, and D3.
5. Put one dropper full of zinc nitrate in wells A4, B4, and C4.
6. Put a small piece of shiny copper metal (you may need to buff it with emery cloth first) in each of the wells containing a solution in the first row (A).
7. Add shiny Mg metal to the solutions in the second row (B), add shiny Pb metal to the solutions in the third row (C), and add shiny Zn metal to the solutions in the fourth row (D). Use a stirring rod to submerge the pieces of metal in the solutions (but don't carry over solution from one well to another). Allow to stand for at least *5 minutes*.
8. Determine if reactions occurred in each well by observing if a new metal has been deposited or the surface of the metal has become coated or discolored.

Disposal

The pieces of solid metal go into waste baskets. All solutions go down the sink followed by plenty of water. Wash the well plate with soap and tap water and then either dry it or rinse it in distilled water.

Data for Part I

Data Table for Activity Series for Some Metals

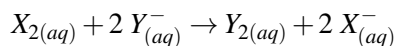
	Cu^{2+}	Mg^{2+}	Pb^{2+}	Zn^{2+}
$\text{Cu}_{(s)}$				
$\text{Mg}_{(s)}$				
$\text{Pb}_{(s)}$				
$\text{Zn}_{(s)}$				

Post-Lab Questions

- Write a balanced net ionic equation for all the reactions that occurred with the metals.
- List the four oxidation half-reactions written in the pre-lab question 1 in order of decreasing ease of oxidation. That is, the one easiest to oxidize should be first and the most difficult to oxidize should be last.
- Write reduction half-reactions for the four ions in the experiment and put them in order of easiest to reduce first and most difficult ion to reduce last. This list should be in reverse order of your activity series in question 2.
- Find a list of standard reduction potentials (in your textbook or elsewhere). How does your list in question 3 compare to the standard reduction potential list?
- Suppose a piece of shiny silver metal had been placed into separate solutions of each of the metallic ions in the lab and did not react with any of them. Write an oxidation half-reaction for silver and place in its proper place in your metals activity series.

Background for Part II

A similar activity series can be constructed for the halogens. The most active halogen is the most easily reduced, accepting electrons from the less active species and replacing it in the halide salt. For example, let X_2 and Y_2 represent two halogens. The sodium salts of these halides would be NaX and NaY . In solution the NaX would produce X^- ions and NaY would produce Y^- ions. If X_2 were the more active halogen, then placing X_2 in a solution containing Y^- ions would result in a reaction in which X_2 received electrons from Y^- ions. (Working with both elemental halogen gases and the halogen ions can be confusing, so in the procedure, the halogen gases are referred to as elemental halogens.)



In water solution, none of the halogen gases or ions are darkly colored enough for an observer to determine exactly what is present in the final solution. The halogen gases, however, do have distinctive colors when dissolved in mineral oil. A technique called solvent extraction is used to separate and identify the products of the reactions in this part of the lab.

Your teacher may have already prepared reference solutions of the halogen gases in mineral oil so that you can see the distinctive colors of Cl_2 , Br_2 , and I_2 when dissolved in mineral oil. You will also want to know if the halide ions produce any color with mineral oil. If the reference solutions are available, you may skip procedure steps 1, 2, 3 and 4. If the reference solutions are not available, you will produce them in steps 1, 2, 3, and 4.

Mineral oil is a non-polar solvent. Halide ions are much more soluble in water (polar) than non-polar solvents. The elemental halogens, Cl_2 , Br_2 , and I_2 , are non-polar molecules and therefore, are much more soluble in non-polar solvents. When water solutions containing both halide ions and elemental halogens are shaken with mineral oil, the elemental halogens present will concentrate in the mineral oil and the ions will concentration in the water. When the water and mineral oil separate (immiscible) the substances dissolved in them are separated. This process is called solvent extraction.

Apparatus and Materials (per lab group) for Part II

- Chlorine water, Cl_2 in water, 3 mL
- Bromine water, Br_2 in water, 3 mL
- Iodine water, I_2 in water, 3 mL
- Mineral oil, 12 mL
- Sodium chloride solution, $NaCl$, 0.10 M, 3 mL
- Sodium bromide solution, $NaBr$, 0.10 M, 3 mL
- Potassium iodide solution, KI , 0.10 M, 3 mL
- Beryl type pipets, 7
- Test tubes, 13 × 100 mm, 12
- Stoppers for test tubes, 12
- Test tube rack

Procedure for Part II

1. As a reference, test to see what color develops when each halogen (elemental) is dissolved in mineral oil. Place one dropper full of chlorine water, one of bromine water, and one of iodine water, into three separate 10 mm test tubes.
2. Add one dropper full of mineral oil to each test tube, stopper the tube, and shake it for 10 seconds.
3. Let the mineral oil rise to the top and record the color that each halogen shows when dissolved in mineral oil. Record your observations in the Part II Data Table.
4. Test to see if halide ions give a color to mineral oil by repeating steps 1, 2, and 3 with solutions of $NaCl$, $NaBr$, and KI with mineral oil. Record your observations.
5. Set up six test tubes in a test tube rack. Label the tubes 1 through 6.
6. You will now react each elemental halogen with the OTHER two halide ion solutions to determine if the ions reduce the halogens. Place one dropper full of $NaBr$ into test tube 1 and one dropper full of KI into test tube 2.
7. Add one dropper full of chlorine water to each of test tubes 1 and 2. Stopper and shake each.
8. Add one dropper full of mineral oil to each of test tubes 1 and 2. Stopper and shake each.
9. When the mineral oil layer has separated, determine its color and decide whether has reaction has occurred. If the color of chlorine appears in the mineral oil, then no reaction has occurred. If either the bromine or the iodine color appears in the mineral oil layer, then there was a reaction.
10. Record both the color and the reaction results for Cl_2 in the Data Table.
11. Repeat the test using bromine water. Add one dropper full of $NaCl$ to test tube 3 and one dropper full of KI to test tube 4. Add one dropper full of bromine water to each of test tubes 3 and 4. Stopper and shake each.
12. Add one dropper full of mineral oil to each of test tubes 3 and 4. Stopper and shake each.
13. When the mineral oil has separated, determine its color and whether or not a reaction occurred. If the color of bromine appears, no reaction has occurred. If either chlorine or iodine color appears, then there was a reaction. Record both color and reaction results in the Data Table.
14. Repeat the test using iodine water. Add one dropper full of $NaCl$ to test tube 5 and one dropper full of $NaBr$ to test tube 6. Add one dropper full of iodine water to each of test tubes 5 and 6. Stopper and shake each.
15. Add one dropper full of mineral oil to each of test tubes 5 and 6. Stopper and shake each.
16. Record the color of the mineral oil layer and the reaction results for I_2 in the Data Table.

Data for Part II

Color Reference Table for Part II

Halogen	Color in Mineral Oil	Halide Ion	Color in Mineral Oil

Result Color/Reaction Table

Reactants	$\text{Cl}_{2(\text{aq})}$	$\text{Br}_{2(\text{aq})}$	$\text{I}_{2(\text{aq})}$
$\text{Cl}^-_{(\text{aq})}$			
$\text{Br}^-_{(\text{aq})}$			
$\text{I}^-_{(\text{aq})}$			

Post-Lab Questions for Part II

1. Why is it assumed that the halide ions will not dissolve in mineral oil?
2. Write net ionic equations for all the reactions that occurred with the halogens.
3. List the reduction half-reactions of the halogens in order of decreasing activity.

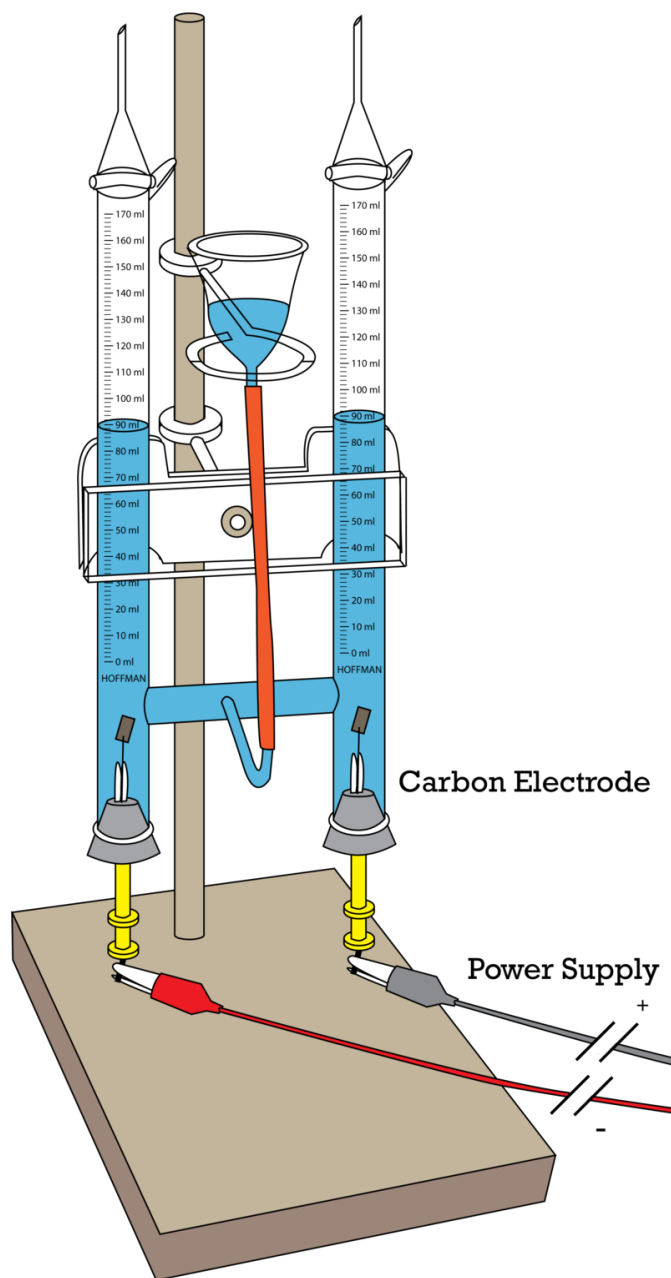
Demonstrations for Chapter 23

The Electrolysis of Water

Brief description of demonstration

A solution of sodium sulfate is electrolyzed using either a Hoffman apparatus. With the Hoffman apparatus, the gas collected in the separate tubes has a ratio of 2:1, and the colors in each tube change to yellow and blue. The gas is later collected into test tubes, and appropriate test for hydrogen and oxygen are done.

Hoffman Electrolysis Apparatus



Apparatus and Materials

- 1.5 L 1.0 M sodium sulfate solution, Na_2SO_4
- 60 mL bromothymol blue solution
- stirring rod
- disposable pipette
- DC power supply capable of delivering 25 volts at 2 amperes
- Hoffman apparatus
- Graphite or platinum electrodes
- 10 mL 0.1 M sulfuric acid, H_2SO_4
- 10 mL 0.1 M sodium hydroxide solution, $NaOH$
- Alligator clip leads for power supply

- 2 test tubes, 15 mm × 125 mm
- wooden splint and matches
- Ring stand
- Utility clamp

Procedure

Add the sodium sulfate solution to a large beaker or other clear container. Add 60 mL of bromothymol blue solution and stir. The solution should be green. If the solution is blue, add a few drops of H_2SO_4 and stir until the solution turns green. If it is yellow, add a few drops of $NaOH$ solution and stir until the solution turns green. This is an indication of a neutral solution. Close the stopcocks on the Hoffman apparatus and fill the bulb with the green solution. Once it is full, open the stopcock of one of the arms slightly, and the arm will fill with solution. Close the stopcock once it is full. Repeat this procedure with the other arm. Connect the wire leads to the electrodes on the Hoffman apparatus, and then connect the leads to the power supply. Turn on the power supply, and adjust the EMF to about 25 volts. Gas will collect into each arm, hydrogen from the negative side (the cathode) and oxygen at the positive side (the anode). The volume of hydrogen will be twice that of the volume of the oxygen, a nice proof of Avogadro's Law and the expression of the empirical formula of water. The solution at the anode turns yellow, and the solution at the cathode turns blue, due to the formation of H^+ and OH^- ions, respectively. After the hydrogen arm is about $\frac{3}{4}$ of the way full, turn off the power supply and disconnect the alligator clips. Perform the standard tests for hydrogen and oxygen by collecting test tubes of each by opening the stopcock of each tube, dispensing each gas into separate test tubes.

Hazards

Even though this demonstration is run at low voltage, the amperes are high, so there is an electrical shock hazard. The acid and base solutions can cause chemical burns. Use caution. **Do not mix the hydrogen and oxygen, as it can cause an explosion hazard.**

Disposal

Pour the solutions down the sink and flush the down with 100 times excess of water.

23.6 Worksheets for Chapter 23

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Balancing Redox Equations Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

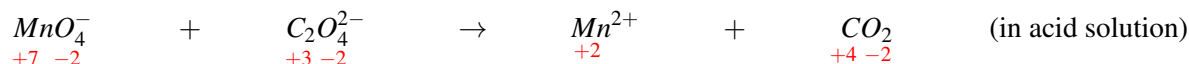
Steps in the balancing redox equations process.

- Determine the oxidation number for all atoms in the reaction.
- Determine which atom is being oxidized and which is being reduced.
- Write a half-reaction for the reduction process, showing the species containing the atom being reduced and the product containing that atom.
- Write a half-reaction for the oxidation process, showing the species containing the atom being oxidized and the product containing that atom.
- If the atoms being oxidized and reduced are not already balanced in the half-reactions, balance them.
- Add the appropriate number of electrons to each half-reaction needed to bring about the reduction and oxidation.
- Balance all other atoms in each half-reaction except H and O .
- Balance the H and O according to either (a) or (b) depending on whether the reaction is acidic or basic.
 - If the reaction is acidic, add H_2O and H^+ . Balance O first by adding H_2O , then balance H by adding H^+ . Charge should now be balanced.
 - If the reaction is basic, add OH^- and H_2O . Balance charge first by adding OH^- , then balance O by adding H_2O . The H should now be balanced.
- Once the half-reactions are balanced, find the lowest common multiple (LCM) for the electrons in the two half-reactions.
- Multiply each half-reaction by a whole number so that the total number of electrons in the reduction half-reaction equals the total number of electrons in the oxidation half-reaction, and they each equal the LCM.
- Add the two half-reactions and cancel those species that are common to both sides.
- Check the equation to be sure that it is balanced by both atoms and charge.

Example of an acidic redox reaction balancing.

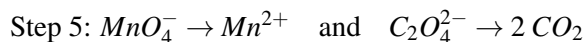
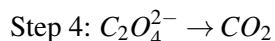
Given skeleton: $MnO_4^- + C_2O_4^{2-} \rightarrow Mn^{2+} + CO_2$ (in acid)

Step 1:

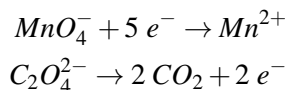


Step 2: Mn^{+7} is being reduced to Mn^{+2} and C^{+3} is being oxidized to C^{+4} .

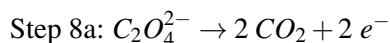
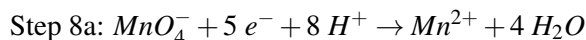
Step 3: $MnO_4^- \rightarrow Mn^{2+}$



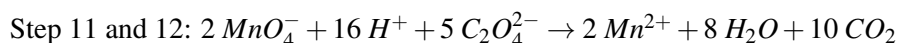
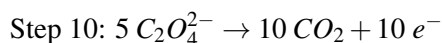
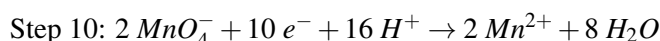
Step 6:



Step 7: All atoms other than H and O are balanced.



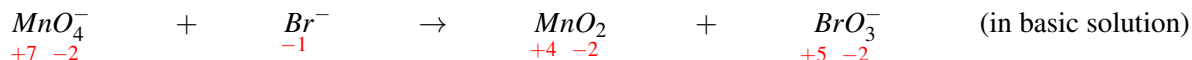
Step 9: The lowest common multiple for the electrons is 10. Therefore, we will multiply the reduction half-reaction by 2 and the oxidation half-reaction by 5.



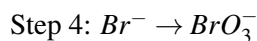
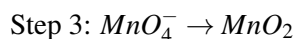
Example of an basic redox reaction balancing.



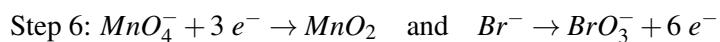
Step 1:



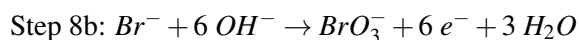
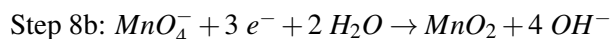
Step 2: Mn^{+7} is being reduced to Mn^{+4} and Br^- is being oxidized to Br^{+5} .



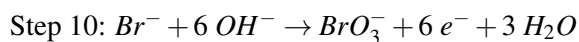
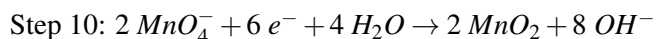
Step 5: Both the atoms being oxidized and the atoms being reduced are balanced in the half-reactions.



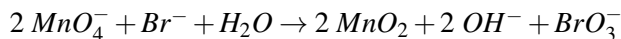
Step 7: All atoms other than H and O are balanced.



Step 9: The LCM for the electrons is 6. Therefore, we will multiply the reduction half-reaction by 2 and the oxidation half-reaction by 1.



Steps 11 and 12 (Cancel electrons, H_2O , and OH^-):



Exercises

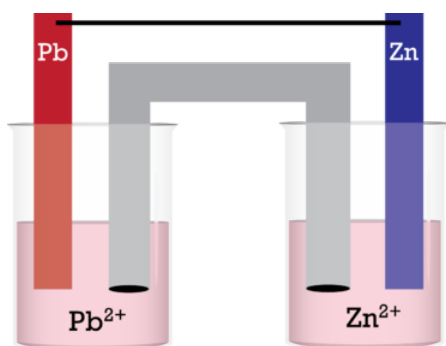
Balance the following redox equations.

- $Br_2 + SO_2 \rightarrow Br^- + HSO_4^-$ (in acidic solution)
- $PbO_2 + Mn^{2+} \rightarrow Pb^{2+} + MnO_4^-$ (in acidic solution)
- $MnO_4^- + SO_3^{2-} \rightarrow MnO_2 + SO_4^{2-}$ (in basic solution)
- $Zn + NO_3^- \rightarrow NH_3 + Zn(OH)_4^{2-}$ (in basic solution)
- $H_2O_2 + Cl_2O_7 \rightarrow ClO_2^- + O_2$ (in basic solution)

Electrochemical Cells Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

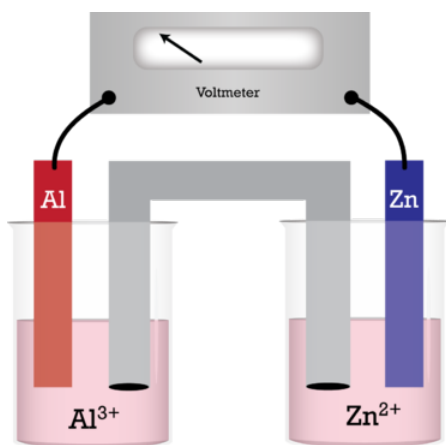


<u>Reduction Half Reactions</u>	<u>E°</u>
$Pb^{3+} + 2 e^- \rightarrow Pb$	-0.13 V
$Zn^{2+} + 2 e^- \rightarrow Zn$	-0.76 V

Use the standard cell sketched above to answer questions 1 - 9.

- Which electrode is the cathode?
 - Pb
 - Zn
 - Neither.
- Which electrode is the anode?
 - Pb
 - Zn
 - Neither.
- At which electrode will oxidation occur?
 - Pb
 - Zn
 - Neither.
- What is the maximum voltage for this standard cell?
 - 0.89 V
 - 0.63 V
 - 0.89 V
 - 0.63 V
 - 0.50 V

5. Which way do the electrons flow in the external circuit?
- From *Pb* to *Zn*.
 - From *Zn* to *Pb*.
 - No electron flow occurs.
6. Which way do cations flow through the salt bridge?
- Toward the *Pb* electrode.
 - Toward the *Zn* electrode.
 - No cation flow occurs.
7. What happens to the cell voltage when the reaction reaches equilibrium?
- Becomes maximum.
 - Drops to zero.
 - Becomes a positive value less than maximum.
8. Which electrode will gain mass as the cell runs?
- Pb*
 - Zn*
 - Neither.
9. What happens to the cell voltage as the cell runs?
- Remains constant.
 - Increases.
 - Decreases.
 - May increase or decrease.



Standard Reduction Half Reactions	E°
$\text{Al}^{3+} + 3 e^- \longrightarrow \text{Al}$	-1.66 V
$\text{Zn}^{2+} + 2 e^- \longrightarrow \text{Zn}$	-0.76 V

Use the standard cell sketched above to answer questions 10 - 21.

10. Which electrode is the anode?
- Al*
 - Zn*
 - Neither.
11. At which electrode does reduction occur?

A. *Al*

B. *Zn*

C. Neither.

12. What is the voltage of this standard cell?

A. 2.42 V

B. -2.42 V

C. 0.90 V

D. -0.90 V

E. 1.80 V

13. Which way do the electrons flow in the external circuit?

A. From *Al* to *Zn*.

B. From *Zn* to *Al*.

C. No electron flow occurs.

14. Which way do anions flow through the salt bridge?

A. Toward the *Al* electrode.

B. Toward the *Zn* electrode.

C. No cation flow occurs.

15. Which electrode loses mass as the cell runs?

A. *Al*

B. *Zn*

C. Neither.

16. How many moles of electrons pass through the external circuit in order for 1.00 mole of atoms to be deposited on the cathode?

A. 6

B. 3

C. 4

D. 2

E. 1

17. If 24 electrons pass through the external circuit, how many atoms of zinc must react?

A. 24

B. 12

C. 8

D. 4

E. 0

18. If 24 electrons pass through the external circuit, how many atoms of aluminum must react?

A. 24

B. 12

C. 8

D. 4

E. 0

19. What will happen to the voltage of the cell if the molarity of Zn^{2+} is increased?

A. Increase.

B. Decrease.

C. Remain the same.

20. What will happen to the voltage of the cell if the molarity of Al^{3+} is increased?

A. Increase.

B. Decrease.

C. Remain the same.

21. What will happen to the voltage of the cell if the salt bridge is removed?

A. Increase slightly.

B. Decrease slightly.

C. Remain the same.

D. Drop to zero.

22. In the two cells in this worksheet, there are a total of three reduction half-reaction indicated, *Al*, *Zn*, and *Pb*. Which of these three metals is most easily oxidized?

A. *Al*

B. *Zn*

C. *Pb*

23. Will a reaction occur if aluminum metal is placed in a solution of Zn^{2+} ?

A. Yes

B. No

24. Will a reaction occur if Pb metal is placed in a solution of Al^{3+} ?

A. Yes

B. No

25. Will a reaction occur if aluminum metal is placed in a solution of Zn^{2+} ?

A. Yes

B. No

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

23.7 Extra Reading for Chapter 23

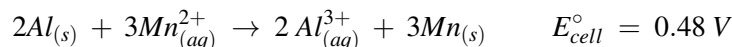
Dependence of Cell Potential on Concentration

Non-Standard Cells

For the most part, the discussion of cell potential in high school chemistry deals with cells under standard conditions. Standard conditions for cells is 25°C , 1.0 atm pressure, and the concentrations of ions is 1.0 M . If you were building a cell to use to do work, you would not build a standard cell. Standard cells are used essentially for teaching or experimentation. They do not have either the maximum voltage, or the maximum capacity, that can be built into Galvanic cells. The advantage of standard cells is that their voltages are precisely predictable and easily calculated.

The cell potential for galvanic cells is closely related to the net movement of materials from reactants to products. A faster net reaction would produce a greater cell potential, and a slower net reaction would produce a smaller cell potential. At equilibrium, the net reaction is zero and therefore, the cell potential would be zero. If the forward reaction rate is increased with no change in the reverse reaction rate, then the net forward reaction is greater, and the cell potential will be greater. If the reverse reaction rate is decreased with no change in the forward rate, then the net forward reaction is greater, and the cell potential will be greater.

Consider the cell composed of the standard half cells of aluminum and manganese.



The E_{cell}° for this reaction is determined when the concentrations of manganese ion and the aluminum ion are both 1.0 M . This is the voltage of this cell at standard conditions. If the concentration of the manganese ion is increased, the forward reaction rate will increase, and the net movement of material in the forward direction will increase. This increase in the net movement of material in the forward direction will cause the cell voltage to be higher. If the concentration of the aluminum ion is decreased, the reverse reaction rate will decrease, and the net movement of material in the forward direction will increase. This increase in the net movement of material in the forward direction causes the voltage to be higher. If the $[\text{Mn}^{2+}]$ concentration is decreased, the forward reaction rate will decrease, and the net movement of material in the forward direction will decrease. Therefore, the voltage of the cell will be lower. If the $[\text{Al}^{3+}]$ concentration is increased, the reverse reaction rate will be increased, and the net movement of material in the forward direction will decrease. Therefore, the voltage of the cell will be lower.

Cells that do not have the concentrations of ions at 1.0 M are called non-standard cells. In the cell described above, $[\text{Mn}^{2+}] > 1.0\text{ M}$ will cause the cell voltage to be greater than 0.48 V and $[\text{Al}^{3+}] > 1.0\text{ M}$ will cause the cell voltage to be less than 0.48 volts . Vice versa would be true if the concentrations were less than 1.0 M .

Cell voltages for non-standard cells can also be calculated using the Nernst Equation.

The Nernst Equation is $E = E^{\circ} - \left(\frac{0.0591}{n}\right) (\log Q)$, where E is the voltage of the non-standard cell. E° would be the voltage of these reactants and products if they were a standard cell, n is the moles of electrons transferred in the balanced reaction, and Q is the reaction quotient. The reaction quotient is the equilibrium constant expression, but when the reaction is not at equilibrium it is called the reaction quotient.

The reaction quotient for the example cell used here is $Q = \frac{[\text{Al}^{3+}]^2}{[\text{Mn}^{2+}]^3}$.

The Nernst Equation for this cell is $E = E^{\circ} - \left(\frac{0.0591}{n}\right) \left(\log \frac{[\text{Al}^{3+}]^2}{[\text{Mn}^{2+}]^3}\right)$.

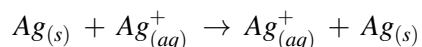
If you follow the mathematics for the case when both ions concentrations are 1.0 M , the reaction quotient would be 1 and the log of 1 is zero. Therefore, the second term in the Nernst Equation is zero and $E = E^\circ$.

Let's take the case of the example cell when $[Mn^{2+}] = 6.0 M$ and $[Al^{3+}] = 0.10 M$.

$$\begin{aligned}
 E &= E^\circ - \left(\frac{0.0591}{n}\right) \left(\log \frac{[Al^{3+}]^2}{[Mn^{2+}]^3}\right) \\
 E &= 0.48 V - \left(\frac{0.0591}{6}\right) \left(\log \frac{[0.10]^2}{[6.0]^3}\right) \\
 E &= 0.48 V - \left(\frac{0.0591}{6}\right) (-4.33) \\
 E &= 0.48 V - (-0.04 V) \\
 E &= 0.52 V
 \end{aligned}$$

Concentration Cells

If we attempt to construct a standard cell from the same two reactants, we do not get a reaction or a cell voltage. Suppose we attempt to build a cell with two silver half-cells.



If this is a standard cell, the half-cell voltage for the oxidation half-reaction is $-0.80 V$ and the half-cell voltage for the reduction half-reaction is $+0.80 V$. Clearly the net voltage is 0. It is possible, however, to produce a voltage using the same two half-reactions if we alter the concentrations of the ions. Such a cell is called a concentration cell and its voltage can be calculated using the Nernst Equation.

The Nernst Equation would look like this:

$E = E^\circ - \left(\frac{0.0591}{n}\right) \left(\log \frac{[Ag^+]_{\text{products}}}{[Ag^+]_{\text{reactants}}}\right)$, where the silver ion concentration in the numerator is the concentration of the silver ion in the products and the silver ion concentration in the denominator is the silver ion concentration in the reactants.

Suppose we build this cell using a concentration of silver ion in the products of 0.010 M and a silver ion concentration in the reactants of 6.0 M . The E° in this case is zero and $n = 1$.

$$\begin{aligned}
 E &= E^\circ - \left(\frac{0.0591}{n}\right) \left(\log \frac{[Ag^+]_{\text{products}}}{[Ag^+]_{\text{reactants}}}\right) \\
 E &= 0 V - (0.0591) \left(\log \frac{[0.010]}{[6.0]}\right) \\
 E &= 0 V - (0.0591)(-2.78 V) \\
 E &= 0 V - (-0.16 V) \\
 E &= 0.16 V
 \end{aligned}$$

23.8 Assessment for Chapter 23

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 24 Nuclear Chemistry C-SE-TE

Chapter Outline

- 24.1 THE DISCOVERY OF RADIOACTIVITY
- 24.2 NUCLEAR NOTATION
- 24.3 NUCLEAR FORCE
- 24.4 NUCLEAR DISINTEGRATION
- 24.5 NUCLEAR EQUATIONS
- 24.6 RADIATION AROUND US
- 24.7 APPLICATIONS OF NUCLEAR ENERGY
- 24.8 WORKSHEETS FOR CHAPTER 24
- 24.9 EXTRA READING FOR CHAPTER 24
- 24.10 ASSESSMENT FOR CHAPTER 24

Lessons and Number of Activities for Lessons

TABLE 24.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. The Discovery of Radioactivity	0	0	0	0
2. Nuclear Notation	0	0	1	0
3. Nuclear Force	0	0	0	0
4. Nuclear Disintegration	0	0	0	0
5. Nuclear Equations	0	0	1	0
6. Radiation Around Us	1	0	0	1
7. Applications of Nuclear Chemistry	0	0	1	1

24.1 The Discovery of Radioactivity

Student Behavioral Objectives

- The students will describe the roles played by Henri Becquerel and Marie Curie played in the discovery of radioactivity.
- The students will list the most common emissions from naturally radioactive nuclei.

Timing, Standards, Activities

TABLE 24.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>The Discovery of Radioactivity</i>	1.0	None

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for The Discovery of Radioactivity (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

24.2 Nuclear Notation

Student Behavioral Objectives

- The students will state the information contained in the atomic number of a nucleus.
- The students will state the information contained in the mass number of a nucleus.
- The students will subtract the atomic number from the mass number to determine the number of neutrons in a nucleus.
- Students will read and write complete nuclear symbols (know the structure of the symbols and understand the information contained in them).

Timing, Standards, Activities

TABLE 24.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Nuclear Notation</i>	1.0	11c

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Nuclear Notation Worksheet

Extra Readings

1. None

Answers for Nuclear Notation (L2) Review Questions

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

24.3 Nuclear Force

Student Behavioral Objectives

- Students will compare the energy released per gram of matter in nuclear reactions to that in chemical reactions.
- Students will express the equation for calculating the change in mass during nuclear reactions that is converted into energy.
- Students will express the relationship between nuclear stability and the nuclei's binding energy per nucleon ratio.

Timing, Standards, Activities

TABLE 24.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Nuclear Force</i>	1.0	11a, 11b, 11g

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Nuclear Force (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

24.4 Nuclear Disintegration

Student Behavioral Objectives

The student will:

- list some naturally occurring isotopes of elements that are radioactive.
- describe the three most common emissions during natural nuclear decay.
- express the changes in the atomic number and mass number of radioactive nuclei when an alpha particle is emitted.
- express the changes in the atomic number and mass number of radioactive nuclei when a beta particle is emitted.
- express the changes in the atomic number and mass number of radioactive nuclei when a gamma ray is emitted.
- express that protons and neutrons are not indivisible and are composed of particles called quarks.
- express the number of quarks that make up a proton or neutron.

Timing, Standards, Activities

TABLE 24.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Nuclear Disintegration</i>	1.5	11d, 11e, 11g

Activities for Lesson 4

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Nuclear Disintegration (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org

to request sample answers.

24.5 Nuclear Equations

Student Behavioral Objectives

- The students will give definitions and examples of fission and fusion.
- The students will classify nuclear reactions as fission or fusion.
- Given a nuclear equation with one species missing, the student will be able to correctly fill in the missing particle.
- Students will write balanced equations for nuclear transmutations.

Timing, Standards, Activities

TABLE 24.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Nuclear Equations</i>	1.5	1f, 11b, 11c

Activities for Lesson 5

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Nuclear Equations Worksheet

Extra Readings

1. None

Answers for Nuclear Equations (L5) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

24.6 Radiation Around Us

Student Behavioral Objectives

- Students will calculate the amount of radioactive material that will remain after an integral number of half-lives.
- Students will describe how carbon-14 is used to determine the age of carbon containing objects.
- Students will qualitatively compare the ionizing power and penetration power of α , β , and γ particles.

Timing, Standards, Activities

TABLE 24.7: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Radiation Around Us</i>	1.0	11c,11e,11f

Activities for Lesson 6

Laboratory Activities

1. Basic Nuclear Lab Kit

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. Carbon-14 Dating

Answers for Radiation Around Us (L6) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

24.7 Applications of Nuclear Energy

Student Behavioral Objectives

The student will:

- trace the energy transfers that occur in a nuclear power reactor from the binding energy of the nuclei to the electricity that leaves the plant.
- define the term “breeder reactor.”
- list some medical uses of nuclear energy.

Timing, Standards, Activities

TABLE 24.8: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Applications of Nuclear Energy</i>	1.0	1f, 11b

Activities for Lesson 7

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Radioactivity Worksheet

Extra Readings

1. Facts and Myths About Civilian Nuclear Power Plants

Answers for Applications of Nuclear Energy (L7) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 24

This website provides animations demonstrating balancing alpha and beta decay equations.

- <http://chemmovies.unl.edu/ChemAnime/BBETAD/BBETAD.html>
- <http://chemmovies.unl.edu/ChemAnime/ALPHAD/ALPHAD.html>

To learn more about Marie Curie, visit:

- http://www.time.com/time/specials/packages/article/0,28804,1848817_1848816_1848808,00.html#ixzz12DfQ1top

The following webpage provides more information about the discovery of radioactivity.

- http://www.chem.duke.edu/jds/cruise_chem/nuclear/discovery.html

This video discusses binding energy, fission and nuclear force.

- <http://www.youtube.com/watch?v=UkLkiXiOCWU>

Part 1 and 2 of A Funny 1950's Science Educational Cartoon Video about the Atom, the chemical elements, isotopes, electrons, protons, neutrons. The video also shows nuclear reactors and the benefits of peaceful uses of nuclear energy.

- <http://www.youtube.com/watch?v=UPb0hFtd27A>
- <http://www.youtube.com/watch?v=ZHfO14BtNxE>

For an introduction to the energy from the nucleus, see

- <http://www.youtube.com/watch?v=YAXpJN-gVR0>

This website provides more information about strong nuclear forces.

- <http://www.wisegeek.com/what-is-the-strong-nuclear-force.htm>

This video provides an overview of Alpha, Beta, Gamma Decay and Positron Emission.

- <http://www.youtube.com/watch?v=3koOwozY4oc>

For an introduction to quarks, see

- <http://www.youtube.com/watch?v=TGrDj5vFefQ>

The following is a video showing nuclear reactions.

- http://www.youtube.com/watch?v=_f4kYYAQC3c

Short animation showing the bending of α , β , and γ radiation under the influence of electrically charged plates.

- <http://www.mhhe.com/physsci/chemistry/chemistryessential/flash/radioa7.swf>

For a demonstration of the mouse trap model of a chain reaction, see

- <http://www.youtube.com/watch?v=HmbzJGf90Xc>

This video is an introduction to natural transmutations.

- <http://www.youtube.com/watch?v=I7WTQD2xYtQ>

These videos chronicle the search for cold fusion.

- <http://www.youtube.com/watch?v=vrs01XsmpLI>
- <http://www.youtube.com/watch?v=WQ1oghHOjHA>
- http://www.youtube.com/watch?v=4u_-h6IGb6E (10:00), <http://www.youtube.com/watch?v=6mBu7-y2mM>
- <http://www.youtube.com/watch?v=QqiJtv5DR-4>

A short animation of nuclear fission can be viewed at:

- http://www.classzone.com/cz/books/woc_07/resources/htmls/ani_chem/chem_flash/popup.html?layer=act&src=qtiwf_act129.1.xml

A short animation of nuclear fusion can be viewed at:

- http://www.classzone.com/cz/books/woc_07/resources/htmls/ani_chem/chem_flash/popup.html?layer=act&src=qtiwf_act130.1.xml

This website has several videos that show the tremendous effects of a nuclear explosion.

- http://www.atomicarchive.com/Movies/index_movies.shtml

For a lecture about types of nuclear decay including the ionizing and penetration power of the three major types of radioactive emissions, see

- <http://www.youtube.com/watch?v=aEBGE1Nm7vc>

The following website provides more information about background radiation.

- [www.hps.org/publicinformation/ate/cat10.html www.hps.org/publicinformation/ate/cat10.html]

The website below displays a map that shows the amount of environmental radiation across the United States.

- [www.radiationnetwork.com/RadiationNetwork.htm www.radiationnetwork.com/RadiationNetwork.htm]

For more details about nuclear power versus other sources of power, you can read the following report.

- *Nuclear Power VS. Other Sources of Power*, Neil M. Cabreza, Department of Nuclear Engineering, University of California, Berkeley, NE-161 Report. Available at <http://www.nuc.berkeley.edu/thyd/ne161/ncabreza/sources>.

The following web site contains a history of the discovery (creation) of the transuranium elements and includes some of the reactions used to produce them.

- <http://www.lbl.gov/abc/wallchart/chapters/08/0.html>

Laboratory Activities for Chapter 24

Basic Nuclear Lab Kit

Listed below is one commercial lab kit for nuclear labs. There are several others available from high school chemistry vendors. All of them with lab manuals.



The Basic Nuclear Lab Set is a complete student station designed for high school and college level nuclear science instruction. The system may be used with or without a computer, and consists of a lab station with built-in Geiger-Mueller counter, scaler with preset time, large 6-decade LED display, sample positioning tray and 11-piece radiation absorber set. The system features a variable high voltage power supply for plateau measurements, serial interface and cable with DATALINK software for data transfer to PC. System includes alpha, beta and gamma radioisotope sources & experiment manuals. Data files are stored in spreadsheet compatible format for graphing and manipulation prior to printing.

To ensure the longest half-life possible, all isotopes are fabricated when the order is placed and shipped directly to you from the NRC licensed nuclear isotope manufacturer in Oak Ridge, TN. The system may be used with Nuclear lab station with GM counter, serial interface and DATALINK software.

Serial cable for PC

Set of three radioisotope sources, 1 each alpha, beta, and gamma disks

Nuclear science experiments manual with sample data. CD version.

Basic Nuclear Lab Kit: \$985.00

Demonstrations for Chapter 24

24.8 Worksheets for Chapter 24

Copy and distribute the lesson worksheets. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Nuclear Notation Worksheet

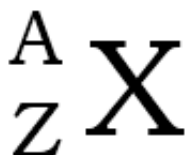
CK12 Foundation Chemistry

Name _____ Date _____

Background

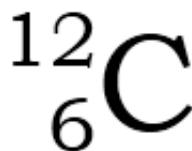
Each element has one or more atoms associated with it. Every atom of a particular element has the same number of protons. A particular atom of an element does not have to have the same number of neutrons. A group of atoms of an element must all have the same number of protons but they may have several different numbers of neutrons. The atomic number of an atom indicates the number of protons in its nucleus and the mass number of an atom indicates the total number of protons and neutrons in its nucleus. Atoms with the same atomic number but a different mass number are called isotopes.

Nuclear notation is a shorthand way of writing information about a particular atom. An example of the accepted form of nuclear notation is shown below.



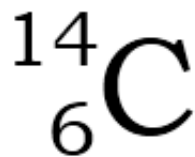
The large “X” represents the symbol for the element. The mass number is represented in the image by the letter “A” and is positioned as a superscript preceding the X. The atomic number is represented by the letter “Z” and is positioned as a subscript preceding the X.

Here is an example of this notation using a carbon isotope that contains six protons and six neutrons in its nucleus.



We see a carbon atom represented with an atomic number of 6, showing this atom, like all carbon atoms, has 6 protons in the nucleus. The mass number of 12 shows the nucleus contains a total of 12 protons and neutrons. Since we know the carbon atom has 6 protons, a subtraction of the atomic number from the mass number indicates the nucleus contains 6 neutrons.

Another isotope of carbon is a carbon nucleus that contains 6 protons and 8 neutrons. The nuclear notation for this isotope of carbon appears below.



In the nuclear notation for this isotope of carbon, the mass number is 2 greater because the nucleus contains 2 more neutrons and the same number of protons.

Exercises

1. An atom contains 3 protons, 3 neutrons, and 3 electrons. What is its atomic number?
2. An atom contains 3 protons, 3 neutrons, and 3 electrons. What is its mass number?
3. Indicate the number of protons, neutrons, and the mass number for the following nuclear symbol.

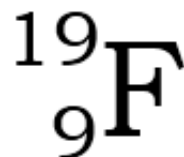


Number of protons = _____

Number of neutrons = _____

Mass number = _____

4. Indicate the number of protons, neutrons, and the mass number for the following nuclear symbol.

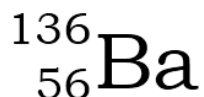


Number of protons = _____

Number of neutrons = _____

Mass number = _____

5. Indicate the number of protons, neutrons, and the mass number for the following nuclear symbol.

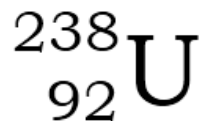


Number of protons = _____

Number of neutrons = _____

Mass number = _____

6. Indicate the number of protons, neutrons, and the mass number for the following nuclear symbol.



Number of protons = _____

Number of neutrons = _____

Mass number = _____

7. Write the nuclear symbol for an isotope of neon whose nucleus contains 10 protons and 10 neutrons.

8. Write the nuclear symbol for an isotope of bromine whose nucleus contains 35 protons and 45 neutrons.

Nuclear Equations Worksheet

CK-12 Foundation Chemistry

Name _____ Date _____

1.



Name the nuclear particle indicated by the nuclear symbol shown above. _____

2.



Name the nuclear particle indicated by the nuclear symbol shown above. _____

3.



Name the nuclear particle indicated by the nuclear symbol shown above. _____

4.



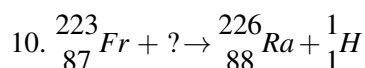
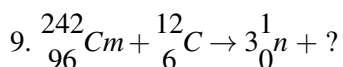
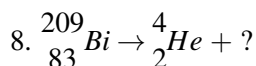
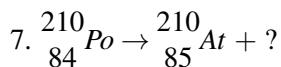
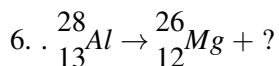
Name the nuclear particle indicated by the nuclear symbol shown above. _____

5.



Name the nuclear particle indicated by the nuclear symbol shown above. _____

In questions 6 – 10, a single nuclear particle is missing. Write the complete nuclear symbol for the missing particle.



Multiple Choice

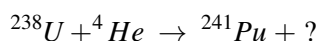
11. An isotope of bismuth, ${}^{209}\text{Bi}$, is bombarded with a proton. The produce of the ensuing reaction is an isotope of element X and two neutrons. What is the atomic number of this isotope of element X?

A. 82 B. 83 C. 84 D. 85 E. None of these.

12. An isotope of bismuth, ${}^{209}\text{Bi}$, is bombarded with a proton. The produce of the ensuing reaction is an isotope of element X and two neutrons. What is the mass number of this isotope of element X?

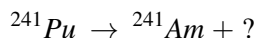
A. 208 B. 209 C. 210 D. 211 E. None of these.

13. Which of the following particles completes this equation?



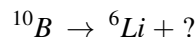
A. beta B. alpha C. proton D. neutron E. None of these.

14. Which of the following particles completes this equation?



A. beta B. alpha C. proton D. neutron E. None of these.

15. Which of the following particles completes this equation?



A. beta B. alpha C. proton D. neutron E. None of these.

16. If ${}^{234}\text{Th}$ undergoes beta decay, which of the following will be the resultant particle?

A. ${}^{234}\text{Ra}$ B. ${}^{230}\text{Th}$ C. ${}^{234}\text{Pa}$ D. ${}^{235}\text{U}$ E. None of these.

17. ^{234}U undergoes alpha decay and the resultant particle undergoes beta decay. What is the final particle after both decays?
A. ^{236}Np B. ^{230}Pa C. ^{232}Ac D. ^{239}Np E. ^{233}Pa
18. Write the nuclear equations for the decay of ^{210}Po if it undergoes two consecutive alpha decays followed by a beta decay.

Radioactivity Worksheet

CK12 Foundation Chemistry

Name _____ Date _____

- What does it mean when an element is radioactive?
A. it means that some of the atoms of the element emit radiation B. it means that some of the atoms of the element are changing into atoms of different elements C. it means that some of the atoms of the element are becoming more stable D. All of the above are true.
- Stable systems have more potential energy than unstable systems.
A. True B. False
- Gamma rays are emitted along with alpha and beta radiation. What do gamma rays account for?
A. lost mass B. lost charge C. lost energy D. All of the above.
- Within a nuclear reactor, the purpose of the moderator is to
A. absorb neutrons in the reactor core B. absorb neutrons in the outer containment structure C. slow down neutrons in the reactor core D. slow down neutrons in the outer containment structure
- Which type of radiation is most similar to high-energy x-rays?
A. alpha B. beta C. neutron D. gamma
- Which of the noble gases is naturally radioactive and has no stable isotopes?
A. Ar B. Kr C. Xe D. Rn
- Which of the following best describes the operation of a cyclotron?
A. an uncharged particle is accelerated to great speeds by a fast moving current of air B. a charged particle is accelerated by changing charges in a series of pipes or tubes C. an uncharged particle is accelerated by alternating charges on adjacent Dees. D. a charged particle is accelerated by alternating charges on adjacent Dees while being subjected to a magnetic field which causes the particle to move in a curved path.
- Which of the following changes occurs in a nucleus when a positron is given off?
A. a proton is produced B. a neutron is lost C. a helium-4 nucleus is lost D. a neutron is split into a proton and an electron E. a neutron is produced
- Which of the following particles could not be accelerated in a linear accelerator?
A. proton B. neutron C. helium nucleus D. electron E. All of the particles listed above could be accelerated in a linear accelerator.
- Which of the following statements is false concerning nuclear radiation?
A. exposure to radiation will always make you radioactive B. solar radiation is more dangerous than nuclear radiation C. radiation is always good for you D. man is responsible for all radiation E. All of the above are false.

11. An unknown type of radiation was giving a reading on a Geiger counter of 2000 counts per second. When a piece of paper was placed between the source and Geiger counter, the reading dropped to 50 counts per second. Based on this observation, you could conclude that the radiation was mostly
A. alpha B. beta C. gamma D. x-ray E. insufficient data
12. Nuclear radiation was discovered by
A. Albert Einstein B. Ernest Rutherford C. Henri Becquerel D. Marie Curie E. Pierre Curie
13. 20.0 *grams* of a radioactive element is prepared in a nuclear reactor. The half-life of the isotope is 3.0*days*. How many days will it take before there is only 2.50 *grams* of the substance remaining?
A. 1.5 *days* B. 3.0 *days* C. 6.0 *days* D. 9.0 *days* E. 12 *days*
14. Element X has only two isotopes. One of the isotopes has a mass number of 190 and the other has a mass number of 194. If the atomic mass of the element is 193.6, which of the two isotopes is most commonly found in nature?
A. 190 B. 194 C. 193.6 *days* D. The two isotopes are equally common. E. Insufficient data to determine.
15. A 12.0 *gram* sample of the isotope ^{131}I decays to 1.5 *grams* in 24.0 *days*. What is the half-life of this isotope?
A. 8.0 *days* B. 16 *days* C. 20. *days* D. 24 *days* E. None of these.
16. Define the following terms.
- A. mass defect
 - B. chain reaction
 - C. critical mass
 - D. fission
 - E. fusion
 - F. beta decay
 - G. ionizing power
 - H. binding energy

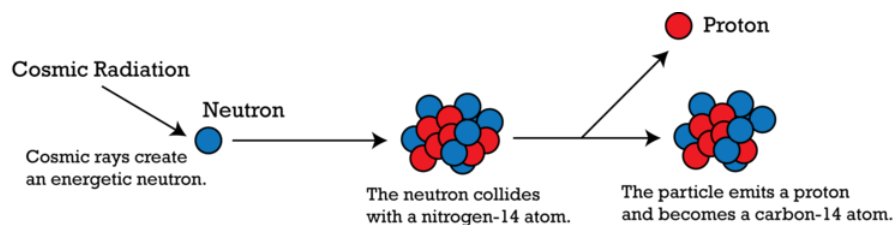
Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

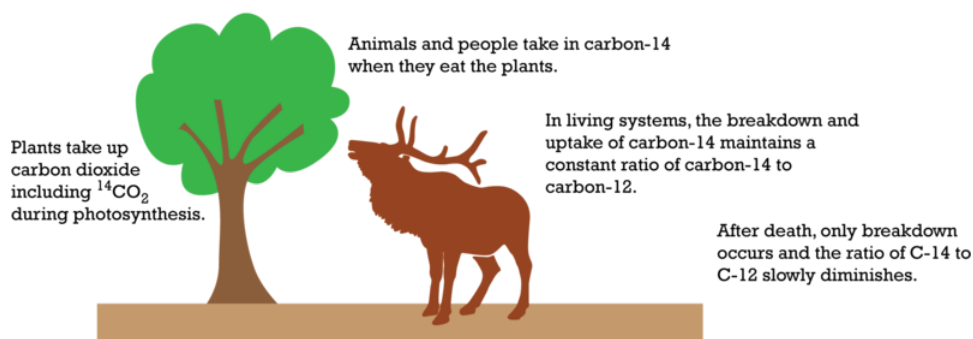
24.9 Extra Reading for Chapter 24

Carbon-14 Dating

If you're a fan of the television CSI shows or other mystery or crime programming, you are probably aware of different means to estimate the timing of the poor unfortunate's demise, dealing with factors such as body temperature, etc. For an archaeologist or an anthropologist, the trail of evidence is much colder. In the late 1940's, Willard Libby of the University of Chicago devised a method to establish the age of even the oldest unearthed fossils based on the remaining amount of radioactive ^{14}C . This isotope is one of three for the element carbon, which is ubiquitous in living systems. Carbon has an atomic number of six, and exists in three nuclear configurations: with six, seven and eight neutrons respectively. Thus the ^{14}C isotope has a nucleus consisting of six protons and eight neutrons. This assembly spontaneously decays into Nitrogen-14 and the release of beta radiation. Radioisotopic carbon has been measured to decay at a constant rate, with half the initial amount remaining, after 5730 years. If it is assumed that the ^{14}C is not replaced, the loss of ^{14}C suggests the time interval since the artifact last exchanged CO_2 with the atmosphere.



The ^{14}C reacts with O_2 in the atmosphere to become $^{14}\text{CO}_2$.



When the ratio of the remaining amount of ^{14}C to ^{12}C is compared to the same ratio in a living organism, the amount of time elapsing since the organism's death can be analyzed.

Thus, over time, in any material that contains the element carbon, the amount of remaining ^{14}C in a sample is an indicator of the age of the artifact.

One of the best-known applications of this technique was in the analysis of Ötzi, the alpine Ice Man. Found in a region straddling the Austrian-Italian border, by hikers in 1991, Ötzi was the name given to the partially mummified remains of a hunter located still frozen into a glacier. Ötzi provided a wealth of information to anthropologists in that he was still dressed with fur boots, and a pack with tools including a copper hatchet and arrows. Researchers

even discovered the menu of his last meal, probably deer meat and wheat bran, by analyzing his stomach contents. Analysis of small tissue samples from his corpse suggest that he lived from 5300 to 5100 years ago, before the Bronze Age. The construction and contents of his clothing and possessions provide an invaluable insight into the culture and technological sophistication of his age.

Other applications of this technology include dating of the Dead Sea scrolls, and analysis of the time period of the cave art found in central Europe. Ocean sediment samples, and even a meteorite believed to have originated on Mars!

Facts and Myths About Civilian Nuclear Power Plants

- MYTH: Nuclear reactors may undergo a nuclear explosion killing tens of thousands of people.
- FACT: Civilian nuclear power plants in the U.S. never contain a supercritical mass of fissionable material and therefore, cannot explode even if operators tried to make them explode.
- MYTH: Nuclear power plants are not safe.
- FACT: 1. The radiation levels measured outside of the containment building of nuclear power plants are essentially the same as background radiation. 2. There was a nuclear accident in 1986 at the Chernobyl nuclear installation in the Soviet Union that resulted in the immediate death of 28 people (mostly employees and fire fighters), the subsequent death of 19 people, and 9 deaths from thyroid cancer apparently due to the accident. The number of injuries due to fall-out radiation from the accident is unknown (at that time, news from the Soviet Union was highly censored). Predictions of numbers of injuries by UNSCEAR (United Nations Scientific Committee on the Effects of Atomic Radiation) were in excess of 4,000 but the number was disputed by the IAEA (International Atomic Energy Agency).

In either a nuclear explosion or a severe nuclear accident, a large amount of radioactive material rises into the upper atmosphere. At some later time, this radioactive material falls back to the earth, usually downwind from where the explosion or accident occurred. This material that falls back to the earth is known as **fall-out**. It is known that the fall-out from the Chernobyl accident was 30 times the fall-out from the Hiroshima and Nagasaki bombings and that 300,000 people were evacuated from the Chernobyl area. It is also known that the fall-out reached as far as Sweden because Swedish workers at a Swedish nuclear power plant began testing positive for radiation, and after a thorough check of their own plant, officials determined the radiation was coming from fall-out from the Chernobyl plant. The greatest threat from fall-out is to children because children are growing and developing rapidly and radiation like all toxic materials have the greatest effect on children. It may still be years before all the effects of the Chernobyl accident are known. A small area called the Chernobyl Exclusion Zone is still closed, but the rest of the fall-out area is now considered safe.

The Chernobyl accident occurred in an early model of Soviet reactor that had no reaction vessel and no containment building. The fuel rods and control rods were inserted into graphite blocks. The graphite blocks worked well as a moderator, but graphite is combustible. When a fire started in the reactor, the employees and emergency workers were unable to control it, and the graphite burned away releasing radiation into the environment. There are some military reactors of this type in the United States, such as the one at Hanford, WA and there have been some radiation injuries at that facility. Civilian nuclear power plants in the U.S., however, do not use the graphite block reactor design. All U.S. civilian nuclear generating plants use LWR reactors, which have the reactor core submerged in a vessel of water and surrounded by a containment building. They also have a series of fail-safe shut down safety measures.

- MYTH: Hundreds of uranium miners die every year from radiation sickness.
- FACT: Hundreds of uranium miners died in the early days of uranium mining but that problem was solved long ago. On the other hand, 5,000 coal miners continue to die every year worldwide due to cave-ins, explosions, and black lung disease.

- MYTH: Nuclear reactors produce a large amount of radioactive waste that will be dangerous for thousands of years.
- FACT: When the percentage of $U - 235$ in fuel rods gets below a certain level, they will no longer function as fuel and must be replaced. Even though the radioactivity is too low to function as fuel rods, they are still extremely dangerous, and must be isolated for a long period of time. Several suggestions have been made for storage of this used fuel, but even though the method is considered safe by nuclear scientists, the people who live in the area where the waste is to be stored strongly oppose having the material stored near them. At the present time, the used fuel rods are still submerged in the reaction vessels where they were replaced. Now that the US government has removed the ban on recycling used fuel, the amount of radioactive waste will be reduced to approximately one-fourth of the present amount. Not only will recycling help with the waste disposal problem, it will also reduce the cost of fuel. The Department of Energy is considering space disposal (rocket the waste into the sun), geological disposal (burying the waste thousands of feet underground in geologically stable areas), transmutation disposal (building a nuclear reactor that will consume nuclear waste; this idea was banned by President Carter, reinstated by President Reagan, and was being investigated by President Bush). For a complete discussion on the handling of nuclear waste, the internet has several dozen sites.
- MYTH: Nuclear reactors are particularly vulnerable to terrorist attack.
- FACT: In 1988, Sandia National Laboratories conducted a test by slamming a military F-4 Phantom jet fighter into a concrete block built to simulate a nuclear reactor containment building. The airplane hit the block at 481 *miles* per hour and while the airplane was demolished, the six-foot thick wall suffered a dent 2.5 *inches* deep. The Turkey Point Nuclear Generating station (near Miami, Florida) suffered a direct hit by hurricane Andrew in 1992. Turkey Point has two fossil fuel units and two nuclear units. The fossil fuel plants suffered \$90 million of damage while the nuclear containment buildings were undamaged.
- COMPLAINT: Some countries may divert nuclear reactor materials to weapon building.
- FACT: True.
- COMPLAINT: When the cooling water from nuclear power plants is dumped back into the original source (river, lake, bay), the temperature of the water over a period of time may be raised several degrees. The amount of oxygen that water will hold in dissolved form is highly dependent on the temperature. Active fish (so-called sport fish) frequently move away from areas where the temperature has increased a few degrees and less active fish (so-called trash fish) move in. To keep everyone happy, the use of cooling towers needs to be greatly increased so that the water returned to its natural source is at the same temperature as when it was taken.
- FACT: True.

24.10 Assessment for Chapter 24

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.

CHAPTER 25 Organic Chemistry C-SE-TE

Chapter Outline

- 25.1 CARBON, A UNIQUE ELEMENT
- 25.2 HYDROCARBONS
- 25.3 AROMATICS
- 25.4 FUNCTIONAL GROUPS
- 25.5 BIOCHEMICAL MOLECULES
- 25.6 WORKSHEETS FOR CHAPTER 25
- 25.7 EXTRA READING FOR CHAPTER 25
- 25.8 ASSESSMENT FOR CHAPTER 25

Lessons and Number of Activities for Lessons

TABLE 25.1: Lessons and Activities for Lessons

Lesson	No. of Labs	No. of Demos	No. of Worksheets	No. of Extra Readings
1. Carbon, A Unique Element	0	0	0	0
2. Hydrocarbons	0	0	1	0
3. Aromatics	0	0	0	0
4. Functional Groups	1	0	0	0
5. Biochemical Molecules	0	1	0	2

25.1 Carbon, A Unique Element

Student Behavioral Objectives

The student will:

- describe the allotropes of carbon.
 - describe the hybridization of carbon.
 - explain how the hybridization of carbon allows for the formation of large number of compounds containing carbon.
-

Timing, Standards, Activities

TABLE 25.2: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Carbon, A Unique Element</i>	1.0	10b

Activities for Lesson 1

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Carbon, A Unique Element (L1) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

25.2 Hydrocarbons

Student Behavioral Objectives

The student will:

- define alkanes as well as name and draw alkanes.
- define alkenes as well as name and draw alkenes.
- define alkynes as well as name and draw alkynes.
- define structural formula.
- define isomers and be able to draw isomers for alkanes, alkenes, and alkynes.
- define substituted halogens as well as name and draw substituted halogens.

Timing, Standards, Activities

TABLE 25.3: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Hydrocarbons</i>	2.0	10d

Activities for Lesson 2

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. Organic Nomenclature

Extra Readings

1. None

Answers for Hydrocarbons (L2) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

25.3 Aromatics

Student Behavioral Objectives

The student will:

- describe the bonding in benzene.
 - define aromaticity.
 - name simple compounds containing benzene.
 - draw simple compounds containing benzene.
-

Timing, Standards, Activities

TABLE 25.4: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Aromatics</i>	1.5	10d

Activities for Lesson 3

Laboratory Activities

1. None

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Aromatics (L3) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

25.4 Functional Groups

Student Behavioral Objectives

The student will:

- identify alcohols, aldehydes, ketones, ethers, organic acids, and esters based on their functional groups.
- name and draw simple alcohols, aldehydes, ketones, ethers, organic acids, and esters.

Timing, Standards, Activities

TABLE 25.5: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Functional Groups</i>	2.0	10e

Activities for Lesson 4

Laboratory Activities

1. Synthesis of Esters

Demonstrations

1. None

Worksheets

1. None

Extra Readings

1. None

Answers for Functional Groups (L4) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

25.5 Biochemical Molecules

Student Behavioral Objectives

The student will:

- describe the basic structure of fatty acids, monosaccharides, and proteins.
- identify the chemical purpose fulfilled by lipids, carbohydrates, and enzymes.
- describe the biological function of hemoglobin and DNA.

Timing, Standards, Activities

TABLE 25.6: Timing and California Standards

Lesson	Number of 60 min periods	CA Standards
<i>Biochemical Molecules</i>	2.0	10a, 10c, 10f

Activities for Lesson 5

Laboratory Activities

1. None

Demonstrations

1. Cuprammonium Rayon

Worksheets

1. None

Extra Readings

1. Trans Fats
2. Ozone's Role in the Atmosphere

Answers for Biochemical Molecules (L5) Review Questions

- Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.

Multimedia Resources for Chapter 25

This website provides an introduction to Organic Chemistry.

- http://www.visionlearning.com/library/module_viewer.php?mid=60

This video describes the history of organic chemistry.

- <http://www.youtube.com/watch?v=HJnlNpMSlbs>

The *learner.org* website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos but there is no charge. The website has one video that relates to this lesson called “Carbon.”

- <http://learner.org/resources/series61.html>

The *learner.org* website allows users to view the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The video called “Molecular Architecture” examines isomers (primarily organic molecules) and how the electronic structure of a molecule’s elements and bonds affects its shape and physical properties.

- http://www.learner.org/vod/vod_window.html?pid=801

This video shows an animation of the formation of a protein.

- <http://www.youtube.com/watch?v=w-ctkPUUpUc>

This video explains how amino acids form the polypeptide backbone structure of proteins.

- http://www.youtube.com/watch?v=0_WaQniUU-g

The *learner.org* website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos but there is no charge. The website has one video that relates to this lesson called “Proteins: Structures and Function.”

- http://www.learner.org/vod/vod_window.html?pid=815

This video is a ChemStudy film called “Synthesis of an Organic Compound.” The film is somewhat dated but the information is accurate.

- <http://www.youtube.com/watch?v=ToSmwYgbvI0>

Laboratory Activities for Chapter 25

Teacher’s Pages for Esterification

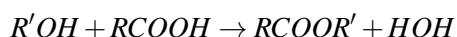
Lab Notes

The following laboratory produces esters by reacting a carboxylic acid with an alcohol. The reaction is catalyzed by the addition of sulfuric acid: the concentrated acid absorbs the water produced, driving the reaction towards producing more products in accordance with Le' Chatelier's principle. Nonetheless, the reaction only yields 60% ester under the best conditions. This is typical of organic reactions. The ester can be purified further by distillation, but using simple distillation is not enough due to reaction reversibility as the distillate is collected. Azeotrope formation is also a problem. Simple column chromatography or HPLC can be used, but these techniques are either not available in the high school laboratory, or not appropriate due to time constraints. Handling these materials is to be done with great caution. The carboxylic acids often have a bad smell, and can leave lingering odors that are offensive even when the majority of the acid has been disposed of. The sulfuric acid is 18 M, and will char tissue if exposed. The odors of the ester must be done by wafting, and should be done after the mixture has cooled. Otherwise, the mole fraction of vapor at higher temperatures tends to be towards the excess acid, which is very irritating. The odor of ester can also be irritating. Students who have symptoms of vapor intoxication should be removed to fresh air. Isoamyl alcohol vapor is extremely toxic, and must be handled in the fume hood. It is deceiving because it has such a pleasant odor.

The water baths should be kept at or slightly below 60°C. Methanol boils at 62°C.

Answers to Pre-Lab Questions

1. The reaction is dehydration synthesis:



2. The sulfuric acid absorbs the water produced, driving the reaction towards products.

Synthesis of Esters

Background Information

Whenever you eat almost any fruit or many other foods, the flavor that you taste is due primarily to an organic compound called an ester. Because esters are easily synthesized, many are produced commercially. Food and beverage companies use them to enhance the flavors in ice cream, sherbet, gelatin desserts, cake mixes, soft drinks, candy and chewing gum. Sometimes, the synthetic ester is used in place of the naturally derived flavoring. Why use only artificial flavoring? Many natural flavoring contain ingredients which when heated develop bitter or sour tastes. Some also have poor shelf lives. Availability and cost are important factors. The demand for flavoring has become so immense that without the synthetic esters there simply would not be enough for food manufacturers. The synthetic flavoring may not taste the same as the natural analogue. The reason is that the ester is only one of many organic components of a well-known flavor. As shown in Table 1, pineapple flavor contains seventeen different compounds in varying amounts. Food and flavoring companies are becoming increasingly adept at imitating nature, so that many artificial flavorings are complex mixtures capable of deceiving even professional food tasters. One method of preparing esters is the reaction of an alcohol with a carboxylic acid in a process called esterification. This reaction requires a catalyst, in this case, sulfuric acid.

Pre-Lab Questions

1. What type of chemical reaction produces esters?
2. What is the role of the sulfuric acid in this experiment?

Purpose

To prepare esters by the reaction of alcohol and carboxylic acid, and to observe and record their odors.

Apparatus and Materials

- Test tubes, 20 mm × 200 mm or similar
- 600 mL beaker
- Electronic Balance
- Disposable pipettes
- 10 mL graduate
- Hot Plate
- An assortment of alcohols: Methanol, octanol, ethanol, isoamyl alcohol
- An assortment of organic acids such as salicylic acid, butyric acid, acetic acid, anthranilic acid.

Safety Issues

Sulfuric acid is highly corrosive and must be handled with caution. Clean up any spillages immediately. If any acid touches your skin, immediately rinse with copious amounts of water.

Acetic and butyric acids have strong, irritating odors. If possible, contain their use to the fume hood.

Procedure

Place the prescribed amount of your carboxylic acid in a large test tube. Add your alcohol and swirl the tube until the material dissolves. Cautiously add 5 drops of concentrated sulfuric acid to the mixture and place the test tube in a hot water bath for 15 – 20 minutes. At this time remove the test tube from the bath and allow it to cool to room temperature. Cautiously draw the vapors from the test tube to your nose by your cupped hand. Identify the ester formed by its aroma. Dispose of the materials into a waste container provided by the teacher. Repeat this procedure with the other acid – alcohol pairings.

TABLE 25.7: short caption

Ester Produced	Acid	Amount	Alcohol	Amount
Isoamyl Acetate	Ethanoic Acid	2 mL	Isoamyl Alcohol	2 mL
Octyl Acetate	Ethanoic Acid	2 mL	Octanol	3 mL
Methyl Salicylate	Salicylic Acid	0.7 g	Methanol	3 mL
Methyl Butyrate	Butyric Acid	2 mL	Methanol	1.75 mL
Ethyl Butyrate	Butyric Acid	2 mL	Ethanol	2.5 mL
Methyl Anthranilate	Anthranilic Acid	0.7 g	Methanol	3 mL

Pineapple Flavoring

- allyl caproate 5%
- oil of sweet birch 1%
- isoamyl acetate 3%
- oil of spruce 2%
- isoamyl isovalerate 3%
- Balsam Peru 4%
- ethyl acetate 15%
- mustard oil 1%
- ethyl butyrate 22%
- Cognac oil 5%
- terpinyl propionate 3%
- orange oil 4%
- ethyl crotonate 5%
- lime oil 2%
- acetic acid 5%

TABLE 25.8: Data for Part I

Ester	Flavor	Observed Odor
Isoamyl Acetate	Banana	
Octyl Acetate	Orange	
Methyl Butyrate	Apple	
Ethyl Butyrate	Pineapple	
Amyl Butyrate	Apricot	
Methyl Salicylate	Wintergreen	
Methyl Anthranilate	Grape	

Post-Lab Questions

1. Based on the odors produced by your esterification reaction, what esters were formed?
2. What combination of carboxylic acid and alcohol were used to produce your ester?
3. Isobutyl formate is the chemical ester used in the production of raspberry flavoring – what combination of carboxylic acid and alcohol would result in this ester?

Demonstrations for Chapter 25

Cuprammonium Rayon

Brief description of demonstration

A blue solution containing dissolved cellulose is discharged with a syringe or a pipette into an acid solution. A thread of rayon is generated.

Materials

- 25 g copper(II) sulfate-5-hydrate, $CuSO_4 \cdot 5H_2O$
- 100 mL distilled water
- 200 mL concentrated ammonium hydroxide solution (15 M)
- 300 mL 0.5 M sulfuric acid, H_2SO_4
- 250 mL beaker
- stirring rods
- magnetic stirrer and stir bar
- Buchner funnel, 11 cm
- 11 cm filter paper
- 1 L vacuum filter flask
- 2 400 mL beakers
- Large disposable pipette or 10 mL syringe

Procedure

Under a fume hood, dissolve 25 g of the copper(II) sulfate-5-hydrate in 100 mL of water using the magnetic stirrer. Slowly add 13 mL of the concentrated ammonium hydroxide solution into the copper sulfate solution. A gelatinous blue precipitate of copper(II) hydroxide will form.

Using a Buchner funnel, filter the precipitate away from the supernatant using vacuum filtration. Wash the precipitate with (3) 50 mL portions of cold distilled water to remove the ammonium sulfate. This washing will take a little while because the particle size is small.

Place the filtered precipitate and the filter itself in a 400 *mL* beaker. Place the beaker onto a magnetic stirrer in the fume hood. Add 150 *mL* of concentrated ammonium hydroxide and stir. The paper and the precipitate should dissolve. Tear up 3 or 4 more pieces of filter paper, and add them to the beaker. The paper will dissolve, though it will take about an hour.

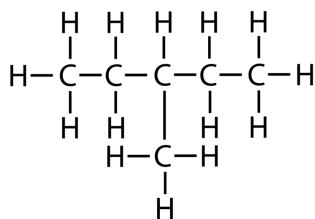
To perform the demonstration, add about 300 *mL* of the sulfuric acid solution to a 400 *mL* beaker. Draw some of the blue ammonia/copper/paper solution into a 10 *mL* syringe or disposable pipette. Place the tip of the pipette or syringe underneath the surface of the acid in the beaker, and inject the blue solution into the acid. A blue strand of cuprammonium rayon will form.

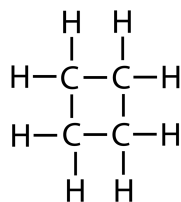
Hazards

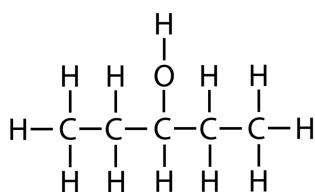
The ammoniacal solution of copper(II) ion is very caustic, and the sulfuric acid is corrosive. Avoid contact or inhalation of the ammonia fumes. Copper compounds are toxic. Rinse the generated rayon strand thoroughly with water before handling.

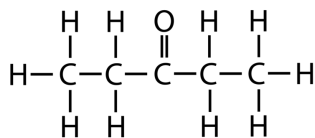
Disposal

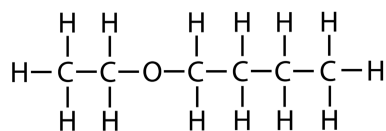
Separately rinse each solution down the drain with 100 times excess of water. Do not pour them into the sink together, or you could clog your pipes with rayon!

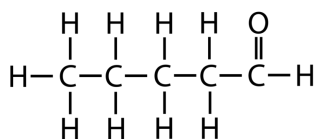


 5.


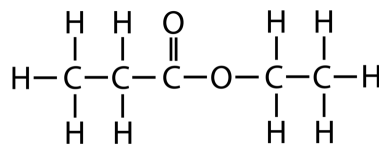
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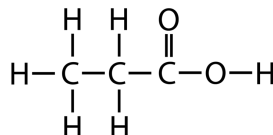
 8.


 9.


10.



11.



Draw the following molecules.

12. 1-Butyne

13. Methoxyethane

14. Butanal

15. 1, 2-Dibromopropane

Answers to Worksheets

- The worksheet answer keys are available upon request. Please send an email to teachers-requests@ck12.org to request the worksheet answer keys.

25.7 Extra Reading for Chapter 25

Trans Fats

Twenty-first century Americans are becoming increasingly cognizant of the role of dietary fats in their long-term health considerations. Many consumers seek to limit the amount of fats in their diets to minimize their risk of developing coronary heart disease. In particular, one category of fats appears to be linked to several contributory mortality risks: Trans Fats. The chemical structure of this class of compounds consists of long hydrocarbon chains, with one or more trans-configured alkene groups within the chain. Naturally occurring animal fats generally consist of fully hydrogenated or saturated fatty acids, lacking alkene $C = C$ bonds.

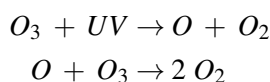
In the 1960s, health concerns about saturated fats led to the popularity of unsaturated or partially hydrogenated fatty acids. Also the lower cost of these mainly vegetable oils increased their adoption. These naturally occurring unsaturated fats are usually liquids, due to their predominantly cis-alkene configuration, which produces a “bent” structure that does not pack well. Consumer demand for solid fats, for example, “spreadable” margarines, lead to the application of chemical hydrogenation. Adding hydrogen atoms across the cis $C = C$ bonds can yield the completely saturated fat, or it can also lead to partial hydrogenation. In this case, the $C = C$ bonds do not become fully saturated, but instead the alkene bonds twist to the trans configuration. Unlike the cis-alkene fats, with their bent shape, the trans fats are more linear in shape and like their fully saturated congeners, they pack more effectively and can be produced as solids. Trans fats are also less likely to be attacked by atmospheric radicals and are therefore less vulnerable to rancidity. Their shelf life increases dramatically.

The link between trans fat consumption and heart disease is strongly supported by many medical studies. Other health effects linked to the use of trans fats include liver dysfunctions; due to the synthetic nature of trans fats, they may not metabolize in the same way as other fats.

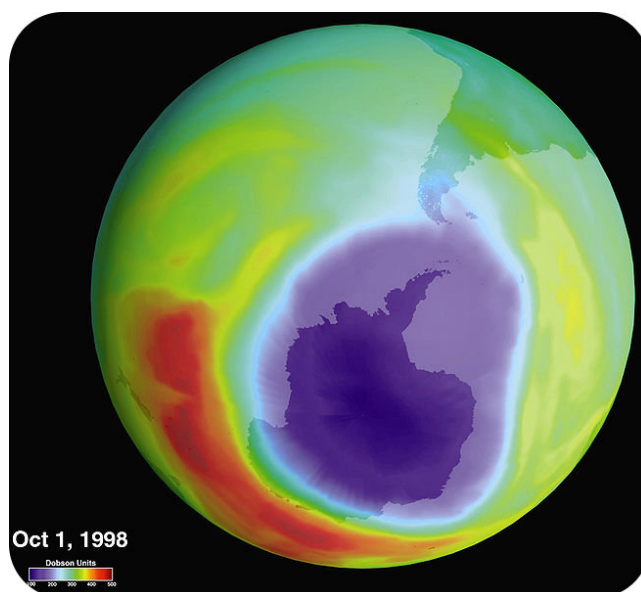
Recently, several municipalities, such as New York City, Philadelphia and San Francisco, have limited or banned outright the use of trans fats in food preparation. The Food and Drug Administration (FDA) now require food manufacturers to list the presence of trans fats on food labels.

Ozone's Role in the Atmosphere

We live on a planet uniquely situated in what astronomers refer to as our solar system's habitable zone. This region can be described as one in which the proper temperature range, elemental composition, and physical conditions have allowed life forms to exist and flourish for billions of years. An appropriate climate range, and the right array of elements, along with sufficient mass for a gravitational, pull enabled Earth to develop a protective atmosphere. The presence of oxygen gas, O_2 , causes incoming space debris to burn up usually before reaching the surface. In a similar fashion, an allotrope of oxygen, called ozone, O_3 , filters out much of the ultraviolet, B radiation (wavelengths between 280 and 320 nm) reaching Earth from the solar system. This type of radiation is a cause for concern in that it is linked with DNA mutations, particularly those associated with skin melanoma and carcinoma.



Nobel Laureate, F. Sherwood Rowland, and his research team at the University of California Irvine, in the 1970s discovered that the amount of ozone found in the stratosphere was diminishing. They found that this depletion could be associated with the cumulative amount of chlorofluorocarbon gases in the atmosphere. These CFC's or as they are otherwise known, Freons, were non-degradable remains of consumer products such as aerosol propellants and refrigerants. Their studies indicated that in the upper atmosphere, ozone was being split in the presence of ultraviolet radiation into diatomic oxygen and oxygen radical atoms, which would, in turn react with the chlorofluorocarbons. The resulting chloride monoxide, ClO , radicals were identified in the upper stratosphere over Antarctica. This location is significant because winter in the Southern Hemisphere (September & October) under the Antarctic vortex, results in the coldest winter temperatures on the planet. Under these conditions, any moisture in that locale exists only in the ice phase. The researchers found that the chlorofluorocarbon/ozone reaction was catalyzed by the presence of certain types of ice crystals in the lower atmosphere. Atmospheric measurements confirmed the accelerated depletion of ozone in the late months of the Antarctic winter, and the increased production of chloride monoxide radicals.



The purple area over Antarctica is low in ozone. The red area is higher in ozone.

Increased ozone depletion has been linked, as mentioned earlier, with an increased risk of DNA mutations. This association has been supported by increasing levels of skin cancer in humans, animals, and even plants, in latitudes where significant ozone depletion is noted. In addition, other conditions, such as diminished immune response and a higher risk of earlier cataract development, appear to be linked with ozone depletion.

Chlorofluorocarbon use has been largely phased out in developed countries, but their relative inertness leads to the situation that their influence will continue to be noted in the stratosphere for the indefinite future. Many manufacturers now substitute HCFC's, (where one or more chlorine atoms have been replaced by hydrogen atoms). These molecules retain many of the desirable properties of chlorofluorocarbons without their contributions to atmospheric destruction.

25.8 Assessment for Chapter 25

- The chapter quiz and answer key are available upon request. Please send an email to teachers-requests@ck12.org to request the material.