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CK-12 Chemistry - Second Edition Labs & Demos

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Richard Parsons

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CHAPTER 1

Introduction to Chemistry JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

1.1 Laboratory Activities

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 cylinder in shape[1] and has a diameter[2] of about 2 *cm*. The length[3] of the candle was initially about 18 *cm* and changed slowly during observation[4], decreasing about 4 *mm* in 20 *minutes*.

The candle is made of a translucent,[5] white[6] solid[7] which has a slight odor[8] 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 axis[11] and protrudes[12] about 10 *mm* above the top of the candle. The wick is made of three strands[13] of string braided[14] 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 itself[16] at the wick.

The burning candle makes no sound.[17] While burning, the body of the candle remains cool to the touch[18] except near the top.[19] Within about 5 *mm* from the top of the candle, it is warm[20] but not hot, and sufficiently soft to mold[21] easily.

The flame flickers[22] in response to air currents and tends to become smoky[23] while flickering. In the absence of air currents, the flame is in the form shown in the picture[24] although it retains some movement[25] at all times.

The flame begins[26] 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* wide[28] and extending about 8 *mm* above[29] the top of the wick, the flame is dark.[30] This dark region is roughly conical[31] in shape.

Around this dark zone and extending about 5 *mm* above the dark zone is a region which emits yellow light,[32] bright[33] but not blinding. The flame has rather sharply defined sides[34] but a ragged[35] top.

The wick is white[36] 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 curls[39] over about 4 *mm* from its end. As the candle becomes shorter, the wick shortens[40] too, so as to extend roughly a constant distance above the top of the candle.

Heat is emitted[41] 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 wet[42] with a colorless[43] liquid and becomes bowl-shaped.[44] Sometimes, the liquid in the bowl drains[45] down the side of the candle, cools, gradually solidifying[46] and attaching itself to the candle.[47]

Under quiet conditions, a stable[48] pool of clear liquid remains in the bowl-shaped top of the candle. The liquid rises[49] slightly around the wick, wetting[50] the base of the wick as high as the base[51] 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 1.1: **Answers for DAZOO**

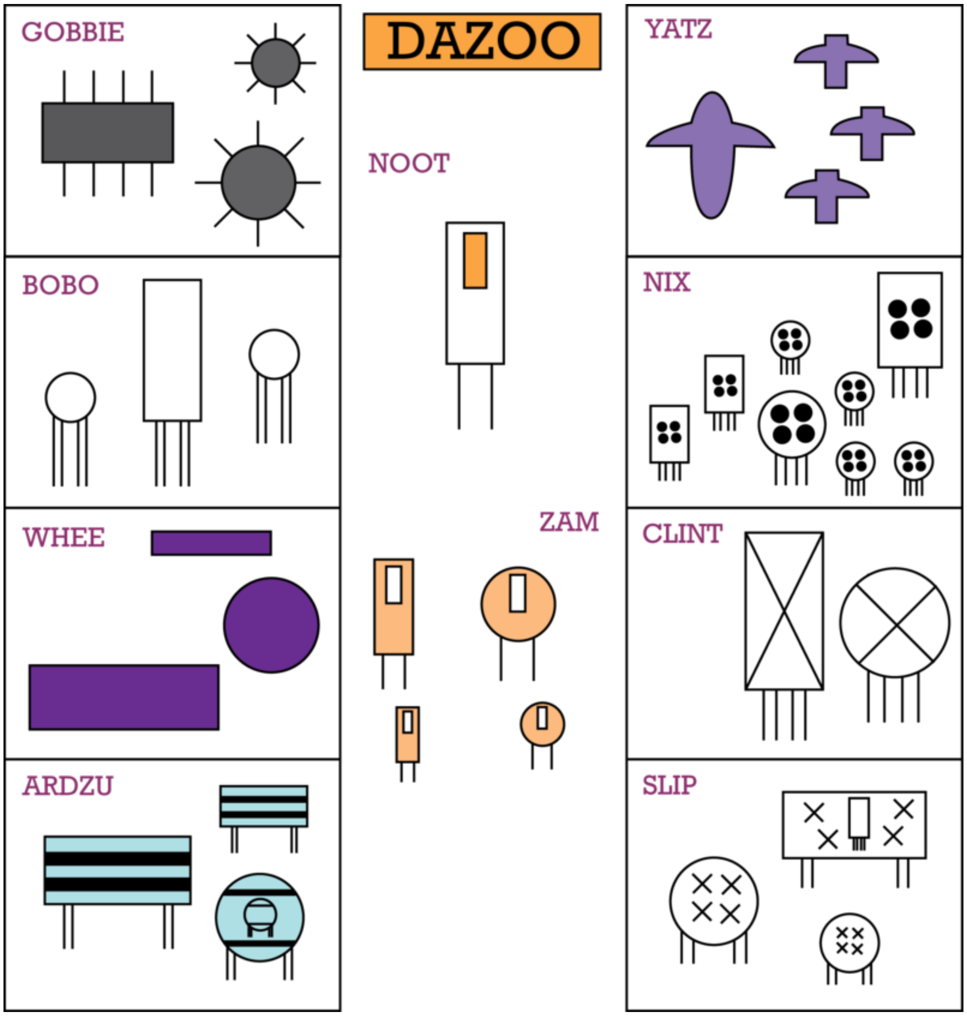
Question Number	Answer	Reason
1.	ZAM	It is a family and they are outside the cages.
2.	NOOT	It is a single male outside the cages.
3.	NIX	It is the largest number of individuals in a cage.

TABLE 1.1: (continued)

Question Number	**Answer**	**Reason**
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.
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?

Chemical History Lab #8211; Early Contributions to Metallurgy

While the rest of the world was searching for the philosopher#8217;s stone or the elixir of life, chemists from the Indian subcontinent were learning a lot about metals. India developed many techniques for the extraction and purification of metals from ores that could be mined out of the ground. A method for production of Wootz steel was developed in Sri Lanka as well. Another important contribution to metallurgy developed in India was the method of identifying different metals based on their flame tests. In fact, this method is still sometimes used today.

WARNING!! This experiment should be performed as a demonstration at the front of the class. It is NOT suggested that students be allowed to do this experiment themselves. Be careful to avoid contact with the 6 M hydrochloric acid, as it is corrosive and can cause severe burns. In addition, remember that the wire loop will get extremely hot. Do not touch the end of the wire loop until you have given it plenty of time to cool. Finally, ALWAYS wear safety goggles.

Materials:

A clean wire loop (Platinum or nickel-chromium works best.)

6 M Hydrochloric acid

A Bunsen burner

Potassium nitrate, KNO_3

Strontium nitrate, $Sr(NO_3)_2$

Barium nitrate, $Ba(NO_3)_2$

Sodium nitrate, $NaNO_3$

Copper(II) nitrate, $Cu(NO_3)_2$

Method:

1. Dip the wire loop into the hydrochloric acid and then heat the wire in the Bunsen burner flame. This step is to ensure that the wire loop is clean. If there is no color produced when the wire is put into the flame, you know that the loop is clean. If, however, you see color in the flame, dip the wire loop into the hydrochloric acid again. Continue this process until the wire loop burns without color.
2. Once you have a clean wire loop, dip the loop into the sample of the first metal salt you want to test; then heat the loop in the Bunsen burner flame. The metal salt should burn with a color that is obvious to the class.
3. Repeat the process for the rest of the metal salts that you wish to try, being sure to clean the wire loop (step one) between each test.

Although everyone will see slightly different colors, the following is a general guideline:

K^+ – lilac/pink

Sr^{2+} – red

Ba^{2+} – pale green

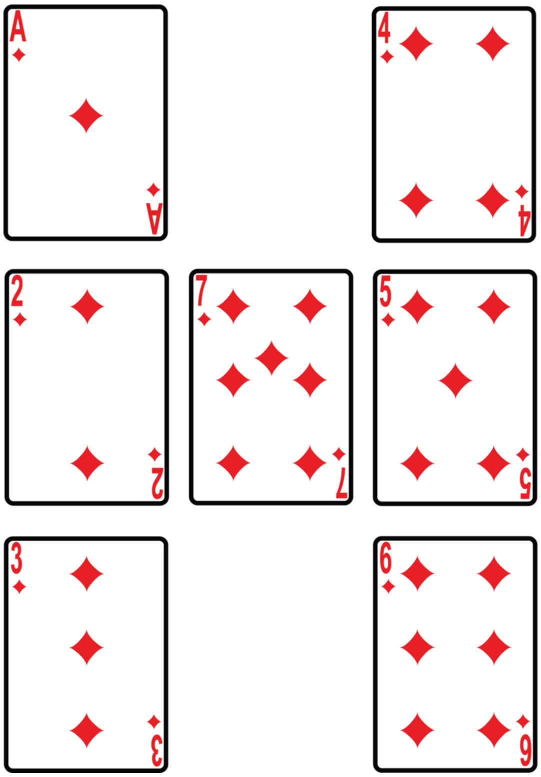
Na^+ – strong persistent yellow

Cu^{2+} – blue-green

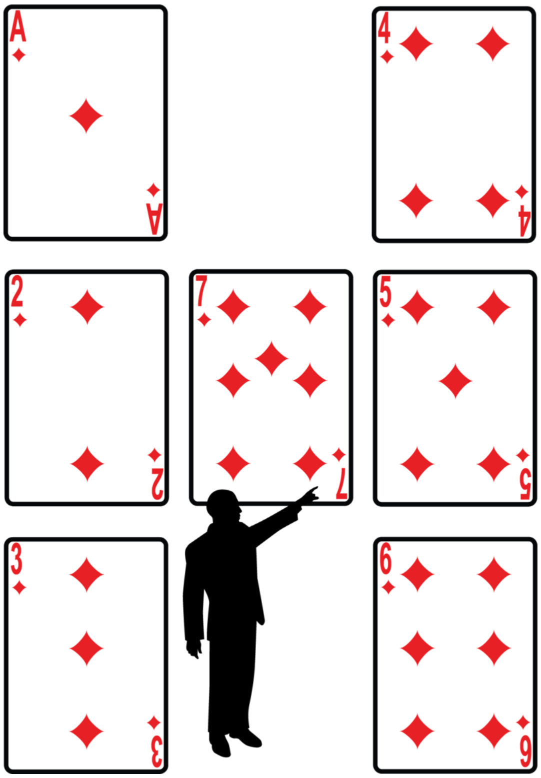
1.2 Demonstrations

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 #8220;psychic#8221;. The teacher draws the seven-of-diamonds set up on the board as shown below.



To begin the game, the #8220;psychic#8221; 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, ***#8220;Is this the card?#8221;*** The psychic responds either #8220;yes#8221; or #8220;no#8221; 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 hypotheesis 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.



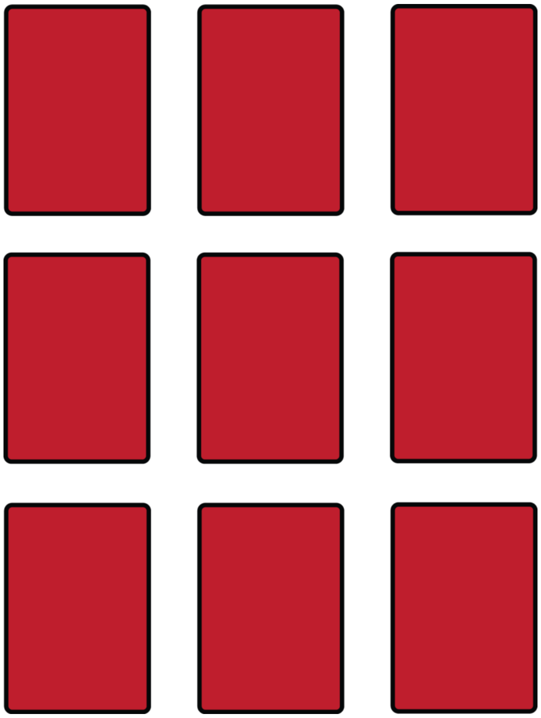
THE TRICK

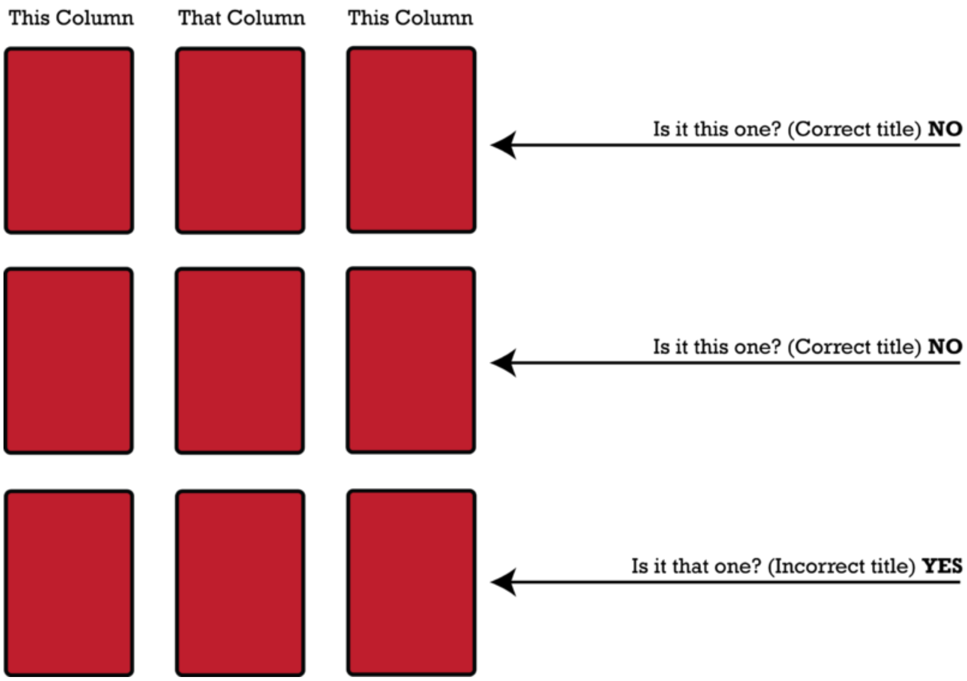
The layout of the cards on the board and the spots on the seven-of-diamonds exactly correspond. The psychic will not know the correct card until the teacher points to the seven-of-diamonds and asks, #8220;Is this the card?#8221; When the teacher points to the seven-of-diamonds, he/she points to the spot 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 #8220;6#8221; in the layout. Therefore, the six-of-diamonds is the selected card for this trial. The psychic continues to say #8220;no#8221; until the teacher points to the six-of-diamonds and then says #8220;yes.#8221; 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.

#8220;This#8221; or #8220;That#8221; 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 #8220;psychic#8221; conspire and assign the two outside columns to be called #8220;this#8221; columns and the middle column to be called a #8220;that#8221; column.

As usual, the psychic leaves the room and the students select one of the squares to be #8220;psychically#8221; identified. The psychic is called back into the room and the teacher proceeds to point at various squares and ask the psychic, #8220;Is it this one?, or #8220;Is it that one?#8221; 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 #8220;no.#8221; If the teacher uses the incorrect name of the column when inquiring, the psychic replies #8220;yes.#8221;





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 #8220;that#8221; column) and ask, #8220;Is it this one?#8221; and the psychic replies #8220;yes.#8221;

1.3 Extra Readings

Historical Comparisons

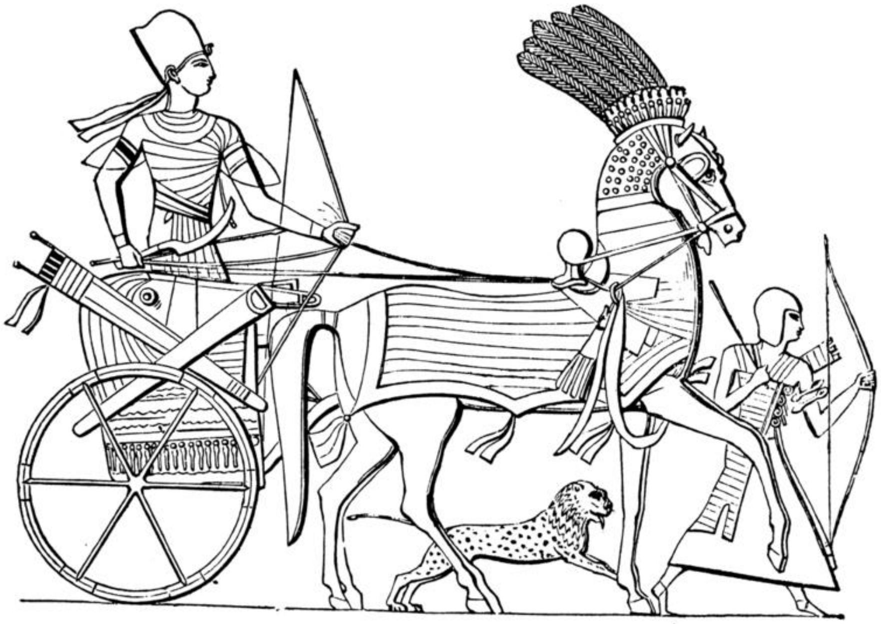
Introduction to Science

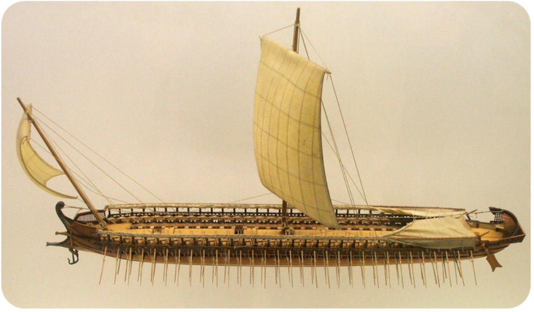
What is science? Is it a list of marvelous inventions and how they work? Or is it a list of theories about matter and energy and biological systems? Or is science a subject that you learn by carrying out activities in a laboratory? Science is all of these, but it is also something even more basic. Science 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** below). Crossing water, people could paddle a boat or have an animal walk beside the river and pull the boat (**Figure** below). These methods of transportation required muscle power, either human muscles or animal muscles.

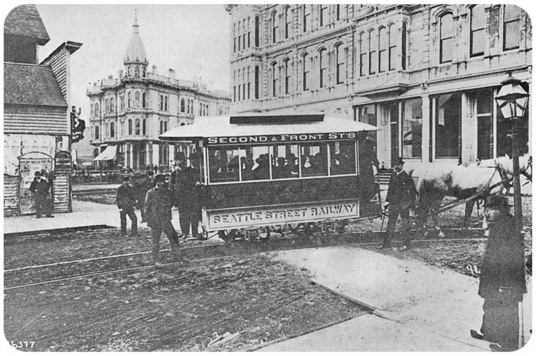




A few societies had designed rowboats or sailboats, which used muscle power or the force of the wind to move the boat. These early means of transportation were very limited in terms of speed and therefore, also limited in terms of distances traveled. The sail and rowboats were used on rivers and inland 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** below). Large cities had streetcars pulled by horses (**Figure** below). 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.



Transportation in 1995

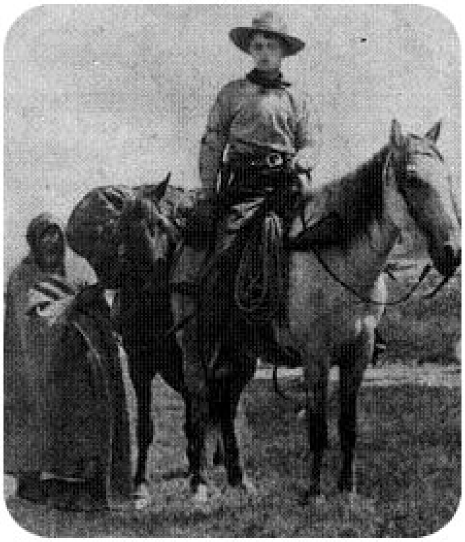
By the year 1995, steam engines, gasoline engines, automobiles, propeller-driven and jet engines, locomotives, nuclear-powered ships, and inter-planetary rocket ships were invented (**Figure** below). In all industrialized countries, almost anyone could own an automobile and travel great distances in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

Communication in 1000 B.C.

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** above). 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** above). The means of communication in 1830 was essentially the same as in 1000 B.C.

Communication in 1995

By the year 1995, the telegraph, telephone, radio, television, optical fibers, and communication satellites were invented (**Figure** above). 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.

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** above). 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 #34;bad odors,#34; 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 #34;bad odors,#34; 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

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** above).

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 (14^{th} 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 16^{th} and 17^{th} 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 Bellarmini 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 #34;scientists#34; claiming that unsupported opinions are #34;facts#34; and we still have people deciding the #34;truth#34; 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 #34;casting spells#34; and a substitute teacher in Florida was fired in 2008 for #34;wizardry#34; after performing a magic trick for his students.

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 #8220;see#8221; 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 Dowsers

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 dowsers. A double blind test requires that neither the dowser nor the judges know the position of the dowsers 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 #8220;sensed#8221; 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 dowsers 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 #8220;Mysterious#8221; Bermuda Triangle

The #8220;Bermuda Triangle#8221; is a triangular expanse of ocean between the three vertices of Bermuda, Puerto Rico, and Miami. The #8220;mystery#8221; 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 whly 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: #8220;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.#8221;

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 #8220;vanished#8221; and #8220;disappeared#8221; rather than #8220;crashed in the ocean and sank#8221; 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 #8220;suffered a loss by disintegration#8221;. 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 #8220;disintegrated#8221;. 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 #8220;fireball which circled his flagship#8221;.

Other readers of the logbook report: Columbus wrote of #8220;a great flame which fell into the sea.#8221; 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 #8220;data#8221; 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.

CHAPTER 2 Measurement in Chemistry JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

2.1 Laboratory Activities

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 2.1: **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.

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

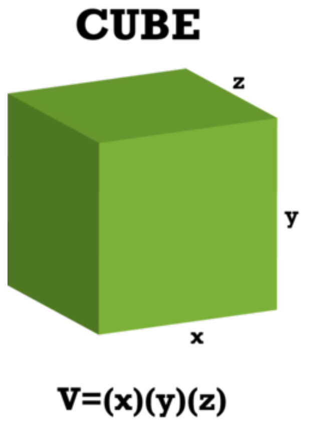
Example: Calculate the density of a piece of lead whose mass is 226 *grams* and whose volume is 20.0 *mL*. Also calculate the density of a sample of gold whose mass is 57.9 *grams* and whose volume is 3.00 *mL*.

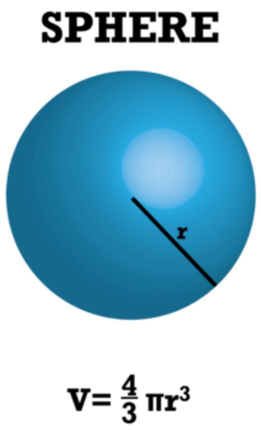
Solution:

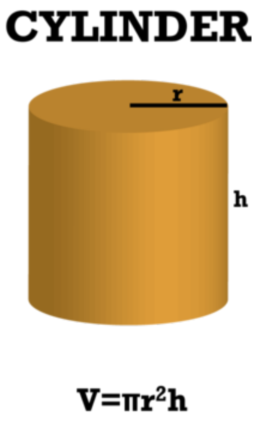
$$\text{Density of Lead} = \frac{mass}{volume} = \frac{226\ g}{20.0\ mL} = 11.3\ \text{g/mL}$$
$$\text{Density of Gold} = \frac{mass}{volume} = \frac{57.9\ g}{3.00\ mL} = 19.3\ \text{g/mL}$$

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.

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

Teacher#8217;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 $-10^\circ C$. When ice and water are both present and in equilibrium, you can be sure the temperature of the mixture is $0^\circ C$.

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 $80^\circ C$ and $82^\circ C$.

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

Near $101^\circ C$

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 $100^\circ C$. Since boiling water is less than $100^\circ C$, cooking in boiling water will take longer.

TABLE 2.2: **Vapor Pressure of Water at Various Temperatures**

Temperature in $^{\circ}C$	Vapor Pressure in mm of Hg	Temperature in $^{\circ}C$	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 $-20^{\circ}C$ to $120^{\circ}C$. 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 $0^\circ C$ 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 is 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 is 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.

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

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

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

TABLE 2.3: Actual Boiling Point of Water versus Various Room Pressures

Room Pressure (mm of Hg)	Boiling Point of Water ($^\circ C$)
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 2.3: (continued)

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

2.2 Demonstrations

Density of Diet Soda vs. Regular Soda

Brief description of demonstration

A can of diet soda and regular soda are place 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.

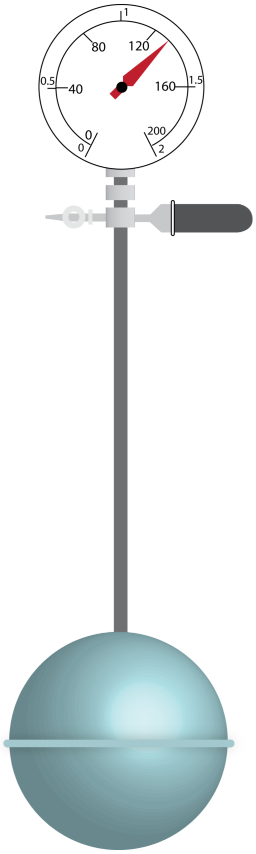
Discussion

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.



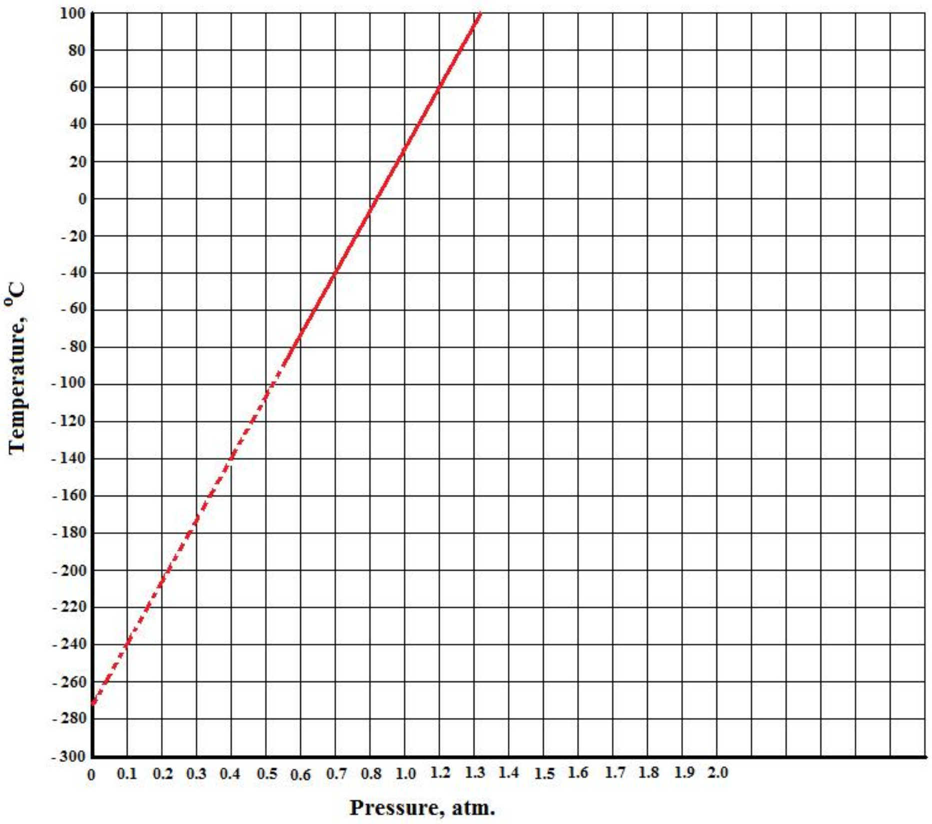
Apparatus and Materials

- 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 #8211; dry ice can be stored in a Styrofoam cooler but do not put the lid on tightly)
- Ethanol #8211; 600 mL
- If you have a mercury thermometer that covers the range $-100°C$ to $+100°C$, 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.
7. Plot a graph of the temperatures and pressures recorded. Make sure that the temperature axis on your graph extends below $-280°C$.

8. 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 #8211; 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.

2.3 Extra Readings

CHAPTER 3

Matter and Energy JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

3.1 Laboratory Activities

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 $AgNO_3$ solution $(1.7\ g/100\ mL)$
- 1 mL of approximately 0.10 M $FeCl_3$ 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 HNO_3 (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 $(13 \times 10\ mm)$
- laboratory burner
- 3 test tubes $(18 \times 150\ mm)$
- ring stand and ring
- clay triangle
- watch glass
- wood splint
- matches
- crucible tongs
- solid $NaCl$
- solid copper turnings
- solid $Cu(NO_3)_2$
- 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 $Cu(NO_3)_2$ 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, $Cu(NO_3)_2$, 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 3.1: 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		

3.2 Demonstrations

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.

3.3 Extra Readings

CHAPTER 4 The Atomic Theory JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

4.1 Laboratory Activities

Early Development of a Theory Lab

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

In today#8217;s lesson, we have discussed two different laws that are often confused #8211; *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#8217;s Atomic Theory. Materials:

Scissors

A copy (preferably in color) of the #8220;Paper Atoms#8221; 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 #8220;paper atoms#8221; will be used in both mini-labs. Notice that the $N\ atoms$ have #8220;14 u#8221; on them. This means that a nitrogen atom has a mass of 14 #8220;atomic mass units.#8221; (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 #8220;16 u#8221; on them. This means that an oxygen atom has a mass of 16 #8220;atomic mass units.#8221;

According to Dalton#8217;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 #8220;matter#8221;.*
2. 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***.
3. 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, #8220;mass#8221; and color. Similarly, all of the oxygen, O atoms are the same size, #8220;mass#8221; and color.*
4. The atoms of different elements differ in mass and other properties. *Notice, however, that the nitrogen, N, atoms are not the same size, #8220;mass#8221; or color as the oxygen, O, atoms.*
5. Atoms of one element can combine with atoms of another element to form #8220;compounds#8221; #8211; new, complex particles. In a given compound, however, the different types of atoms are always present in the same relative numbers. *We#8217;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**.*

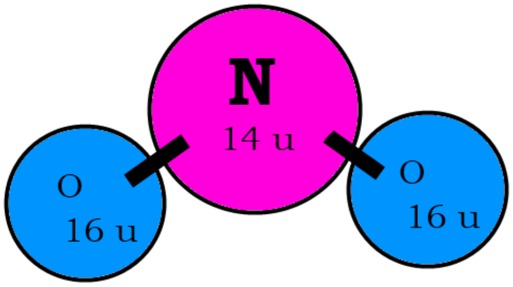
Method:

We#8217;ll consider the Law of Definite Proportions by #8220;making#8221; the substance #8220;NO_2#8221;, or #8220;nitrogen dioxide#8221;. Draw the following table in your notebook.

TABLE 4.1: 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			

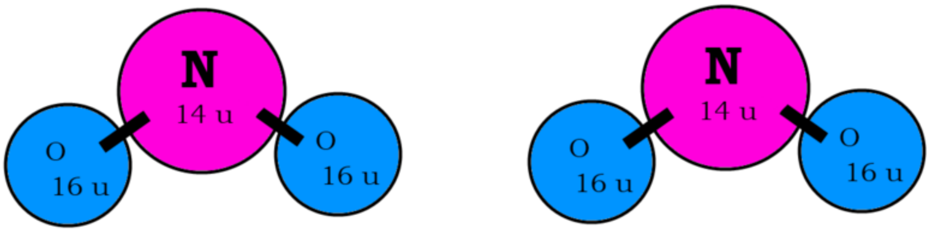
1. Begin by constructing a single NO_2 #8220;molecule#8221; (#8220;molecule#8221; is a word given to a complex particle formed by joining several atoms together) by taping an oxygen atom to either side of one of your nitrogen atoms.



2. Add up the total mass of N *atoms* in your one NO_2 #8220;molecule.#8221; This is easy, of course #8211; 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 #8220;Total mass of nitrogen (u).#8221;

3. Add up the total mass of O *atoms* in your one NO_2 #8220;molecule#8221;. This is slightly more challenging #8211; 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 #8220;Total mass of oxygen (u)#8221;.

4. 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 #8220;molecules.#8221; Since you now have 2 nitrogen atoms, you have a total mass of $14\ u + 14\ u = 28\ u$. Fill this number into the second row of your table under the column heading #8220;Total mass of nitrogen (u).#8221;

6. Add up the total mass of O *atoms* in *both* NO_2 #8220;molecules.#8221; Because you now have 4 oxygen atoms, you have a total mass of $16\ u + 16\ u + 16\ u + 16\ u = 64\ u$. Fill this number into the second row of your table under the column heading #8220;Total mass of oxygen (u).#8221;

7. Continue in this manner, filling out the table as you go, until you#8217;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#8217;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 $4 \times 16\ u = 64\ u$.)

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#8217;s Atomic Theory:

1. Matter is made of tiny particles called atoms. *We made nitrogen dioxide by assuming that it was composed of N atomsand 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 #8220;mass#8221; of* $14\ u$. *Similarly, all of the oxygen, O, atoms were the same size and color, and they all had a #8220;mass#8221; 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 #8220;mass#8221; as the oxygen, O, atoms.*
5. Atoms of one element can combine with atoms of another element to form #8220;compounds#8221; #8211; 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

#8220;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*.#8221;

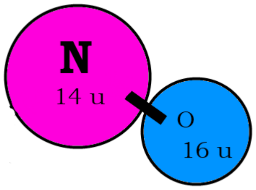
Method:

You#8217;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#8217;ll consider the Law of Multiple Proportions by #8220;making#8221; the substances #8220;NO#8221;, or #8220;nitrogen monoxide#8221;, NO_2, or #8220;nitrogen dioxide#8221;, NO_3, or #8220;nitrogen trioxide#8221; and NO_4 or #8220;nitrogen tetraoxide#8221;. In each case, we#8217;ll use five nitrogen atoms, meaning we#8217;ll always have the same mass of nitrogen. Draw the following table in your notebook.

TABLE 4.2: **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.

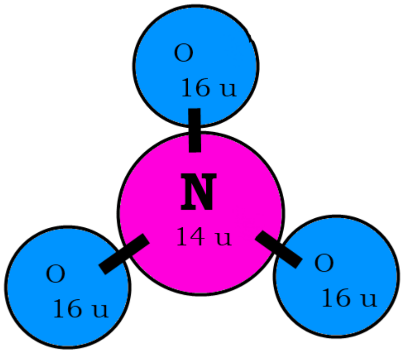


2. Add up the total mass of $O\ atoms$ in *all five* NO molecules. This gives a total of $16\ u + 16\ u + 16\ u + 16\ u + 16\ u = 80\ u$. Fill this number into the first row of your table under the column heading #8220;Total mass of oxygen (u)#8221;.

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.

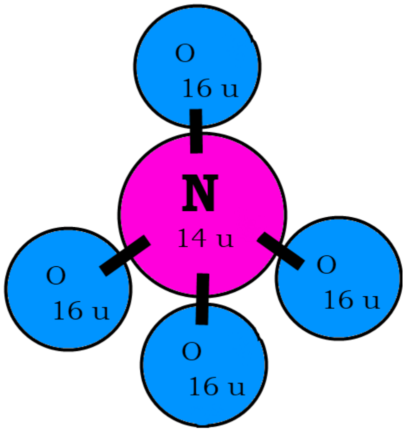
4. Add up the total mass of $O\ atoms$ in *all five* NO_2 molecules. Fill this number into the second row of your table under the column heading #8220;Total mass of oxygen (u)#8221;. (Note: If you#8217;d like to save time at this point, you do not actually need to construct all 5 NO_2 molecules, since that#8217;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 #8220;Total mass of oxygen (u)#8221;.

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* NO_4 molecules. Fill this number into the fourth row of your table under the column heading #8220;Total mass of oxygen (u)#8221;.

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 $14\ u + 14\ u + 14\ u + 14\ u + 14\ u = 70\ u$. 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 NO_2 to the mass of oxygen in NO for 70 u of N:

= (total mass of oxygen in NO_2) ÷ (total mass of oxygen in NO)

= (________________________) ÷ (________________________)

= ________________

b. The ratio of the mass of oxygen in NO_3 to the mass of oxygen in NO for 70 u of N:

= (total mass of oxygen in NO_3) ÷ (total mass of oxygen in NO)

= (________________________) ÷ (________________________)

= ________________

c. The ratio of the mass of oxygen in NO_4 to the mass of oxygen in NO for 70 u of N:

= (total mass of oxygen in NO_4) ÷ (total mass of oxygen in NO)

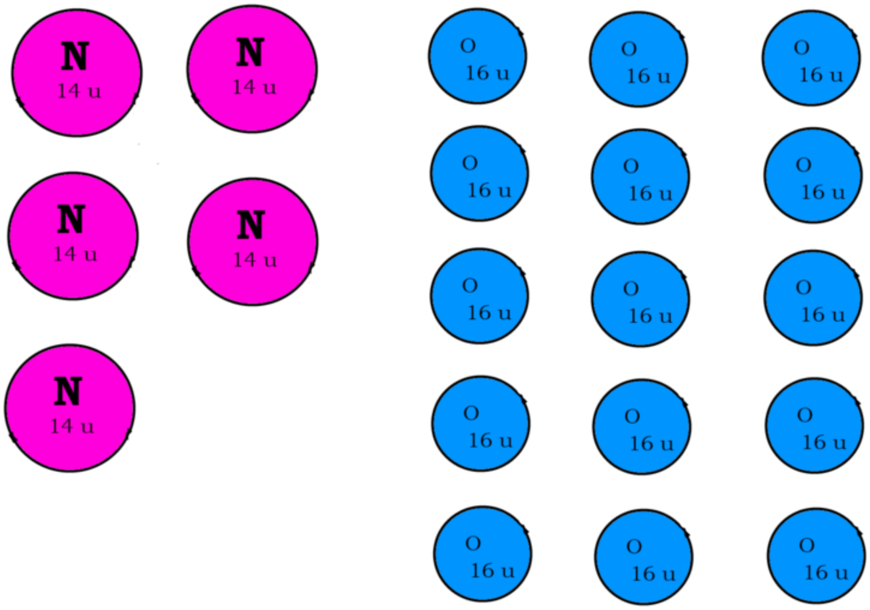
= (________________________) ÷ (________________________)

= ________________

What do you notice about the ratios? Are they small whole numbers? Does the Law of Multiple Proportions hold for the formation of NO, NO_2, NO_3 and NO_4? Notice, again, how the Law of Multiple Proportions occurs naturally provided we assume that NO, NO_2, NO_3 and NO_4 are formed according to rules in Dalton#8217;s Atomic Theory.

To the Teacher: We could do the same calculations as above, but compare the mass of oxygen in NO_2 to the mass of oxygen in NO_3, or the mass of oxygen in NO_3 to the mass of oxygen in NO_4. This wouldn#8217;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



Further Understanding of the Atom #8211; Rutherford#8217;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#8217;s Gold Foil Experiment. In this lab, we#8217;ll try to recreate the principle that Rutherford used in his Gold Foil experiment, but we#8217;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

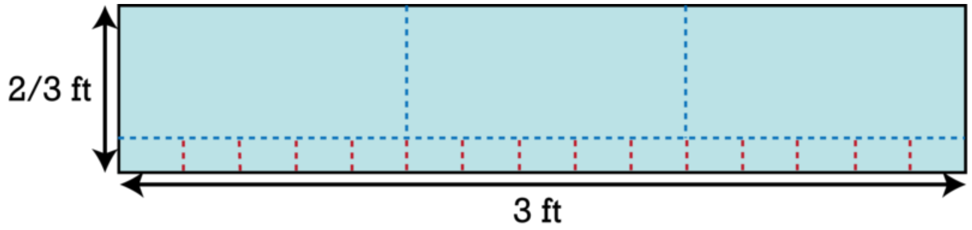
glue gun

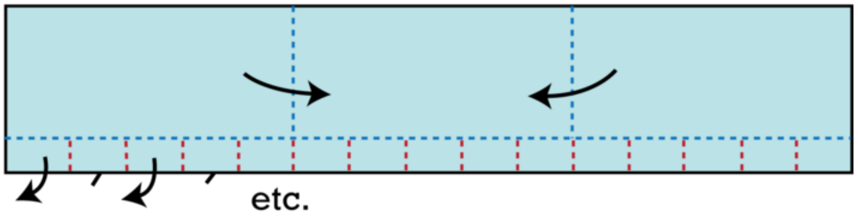
spice jar (approximately 6 *inches* tall and 2 *inches* wide) filled with sand or some other heavy substance

Method:

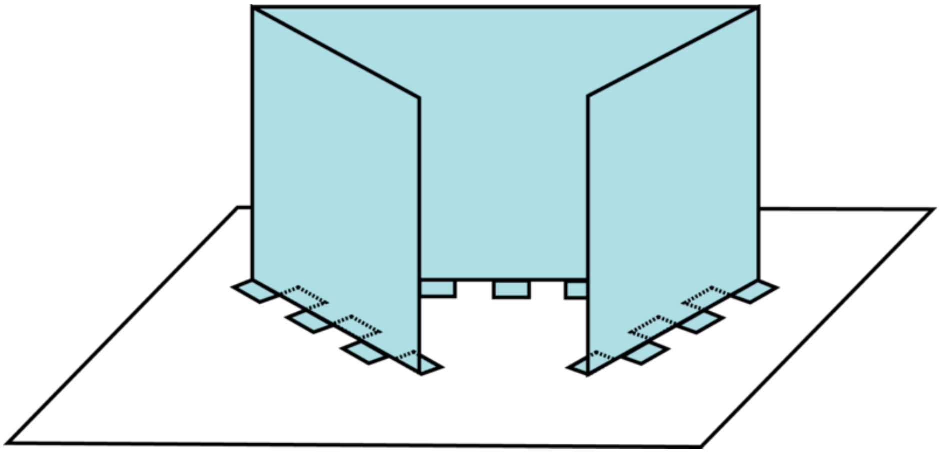
Pre-lab preparation

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#8217;t, there won#8217;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#8217;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 #8211; 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#8217;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.

Energy Lab #8211; 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#8217;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.

4.2 Demonstrations

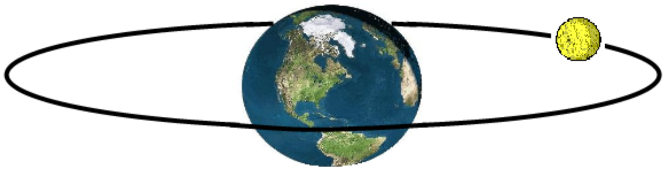
4.3 Extra Readings

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 #8211; 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 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 #8211; living and non-living #8211; would be at a standstill; nothing would move, nothing would live. Energy is considered to be the #8220;mover of matter#8221;. 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 #8211; or #8211; 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 #8220;ability to make matter move#8221;. 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 #8211; 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 #8211; 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.

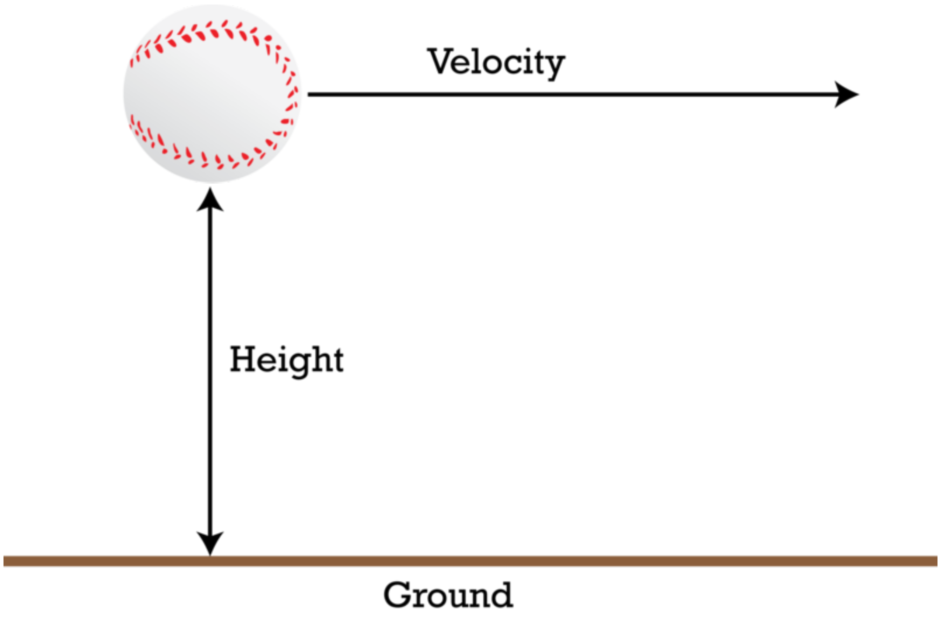
This baseball flying through the air has both kinetic and potential energy. The kinetic energy is due to its motion and the potential energy is due to the balls#8217; height above the ground.

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.

$$PE = mgh$$



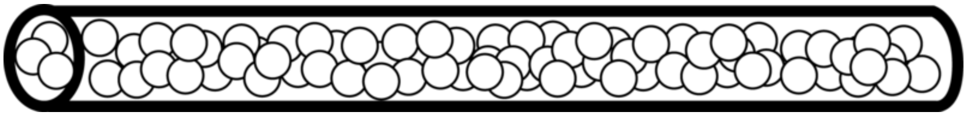
Potential energy is always present when two objects are attracted or repelled and are held in position. The most obvious case is an object that is held above the earth. The object is attracted to the earth by gravity but is kept from falling (gravitational potential energy). This same type of energy 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

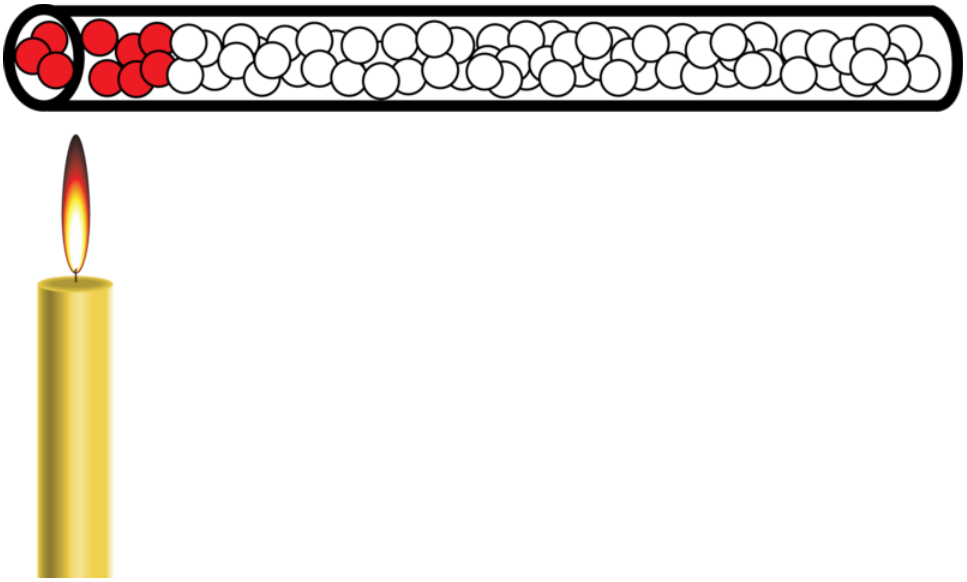
Scientists use three words to indicate the different methods by which energy moves from place to place. These three words are conduction, convection, and radiation.

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 #8220;hotter#8221; molecules to the #8220;colder#8221; molecules. Consider an object such as an iron bar #8211; 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.

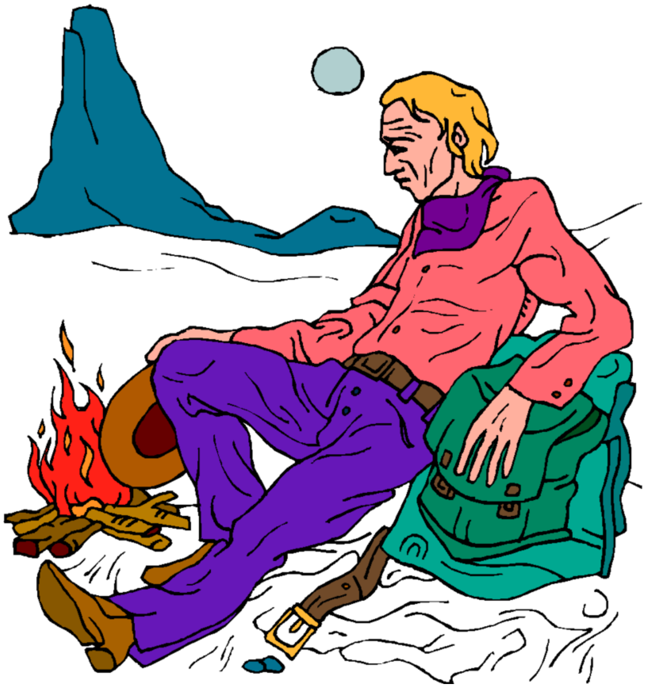
Convection

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 #8211; 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.

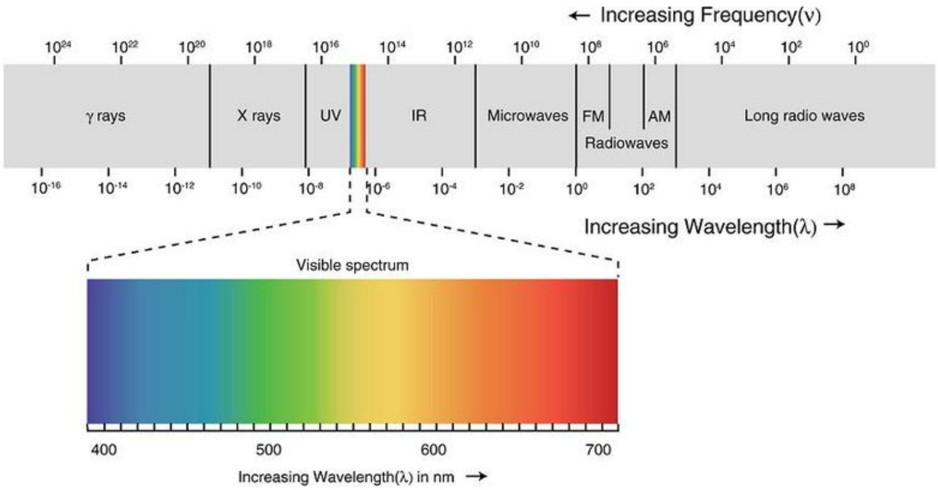


This is the way that light from the sun travels through the vacuum of outer space and arrives at the earth. This type of energy is called electromagnetic radiation. There are various levels of energy for electromagnetic radiation. One of the lower energy forms of EMR are radio waves. As the energy of 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 #8211; 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 #8220;eyes#8221; 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.

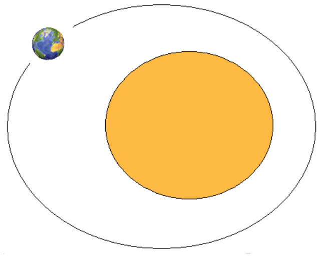
The size of the force of gravity is dependent on the masses of the two attracting objects and also on the distance between the centers of the two objects. The amount of matter in an object is called its mass and whenever the mass of one or both of the objects is increases, the force of gravity between 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 $10,000,000,000,000,000,000,000,000$ *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).



Timeline for the Discovery of the Elements

TABLE 4.3: 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	#8211; 1745
Pt, Co		
Simple Metallurgy	1745	#8211; 1755
Ni		

TABLE 4.3: (continued)

Method of Discovery	Year of Discovery	Element Symbols
	1755	#8211; 1765
Electrochemistry F, Mn, Cl, O, N, H	1765	#8211; 1775
Electrochemistry Te, W, Mo	1775	#8211; 1785
Electrochemistry Y, Ti, Sr, U	1785	#8211; 1795
Electrochemistry Ir, Os, Rh, Pd, Ce, Ta, Nb, Be, Fr, V	1795	#8211; 1805
Electrochemistry I, B, Mg, Ca, Ba, K, Na	1805	#8211; 1815
Electrochemistry Zr, Si, Se, Cd, Li	1815	#8211; 1825
Electrochemistry Th, Br, Al	1825	#8211; 1835
Electrochemistry Ru, Er, Tb, La	1835	#8211; 1845
	1845	#8211; 1855
Electrochemistry In, Tl, Rb, Cs	1855	#8211; 1865
	1865	#8211; 1875
Separation Techniques Tm, Ho, Sc, Sm, Yb, Ga	1875	#8211; 1885
Separation Techniques Ar, Ge, Dy, Gd, Nd, Pr	1885	#8211; 1895
Separation Techniques Rn, Ac, Ra, Po, Xe, Ne, Kr, Eu, He	1895	#8211; 1905
Separation Techniques Lu	1905	#8211; 1915
Separation Techniques Pa, Hf	1915	#8211; 1925
Separation Techniques Re	1925	#8211; 1935
Nuclear Synthesis Cm, Pu, Np, At, Fr, Tc	1935	#8211; 1945
Nuclear Synthesis Cf, Bk, Pm, Am, Es, Fm, Md	1945	#8211; 1955
Nuclear Synthesis No, Lr, Rf, Db, Sg, Bh, Hs, Mt	1955	#8211; 1965
Nuclear Synthesis Ds, Rg, Cp, 114, 116, 118	1965	#8211; Present

CHAPTER 5 The Bohr Model of the Atom JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

5.1 Laboratory Activities

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 #8216;white#8217;. 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 #8216;stirred#8217; 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 #8216;white#8217; really isn#8217;t a #8216;color#8217;, but rather #8216;all colors#8217; 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:

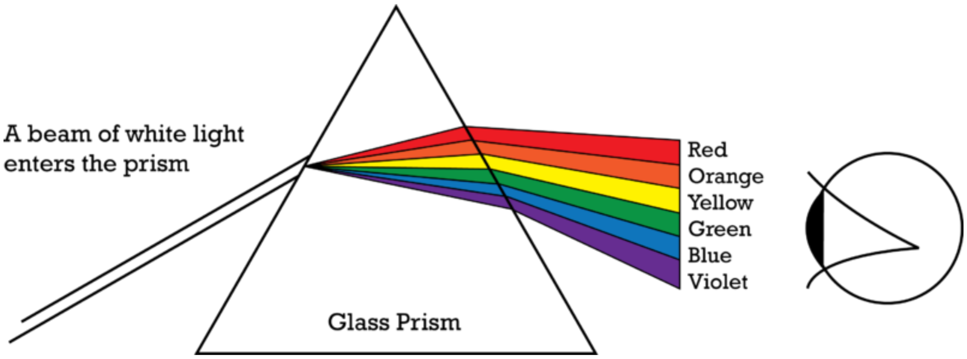
Start with the color wheel at rest, noticing how it contains all of the colors in the rainbow. Begin spinning the color wheel, slowly at first, and then, with increasing speed. Notice how, once it is spinning fast enough, the colored wheel turns white. Again, this should convince you that #8216;white#8217; really isn#8217;t a #8216;color,#8217; but rather, all colors at once.

5.2 Demonstrations

Atomic Spectra Viewed Through a Diffraction Grating Demo

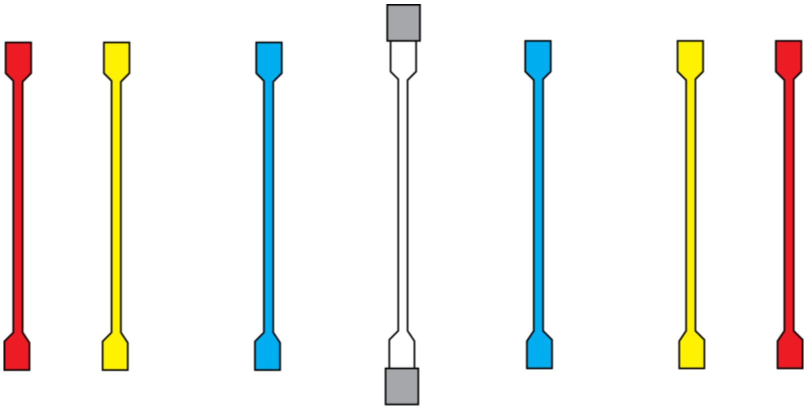
Background:

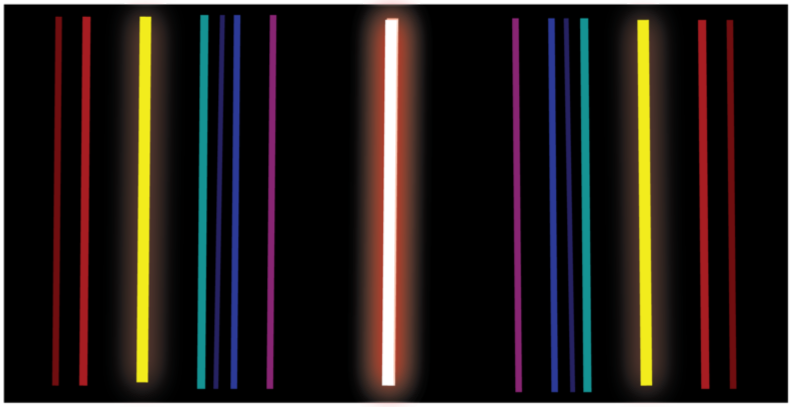
The speed of light in a vacuum is 3.0×10^8 $meters/second$. All frequencies of light in a vacuum travel at the same speed. When light passes through media other than a vacuum, such as glass or water, however, all frequencies do not travel at the same speed. In general, light travels at a speed of about 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#8217;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 #8211; 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 exited 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.

5.3 Extra Reading

CHAPTER 6 The Quantum Mechanical Model of the Atom JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

6.1 Laboratory Activities

6.2 Demonstrations

6.3 Extra Readings

CHAPTER 7

The Electron Configuration of Atoms JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

7.1 Laboratory Activities

The Electron Spin Quantum Number Lab #8211; Diamagnetic Levitation

In today#8217;s lesson, 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.

Mini-lab: Standing Waves in a Jump Rope

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!

7.2 Demonstrations

7.3 Extra Readings

CHAPTER 8 Electron Configuration and the Periodic Table JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

8.1 Laboratory Activities

8.2 Demonstrations

8.3 Extra Readings

CHAPTER 9

Chemical Periodicity JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

9.1 Laboratory Activities

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 #8211; 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 9.1: Data

	With Magnet	Without Magnet
Mass of the MnO_2/capsule/cup assembly		
Mass of the Mn_2O_3/capsule/cup assembly		
Mass of the $KMnO_4$/capsule/cup assembly		

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 $KMnO_4$?

9.2 Demonstrations

9.3 Extra Readings

CHAPTER 10 Ionic Bonds and Formulas JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

10.1 Laboratory Activities

10.2 Demonstrations

10.3 Extra Readings

CHAPTER 11

Covalent Bonds and Formuas JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

11.1 Laboratory Activities

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 11.1: **Answers to Pre-Lab Questions**

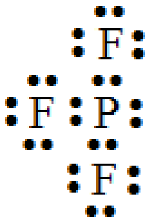
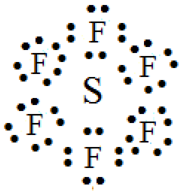
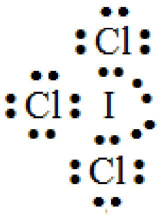
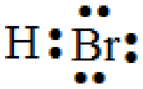
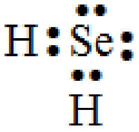
Formula	Electron-Dot Formula
PF_3	
SF_6	
ICl_3	

TABLE 11.1: (continued)

Formula	**Electron-Dot Formula**
HBr	
H_2Se	

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 place 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.

PF_3

SF_6

ICl_3

HBr

H_2Se

Purpose:

Models built according to the predictions of the VSEPR theory illustrate the reular 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 11.2: **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.**

11.2 Demonstrations

11.3 Extra Readings

CHAPTER 12 The Mole Concept JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

12.1 Laboratory Activities

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 #8211; 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:

$$Mg_{(s)} + N_{2(g)} + O_{2(g)} \rightarrow MgO_{(s)} + Mg_3N_{2(s)}$$
$$MgO_{(s)} + Mg_3N_{2(s)} + H_2O_{(L)} \rightarrow MgO_{(s)} + Mg(OH)_{2(s)} + NH_{3(g)}$$
$$MgO_{(s)} + Mg(OH)_{2(s)} \rightarrow MgO_{(s)} + H_2O_{(g)}$$

Pre-Lab Questions

1. If the mass percent of each element in a compound is known, what steps are taken to determine the compound#8217;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

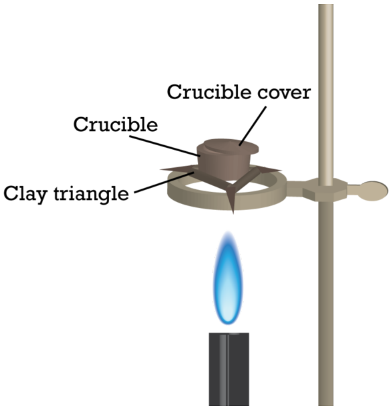
1. Heat the empty crucible and lid for about 3 *minutes* to remove water, oils, or other contaminants and to make sure there are no cracks. The bottom of the crucible should glow red-hot for about 20 *seconds*. Remove the flame and cool the crucible with lid.

2. Record the mass of crucible and lid once it has cooled. Handle the crucible with tongs.

3. 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.

4. Record the mass of the magnesium ribbon, lid and crucible.

5. 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.



6. 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.

7. Heat until all the magnesium turns into gray-white powder (around 10 *minutes*).

8. Stop heating and allow the crucible, lid and contents to cool.

9. 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.

10. 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*.

11. Allow the crucible to cool and then record the mass of the crucible, lid and contents.

12. 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

1. 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

2. Determine the number of moles of magnesium used. Remember: $\frac{\text{mass}}{\text{atomic weight}}$ = number of moles. The atomic weight of magnesium is 24.3 g/mol.

Number of moles of magnesium = _______ mole

3. 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 #8211; 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 #8211; 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 12.1: 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 $CuSO_4\left(\frac{7}{6}\right)$	

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?

12.2 Demonstrations

12.3 Extra Readings

CHAPTER 13 Chemical Reactions JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

13.1 Laboratory Activities

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 13.1: **Data Table**

	1	2	3	4	5	6	7	8
A								

TABLE 13.1: (continued)

	1	2	3	4	5	6	7	8
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 though F8, add 5 drops of 0.1 $mol/L\ AgNO_3$.

Part 8: In wells G1 though G8, add 5 drops of 0.1 $mol/L\ HNO_3$.

Part 9: In wells A1 through to G1, add 5 drops of 0.1 $mol/L\ Na_2CO_3$.

Part 10: In wells A2 though G2, add 5 drops of 0.1 $mol/L\ Na_2SO_4$.

Part 11: In wells A3 though G3, add 5 drops of 0.1 $mol/L\ NaCl$.

Part 12: In wells A4 though G4, add 5 drops of 0.1 $mol/L\ NaI$.

Part 13: In wells A5 though G5, add 5 drops of 0.1 $mol/L\ Na_2CrO_4$.

Part 14: In wells A6 though G6, add 5 drops of 0.1 $mol/L\ Na_2Cr_2O_7$.

Part 15: In wells A7 though G7, add 5 drops of 0.1 $mol/L\ NaOH$.

Part 16: In wells A8 though 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 pro-beware.

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

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.

$$SO_{3(g)} + H_2O_{(L)} \rightarrow H_2SO_4(aq)$$

(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.

$$CaCO_{3(s)} + heat \rightarrow CaO_{(s)} + CO_{2(g)}$$

(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.

$$Zn_{(s)} + 2\ HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

(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.

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{2(aq)}$$

(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.

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)} + heat + light$$

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 13.2: Observations

Reaction	Observations	Reaction Type
Fe and $CuSO_4$		
$Pb(NO_3)_2$ and KI		
$CuSO_4 \cdot 5H_2O$ and heat		
Mg and HCl		
H_2O_2 and heat		
Electrolysis of H_2O		
$NaClO_2$ 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.

13.2 Demonstrations

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, #8216;Wow, look at that#8217; 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 #8211; 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 $18\ MH_2SO_4$. 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 $3.4\ gMnSO_4 \cdot H_2O$ 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 $mm \times 150$ 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:

$CaC_2 + H_2O \rightarrow Ca(OH)_2 + C_2H_2$ (ethyne)

$C_2H_2 + O_2 \rightarrow C +$ many reaction products (Soot)

$C_2H_2 + O_2 \rightarrow CO_2 + H_2O$

13.3 Extra Readings

CHAPTER 14

Stoichiometry JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

14.1 Laboratory Activities

Teacher's Pages for Mass-Mass Relationships in a Chemical Change

Lab Notes

The setup for this lab is easy. Prepare the 6 $MHCl$ 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 #8211; 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:

$$C + O_2 \rightarrow CO_2$$
$$12g\ 32g = 44g$$

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:

$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$$

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 $2.0\ g \pm 0.10\ g$ of $NaHCO_3$ into the evaporating dish, and record the mass.
3. Measure 5.0 mL of 6 $MHCl$ 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 $NaHCO_3$ ____________________g

Mass of dish + watch glass + $NaCl$ ____________________g

Moles of $NaHCO_3$ ____________________g

Moles 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.

14.2 Demonstrations

Endothermic Reaction Demonstration

Brief description of demonstration

Two dry powders are mixed together. The resulting powders are shaken to mix, becoming at first doughy, then slushy, then a milky suspension. The flask gets extremely cold. A strong odor of ammonia is emitted.

Materials

- 32 g barium hydroxide-8-hydrate, $Ba(OH)_2 \cdot 8H_2O$
- 11 g ammonium chloride, NH_4Cl
- $2-250$ mL Erlenmeyer flasks
- wash bottle
- small block of non porous wood, $10\ cm \times 10\ cm \times 1-2\ cm$

Procedure

Mass 32 g of barium hydroxide-8-hydrate into a 250 mL flask. Make sure it is finely ground. Mass 11 g of ammonium chloride into the other 250 mL flask. Both substances are hygroscopic, so if the demonstration is to be prepared long before the class, hand stopper the flasks. 17 g of ammonium nitrate may be substituted for the ammonium chloride. Pour the ammonium chloride into the barium hydroxide-8-hydrate, and mix the two powders with a vigorous swirling motion. The mixture will thicken: as it is mixed, more of the water of hydration from the barium hydroxide is released, and the viscosity drops to that of milk. A strong odor of ammonia is emitted. The flask will drop in temperature from room temperature to below $-20^\circ C$ within two minutes. At this point, a few drops of water are added to the surface of the wooden block, and the flask placed on top of it. Let it stand for 30 *seconds* or so. The wooden block will freeze to the flask.

Hazards

The barium hydroxide is a strong base, and barium salts are poisonous. Do not inhale the dust or let the powders get on the skin. Ammonium salts are acidic when mixed with water, and are skin irritants. The ammonia vapor is irritating. Ammonium nitrate is a strong oxidizing agent and an explosive if heated.

Disposal

Rinse the resulting mixture down the sink with a 100 fold excess of water.

Gummi Bear Oxidation (Exothermic Reaction) Demo

Brief description of demonstration

A Gummi Bear candy is dropped into a test tube containing melted potassium chlorate. The ensuing reaction is incredibly exothermic, emitting heat, light, and steam.

This reaction produces a large quantity of heat, flame, and smoke (mostly water vapor). It should be done in a well ventilated room, or, preferably, in a fume hood. Potassium chlorate should be used with caution. It is a strong oxidizing agent, especially when molten. Keep all combustible materials away from the reaction area. Make sure the test tube used is scrupulously clean and the mouth is pointed away from the audience.

Wear goggles and a lab coat.

There are several websites where you can watch this demo . . . type Gummi Bear oxidation into your browser.

Materials

- one 25×150 *mm* Pyrex test tube
- one ring stand with clamp for test tube
- $5-7$ *grams* potassium chlorate
- Bunsen burner
- one candy Gummi bear (or any other candy - M&M#8217;s, gum drops etc.)

Procedure

1. Set up the stand and clamp, and support the test tube in the clamp in a slightly off-vertical position.
2. Add $5-7$ *grams* of potassium chlorate to the test tube (about 1 *cm* in depth).
3. Gently heat the tube with the burner until the potassium chlorate is completely molten. Bubbles of oxygen will begin to form.
4. Remove the burner and use crucible tongs to drop in the Gummi bear, and stand back!

Hazards

On rare occasions, the test tube may crack. It is possible for some molten material to be ejected from the end of the test tube so make sure there are no people in the direction the tube is pointed.

Disposal

The test tube is very difficult to clean and probably should be discarded.

Discussion

This demonstration illustrates the vast amount of energy which is available from the oxidation of carbohydrates, such as sugar. An excess of oxygen, generated by the decomposition of potassium chlorate, will react with the glucose in a Gummi bear, releasing a large amount of energy quickly and dramatically.

Sugar is a fuel that we use for energy. We can also use it to fuel a chemical reaction. In this reaction, the sugar from the Gummi bear reacts with the oxygen produced from the decomposition of potassium chlorate. Purple sparks form, which are a result of the potassium ion (it has nothing to do with the color of Gummi bear you use).

When heated, potassium chlorate decomposes, producing sufficient oxygen to ignite the sugar in the Gummi bear. Since the oxidation of the sugar is very exothermic, sodium chlorate continues to decompose to oxygen, and the rate of combustion becomes very rapid. As we heat the Potassium Chlorate it starts to decompose as follows:

$$2\ KClO_{3(s)} \rightarrow 2\ KCl_{(s)} + 3\ O_{2(g)}$$

This produces oxygen which oxidizes the sugar (glucose) in the Gummi bear. This oxidation is incredibly exothermic $(-5000\ kJmol-1)$. The reaction is:

$$C_6H_{12}O_{6(s)} + 6\ O_{2(g)} \rightarrow 6\ CO_{2(g)} + 6\ H_2O_{(g)}$$

Some questions you might wish to ask students:

1. What is a decomposition reaction?
2. Write a balanced reaction for the decomposition of potassium chlorate.
3. How could we test for the presence of oxygen in this decomposition reaction?
4. Why is it necessary to heat this reaction?
5. What is the difference between an endothermic and an exothermic reaction?

14.3 Extra Readings

CHAPTER 15 The Behavior of Gases JRS - Labs Demos

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

15.1 Laboratory Activities

Teacher#8217;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. $P_{H2} = P_{TOTAL} - P_{H2O} = 755.0\ mm\ of\ Hg - 23.7\ mm\ of\ Hg = 731.3\ mm\ of\ Hg$
2. $V_2 = \frac{P_1V_1T_2}{P_2T_1} = \frac{(730.\ mm\ of\ Hg)(137\ mL)(273\ K)}{(760.\ mm\ of\ Hg)(298\ K)} = 121\ mL$
3. $\frac{0.51\ g}{0.250\ L} = \frac{x\ g}{22.4\ L} \qquad x = 45.7\ g$

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. $V_2 = \frac{P_1V_1T_2}{P_2T_1} = \frac{(775\ mm\ of\ Hg)(498\ mL)(273\ K)}{(760.\ mm\ of\ Hg)(373\ K)} = 372\ mL$

$$\frac{2.31\ g}{0.372\ L} = \frac{x\ g}{22.4\ L} \qquad x = 139\ g$$

Therefore, the molar mass is 139 g/mol.

3. The molar mass of C_4H_{10} is 58.0 g/mol.

moles $C_4H_{10} = \frac{3.60\ g}{58.0\ g/mol} = 0.0621\ moles$

4. $of\ molecules = (moles)(molecules/mole) = (0.0621)(6.02 \times 10^{23} = 3.74 \times 10^{22})\ molecules$

5. $V = \frac{nRT}{P} = \frac{(0.0621\ mol)(0.0821\ L{\cdot}atm/mol{\cdot}K)(273\ K)}{(1.00\ atm)} = 1.39\ liters$

6. $2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$

$\frac{2\ mols\ C_4H_{10}}{13\ mols\ O_2} = \frac{0.0621\ mol\ C_4H_{10}}{x\ mol\ O_2} \qquad x = 0.404\ mol\ O_2$

$grams\ O_2 = (0.404\ mol)(32.0\ g/mol) = 12.9\ grams$

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. mL of hydrogen gas was collected over water at a pressure of 755.0 $mm\ of\ Hg$ and a temperature of $25^\circ C$. If the vapor pressure of water at $25^\circ C$ 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. $mm\ of\ Hg$ and $25^\circ C$, what volume will it occupy under standard conditions?
3. A 250. 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
- $250-300$ 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 = _________$^\circ C$
6. Atmospheric pressure in the lab = ____________mm of Hg
7. Vapor pressure of water at this temperature = __________ mm of Hg

TABLE 15.1: Table of water vapor pressures at normal lab temperatures.

Temperature, $^\circ C$	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.

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

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.$^\circ C$ 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?

15.2 Demonstrations

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#8217;s alive, which can be startling to some. Be aware of over-reaction, and clear away any breakable items that may come into harms#8217; way.

Disposal

Throw the used balloons away.

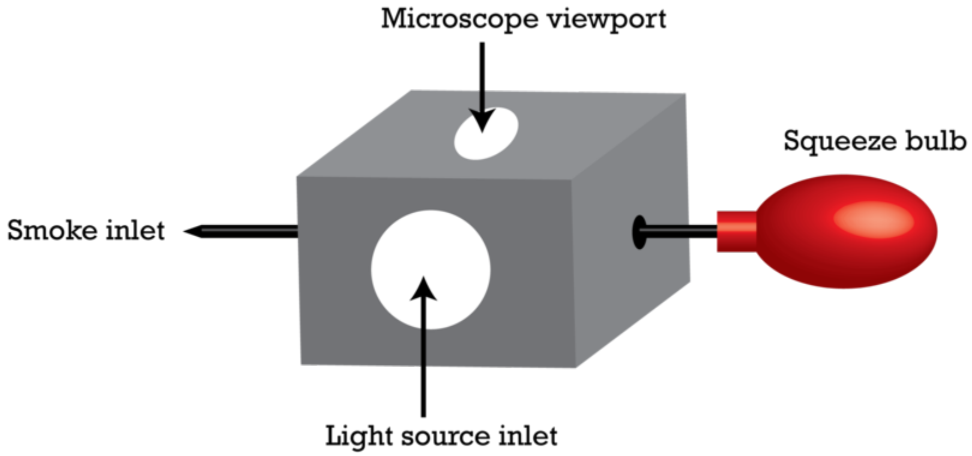
Discussion

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.

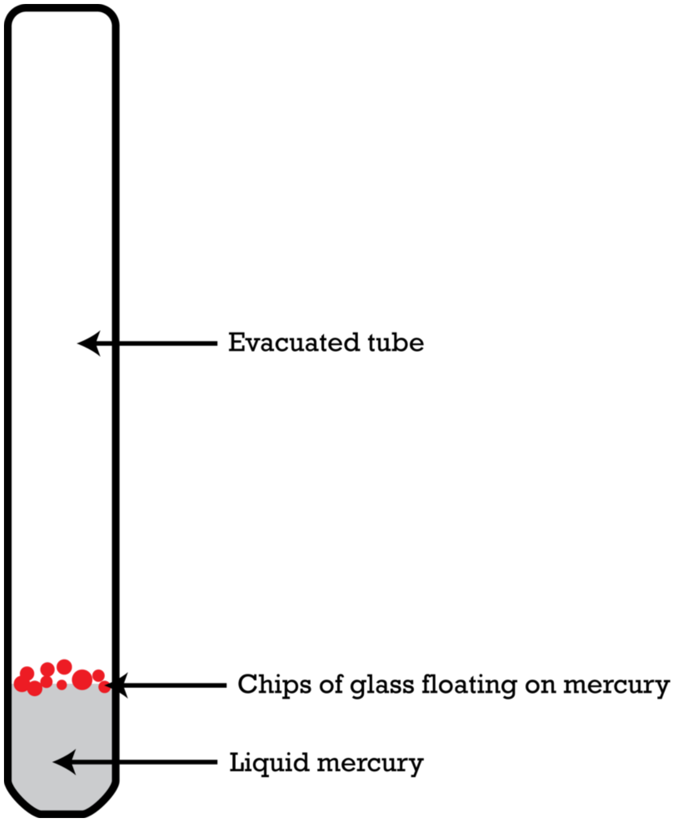
Discussion

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#8217;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#8217;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 #8211; 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).

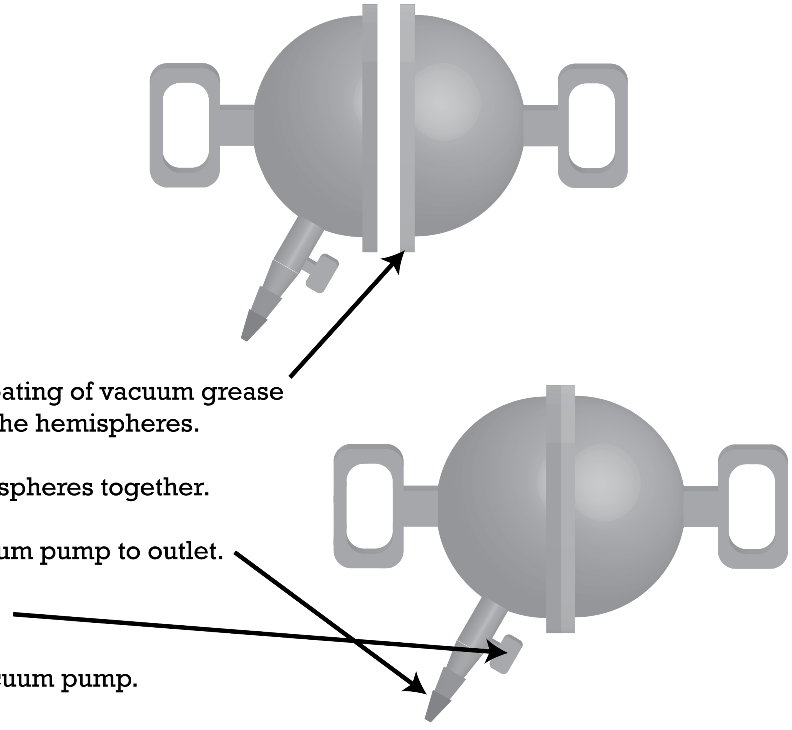
Purpose

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#8217;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#8217;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.

15.3 Extra Readings

CHAPTER 16 Condensed Phases: Solids and Liquids JRS - Labs Demonstrations

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

16.1 Laboratory Activities

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 $\pm 5°C$. 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 #8220;normal boiling point#8221; 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 $0^\circ C$. 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 $110^\circ C$. The first group to do it wins!

TABLE 16.1: Data

Time, in Minutes	Temperature, in $^\circ C$	Time, in Minutes	Temperature, in $^\circ C$
Start		13	
1		14	
2		15	
3		16	
4		17	
5		18	
6		19	

TABLE 16.1: (continued)

Time, in Minutes	**Temperature, in $^\circ C$**	**Time, in Minutes**	**Temperature, in $^\circ C$**
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?

16.2 Demonstrations

Boiling Water in a Paper Cup

Brief description of demonstration

Water is placed in a non #8211; 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 #8211; 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 $4.18\ J/g\cdot{}^\circ C$ vs. less than $1\ J/g\cdot{}^\circ C$ for other materials. Thus it can absorb a lot of heat energy before changing its temperature. Once it reaches its boiling point of $100^\circ C$, 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 $233^\circ C$, 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 it#8217;s hot contents.

16.3 Extra Readings

CHAPTER 17 Solutions and Their Behavior JRS - Labs Demonstrations

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

17.1 Laboratory Activities

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#8217;t use all your sample on the first test. Look for things like whether the substance is magnetic, whether is 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 17.1: 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#8217;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 MNH_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 $MHNO_3$ 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 $MBa(NO_3)_2$ 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 $MKSCN$ Dissolve 4.86 $grams$ of $KSCN$ in sufficient water to make 500 mL of solution.

0.10 $MAgNO_3$ 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 $MBa(NO_3)_2$; 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 $MFe(NO_3)_3$; 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 $MLiNO_3$; Dissolve 3.45 $grams$ of $LiNO_3$ in sufficient water to make 500 mL of solution. Potassium, K^+, use 0.1 $MKNO_3$; Dissolve 5.06 $grams$ of KNO_3 in sufficient water to make 500 mL of solution.

Silver, Ag^+, use 0.1 $MAgNO_3$; Dissolve 8.5 $grams$ of $AgNO_3$ in sufficient water to make 500 mL of solution.

Sodium, Na^+, use 0.1 $MNaNO_3$; Dissolve 4.25 $grams$ of $NaNO_3$ in sufficient water to make 500 mL of solution.

Carbonate, CO_3^{2-}, use 0.1 MNa_2CO_3; Dissolve 5.30 $grams$ of Na_2CO_3 in sufficient water to make 500 mL of solution.

Chloride, Cl^-, use 0.1 $MNaCl$; 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 MNa_2SO_4; Dissolve 7.10 $grams$ of Na_2SO_4 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

1. $AgNO_{3(aq)} + NaI_{(aq)}$

Yes, precipitate forms, AgI

2. $Pb(NO_3)_{2(aq)} + CaCl_{2(aq)}$

Yes, precipitate forms, $PbCl_2$

3. $NH_4NO_{3(aq)} + CaCl_{2(aq)}$

No precipitate forms

4. $Sr(NO_3)_{2(aq)} + K_2SO_{4(aq)}$

Yes, precipitate forms, $SrSO_4$

Qualitative Ion Testing Lab

Background:

How are unknown chemicals analyzed? One method is by making comparisons to #8220;known#8221; 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 #8211; 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.

$$KI_{(s)} \rightarrow K^+_{(aq)} + I^-_{(aq)} \qquad \text{Equation 1}$$

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.

$$Pb(NO_3)_{2(s)} \rightarrow Pb^{2+}_{(aq)} + 2\,NO^-_{3(aq)} \qquad \text{Equation 2}$$

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.

$$2\,K^+_{(aq)} + 2\,I^-_{(aq)} + Pb^{2+}_{(aq)} + 2\,NO^-_{3(aq)} \rightarrow PbI_{2(s)} + 2\,K^+_{(aq)} + 2\,NO^-_{3(aq)} \qquad \text{Equation 3}$$

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.

$$Pb^{2+}_{(aq)} + 2\ I^{-}_{(aq)} \rightarrow PbI_{2(s)} \qquad \text{Equation 4}$$

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.

$$Cu(OH)_{2(s)} + 4\ NH_{3(aq)} \rightarrow Cu(NH_3)_4^{2+(aq)} + 2\ OH^{-}_{(aq)} \qquad \text{Equation 5}$$

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.

$$CaCO_{3(s)} + 2\ H^{+}_{(aq)} \rightarrow Ca^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(L)} \qquad \text{Equation 6}$$

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#8217;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 MNH_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 #8211; 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 $MAgNO_3$ 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 $MHNO_3$ 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 MNH_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 $MBa(NO_3)_2$ 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 $MHNO_3$ 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 $MHNO_3$ to 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 #8211; 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 #8211; Identification of an Unknown Salt

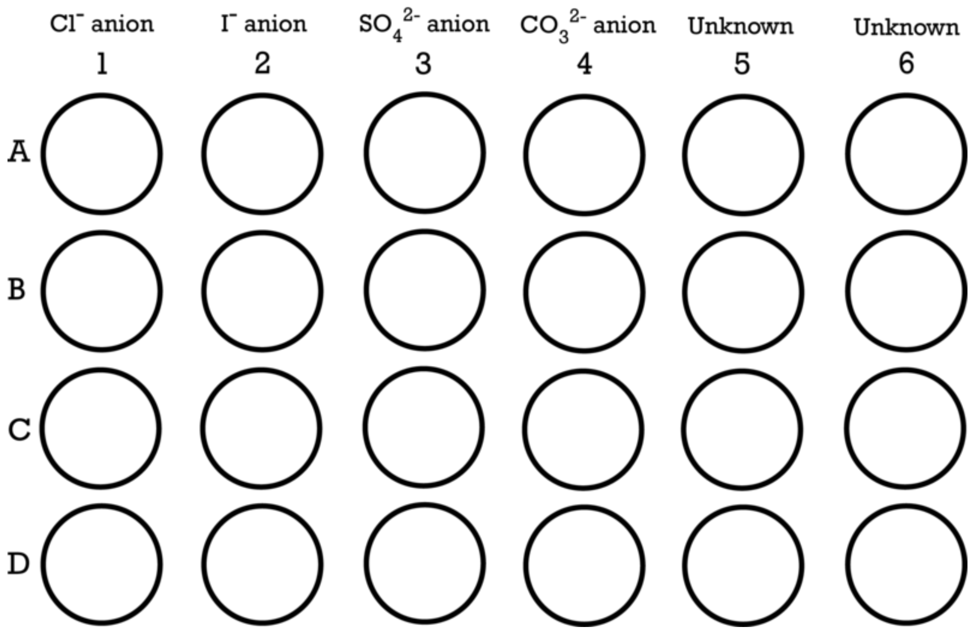
Note: the unknown salt is made up of one of the cations and one of the anions previously tested.

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.

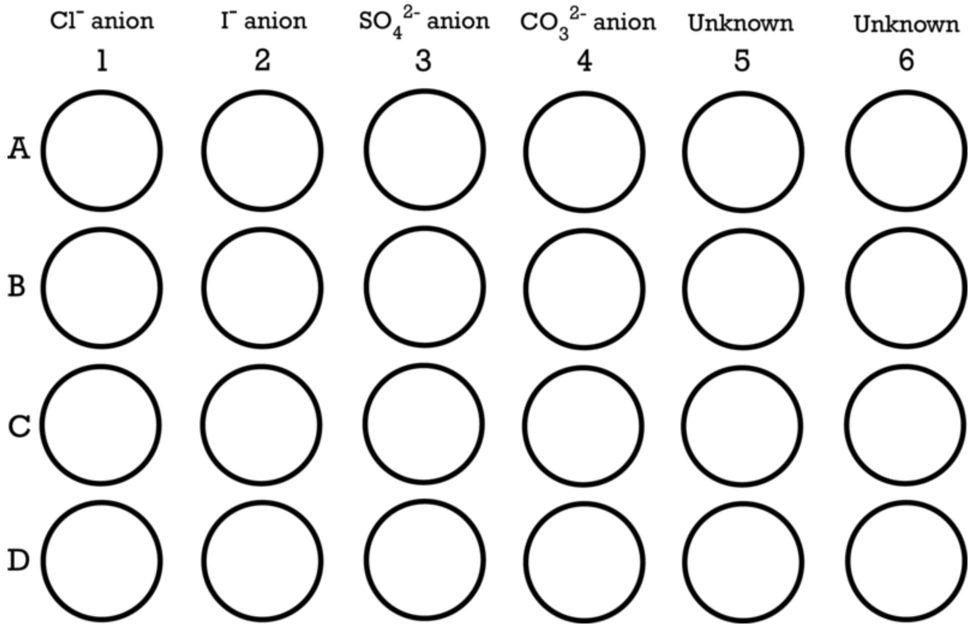
Data for Anion Testing

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.



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.



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 #8211; 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.

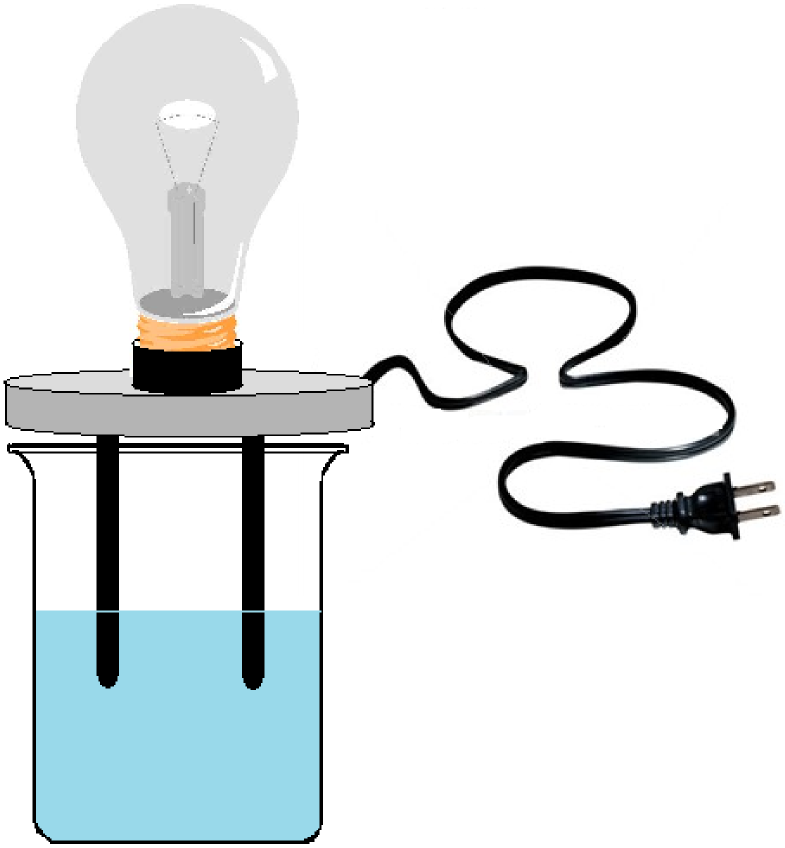
Solubility Rules

- 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+}.

17.2 Demonstrations

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 propanoic acid, $HC_3H_5O_2$
- 100 mL 0.1 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.

17.3 Extra Readings

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 #8220;dissolving#8221;, 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 #8216;n-type#8217; semiconductor). Similar amounts of arsenic dissolved in silicon have equally large effects, but result in different electrical characteristics (the material becomes a #8216;p-type#8217; 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.

CHAPTER 18

Chemical Kinetics JRS - Labs Demonstrations

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

18.1 Laboratory Activities

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

18.2 Demonstrations

Catalytic Oxidation of Alcohol with Copper Demonstsration

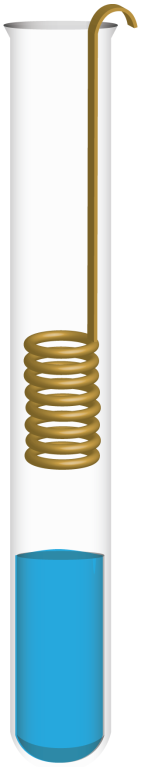
Brief description of demonstration:

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.

$$RCH_2OH \rightarrow RCHO + H_2O$$
$$\text{Alcohol heat} + CuO \text{ Aldehyde + water}$$

Hydrogen Peroxide Decomposition Catalyzed by Manganese Dioxide

Description of Demonstration

Hydrogen peroxide decomposes extremely slowly at room conditions to produce water and oxygen gas.

$$2\ H_2O_{2(aq)} \rightarrow 2\ H_2O_{(L)} + O_{2(g)}$$

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 #8211; 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.

18.3 Extra Readings

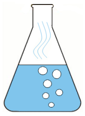
CHAPTER 19 Chemical Equilibrium JRS - Labs Demonstrations

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

19.1 Laboratory Activites

Teacher#8217;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 ·*H_2O</math>ina* 250. ml *volumetricflaskandfilltotheline*.

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.

$$C + Y \leftrightharpoons 3B$$

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 places 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

$$Fe^{3+} + SCN^- \leftrightharpoons Fe[SCN]^{2+}$$

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 $KCSN$ 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 $KCSN$ 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 19.1: **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?

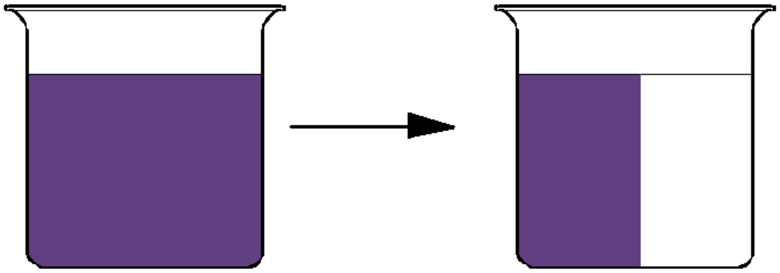
7. Given the information that phosphate ion removes iron (III) ion from solution, describe what happened when phosphate ion was added to test tube 4.

19.2 Demonstrations

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:

1. 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 #8211; so prepare it only when you are ready to use it.)
- 100. mL of 1.0 MH_2SO_4
- sufficient distilled water to bring the total volume to 600 mL.
- mix thoroughly

2. 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 $Fe2+$ ions to $Fe3+$ 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.

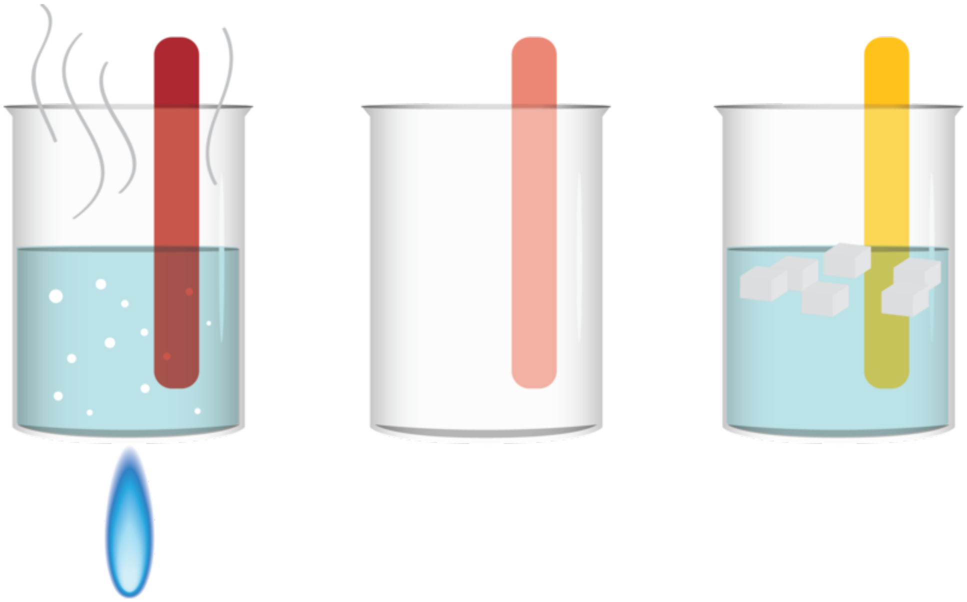
Equilibrium Between Nitrogen Dioxide and Dinitrogen Tetroxide Demonstration

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.

$$2\,NO_{2(g)} \leftrightharpoons N_2O_{4(g)}$$

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.

19.3 Extra Readings

CHAPTER 20

Acids-Bases JRS - Labs Demonstrations

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

20.1 Laboratory Activites

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:
    a. $CuSO_4 \rightarrow Cu^{2+} + SO_4^{2+}$
    b. $Ca(NO_3)_2 \rightarrow Ca^{2+} + 2\ NO_3^-$
    c. $K_3PO_4 \rightarrow 3\ K^+ + PO_4^{3-}$
    d. $KCl \rightarrow K^+ + Cl^-$
    e. $KBr \rightarrow K^+ + Br^-$
    f. $Na_2S \rightarrow 2\ Na^+ + S^{2-}$
    g. $(NH_4)_2CO_3 \rightarrow 2\ NH_4^+ + CO_3^{2-}$
    h. $Na_2CrO_4 \rightarrow 2\ Na^+ + CrO_4^{2-}$
    i. $MgBr_2 \rightarrow Mg^{2+} + 2\ Br^-$
    j. $NaCl \rightarrow Na^+ + Cl^-$

2. $NaC_6H_5CO_2 + H_2O \rightarrow HC_6H_5CO_2 + Na^+ + OH^-$ solution will be basic

Lab #8211; 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:

$$NH_4^+ + H_2O \rightarrow NH_4OH + H^+$$

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:

$$C_2H_3O_2^- + H_2O \ \rightarrow \ HC_2H_3O_2 + OH^-$$

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

1. Write dissociation equations for the following salts:
    a. Copper(II) sulfate
    b. Calcium nitrate
    c. Potassium phosphate
    d. Potassium chloride
    e. Potassium bromide
    f. Sodium sulfide
    g. Ammonium carbonate
    h. Sodium chromate
    i. Magnesium bromide
    j. Sodium chloride

2. 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 20.1: **Data Table**

Well #	Salt	Effect on Indicator	pH	Original Acid	Strong or Weak Acid	Original Base	Strong or Weak Base
1							
2							
3							
4							
5							
6							
7							
8							

TABLE 20.1: (continued)

Well #	Salt	Effect on pH Indicator	Original Acid	Strong or Weak Acid	Original Base	Strong or Weak Base
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.

$$Cu^{2+} + SO_4^{2-} + 2\,H_2O \rightarrow$$
$$Ca^{2+} + 2\,NO_3^{-} + 2\,H_2O \rightarrow$$
$$3\,K^{+} + PO_4^{3-} + 3\,H_2O \rightarrow$$
$$K^{+} + Cl^{-} + H_2O \rightarrow$$
$$Na^{+} + Br^{-} + H_2O \rightarrow$$
$$2\,Na^{+} + S^{2-} + H_2O \rightarrow$$
$$2\,NH_4^{+} + CO_3^{2-} + H_2O \rightarrow$$
$$2\,Na^{+} + CrO_4^{2-} + 2\,H_2O \rightarrow$$
$$Mg^{2+} + 2\,Br^{-} + 2\,H_2O \rightarrow$$
$$Na^{+} + Cl^{-} + H_2O \rightarrow$$

2. How do your observations and pH readings compare with the expected results based on the equations for the hydrolysis reactions?

3. What is a spectator ion? Name the spectator ions present in each hydrolysis reaction in this experiment.

4. 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.

5. 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.

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:

$pH \;=\; -log\,[H^+]$ where $[H^+]$ = the concentration of H^+ 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 x 10^{-14}, what is the chemical#8217;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 #188; 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?

20.2 Demonstrations

20.3 Extra Readings

CHAPTER 21 Neutralization JRS - Labs Demonstrations

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

21.1 Laboratory Activities

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.

$$HOH_{(L)} \rightarrow H^+_{(aq)} + OH^-_{(aq)}$$

The extent of this dissociation is 1.0×10^{-7} $moles/liter$ (at 25°C). Therefore, in all neutral water (and neutral water solutions), the concentration of hydrogen ions is 1.0×10^{-7} M and the concentration of hydroxide ions is 1.0×10^{-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 #8220;power of hydrogen#8221; and established the equation, $pH = -log\ [H^+]$.

In a neutral solution, the hydrogen ion concentration is 1.0×10^{-7} M and therefore, the pH is 7. If the concentration of hydrogen ions is 1.0×10^{-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, phophates, and hemoglobin. The kidneys play a role by eliminating excess electrolytes.

TABLE 21.1: **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
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 #8220;acid-base indicators#8221;. 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 #8220;universal indicator#8221;.

TABLE 21.2: **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#8217;s Universal Indicator), is made by mixing a number of indicators that all change color at different pHs. 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 #8211; 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 phenolophthalein.

Data for Part II

1. Describe the color changes and the pHs around the color change pH for methyl orange.
2. Describe the color changes and the pHs around the color change pH for bromthymol blue.
3. Describe the color changes and the pHs 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 #8211; 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 #8211; 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 CO_2. (The CO_2 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 #8211; Potassium Hydrogen Phthalate #8211; 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 #8211; 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 #8211; student teams:

TABLE 21.3: 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

1. 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.

2. 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.

$$\begin{aligned} Moles\ of\ NaOH \text{ added} &= M \times L = (0.1\ M)(0.040\ L) = 0.004\ moles\ NaOH \text{ added: thus} \\ Moles\ KHP \text{ needed} &= 0.004\ moles \\ Grams\ of\ KHP \text{ needed} &= (moles)(g/mole) = (0.004\ mol)(204.23\ g/mol) = 0.8\ grams\ KHP \end{aligned}$$

3. $Molarity\ of\ KHP = \frac{moles\ KHP}{liters\ of\ solution}$

$Moles\ of\ KHP = \frac{grams\ KHP}{g/mol\ of\ KHP} = \frac{0.759\ g}{204.2\ g/mol} = 0.00372\ mol\ KHP$

$Molarity = \frac{0.00372\ mol\ KHP}{0.0500\ L} = 0.0743\ M\ KHP$

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 it#8217;s molecular weight :

$$Moles\ of\ KHP = \frac{0.521\ g}{204.2\ g/mol} = 0.00255\ mols\ KHP$$

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:

$$\begin{aligned} M &= \frac{moles}{L} \\ \text{so } L &= \frac{moles}{M} = \frac{0.00255\ mol}{0.102\ mol/L} = 0.0213\ L = 21.3\ mL \end{aligned}$$

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, $KHC_8H_4O_4$ (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:

$$KHC_8H_4O_4 + NaOH \rightarrow KNaC_8H_4O_4 + H_2O$$

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, H_2A, two acidic hydrogens, and a triprotic acid, H_3A, three acidic hydrogens. The stoichiometries of the reactions are shown below. You will be told whether your unknown acid is monoprotic, diprotic, or triprotic.

$$\begin{aligned} \text{Monoprotic } & HA + NaOH \rightarrow NaA + H_2O \\ \text{Diprotic } & H_2A + 2NaOH \rightarrow Na_2A + 2H_2O \\ \text{Triprotic } & H_3A + 3NaOH \rightarrow Na_3A + 3H_2O \end{aligned}$$

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 $KHP = 204.23\ g/mol$)
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 $MNaOH$ 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 $MNaOH$ with distilled water. (Calculate, ahead of time, how much water and how much 6 $MNaOH$ will be needed.)

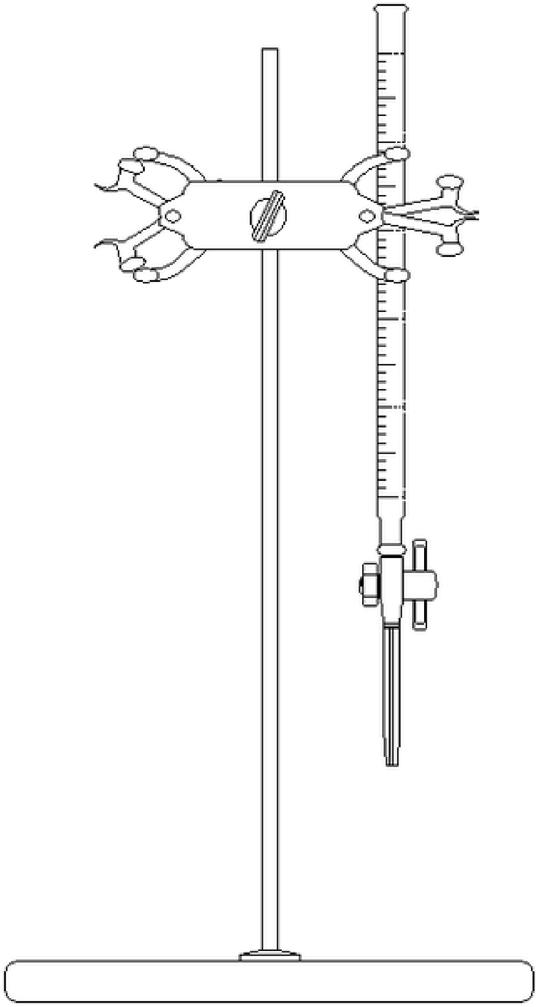
WARNING: Concentrated sodium hydroxide is corrosive and causes severe burns. Handle with care. Dilute and wash up spills with plenty of water. Wash affected skin with water until it no longer feels slippery, but feels #8220;squeaky#8221; clean.

Store the solution in your plastic bottle, label it #8220;0.1 $MNaOH$#8221;, 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 $MNaOH$. 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.

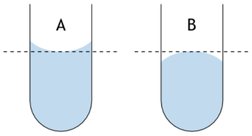
Rinse the inside of a CLEAN buret three times with small quantities of your 0.1 M sodium hydroxide solution (called #8220;rinsing in#8221; with the solution to be used in the buret). Drain the rinses though the stopcock and tip. Do not forget to rinse liquid through the tip, to replace water there. Fill the buret above the 0.0 mL mark with 0.1 $MNaOH$, 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

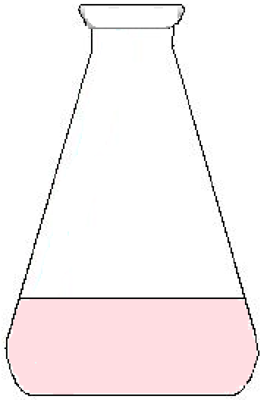
- To read the buret accurately, hold a white card with a black stripe behind the buret, with the black stripe below the meniscus, and the meniscus itself in front of the white region above the black stripe (see illustration). The meniscus will appear black against the white card. Keeping your eye level 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 21.4: **Data Table for Part I**

	Trial 1	**Trial 2**	**Trial 3**
Initial Reading, $NaOH$ buret	mL	mL	mL
Final Reading, $NaOH$ buret	mL	mL	mL
Volume of $NaOH$ added	mL	mL	mL
Grams of acid standard	g	g	g
Moles of acid standard	mol	mol	mol
Molarity of $NaOH$	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 phenophthalein 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 21.5: 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] Moles of acid present = $\frac{moles\ NaOH}{number\ of\ H^{+} ions\ per\ acid\ molecule}$

[2] Molar mass of acid = $\frac{grams\ acid}{moles\ acid}$

Average Molar Mass = _______g/mol

21.2 Demonstrations

21.3 Extra Readings

CHAPTER 22

Thermochemistry JRS - Labs Demonstrations

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

22.1 Laboratory Activities

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

1. $q = mC\Delta t = (500.\ g)(4.18\ J/g \cdot {}^\circ C)(3.5^\circ C) = 7315\ J = 7.3\ kJ$
2. 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

1. (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.

$$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)} \qquad \Delta H_1 =$$

Reaction 2: The reaction of solid sodium hydroxide with dilute hydrochloric acid.

$$NaOH_{(s)} + H^+_{(aq)} + Cl^-_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(L)} \qquad \Delta H_2 =$$

Reaction 3: The reaction of sodium hydroxide solution with dilute hydrochloric acid.

$$Na^+_{(aq)} + OH^-_{(aq)} + H^+_{(aq)} + Cl^-_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(L)} \qquad \Delta H_3 \ =$$

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 $3.5^\circ C$, how many Joules of heat were absorbed?
2. 100. mL of 0.100 $MNaCl$ contains 0.0100 $moles$ of $NaCl(moles = M \times L)$. If 100. mL of 0.100 $MNaCl$ 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 $MNaOH$ solution (prepared the day before so it has time to cool)
- 0.50 $MHCl$ solution
- 0.25 $MHCl$ 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 $MHCl$.

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 $MHCl$ solution and pour it into your calorimeter. Accurately measure 100. mL of 0.50 $MNaOH$ 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 = _______________ g

Final temperature of solution = ____________ g

Δt = ____________ $^\circ C$

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 = _____________$^\circ C$

Final temperature of solution = _____________$^\circ C$

Δt = ____________ $^\circ C$

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 = _____________$^\circ C$

Initial temperature of $NaOH$ solution = ____________$^\circ C$

Average initial temperature = _____________$^\circ C$

Final temperature of mixture = _____________$^\circ C$

Δt = ____________ $^\circ C$

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.

$$\%\ error = \frac{true\ value - experimental\ value}{true\ value} \times 100 =$$

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?

22.2 Demonstrations

22.3 Extra Readings

CHAPTER 23 Electrochemistry JRS - Labs Demonstrations

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

23.1 Laboratory Activities

Teacher#8217;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#8217;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.

$$\begin{aligned}
Cu_{(s)} &\rightarrow Cu^{2+} + 2e^- \\
Zn_{(s)} &\rightarrow Zn^{2+} + 2e^- \\
Mg_{(s)} &\rightarrow Mg^{2+} + 2e^- \\
Pb_{(s)} &\rightarrow Pb^{2+} + 2e^-
\end{aligned}$$

2. Write reduction half-reactions for the three halogens, Cl_2, Br_2, and I_2.

$$\begin{aligned}
Cl_{2(aq)} + 2\ e^- &\rightarrow 2\ Cl^-_{(aq)} \\
Br_{2(aq)} + 2\ e^- &\rightarrow 2\ Br^-_{(aq)} \\
I_{2(aq)} + 2\ e^- &\rightarrow 2\ I^-_{(aq)}
\end{aligned}$$

Answers to Post-Lab Questions

1. Write a balanced net ionic equation for all the reactions that occurred with the metals.

$$Mg_{(s)} + Zn^{2+}_{(aq)} \rightarrow Zn_{(s)} + Mg^{2+}_{(aq)}$$
$$Mg_{(s)} + Pb^{2+}_{(aq)} \rightarrow Pb_{(s)} + Mg^{2+}_{(aq)}$$
$$Mg_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Mg^{2+}_{(aq)}$$
$$Zn_{(s)} + Pb^{2+}_{(aq)} \rightarrow Pb_{(s)} + Zn^{2+}_{(aq)}$$
$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$$
$$Pb_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Pb^{2+}_{(aq)}$$

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.

$$Mg_{(s)} \rightarrow Mg^{2+} + 2e^-$$
$$Zn_{(s)} \rightarrow Zn^{2+} + 2e^-$$
$$Pb_{(s)} \rightarrow Pb^{2+} + 2e^-$$
$$Cu_{(s)} \rightarrow Cu^{2+} + 2e^-$$

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.

$$Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$$
$$Pb^{2+} + 2e^- \rightarrow Pb_{(s)}$$
$$Zn^{2+} + 2e^- \rightarrow Zn_{(s)}$$
$$Mg^{2+} + 2e^- \rightarrow Mg_{(s)}$$

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.

$$Mg_{(s)} \rightarrow Mg^{2+} + 2e^-$$
$$Zn_{(s)} \rightarrow Zn^{2+} + 2e^-$$
$$Pb_{(s)} \rightarrow Pb^{2+} + 2e^-$$
$$Cu_{(s)} \rightarrow Cu^{2+} + 2e^-$$
$$Ag_{(s)} \rightarrow Ag^{+} + e^-$$

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.

$$Cl_{2(aq)} + 2\,Br^- \rightarrow 2\,Cl^-_{(aq)} + Br_{2(aq)}$$
$$Cl_{2(aq)} + 2\,I^- \rightarrow 2\,Cl^-_{(aq)} + I_{2(aq)}$$
$$Br_{2(aq)} + 2\,I^- \rightarrow 2\,Br^-_{(aq)} + I_{2(aq)}$$

3. List the reduction half-reactions of the halogens in order of decreasing activity.

$$Cl_{2(aq)} + 2\,e^- \rightarrow 2\,Cl^-_{(aq)}$$
$$Br_{2(aq)} + 2\,e^- \rightarrow 2\,Br^-_{(aq)}$$
$$I_{2(aq)} + 2\,e^- \rightarrow 2\,I^-_{(aq)}$$

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.

$$Ca_{(s)} \rightarrow Ca^{2+} + 2e^-$$
$$Fe_{(s)} \rightarrow Fe^{2+} + 2e^-$$
$$Au_{(s)} \rightarrow Au^{3+} + 3e^-$$

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:

$$Ca_{(s)} + Fe^{2+} \rightarrow Ca^{2+} + Fe_{(s)} \qquad \text{Reaction 1}$$
$$3\,Ca_{(s)} + 2\,Au^{3+} \rightarrow 3\,Ca^{2+} + 2\,Au_{(s)} \qquad \text{Reaction 2}$$

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:

$$Fe_{(s)} + Ca^{2+} \rightarrow Fe^{2+} + Ca_{(s)} \qquad \text{Reaction 3 (Does not occur)}$$
$$3\,Fe_{(s)} + 2\,Au^{3+} \rightarrow 3\,Fe^{2+} + 2\,Au_{(s)} \qquad \text{Reaction 4}$$

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.

$$Au_{(s)} + Ca^{2+} \rightarrow \text{No reaction} \qquad \text{Reaction 5}$$
$$Au_{(s)} + Fe^{2+} \rightarrow \text{No reaction} \qquad \text{Reaction 6}$$

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, Cl_2, Br_2, and I_2.

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#8217; 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.

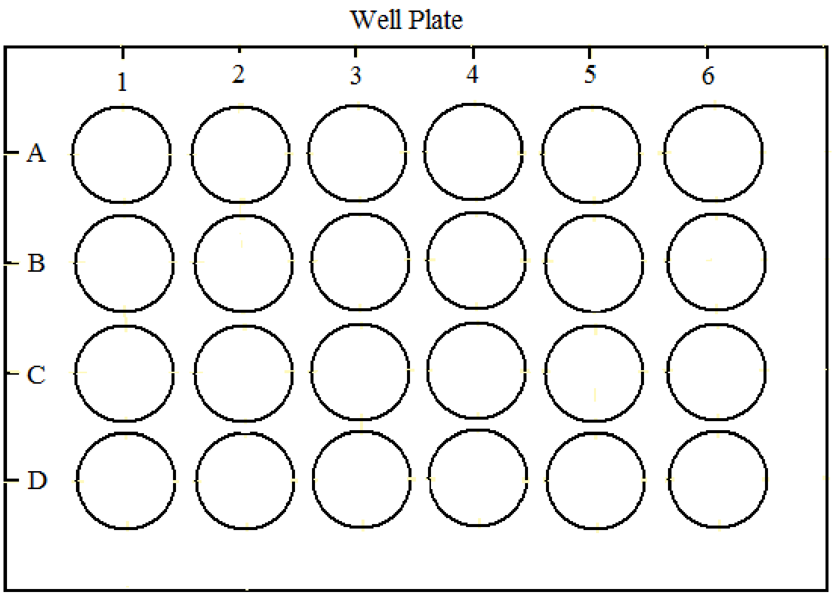
Materials and Materials (per lab group) for Part I

- 4 - copper foil, $6 \times 6\ mm$ squares
- 4 - zinc foil, $6 \times 6\ mm$ squares
- 4 - magnesium foil, $6 \times 6\ mm$ squares
- 4 - lead foil, $6 \times 6\ mm$ 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 #8211; 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 many 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#8217;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+}
$Cu_{(s)}$				
$Mg_{(s)}$				
$Pb_{(s)}$				
$Zn_{(s)}$				

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.
4. 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?
5. 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.)

$$X_{2(aq)} + 2\,Y^-_{(aq)} \rightarrow Y_{2(aq)} + 2\,X^-_{(aq)}$$

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. It 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	$Cl_{2(aq)}$	$Br_{2(aq)}$	$I_{2(aq)}$
$Cl^-_{(aq)}$	▨		
$Br^-_{(aq)}$		▨	
$I^-_{(aq)}$			▨

Post-Lab Questions for Part II

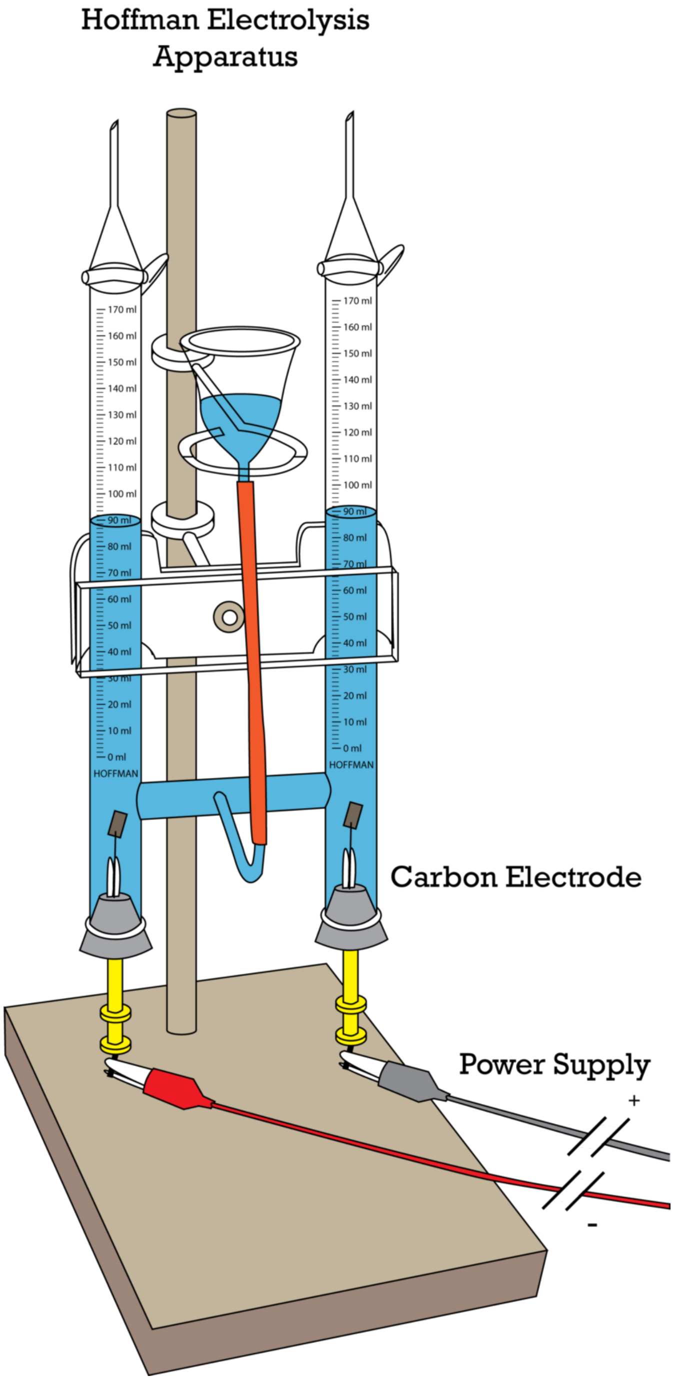
1. Why is it assumed that the halide ions will not dissolve in mineral oil?
2. Write net ionic equations for all the reactions that occurred with the halogens.
3. List the reduction half-reactions of the halogens in order of decreasing activity.

23.2 Demonstrations

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.



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 tunes, 15 $mm \times 125$ 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#8217;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.

Turning a Copper Coin to #8220;Silver#8221; and #8220;Gold#8221;

Teaching notes

- It is advisable to carry out a trial experiment before performing the demonstration for your classes.
- If the mixture of sodium zincate and zinc powder is cloudy, allow it to cool, and then filter off the zinc powder to leave a clear filtrate. Place a small piece of zinc in the liquid as a substitute for the zinc powder.

Brief description of demonstration

A #8220;copper#8221; coin is dipped into a solution of sodium zincate in contact with zinc metal. The coin is plated with zinc and appears silver in color. The plated coin in held in a Bunsen burner flame for a few seconds and the zinc and copper from an alloy of brass. The coin now appears gold.

Apparatus and Materials

- Eye protection: goggles
- Disposable gloves (preferably nitrile)
- Beaker, 250 *mL*
- Bunsen burner
- Hot plate
- Tripod and wire gauze
- Pair of tongs or long forceps
- Glass stirring rod
- Zinc powder (highly flammable), 5 *g*
- Sodium hydroxide pellets (corrosive), 24 *g*
- Steel wool
- Distilled water
- Copper coins (American pennies will do or other copper coins)

Procedure

Before the demonstration

1. Dissolve 24 *g* of $NaOH$ in 100 *mL* of distilled water in a 250 *mL* beaker, stirring continuously. The solution will get warm and is corrosive.

2. Heat the solution to boiling on a hot plate. (Note: it a Bunsen burner is used to heat the solution, the Bunsen burner MUST be turned off before the next steps. Both zinc powder and evolved hydrogen in the next step are highly flammable.)

3. Add 5 *g* of zinc powder carefully. The solution will fizz as some of the zinc dissolves forming sodium zincate and giving off hydrogen gas. NO FLAMES ALLOWED!

4. Clean a copper coin with steel wool until it is shiny.

The demonstration

5. Drop the cleaned coin into the hot solution containing sodium zincate and the remaining zinc powder. The coin must make contact with the powdered zinc at the bottom of the solution. If necessary, use a glass rod to move the coin until it touches the un-reacted zinc.

6. Leave the coin in the solution until it is plated with a shiny coat of zinc. This will take about 2 or 3 *minutes*. Leaving the coin too long may causes lumps of zinc to stick to it.

7. Remove the plated coin with tongs or long forceps and rinse it under running tap water to remove traces of sodium hydroxide, sodium zincate, and any powdered zinc.

8. Show the #8220;silver#8221; coin to the audience.

9. Using tongs or forceps, hold the plated coin in the upper part of a roaring Bunsen flame for few seconds until the surface turns gold. Turn the coin so that both sides are heated equally. Overheating will cause the coin to tarnish.

10. Allow the coin to cool and show it to the audience.

Disposal

Solutions may be poured down the sink with 100 fold excess of water.

Discussion

The reaction between zinc and sodium hydroxide to form sodium zincate is:

$$Zn_{(s)} + 2\ NaOH + 2\ H_2O \rightarrow Na_2[Zn(OH)_4]_{2(aq)} + H_{2(g)}$$

The plating reaction is an electrochemical cell. The reaction will not take place unless the zinc and copper are in contact.

The coating of zinc gives the appearance that the coin is now silver.

On heating the coin in the Bunsen burner flame, brass is formed by the zinc migrating into the surface layer of the copper. This gives a gold appearance to the coin. Brass is an alloy of copper containing between 18% and 40% zinc.

23.3 Extra Readings

CHAPTER 24

Nuclear Chemistry JRS - Labs Demonstrations

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

24.1 Laboratory Activities

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

24.2 Demonstrations

24.3 Extra Readings

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 in 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).[5,6]

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[7] 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.[7]

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[8]. 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.

CHAPTER 25

Organic Chemistry JRS - Labs Demonstrations

Chapter Outline

SOURCE: Chemistry Labs, Demonstrations, and Extra Readings

25.1 Laboratory Activities

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#8217; Chatelier#8217;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 $60^\circ C$. Methanol boils at $62^\circ C$.

Answers to Pre-Lab Questions

1. The reaction is dehydration synthesis:

$$R'OH + RCOOH \rightarrow RCOOR' + HOH$$

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 $mm \times 200$ 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 #8211; alcohol pairings.

TABLE 25.1: 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 25.2: **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 #8211; what combination of carboxylic acid and alcohol would result in this ester?

25.2 Demonstrations

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!