Green Chemistry

Laboratory Manual

First Year High School
Chemistry Course
Teacher's Manual

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Getting Acquainted

TN Standard ***2.1:*** *The student will investigate the characteristics of matter.*

Welcome to the chemistry laboratory!

Going to a new school can be very exciting and challenging. There are always a lot of things to learn, such as the location of your classes, the library, cafeteria, and gym. You also have new rules to learn and what some things are called. When you walk into a chemistry laboratory the first time, there are also many new things to learn. At first it may seem very confusing and dangerous. But it is really just like going to someplace new. You will need to learn the location of certain items, the rules, and what some of the things are called. After you get acquainted with your surroundings, the chemistry lab will no longer seem confusing or even dangerous.

Introduction

Playing a game is more enjoyable than just watching or reading about it. This is also true about learning chemistry. Getting to actually do experiments is a lot of fun. But chemistry can also be dangerous, and getting hurt is definitely not fun! For this reason there are safety rules to follow and protective equipment to know its location and how to use it.

This manual has "*Green Chemistry*" in the title. This means that when compared to similar experiments, the ones in this manual are generally safer and produce less hazardous waste. Each of these experiments has been performed safely by students. However, safer and less hazardous waste does not mean accidents, injuries, or damage cannot happen. The authors and publisher take no responsibility for any that may occur.

Since accidents can happen, you never know when you may need to use some of the safety equipment. A brief description on how to use safety equipment often found in a chemistry laboratory is given below.

Safety equipment

1. ***Safety Shower*** **–** A safety shower is used when a hazardous chemical is spilled on a person where they are unable to rinse it off thoroughly in the sink. It can also be used if a person's clothes catch on fire. Most safety showers are operated by pulling a chain. If a chemical is spilled on you so that you need to wash it off in a safety shower, let your teacher know and get there fast! The sooner the chemical is rinsed off, the less damage it will do. All of your contaminated clothes will need to be removed. You can start doing this on the way to the shower. This is not a time for modesty! Your teacher will clear the room if necessary.

2. ***Eye Wash*** – An eye wash is used if a harmful chemical is splashed into your eyes or face. It is usually operated by pushing forward on the handle.

3. ***Fire Extinguisher*** – A fire extinguisher is used to put out small to medium fires.

4. ***Fire Blanket*** – A fire blanket is wrapped around a person to smother the fire. It is used whenever a person cannot get to a safety shower.

5. ***Laboratory Fume Hood*** - A laboratory fume hood removes harmful gases and fumes that are sometimes present when doing an experiment. You should always work in a fume hood whenever you are working with corrosive, noxious, or flammable materials.

6. ***Safety Goggles*** – Safety goggles are worn to protect your eyes. They should be worn at all times when you are in the chemistry laboratory. Even if you are not working with chemicals, someone else might be. They could splash a chemical in your eyes.

7. ***Gloves*** – At times you may need to wear gloves to protect your hands from harmful chemicals or hot objects. The type of glove needed will depend on the application. For example, oven mitts are worn to remove hot objects from an oven while latex gloves are often used when working with acids and bases. It is very

important not to touch your work area with gloves that have been contaminated with harmful chemicals.

Besides knowing where the safety equipment is located and how and when to use it, there are general safety rules you will also need to follow in the laboratory. Some of the common safety rules are listed below. Your teacher may have a few others.

Chemical Laboratory Safety Rules

1. Always wear safety glasses or goggles. Never wear contact lenses.
2. Never attempt unauthorized experiments.
3. Never work alone in the laboratory.
4. Never bring food, drink, or tobacco into the laboratory.
5. Always keep your work area free of clutter.
6. Always wear sensible clothing. This includes no bare midriffs or open-toe shoes.
7. Know the location of and how to use safety equipment in your laboratory. This includes safety showers, fire blankets, fire extinguishers, and eye wash fountains.
8. Always read the experiment before doing it.
9. Report any injury (no matter how small) to the instructor.
10. Always wash hands before leaving the lab.
11. Tie back long hair.
12. Never run or play practical jokes in the laboratory.
13. Place broken glass in a broken glass container, never in a trash can.

In addition to following the general safety rules, chemicals need to be handled properly. Listed below are some guidelines on how to handle chemicals properly.

Handling Chemicals Properly

1. Always add acids to water, never water to acids.
2. Work with small containers when dispensing liquid reagents.
3. Never return unused chemicals to stock bottles.
4. Dispose of used chemicals in the proper waste containers.
5. Always clean the work area, and put away extra equipment when laboratory work is completed.
6. Never leave anything unattended while it is being heated or is reacting rapidly.
7. Take a container to where chemicals are located. Do not take bottles containing chemicals to your work area.

8. Never carry out a reaction or heat a substance in a closed system. Pressure can increase within the system causing glassware to break and/or chemicals to forcefully escape.
9. Always use the fume hoods whenever working with corrosive, noxious, or flammable materials.
11. Always be careful when working with previously heated objects.
12. Always replace stoppers or lids on reagent bottles.
13. Weigh chemicals in containers or on paper provided for that purpose. Never weigh chemicals by placing them directly on the pan of the balance.
14. Report spills to your teacher who will give information as to how to safely clean up the spill.
15. Label all chemicals clearly and completely.
16. Read labels carefully before using chemicals.
17. Always lubricate glass tubing or thermometers before inserting them into rubber stoppers.

Even though you follow all of the safety rules, accidents can still happen. This is why it is so important to know what to do for each type of accident. Remember that no matter how minor an accident or injury may seem, be sure to report it to your teacher. Listed below are ways to respond to accidents that occasionally may occur in the chemical laboratory.

How to Respond to Accidents

1. ***Chemical Spills on the bench or floor* –** If the spill involves volatile or flammable materials, make sure ALL flames in the lab are extinguished and spark-producing equipment is shutdown. Ask your teacher how to properly clean up the chemical.

2. ***Hazardous chemical spills on a person* –** Immediately notify the lab instructor. If it is a large area, the typical course of action is to remove all contaminated clothing while the person is under the safety shower. If it is a small area, flush the area immediately with a large amount of water and then wash it with soap. Medical assistance may be necessary.

3. ***Chemicals spills in the eyes* –** If a harmful chemical is splashed on your face and/or in your eyes, immediate attention is critical. Call for help and get to the nearest eyewash. If the chemical splashes on your face, and you have goggles on, KEEP the goggles on. Remove the chemical from your face before you remove the goggles. If a chemical gets in your eyes, hold your eyes open in the eyewash for at least 20 minutes. Even though you should not be wearing contact lenses in the lab, if you are,

rinse your eyes thoroughly, remove your contacts, and continue to rinse your eyes. A doctor should examine your eyes as soon as possible.

4. ***Ingestion of chemicals –*** Notify the instructor so immediate and appropriate action can be taken for the specific substance ingested.

5. ***Burns –*** Flush the area with cool running water for 20 minutes. Medical attention may be necessary.

6. ***Cuts and wounds*** –- Avoid contamination; notify instructor so appropriate action can be taken.

7. ***Fire –*** Fires in a laboratory are often contained in a piece of glassware called a beaker. You should not move a beaker that has a chemical burning in it. Instead, control the fire where it occurs. These fires can usually be put out simply by covering the mouth of the beaker with a thin curved piece of glass called a watch glass and turning off the source of the flame. Be sure to report it immediately to your teacher. If the fire cannot be extinguished by covering it, your teacher can often put it out with a fire extinguisher. However, if the fire is too large, your teacher will have to clear the building and call the fire department.

   Clothing fires can be extinguished in a safety shower if it is close by. If it is not very close, you will need to STOP, DROP, and ROLL to quickly smother the fire.

Bunsen Burner Safety

A Bunsen burner is often used as a source of heat in the chemistry laboratory. It works by burning a mixture of natural gas and air and to produce a flame. Since a flame is present extra precautions must be taken. Safety rules to follow when using a Bunsen burner are listed below.

1. Always wear safety goggles when using a Bunsen burner.
2. Never use a Bunsen burner to heat organic solvents such as alcohols. Instead use a heating mantle or hot plate in a hood to heat organic solvents.
3. Never light a Bunsen burner when organic solvents are present.
4. Always check the tubing you will use to connect the Bunsen burner to the gas source for cracks. Replace the tubing if it will cause a gas leak.

5. If you are using a striker to light the burner, make sure it is operating properly before turning the gas on. Replace the striker's flint tip if it is worn to where there is no flint left.

6. If you are using a match to light the Bunsen burner, never place the lit match directly over the top of the burner. Bring the flame to the side of the burner tube below the top and then gradually raise the flame up until the burner lights. Be sure the match is completely extinguished before disposing of it.

7. Turn off the gas immediately if the flame goes out or you smell gas after the Bunsen burner is lit.

When you follow the safety rules the chemistry laboratory can be a very interesting place to explore and learn.

Objectives

- To learn how work safely in the chemical laboratory.
- To learn when and how to use the safety equipment in the chemical laboratory.
- To learn the names of the equipment used in the experiments.
- To learn how to properly use a Bunsen burner.

Date:________ Name:_______________________ Period_________

Pre-lab Questions

1. What should you always wear to protect your eyes when you are in the chemistry laboratory?

   Goggles or eye protection should always be worn.

2. Do you add acid to water or water to acid?

   Acid to water

3. Where should you dispose of broken glass?

   Broken glass should be carefully placed in a broken glass container; never in the trash.

4. What should you do if you spill a chemical on you hand?

   You should immediately notify the lab instructor. Because it is a small area, flush it immediately with a large amount of water and then wash with soap.

5. What should you immediately do if you are using a Bunsen burner and the flame goes out or you smell gas?

   You should quickly turn off the gas.

Now that you are acquainted with safety and the chemistry laboratory, you need to sign an agreement to practice safety whenever you are in the lab. Be sure to sign and date both of the Safety Agreements on the next page. Give one of them to your teacher and keep the other one.

SAFETY AGREEMENT

I have carefully read the safety instructions, and listened to the safety lecture presented by my chemistry teacher. I understand the importance of practicing chemical safety at all times and my right to know about the materials used in the lab. I recognize my responsibility to follow these practices and precautions while I am present in the laboratory. **When I am in the laboratory I will wear the recommended eye and personal protection; I will follow the recommended procedures for working safely in the laboratory.**

Name ____Date_______

Signature____

Course name ____

SAFETY AGREEMENT

I have carefully read the safety instructions, and listened to the safety lecture presented by my chemistry teacher. I understand the importance of practicing chemical safety at all times and my right to know about the materials used in the lab. I recognize my responsibility to follow these practices and precautions while I am present in the laboratory. **When I am in the laboratory I will wear the recommended eye and personal protection; I will follow the recommended procedures for working safely in the laboratory.**

Name ____ Date_____

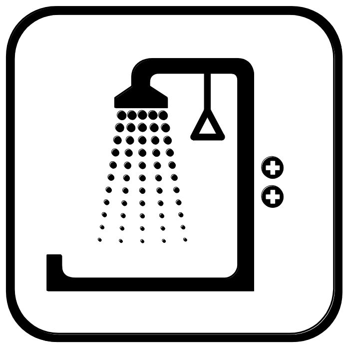
Signature ____

Course name ____

Procedure

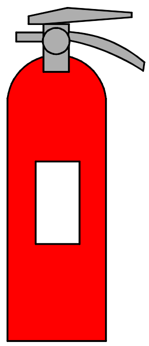
Part 1: Safety First!

1. Look around the lab and see if you can find these safety items. Write down their locations in the spaces provided.



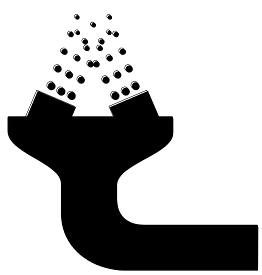
Safety Shower with Pull Chain

Location____________________



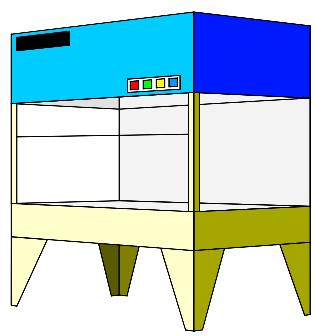
Fire Extinguisher

Location________________________



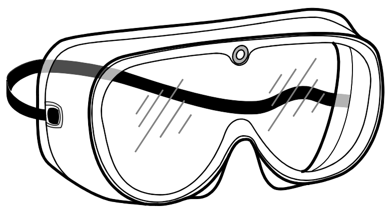
Eyewash Facility

Location____________________



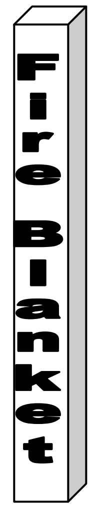
Laboratory Fume Hood

Location________________________



Safety Goggles

Location__________________



Fire Blanket

Location ____________________

2. List any other safety equipment or safety related items that are in the laboratory. Briefly describe their use.

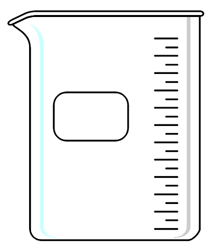
Part 2: What is it?

1. A chemical laboratory contains special equipment to use while you are performing an experiment. Locate each of these items in your lab drawer or the place your teacher tells you to look. Place a check by it when you find it.

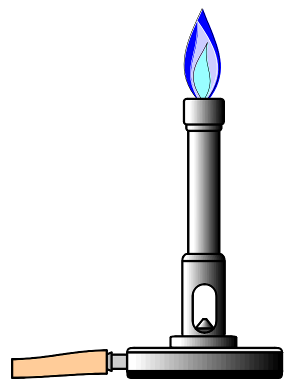
Name_______________________________________

Lab Partner__________________________________

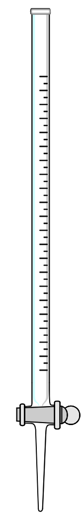
Lab Drawer Number_______



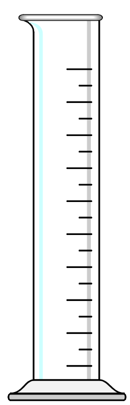
Beakers
50 mL _____ _____
150 mL _____ _____
250 mL _____ _____
600 mL _____ _____



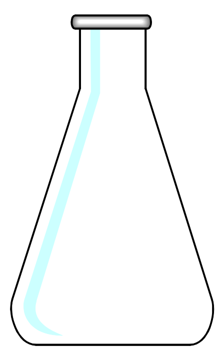
Bunsen Burner _____ _____



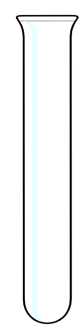
Buret _____ _____



Graduated Cylinders
10 mL _____
25 mL _____
100 mL _____

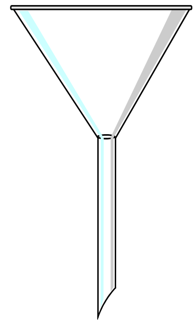


Erlenmeyer Flasks
125 mL_____ _____
250 mL_____ _____

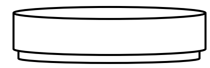


Test tubes
6 Small_____ _____
11 Medium_____ _____
5 Large_____ _____
2 Extra Large_____ ___

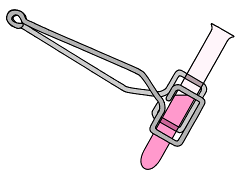
Iron Ring____ ____



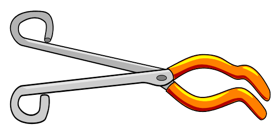
Funnel____ ____



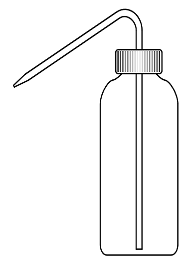
Petri Dish____ ____



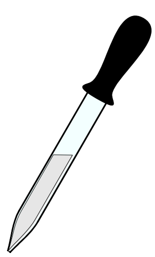
Test Tube Holder_____ ____



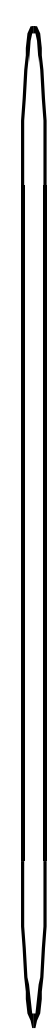
Tongs_____ ____



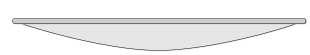
Water Bottle_____ _____



Eye Dropper____ ____



Stirring Rod____ ____



Watch Glass____ ____

2. Sketch a picture and name any other items that are located in your laboratory drawer or where the other items are located.

Part 3: Lighting a Bunsen Burner

1. Make sure the gas valve located at the base of the Bunsen burner turns easily. Also make sure you can adjust the air intake by freely turning the cylinder area with the holes that let air in.

2. Make sure the hose for your Bunsen burner does not have cracks and holes where gas may leak out. After you are sure it is safe to use, connect one end to the Bunsen burner and the other to the gas source.

3. Turn on the gas and light the Bunsen burner by either using a striker or a flame source such as a match. If using a match, bring the flame up from the side to the top of the burner head. Caution: If after you light the flame you still smell gas, turn the gas off immediately!

4. Adjust the gas and air intake until you have a 2-3 inch blue flame that has a lighter blue inner cone. The top of the blue inner cone is the hottest part of the flame.

5. Turn off the gas. Let the Bunsen burner cool to room temperature. Remove the tubing from the gas source and place the Bunsen burner back where it belongs.

Lab
2

Introduction to Significant Figures and Measurements

*TN Standard **2.2** : The student will explore the interactions of matter and energy.*

Why is it really important to make accurate and precise measurements?

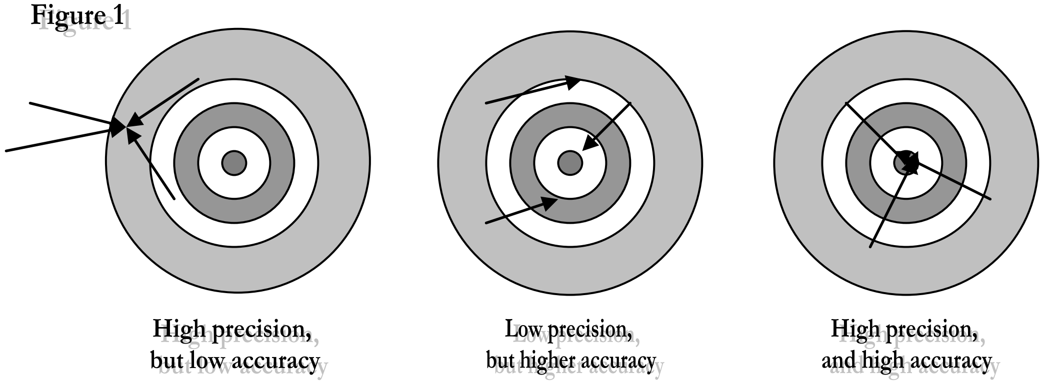
The loss of the Mars Climate Orbiter spacecraft in late September of 1999 was a tragic blow to NASA's Mars exploration project. The $320 million dollar mission ultimately failed due to a miscalculation in the conversion of English measurements to metric measurements. NASA has reported the spacecraft's engine appeared to have burned up while in orbit mostly likely due to a miscalculation of altitude. The problem occurred when the supervisors of the project had measured in miles and feet, instead of kilometers and meters. During the project's development all directions had specified that the metric system be used as the standard measurement. The mistake of using improper units cost NASA millions of dollars, much embarrassment, and a wealth of potential information.

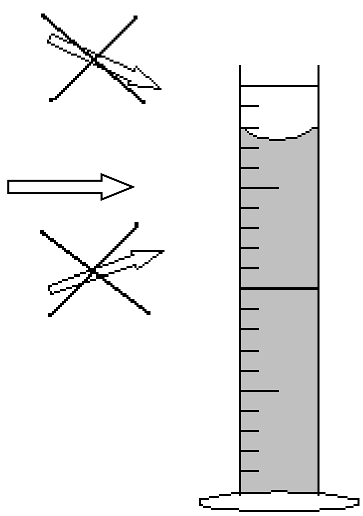
Introduction

Throughout this lab, we will learn how to make measurements with both accuracy and precision while using significant figures and converting units. **Precision** is the agreement between repeated measurements under same conditions. **Accuracy** is how close a measurement is to an accepted measurement.

For example, you are throwing darts on a dartboard and all your arrows go to the same place on the dartboard, but they are nowhere near the bull's eye. You are precise in your technique of hitting the dartboard, but your accuracy is very bad. In this same manner, we will use techniques in the lab that can be repeated, and if under the same conditions, will give a high precision. However, as chemists, this is not good enough because we want our data to be accurate or as close as possible to the accepted value. If we use our thermometer and find that water boils at 94º Celsius, that is not accurate because the accepted temperature at which water boils at 100º Celsius. As a chemist, getting correct measurements is important, as well as recording the correct digits.

Figure 1





The use of **significant figures** enables a chemist to decide whether certain digits in a measurement or calculation are important enough to include in data. When it comes to measurements, the rule of thumb is to record the certain digits and estimated one further decimal place. For example, you are measuring the amount of liquid in millimeters in a graduated cylinder (Figure 2). According to Figure 2 you see that there is at least1 mL. For the tenths place you can see that the meniscus lies above the 1.7 mL mark. Since there is a little more than 1.7 mL, you are allowed one guess. It is about half way between 1.7 and 1.8, so you might guess 1.75 mL. In this measurement there are three significant figures- two are certain and one is estimated.

Is zero significant when using significant figures? Sometimes it is, but not when it is simply used as a placeholder in a measurement. For example, suppose a watermelon has a mass of 1050 grams. The number of significant figures includes the 1, the zero in between, and the 5. The zero on the end could be simply acting as a placeholder to allow one to know that it is 1050 and not 105 or it could be the actual value. Scientists will often use exponential notation so that others can tell if the

Figure 2
When measuring a liquid, there is a certain place that one must measure - the bottom of the meniscus. **The meniscus is the curved line that a liquid makes when placed in a narrow container.** When looking for the bottom of the meniscus, one must look straight at it. When one's line of sight is too high, then the reading that is received is too low. When one's line of sight is too low, then the reading received is too high.

last zero is significant. If it is not, a scientist would write the number as 1.05 x 10^3 g. However, if the last zero is significant, the scientist would write the number as 1.050 x 10^3 g. Zeros on the front of numbers, such as 0.00256, are not significant figures. The number 0.00256 has three significant digits. Remember the one zero before and the two zeros after the decimal act as placeholders.

Now that we understand what to include as a significant figure, how do we use the significant figures in calculations? When it comes to multiplication and division, the rule is the sum must have the same number of significant figures as the number with the lowest of number of significant figures in the calculations. When it comes to addition and subtraction the number with the largest uncertainty or number with least decimal places is the deciding factor.

Example 1: Significant Figures: How many significant figures are in each of the numbers?

A. 0.000256401 m
B. 17.0067 kg
C. 300 mm
D. 12,511 L
E. 2.020 cm

Answers:

A. 6 significant figures (the digits before the number do not count)
B. 6 significant figures (the middle zeros count)
C. 1, 2, or 3 significant figures (We could represent this in exponential notation. For example, 3 x 10^2 would represent 1 significant figure, 3.0 x 10^2 would represent 2 significant figures, and 3.00 x 10^2 would represent 3 significant figures.)
D. 5 significant figures
E. 4 significant figures (the final digit in a decimal place counts).

Example 2: Addition/Subtraction List the number of significant figures in the sum of the numbers.

A.	**B.**	**C.**
425.35	220. 0	121
27.6	56.4565	0.00365
+155	5.28795	+ 57.8
	+ 350.56	

Answer A: The answer is 607.95. However, this is not reported in the correct number of significant figures. The determining factor that decides the number of significant figures is the term with the fewest decimal places. In **A** this is the 155. It contains the lowest number of decimal places or the greatest amount of uncertainty. Therefore, the answer is reported to the ones place since that is the least place known for all three of the numbers being added.

Answers: **A**. 608 **B**. 632.3 **C**. 179

Example 3: Multiplication/Division Give the answers to the following questions with the correct number of significant figures.

A. 6.3 x 56.3 = **B.** 62.3/68 = **C**. 1.65 x 1.54 x .57 =

Answer A: The answer is 354.69, however this is not reported in the correct number of significant figures. The deciding factor is 6.3 simply because it has the fewest number of significant figures. Therefore, we must round our answer off to **3.5×10^2**. This answer has two significant figures just as the deciding factor.

Answers: A. 350 or 3.5×10^2 **B.** 0.92 **C.** 1.4

When making measurements we must understand that not all of what we measure is 100% correct. There is a degree of uncertainty that we must include and this is why significant figures are so crucial. For example, you make a measurement of the width of a cube to be 2.36 cm. Your lab partner makes a measurement of the same cube to be 2.38 cm. Well what is the correct answer? The correct answer can be both because both answers take into consideration the error of uncertainty. Remember when using significant figures the last digit is a guess.

Units are another crucial issue that effect what we measure. If we didn't have a standard unit of measurement what would we measure with? You could e-mail your friend in Germany and tell them that you weigh 123. However, this would tell your friend nothing because he doesn't understand what you're comparing. For all your friend knows you could be comparing it to 123 horses. Ok, well your friend might not, compare your weight to 123 horses, but he would probably compare it to kilograms because in Germany that is the standard measurement. Needless to say units are important. We must always be sure to use them and be very careful when converting.

Converting units can be tricky. When we convert, we must be sure that the units **cancel out** and that you are left with the desired unit.

Example 4: Unit conversions - You measure the length of your foot in order to buy shoes and find it to be 10 inches long. The shoe company you are ordering from wants to know your foot length in centimeters. Convert 10 inches to centimeters. (**Hint**: 1 cm = 0.39370 in. or 2.54cm = 1 in.)

$$10\ \cancel{\text{inches}} \times \frac{1\text{cm}}{0.39370\ \cancel{\text{in.}}} = 25\text{ cm}$$

(Not $10\text{ in} \times \frac{0.39379\text{ in}}{1\text{ cm}}$ Since the units would not cancel.)

OR

$$10\ \cancel{\text{inches}} \times \frac{2.54\text{ cm.}}{1\ \cancel{\text{in}}} = 25\text{ cm}$$

Always be sure that your units cancel out!

Metric- English Equivalents

1 kg = 2.2 lb	1km = 0.62 mi.	3.744 L = 1 gal
907.185 kg = 1 ton	1 m = 39.37 in.	1 L = 1.06 qt
28.3 g = 1oz.	1m = 1.0963 yd.	250 mL = 1c
453.59 g = 1 lb	1 cm = 0.39370 in.	°C = 5/9 x (°F – 32)
4.184 J = 1 cal	101,325 Pa = 1 atm	

The following tables contain some general equivalents and conversions commonly used in the chemistry laboratory.

Prefixes with SI Units

Prefix	**Symbol**	**Meaning**	Scientific Notation
giga-	G	1,000,000,000	10^9
mega-	M	1,000,000	10^6
kilo-	k	1,000	10^3
hecto-	h	100	10^2
deka-	da	10	10^1
-----	--	1	10^0
deci-	d	0.1	10^{-1}
centi-	c	0.01	10^{-2}
milli-	m	0.001	10^{-3}
micro-	µ	0.000 001	10^{-6}
nano-	n	0.000 000 001	10^{-9}
pico-	p	0.000 000 000 0001	10^{-12}

D ensity is somet hing that you will learn more in detail about in a later lab. However, we will briefly touch the subject here. **Density** is the mass of a substance divided by its volume.

$$\textbf{Density} = \frac{\textbf{Mass}}{\textbf{Volume}}$$

Density has the unit grams/milliliter (g/mL). Water is a special standard for density as well as several things in chemistry. In this experiment you will calculate the density of water and compare your value to the accepted value. Therefore, we need to know the accepted value for the density of water. It is 1.00 g/mL at 20^{o}C.

Objectives

- To learn how to use significant figures and understand their importance.
- To effectively make measurements useful in a chemistry lab.
- To convert measurements relevant to chemical lab procedures and calculations.
- To incorporate the relationship between volume and mass.

Date:________ Name:________________________ Period:________

Pre-lab Questions

1. What is the importance of significant figures in the science of chemistry?

The use of significant figures enables a chemist to decide whether certain digits in a measurement or calculation are important enough to keep or regard in the data.

2. State the number of significant figures

A. 12.01

4

B. 56

2

3. Answer the following questions in the correct number of significant figures.

A. 532.02 +50.2 +5601 =

6183

B. 56.00-5.3 =

50.7

C. 6.3 x 56.3 =

3.5×10^2

D. 58/63 = 0.92

0.92

4. Covert the following units into the given unit of measurement.

A. 153 lbs.→ kilograms

153 lbs x (1 kg/2.2 lbs) = 69.5 kg

B. 62.0 in.→ centimeters

62.0 in x 2.54 cm/ 1 in = 157 cm

5. Calculate the density of an object whose mass is 43 g and volume is 56.0 mL.

Density = Mass/Volume

Density = 43 g/56.0 mL = 0.77 g/mL

Preparation Materials

General Supplies	Laboratory Equipment
Cubic block with one open end in which interior can be measured	Balance
Metric ruler	10 mL graduated cylinder
Water	

Procedure

SAFETY FEATURES

- Measure Accurately
- Safety Goggles

MATERIALS NEEDED

- An cubic block with one open end
- metric ruler
- Balance
- 10 mL Graduated Cylinder
- Water

Part 1: Density of Water Using a Cube

1. Obtain a symmetrical cube that has an open end. Measure the height, width, and length of the cube in cm and record these measurements in the data table.
2. Use a balance to obtain the initial mass of the cube. Record it in the data table.

3. Fill the cube **completely** with water. Obtain the mass of the cube with the water. Record it in the data table. **CAUTION: Water spills can cause a slick floor and work area.**

4. Pour out the water and dry the cube. Repeat steps 1 through 3 with the same cube.

Part 2: Density of Water Using a Graduated Cylinder

5. Weigh a 10 mL graduated cylinder. Record the mass in the data table.

6. Fill the graduated cylinder with water to ~8 mL. Measure the amount of water accurately. Remember to measure at the bottom of the meniscus, the circular bottom of the water. **HINT:** Refer to Figure 2 in the Introduction.

7. Weigh the graduated cylinder and water and record this mass in the data table.

8. Empty the water and dry the cylinder. Repeat steps 5 through 7 with the same cylinder.

9. Clean up and prepare for calculations.

Date:________ Name: ___________________________ Period:_________

Lab Partner: ___________________________________

Data

* Be sure to report all data in significant figures and indicate units of measurement.

Part 1 Data Table: Density of Water Using a Cube

Container	Cube	Cube
Mass of cube	10.46 g	10.48 g
Mass of water and cube	46.95 g	47.63 g
Mass of water (Mass of water and cube minus the mass of the cube)	36.49 g	37.15 g
Dimensions (cm)	Height = 3.40 cm Width = 3.35 cm Length = 3.39 cm	Height = 3.40 cm Width = 3.35 cm Length = 3.39 cm

Part 2 Data Table: Density of Water Using a Graduated Cylinder

Container	Graduated Cylinder	Graduated Cylinder
Mass of graduated cylinder	36.48 g	36.85 g
Mass of water and graduated cylinder	44.57 g	44.55 g
Mass of water (mass of water and cylinder minus the mass of the cylinder)	7.76 g	7.70 g
Volume of water in mL	7.89 mL	7.95 mL

Calculations

Part 1: Density of Water Using a Cube

1. Calculate the density of water for both trials. Remember 1 cm^3 =1 mL. Find the average of the two densities. Be sure to use the correct number of significant figures. Show your calculations.

Trial 1

Volume of cube (L x W x H) = **3.40 cm x 3.35 cm x 3.39 cm**

Volume = 38.6 cm^3

Density = mass/volume

Density = **36.49g / 38.6 cm^3 = 0.945 g/mL**

Trial 2

Volume of cube (L x W x H) = ***3.40 cm x 3.35 cm x 3.39 cm***

Volume = 38.6 cm^3

Density = mass/volume

Density = ***37.15g / 38.6 cm^3 = 0.962 g/mL***

Average of the densities = ***(0.945 g/mL+ 0.962 g/mL)/2 = 0.954 g/mL***

2. Use the equation below to calculate the percent error. **Remember the accepted density for water is 1.00 g/mL**. Your average density of water is the experimental density. Show your calculation.

$$\text{Percent Error} = \frac{|\text{Accepted} - \text{Experimental}|}{\text{Accepted}}$$

$$\textbf{\% Error} = \frac{|\,1.00\ g/mL - 0.954\ g/mL\,|}{1.00\ g/mL} \times 100 = 4.60\ \%$$

Part 2: Density of Water Using a Graduated Cylinder

1. Calculate the density for both trials. Find the average of the two densities. Be sure to use the correct number of significant figures. Show your calculations.

Trial 1

Density = mass/volume

Density = ***7.76 g / 7.89 mL = 0.984 g/mL***

Trial 2

Density = mass/volume

Density = ***7.70 g / 7.95 mL = 0.969 g/mL***

Average of the densities = ***(0.984 g/mL+ 0.969 g/mL)/2 = 0.977 g/mL***

2. Now we are going to find the percent error. **The accepted density for water is 1.0 g/mL.** Your average density of water is the experimental density. Use these measurements to determine the percent error.

$$\text{Percent Error} = \frac{|\text{Accepted} - \text{Experimental}|}{\text{Accepted}} \times 100$$

$$\%\ \text{Error} = \frac{|1.00\text{g/mL} - 0.977\ \text{g/mL}|}{1.00\ \text{g/mL}} \times 100 = 2.30\%$$

Post Lab Questions

1. Define the difference between precision and accuracy.

   Precision is the agreement between repeated measurements under the same conditions. Accuracy is how close a measurement is to an accepted measurement.

2. Were your measurements precise in Part 1 and in Part 2? Explain your answer.

Yes, the densities from the two trials were close in number.

3. Were your measurements accurate in Part 1 and in Part 2? Explain your answer.

Yes, the densities from the two trials were close to the density of water in both parts. Also the percent error was low for both parts.

Lab

3

Properties of Matter: Density

*TN Standard **2.1**: The student will investigate the characteristics of matter.*

Have you ever noticed that when you take a deep breath before going underwater you float better?

Chemical compounds can differ in their chemical and/or physical properties. Some physical properties include melting point, boiling point, molecular mass, density, and sometimes the solubility. For water, the melting point is 0^0 C, the boiling point is 100^0 C, the molar mass is 18.0 g/mol, and the density is 1.0 g/mL at 20^0 C. The solubility of water was not mentioned because it is usually the solvent. We can use the physical properties of substances to help identify what different things are made from.

Introduction

The answer to the introductory question relates to the physical property of density. Density is the ratio of mass to volume of a substance at a certain temperature. It is determined by using the formula:

$$\textbf{Density} = \frac{\textbf{Mass}}{\textbf{Volume}}$$

Density decreases when you take a deep breath. The air is less dense than the lungs are normally which causes you to float better.

The density of a substance will change a small amount when the temperature changes. It will almost always increase when the temperature decreases. Water in the temperature range of 0^oC to 4^oC is an exception to this. Water's density is actually a little higher at 4 ºC than it is at 0ºC. This explains why ice will float in a glass of water instead of sinking to the bottom.

The physical property of density is often used by chemists as an analytical tool. A chemist may check the density of a part to help verify it is made out of the correct material. For example, if a washer of made out of aluminum was set next to washers made out of galvanized or stainless steel of the same size, they may all look so much alike that you could not tell them apart. Also different types of plastics such as polystyrene and polypropylene can also look alike. But each part made out of a particular material has specific uses. In many cases if a part is used that is made out of the wrong type of material it could fail quickly and have disastrous results.

There are several methods that are frequently used to determine if a part is made out of a material that has the correct density. In this lab you will use two different methods. In the first part of this lab you are asked to find what material a sample of metal washers are made from. To do this you will first find the mass of the washers and then their volume by how much water they displace. After you determine their mass and volume, you will calculate their density and compare this to a table that gives the density of several metals.

In the second part of this lab you will observe that different plastics have different densities through using a density gradient column. In a density gradient column the density of liquid in the column is gradually varied from one with a higher density at the bottom of the column to one with a lower density at the top of the column. After you make a density gradient column you will drop pieces of plastics of known compositions in it and observe where they are land.

Objectives

- List physical properties of substances.
- Determine the densities of items, using the density formula.
- Apply the density concept by making a density gradient column.

Date:_________ Name:__________________________ Period:_________

Pre-lab Questions

1. What can cause the density of the same substance to be different?

   Density varies with the temperature of the substance.

2. A gold colored ring was found to weigh 2.542 g at 20ºC. Its volume was found by dropping it in a graduated cylinder that was half full of water and seeing how much water was displaced. The ring displaced 0.33 mL of water. The density of 14 carat gold is 12.9 to 14.6 g/mL at 20ºC. Is the ring made out of gold? Explain your answer.

   Density = Mass/Volume = 2.542 g/ 0.33 mL = 7.7 g/mL

   Since the ring's density is less than the density for 14 carat gold it is not made out of gold.

3. List two possible hazards of the laboratory experiment.

   Glass breakage, water spills and slips

Preparation Materials

General Supplies	Laboratory Equipment	Chemicals

3/8 inch diameter (#6) washers	Balance	30% Calcium chloride solution (~8 mL per group) Prep: 30.0 g $CaCl_2$ Diluted to 100 mL with H_2O
Deionized or distilled Water	10 mL and 25 mL graduated cylinders	Ethanol (~ 8 mL per group)
Mixture of Plastics: Coffee stirrers, handle of plastic silverware, plastic cup, Styrofoam cup (approximately 1 cm in length)	50 mL beaker	

Procedure

SAFETY FEATURES

- Glassware
- Safety Goggles

MATERIALS NEEDED

Part 1: Determination of Density

1. Weigh out 10 washers and record their mass.

2. Determine the volume of the washers by their water displacement in a graduated cylinder. Fill a 10.00 mL graduated cylinder up to the 5.00 mL mark with deionized water. Tilt the graduated cylinder to provide a slope for the washers to come down. **GENTLY** drop the washers into the graduated cylinder so that no water is splashed out. Gently tap the graduated cylinder until there are no more air bubbles around the washers. Record the difference between the initial and final volume of water in the graduated cylinder to the 0.01 mL place.

- 3/8" diameter washers
- 30% Calcium Chloride Solution
- Ethanol
- Distilled water
- Balances
- 10 and 25 mL Graduated Cylinders
- 50 mL beaker
- Dropper
- Mixture of plastics: coffee stirs, plastic silverware and cup, Styrofoam cup

3. Determine the density of the washers using the formula: D = m/v.

4. Compare the density of your washers to the densities of different metals found in Table1: Densities of Some Common Metals. Determine which metal your washers are made out of and record this in the data section.

Table 1: Densities of Some Common Metals

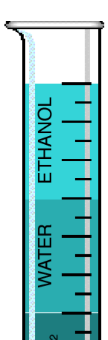
Type of Material	***Density at $20^{\circ}C$***
Aluminum	2.70 g/mL
Brass	8.40 g/mL
Copper	8.92 g/mL
Lead	11.34 g/mL
Silver	10.50 g/mL
Steel	7.80 g/mL
Zinc	7.14 g/mL

Part 2: Density of plastics

1. Obtain a piece of each of the different plastics.

2. Make a density column in the 25 mL graduated cylinder.

a. First pour 8 mL of 30% calcium chloride solution into the graduated cylinder. This solution has a density of 1.28 g/mL. **HINT:** It is easier to pour in almost the right amount using



smaller glassware such as a 50 mL beaker and then use a dropper to finish adding the right amount.

b. Measure 8 mL of distilled water into a 10 mL graduated cylinder. Slowly add the distilled water into the density column so that the water layer rests on top of the calcium nitrate layer. Water has a density of 1.00 g/mL.

c. Measure 8 mL of ethanol into a 10 mL graduated cylinder. Slowly add the ethanol into the density column so that the ethanol rests on top of the water layer. Ethanol has a density of 0.80 g/mL.

3. Gently drop each piece of the plastic in the plastic mixture one-by-one into the graduated cylinder.

4. Watch where the plastic stops in the graduated cylinder.

5. In the data section, record the level each type of plastic comes to.

HINT: Make observations quickly, but watch the objects for a couple minutes. They may make some small movement up and down the column**.**

6. Clean-up the work area and dispose of column down the sink with plenty of water.

Date:________ Name: ____________________________ Period:_________

Lab Partner: ____________________________________

Data and Calculations

Part 1.

1. Mass of sample washers: ______*4.234 g*______________

2. Initial volume of water: ________*5.00 mL*______________

3. Final volume of water: _________*6.55 mL*______________

4. Change in volume of water: ____*1.55 mL*______________

5. Density of washers:

Density = g/mL

Density = (4.234 g / 1.55 mL) = 2.73 g/mL

Part 2. Observations of plastics:

Type of Plastic	Location on Column (mL)
Styrofoam	*24.0 mL*
Styrofoam	*24.0 mL*
Plastic coffee stir	*22.0 mL*
Plastic cup	*13.5 mL*
Plastic fork	*7.5 mL*

Post Lab Questions

1. What material did you determine your washers were made from? Why did you decide your sample washers were made out of this metal?

It was determined the washers were made from aluminum. This was concluded by comparing the density calculated to the actual density of aluminum found in Table 1.

2. Why was the calcium chloride solution poured into the density column first?

The solutions should be added into the density column going in order from the densest solution to the least dense solution. Since calcium chloride was the densest solution used in this gradient column, it should be poured first. If it was not it would displace the other solutions as it gradually sinks to the bottom of the column.

3. Which type of plastic was most dense? Which type was the least dense? Explain.

The segment of the plastic fork was the densest; it sunk to the lowest level in the cylinder. The sample of Styrofoam was the least dense; it did not sink, it floated.

4. How could you use a density gradient column to help identify the type of plastic an object is made from?

A density column could be used as a tool to determine the density of an object. Comparing the known density of the substance the object was trapped can help determine the unknown identity of the plastic.

Lab 4

Separation of a Mixture

*TN Standard **4.1**: The student will investigate the characteristics of solutions.*

How does starting your morning out right relate to the separation of caffeine from a caffeine pill?

It is a lazy Saturday morning and you've just awakened to your favorite cereal "Morning Trails" and coffee. As you pour your bowl of cereal you remember the part about your favorite cereal that you don't like - the almonds. "Morning Trails' contains all of your favorite things like granola, raisins, oats, and crispy flakes; however, in your disgust for almonds, you pick them out. Then you pour the sugar and creamer into your coffee and enjoy your meal. Well how can breakfast relate to mixtures? You use different characteristics within the cereal to pick out what you don't like. You can see the appearance, the taste, the smell, and the way that is feels. In the same way we can use different characteristics to separate other mixtures.

Introduction

Often a mixture contains two or more substances that can be separated from each other by their physical properties. There are two categories into which mixtures can be divided: heterogeneous and homogeneous. A heterogeneous mixture is a mixture that contains visibly different parts, much like your cereal. A homogeneous mixture is a mixture than does not contain visibly different parts, much like your coffee after the creamer and sugar has been stirred in.

Whether heterogeneous or homogeneous, chemists often need to separate the components of a mixture. Many techniques have been developed for separating a mixture. These techniques are based on the different characteristics of the compound or element in the mixture.

One technique used for the separation of a heterogeneous mixture is filtration. This is used when a heterogeneous mixture contains both a solid and a liquid component. In filtration, a mixture is poured through a piece of filter paper seated in a funnel. The solid components of the mixture are trapped on the paper while the liquid components drain into a receiving container.

Other common techniques used for separating mixtures by physical changes include distillation, chromatography, and crystallization. Distillation takes advantage of the characteristic of compounds that have different boiling points. For example, in a mixture of two liquids with two different boiling points, the one with the lower boiling point will boil first and undergo a phase change into a gas. The gas will then travel to a separate container to cool and change form back into a liquid again. Certain methods of distillation can separate liquids in a mixture that vary only by a few degrees Celsius. Chromatography is a method that can be used to separate a mixture whose components have different properties that allow them to move at different rates along such items as a strip of paper or a column.

Crystallization is a technique that is often used to produce solids with very high purity. In this lab we will separate the components of a caffeine pill using crystallization. The caffeine pill is a mixture that is made up of about half caffeine and half binders and fillers. Binders and fillers hold the pill together to help it reach the appropriate part of your body before becoming active. Such ingredients in binders and fillers may include cellulose, dyes, starch, as well as many other ingredients to accompany the active ingredient on its way through your body.

In this lab, we will start by grinding up caffeine pills and dissolve the mixture in hot water. A large amount of caffeine will dissolve in hot water, but only a very small amount will dissolve in cold water. We can use this solubility difference to separate the impurities that do not dissolve in hot water from the caffeine by filtering out the insoluble impurities. The solution containing the caffeine will then be allowed to cool in an ice bath. Now when the solution is filtered almost all of the solid caffeine will remain in the filter paper and the impurities that are soluble in cold water will remain in the filtrate. To remove more soluble impurities, the caffeine crystals are then rinsed with only a very small amount of ice cold water. Since only a very small amount of ice cold water is used, only a really small amount of caffeine will dissolve and be lost.

Objectives

- To use physical properties to separate the components of a mixture.
- To use laboratory techniques to separate the components of a mixture.

Date:________ Name:_______________________ Period_________

Pre-lab Questions

1. Name two homogeneous mixtures and two heterogeneous mixtures that you come across daily.

   Two examples of homogeneous mixtures are coffee and air. Two examples of heterogeneous mixtures are cereal and salad.

2. How would you separate a solid mixture of sand and salt?

   Since a mixture of sand and salt is a heterogeneous mixture, filtration could be used. First water could be added and the mixture stirred. The solution could then be filtered through filter paper. The sand would remain in the filter paper and the salt would be dissolved in the water that filters through. You could then evaporate off the water.

3. Why must you rinse the final caffeine crystals remaining in the filter paper with only a very small amount of ice cold water?

   Caffeine is slightly soluble in ice cold water and so the more water you use, the more crystals you will lose.

Preparation Materials

General Supplies	Laboratory Materials
4 – 200 mg Caffeine pills (per group) (Use Vivarin© or similar generic brands such as Equate Stay Awake©. DO NOT use NoDoz© or similar products that contain benzoic acid.)	2 medium sized test tubes (18 x 150 mm)
Ice	2 600 mL beakers (400 mL beakers can be used instead)
	Hot plate
	2 Filtering funnels
	Fast filter paper (Whatman 41)
	Ring and Ring stand
	Mortar and pestle
	Stirring rod
	10 mL graduated cylinder
	Balance
	Test tube holder
	Watch glass

Procedure

SAFETY FEATURES

- Measure Accurately
- Safety Goggles

MATERIALS NEEDED

- 4-200mg Caffeine Pills (Use: Generic product Equate Stay Awake® or name brand, Vivarin®. **DO NOT** use NoDoz® or similar products that contain benzoic acid.)
- 2 medium sized test tubes
- 2-600 mL beakers
- A hot plate
- Ice
- 2 filtering funnels
- Ring and stand
- Fast filter paper
- Mortar and pestle
- Stir Rod
- Small Spatula
- 10 mL graduated cylinder
- Test tube holder
- Wash bottle
- Balance
- Watch glass

1. Fill a 600 mL beaker or the largest available beaker about half full with water.

2. Place the beaker on a hot plate. Set the heat so as to make a gently boiling hot water bath. **CAUTION: Hot plates can become very HOT very FAST.**

3. Prepare an ice water bath in another large beaker, and place a wash bottle filled with deionized water in the bath.

4. Obtain four 200 mg caffeine pills, and use a mortar and pestle to grind the pills into a fine powder.

5. Place all of the ground powder into a test tube. **HINT**: Pour all of the powder onto a piece of paper. Shape the paper so it will allow the powder to funnel into the test tube.

6. Add 10 mL of water to the powder in the test tube and stir with a stirring rod. **CAUTION: The bottom test tubes can be easily broken when too much force is applied with a stirring rod.**

7. Use a test tube holder to place the test tube in the hot water bath while it is heating to a boil. After the water reaches a **gentle boil**, let the test tube stay in the bath for an additional 5 minutes. Stir the solution in the test tube frequently while it is being heated.

8. Attach a ring to a ring stand. Place a filtering funnel in a ring that is attached to a ring stand. Fold a piece of filter paper in half and then in quarters. Place the filter paper in the funnel so that three quarters are open on one side and one quarter is on the opposite side.

9. Place a clean test tube under the funnel. It can be held in place by setting it in a clean beaker or a test tube rack. Seat the filter paper into the funnel with a very small amount of water. Discard any water that filters into the test tube.

10. Use a test tube holder to remove the test tube from the hot water bath. Pour the **hot** contents in the test tube into the filtering funnel. Allow all the liquid to filter into the test tube. **CAUTION: This solution is extremely hot. Do not touch it directly with your hands.**

Helpful HINT:
If crystals do not form within a few minutes have the student try scratching the side of the test tube in the solution. If this does not work, the water could be too acidic. Adding a few drops of NaOH can alter the pH and solve the problem.

11. Allow the hot solution to cool for a couple of minutes. Place the test tube with the **cooled filtered** liquid into the ice water bath. Continue to cool the solution until generous amounts of crystals form. This should take about 5 minutes. If crystals do not form, try scratching the sides of the test tube with a glass stirring rod until the crystals begin to form.

12. Pre-weigh a sheet of filter paper and a watch glass. Record each mass in the data section.

13. Set the pre-weighed filter paper into another **clean** filtering funnel. Place a clean test tube under the funnel. Hold the test tube in place as you did previously. Seat the filter paper into the funnel by moistening the filter paper with a very small amount of water. Discard any water that filters into the test tube.

14. Transfer the crystals into the filtering funnel. Use **1-2 mL** of **ice cold** distilled water to rinse the crystals. **DO NOT USE MORE THAN 1-2 mL TOTAL OF ICE COLD WATER TO TRANSFER AND RINSE THE CRYSTALS OR SOME OF YOUR CRYSTALS WILL DISSOLVE.**

15. Carefully remove the filter paper containing the crystals and place it on the pre-weighed watch glass. Set the crystals aside to dry. **DO NOT THROW THE CRYSTALS AWAY**. Allow the crystals to dry until the next lab period or for an extended period of time.

16. Clean up by washing any dirty glassware. Be sure that you do not throw away the crystals drying on the filter paper!

17. During the next lab period obtain the mass of the dry crystals and the filter paper. Record the mass in the data section. **DO NOT THROW THE CRYSTALS AWAY. THEY WILL BE USED FOR A LATER LAB.**

Date:________ Name: ___________________________ Period:_________

Lab Partner: ___________________________________

Data and Calculations

1. Pre-weighed filter paper mass ____*0.800 g*______
2. Pre-weighed watch glass mass ____*34.500 g*______
3. Filter paper, watch glass, and caffeine mass ____*36.280 g*______
4. Mass of caffeine obtained ____*0.900 g*_______
5. Calculate the percent recovery using the formula below:

$$\frac{\text{Actual (amount of caffeine obtained)}}{\text{Theoretical Yield (0.200g x \# of pills used)}} \times 100\% = \_\_\_\_\_\_\_\_\_\_ \%$$

$$\frac{0.900\text{ g}}{(0.200\text{g} \times 4)} \times 100\% = 113\%$$

Table 1. Observations for Lab

Procedures	Observations
1. Caffeine pill and water in hot water bath.	***Some of the solid dissolves in hot water.***
2. Caffeine water in ice	***Precipitate or crystals form in liquid.***
3. Caffeine crystals	***Crystals formed appear yellow and lumpy.***

Post Lab Questions

1. What important characteristic of caffeine did you use to separate it from the other pill components?

   Caffeine's solubility in hot water, but very little in cold water is the characteristic this separation was based on. The other components either did not dissolve in the hot water or were very soluble in the cold water.

2. Evaluate your percent recovery. Give an explanation if your recovery was more than 100% or very low? How could you improve your results?

   My percent recovery was 113% and is obviously more than 100%. Since my percent recovery is greater than 100%, the caffeine recovered is not pure. To improve my results, the caffeine could be purified again using crystallization or another technique.

Lab 5

Chemical and Physical Changes of Caffeine

*TN Standard **2.1**: The student will investigate the characteristics of matter.*

What does chemistry and superheroes have in common?

Movie producers often formulate an action movie where an everyday person will dramatically turn into a superhero. In the movies when a crisis hits, a normal guy or girl will disappear in a flash. Suddenly he or she will radically reappear disguised as a superhero who then destroys the horrifying villain and saves the day! In slightly the same way, molecules can transform their identity. Molecules can change into something that may appear physical different or chemically act very differently. Sometimes these changes occur in appearance, while other times, these changes make a different molecular substance.

Introduction

Within the science of chemistry there are many types of changes that can take place. These changes can be categorized into two primary types, physical changes and chemical changes.

A **physical change** alters the state of a substance without altering the identity of the substance. This includes dissolving, crushing, tearing, heating, melting, evaporating, sublimation or cooling.

A **chemical change** alters the identity of the molecular identity of the substance. An example is iron rusting. The iron metal changes to iron oxide by adding oxygen to its chemical formula.

This experiment involves first changing caffeine physically through the process of sublimation. **Sublimation** is when a solid goes directly into a vapor state without going through a liquid state. The next part of this experiment involves changing caffeine chemically by adding hydrochloric acid (HCl) to it. The caffeine changes into a new chemical that can be represented as caffeine·HCl. It is very soluble in water. You will then change the caffeine·HCl back to caffeine by adding the base NaOH to it.

Objectives

- To recognize the differences between chemical and physical changes.
- Determine if a change is physical or chemical.

Date:________ Name:______________________ Period________

Pre-lab Questions

1. Define physical change and list 2 examples of types of physical changes.

A physical change alters the state of a substance without altering the identity of the substance. Examples of physical changes include melting, dissolving, or subliming a substance.

2. Define chemical change and list 2 examples of chemical changes.

A chemical change involves the alteration of the molecular identity of a substance. Examples of chemical changes include rusting iron or burning wood.

3. Define sublimation.

Sublimation is when a solid goes directly into a gas without going through a liquid phase.

Preparation Materials

Laboratory Equipment	Chemicals
1 medium sized test tube (18 x 150 mm)	Caffeine crystals formed in the Separation of a Mixture Lab (or 0.8 g of purchased caffeine per group)
1 Pyrex Petri dish (per group)	5 mL 3M hydrochloric acid (HCl) per group Prep: 24.9 mL concentrated HCl (37%) per 100 mL
10 mL graduated cylinder	6 -7 mL 6M sodium hydroxide (NaOH) per group Prep: 24.0 g NaOH per 100 mL water
Hot plate	
Stirring rod	

Procedure

SAFETY FEATURES

- Safety Gloves
- Safety Goggles
- Corrosive

MATERIALS NEEDED

- 0.8 gram of caffeine $C_8H_{10}N_4O_2$
- Stir rod
- 5 mL 3M hydrochloric acid (HCl)
- 6-7 mL 6M sodium hydroxide (NaOH)
- 1 medium sized test tube
- 1 Pyrex Petri dish (both top and bottom)
- 10 mL graduated cylinder
- Hot plate

Helpful HINT: ***Teachers may need to pre-test the hot plates to determine the appropriate setting for sublimation. Hot plates will vary. Ours generally use a half way setting. The crystals that sublime to the top of the Petri dish should be white and spindly. If they are tan or brown, the hot plate setting was too high.***

1. After the caffeine crystals obtained from the "Separation of Mixtures" lab have been weighed and recorded in the previous lab data section, divide them in half. Or if they are not available, obtain ~0.8 g of caffeine. You will now use half of the mixture to show a chemical change and half of the mixture to show a physical change.

Part 1: Physical Change

2. Place half of the mixture or ~0.4 g of caffeine into the bottom part of a Petri dish and spread it out evenly.

3. Place the top on the Petri dish and place the Petri dish on a hot plate. Gradually increase the heat until the caffeine starts to sublime. **HINT:** Slowly increase the heat setting! If heating occurs too fast, the caffeine will turn brown or only burn and not sublime.

4. Allow the caffeine to sublime until either no more crystals are forming on the top of the Petri dish or the top of the Petri dish can no longer hold all of the crystals. Turn off the hot plate and **allow the Petri dish to cool** to room temperature. **CAUTION: Do not try to remove the top of the Petri dish immediately. It is extremely hot and can cause severe injury or release caffeine vapor into the air**. **HINT**: To save time you can now go to the Chemical Changes section. After you have completed these steps, come back to step 5 of Part 1.

5. After you are sure the Petri dish has cooled to room temperature, carefully remove the Petri dish lid. Record your observations in the data table.

Part 2: Chemical Change

6. Place the other ½ of the obtained caffeine (~0.4 g) into a medium to large size test tube.

7. Add 5 mL of 3M hydrochloric acid (HCl) into the test tube. **CAUTION: 3M HCl is a strong acid and can irritate the skin.**

8. Use a stirring rod to stir the contents until almost all of the crystals are dissolved. If the crystals are not dissolved, add additional 3M HCl dropwise until most of the crystals are dissolved. Record your observations in the data section.

9. Measure between 6 – 7 mL of 6M sodium hydroxide solution (NaOH) into a 10 mL graduated cylinder. Add it into the test tube.

10. Patiently stir the contents in the test tube with a stirring rod until the crystals reform. If the crystals do not reform within a few minutes, add more 6M NaOH dropwise until a large number of crystals have formed. Note all observations in the data table. **CAUTION: 6M NaOH is a strong base and can irritate the skin.**

11. To clean up, use vinegar to neutralize the crystals and solution in the test tube to a pH between 4 and 8 and then wash it down the drain.

Date:________ Name: ___________________________ Period:_________

Lab Partner: ___________________________________

Data

Procedure	Observations
Part 1: Physical Change Sublimation	*Yellow crystals from the bottom of the Petri dish reform on the top of the dish. They are now white and hanging down in a pointed formation.*
Part 2: Chemical Change Addition of HC l	*Yellow crystals dissolved into the solution.*
Addition of NaOH	*A white precipitate or crystals formed in the solution.*

Post Lab Questions

1. The following are often changes that can take place. Define each as a chemical or physical change.

   A. Igniting a Bunsen burner ***Chemical change***

   B. Boiling water ***Physical change***

   C. Making popsicles ***Physical change***

   D. Melting ice cream ***Physical change***

   E. Rusting Iron ***Chemical change***

   F. Subliming moth balls ***Physical change***

2. Describe the process of sublimation and the change the caffeine underwent.

   During sublimation the caffeine became a gas without first becoming a liquid. It collected on top of the Petri dish. Sublimation is only a physical change, because its molecular composition is not altered.

3. Why is adding an acid to the caffeine considered a chemical change?

When adding an acid to the caffeine it becomes a chemically different compound that can be called caffeine.HCl. Therefore, it is a chemical change.

4. Why did adding a NaOH solution cause the caffeine crystals to reform?

When the NaOH was added the caffeine·HCl changed back into caffeine. The caffeine is not nearly as soluble in water as caffeine·HCl so the crystals reformed.

Lab 6

The Calorimetry of Junk Food

*TN Standard **2.2**: The student will investigate the characteristics of matter and the interaction of matter and energy.*

Have you ever roasted marshmallows at a campfire?

While roasting marshmallows, have you ever caught one on fire? A burning marshmallow releases energy stored within the marshmallow as heat. This energy released is measured in the form of calories.

Introduction

You have probably heard that food has calories, but do you know what a calorie is? A calorie is the amount of energy it takes to raise the temperature of one gram of water by one degree Celsius. Heat is a form of energy, and measuring the amount of heat given off by burning food is one way to determine how much energy (calories) that food contains. The following equation is used to determine the amount of energy the food contains.

$$\mathbf{E = mc\Delta T}$$

In this equation **E** is the energy in calories, **m** is the mass in grams of what is being heated, **c** is the specific heat in cal/g · ºC, and **ΔT** is the change in temperature in ºC.

Specific heat (**c**) is a physical property of a substance. It is the amount of heat needed to raise the temperature of 1.00 g of a substance by 1ºC. The specific heat is known for many common substances such as water (1.00 cal/g · ºC), glass (0.22 cal/g · ºC), and copper (0.092 cal/g · ºC). You can see that water has a higher specific heat than glass or copper. This means it takes more energy to raise the temperature of water than it does to raise the temperature by the same amount for the same mass of glass or copper. The high specific heat of water is what makes it good to use in cold compresses.

In this lab you will explore which junk food has the most calories. You will do this by seeing which junk food sample gives off the most heat when it is burned. Since the specific heat of junk foods is not commonly known, the calories will have to be found by finding out how much the temperature of a known volume of water changes when heated by burning the food. The density of water is known to be 1.00 g/mL at room temperature. The mass of the water heated can be determined by using the formula:

Mass = Density x Volume

Food calories are reported as kilocalories. The unit Calories is capitalized to mean kilocalories. This means the calories found must be divided by 1000 to convert them to Calories.

You now have all the information needed to find the calories in everyone's favorite fluffy snack. For example, if 0.500g of a mini marshmallow is burned underneath a can containing 100.0 mL of water and the water temperature changes from 24°C to 28.5°C, the number of Calories per gram contained in the marshmallow can be calculated by the following steps:

1. Use the density of water to determine its mass:

   Mass of water = Density x Volume

   m = (1.00 g/mL)*(100.0 mL) = 100 g

2. Look up the constant for the specific heat of water:

   c = 1.00 cal/g°C

3. Determine the change in temperature:

   ΔT = (28.5°C – 24.0°C) = 4.5°C

4. Calculate the number of calories contained in the marshmallow:

   E = mcΔT

   E = (100.0 g)*(1.00 cal/g°C)*(4.5°C) = 450 cal

5. Calculate the number of calories per gram of marshmallow:

$$\text{calories/g marshmallow} = \frac{450\ \text{cal}}{0.500\ \text{g}} = 900\ \text{cal/g}$$

6. Convert to Calories (kilocalories) per gram of marshmallow:

$$\text{Calories/g marshmallow} = \frac{900\ \text{cal}}{\text{g}} \times \frac{1\ \text{Cal}}{1000\ \text{cal}} = 0.90\ \text{Calories/g}$$

The experimental method you will use gives a good comparison between two junk foods. Be sure to treat the samples the same so that you will get a more accurate comparison.

The experimental number of calories can also be compared to the actual number of calories reported by the manufacturer on the package. A significant difference will be found between the experimental Calories you find and the actual Calories reported by the manufacturer. This is because the method you will use loses much of the heat generated by the burning food to the surrounding air. The methods used by food manufacturing companies are much more sensitive and analytical.

Objectives

- Gain applicable knowledge about calories.
- Compare the calorie content of food sample.

Date:________ Name:_______________________ Period_________

Pre-lab Questions

1. In terms of food, what is a calorie?

   A calorie is a measurement of the amount of energy that a food contains.

2. A piece of chocolate chunk cookie was tested the same way as is used in this procedure. The cookie sample had a mass of 0.851 g and the temperature of a 100.0 mL of water increased by 7.87 °C. How many Calories per gram did the cookie contain?

Mass of water = Density of water x Volume of water used

= 1.00 g/mL x 100.0 mL

Mass of water = 100 g

$$E = mc\Delta T$$

$$E = (100.0\ g) * (1\ cal/g^{\circ}C) * (7.87^{\circ}C) = 787\ cal$$

$$calories/g = 787/0.851\ g = 925\ calories/g$$

$$Calories/g = \frac{926\ cal}{g} \times \frac{1\ Cal}{1000\ cal} = 0.925\ Cal/g$$

Preparation Materials

General Supplies	Laboratory Equipment

12 oz. Aluminum can with tab (1 per group)	Bunsen burner
Large non-coated Paper clips (2 per group)	Striker / match
Clay (ping pong ball size for each group)	Thermometer
Junk Foods good for this experiment: (recommended sample size: 1 cm^2) *Mini marshmallows Cheetos (or cheese puffs) Potato chips	100 mL graduated cylinder
Water	Balance
	Ring stand
	Iron ring
	Clamp
	Split Rubber Stopper

Procedure

SAFETY FEATURES

- Glassware
- Safety Goggles

MATERIALS NEEDED

- 2 Large non-coated paper clips
- Aluminum can (12 oz.)
- Clay
- Junk Foods
- Bunsen burner
- Striker or match
- Thermometer
- Split Rubber Stopper
- Ring stand
- Clamp
- Small iron ring
- Distilled water
- 50 mL beaker
- Balance
- 100 mL graduated cylinder

1. Measure between 75-100 mL of distilled water into a 100 mL graduated cylinder. Record the exact amount of water in the data table.

2. Carefully pour the measured water into an empty aluminum can that still has a tab attached.

3. Attach a small iron ring onto a ring stand. Partially unbend a large paper clip and hang it on the front of the iron ring. Hang the aluminum can by its tab on the other end of the paper clip (see Figure 1).

4. Insert a thermometer into a split rubber stopper and attach it to the ring stand by using a clamp. Lower the thermometer into the aluminum can so that it is in the water but does not touch the bottom of the can.

5. Gently straighten the outside fold of another large paper clip. This will be used as a holder for the food sample.

6. Insert the end of the paper clip that is still folded into a hunk of clay. The hunk of clay should be approximately the size of a ping pong ball.

7. Tear a junk food sample into a piece that is ~1 cm^2. Determine the mass of the piece of junk food and record it in the data table. Remember NO eating in the lab!

8. Insert the straightened end of the food holder into the sample. **HINT**: If this does not hold the sample you can make a loop at the end of the paper clip to rest the sample in.

9. Place a 50 mL beaker approximately half full of water near the apparatus to extinguish smoke after the sample has finished burning.

10. Record the initial temperature of the water inside the aluminum can in the data table.

11. Light the Bunsen burner. Adjust it until there is a fairly small, blue flame. **HINT**: To decrease heat loss, the Bunsen burner and apparatus should be located relatively close to one another. **CAUTION**: **Burns can occur with the use of flames**.

12. Hold the clay end of the sample holder and carefully bring the sample into the flame until it ignites. **HINT**: The sample should be held in the flame for a few seconds to assure the sample will burn strongly.

13. Immediately and carefully move the food approximately 1 cm away from the bottom of the aluminum can in order to minimize the amount of heat lost. If the food sample goes out before it is completely burned or is producing only a little flame and excessive smoke, quickly relight it in the Bunsen burner and place it back under the aluminum can. **CAUTION**: **Excessive smoke can result from the ignited sample and can be a respiratory irritant. If there is excessive smoke the sample should be relit immediately.**

14. Watch the thermometer while the food is burning. Record the maximum temperature that is reached.

15. Immediately after the sample has completely burned, dip it using the food holder into the beaker of water and wait for it to cool. Turn off the Bunsen burner.

16. Place the remains of the sample in the trash.

17. Wash end of the paper clip, and then dry it with a paper towel.

18. Repeat the procedure from steps 7-17 using a different type of food. (**Hint:** Use a sample size that has a similar mass to the previous sample.)

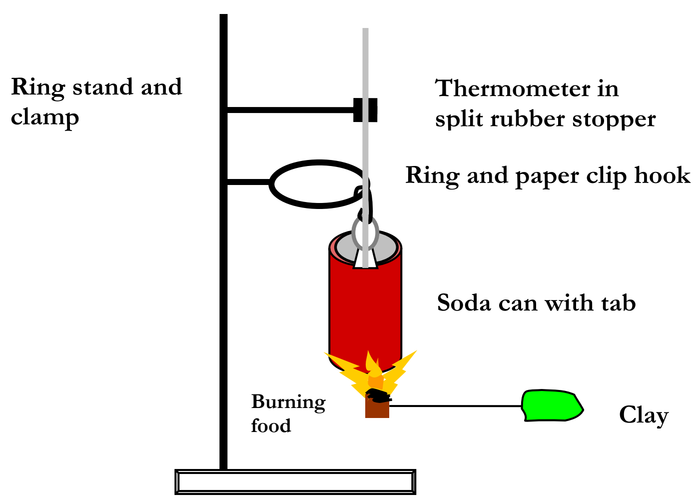


Figure 1: Apparatus for Calorimetry Experiment

Date:________ Name: ___________________________ Period:_________

Lab Partner: ___________________________________

Data

Junk Food	*Volume of Water (mL)*	*Mass of Junk Food (g)*	*Initial Temperature (˚C)*	*Maximum Temperature (˚C)*	*Δ T (˚C)*
A *Cheeto*	100.0 mL	0.520 g	30.0°C	38.9°C	8.9°C
B *Mini marshmallow*	100.0 mL	0.470 g	24.8°C	28.5°C	3.7°C

Calculations

1. Assuming that all the heat from each food went into heating the water in the aluminum can, calculate how many **Calories per gram** each sample contained. **HINT:** Food calories are measured in kilocalories!

❖ **Food Sample A**

Cheeto:

$E = mc\Delta T$

$m_{\text{(mass of water)}} = 100.0 \text{ mL} \times 1.0 \text{ g/mL} = 100 \text{ g}$

$E = (100.0 \text{ g})^* (1 \text{ cal/g}^\circ \text{C})^*(8.90^\circ \text{C}) = 890 \text{ cal}$

calories/g = 890/0.520 g = 1712 calories/g

$\text{Calories/g} = \frac{1712 \text{ cal}}{\text{g}} \times \frac{1 \text{ Cal}}{1000 \text{ cal}} = 1.71 \text{ Cal/g}$

❖ **Food Sample B**

Mini marshmallow: ***E= mcΔT***

m *(mass of water)* ***= 100.0 mL x 1.0 g/mL = 100 g***

E = (100.0 g)* (1 cal/g°C)*(3.7°C) = 370 cal

calories/g = 370/0.470 g = 787 calories/g

Calories/g = 1712 cal / g x 1 Cal / 1000 cal = 0.787 Cal/g

$$\text{Calories/g} = \frac{1712\ \text{cal}}{\text{g}} \times \frac{1\ \text{Cal}}{1000\ \text{cal}} = 0.787\ \text{Cal/g}$$

2. Using the nutrition facts and serving size reported by the manufacture, calculate how many **Calories per gram** each food contained reported by company.

- **Food Sample A**

Frito Lay Cheetos Crunchy:

(160 Calories/serving)/(28 grams/serving) = 5.7 Cal/g

- **Food Sample B**

Kraft Jet-Puffed FunMallows Miniature Marshmallows:

(100 Calories/serving)/(30 grams/serving) = 3.3 Cal/g

Post Lab Questions

1. Which food sample had the most calories per gram? Was this what you expected? Why or why not?

The Cheeto had more calories than the marshmallow. This conclusion was what I expected, because it had the largest flame.

2. How does each sample that you tested compare with the manufacturers' data?

The experimental number of calories of each sample found was much less than the actual number of calories reported.

3. If your data was not the same as that of the manufacturers, what do you think is the largest source of error?

The major factor is heat loss. Heat is lost into the air around the apparatus.

4. Was the sample you experimentally found to have the most Calories per gram the same as the sample you calculated to have the most Calories per gram based on the manufacturers' labels? If it was not, why do you think it was different?

I found the Cheetos to have a little more than twice the Calories per gram than the marshmallow. This is in fairly close to the ratio calculated from the manufacturers' labels.

Lab 7

Thermodynamics: Cold Packs vs. Hand Warmers

*TN Standard **2.2**: The student will explore the interactions of matter.*

Have you ever needed to place a cold pack on a sprained muscle?

It's the final seconds of the division championship basketball game, and your team is behind by one point. One of your team's players takes a shot and scores. The game is over, and your team won! But something is wrong. The player is sitting on the floor, and appears to be in a lot of pain. The coach quickly brings a cold pack to the player, squeezes it, and places it on the swelling ankle. The bag immediately becomes cold, but why?

Introduction

Chemical processes can give off or take in energy either through a chemical reaction or a physical change. When a reaction or physical change gives off energy and therefore increases in temperature it is called an exothermic process. To remember exothermic, think of exiting which is leaving or going out. An endothermic process does just the opposite. It takes in energy from the surroundings which causes a decrease in temperature.

Thermodynamics is the study of changes in energy in chemical processes and how temperature influences those changes. Enthalpy is the amount of energy during the reaction, or the amount of heat in a substance. The amount of heat in a substance cannot be measured directly, but changes in heat can be observed. For example, if you stick your finger in a glass of tap water while inside, then you go outside on a cold day and put your finger into a puddle of water, you can easily notice that the puddle of water is uncomfortably colder. You probably could not, however, give the exact energy of the puddle of water outside. In short, enthalpy can be compared and changes can be recorded, but it cannot be exactly determined.

Enthalpy is a quantity of energy contained in the process. The direction the energy moves determines whether the process is considered endothermic or exothermic. Energy is absorbed in an endothermic reaction or physical change. An example of an endothermic process is what occurs in an instant cold pack like the ones used to decrease the swelling caused from a sports injury. The decrease in temperature is a result of the physical process of ammonium nitrate, NH_4NO_3, dissolving in water. During the process of the ammonium nitrate dissolving, heat energy is taken in. This causes a decrease in temperature.

In contrast, energy is released in an exothermic process. An example of an exothermic process is the reaction that occurs in hand warmers. The increase in temperature is the result of the chemical reaction of iron rusting.

$$4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)$$

Iron usually rusts fairly slowly so the heat given off in the reaction is not noticed. Common table salt is added to the iron filings in hand warmers as a catalyst to speed up the rate of the reaction. Hand warmers also have a permeable plastic bag that regulates the flow of air into the bag. This allows just the right amount of oxygen in so that the desired temperature is maintained for a long period of time. The other ingredients that are found in hand warmers include a cellulose filler, carbon to disperse the heat, and vermiculite to insulate and retain the heat.

In this lab you will observe the temperature changes for cold packs and hand warmers. Since temperature is defined as the average kinetic energy of the molecules, changes in temperature indicate changes in energy. You will use simply a Styrofoam cup as a calorimeter to capture the energy. The customary lid will not be placed on the cup since ample oxygen from the air is needed for the hand warmer ingredients to react within a reasonable amount of time.

Objectives

- Understand the difference between endothermic and exothermic processes.
- Gain a knowledge of enthalpy.

Date:________ Name:______________________ Period:_________

Pre-lab Questions

1. Define enthalpy.

Enthalpy is the amount of energy during the reaction or the amount of heat in a substance.

2. What is the relationship between the enthalpy of a reaction and its classification as endothermic or exothermic?

The enthalpy of the reaction is the quantity of energy involved in the reaction, but endothermic or exothermic describes the direction of the energy, whether the energy is being "given off" as in exothermic reactions or "taken in" as in endothermic reactions.

3. In instant hot compresses calcium chloride dissolves in water and the temperature of the mixture increase. Is this an endothermic or exothermic process?

It is an exothermic process.

Preparation Materials

General Supplies	**Laboratory Equipment**
Calorimeter (one small Styrofoam cup)	Thermometers (Digital work best)
Cold pack (1/4 of the NH_4NO_3 contents or 10.0 g per group)	Stirring rod
Hand warmer (1/3 of inner contents or 6.0 g per group)	10 mL graduated cylinder
Water (10 mL per group)	Spatula
	Balance Stop Watch

Procedure

SAFETY FEATURES

Safety Goggles

MATERIALS NEEDED

- Calorimeter (one Styrofoam cup)
- Thermometer (digital is preferred)
- Stirring rod
- 10 mL graduated cylinder
- Distilled water
- Balances
- Spatula
- 1/3 contents of a hand warmer
- ¼ contents of a cold pack
- Stop Watch

Helpful HINT: ***Any liquid contained in the inner contents of the cold pack should be decanted from the solid and discarded.***

Part 1. Cold Pack Experiment

1. Measure 10 mL of distilled water into a 10 mL graduated cylinder.

2. Place ~1/4 (or ~10.0 g) of the ammonium nitrate crystals found in the solid inner contents of a cold pack into a Styrofoam cup. The Styrofoam cup is used as a simple calorimeter.

3. Place a thermometer and a stirring rod into the calorimeter (Styrofoam cup). **CAUTION: Hold or secure the calorimeter AND the thermometer to prevent breakage.**

4. Pour the 10 mL of water into the calorimeter containing the ammonium nitrate, (NH_4NO_3).

5. Immediately record the temperature and the time.

6. Quickly begin stirring the contents in the calorimeter.

7. Continue stirring and record the temperature at thirty second intervals in the corresponding Data Table. You will need to stir the reaction the entire time you are recording data.

8. Collect data for at least five minutes and until after the temperature reaches its minimum and then begins to rise. This should take approximately 5 to 7 minutes.

9. Record the overall minimum temperature in the appropriate place on the Data Table.

Part 2. Hand Warmer Experiment

1. Wash and dry the Styrofoam cup and thermometer. **HINT:** Remember to rinse them with distilled water before drying.

2. Carefully place and hold the thermometer in the cup.

3. Cut open the inner package of a hand warmer and quickly transfer ~ 1/3 of it into the calorimeter. Immediately record the initial temperature of

Helpful HINT: ***The hand warmers are activated immediately after they are opened. This means you must quickly transfer ~ 1/3 of the contents into 3***

the contents and begin timing the reaction. **HINT:** Data collection should start quickly after the package is opened because the reaction will be activated as soon as it is exposed to air.

4. Quickly insert the stirring rod into the cup and begin stirring the contents in the calorimeter.

5. Continue stirring and record the temperature at thirty second intervals in the Data Table. You will need to stir the reaction the entire time you are recording data.

6. Let the reaction continue for at least five minutes and until the temperature has reached its maximum and then fallen a few degrees. This should take approximately 5 to 7 minutes.

7. Record the overall maximum temperature in the appropriate place on the Data Table.

Name:______________________________Date:__________Period:_____

Lab Partner:_________________________

Data

Part 1. Cold Pack Experiment

Table 1. Cold Pack

Minimum Temperature (^{0}C) : -1.8 °C

Time (sec)	Temp. (^{0}C)	Time (sec)	Temp. in (^{0}C)
Initial	22.3	240	1.6
30	-0.7	270	1.7
60	-1.6	* 300	2.0
90	-1.8	330	
120	-0.8	360	
150	0.4	390	
180	0.9	420	
210	1.4	450	

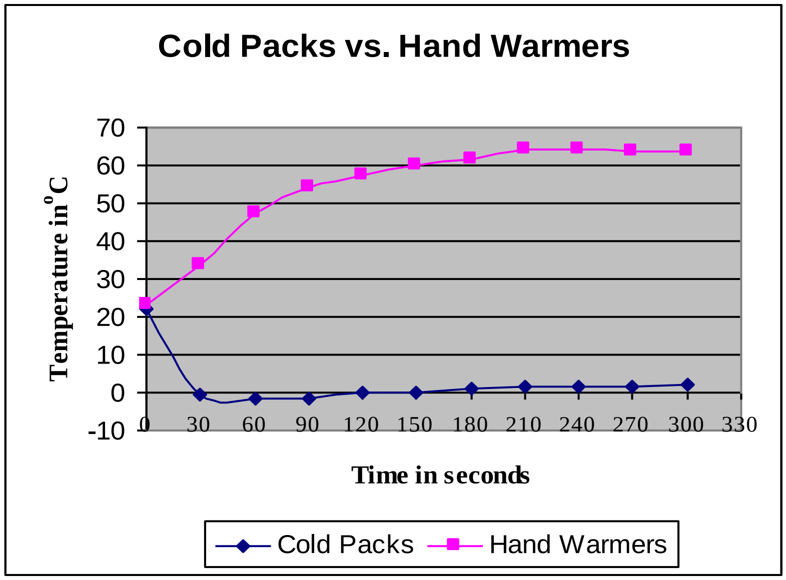
Part 2. Hand Warmer Experiment

Table 2. Hand Warmer

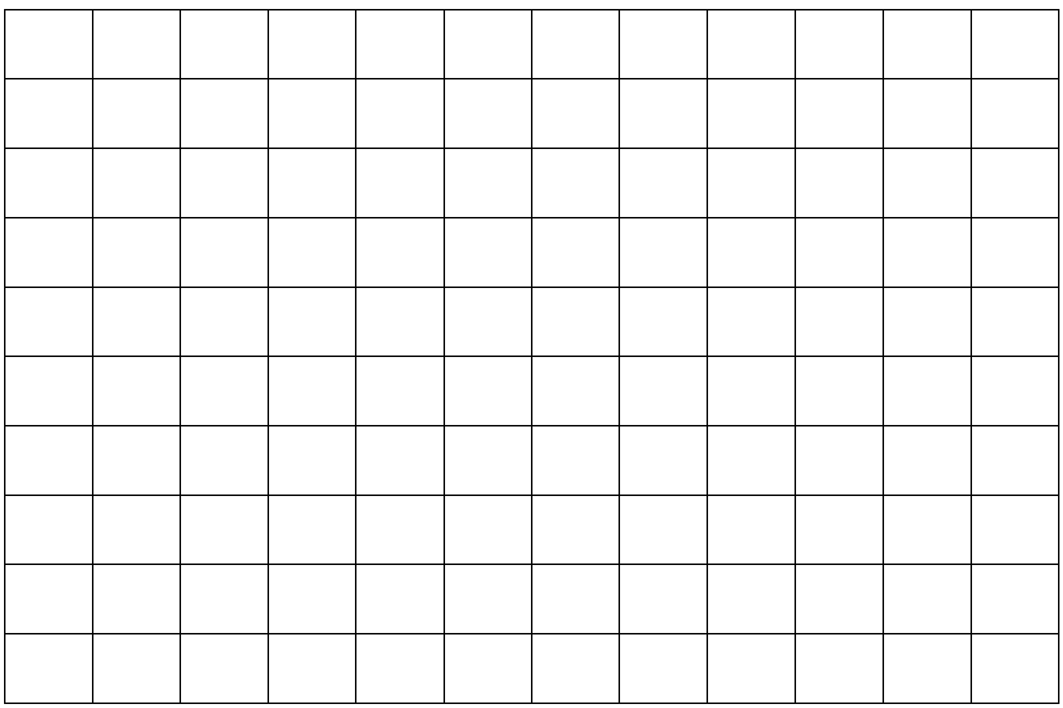
Maximum Temperature (^{0}C) :__64.2 °C___

Time (sec)	Temp. (^{0}C)	Time (sec)	Temp. (^{0}C)
Initial	23.0	240	64.0
30	33.8	270	63.8
60	47.2	* 300	63.6
90	54.3	330	
120	57.5	360	
150	59.8	390	
180	61.76	420	
210	64.2	450	

- Graph the data from Data Tables 1 and 2 as two separate lines on the same graph. Be sure to title your graph and label each axis. Include the units in the axis labels. An example is shown below.



Use the following grid to graph your data:



Calculations

Calculate the overall temperature change for cold and hot pack substance. **HINT**: This is the difference in the maximum temperature and minimum temperature of each.

Cold pack: ***22.3 – (-1.8) = 24.1°C***

Hand warmer: ***64.2 – 23.0 = 41.2°C***

Post Lab Questions

1. Which pack works by an exothermic process? Use experimental data to support your answer.

   The hand warmer works by an exothermic process because heat was given off by the reaction. This caused the temperature to rise significantly.

2. Which pack works by an endothermic process? Use experimental data to support your answer.

   The cold pack works by an endothermic process because the minimum temperature was much less than the initial temperature. This means heat was taken in by the cold pack.

3. Which pack had the greatest change in enthalpy? How do you know?

   The hand warmers had the greatest change in enthalpy. This was concluded since the greatest temperature change was found when the hand warmer ingredients were stirred.

Lab
8

Electron Configuration: Chemistry of Fireworks

*TN Standard **1.5**: Relate the spectral lines of an atom's emission spectrum to the transition of electrons between different energy levels within an atom.*

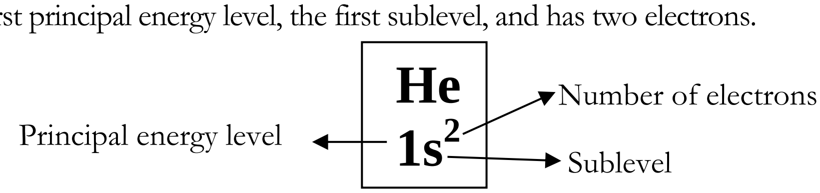
Have you ever wondered why fireworks burst into different colors?

Atoms radiate light when a change in energy state of the electrons occurs. Various salt compounds can be added to produce a specific color. It is on this concept that fireworks and their awesome colors and brilliance are based. The main ingredients of fireworks are potassium chlorate or perchlorate, charcoal, and sulfur. With each different ingredient or compound, a different color can be produced to give a beautiful multi-colored light show.

Introduction

The energy of atoms is quantized, meaning there are specific energy levels. This is similar to climbing up a ladder. As you climb up the rungs, you cannot stop between rungs. Your foot must always rest on a rung. In a similar fashion, the changes in energy of electrons must be whole energy level changes, not between two levels.

Picture your foot is on the first rung of a ladder. This is like an electron being in the first principle energy level (n=1). Scientists invented a naming system for describing the location of a specific electron in an atom. Each principle energy level has a sublevel or sublevels (s, p, d, or f). The system for naming is called the electron configuration. For example, the electron configuration for helium is $1s^2$ because it is in the first principal energy level, the first sublevel, and has two electrons.



Electrons fill the lowest energy levels first. An atom in an unexcited state is said to be in the ground state. The first sublevel in any orbital is the “s” sublevel. It can hold up to 2 electrons. The next sublevel filled is the “p” sublevel. It is able to hold up to 6 electrons. Continuing the “d” sublevel can hold up to 10 electrons and “f” sublevel can hold up to 14 electrons. The order in which the sublevels are filled is shown in the following figure.

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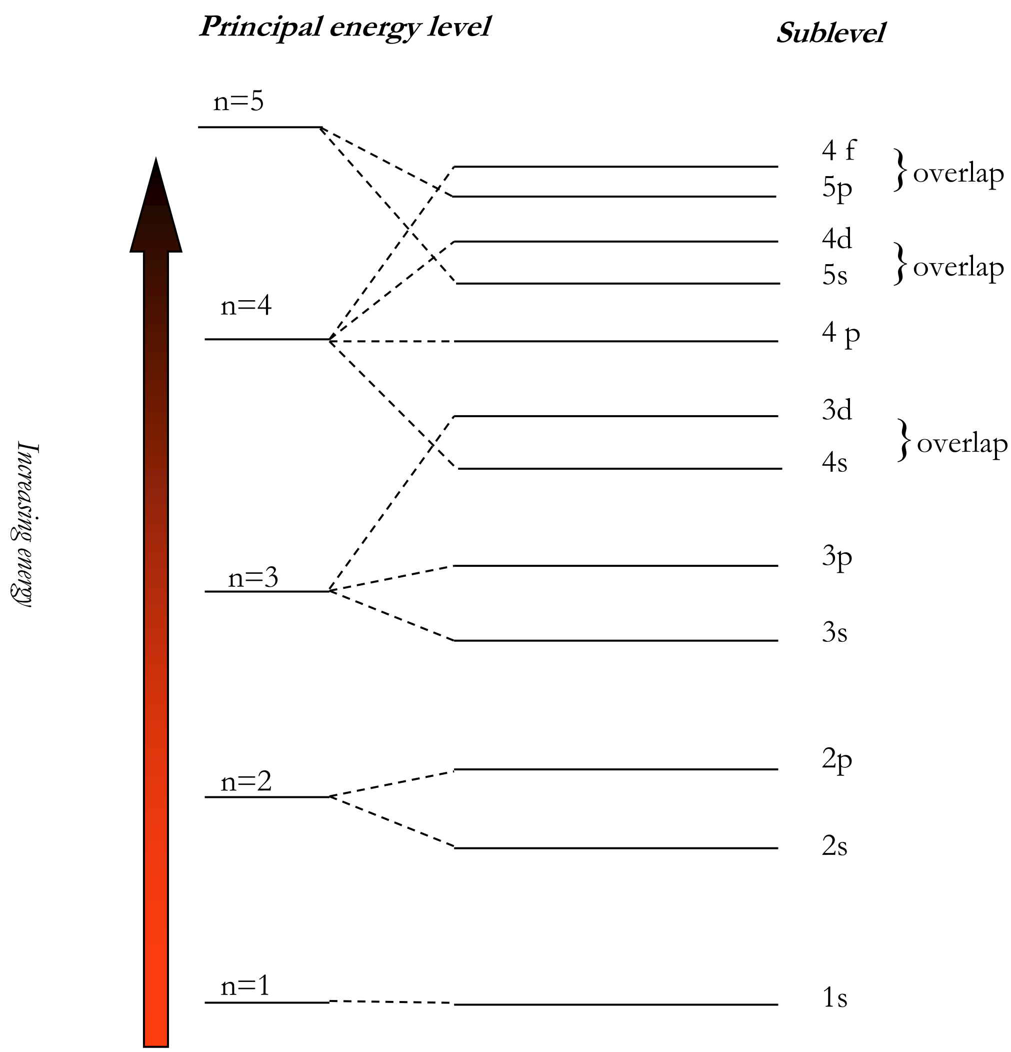


When an atom receives additional energy from an outside source that is equal to the amount of energy a single electron needs to reach a higher energy orbital or level, the electron will absorb that energy and go to the higher energy level. It will not stay at this higher energy level. Instead it will emit or release the exact amount of energy it absorbed as a wavelength of light and return to its previous lower energy level. The outside energy source is usually some form of heat or light.

In this experiment the flame from a Bunsen burner is the outside energy source. The burner emits a broad range of energy, but the electrons of the atom being heated will only absorb specific amounts of energy.

Electrons of different elements absorb and emit different amounts of energy. This is how fireworks produce different colors of light. The color of the light is related to the wavelength of the light emitted. Each color has its own wavelength. A shorter wavelength will emit higher energy than a longer wavelength. The following table summarizes several color shades and their wavelength ranges.

Color Shades	**Wavelength Range** (nm)
Red	650 – 750
Orange	595 – 650
Yellow	580 – 595
Yellow - Green	560 – 580
Green	500 -560
Green - Blue	490 – 500
Blue – Green	480 – 490
Blue	435 – 480
Violet	400 – 435

Objectives

- Obtain a general knowledge of what produces different colors in fireworks.
- Observe energy emitted from different energy levels when salt compounds are ignited.

Date:________ Name:______________________ Period_________

Pre-lab Questions

1. What is an electron configuration?

   It is the system used to name each principle level and its sublevels.

2. How is light emitted or produced?

   Light is produced when an atom changes from a higher to a lower energy level.

3. Why is it extremely important to be attentive when holding the wire wand in the flame?

   The wire will become hot and REMAIN hot for several minutes after removing it from the flame. It could cause burns to skin and clothing.

Preparation Materials

General Supplies	Laboratory Equipment	Chemicals
Wire wand (nichrome wire with loop)	7 small test tubes or vials (13 x 100 mm) (per group)	5% LiCl Lithium chloride Prep: 5 g per 100 mL distilled water
Distilled Water	Bunsen burner (If a Bunsen burner is not available, this lab can be done as a demonstration with a propane torch or butane lighter as the burner.)	5% NaCl Sodium chloride Prep: 5 g per 100 mL distilled water
Wax pencil / Sharpie	Striker / match	5% KCl Potassium chloride Prep: 5 g per 100 mL distilled water
	10 mL graduated cylinder	5% $CaCl_2$ Calcium chloride Prep: 5 g per 100 mL distilled water
	Spatula	5% $SrCl_2$ Strontium chloride Prep: 5 g per 100 mL distilled water (For an even greener experiment this can be omitted.)
		Activated charcoal (C)
		1*M* HCl Hydrochloric acid Prep: 8.3 mL Concentrated HCl (37%) per 100 mL distilled water

Procedure

SAFETY FEATURES

- Safety Goggles
- Glassware

MATERIALS NEEDED

- Seven small test tubes or vials
- Wax pencil or Sharpie
- 5% LiCl Lithium Chloride
- 5% NaCl Sodium chloride
- 5% KCl Potassium chloride
- 5% $CaCl_2$ Calcium chloride
- 5% $SrCl_2$ Strontium chloride
- Activated charcoal (C)
- 10 mL Graduated cylinder
- 1M HCl Hydrochloric Acid
- Bunsen burner
- Striker or lighter
- Nichrome Wire wand

1. Label 5 small test tubes LiCl, NaCl, KCl, $CaCl_2$, $SrCl_2$. Pour a small amount (about 20 drops) of the corresponding 5% solutions into each test tube.

2. Label another test tube C and place a small amount of activated charcoal into it.

3. Label yet another small test tube HCl. Pour approximately 5 mL of 1 *M* HCl into it. This solution is for cleaning the nichrome wire. **CAUTION: HCl is a strong acid and can irritate the skin.**

4. If it has not already been done for you, make a small (2 mm in diameter) loop in the end of the nichrome wire by gently bending it.

5. Carefully light the Bunsen burner and if necessary, adjust the flame. **CAUTION: A Bunsen burner can cause fire or burns to skin and/or clothing.**

6. Hold the nichrome wire at the very end of the non-looped end in order to avoid burns. Clean the end of the nichrome wire by dipping it into the hydrochloric acid, then heating it in the flame. **CAUTION: The loop will remain extremely hot for several minutes following flaming. Do not touch the loop!**

7. REPEAT this cleaning until the flame is faintly yellow when the wire is introduced. Other colors indicate impurities. If the hydrochloric acid solution in the beaker becomes tinted, dispose of it properly by neutralizing it with baking soda to a pH between 4-8 and rinsing it down the sink, and then replace it with clear 1 *M* hydrochloric acid solution.

8. After the nichrome wire loop is clean and has briefly cooled, dip it into the LiCl solution.

9. Bring the loop into the flame. Record observations about what happened, especially any color changes.

10. Clean the loop as you did in steps 6 and 7 until the color change is no longer seen.

11. Repeat steps 8 through 10 for all solutions and the activated charcoal.

12. Clean the nichrome wire as before when you are finished. Dispose of 1M HCl by neutralizing it with baking soda to a pH between 4-8 and rinse it down the sink. Be sure to wash your hands before leaving the lab since chemicals can be harmful if ingested.

Date:________ Name: ___________________________ Period:_________

Lab Partner: ___________________________________

Data

Results of Firework Material Ignition	
Substance	**Observations**
Lithium chloride	*Bright red flame*
Sodium chloride	*Large bright yellow flame*
Potassium chloride	*Light purple flame*
Calcium chloride	*Orange flame*
Strontium chloride	*Bright red flame*
Activated charcoal	*Bright orange flame with sparks*

Post Lab Questions

1. Write out the electron configurations of each of the metals of the salt compounds used and of carbon. Potassium is already done as an example for you. **HINT:** The periodic table is very helpful and can be used as guide.

Element	Electron Configuration
K	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Li	$1s^2\ 2s^1$
Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Ca	$1s^2\ 2s^2\ 2p^6\ 3s^2 3p^6\ 4s^2$
Sr	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$
C	$1s^2\ 2s^2\ 2p^2$

2. What is the approximate wavelength of light emitted of each of the salts?

Salt	Color	Wavelength
LiCl	Reddish-violet	425 nm
NaCl	Yellow	590 nm
KCl	Reddish-violet	425 nm
$CaCl_2$	Orange	625 nm
$SrCl_2$	Orange	625 nm
C	Yellow	590 nm

3. Why does a salt compound give off light or a colored flame?

It is returning to its previous lower energy level by releasing the exact amount of energy it has absorbed from the flame as a particular wavelength of light.

4. Did lithium and strontium give off similar colors? Why or why not?

The two emitted a similar colored flame because they have electrons that absorb and emit energy at wavelengths that are close to each other.

Lab 9

Finding a Laser Pointer's Wavelength of Light

*TN Standard **1.5**: The student will relate the spectral lines of an atom's emission spectrum to the transition of electrons between different energy levels within an atom.*

Why do we see a beautiful rainbow paint the sky just after it rains?

The gentle rain that has been falling gradually comes to an end. The gray sky is replaced with a light blue color and a few wispy white clouds. Off in the distance a beautiful rainbow can be seen. People have wondered throughout history what causes the colors of the rainbow and why are the colors always in a certain order?

Introduction

The first known scientist to uncover the true nature of light was Sir Isaac Newton. In 1666 he closed his window shutters until there was only a small hole that he could direct in such a way that a beam of sunlight would shine through a glass prism. When he did this, he observed the white sunlight became a band of the colors of a rainbow. From this, Newton theorized that white light consists of all the colors of the rainbow.

Although we see sunlight (or white light) as a single color, it actually is not. Instead it is composed of a broad **range** of colors. A rainbow is seen when the visible light from the sun is spread out to show off all of the beautiful colors that make up the sunlight.

Visible light is commonly looked at as a wave phenomenon. It is characterized by a wavelength. **Wavelength** is defined as the distance between adjacent peaks (or troughs) as shown in Figure 1. It is given as a length such as meters, centimeters or nanometers (10^{-9} meters).

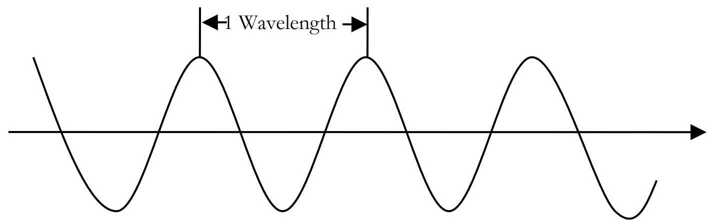


Figure 1: Wavelength Illustration

Wavelengths

An obvious difference between certain compounds is the color they emit. The human eye analyzes light reflected from the surface of a solid or passing through a liquid. Visible wavelengths cover a distance range from approximately 400 to 800 nm. The longest visible wavelength is red and the shortest is violet. The wavelengths of what we see as particular colors in the visible portion of the spectrum are listed in Table 1.

Table 1: Table of Light Colors

Wavelength (nm)	Light Color
400 - 435	Violet
435 - 480	Blue
480 - 490	Blue-green
490 - 500	Green-blue
500 - 560	Green
560 - 580	Yellow-green
580 - 595	Yellow
595 - 650	Orange
650 - 750	Red

Electromagnetic Spectrum

The visible spectrum is just a small part of the total radiation spectrum. The sun not only provides us with the radiation we can see known as visible light, but also with other forms of electromagnetic radiation that we cannot see such as infrared light, ultraviolet light, and cosmic rays. This radiation can be detected by instruments designed to do so. This electromagnetic spectrum ranges from very short wavelengths (including gamma and x-rays) to very long wavelengths (including microwaves and broadcast radio waves). Figure 2 displays many of the important regions of this spectrum. It can be seen that the longer the wavelength, the lower the energy.

←Gamma↓ X-Rays↓ UV ↓V↓ IR ↓ Microwave ↓ Radio →

Wavelength (m) |___|___|___|___|___|___|___|___|___|___|___|___|___|___|

10^{-13} 10^{-11} 10^{-9} 10^{-7} 10^{-5} 10^{-3} 10^{-1} 10

Energy (kcal) |___|___|___|___|___|___|___|___|___|___|___|___|___|___|

10^{8} 10^{6} 10^{4} 10^{2} 10^{0} 10^{-2} 10^{-4} 10^{-6}

Figure 2: The Electromagnetic Spectrum

The wavelengths of light can be spread out by using a prism or a **diffraction grating**. A diffraction grating has thousands of parallel lines etched into it. The closer the grooves are etched, the further the spectrum will spread out. A CD is like a diffraction grating in that it has many parallel lines etched into it. This is why a CD creates a spectrum when light is reflected off of it.

If when the light is shone through a diffraction grating the colors blend one into the other the spectrum is called a **continuous spectrum**. The spectrum seen as a rainbow is a continuous spectrum. If the spectrum has instead distinct isolated lines of color it is called a **bright line spectrum.** The light that shines out of a laser pointer is an example of this since it has only one distinct color.

Not only can a diffraction grating be used to see the spectrum on a particular light, it can also be used to calculate the wavelength of a light. To do this light is shone directly through a diffraction grating whose distance between lines is known. The distance between the center spot and the first side spot (**X**) and the distance between the diffraction grating and the first side spot (**L**) are measured. The modified "grating equation" shown below can then be used to calculate the light's wavelength(s):

$$n\lambda = d\ X/L$$

where, λ = wavelength
d = the distance between the lines on the diffraction grating
n = the order of the image (most of the time n = 1 since the measuring is made to the first side spot)
X = the distance between the center spot and first side spot
L = the distance between the diffraction grating and the first side spot

When you use this equation it is easier to keep all calculations in millimeters and convert the answer to nanometers (10^{-9} meters) in the end.

In this experiment you will actually use the modified grating equation to calculate the wavelength of light a laser pointer is emitting. An example of how to use this is given below:

Example 1: Calculating the Wavelength of a Laser Pointer's Light

You shine a laser pointer light through a diffraction grating that has 500 lines/mm and find the distance between the center spot and first side spot is 40 mm and the distance between the diffraction grating and the first side spot is 126 mm. To find the laser pointer light's wavelength you need to first calculate **d**.

d = 1/500 mm = 0.002 mm

Next solve the modified grating equation for λ.

$$n\lambda = d\ X/L$$

$$1\lambda = [(0.00200\text{ mm})*(40.0\,\cancel{\text{mm}}) / (126.0\,\cancel{\text{mm}})]$$

$$\lambda = 0.000635\text{ mm}$$

Now convert mm to nm.

$$\lambda = 0.000635\,\cancel{\text{mm}}\ (1\text{ nm}/10^{-6}\,\cancel{\text{mm}})$$
$$\lambda = 635\text{ nm}$$

The red light emitted by laser pointers consists of waves that are shorter than a millionth of a meter. Even though the red color of the light emitted by different laser pointers may all seem the same, they are not necessarily the same. Laser pointers have been made that emit red light at different wavelengths between 633 to 690 nm.

Objectives

- To learn the difference between **continuous** and **line spectra**.
- To determine the wavelength measurement of a red laser light.

Date:________ Name:______________________ Period_________

Pre-lab Questions

1. What is a diffraction grating?

   It is an apparatus that has thousands of parallel lines etched into it. This causes light to spread out so that the colors that make up the light can be seen.

2. You see a girl with a beautiful diamond ring that has the sun shining brightly on her finger. If purple and green are reflecting onto a piece of paper near her hand, what wavelengths of light do you suspect are reflecting off of her diamond?

   Since purple is being reflected I would suspect a wavelength of 400-435 nm and due to green I would suspect a wavelength of 500-560 nm.

3. A laser light was shone through a diffraction grating whose lines were 1/1000 mm apart. The distance was measured between the center spot and the first side spot and found to be 99 mm. The distance from the diffraction grating to the first side spot was found to be 154 mm. Calculate the wavelength of light in nm that the laser pointer was emitting.

   d = 1/1000 mm = 0.001 mm

   1λ= [(0.001 mm)*(99 mm) / (154 mm)]

   λ = 0.000643 mm

   λ= 0.000643 mm (1 nm/10^{-6} mm) = 643 nm

Preparation Materials

General Supplies	Laboratory Equipment
Laser pointer	500 and 1000 lines/mm Diffraction gratings (these can be purchased from Rainbow Symphony, Inc., (818) 708-8400 www.rainbowsymphony.com)
String	Ring and ring stand
Rulers	Clamp
White paper	
Flashlight	

Procedure

SAFETY FEATURES

Safety Goggles

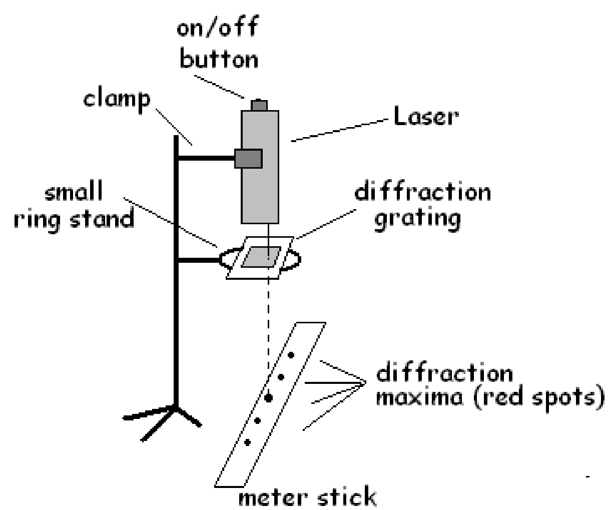
MATERIALS NEEDED

- Laser pointer
- Diffraction grating
- Ring and ring stand
- Clamp
- String
- Ruler
- White paper
- Flashlight

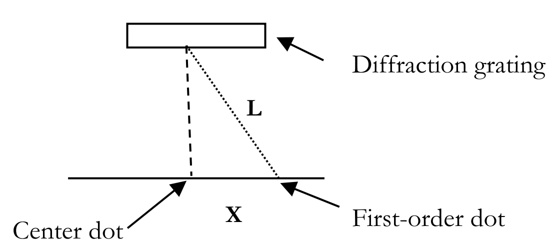
Part 1: Measuring the Wavelength of Laser Light

1. Set up an apparatus as shown in the following figure. This is done by first placing a piece of white paper on the base of a ring stand. Adjust a small ring on the ring stand so that it is approximately 200 mm from the paper.

2. Place a 1000 line/mm diffraction grating on top of the ring. Clamp a laser pointer so it is located just above the diffraction grating. Its red beam will point down through the diffraction grating onto the paper. **CAUTION: Do point the laser pointer into anyone's eyes.**



3. Place the diffraction grating slide so that the laser light will shine through the 1000 lines/mm diffraction grating. Turn on the laser light and record your observations. You should see a central bright red dot and two somewhat dimmed red dots to both sides of this dot (some distance away). These dots are the first-order diffraction image of the laser beam. You may also see two even dimmer red dots even further away. These dots are the second-order diffraction image of the laser beam. Turn off the laser pointer.



4. Replace the 1000 lines/mm diffraction grating slide with the 500 lines/mm slide. Turn on the laser light and make observations on how the projected image has changed. Record your observations in the data section.

5. When shining the laser pointer through the 500 lines/mm diffraction grating, mark the location of the central dot and the first order dot with a small "x" on the white paper. Measure the distance between the central dot and the first order dot. This is distance **X.** Record this measurement in millimeters in the data section. **HINT**: Make sure the measurement is about the same to both of the first order dots. If it is not, add the two **X** measurements and divide by two. Use this value for **X**.

6. Measure **L** by stretching a string from the first-order dot to where the laser light passes through the diffraction grating. Use a ruler to determine the length of the string. **HINT**: Tie a knot in one end of the string to help hold it down with your finger. Also, make sure the measurement is about the same to both of the first order dots. If it is not, add the two **L** measurements and divide by two. Use this value for **L**.

Part 2: Observing the Wavelengths of a Flashlight

1. Replace the laser light with a flashlight. **HINT**: In the set-up a clamp may not be needed to secure the flashlight, it probably can be balanced on its own.

2. Record your observations. Turn off the flashlight.

Date:________ Name:___________________________ Period:_________

Lab Partner: ___________________________________

Data

Part 1. Laser pointer

Observations for the laser pointer shone through 1000 and 500 lines/mm diffraction gratings:

The number of dots increased and became closer together with the 500 lines/mm diffraction grating.

Measurement for **X** using the 500 lines/mm diffraction grating: _***77.0 mm***___

Measurement for **L** using the 500 lines/mm diffraction grating: _***231.0 mm***___

Part 2. Flashlight

Observations for the flashlight:

A continuous spectrum of colors is seen starting with violet and ending with red. Several dots appear. They are larger in size and reflect many colors in a single dot.

Calculations

Part 2. Laser pointer

Calculate the wavelength of the **laser light** using the modified "grating equation":

$$n\lambda = d\,X/L$$

where

λ = wavelength
d = the distance between the lines on the diffraction grating
n = the order of the image (in this case n = 1)
X = the distance between the center spot and side spot
L = the distance between the diffraction grating and the side spot

Remember to keep all calculations in millimeters until the end and then convert the answer into nanometers (1 nm/10^{-6} mm). Show your calculations.

d = 1/500 mm

1λ= [(0.002 mm)*(77.0 mm) / (231.0 mm)]

λ = 0.000667 mm

λ= 0.000667 mm (1 nm/10^{-6} mm) = 667 nm

Post Lab Questions

1. Is the value you obtained for the wavelength of the laser light consistent with the range of wavelengths for red light? Is this what you expected?

Yes, 667 nm was consistent with the range of wavelengths for red from 650-750 nm. It was expected to be similar to red, because the laser pointer shines light that is red in color.

2. Why was the pattern seen when shining the flashlight through the diffraction grating different than when shining a laser pointer light through it?

The light from the flashlight is made up of light from many different wavelengths, whereas the light from a laser pointer is made up of a single wavelength of light.

Lab

10

Magnets, Marshmallows, and Molecular Models

*TN Standard **2.2**: The student will explore the interactions of matter and energy.*

Why can't you play basketball with a football?

The obvious answer is because a football isn't the right shape. A football can't be dribbled, and would be very difficult to shoot. On the other hand a fairly large spherical ball can be easily dribbled and is needed to successfully play basketball. The shape of a molecule is also extremely important in how it can be used. Just like sports need objects of a particular shape, certain reactions or applications need molecules of a particular shape.

Introduction

The shapes of the different objects used in sports can be described in a few words. For example, an arrow can be described as straight or linear, a hockey puck as a flat disc, a football as oblong, and a basketball as spherical. Molecules also have definite shapes that can be described in a few words. A few of the words used to describe molecular shapes are linear, bent, trigonal planar, pyramidal, and tetrahedral. Most of these words are very similar to common words.

But how do you know the shape of a particular molecule since you can't see it? Fortunately a model has been developed that helps us predict the shapes of molecules. It is called the **VSEPR** model. VSEPR stands for valence-shell electron-pair repulsion.

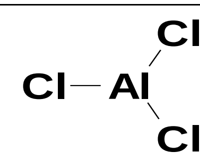
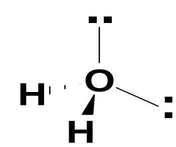
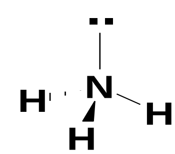
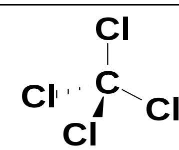
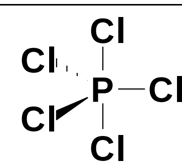
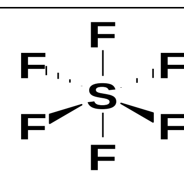
When you play a team sport one of the first things you need to decide is who is going to be the captain. When using the VSEPR model a central atom has to first be decided. A central atom is usually the one with the most bonds to it. A single bond counts as one bond, a double bond counts as two bonds, etc. For example, the carbon atom in methane, CH_4, is the central atom since it has 4 bonds and each hydrogen atom has just one bond. In water, H_2O, the oxygen has two bonds so it is chosen as the central atom. In carbon dioxide, CO_2, the carbon has 2 double bonds or 4 total bonds so it is chosen as the central atom. Sometimes there are two different atoms that tie in having the most bonds. When this happens, you chose one of these to be the central atom. For example, the oxygen atoms in hydrogen peroxide, HOOH, each have 2 bonds to them whereas the hydrogen atoms each have only one bond. In this case, one of the oxygen atoms is chosen.

After the central atom is decided, the number of regions of electron density around it is determined. A region of electron density is a negatively charged region. The number of regions can be different than the number of bonds. A single, double, or triple bond **or** a pair of unshared electrons counts as one region of electron density. The carbon in methane, CH_4, has four single bonds so it has four regions of electron density. However in H_2O, the oxygen in water has two single bonds and two pairs of unshared electrons. This means it has four regions of electron density.

When the same poles of magnets are brought together they repel each other. The VSEPR model is based on a similar concept. The valence electrons are involved in bonding or are unshared. They create a **negative region that repels the other negative regions**. In the VSEPR model, the regions of electron density are placed the same distance from the central atom, but as far away from each other as possible.

The last thing to be done is to describe the shape the atoms make. The unshared electrons influence the shape, but only the resulting shape of the atoms is described. Certain terminology is used for this. The shape is called the molecular geometry. Listed in Table 1 are some common molecular shapes and the words used to describe them.

Table 1: Molecular Models

Example Atom (central atom is underlined)	**Number of Bonds on Central Atom** (multiple bonds count as 1)	**Unshared electron pairs on central atom**	**Regions of Electron Density**	**Molecular Geometry**	**Structure**
$\underline{C}O_2$	2	0	2	Linear	
$\underline{Al}Cl_3$	3	0	3	Trigonal Planar	
$H_2\underline{O}$	2	2	4	Bent	
$\underline{N}H_3$	3	1	4	Trigonal pyramidal	
$\underline{C}Cl_4$	4	0	4	Tetrahedral	
$\underline{P}Cl_5$	5	0	5	Trigonal bipyramidal	
$\underline{S}F_6$	6	0	6	Octahedral	

All of this can be summarized in 4 steps.

1. Determine the central atom.
2. Determine the regions of electron density around the central atom by adding the number of bonded atoms and the number of unshared electrons.
3. Place the regions the same distance away from the central atom and as far as possible away from each other.
4. Describe the shape only the atoms make.

Objectives

- To understand why molecules have a particular shape.

- To be able to determine the shapes of molecules using the VSEPR model.

Date:________ Name:_______________________ Period:_________

Pre-lab Questions

1. Identify the central atom and determine how many regions of electron densities there are for the central atom in each of the following molecules:

Molecule	Central Atom	Regions of Electron Density
$BeCl_2$	***Be***	***2***
BH_3	***B***	***3***
CBr_4	***C***	***4***

2. Explain why the shape of H_2O is bent and not linear.

Even though H_2O has 2 single bonds, it also has 2 pairs of unshared electrons. This gives it 4 regions of electron density. When you look at the resulting shape of only the atoms, it is bent.

Preparation Materials

General Supplies
3 bar magnets
Toothpicks
Large and mini marshmallows
Protractor

Procedure

MATERIALS NEEDED

- 3 bar magnets
- Toothpicks
- Large and mini marshmallows
- Protractor

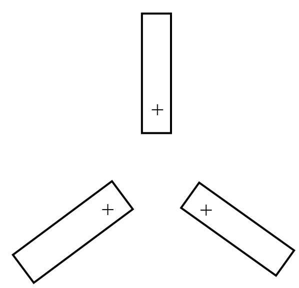
Part 1. Magnets

1. Bring together the **same** poles of two bar magnets.

2. Observe and describe what happens.

3. Explore the linear shape of molecules by bringing the **same** poles as close together as you can in a straight line. Record your observations.

4. Explore why a molecule makes a trigonal planar shape by placing the **same** pole of 3 bar magnets as the points of an equilateral triangle with the rest of the magnet behind them.



5. Slowly bring the same pole of the 3 magnets as close as it is possible **while maintaining the equilateral triangular appearance**. Now try to find another shape that will be as stable with all of the magnets at least as close as they were in the equilateral triangle position. Record your observations.

Part 2. Marshmallows

6. Make a model of a molecule made up of two atoms such as H_2 by attaching a marshmallow to a both ends of a toothpick. Describe the shape.

7. Make a model of a linear molecule such as $CaCl_2$ that is made up of 3 atoms by the following:

a. First attach two toothpicks to a larger marshmallow representing the central atom. Locate them so that they are as far away from each other as possible.
b. Next attach 2 smaller marshmallows to the ends of the two toothpicks.
c. Describe the shape.

d. Determine the bond angle from one of the mini marshmallows to the large marshmallow to another mini marshmallow using a protractor. Record the bond angle in the data section.

8. Make models of the other molecular shapes listed in **Table 1** by the following:

a. First attach the same number of **toothpicks** as shown in the regions of electron density on a **large** marshmallow.
Locate them so that they are as far away from each other as possible.
b. Next attach the same number of **mini** marshmallows as shown in the number of bonds column to the ends of the toothpicks.
c. Remove the toothpicks that do not have small marshmallows attached.
d. Describe and record the shape.
e. Use a protractor to determine and the bond angle from one of the mini marshmallows to the large marshmallow to another mini marshmallow. Record the bond angle in the data section.

Date:________ Name: ___________________________ Period:_________

Lab Partner: ___________________________________

Data

Part 1. Magnets

Step from Procedure	Observations
Step 1	*The same poles will not come together without force. There is repulsion.*
Step 3	*Bringing the same poles together in a linear fashion repulse the other in an upward or downward direction.*
Step 5	*Of the formations explored, the equilateral triangle is the one that allows them to come the closest together.*

Part 2. Marshmallows

Molecular Shape	Observations	Bond Angle
Linear (2 atoms)	*The large and small "atoms" form a line in the same plane.*	*180°*
Linear (3 atoms)	*The large and small "atoms" form a line in the same plane.*	*180°*

Trigonal Planar	*The three small "atoms" forms a triangle shape around the central "atom". All the atoms are in the same plane.*	*120°*
Bent	*The large and small "atoms" form a triangle shape.*	*~104°*
Trigonal pyramidal	*Three "atoms" form a triangle around the bottom of the central "atom"*	*~107°*
Tetrahedral	*The shape of the "atoms" forms a triangle on four planes.*	*~109°*
Trigonal bipyramidal	*Three small "atoms" form a triangle around the central "atom" at 120° with two more "atoms" at 90°, one above and one below the plane.*	*90° & 120°*
Octahedral	*Four "atoms" form a square around the central "atom" in the same plane with two more atoms, one above and one below the plane.*	*90°*

Post Lab Questions

1. In Part 2, did your models fit the molecular shape description or match the given geometry in the introduction? Explain why or why not.

Some models did not result with the exact descriptions and expected bond angles, but most were quite close. The various differences can be attributed to the marshmallows not actually repelling each other as electrons do.

2. Predict the shapes and bond angles of the following molecules:

a. $BeCl_2$

Linear and 180°

b. BH_3

Trigonal planar and 120°

c. CBr_4

Tetrahedral and 109.5°

Lab

11

Sugar or Salt? Ionic and Covalent Bonds

TN Standard ***3.1:*** *The student will investigate chemical bonding.*

Have you ever accidentally used salt instead of sugar?

Drinking tea that has been sweetened with salt or eating vegetables that have been salted with sugar tastes awful! Salt and sugar may look the same, but they obviously taste very different. They are also very different chemically. Salt is made up of sodium and chloride and is ionically bonded. Sugar, on the other hand, is composed of carbon, oxygen, and hydrogen and has covalent bonds.

Introduction

A salt molecule is made up of one sodium ion and one chloride ion. For salt to be made, the sodium atom must lose an electron and become a sodium ion. When sodium loses an electron it becomes a Na^+ and is called a **cation**.

$$\mathbf{Na \rightarrow Na^+ + e^-}$$

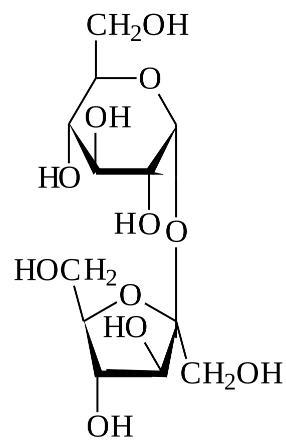
The chlorine atom must add the sodium's electron and become a chloride ion. When it adds the sodium's lost electron it becomes Cl^- and is called an **anion**.

$$\mathbf{Cl + e^- \rightarrow Cl^-}$$

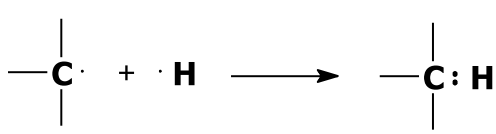
The sodium ion is then attracted to the chloride ion and a bond is formed from the attraction between a positive and negative ion. This type of bond is called an **ionic bond**. Ionic bonds usually form between metals and non-metals.

$$Na^{+} + Cl^{-} \rightarrow Na^{+}Cl^{-}$$

Table sugar or sucrose differs from salt in the bonding between its atoms. The atoms in sugar do not form ions; instead, they share their electrons. The type of bond that forms from the sharing of electrons between the atoms of the table sugar is a **covalent bond.** Table sugar has a much more complex chemical structure than salt. It looks like this:



A bond forms between one of the carbon atoms and one of the hydrogen atoms when one of the valence electrons of the carbon atom combines with one of the valence electrons of the hydrogen atom. This forms an electron pair.



(This is normally written C-H.)

Ionically bonded compounds behave very differently from covalently bonded compounds. When an ionically bonded compound is dissolved in water it will conduct electricity. A covalently bonded compound dissolved in water will not conduct electricity. Another difference is that ionically bonded compounds generally melt and boil at much higher temperatures than covalently bonded compounds.

In the first part of this lab you will investigate how ionically bonded and covalently bonded substances behave differently in their conduction of electricity. You will do this by using a simple anodizing apparatus. A stainless steel screw and an iron nail will be used for the electrodes. In an anodizing apparatus, the water must contain enough ions to conduct electricity. Then the water will react to form hydrogen and oxygen gases.

$$2H_2O \rightarrow 2H_2 + O_2$$

Since the stainless steel screw is not very reactive, bubbles can be seen coming off of it. The iron nail will react with the oxygen to form iron oxide which is commonly called rust. This can be seen on the nail after the reaction proceeds for several minutes.

In the second part of this lab you will explore the differences in melting points between ionically bonded and covalently bonded compounds. You will do this by placing a small amount of sugar in one small test tube and heating it at different heights over a Bunsen burner. You will then repeat this procedure using salt instead of sugar.

Objectives

- To understand the difference between ionic and covalent bonding.
- To link ionic and covalent bonding with their physical properties.

Date:________ Name:______________________ Period_________

Pre-lab Questions

1. What is an ionic bond?

   An ionic bond is a bond formed from the attraction between a positive and negative ion.

2. What is a covalent bond?

   A covalent bond forms when the valence electrons of two atoms are shared.

3. Do you think sugar or salt will melt at a higher temperature? Explain your answer.

   The salt should melt at a higher temperature since it has an ionic bond, while sugar has covalent bonds. Ionically bonded compounds generally melt at higher temperatures than covalently bonded compounds.

Preparation Materials

General Supplies (all per group)	Laboratory Equipment	Chemicals
2 packets of sugar (~3 g each)	150 mL Beaker	Deionized water
2 packets of salt (~0.65 g each)	2 Small test tubes	
2 wire leads with alligator clips on both ends	Test tube holder	
~16d iron nail (not galvanized)	Bunsen burner	
~2 inch long stainless steel screw	Striker	
9-volt battery	Ruler	
2 rubber bands	Test tube rack	
	Stirring rod	

Procedure

SAFETY FEATURES

Safety Goggles

MATERIALS NEEDED

- 2 packets of sugar
- 2 packets of salt
- 9-volt battery
- 2 rubber bands
- Deionized water
- 1 Iron nail
- 2 Wire leads with alligator clips on each end
- 1 Stainless steel screw
- 150 mL beaker
- Stirring rod
- 2 small test tubes
- Spatula
- Test tube holder
- Striker
- Bunsen burner
- Sharpie marker
- Ruler

Part 1. Nail Test for Ionic Bonding

1. Rinse a clean 150 mL beaker several times with deionized water to prevent contamination from ions that may be on the beaker. Fill the beaker about ¾ full with deionized water.

2. Pour a packet of sugar (~3 g) into the 150 mL beaker. Stir the solution with a clean stirring rod until the sugar is dissolved and the solution is well mixed.

3. Stretch two rubber bands around the 150 mL beaker. Be careful not to spill any of the solution. The rubber bands should loop from the top to the bottom of the beaker. Position the 2 rubber bands next to each other (Figure 1). **HINT:** Do not position the bands around the circumference of the beaker.

4. Attach the alligator clip on one end of a wire lead to just underneath the flat head of an iron nail. Place the iron nail between the 2 rubber bands on one side of the 150 mL beaker. The end of the nail should be in the solution while the head with clip is resting on the rubber bands.

5. Attach the alligator clip on one end of another wire lead to just below the head of the stainless steel screw. Place the screw between the 2 rubber bands on the opposite side of the 150 mL beaker. Make sure the end of the screw is in the solution and the head with the clip is resting on the rubber bands.

6. Connect the alligator clip on the other end of the wire lead that is attached to the nail to the (+) terminal of a 9-volt battery. **CAUTION: Be careful when using energy sources such as batteries around water.**

7. Connect the alligator clip on the other end of the wire lead that is attached to the screw to the (-) terminal of a 9-volt battery. **CAUTION: Be careful when using energy such as batteries around water.**

8. Allow the apparatus to stand for two minutes and make observations. Record your observations in Part 1 of the data section.

9. Thoroughly clean the glassware, nail and screw with deionized water.

10. Repeat the procedure using a packet of salt (~0.65 g) instead of the sugar.

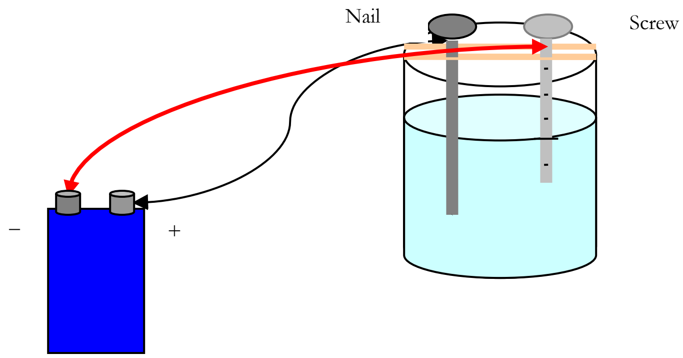


Figure 1. Apparatus for Part 1

Part 2. Melting points

1. Place a spatula tip full of sugar into a small test tube. The sugar should just coat the bottom of the test tube. **CAUTION**: **Be sure the test tube does not have any small cracks or chips in it.**

2. Light a Bunsen burner and adjust the flame to where it is approximately 2 -3 inches tall and has a blue inner cone. **CAUTION**: **Long hair should be tied up and loose clothing restrained when around an open flame to prevent fire and burns**. **Be sure you are wearing your safety goggles.**

3. Place the test tube containing the sugar in a test tube holder. Hold the test tube at a slight angle approximately 4 inches above the top of the burner. Continue to hold the test tube in the flame until the sugar just begins to melt. If it does not melt after ~15 seconds, go on to step **4**. If it has melted, go to step **6**. **HINT:** If you keep the sugar in the flame until it turns dark brown or black, you will not be able to clean the test tube.

4. If the sugar did not melt in step 3, lower the test tube until the bottom of the test tube is touching the top of the flame. Hold it there until the sugar just begins to melt or for ~15 seconds.

5. If the sugar still does not melt, lower the test tube until the bottom of the test tube is located at the top of the inner blue cone of the flame. This is the hottest part of the flame. Hold it there until the sugar just begins to melt or for ~15 seconds.
6. Allow the test tube to cool to room temperature before touching it. **CAUTION**: **The test tube will be very hot and can burn your skin if touched before it cools. Hint:** After the test tube has cooled for a few

seconds, place it in a beaker or wire test tube rack to finish cooling and continue with the procedure.

7. Record your observations in the data section.

8. Repeat the procedure using salt instead of the sugar.

9. Turn off the Bunsen burner. Make sure the test tubes have cooled to room temperature before touching them. **CAUTION**: **The test tube will be very hot and can burn your skin if touched before it cools.**

10. Record your observations in the data section.

11. Clean-up: The sugar and salt solutions can be poured down the drain. Rinse the beaker, screw, nail, and stirring rod several times with deionized water. Clean the test tubes with water first and then rinse them with deionized water. They may need to soak for a few minutes in hot water in order to remove the melted substances.

Date:________ Name: __________________________ Period:_________

Lab Partner: _________________________________

Data

Part 1.

- Observations for the sugar solution:

  Only a very slight bubbling around the screw was noticed, but the overall appearance of the solution changed very little.

- Observations for the salt solution:

  Bubbling was seen around the screw. A small amount of dark precipitate formed around the nail, and the color of the solution darkened a little and clouded slightly.

Part 2.

- Observations for the melting of sugar:

  The sugar melted after only a few seconds when the test tube was held 4 inches above the top of the burner.

- Observations for the melting of salt:

  The salt took much longer to melt. It did not melt until it was placed in the inner blue cone of the flame.

Post Lab Questions

1. Why is deionized water instead of tap water used in Part 1?

Tap water contains a small amount of ions in it, so it will conduct electricity. This would give an incorrect result.

2. In Part 1, why did you not observe a stream of bubbles coming off of the stainless steel screw with the sugar solution?

Sugar has covalent bonds, so when it is dissolved in the water there are no ions present to conduct electricity.

3. Why do you think you may see a few bubbles forming in Part 1 with the sugar solution?

The beaker may have contaminated the water if it was not cleaned and rinsed well enough with deionized water. Also the deionized water may have been slightly contaminated and had some ions still in it.

4. In Part 1, why did you observe a stream of bubbles coming off of the stainless steel screw and rust forming on the nail with the salt solution?

Salt has ionic bonds, so when it is dissolved in the water, there are enough ions present to conduct electricity. The water reacted to form hydrogen and oxygen gases. The oxygen gas reacted with the iron nail to form rust.

5. In Part 2, which of the substances had the lower melting point? Was this what you expected? Explain your answer.

The sugar melted above the flame where temperatures are cooler than in the flame. This demonstrated that sugar had the lower melting point. This was expected because sugar has covalent bonds, and salt has an ionic bond. Covalently bonded substances generally melt at lower temperatures than substances with ionic bonds.

Lab 12

Chemical Bonding: Polarity of Slime and Silly Putty

*TN Standard **3.1**: Investigate chemical bonding. Students will distinguish between polar and non-polar molecules.*

Have you ever read the newspaper using silly putty?

Newsprint can be transferred to silly putty. This oddity is due to the chemical characteristic called polarity. Polarity is based on two primary factors, electronegativity and the shape of the molecule. Polarity is an important aspect of chemistry and it is everywhere. Loads of household substances are examples of both polar and nonpolar molecules. To explore polarity, let's experiment with two favorite toys—slime and silly putty!

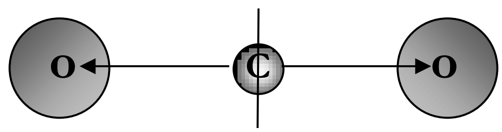
Introduction

The root word for polarity is "pole". There are strong attractions in a molecule that induce poles, similar to the North and South poles of the earth or of a magnet. A molecule can be polar too. Whether a molecule is polar or nonpolar is primarily based on its electronegativity and its shape.

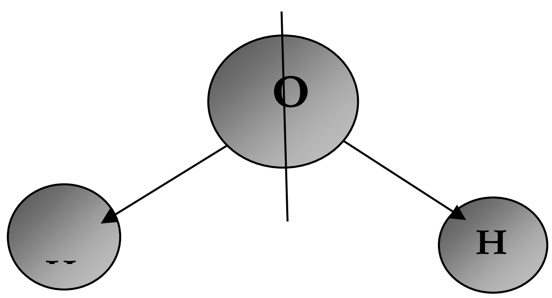
Electronegativity is a term that describes the attraction an atom has for electrons. Fluorine has the strongest attraction for electrons, and therefore has the greatest electronegativity. Elements on the right side of the Periodic Table close to where fluorine is located have larger electronegativities than elements located on the left side of the Periodic Table. If all of the atoms in a molecule have similar electronegativities, the molecule is non-polar. Hexane is an example of a non-polar molecule. It has only carbon and hydrogen atoms. The electronegativities of carbon and hydrogen atoms are very similar. They have only a 0.4 difference in their electronegativity values.

If there is a large difference between the electronegativities of two covalently bonded atoms, the bond is polar. For a covalent bond to be considered polar, the atoms bonded together must have an electronegativity difference between 0.5 and 1.9. But even if a molecule has a bond between two atoms that has this much electronegativity difference, it does not automatically mean that the molecule is polar. The shape of a molecule also has to be considered.

To understand how the shape of the molecule affects the polarity, imagine the atoms within a molecule playing tug-of-war. If there is an equal pull on each side, the rope does not move. Similarly, if there are two polar bonds that have equal electronegativity differences, they each pull with the same force, and the overall pull is zero.



An example of this is CO_2. In this case, there are two polar bonds, but the molecule is non-polar because the overall pull is zero. If there is an unequal net pull, however, the molecule is polar. Water is an example of this. It has a bent shape because the pull is unequal.



Polar molecules such as water, vinegar, or ethanol dissolve other polar molecules. Non-polar molecules, such as oil or gasoline, dissolve other non-polar molecules. This trend is known as **"like dissolves like."** Since sugar dissolves in water, is sugar polar or non-polar? If you said polar, you are correct. Sugar is a polar molecule since it is dissolved water, a polar molecule.

Paper chromatography is a laboratory technique that uses **"like dissolves like"** to separate different chemicals. A sample is first spotted on an absorbent paper such as filter paper and allowed to dry. A solvent is then allowed to travel up the paper by capillary action. The components of the sample most like the solvent will stay dissolved in the solvent more than on the paper and travel the farther up the paper. The components that are least like the solvent will travel the least or not at all.

Some inks are polar while others are non-polar. A polar substance will pick up water soluble inks by dissolving the ink. Likewise a non-polar substance will pick up non-polar inks by dissolving the ink. In this lab you will use inks to identify slime and silly putty as polar or non-polar. You will also use paper chromatography to verify the inks are correctly identified as polar or non-polar.

Objectives

- Compare and contrast the chemical bonding properties of slime and silly putty.
- Demonstrate knowledge gained about polar and non-polar bonding through supporting or rejecting his/her hypothesis about which inks will be absorbed.
- Learn about the technique of chromatography and how it is often used to separate the components of a mixture based on their polarity.

Name:______________________________ Date:__________ Period:________

Pre-lab Questions

1. What two conditions are considered when determining whether a molecule is polar or non-polar?

   The first condition is whether or not a molecule has polar bonds. The second condition is the shape of the molecule.

2. What determines if a bond is polar?

   For a bond to be considered polar there must be a large difference in the electronegativities of the bonded atoms.

3. List examples of polar molecules.

   Water, vinegar, ethanol, sugar

4. List examples of non-polar molecules.

   Oil, gasoline

5. What is the rule when using polar and non-polar solvents?

   "Like dissolves like" is the usual trend.

Preparation Materials

General Supplies	Laboratory Equipment	Chemicals
Notebook /Scrap paper	10 mL Graduated cylinder	Borax solution (4%) Prep: 4 g diluted to 100 mL with H_2O
Newspaper	100 mL Graduated cylinder	Guar gum (0.5 g per group)
Variety of ink : Uni-ball pens, Sharpies, Highlighters, and Dry eraser markers	250 mL Beaker	
Zip lock bag (1 per group)	Stirring rod	
Distilled water	Balance	
Silly putty	Filter paper (Whatman 41 Ashless Circles 90 mm) (1 & ½ sheets per group)	
	Spatula	

Procedure

SAFETY FEATURES

- Glassware
- Safety Goggles

MATERIALS NEEDED

- Borax solution (4%)
- Guar gum
- Silly putty
- Notebook paper
- Newspaper
- Variety of inks: Uni-ball pens, Sharpies, Highlighters, and Dry erase markers
- 10 and 100 mL Graduated Cylinder
- 250 mL beaker
- Filter paper
- Stirring rod
- Spatula
- Balances
- Distilled water
- Zip lock plastic bag

Teaching Alternative:
Part 1 can be eliminated by purchasing slime. However, the purchased slime should be a light colored so the ink dissolved can be easily detected.

Part 1. Making Slime

1. If slime is provided go to Part 2. If it is not provided, weigh out 0.5 g of guar gum into a 250 mL beaker.

2. Measure 50.0 mL of distilled water into a 100 mL graduated cylinder and pour it into the 250 mL beaker that contains the guar gum.

3. Rapidly stir the mixture with a stirring rod for at least 3 minutes and until the guar gum is dissolved.

4. Measure 4.00 mL of a 4% borax solution into a 10 mL graduated cylinder and add it to the guar gum and water.

5. Stir the solution until it becomes slime. This will take a few minutes. If the slime remains too runny, add 1.0 mL of the 4.0% borax solution and continue to stir until the slime is the right consistency.

6. Once you are satisfied with the slime, pour it into your hands. Be sure not to drop any of it on to the floor.

7. Manipulate the slime in your hands. Write down observations made about how slime pours, stretches, breaks, etc. **CAUTION: Slime is slippery, and if dropped, it can make the work area slick.**

8. Place the slime back into the beaker and WASH YOUR HANDS.

Part 2. Slime and Silly Putty Ink Tests

1. On a piece of notebook paper make one 20-25 mm long mark of each of the inks you are testing. Space the marks at least one inch apart. Use a pencil to label each mark with its description.

- **Water soluble inks** include a highlighter and a Uni-ball pen.
- **Water insoluble inks** include a Sharpie pen/marker, newsprint, and a dry-erase marker.

Helpful HINT:
Two of each type of ink, two soluble ink samples and two insoluble ink samples, should be sufficient for this experiment.

2. While the inks are drying, select a passage or a picture in the newspaper to test with the slime.

3. Break off a small piece that is 3-5 cm in diameter of slime. Gently place this piece on top of the newspaper ink, and then carefully pick it up again.

4. Observe and record whether or not the ink was picked up onto the slime.

5. Break off another small piece of slime. Gently place it on top of the first dried ink on the notebook paper, and then carefully pick it up. Repeat this for each of the inks.

6. Observe and record which inks were picked up (dissolved) by the slime in Table 1. Repeat this ink testing two more times for accuracy.

7. Store the slime in zip lock plastic bag.

8. Before performing ink tests on silly putty, in the data section hypothesize which inks the silly putty will pick up.

9. Perform ink tests on silly putty in the same manner as above.

10. Record results in Table 2.

Part 3. Chromatography of Ink Samples

1. Use a pencil or scissors to poke a **small** hole in the center of a piece of filter paper (see Figure 1).

2. Spot the filter paper evenly spaced ~3 cm from the small hole with the two insoluble inks and the two soluble inks that were used in Part 2.

3. Obtain a ½ piece of filter paper. Fold the paper in half several times so that it makes a narrow wick.

4. Insert the wick into the hole of the spotted paper so that it is above the top of the filter paper by ~2 cm.

5. Fill a 250 mL beaker 3/4 full with water.

6. Set the filter paper on top of the beaker so that the bottom of the wick is in the water. The paper should hang over the edge of the beaker with the spotted side up.

7. Allow water to travel until it is ~1 cm from the edge of the filter paper. Remove the filter paper from the beaker.

8. Observe which inks moved from where they were originally spotted. Record your observations in Part 3 of Data.

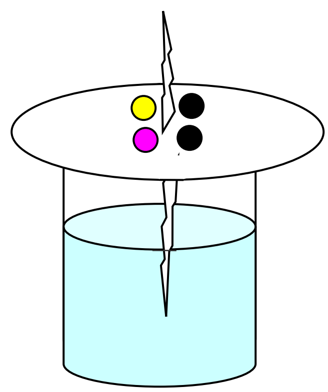


Figure 1: Chromatography Apparatus

Name:______________________________ Date:__________ Period:________

Lab Partner:_________________________

Data

Part 1.

- Observations for Slime (if slime synthesized)

The slime is slightly runny and light colored.

Part 2.

Table 1. Results of Ink Testing for Slime

Name of Ink	Picked up (dissolved)			Did not pick up (dissolve)		
	Test 1	Test 2	Test 3	Test 1	Test 2	Test 3
Newsprint				X	X	X
Highlighter	X	X	X			
Uni-ball pen	X	X	X			
Sharpie marker				X	X	X
Dry-erase marker				X	X	X

❖ Hypothesis of which inks silly putty will pick up:

The silly putty will pick up ink from the newsprint and Sharpie and Dry-erase markers.

Table 2. Results of Ink Testing for Silly Putty						
Name of Ink	Picked up (dissolved)			Did not pick up (dissolve)		
	Test 1	Test 2	Test 3	Test 1	Test 2	Test 3
Newsprint	X	X	X			
Highlighter				X	X	X
Uni-ball pen				X	X	X
Sharpie marker	X	X	X			
Dry-erase marker	X	X	X			

Part 3.

- Observations of inks following chromatography:

The water soluble inks traveled on the filter paper and their components were separated. The water insoluble inks did not travel on the filter paper.

Post Lab Questions

1. Did the slime pick up water soluble or water insoluble inks? From these results what can use conclude about the polarity of slime?

The slime picked-up or dissolved the Uni-ball pen and highlighter, which are water soluble inks. From this result and the rule "like dissolves like" it can be determine that slime is polar just the water soluble inks are polar.

2. Explain how you determined your hypothesis about whether or not silly putty would pick up water. Was your hypothesis correct?

The drawing of my hypothesis resulted from the early test of newsprint. The newsprint test revealed that it was only dissolved by silly putty and not by slime. However, following the test with slime it was found that slime dissolved water soluble inks. This means it was polar. Therefore the silly putty must be the opposite or non-polar.

3. Were the inks you used properly classified as soluble and insoluble? Explain your answer.

Yes. Water soluble inks spotted on the filter paper should dissolve in the water and travel and separate on the filter paper. Water insoluble inks should not dissolve in the water and travel on the filter paper. The inks classified as water soluble (highlighter and Uni-ball pen) traveled and on the paper and the water insoluble inks (dry-erase marker and Sharpie marker) did not.

Lab

13

Qualitative Tests for Fluoride Ions in Mouth Rinses

*TN Standard **2.1**: The student will investigate the characteristics of matter.*

Have you ever wondered how toothpaste helps prevent cavities?

Remember your last visit to the dentist to have you teeth cleaned and a check-up? You most likely anxiously waited as the dentist finished checking your teeth, and breathed a big sigh of relief if he pronounced, "No cavities!" But what causes cavities anyway, and how does brushing your teeth help to prevent them?

Introduction

Your teeth like many other bones in the body are mostly made of a substance called hydroxyapatite. The empirical formula of hydroxyapatite is $Ca_{10}(PO_4)_6(OH)_2$. Plaque organisms that live in your mouth produce acids that dissolve the minerals, such as the calcium and phosphorous found in hydroxyapatite. We need something to help replace those lost minerals. The active ingredient in toothpaste to do this is fluoride. It has been found that the hydroxyapatite in your teeth can easily replace lost minerals with fluoride. A tooth's crystal structure that has incorporated fluoride is more resistant to decay than the original structure. That's why we need to brush our teeth with toothpaste that contains fluoride each day!

Fluoride is an ion. Ions have very different chemical and physical properties than atoms. Fluorine (F_2) as an element is a very poisonous gas. Fluoride (F^-), as an ion, is safe enough to add to drinking water and toothpaste to help prevent cavities. Throughout history, scientists have developed many ways to test for the presence of different types of ions. These are called **qualitative** tests. In many of these tests, one ion replaces another to give a solid precipitate. There are also some useful generalizations that can help a chemist know whether or not a particular chemical is soluble in water. The solubility of compounds in water can also be found in handbooks.

Some mouth rinses also contain fluoride, usually as sodium fluoride. It is soluble in water. This relates to one of the useful solubility rules for ionic compounds. This is that compounds containing Li^+, Na^+, K^+, or NH_4^+ are soluble in water. In this lab you will determine which one of two mouth rinses would be better at preventing cavities by replacing lost minerals with fluoride. You will do this by determining which mouth rinse contains fluoride.

To qualitatively test for fluoride ions in two different mouth rinses, you will use calcium nitrate. This relates to another one of the solubility rules. It is that nitrates (NO_3^-) and acetates (CH_3COO^-) are water soluble. This means calcium nitrate is soluble in water. A solution of calcium nitrate will react with fluoride to make calcium fluoride. Calcium fluoride is not very soluble in water and so it forms a solid precipitate.

You can see that nitrates (NO_3^-), acetates (CH_3COO^-), and ammonium ions (NH_4^+) all have a charge like the fluoride ion, but they are made up of more than one atom. Ions that are made up of more than one atom are called **polyatomic ions.** They exist as a group and have a certain charge.

Chemists have a short way to write out what is happening in a reaction. It is called a **chemical equation**. The chemicals that are reacting are called the reactants and placed on the left side of an arrow. The chemicals that are made are called the products and placed on the right side of the arrow. The reaction that you will be doing can be written as the following chemical equation:

$$\mathbf{2\,NaF_{(aq)} + Ca(NO_3)_{2(aq)} \rightarrow CaF_{2(s)} + 2\,NaNO_{3(aq)}}$$

The (aq) by the chemical formula means the substance is dissolved in water, and the (s) means the substance is a solid.

This chemical equation has also been balanced. This means there is the same number of each of the different atoms on the left side of the arrow where the reactants are written as there is on the right side of the arrow where the products are written. This may not be very clear when you first look at it. Notice that there is a 2 in front of the NaF and no number in front of the $Ca(NO_3)_2$. A number in front of a molecule, such as the 2 in front of the NaF, is called a **coefficient**. If there is no coefficient in front of a molecule it has a coefficient of 1, the 1 is just not written. One molecule of a substance does not always react with just one molecule of another substance. The coefficients tell us the ratio of how the molecules react. In this reaction 2 NaF react with 1 $Ca(NO_3)_2$.

A subscript after an element's symbol in a chemical formula means that there are that many of that particular element in the compound. CaF_2 has one Ca and two F's. This is also true for polyatomic ions. In $Ca(NO_3)_2$, the NO_3^- is in parenthesis and has a 2 as a subscript at the end of it. This means there are two of them in calcium nitrate. You can think of it as $Ca(NO_3)(NO_3)$. $Ca(NO_3)_2$ has one Ca, two N's and six O's. If a molecule with a polyatomic ion in it does not have it in parenthesis it means that there is only one of them for each molecule. The $NaNO_3$ has only one NO_3^-.

Now that you know how a chemical equation is written, count the number of each atom on the reactants side and on the products side for the reaction that you are going to be doing. Did you get 2 Na, 2 F, 1 Ca, 2 N, and 6 O on each side? If so, you are correct! Since there is the same number of each atom on both sides of the arrow, the chemical equation is balanced.

Objectives

- Understand how fluoride prevents cavities.
- Understand qualitative tests for ions that are based on solubility.
- Learn that Na^+, K^+, NH_4^+, NO_3^-, and CH_3COO^- are water soluble.
- Recognize a balanced chemical equation
- Qualitatively test for fluoride ions in mouth rinse.

Name:______________________________ Date:__________ Period:________

Pre-lab Questions

1. Name the chemical that makes up teeth.

   Hydroxyapatite

2. How does plaque harm teeth?

   Plaque is an organism that lives in the mouth and produces acids that dissolve away healthy minerals in the teeth.

3. Why does fluoride promote healthy teeth?

   Fluoride is known to replace minerals in teeth such as calcium and phosphorus that have been dissolved by the acids. Teeth with fluoride in their structure are more resistant to tooth decay.

4. List two solubility rules that are used in this experiment.

   1. Compounds containing Li^+, Na^+, K^+, or NH_4^+ are soluble in water.

   2. Nitrates (NO_3^-) and acetates (CH_3COO^-) are soluble in water.

Preparation Materials

***Helpful HINT:* Generic brands work VERY well.**

***Helpful HINT:* ACT contains fluoride, where Scope and Listerine do not.**

General Supplies	Laboratory Equipment	Chemicals
Mouth Wash – 2 different types: ACT & Scope or ACT & Listerine (Generic can be used)	3 medium sized test tubes (18 x 130 mm) (per group)	0.1 *M* Sodium fluoride, NaF (10 mL per group) Prep: 0.420 g per 100 mL H_2O
	10 mL graduated cylinder	1 *M* Calcium nitrate, $Ca(NO_3)_2$ (9 mL per group) Prep: 16.408 g per 100 mL H_2O
	Wax pencil/Sharpie	
	Stirring rod	

For a greener alternative, 1*M* calcium acetate can be substituted for the 1*M* calcium nitrate.

The children's fluoride rinse can be added as another sample to test. Although it gives a precipitate, it is considerably less of one.

Procedure

SAFETY FEATURES

Safety Goggles

MATERIALS NEEDED

- Two different mouth rinses
- 0.1 *M* Sodium Fluoride, NaF
- 1 *M* Calcium Nitrate, $Ca(NO_3)_2$
- 3 medium-sized test tubes
- 10 mL Graduated Cylinder
- Stirring rod
- Wax pencil or Sharpie

1. Label the three medium test tubes with a wax pencil or Sharpie: **NaF**, **A**, and **B**.

2. Pour 10 mL of 0.1 *M* NaF into the test tube marked **NaF**. This will serve as the positive control.

3. Pour 10 mL of Rinse A into the test tube marked **A.** **HINT**: If using the same graduated cylinder, rinse WELL to prevent cross contamination.

4. Pour 10 mL of Rinse B into the test tube marked **B**.

5. Pour 3 mL of 1 *M* $Ca(NO_3)_2$ solution into each of the test tubes. Gently stir each test tube with a stirring rod to mix. Be sure to clean your stirring rod each time before placing it in a solution. **CAUTION: Mixing should be done gently to prevent glass breakage and injury.**

6. Observe the reactions for at least 10 minutes to insure it is finished. Compare the positive test for fluoride with the two mouth rinses. **HINT:** A positive test is indicated by a cloudy appearance of the solution. The precipitate formed can be more easily seen if the test tube is held up to the light. The precipitate will eventually settle to the bottom of the test tube.

7. Record all observations in the data section.

8. To clean up, you can rinse the small amount of precipitate down the drain.

Name:______________________________ **Date:**__________ **Period:**________

Lab Partner:_________________________

Data

- Observations of NaF and $Ca(NO_3)_2$:

A precipitate formed quickly and after time settled to the bottom of the tube.

- Observations of Rinse A and $Ca(NO_3)_2$:

A precipitate formed quickly and after time settled to the bottom of the tube. This was similar to the NaF standard.

- Observations of Rinse B and $Ca(NO_3)_2$:

A precipitate did not form. After adding the $Ca(NO_3)_2$ no changes in the solution were observed.

Post Lab Questions

1. Did either of the mouth rinses contain fluoride? How did you know?

Rinse A was found to contain fluoride. This was determined by comparing the mouth rinses to the standard after the sodium fluoride was added.

2. Which mouth rinse would be better at fighting cavities? Why?

The mouth rinse that had a positive test for fluoride, rinse A, would be the best because it strengthens the crystal structure of the teeth.

3. Based on the solubility generalizations learned in this lab, could you use potassium nitrate to test for fluoride in mouth rinses? Explain your answer.

No. Potassium fluoride precipitate would not form since compounds containing potassium ions are soluble in water.

Lab

14

Types of Chemical Reactions

*TN Standard **3.2**: The student will analyze chemical reactions.*

Chemical reactions and laundry: what's the connection?

Laundry involves many different colors and fabrics that require special treatments for washing and drying. You simply can't wash a white dress shirt with a bright red sweater. Why? Because the colors will blend together and you won't exactly have a white shirt anymore. You must sort your laundry into colors, darks, and whites. In a similar fashion, chemical reactions can be sorted into categories based on the characteristics of the reactions.

Introduction

There are many types of reactions, and many ways you can classify reactions. A reasonably easy way is to divide the reactions into five categories. They are combustion, synthesis, single replacement, double replacement, and decomposition. Most reactions can be placed into these categories, including the reactions we will observe in this lab.

The first reaction you will perform is a **combustion** reaction. During combustion, a hydrocarbon and oxygen break into two simple compounds, water vapor and carbon dioxide gas. A hydrocarbon is a molecule that contains carbon and hydrogen atoms. To illustrate, natural gas, the gas used for most Bunsen burners, will combust with the oxygen in the air. Natural gas is mainly methane gas (CH_4). The equation for this reaction is:

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

In the chemical reaction given above, all of the reactants and products are given, but is the equation balanced? From the law of conservation of mass, we know that atoms are never destroyed or created. In the chemical equation above there is 1 C on both sides, but there is 4 H and 2 O on the left side (reactant side) of the arrow and 2 H and 3 O on the right side (product side) of the arrow. This means that the equation is not balanced. If you answered no, you are correct!

There are 4 steps used to balance a chemical equation.

1. Count the number of atoms of each element on both the reactant and the product sides.

2. Determine which atoms are not the same for both sides.

3. Balance one element at a time by changing the coefficients for the molecules in the reaction and not their chemical formulas.

4. After you think the chemical equation is balanced, check it as in step 1.

Now let's use these 4 steps to balance the chemical equation.

Step 1

Reactants	***Products***
1 C	1 C
4 H	2 H
2 O	3 O

Step 2

The hydrogen and oxygen are not balanced.

Step 3

Insert a 2 before the O_2 on the reactants side and a 2 before the H_2O on the products side.

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

Step 4

Reactants	*Products*
1 C	1 C
4 H	4 H
4 O	4 O

Now the chemical equation is balanced! The rest of the chemical equations for the reactions that you will do are given in this introduction as unbalanced. Try to balance each one.

The next type of reaction you will perform is a **synthesis** reaction. A synthesis reaction takes two or more substances and combines them to create a more complex substance. A general reaction equation for this type of reaction is:

$$\mathbf{A + B \rightarrow C}$$

The unbalanced chemical equation for the synthesis reaction that you will do is:

$$\mathbf{Mg_{(s)} + O_2 \rightarrow MgO_{(s)}}$$

In this reaction magnesium ribbon is heated and the oxygen in the air combines with it to form a single product, the white powder of magnesium oxide.

Following the synthesis reaction you will perform a **single replacement** reaction. This type of reaction takes place when a more reactive component replaces a component within a compound. This reaction involves two or more reactants and produces two or more products. An example of a single replacement reaction is when a metal replaces another metal. A general reaction equation for this type of reaction is:

$$\mathbf{A + BC \rightarrow AC + B}$$

The unbalanced chemical equation for the single replacement reaction that you will do is:

$$\mathbf{Mg_{(s)} + HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2\,(g)}}$$

This reaction also uses magnesium ribbon, but here magnesium is a more reactive component that is replacing the less reactive component hydrogen. The magnesium chloride stays in solution and hydrogen gas is given off. When this reaction is allowed to go to completion, the magnesium ribbon disappears.

The fourth type of reaction you will perform is a **double replacement** reaction. It involves two different ionic compounds. The ionic compounds exchange components in the reaction. Most single or double replacement reactions take place in an aqueous solution where the free ions can float around and react. The general format for the reaction is:

$$AB + CD \rightarrow AD + CB$$

The unbalanced chemical equation for the double replacement reaction that you will do in this lab is:

$$Zn(C_2H_3O_2)_{2\,(aq)} + Na_3PO_{4\,(aq)} \rightarrow NaC_2H_3O_{2\,(aq)} + Zn_3(PO_4)_{2\,(s)}$$

In this reaction the zinc and sodium change places. The zinc bonds to phosphate ions and sodium bonds to acetate ions. Zinc phosphate is not soluble in water like the two reactants. After the two reactants are mixed, a precipitate of zinc phosphate is seen.

A **decomposition** reaction is the last type of reaction you will do. It is much like a synthesis reaction running in reverse. In this type of reaction, a more complex compound breaks down into a less complicated compound or elements.

$$C \rightarrow A + B$$

The unbalanced chemical equation for the decomposition reaction that you will do in this lab is:

$$(NH_4)_2CO_{3\,(s)} \rightarrow NH_{3\,(g)} + H_2O_{(g)} + CO_{2(g)}$$

As shown, the ammonium carbonate decomposes when heated to form three gases: ammonia, water vapor, and carbon dioxide.

These five categories of reactions will give you a good foundation to understand reaction processes throughout chemistry.

Objectives

- Observe chemical reactions, while identifying the products and reactants.
- Classify the type of chemical reaction.
- Balance chemical equations.

Date:________ Name:_______________________ Period:_________

Pre-lab Questions

1. Michelangelo used fresco painting when he painted the Sistine Chapel. Fresco painting involves most of the types of chemical reactions you just studied. Listed below are some of the reactions used in creating a Fresco painting. Identify the type of chemical reaction used for each step and balance the chemical equation if needed.

a. Initially some sort of heat must be generated. Propane is an example of a common fuel source used for heating.

C_3H_8 + ***5*** $O_2 \rightarrow$ ***3*** CO_2 + ***4*** H_2O

This is a ___________***combustion***_____________ reaction.

b. Next quicklime (calcium oxide) is made by roasting calcium carbonate (limestone).

$CaCO_3 \rightarrow CaO + CO_2$

This is a ___________***decomposition***_____________ reaction.

c. The quicklime is slaked to form lime plaster.

$CaO + H_2O \rightarrow Ca(OH)_2$

This is a ___________***synthesis***_________________ reaction.

d. The lime plaster is cured or dried.

$Ca(OH)_2 + CO_2 \rightarrow Ca(OH)(HCO_3)$

This is a ____________***synthesis***_________________ reaction.

This quickly continues to react to form calcium carbonate and water.

$Ca(OH)(HCO_3) \rightarrow CaCO_3 + H_2O$

This is a ____________***decomposition***_______________ reaction.

e. Frescos will deteriorate over time when exposed to the damp, acidic environments that is the typical atmosphere of modern urban cities.

$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2$

This is a ________***double replacement***______________ reaction.

2. Balance each of the chemical equations you will be doing in this laboratory exercise.

Combustion

$CH_4(g)$ + _***2***_ $O_2(g)$ → $CO_2(g)$ + __***2***_ $H_2O(g)$

Synthesis

_***2***_ Mg (s) + O_2 → _***2***_ MgO (s)

Single Replacement

Mg (s) + _***2***_ HCl(aq) → $MgCl_2$ (aq) + H_2 (g)

Double Replacement

_***3***_ $Zn(C_2H_3O_2)_2$ (aq) + _***2***_ Na_3PO_4 (aq) → _***6***_ $NaC_2H_3O_2$ (aq) + $Zn_3(PO_4)_2$ (s)

Decomposition

$(NH_4)_2CO_3$ (s) → _***2***_ NH_3 + H_2O + CO_2 (g)

Preparation Materials

Laboratory Equipment	**Chemicals**
Crucible	Magnesium ribbon (~4 cm per group)
Crucible tongs	1 M HCl (2 mL per group) Prep: 8.39 mL HCl (37%) per 100 mL H_2O
Ring stand, ring, and clay triangle	0.1 M Zinc acetate ($Zn(C_2H_3O_2)_2$) (2 mL per group) Prep: If using Zinc acetate dihydrate, use 2.195 g $Zn(C_2H_3O_2)_2 \cdot 2\ H_2O$ per 100 mL H_2O
Bunsen burner and striker or match	0.1 M Sodium phosphate tribasic (Na_3PO_4) (2 mL per group) Prep: 1.639 g Na_3PO_4 per 100 mL H_2O or if using Sodium phosphate dodecahydrate, use 3.801 g $Na_3PO_4 \cdot 12\ H_2O$ per 100 mL H_2O
3 small test tubes (~13 x 100 mm)	Ammonium carbonate $(NH_4)_2CO_3$ (~ 0.02 gram per group)
10 mL graduated cylinder	
Test tube rack (or small beaker) and holder	

Procedure

SAFETY FEATURES

- Safety Gloves
- Safety Goggles
- Corrosive

MATERIALS NEEDED

- 3 cm Magnesium ribbon (Mg)
- Bunsen burner
- Crucible tongs
- Small Crucible
- 2mL 1M HCl
- Ring Stand
- Clay Triangle
- 3 Test Tubes
- Test tube holder
- Test tube rack
- 10 mL graduated cylinder
- 2 mL 0.1M zinc acetate ($Zn(C_2H_3O_2)_2$)
- 2 mL 0.1M sodium phosphate tribasic (Na_3PO_4)
- Ammonium carbonate powder, $(NH_4)_2CO_3$

Part 1: Combustion

$$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$$

1. Light a Bunsen burner and observe the flame. (The ignition of the flame is a reaction between natural gas and the oxygen in the air you breathe.)

2. Record your observations in the data table from when the gas is turned on until it is turned off.

Part 2: Synthesis A + B → C

$$Mg\,(s) + O_2 \rightarrow 2\,MgO\,(s)$$

1. Attach a small ring to a ring stand. Place a clay triangle on the ring.

2. Move a Bunsen burner (not yet ignited) directly underneath the clay triangle. Adjust the ring so that the clay triangle is about 3 inches above the burner.

3. Place a ~2 cm piece of magnesium ribbon in a small crucible. Make and record observations in the data section about the appearance of the magnesium ribbon.

4. Place the crucible so that it is sitting upright in the clay triangle. Ignite the Bunsen burner and heat the crucible for about 3 minutes.

5. Turn off the Bunsen burner and allow the crucible to cool to room temperature. This will take several minutes. **CAUTION: The crucible is extremely hot and fragile after heating. Burns and breakage can easily occur.**

6. After the crucible has cooled, observe the product. Record your observations in the data table.

A. Alternative DEMONSTRATION IDEA**: Set up a Bunsen burner with an evaporating dish set at the base of the burner. Use crucible tongs to place about 5cm of magnesium ribbon in a flame until it catches fire. Next, place the burning ribbon over the evaporating dish.* ***CAUTION: Tell the students to not look directly at the magnesium ribbon while it is burning. Be sure to hold the magnesium away from you. *Place the remaining material of the magnesium in the evaporating dish. Let the students examine its products and record their observations in the data table.*

Part 3: Single Replacement A + BC → AC + B

$Mg\ (s) + HCl(aq) \rightarrow MgCl_2\ (aq) + H_2\ (g)$

1. Place a small test tube in a test tube rack or small beaker.
2. Use a 10 mL graduated cylinder to measure out ~2 mL of 1 M HCl and pour it into the test tube. **CAUTION: HCl can be corrosive to skin and clothing.**
3. Now carefully drop a ~1 cm piece of magnesium ribbon into the test tube with the 1 M HCl. Make observations until the reaction is finished. Record your observations in the data table.

B. DEMONSTRATION IDEA**: Do steps 1-3 as written. After adding the magnesium metal, place a rubber stopper so that it fits **loosely** in the test tube. Light a match. As soon as the pressure makes the stopper pop off, place the lit match just inside the mouth of the test tube. Have the students listen for a distinct pop! This can be repeated several times.* ***CAUTION: Be sure the test tube is away from your body and face.

Part 4: Double Replacement AB + CD → CB + AD

$Zn(C_2H_3O_2)_2\ (aq) + Na_3PO_4\ (aq) \rightarrow NaC_2H_3O_2\ (aq) + Zn_3(PO_4)_2\ (s)$

1. Pour ~2 mL of 0.1 M zinc acetate ($Zn(C_2H_3O_2)_2$) into a clean test tube.
2. Add ~2 mL of 0.1 M sodium phosphate tribasic (Na_3PO_4) into the test tube.
3. Record your observations before and after the addition of Na_3PO_4 in the data table.

Part 5: Decomposition AB → A + B

$$(NH_4)_2CO_3\ (s) \rightarrow NH_3 + H_2O + CO_2\ (g)$$

> ***Helpful HINT:***
> ***A small test tube with the approximate amount of ammonium carbonate could be prepared for students to look at to see how much sample to add to the test tube.***

1. Place a spatula tip full (~0.02 g) of ammonium carbonate $(NH_4)_2CO_3$, powder into a small test tube.

2. Light a Bunsen burner and adjust the flame to where it is approximately 2 - 3 inches tall. **CAUTION**: **Long hair should be tied up and loose clothing restrained when around an open flame to prevent fire and burns**. **Be sure you are wearing your safety goggles.**

3. Use a test tube holder to hold the test tube containing the ammonium carbonate at a slight angle in the flame and pointed away from you and other students. Continue to heat the sample until the reaction is finished. **Hint:** Remember the products of this reaction are all gases.

4. Allow the test tube to cool to room temperature before touching it. **CAUTION**: **The test tube will be very hot and can burn your skin if touched before it cools. Hint:** After the test tube has cooled for a few seconds, place it in a beaker or wire test tube rack to finish cooling.

5. Record your observations in the data table.

Date:________ Name: __________________________ Period:_________

Lab Partner: __________________________________

Data

Reaction	Before the Reaction	After the Reaction
Combustion	*Unreacted gases are present. (No flame)*	*A blue flame was seen when ignited.*
Synthesis	*The Mg ribbon is a solid.*	*A white residue appeared on the Mg ribbon.*
Single Replacement	*The Mg ribbon is a solid and the 1M HCl is a clear liquid.*	*Bubbling from gas formation resulted and the magnesium disappeared.*
Double Replacement	*Both reactants are clear solutions.*	*A white precipitate formed and caused a cloudy or milky appearance of the solution.*
Decomposition	*The ammonium carbonate is a white powder. There is a slight ammonia odor.*	*A white gas formed and the ammonium carbonate eventually disappeared. A condensate formed in the upper part of the test tube.*

Post Lab Questions

1. Write the combustion reaction that occurs when a butane lighter is lit. Butane is C_4H_{10}. Be sure to balance the reaction equation.

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

2. If you had weighed the magnesium ribbon before heating it and then weighed it after the MgO had formed, would you expect an increase or a decrease in mass? Explain your answer.

You would expect an increase in mass. The magnesium added oxygen to form the product so the mass would increase.

3. Balance the following equations and identify the type of reaction.

A. $BaCl_2\ (s) + K_2SO_4\ (aq) \rightarrow BaSO_4\ (s) +$ **2** $KCl\ (aq)$

Double replacement reaction

B. **2** $KClO_3(s) \rightarrow$ **2** $KCl(s) +$ **3** $O_2(g)$

Decomposition reaction

C. **2** $H_2\ (g) + O_2\ (g) \rightarrow$ **2** $H_2O\ (l)$

Synthesis reaction

D. $F_2\ (g) +$ **2** $LiCl\ (aq) \rightarrow$ **2** $LiF\ (aq) + Cl_2\ (l)$

Single replacement reaction

Lab

15

ReDox Reaction: Can Aluminum Become Magnetic?

*TN Standard **3.2:** The student will analyze chemical reactions.*

Have you ever wondered why some rings with age will turn a finger green and others won't?

The ring you just bought a couple of weeks ago is already turning your finger green. Another ring that you have had for years still looks almost new. Why is this? Knowing how reactive different metals are is extremely important. It is one of their properties that helps us to decide which metal to use for a particular application. Some metals are so reactive that they will chemically react with water. Other metals will only react with acids, and others are so non-reactive that they will not react even with strong acids. This is why jewelry is often made out of gold and not ordinary iron. Iron is more reactive, and will gradually rust just by being exposed to the oxygen and moisture in the air. On the other hand, gold is often used in jewelry since it is very non-reactive.

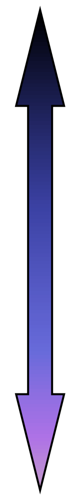
Introduction

You know that gold is less reactive than iron, but what about other metals? How do you know which ones are more reactive. A list of metals has been made that gives their order of activity. This list is called the activity (or the electromotive) series of metals. It is based on experimental evidence. Metals higher on the list give up their valence electrons more easily than the metals below them. This means that any metal in this list will displace a metal below it in its aqueous solution, causing the metal below to precipitate out of the solution. A partial activity series list is shown below.

Table 1: Partial List of the Activity Series of Metals

Metal	Reaction with hydrogen sources
Lithium	
Potassium	Release hydrogen from cold water, steam, and acids
Barium	
Calcium	
Sodium	
Magnesium	
Aluminum	Release hydrogen from steam and acids
Zinc	
Chromium	
Iron	
Cobalt	
Nickel	Release hydrogen from acids
Tin	
Lead	
Hydrogen	
Copper	Do not release hydrogen from acids
Mercury	
Silver	
Platinum	
Gold	

Lose electrons easily
(more easily oxidized)



Do not lose electrons easily
(less easily oxidized)

In Part 1 of this lab, you will observe the reaction of iron (III) chloride and aluminum foil. The reaction equation is as follows:

$$FeCl_{3(aq)} + Al_{(s)} \longrightarrow Fe_{(s)} + AlCl_{3(aq)}$$ **Equation 1**

Since aluminum is higher on the Activity Series of Metals list it gives up its valence electrons easier than the iron. In this reaction, the aluminum transfers its valence electrons to the iron. This causes the iron in the iron chloride solution to become iron metal and precipitate out while the aluminum metal forms aluminum ions and goes into solution. The aluminum is said to be more electropositive than the iron that it displaces.

Reactions like this one that involve the transfer of electrons are called oxidation reduction reactions, or redox for short. The metal losing electrons is being oxidized and the metal gaining electrons is being reduced. In this reaction, since aluminum is losing electrons, it is being oxidized. Iron is gaining electrons so it is being reduced.

But how do we know just by looking at this reaction equation that it is an oxidation reduction reaction and that electrons are being transferred? To know this the oxidation number must first be assigned to each of the atoms involved on both sides of the reaction equation. There are rules for assigning the oxidation numbers. One of the rules that applies for this reaction is that the oxidation number of an element is always zero. This means that the aluminum metal (Al) on the left side of Equation 1 and the iron (Fe) on the right side of the equation have oxidation numbers of 0.

Another rule that applies to this reaction is that when an atom exists as a simple ion in a substance the oxidation number is the same as its charge. This tells us that the oxidation number for the Cl in both the $FeCl_3$ and the $AlCl_3$ is -1 and that the Al in $AlCl_3$ is +3.

The next rule that applies is that for a neutral compound the sum of the oxidation numbers should equal 0. A polyatomic ion's oxidation number equals the charge on the ion. Since $FeCl_3$ does not have an overall charge, the sum of the oxidation numbers should equal zero. We already determined that each Cl has an oxidation number of -1. Since there are three of them, this adds up to -3. For the sum of the oxidation numbers in $FeCl_3$ to equal 0, the Fe must be +3.

Sometimes it is helpful to write the oxidation numbers over the atoms as is shown in Equation 2.

$$\overset{+3}{Fe}\overset{-1}{Cl}_{3(aq)} + \overset{0}{Al}_{(s)} \longrightarrow \overset{0}{Fe}_{(s)} + \overset{+3}{Al}\overset{-1}{Cl}_{3(aq)} \qquad \textbf{Equation 2}$$

The activity series of metal also indicates how easily a metal will release hydrogen. Metals at the top portion of the list release hydrogen merely by being placed in cold water. As you go down the list more harsh conditions are required to release hydrogen. Metals towards the bottom do not release hydrogen even when placed in acids.

In Part 2 of this laboratory exercise you will compare the reactivity of magnesium, aluminum, and iron in a weak hydrochloric acid solution.

Objectives

- Observe an oxidation-reduction reaction.

- Use the properties of the product to verify what it is.
- To rank the reactivity of certain metals in a weak acid, and compare it to their order in the Activity Series of Metals.

Date:________ Name:______________________ Period:_________

Pre-lab Questions

1. What is the oxidation state for each atom in the following reaction:

$$4Fe + 3O_2 \longrightarrow 2Fe_2O_3$$

a) Elemental iron (Fe) _____**0**_____________

b) Elemental oxygen (O_2) _____**0**___________

c) One iron atom in Fe_2O_3 ________**+3**_______

d) One oxygen atom in Fe_2O_3 _____**-2**________

2. Which element was oxidized and which element was reduced in the above reaction equation?

a) Element oxidized _____**Fe**_____________

b) Element reduced ____**O**______________

3. From the Activity Series of Metals, determine the order of reactivity of the following metals: Ni, Au, Fe, Ca, Zn, and Al.

Most Reactive → Least Reactive

Ca → Al → Zn → Fe → Ni → Au

Preparation Materials

General Supplies	Laboratory Equipment	Chemicals
Aluminum foil	1 large test tube (30 x 150 mm)	3 M Iron (III) chloride ($FeCl_3$) Prep: 8.0 g $FeCl_3$ per 100 mL water
Magnet (stir bar retriever)	4 medium sized test tubes (18 x 150 mm)	Distilled water
Al screws, rivets or washers	Test tube rack and test tube holder	Mg ribbon (~3 cm per group)
Iron nails, screws, or washers that that are not galvanized	10 mL graduated cylinder	3 M HCl (3 mL per group) Prep: 24.90 mL HCl (37%) per 100 mL H_2O
Sand Paper	600 mL beaker	
	Hot plate	
	Sand paper or triangular files	

Procedure

SAFETY FEATURES

Safety Goggles

MATERIALS NEEDED

- 3M Iron (III) Chloride Solution ($FeCl_3$)
- 10 mL Graduated Cylinder
- Large test tube
- 4 Medium test tubes
- Test tube rack
- Test tube holder
- Aluminum foil
- 600 mL beaker
- Hot plate
- Magnet
- Mg ribbon
- Aluminum screw rivet or washer
- Iron screw or washer
- 3 M HCl
- Sand Paper
- Baking soda

Part 1. Iron Test for Magnetism

1. Prepare a hot water bath by filling a 600 mL beaker two-thirds full with distilled water and heating it on a hotplate.

2. Pour 5 mL of the 3 M $FeCl_3$ solution into a large test tube.

3. Check to see if there is anything magnetic in the solution by first stirring a magnet in the solution. Remove the magnet and record your observations.

4. Tear a 2 cm by 2 cm square of aluminum foil into 5 strips, so that the aluminum foil will easily fit into the test tube.

5. Test the aluminum to see if they are magnetic by bringing a magnet toward the aluminum. Record all observations.

6. Drop the aluminum foil pieces into the test tube. Use a stirring rod to submerge them into the solution.

7. Use a test tube holder to place the test tube into the hot water bath.

8. In the data section, record the time that heating began.

9. Let the mixture heat until a color change is noted. This should take 10-15 minutes.

10. Continue to Part 2 to use lab time efficiently. Complete Part 1 after you finish Part 2

11. Record observations of the reaction.

12. Record the time that the reaction was completed.

13. Use a test tube holder to very carefully remove the test tube from the hot water bath and place it in the test tube rack. **CAUTION: Be attentive when transporting hot items. Spills and burns could occur.**

14. Let the solution cool for a few minutes.

15. Check to see if there is anything magnetic in the solution by first stirring a magnet in the solution. Remove the magnet and record your observations.

Part 2. Observation of the Reactivity Series of Metals

> ***Teaching Alternative:***
> ***If calcium or sodium metal is available, each can be demonstrated using water as the solvent. Both are high in the reactivity series and will react with water as readily as the other metals do with hydrochloric acid.***

1. Label the 3 test tubes with **Mg**, **Al**, and **Fe** and place them in a test tube rack.

2. Place ~2 cm of Mg ribbon in the test tube labeled **Mg**.

3. Lightly sand an aluminum item and an iron item that is not galvanized to remove any coating and place them in the appropriate test tubes.

4. Add ~1 mL of 3M HCl to each test tube.

5. Make initial observations and continue recording observations after one minute, 3 minutes, and 5 minutes. **HINT**: The most notable observations are how quickly bubbling occurs and how violently the bubbling of each continues.

6. After observing the reactions for 5 minutes, rank the metals in order of their reactivity. Compare your results with the actual reactivity series.

7. For clean-up, decant the acid into a waste beaker. Neutralize the waste solution with baking soda and pour it down the drain. Rinse the metals left with water. Dry the metals and place them in the designated container for later use.

Date:________ Name: __________________________ Period:_________

Lab Partner: _________________________________

Data

Part 1.

- Observations of magnetic test for $FeCl_3$ solution before the aluminum was added:

  No magnetic attraction (negative test)

- Observations of magnetic test for aluminum:
  No magnetic attraction (negative test)

- Observations of reaction:

  During the reaction a color change of the solution occurred.

Table 1.

Time reaction began	Time reaction ended
12:00	***12:10***
Initial color of solution	Final color of solution
Rusty dark brown	***Dull light brown***

- Observations of magnetic test for the product after the solution was heated:
  The product was attracted to the magnet. It was a positive magnetic test.

Part 2.

Table 2.

Time (minutes)	Observations
Initial	***All 3 metals are bubbling.***
1	***Mg ribbon is reacting more violent than the others.***
3	***Al chip is bubbling more rapidly than the Fe.***
5	***Bubbles are being produced by each metal, but Mg is the most violent followed by the Al.***

Summarize observations and list the order of reactivity of the metals that you observed:

The Mg ribbon reacted the most violent. While the Al chip produced more bubbles more rapidly than the Fe sample.

Mg → Al → Fe

Post Lab Questions

1. Based on what you observed, what product was formed in Part 1? How do you know?

   Iron was produced. This was concluded from the magnetism test. The product was magnetic.

2. Did the order of reactivity you determined in Part 2 match the order given in the Activity Series of Metals? Explain.

   Mg → Al → Fe

   The order in which the metals reacted matched the order given in the Activity Series of Metal. This conclusion is based on the amount of bubbles or gas produced and the rate the bubbles were produced.

3. Do you think the following reaction would occur? Explain your answer.

   $$FeCl_{2(aq)} + Cu_{(s)} \longrightarrow Fe_{(s)} + CuCl_{2(aq)}$$

   No, this reaction would not occur. Copper is lower in the activity series of metals.

Lab

16

The Mole: Avogadro's Number

*TN Standard **3.3:** The student will explore the mathematics of chemical formulas and equations.*

How many avocados and artichokes are there in a mole: "avocados number"?

A recipe calls for two avocados and two artichokes. Would you say that equal amounts of avocados and artichokes were used? The answer to that question is relative to what you may consider "the same amount". If you consider the quantity, yes, there were two of each. What if instead the recipe told you to use 250 grams of artichokes and avocados? Would you say the same amount of avocados as artichokes were used? There might be one and a half artichokes used for every two avocados, but there is still the same amount of avocados and artichokes.

Introduction

In order to solve this problem one must be specific in the instructions. The term "the same amount" is being used in one sense to compare a discrete amount of the avocados and artichokes and in another sense to compare their mass. This is the same as it is with atoms and molecules. If you are suppose to put the same amount by mass of two different molecules into a beaker, all you would have to do is weigh the same amount out of each of the substances and then transfer them into the beaker.

The term "same amount" becomes a more complex issue when comparing things that are too small to count, such as molecules, and you have to put the same amount by number into a beaker. Scientists have used the term "mole" as a unit equal to 602,214,199,000,000,000,000,000 or 6.02×10^{23} items. This can be of anything, but it is usually referring to molecules or atoms. A mole is a term used similar to a dozen. Just like there are 12 eggs in a dozen, there are 6.02×10^{23} molecules in a mole.

An experiment calls for 1 mole of NaCl, so all you have to do is count out 6.02×10^{23} molecules and you're good to go! Not quite, it would take you 200 trillion life times if you lived to be a 100 years old each life time and counted 1 particle a second! The earth's age is not even close to 200 trillion life times old. Needless to say there has to be a better way to count molecules. Scientist found the accepted amount of grams per each mole of an element. This information is called the atomic weight and it is conveniently placed under the symbol of each element in the periodic chart.

So how do you measure 1 mole of NaCl? A molecule of salt is made up of one sodium atom and one chloride atom. You look up sodium as Na on the Periodic Table and find its atomic weight is 22.99 amu. This means you would have to weigh out 22.99 grams of sodium to get one mole of sodium. This is 6.02×10^{23} molecules of sodium. But to get 1 mole of NaCl you also need to look up the atomic weight of Cl on the Periodic Table. It is 35.45 amu. This is how many grams of chloride you would have to weigh out to get one mole of chloride atoms. Since one mole of salt contains one mole of sodium atoms and one mole of chloride atoms you add these two masses.

1 Na = 22.99 g/mole
+ 1 Cl = 35.45 g/mole
58.44 g/mole of NaCl
This is how to determine the **Molar Mass**

$$1 \text{ mole NaCl} \times \frac{58.44 \text{ grams NaCl}}{1 \text{ mole NaCl}} = \mathbf{58.44 \text{ grams NaCl}}$$

What if we weighed out 1.00 grams of NaCl. How many molecules is this? We know how many grams there are in a mole, but grams are NOT equal to molecules! How do we determine the number of molecules in a gram of salt? For this we also use the molar mass of salt only we flip it and say 1 mole of NaCl per 58.44 g of NaCl.

$$1.00 \text{ g NaCl} \times \frac{1 \text{ mole NaCl}}{58.44 \text{ g NaCl}} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mole NaCl}} = \mathbf{1.03 \times 10^{22} \text{ molecules NaCl}}$$

Here are some constants that can be helpful!
6.02×10^{23} molecules = 1 mole = 1 molar mass (g/mol)

Through this lab procedure, we will determine the experimental value for Avogadro's number. Cinnamon will be floating, evenly distributed, on the surface of some water in a Petri dish. Joy dishwashing liquid is about 1% sodium stearate and a solution of it with a known concentration will be dropped onto the water. The sodium stearate molecules will then line up in a single layer and will spread out, pushing the cinnamon toward the edges of the Petri dish, allowing the surface area to be determined. We will assume that each molecule takes up 0.21 nm^2 of surface area, and that there is no space between the molecules.

Objectives

- To understand the importance of Avogadro's number.
- To formulate an approximate estimate of Avogadro's number.

Date:________ Name:______________________ Period:_________

Pre-lab Questions

1. How many pennies are there in one mole of pennies?

   There are 6.02×10^{23} pennies in one mole.

2. How many grams of H_2O do you need to weight out to have 1 mole of H_2O?

   2 H = 2 x 1.01 = 2.02

   + 1 O = 1 x 16.00 = 16.00

   18.02 g / mole H_2O

3. How many molecules of water are there in one mole of H_2O?

   There are 6.02×10^{23} molecules in one mole of H_2O.

4. How many moles of H_2O are there in 1.0 g of H_2O?

   $$1.0 \text{ g } H_2O \times \frac{1 \text{ mole } H_2O}{18.02 \text{ g } H_2O} = 0.055 \text{ mole } H_2O$$

5. How many molecules of H_2O are there in 1.0 g of H_2O?

$$0.055 \text{ mole } H_2O \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mole } H_2O} = 3.3 \times 10^{22} \text{ molecules } H_2O$$

Preparation Materials

General Supplies	Laboratory Equipment
Ruler	100 mL graduated cylinder
Ground Cinnamon	10 mL graduated cylinder
Joy dishwashing liquid	Petri dish (bottom only)
	Glass dropper
	50 mL beaker
	Stirring Rod

Procedure

SAFETY FEATURES

Safety Goggles

MATERIALS NEEDED

- Glass dropper
- Petri dish (bottom)
- Dilute Joy dishwashing liquid
- Ruler
- 100 mL graduated cylinder
- Ground cinnamon
- 10 mL graduated cylinder
- Stirring rod
- 50 mL beaker

Part 1: Preparing the Sodium Stearate Solution

1. Measure exactly 1.50 ml of Joy dishwashing liquid into a 10 mL graduated cylinder.

2. Fill a wash bottle with distilled water. Gently rinse the 1.50 mL of dishwashing liquid with distilled water and pour it into a 100 mL graduated cylinder. Gently rinse the 10 mL graduated cylinder several times to make sure all the dishwashing liquid has been transferred to the 100 mL graduated cylinder. **HINT:** Try not to create suds.

3. Add enough additional distilled water to get to 100 mL.

4. Gently stir the solution with a stirring rod until it is mixed well.

Part 2: Calibrating a Dropper

1. Fill a 50 mL beaker approximately half full with water. Use your dropper to fill a 10 mL graduated cylinder to 1.00 mL with water. **HINT:** Make sure the 10 mL graduated cylinder is clean of dishwashing liquid.

2. Next, draw up water from the 50 mL beaker into the dropper. Count the drops, while you add water dropwise until you reach the 2.00 mL mark. Hold the dropper consistently at a 45^{o} angle and drop at a rate of about one drop per second. **HINT**: It should be ~25 drops. If you feel that your measurement is incorrect, repeat as many times as necessary to achieve consistent readings. Then take the average of the readings.

3. Record the number of the drops in the data section.

4. Repeat calibration for a second trial. Record the number of drops in the data section. Average the two results and record this value.

Part 3: Calculating the Number of Molecules

1. Rinse and then fill a Petri dish ¾ full with distilled water. Allow the water to settle and remain motionless.

2. Lightly sprinkle cinnamon onto the surface of the water in the Petri dish. **HINT:** Add just enough to barely cover the water.

3. Draw up the dishwashing liquid solution with the calibrated dropper. Hold the dropper at a 45^{o} angle over the center of the Petri dish. Slowly deliver one drop of the solution. **HINT:** A clear circle should form, spreading the cinnamon outward.

4. Quickly use a ruler to measure the diameter of the cleared circle in cm.

5. Record the diameter in the data section. Wash out the Petri dish.

Date:________ Name: __________________________ Period:_________

Lab Partner: ___________________________________

Data

Part 2: Calibrating a Dropper

1. The number of drops to reach from the 1.00 mL to the 2.00 mL mark:

   Trial 1 _____25_____ Trial 2 _____25_____

2. The number of drops on average per one milliliter:____25____

Part 3: Calculating the Number of Molecules

1. The diameter of the circle formed: ____6.00 cm_____

Calculations

1. Calculate the surface area of the circle formed ($\pi d^2/4$) :

$$\frac{(3.14) * (\text{diameter})^2}{4} = \underline{28.3\ cm^2}$$

2. Calculate the number of molecules on the top layer. We must convert the surface area in centimeters squared to nanometers squared and then multiply that by the surface area of a sodium stearate molecule.

   Convert the surface area of the circle formed (#1) to molecules per layer:

_28.3_ cm^2	1 m^2	1 x 10^{18} nm^2	1 molecule
Top layer (from #1 Calculations)	10 000 cm^2	1m^2	0.210 nm^2

= __1.35 x 10^{16}_ $\frac{\text{molecules}}{\text{top layer}}$

3. Calculate the concentration of grams of sodium stearate per milliliter of diluted solution. To do this, multiply the concentration of sodium stearate in the Joy dishwashing liquid (1 g sodium stearate/100 mL Joy) by the dilution of the solution.

1 g sodium stearate	1.50 mL Joy	
100 mL Joy	100 mL Joy diluted solution	= **1.50 x 10^{-4}** g sodium stearate / mL Joy diluted solution

4. Calculate the number of moles of sodium stearate in the single layer. To do this, first take the number of drops used to achieve the monolayer (1 drop) and convert it to mL using the calibrated number of drops per mL. Then multiply the number of grams of sodium stearate per milliliter of solution. Finally, convert to moles through the molar mass of sodium stearate. **HINT**: The molar mass of sodium stearate is 296.4 g/mol.

1 drop (# of drops added to the Petri dish)	1 mL solution	**1.50 x 10^{-4}** g sodium stearate (from #3, Calculations)	1 mol
Top layer	**25** avg. # calibrated drops (from #2, Part 2 Data)	1 mL Joy diluted solution	296.4 g (molar mass of sodium stearate)

= **2.02 x 10^{-8}** mol/Top Layer

5. Finally, we can calculate the Avogadro's number through the comparison of molecules of sodium stearate in the top single layer to the moles of sodium stearate in the monolayer.

Avogadro's number (experimental) = $\frac{1.35 \times 10^{16} \text{ molecules(from \#2)}}{2.02 \times 10^{-8} \text{ mole (from \#4)}}$

Avogadro's number (experimental) = 6.68×10^{23} molecules / mole

This number will probably be somewhat different. *If one gets 10^{23} it is considered a good result!

Post Lab Questions

1. **1.** Why do you think that Avogadro's exact number, 6.02×10^{23}, was probably not the number you obtained?

   Results may not have been exact due to several factors including the calibration of the dropper, miss counting during the experiment, inaccurate measurements, poor laboratory technique, and/or rounding of numbers during calculations.

2. **2.** How many moles is 0.289 g of methane (CH_4)?

1 C = 12.01 g/mole

+ 4 H = 4.03 g/mole

Molar mass = 16.04 g/mole of CH_4

(0.289 g CH_4) (1mole CH_4 /16.04 g CH_4) = 0.0180 mole of CH_4

3. How many moles are 1,000,000,000 molecules of H_2O_2?

Moles H_2O_2 = (1.00 x 10^9 molecules)(1 mol/6.02 x 10^{23} molecules)

=1.66 x 10^{-15} moles

4. How many grams are 1,000,000,000 molecules of $H_2 0_2$?

2 O = 32.00 g/mole
<u>+ 2 H = 2.02 g/mole</u>
Molar mass = 34.02 g/mole of $H_2 0_2$

Grams H_2O_2 = 1.66 x 10^{-15} moles $H_2 0_2$ x 34.02 g/mole of $H_2 0_2$

= 5.65 x 10^{-14} g $H_2 0_2$

Lab

17

The Periodic Table: Its Trends and Uses

*TN Standard **1.2**: Investigate the basic organization of the modern periodic table, including atomic number and atomic properties.*

What is the importance of an element's location in the Periodic Table?

The periodic table is a keeper of infinite facts of the science of chemistry. It places every individual element in a specific location on the chart. Its organization allows one who knows how to interpret it access to a wealth of knowledge about each element. Without the periodic table, chemists would have to memorize a lot of details and facts. Now memorizing is very difficult and not a lot of fun. But chemistry is fun! So, let's blow up some balloons to show how the periodic table can be a party.

Introduction

The periodic table is organized similar to a typical data table. It contains individual cells that are arranged logically in rows and columns according to specific trends. The rows on the periodic table are called periods, while the columns are referred to as families or groups.

Each cell contains information about a particular element. In the middle of the cell is a one or two letter abbreviation for a particular element. For example Fe is the abbreviation for iron and Na stands for sodium. The abbreviation is called the element's symbol. Above an element's symbol is the atomic number. It is a whole number that tells us how many protons are in the element. Each element has its own unique number of protons. Below an element's symbol is the atomic weight in atomic mass units (amu). It tells us how many grams of that element must be weighed out to have a mole or 6.02×10^{23} atoms of it. For example, the atomic weight of sodium is 22.99 amu. If we weigh out 22.99 grams of sodium we will have 6.02×10^{23} atoms or one mole of sodium. Because of this, the units are often given as grams/mole instead of amu.

The elements are arranged in such a way that both chemical and physical trends run across the rows and down the columns. Trends found in the periodic table can demonstrate uniform increases or decreases or represent similarity. As you go across a period moving left to right, you will find that the atomic number of the elements increase by one. This means the number of protons in the elements increase by one. On the other hand when moving down a group of the periodic table you will find that the number of electrons in the outer shell remains the **same**. This is why the atoms within a group have similar chemical behavior.

Atoms are often bonded together to form compounds. The atomic weight of each atom in a compound can be added together to find the mass of one mole of that compound. For example, to find the mass of one mole of Li_2CO_3 you would add the mass of two lithium atoms, one carbon atom, and 3 oxygen atoms.

2 Li = 2 x 6.94 = 13.88 g/mole

1 C = 1 x 12.01 = 12.01 g/mole

3 O = 3 x 16.00 = 48.00 g/mole

13.88 g/mole + 12.01 g/mole + 48.00 g/mole = 73.89 g/mole

This means if you weigh out 73.89 gram of Li_2CO_3, you would have one mole (or 6.02×10^{23} molecules) of Li_2CO_3. But in this experiment you only want 0.025 moles of Li_2CO_3. To find out how many grams you need to weigh out to have 0.025 moles of Li_2CO_3, you simply multiply the molecular weight of lithium carbonate by 0.025 moles.

73.89 g/~~mole~~ x 0.0250 ~~moles~~ = 1.847 g Li_2CO_3

Notice that the moles cancel out leaving grams.

In this experiment you will investigate Group I elements by reacting their carbonates with HCl to produce water, a salt, and carbon dioxide gas (CO_2). The CO_2 gas will be captured by a balloon which will cause the balloon to inflate. The reaction for the lithium carbonate is shown below:

$$Li_2CO_3 + 2\,HCl \rightarrow 2LiCl + H_2O + CO_2$$

Since you are using only Group I carbonates, the compounds are similar and will react in a similar manner. For instance the carbonates all have the same ratio of two of the Group I element to one CO_3. Also they will all form water, a salt, and carbon dioxide gas when HCl is added to them.

You will first compare the reactions when the same fraction of a mole of Li_2CO_3, Na_2CO_3, and K_2CO_3 react with the same amount of HCl solution. You will then compare the reactions when the same amount of grams of Li_2CO_3 and K_2CO_3 react with the same amount of HCl solution.

Objectives

- To understand the periodic chart and its uses.
- To relate the characteristics **within** a group on the periodic chart.
- To understand the use of molecular weights and molar ratios in chemical reactions.

Date:________ Name:______________________ Period:_________

Pre-lab Questions

1. Complete the three reactions shown below that are taking place throughout this experiment.

a. Li_2CO_3 + 2 HCl → ***2LiCl + H_2O + CO_2***

b. Na_2CO_3 + 2 HCl → ***2NaCl + H_2O + CO_2***

c. K_2CO_3 + 2 HCl → ***2KCl + H_2O + CO_2***

2. Calculate the molecular weights of Na_2CO_3 and K_2CO_3. (See the Introduction for the Li_2CO_3 example.)

Na_2CO_3:

2 Na = 2 x 22.99 = 45.98 g/mole

1 C = 1 x 12.01 = 12.01 g/mole

3 O = 3 x 16.00 = 48.00 g/mole

45.98 + 12.01 + 48.00 = 105.99 g/mole

K_2CO_3:

2 K = 2 x 39.10 = 78.20 g/mole

1 C = 1 x 12.01 = 12.01 g/mole

3 O = 3 x 16.00 = 48.00 g/mole

78.20 + 12.01 + 48.00 = 138.21 g/mole

3. Calculate how many grams of Na_2CO_3 and K_2CO_3 you will need to weigh out to have 0.0250 mol of each of the substances. (See the Introduction for the Li_2CO_3 example.)

Na_2CO_3:

105.99 g/~~mole~~ x 0.0250 ~~moles~~ = 2.650 g Na_2CO_3

K_2CO_3:

138.21 g/~~mole~~ x 0.0250 ~~moles~~ = 3.455 g K_2CO_3

Preparation Materials

General Supplies	Laboratory Equipment	Chemicals
4 Balloons (per group)	4 - 125 mL Erlenmeyer flask (per group)	Li_2CO_3 Lithium carbonate (Do not use hydrate.)
Sharpie Marker	100 mL graduated cylinder	Na_2CO_3 Sodium carbonate (Do not use hydrate.)
~50 cm string (per group)	Balance	K_2CO_3 Potassium carbonate (Do not use hydrate.)
Ruler	Spatula	~200 mL 1*M* HCl Hydrochloric acid (per group) Prep: 83 mL concentrated HCl (37%) per 1 L H_2O
Baking soda		

For a greener alternative 20 mL of saturated citric acid solution can be substituted for each 50 mL of 1 *M* HCl. An example of how to change the reaction equation is:

$$3\ Li_2CO_3 + 2\ H_3C_6H_5O_7 \rightarrow 2\ Li_3C_6H_5O_7 + 3\ H_2O + 3\ CO_2$$

Procedure

SAFETY FEATURES

- Safety Gloves
- Measure Accurately
- Safety Goggles

MATERIALS NEEDED

- 4 Balloons
- 1M HCl, Hydrochloric Acid
- Lithium Carbonate Li_2CO_3
- Sodium Carbonate, Na_2CO_3
- Potassium Carbonate, K_2CO_3
- 4 - 125 mL Erlenmeyer Flasks
- Sharpie Marker
- Balance
- 100 mL Graduated Cylinder
- Spatula
- String (~ 50 cm)
- Ruler

Helpful HINT:* *The carbonates may need to be dried in an oven overnight and stored in a desiccator (if possible) before they are used.

1. Fill in #1 in the Data and Calculations section. Be sure to have your teacher check your masses before weighing out them out.

2. Blow up 4 balloons to approximately the same size and then let the air out. Label each balloon with a Sharpie marker. Label the first one Li_2CO_3, the second Na_2CO_3, the third K_2CO_3, and the fourth 2nd K_2CO_3.

3. Place the calculated number of grams of 0.025 moles of Li_2CO_3, Na_2CO_3, and K_2CO_3 in the appropriate balloons. Place 1.847 g of K_2CO_3 into the balloon marked 2nd K_2CO_3 . **HINT:** Place the entire balloon on the balance and tare (zero) it. Add the substance into each balloon until the correct amount is placed in balloon. Make sure the entire balloon is on the balance pan. **HINT:** If using the same spatula for transferring different chemicals, be sure to clean the spatula before placing it into a different substance.

4. Pour 50 mL of 1M HCl into each of the four 125 mL Erlenmeyer flasks. **CAUTION: HCL is corrosive and a respiratory irritant.**

5. Attach the four balloons to the Erlenmeyer flasks without turning the balloons upright or allowing any of the powder to fall into the flasks. One student should securely hold the flask upright, while his or her partner attaches the balloon. **CAUTION: Remember safety goggles! Make sure the balloons are well secured to the flask. If not, a balloon could fly off and cause injury**.

6. Turn only one balloon upright at a time. Gently shake the balloon until all of the carbonate falls into the flask. Record your observations of the reaction in Data, Table 1.

7. After each reaction is complete, quickly measure the circumference of the largest part of the balloon. Do this by using a string to determine the distance around the largest part of the balloon, and then measure the corresponding length of the string with a ruler. Record the length in Data, Table 1.

8. Repeat steps 6 and 7 with each flask.

Helpful HINT:
To dispose of the student's waste, add a small piece of pH paper to the waste beaker. Set the beaker in the sink. Stir in small amounts of baking soda until the pH is between 4 and 8. It is then safe pour down the drain.

9. To clean up, first remove the balloons from the flasks and throw them in the trash. **CAUTION: When removing the balloons avoid direct inhalation of gases. Hold the balloons away from your or another student's face and body. Remove the lip of the balloon that is opposite from your body first.** Place all reaction solutions in a designated waste beaker.

Date:________ Name: ___________________________ Period:_________

Lab Partner: ____________________________________

Data and Calculations

1. Fill in the masses of 0.0250 moles of each of the carbonates you are going to use in the appropriate spaces below. These are found in the Introduction or were calculated in the pre-lab section. Check these masses with your teacher before you weigh out the material.

Mass of Li_2CO_3: ___*1.847 g*___

Mass of Na_2CO_3: ___*2.650 g*___

Mass of K_2CO_3: ___*3.455 g*___

2. Calculate the number of moles there are in the 1.847 g sample of K_2CO_3 and record this value in the Data Table.

$$1.847\ \cancel{g\ K_2CO_3} \times \frac{1\ mol\ K_2CO_3}{138.21\ \cancel{g\ K_2CO_3}} = \underline{\quad 0.0134 \quad}\ mol\ K_2CO_3$$

3. Complete the following Data Table:

Table 1: Observations and Data for the Carbonates Reactions with HCl

Substance	Moles of Substance	Grams of Substance actually used	Observations of the Reaction	Inflated Balloon Size
Li_2CO_3	0.025 mol	*1.849 g*	*During the reaction the solution is white and bubbling. Afterwards the solution is clear.*	*~ 31.0 cm*
Na_2CO_3	0.0250 mol	*2.649 g*	*During the reaction the solution is white and bubbling. Afterwards the solution is clear.*	*~ 31.0 cm*
K_2CO_3	0.0250 mol	*3.457 g*	*During the reaction the solution is white and bubbling. Afterwards the solution is clear.*	*~ 30.0 cm*
2nd K_2CO_3 (1.847 g)	*0.0134 mol*	*1.850 g*	*During the reaction the solution is white and bubbling. Afterwards the solution is clear.*	*~ 23.0 cm*

Post Lab Questions

1. What was the source of the gas that causes the balloons to inflate?

   The source was carbon dioxide (CO_2) gas that formed during the reaction.

2. Did a trend of similarity appear in the data and/or observations among the **three** reactions with 0.0250 moles of each substance?

   Yes, the measurements from the first three flasks were relatively close, differing by less than five centimeters. The observations of the reactions were the same.

3. Considering the organization of the periodic table, explain why there were similarities among the reactions?

   Lithium, sodium, and potassium are all members of Group I of the Periodic Table. This means they would all react with HCl in a similar manner.

4. In the experiment you used the same number of moles of each of the first three salt carbonates. Explain what happened when you put the same amount of **grams** of the potassium carbonate in the fourth balloon as you used of the lithium carbonate.

   The fourth balloon had a circumference that was much less than the first three. The same amount of grams does not equal the same amount of molecules. 1.847 g of potassium carbonate is just a little over half the number of moles (or molecules) of the carbonates used in the first three balloons. This meant less CO_2 was formed so the balloon did not inflate as much as the balloons in the previous reactions.

Lab

18

Stoichiometry: Synthesis of Garden Lime

TN Standard ***3.3:*** *The student will explore the mathematics of chemical formulas and equations.*

Have you ever wondered why hot dogs are sold in packages of 10, but hotdog buns are sold in packages of 8?

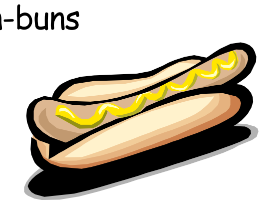
This is an example of a ratio. ***Stoichiometry*** is the scientific word for ratios of moles. It is the study of the quantitative relationships between reactants and products in a chemical reaction. These relationships are often referred to as mole-to-mole ratios. Mole-to-mole ratios help scientists know how much of each reactant is needed to produce a sufficient amount of product. In this lab, you will observe the importance of mole-to-mole ratios when making the main component in lime that is put on a garden!

Introduction

The manufacturers of hotdogs and hotdog buns would not make very good chemists. Hotdogs are sold in packages of 10 where hotdog buns are sold in packages of only 8. If one package of each is purchased, how many hotdog-on-buns can you make? The answer of course is 8. This means there are 2 hotdogs left over. Now examine the equation below and try to fix their mistake. How many packages of hotdogs and hotdog buns must you purchase in order to not waste any hotdogs? How many hotdog-on-buns does this give you?

____10 Hotdogs + ____ 8 Hotdog buns → ____ Hotdog-on-buns

If you buy 4 packages of hotdogs and 5 packages of buns you would get 40 hotdog-on-buns.



Balancing a chemical equation is done the same way. To balance a chemical equation you place numbers, called coefficients, in front of each type of molecule. Just like you cannot change how many hot dogs there are in a package of hotdogs, you cannot change how many of each type of atom there is in a chemical substance. You can however change the ratio of the chemical substances so that the same number of each type of atom is on both sides of the reaction equation. In the hotdog example there are 40 hotdogs and 40 hotdog buns on both sides of the arrow.

But what if instead you buy only 2 packages of hotdogs, and 3 packages of hotdog buns, which would you run out of first? You would have a total of 20 hotdogs and 24 hotdog buns so you would run out of the hotdogs first. This means you could only make 20 of the hotdog-on-buns final product. The number of hotdogs limits the number of hotdog-on-buns you can make.

This is similar to what in chemistry is called the **limiting reagent**. By using stoichiometry it is possible to calculate the limiting reagent of a reaction. Knowing which reactant will be used up first allows chemist to calculate how much product can theoretically be made. This is often compared to the amount that is actually made through calculating what is called the percent yield. The percent yield is found by using the following equation:

$$\text{Percent Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

Using the hotdog example, the percent yield would be like only making 19 hotdogs when there were enough hotdogs and buns to make 20 hotdog-on-buns. In this example the percent yield would be calculated as follows:

$$\text{Percent Yield} = \frac{\text{19 Actual Hotdog-on-buns}}{\text{20 Theoretical Hotdog-on-buns}} \times 100$$

$$\text{Percent Yield} = 95\%$$

Chemist will often try making a desired product many different ways. They will then compare the percent yields to determine which of the methods tried gives the greatest yield.

In this laboratory exercise you will explore the concepts of limiting reagents and percent yields through making calcium carbonate, the main component in lime. Lime is used in many different industries, including farming. Farmers spread it on soil to adjust its pH so that their crops will be more productive.

Calcium carbonate, $CaCO_3$, can be made by mixing $CaCl_2$ and K_2CO_3 in water. The trick is to be sure to use the correct amount of each. For this, stoichiometry is used. Look at the following reaction and balance it to determine the mole-to-mole ratio that would be necessary to have the same number of each kind of atom on both sides of the equation.

$$\_\_\_\_ CaCl_{2(aq)} + \_\_\_\_\_ K_2CO_{3(aq)} \rightarrow \_\_\_\_\_ KCl_{(aq)} + \_\_\_\_\_ CaCO_{3(s)}$$

When this chemical equation is balanced there is one mole of calcium chloride for every 1 mole of potassium carbonate and together these quantities make 2 moles of potassium chloride and one mole of calcium carbonate. The potassium chloride is soluble in water so it remains in solution. The calcium carbonate will precipitate out as a white solid.

$$CaCl_{2(aq)} + K_2CO_{3(aq)} \rightarrow 2KCl_{(aq)} + CaCO_{3(s)}$$

What if instead you added 2 moles of $CaCl_2$ and 1 mole of K_2CO_3. How much $CaCO_3$ could you make? The answer is still only one mole of $CaCO_3$. The K_2CO_3 is the limiting reagent since it limits how much of the product can be made.

The following calculation illustrates how the limiting reagent and percent yield are determined for the above reaction. Determination of the limiting reagent can be complex. Although this is often performed in a single combined step, it can also be broken into several steps. In this example it is broken into several steps for simplification.

Example: A procedure calls for 8.25 g of $CaCl_2$ to be dissolved in water and reacted with 5.00 g of K_2CO_3. How many grams of $CaCO_3$ can theoretically be made? If 3.25 g of $CaCO_3$ were actually made, what is the percent yield?

a. First write the balanced reaction equation:

$$CaCl_2 + K_2CO_3 \rightarrow 2KCl + CaCO_3$$

b. Calculate the formula weight for each reactant:

	2 K = 78.20 g
Ca = 40.08 g	3 O = 48.00 g
+ 2 Cl = 70.90 g	+ C = 12.01 g
110.98 g $CaCl_2$	138.21 g K_2CO_3

c. Calculate the number of moles used of each reactant:

$CaCl_2$: (8.25 ~~g~~)(1 mol/110.98 ~~g~~) = 0.0743 mol

K_2CO_3: (5.00 ~~g~~)(1 mol/138.21 ~~g~~) = 0.0362 mol

d. Determine how many moles of the product could be made by each of the reactants.

$CaCl_2$: (0.0743 ~~mol $CaCl_2$~~) $\frac{(1 \text{ mol } CaCO_3)}{(1 \text{ mol } CaCl_2)}$ = 0.0743 mol $CaCO_3$

K_2CO_3: (0.0362 ~~mol K_2CO_3~~) $\frac{(1 \text{ mol } CaCO_3)}{(1 \text{ mol } K_2CO_3)}$ = 0.0362 mol $CaCO_3$

e. Determine the limiting reagent. It is the reactant that will make the least number of moles of $CaCO_3$.

K_2CO_3 is limiting!

This means that only 0.0362 moles of $CaCO_3$ can be made when 8.25 g of $CaCl_2$ and 5.00 g of K_2CO_3 are mixed in water. In this example, the reactant that had the lowest mass in the procedure turned out to be limiting. However, this is not always true! The only way to know for sure which reactant is limiting is to calculate it.

f. Determine how many grams of $CaCO_3$ can theoretically be produced:

First calculate the formula weight for the product, $CaCO_3$:

Ca = 40.08 g/mol
3 O = 48.00 g/mol
+ C = 12.01 g/mol
100.09 g $CaCO_3$ / mol

Next calculate the theoretical yield by calculating how many grams there are in 0.0362 moles of $CaCO_3$:

Theoretical yield = (0.0362 ~~mol $CaCO_3$~~)(100.09g $CaCO_3$/ ~~mol~~)

= 3.62 g $CaCO_3$

g. Calculate the percent yield:

$$\text{Percent Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

$$= \frac{3.25\ g}{3.62\ g} \times 100$$

$$\text{Percent Yield} = 89.8\%$$

This means that in this example only 89.8% of what theoretically could have been made was actually made.

Objectives

- Demonstrate the use of stoichiometry to synthesize calcium carbonate.
- Practice using a balance and proper lab techniques.
- Find the limiting reagent, the theoretical yield, and the percent yield.

Date:________ Name:_______________________ Period:_________

Pre-lab Questions

1. What is a limiting reagent?

A limiting reagent is the reactant that will be used up first in the reaction, and therefore limits the theoretical amount of product that can be made.

2. A student used 7.15 g of $CaCl_2$ and 9.25 g of K_2CO_3 to make $CaCO_3$. The actual yield was 6.15 g of $CaCO_3$. Calculate the limiting reagent and the percent yield.

$$CaCl_2: (7.15\ \cancel{g})\left(\frac{1\ \cancel{mol\ CaCl_2}}{110.98\ \cancel{g}}\right)\left(\frac{1 mol\ CaCO_3}{1\ \cancel{mol\ CaCl_2}}\right) = 0.0644\ mol\ CaCO_3$$

$$K_2CO_3: (9.25\ \cancel{g})\left(\frac{1\ \cancel{mol\ K_2CO_3}}{138.21\ \cancel{g}}\right)\left(\frac{1\ mol\ CaCO_3}{1\ \cancel{mol\ K_2CO_3}}\right) = 0.0669\ mol\ CaCO_3$$

$CaCl_2$ is limiting!

$$\text{Theoretical yield} = (0.0644\ \cancel{mol})(100.09\ g\ CaCO_3/\cancel{mol})$$

$$= 6.45\ g$$

$$\text{Percent yield} = \frac{6.15\ g}{6.45\ g} \times 100$$

$$= 95.3\%$$

Preparation Materials

Laboratory Equipment	Chemicals
100 mL graduated cylinder	Calcium chloride ($CaCl_2$) 1-3 g per group
250 mL beaker	Potassium carbonate

	(K_2CO_3) 2.5 g per group
250 mL Erlenmeyer flask	Distilled or deionized water
50 mL beaker	Ethanol (if available)
Watch glass	
Ring and ring stand	
Medium filter paper (Whatman 40)	
Funnel	
Wash bottle	
Stirring rod (with rubber policeman if available)	
Spatula	
Balance	

Procedure

SAFETY FEATURES

Glassware

Safety Goggles

1. Weigh into a 250 mL beaker the amount of calcium chloride ($CaCl_2$) **specified for your group** in the chart below. Record the exact mass you weigh out in the data section.

Group A	~1.00 g $CaCl_2$
Group B	~2.00 g $CaCl_2$
Group C	~3.00 g $CaCl_2$

MATERIALS NEEDED

- Calcium Chloride
- Potassium Carbonate
- Balances
- Distilled water
- 100 mL Graduated Cylinder
- Watch glass
- Filter paper
- Funnel
- 250 mL beaker
- Spatula
- Wash bottle
- 250 mL Erlenmeyer flask
- 50 mL beaker
- Stirring rod with rubber policeman
- Oven
- Ethanol
- Ring and Ring Stand

2. Measure 50 mL of distilled water into a 100 mL graduated cylinder. Pour the water into the 250 mL beaker with the calcium chloride.

3. Stir the solution with a stirring rod until all of the calcium chloride is dissolved.

4. Weigh out ~2.50 g of potassium carbonate (K_2CO_3) in a 50 mL beaker. Record the exact mass in the data section.

5. Measure 25 mL of distilled water into a 100 mL graduated cylinder. Add the water into the 50 mL beaker containing the potassium carbonate.

6. Stir the potassium carbonate in the distilled water with a stirring rod until it is all dissolved.

7. Pour the K_2CO_3 solution into the 250 mL beaker that has the $CaCl_2$ solution. Rinse the beaker that contained the K_2CO_3 with a few mL of water and add this to the $CaCl_2$ solution. Stir the mixture.

8. As soon as the reaction begins, record your observations in the data section. Continue stirring until you see no more precipitate forming.

9. Set up the filtering apparatus as shown in Figure 9.1. Secure the Erlenmeyer flask to a ring stand. Rest the funnel in it. **HINT:** Do NOT begin filtering yet!

10. Zero the balance and **weigh a piece of medium flow filter paper and a watch glass**. Record the masses of both items in the data section.

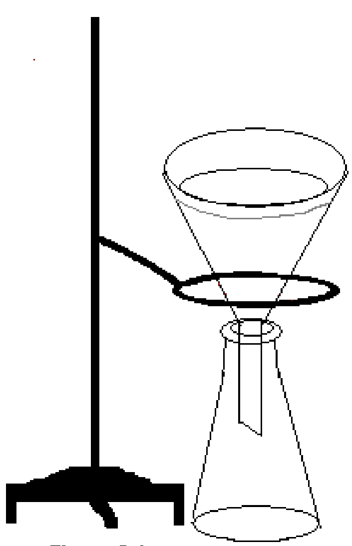


Figure 9.1

11. Fold the filter paper into quarters. Open the filter paper so that a funnel is formed. **HINT:** Three sides will be together and the other side will open out.

12. Place the paper funnel into the glass funnel and hold it in place with your fingers. Pour a small amount of distilled water onto the filter paper to secure it, so that you do not have to hold it in place.

13. Filter the mixture by pouring it into the filter paper in the funnel. Use the stirring rod and distilled water in a wash bottle to transfer the entire solid into the filter paper. **HINT**: For best results, be sure to transfer all of the precipitate into the filter paper. Use a rubber policeman if it is available to help with the transfer.

Helpful HINT*: *Filtering technique is very important in this procedure. You may need to review folding and seating the filter paper.

14. Rinse the remaining solid in the filter paper twice with distilled water from a wash bottle to rinse off excess potassium chloride (KCl). After all the liquid has filtered through, rinse the product with approximately 5 mL of ethanol to aid in its drying. Allow the ethanol to completely finish filtering through the paper.

15. Remove the filter paper carefully so as to not lose any product. Gently unfold the filter paper and lay it flat on the pre-weighed watch glass to dry. **HINT**: Do not throw your product away after you weigh it!

16. Allow the product to air dry until the next laboratory period.

17. In the next laboratory period, weigh the dry product on the filter paper and watch glass. Record the total mass in the data section. Calculate the mass of the product.

18. Complete the data and calculations by completing each step and comparing your results with the other groups.

19. Place the dried product in a designated container for later use in the ***Rates of Reaction*** experiment. Wash any dirty glassware

Date:________ Name: ___________________________ Period:_________

Lab Partner: ___________________________________

Data

Record the data for your group.

Group __*A*__/__*B*__/__*C*__ (Examples are given for each of the groups.)

Mass of $CaCl_2$ __*A: 1.006 g*__/__*B: 2.010 g*__/__*C: 3.003 g*__

Mass of K_2CO_3 __*A: 2.503 g*__/__*B: 2.505 g*__/__*C: 2.501 g*__

Mass of filter paper__*A: 0.821 g*__/__*B: 0.742 g*__/__*C: 0.803 g*__

Mass of watch glass__*A: 38.487 g*__/__*B: 35.564 g*__/__*C: 34.731 g*__

Mass of product, filter paper, and watch glass

__*A: 40.160 g*__/__*B: 37.972 g*__/__*C: 37.230 g*__

Mass of dry product__*A: 0.852 g*__/__*B: 1.666 g*__/__*C: 1.696 g*__

Observations for the reaction:

A white precipitate was formed as soon as the two solutions were mixed.

Table 1. Final Results for EACH group

Group	Actual Mass of $CaCl_2$ (g)	Actual Mass of K_2CO_3 (g)	Mass of Dry $CaCO_3$ Obtained (g)

Group A	1.006 g	2.503 g	0.852 g
Group B	2.010 g	2.505 g	1.666 g
Group C	3.003 g	2.501 g	1.696 g

Calculations

1. Determine the limiting reagent for each group. Show your calculations. (**Hint:** See the example in the Introduction.)

Group A:

$CaCl_2$: (1.006 ~~g~~)(1 ~~mol $CaCl_2$~~)(1mol $CaCO_3$) = 0.009065 mol $CaCO_3$
(110.98 ~~g~~) (1 ~~mol $CaCl_2$~~)

K_2CO_3: (2.503 ~~g~~)(~~1 mol K_2CO_3~~)(1 mol $CaCO_3$)= 0.01811 mol $CaCO_3$
(138.21 ~~g~~) (~~1 mol K_2CO_3~~)

$CaCl_2$ is the limiting reagent.

Group B:

$CaCl_2$: (2.010 ~~g~~)(1 ~~mol $CaCl_2$~~)(1mol $CaCO_3$)= 0.01811 mol $CaCO_3$

$$(110.98\ \cancel{g})\ (1\ \cancel{mol\ CaCl_2})$$

$$K_2CO_3: (2.505\ \cancel{g})\left(\frac{\cancel{1\ mol\ K_2CO_3}}{138.21\ \cancel{g}}\right)\left(\frac{1\ mol\ CaCO_3}{\cancel{1\ mol\ K_2CO_3}}\right) = 0.1812\ mol\ CaCO_3$$

They are almost equal, but $CaCl_2$ is slightly limiting.

(Note: These should be about the same and either $CaCl_2$ or K_2CO_3 can be the limiting reagent depending on their initial masses.)

Group C:

$$CaCl_2: (3.003\ \cancel{g})\left(\frac{1\ \cancel{mol\ CaCl_2}}{110.98\ \cancel{g}}\right)\left(\frac{1 mol\ CaCO_3}{1\ \cancel{mol\ CaCl_2}}\right) = 0.02706\ mol\ CaCO_3$$

$$K_2CO_3: (2.501\ \cancel{g})\left(\frac{\cancel{1\ mol\ K_2CO_3}}{138.21\ \cancel{g}}\right)\left(\frac{1\ mol\ CaCO_3}{\cancel{1\ mol\ K_2CO_3}}\right) = 0.01810\ mol\ CaCO_3$$

K_2CO_3 is the limiting reagent.

2. Calculate the theoretical yield of $CaCO_3$ that could be produced by each group and then fill in Table 2.

$$A: (100.09\ g/\cancel{mol})(0.009065\ \cancel{mol}) = 0.9073\ g\ CaCO_3$$

$$B: (100.09\ g/\cancel{mol})(0.01811\ \cancel{mol}) = 1.813\ g\ CaCO_3$$

$$C: (100.09\ g/\cancel{mol})(0.01810\ \cancel{mol}) = 1.812\ g\ CaCO_3$$

Table 2: Comparison of Theoretical and Actual Yields for $CaCO_3$

Group	Limiting Reagent	Theoretical Yield of $CaCO_3$	Actual Yield of $CaCO_3$
Group A	$CaCl_2$	0.9073 g	0.852 g
Group B	$CaCl_2$ or K_2CO_3	1.813 g	1.666 g
Group C	K_2CO_3	1.812 g	1.696 g

3. Find the **percent yield** your group obtained for the $CaCO_3$.

A: $\frac{0.852\ g}{0.9073\ g} \times 100\% = 93.9\%$

B: $\frac{1.666g}{1.813\ g} \times 100\% = 91.89\%$

C: $\frac{1.696\ g}{1.812\ g} \times 100\% = 93.60\%$

Post Lab Questions

1. Compare your results with those of the other two groups. How does the amount of grams of $CaCO_3$ they obtained compare to your results?

   Group A: My group had a yield that was much less than Groups B and C.

   Group B and C: My group had a yield that was much more than Group A, but similar to Group C/B.

2. Were the results of the other two groups as you expected? Why or why not?

Yes, the results were as expected. Group A used about half the amount of $CaCl_2$ than the other groups and it was the limiting reagent for Group A. Group B used about the same amount of moles of both reagents so you would expect them to make about twice the amount of product as Group A. Group C made about the same amount of product as Group B since K_2CO_3 was the limiting reagent and both groups used about the same amount of it.

3. Predict what would happen if 6.0 grams of $CaCl_2$ were used for the reaction and the amount of K_2CO_3 remained the same.

Increasing the amount of $CaCl_2$ would not affect the results since potassium carbonate would be the limiting reagent.

Lab

19

Ideal Gas Law: Finding Percent H_2O_2 with Carrot Juice

*TN Standard **3.3**: The student will explore the mathematics of chemical formulas and equations.*

Have you ever opened a container of milk after the expiration date and found that it had gone bad?

It is very easy to tell if milk has gone bad. It looks, smells, and tastes awful. The putrid odor lets you know a gas is being formed as it decomposes. Many items you purchase are not nearly as easy to tell if they have degraded. Some form a gas that is impossible to detect just by smelling it. Hydrogen peroxide, a common household item that is used to clean minor cuts, is like this. The bottle you buy in the store says it contains 3% hydrogen peroxide, but it will very slowly decompose over time to form water and oxygen gas. Since we breathe oxygen every second, we can't easily detect this. But we can use the ideal gas law and carrot juice to find this out!

Introduction

The ideal gas law is very valuable when dealing with gases since it establishes a relationship between temperature, pressure, volume, and amount of a gas.

Ideal gas law equation: $PV= nRT$

In this equation:

P is the gas pressure in atmospheres

V is the volume of the gas in liters

n is the number of moles of the gas

R is the constant value of 0.0821 L*atm/mol*K

T is for the temperature of the gas in Kelvin.

Since hydrogen peroxide forms oxygen gas when it decomposes, we can use the ideal gas law check the percent hydrogen peroxide in a bottle of it purchased at the store. To find this out we need to take a small sample out of the bottle and accelerate its decomposition through using a catalyst.

$$2\ H_2O_2 \xrightarrow{\text{catalyst}} 2\ H_2O\ +\ O_2$$

In this experiment, we will use carrot juice to accelerate the decomposition of the hydrogen peroxide into water and O_2 gas. Carrots contain the enzyme catalase which is a catalyst for this reaction. You will add carrot juice to a known amount of hydrogen peroxide and quickly seal off the system so that the O_2 gas formed is collected in a graduated cylinder. After measuring the total volume of gas produced, its temperature, and the atmospheric pressure, the ideal gas law can then used to calculate how many moles of O_2 gas is formed. We can do this by solving the ideal gas law equation for **n**.

$$\mathbf{n = \frac{PV}{RT}}$$

Once the number of moles of O_2 gas is calculated, the percent of H_2O_2 there is in the solution can be determined. To do this, the theoretical number of moles of O_2 there would be if the solution was 100% hydrogen peroxide needs to first be calculated. This can be found by using the following equation:

Theoretical moles of O_2 =

$$\text{mL } H_2O_2 \text{ used} * H_2O_2 \text{ density} * \frac{1\text{mol } H_2O_2}{?\text{ g } H_2O_2} * \frac{1\text{ mol } O_2}{2\text{ mol } H_2O_2}$$

For this experiment:

mL H_2O_2 used is the volume of H_2O_2 you actually use (~5 mL).

H_2O_2 density is 1.02 g/mL

1 mol H_2O_2/g H_2O_2 is the reciprocal of the molar mass of H_2O_2 (flipped).

The molar mass of H_2O_2 is 34.0 g/mol, so this is 1 mol H_2O_2/34.0 g H_2O_2.

1 mol O_2/2 mol H_2O_2 is used since 2 moles of H_2O_2 produces 1 mole of O_2

The units in the entire equation cancel to give moles of O_2.

The percent hydrogen peroxide can now be found. To do this divide (**n**), the actual number of moles you calculated, by the **theoretical moles of O_2** there would be if the hydrogen peroxide were 100%. This number is then multiplied by 100%.

$$\% H_2O_2 = \frac{\text{Actual moles } O_2 \text{ (n)}}{\text{Theoretical moles of } O_2} * 100\%$$

This value can now be compared to the 3% hydrogen peroxide shown on the label to see if any decomposition has occurred.

Objectives

- Use the ideal gas law to determine the percentage of hydrogen peroxide in a commercially available hydrogen peroxide solution.
- Observe how a catalyst affects a reaction.
- Determine the decomposition rate of the hydrogen peroxide solution.

Date:________ Name:______________________ Period:_________

Pre-lab Questions

1. What is it in carrot juice that aids in the decomposition of hydrogen peroxide?

   The enzyme catalase acts as a catalyst to accelerate the decomposition of hydrogen peroxide.

2. List the ideal gas law and define each term with units.

   PV = nRT

   P = Pressure of the gas (atm)

   V = Volume of the gas (L)

   n = Number of moles of the gas (mol)

   R = Gas constant (0.0821 L*atm/mol*K)

   T = Temperature of the gas (K)

3. How many moles of O_2 were produced in a decomposition reaction of H_2O_2 if the barometric pressure was 0.980 atm, the temperature was 298 K and the volume of O_2 gas collected was 0.0500 L?

   n = PV / RT

   n=(0.980atm)(0.0500L)/(0.0821L*atm/mol*K)(298 K)

   n = 2.00 x 10^{-3} moles

4. If you decomposed 10.00 mL of 100% H_2O_2, how many moles of O_2 could you theoretically obtain?

Theoretical moles O_2 =

$$10.00\ \text{mL} * \frac{1.02\ \text{g}}{\text{mL}} * \frac{1\ \text{mol}\ H_2O_2}{34.0\ \text{g}} * \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2O_2}$$

Theoretical moles O_2 = 0.15 moles O_2

Preparation Materials

Teaching Preparation: The carrot juice MUST be fresh! It should be prepared just prior to the procedure. The carrot/water ratio can be adjusted to accommodate class size. A more concentrated carrot juice works faster.

General Supplies	Laboratory Equipment
Fresh carrot juice: For preparation add 1 large carrot and ~100 mL distilled water into a blender. Liquefy and then pour through a strainer into a beaker.	100 mL graduated cylinder
Over the counter 3% hydrogen peroxide solution	125 mL Erlenmeyer flask
Thermometer	1 hole rubber stopper with bent glass tube that fits Erlenmeyer Flask
Distilled or deionized water (tap water can be used if needed)	600 mL beaker
Barometer (If a barometer is not available you can use the barometric pressure given on the internet for your area. This will be close enough.)	U tube (Refer to procedure at end of experiment)
	Latex tubing that snugly fits over the glass tubing
	Large ring or clamp and ring stand
	Dropper

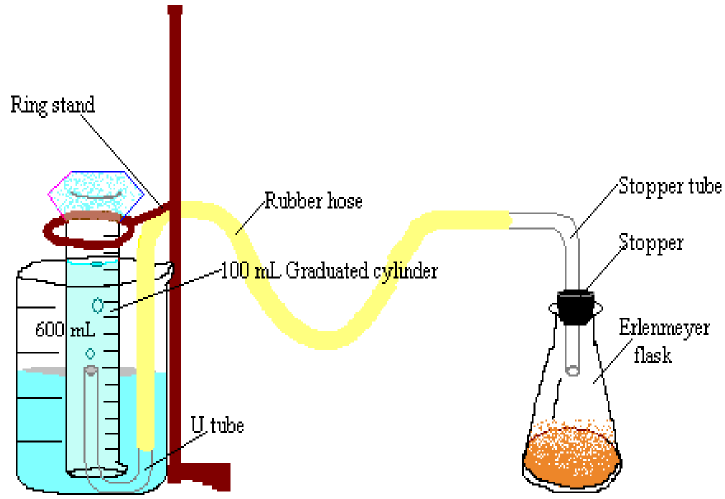


Figure 1. Gas Collection Apparatus

Procedure

1. Assemble an apparatus as shown in Figure 1. First attach one end of the rubber hose to the U tube, and the other end to the glass stopper tube. **HINT:** The figure is very helpful, USE IT!

2. Check to make sure the stopper fits the Erlenmeyer flask. **HINT:** Make sure it is a tight fit in order to prevent O_2 from leaking.

3. Fill the 600 mL beaker with ~400 of distilled water.

4. Fill 100 mL graduated cylinder as close to ½ inch from the top as possible.

5. Without spilling any water, cover the opening of the graduated cylinder with two or three fingers, and turn it upside down into the 600 mL beaker already containing 400 mL of water. DO NOT remove your fingers from the opening until the opening is submerged under the water in the beaker.

6. Insert the open side of the U tube under the graduated cylinder while keeping its opening under water. You want as little air as possible to be in the graduated cylinder.

7. Secure the upside – down graduated cylinder to the ring stand with a clamp

SAFETY FEATURES

Safety Goggles

MATERIALS NEEDED

- Fresh carrot juice
- Hydrogen peroxide
- Distilled Water
- Large ring or clamp
- Ring Stand
- 10 and 100 mL Graduated Cylinders
- U tube
- Stopper with bent glass tube
- Dropper
- 600 mL beaker
- Tubing
- 125 mL Erlenmeyer flask
- Thermometer
- Barometer

or large enough ring to fit over the base of the graduated cylinder.

8. Record the volume of air in the graduated cylinder (the line at which the water reaches in the cylinder) in the data section in Table 1. **HINT:** You may need to very carefully add a little air to get the initial volume of air to where it can be read. This can be done by very briefly tipping the graduated cylinder up above the water. Be sure you have no more than ~10 mL of air in the graduated cylinder before going to the next step.

9. Use a 10 mL graduated cylinder and dropper to measure out 5.00 mL of hydrogen peroxide.

10. Pour the hydrogen peroxide into the 125 mL Erlenmeyer flask.

11. Clean the 10 mL: graduated cylinder by rinsing it al least three times with distilled water. Dispose of the rinse down the drain.

12. Take the temperature of the water in the 600 mL beaker, and record it in the data section. Also, determine the barometric pressure in the lab, and record it in the data section. **HINT:** This may be provided by your teacher.

13. Measure out ~5.00 mL of carrot juice using the rinsed 10 mL graduated cylinder. **HINT:** Do not immediately pour the carrot juice into the Erlenmeyer flask by twisting it down into the flask.

14. Prepare to put the stopper on the Erlenmeyer flask.

15. Quickly pour the 5.0 mL of carrot juice into the Erlenmeyer flask. Immediately place the stopper snugly in the opening of the Erlenmeyer flask by twisting it down into the flask.

16. Record the time in the data section under 1.

17. Swirl the Erlenmeyer flask to mix the two solutions together.

18. You will begin to see bubbles coming up into the 100 mL graduated cylinder. **HINT:** If O_2 gas does not almost immediately replace the water in the graduated cylinder, you probably did not put the stopper on tightly enough or you have a leak in your tubing. You will need to correct the problem and start over.

19. Continue to swirl the Erlenmeyer flask and let the reaction run until **no more** bubbles form. This should take approximately 6 – 10 minutes to assure

the reaction has gone to completion. **HINT:** Catalase works best around the temperature of the human body. You can speed the reaction up by warming the Erlenmeyer flask with your hands.

20. Record the time when the reaction is finished in table 2 of the data section.

21. Record the final volume of air in Table 1 of the data section. Remember to read it at eye – level and measure from the bottom of the meniscus.

Date:________Name:__________________________ Period:_________

Lab Partner: __________________________________

Data

Water Temperature _____23.4___________ °C

Barometric Pressure _____753___________ mm Hg

Table 1.

Initial volume of air (mL)	Final volume of air after reaction (mL)	Volume of O_2 Collected (Final volume – Initial volume of air)
5.7 mL	59.9 mL	54.2 mL

Table 2.

Time reaction started	Time reaction ended	Reaction Time (sec)
10:46	10:54	8:00 minutes OR 480 sec

Calculations

The goal is to find the percentage of hydrogen peroxide in the solution! This can be found by working through the following steps.

1. Convert the temperature of the water from °C to Kelvin (K).

Use the equation: **K = ^{0}C + 273.** This will be equal to **T** or the temperature in Kelvin.

T = ___23.4_____°C + 273 = ____296.4____ **K**

2. If necessary, convert the barometric pressure in the room from mm Hg to atmospheres (atm.)

Divide the measured pressure from the data section by 760 mm Hg. This will be equal to **P** or the pressure in atmospheres.

$$P = \underline{\quad 753 \quad} \text{ mm Hg} * \frac{1 \text{ atm}}{760 \text{ mm Hg}} = \underline{\quad 0.991 \quad} \text{atm}$$

3. Convert the volume of oxygen from mL to liters (L).

Divide the "Volume of O_2 Collected" by 1000. (The volume collected was measured in mL in the data section.) This will be equal to **V** or the volume in liters.

$$V = \underline{\quad 54.2 \quad} \text{ mL} * \frac{1\text{L}}{1000 \text{ mL}} = \underline{\quad 0.0542 \quad} \text{ L}$$

4. The gas law has a constant **R** (no conversion needed):

R = 0.0821 L * atm / mol K

5. Rearrange the ideal gas law to solve for **n.**

$$n = \frac{PV}{RT}$$

6. You are now ready to solve for the number of moles of O_2. Be sure the units cancel so that you end up with only the moles of O_2 left.

$$\textit{Actual number of moles of } O_2 \text{ (n)} = \frac{0.991 \text{ atm} * 0.0542 \text{ L}}{0.0821 \frac{\text{L*atm}}{\text{mole * K}} * 296.4 \text{ K}}$$

***Actual* number of moles of O_2 (n) = ____0.00221____ moles**

7. Calculate the **theoretical number of moles of O_2** there would be if the hydrogen peroxide were 100%, and not an aqueous solution.

Theoretical moles of O_2 =

$$H_2O_2 \text{ volume} * H_2O_2 \text{ density} * \frac{\text{mol } H_2O_2}{?\ \text{g } H_2O_2} * \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O_2}$$

For the above equation:

H_2O_2 volume (the actual mL of H_2O_2 you used) ___________________

H_2O_2 density is 1.02 g/mL

$\frac{\text{mol } H_2O_2}{\text{g } H_2O_2}$ is the reciprocal of the molar mass of H_2O_2 (flipped it) so the units will cancel

Calculate the molar mass of H_2O_2:

Molar mass of H_2O_2 = ____*34.0*____ **g H_2O_2/1 mol H_2O_2**

Molar mass of H_2O_2 reciprocal = *1mol H_2O_2/34.0 g H_2O_2*

Now you have all of the information needed to solve the equation for the theoretical moles of O_2. All you need to do is fill in the blanks and do the calculations.

Theoretical moles of O_2 =

$$\underline{5.00 \text{ mL}}\ H_2O_2 * 1.02 \text{ g/mL} * \frac{1 \text{ mol } H_2O_2}{\underline{34.0}\ \text{g}} * \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O_2}$$

Theoretical **moles (n) of O_2 =** ___*0.0750*___**moles O_2**

8. Find the percent hydrogen peroxide.

% H_2O_2 = *Actual moles O_2* * 100%

Theoretical moles O_2

% hydrogen peroxide = _0.0221moles_ * 100%
0.0750 moles

% hydrogen peroxide = ____2.93_____ %

9. You can also easily determine the reaction rate. To do this, divide the total volume of oxygen collected by the total time of the reaction.

Reaction rate = $\frac{\textbf{Volume of } O_2}{\textbf{Reaction Time}}$ = $\frac{54.2 \text{ mL}}{480 \text{ sec}}$

Reaction rate = _____0.113_____ mL/sec

Post Lab Questions

1. Was the calculated percentage of hydrogen peroxide close to the same as the percentage on the label?

I experimentally found the bottle contained 2.93% H_2O_2 which is very close to the value of 3% found on the hydrogen peroxide label.

2. Considering that catalysts are not consumed in a reaction, how do you think increasing the amount of catalyst would affect the reaction rate for the decomposition of hydrogen peroxide?

If the amount of catalyst was increased, the reaction rate would be increased up to a maximum point.

Supplementary Teaching Procedure

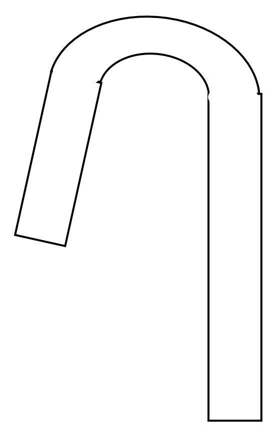
Procedure for Making a U Tube

1. Light a Bunsen burner and adjust the flame so that there is a medium flame with a blue inner cone.

2. Place a 6 inch piece of ~7 mm O.D. glass tube that has flame polished ends over the flame so that the flame contacts the glass from about 2 to 3 inches from the end.

3. Very carefully heat the glass tube in the flame until it becomes softened. Do not heat it so long that you burn your fingers.

4. Bend the tube so that it approximates the U tube pattern shown below.



U Tube Pattern

Lab

20

Rates of Reaction

TN Standard ***3.2****: The student will analyze chemical reactions.*

Why is it easier to ***sweeten*** *tea from hot water and granular sugar than with cold water and sugar cubes?*

As you begin to stir the cold tea and the square sugar cubes, the sugar particles sluggishly separate from the cube while following the flow of your spinning stirrer. Eventually, after much stressful stirring, the little sugar crystals make their way into solution. Suddenly, you notice that your friend has almost finished his sweet tea and has returned to see if you are through making yours so that he could have some more. You complain that you've been stirring forever just to dissolve the sugar in your cold tea. Your friend responds, "I only had to stir mine for a few seconds, and it was good to drink!" Why do you think your friend's sweet tea was finished so much faster? Well, his sweet tea was made from hot water and granular sugar.

Introduction

A reaction rate is the time that it takes for the reactants to be changed into products. Reaction rates can be affected by several factors which include the following: nature of reactants, surface area, concentration, temperature, pressure, and the presence of a catalyst. Whether a reaction rate will increase or decrease depends on the rate that the molecules involved effectively collide to result in a reaction. A reaction rate is given as the change of the concentration of a reactant or product in a certain amount of time. Various units are used to describe it.

Throughout this laboratory exercise, you will use calcium carbonate and hydrochloric acid to discover how temperature, surface area, and concentration affect reaction rates. Calcium carbonate is the main compound found in marble. Marble is often used to make statutes or as decorative rock chips in flower beds. Hydrochloric acid reacts with calcium carbonate to form calcium chloride, carbon dioxide gas, and water. This is similar to how acid rain degrades marble statutes.

In this laboratory exercise you will compare how two different surface areas of calcium carbonate, a powder and a solid rock piece, react with different concentrations of hydrochloric acid at various temperatures. The powder has a high quantity of surface area. In contrast, the solid, a crystallize rock, has a much lower amount of surface area. You will record how long it takes for each reaction to go to completion, and then calculate and compare the reaction rates.

Objectives

- To understand how temperature, surface area, and concentration affect the rate of reactions.

- To relate reaction rates on a molecular level.

Date:________ Name:______________________ Period:_________

Pre-lab Questions

1. Name five factors that can affect the rate of a reaction.

   Surface area, temperature, concentration, pressure, and presence of a catalyst

2. Complete and balance the reaction between $CaCO_3$ (s) and HCl(aq).

   $CaCO_3$ (s) + ***2*** HCl(aq) ⟶ ***$CaCl_2$(aq) + CO_2(g) + H_2O(aq)***

3. In the opening paragraph it took more time to make sweet tea with cold water and sugar cubes than to sweeten hot tea with granular sugar. Why?

   Both temperature and surface area affect the rate at which the sugar would dissolve. The greater temperature and smaller surface area of the grains of sugar made dissolving the sugar occur much faster in the hot tea.

4. What is the primary factor that determines whether a reaction rate increases or decreases?

   The primary factor that determines whether the rate of a reaction increases or decreases is the rate at which the molecules that are involved effectively collide to result in a reaction.

Preparation Materials

General Supplies	Laboratory Equipment	Chemicals

Stopwatch	4 Large test tubes (30 x 150 mm)	$CaCO_3$ rocks (Marble chips can be used. They are available in the garden section of some home improvement stores. They will need to be broken into smaller pieces by hitting them with a hammer.)
Ice	50 mL Beaker	Calcium carbonate powder $CaCO_3$ (You can use the product made in the Stoichiometry Lab or pulverize the marble chips with a hammer if the powder is not available.)
Sharpie marker	10 mL graduated cylinder	4 M HCl, Hydrochloric acid (30 mL per group) Prep: 332 mL concentrated HCl (37%) per 1 L H_2O
	Large beakers for ice bath and hot water bath	2 M HCl, Hydrochloric acid (20 mL per group) Prep: 166 mL concentrated HCl (37%) per 1 L H_2O
	Hot plate	
	Balance	
	Test tube rack and holder	
	Stirring rod	

Procedure

SAFETY FEATURES

- Safety Gloves
- Safety Goggles

MATERIALS NEEDED

- 4 M HCl, Hydrochloric Acid
- 2 M HCl, Hydrochloric Acid
- Calcium Carbonate powder and rock, $CaCO_3$
- Hot plate
- 4 Large test tubes
- Test tube rack
- Test tube holder
- 50 mL Beaker
- Sharpie Marker
- Balance
- 10 mL Graduated Cylinder
- Stopwatch
- Ice
- Stir rod
- Large beaker

1. Fill a large beaker 2/3rds full with deionized water and place it on a hot plate. Heat the water to just boiling. **CAUTION: Heat the water to just boiling. Not a rolling boiling, which can easily boil over and cause burns.**

2. While the water is heating, label 4 large test tubes 1, 2, 4, and 5.

3. Weigh out approximately 0.25 - 0.30 grams of a piece of $CaCO_3$ rock. Record this mass in the data section. Place the rock in test tube **1**.

4. Weigh out three more pieces of approximately 0.25 grams of $CaCO_3$ rock. These should be as close as possible to the mass of the first rock sample. Place the pieces of rock into test tubes **2, 4 and 5**. Record each of their masses in the data section.

5. Into a 50 mL beaker, weigh out an amount of $CaCO_3$ powder as close as possible to the amount of the previously weighed pieces of marble rock. Record the mass in the data section.

Reaction #1

6. Measure 10 mL of 4M HCl in a 10 mL graduated cylinder. Add this to the weighed out piece of $CaCO_3$ rock in test tube **1,** and place it immediately in an ice bath. Record the starting time. Check on this reaction frequently and record the time when the reaction is no longer fizzing. Record observations in the data table.

Reaction #2

7. Measure 10 mL of 2M HCl in a 10 mL graduated cylinder. Add this to the weighed out piece of $CaCO_3$ rock in test tube **2,** and place it in the test tube rack. Record the starting time. Check on this reaction frequently and record the time when the reaction is no longer fizzing. Record observations in the data table.

Reaction #3

8. Measure 10 mL of 2M HCl in a 10 mL graduated cylinder. Add this to the weighed out $CaCO_3$ powder in the 50 mL beaker. Use a stopwatch to time the reaction from when the HCl is poured onto the powder and until the reaction is done fizzing. Record this time in the data table. Record observations in the data table.

Reaction #4

9. Measure 10 mL of 4M HCl in a 10 mL graduated cylinder. Add this to the weighed out piece of $CaCO_3$ rock in test tube **4,** and place it in the test tube rack**.** Record the starting time. Check on this reaction frequently and record the time when the reaction is no longer fizzing. Record observations in the data table.

Reaction #5

10. Measure 10 mL of 4M HCl in a 10 mL graduated cylinder. Add this to the weighed out piece of $CaCO_3$ rock in test tube **5,** and place it in the previously started hot water bath**.** Record the starting time. Check on this reaction frequently and record the time when the reaction is no longer fizzing. Record observations in the data table.

11. To clean up, combine all solutions in a beaker and neutralize with baking soda. Pour the neutralized solution down the drain.

Date:________ Name: ___________________________ Period:_________

Lab Partner: ___________________________________

Data

Complete the data table.

Substance, Reaction #	Variable	Mass of the $CaCO_3$ (g) (these should be close)	Time of the reaction (sec) (start/stop)	Observations
$CaCO_3$ Rock #1	Iced, 4M HCl	0.320 g	(24 minutes)= 1440 sec	Bubbling at a slow rate.
$CaCO_3$ Rock #2	Room Temperature, 2M HCl	0.313 g	(13 minutes)= 780 sec	Bubbling faster than reaction 1.
$CaCO_3$ Powder #3	Room Temperature, 2M HCl	0.303 g	12.73 sec	Reaction went quickly.
$CaCO_3$ Rock #4	Room Temperature, 4M HCl	0.324 g	(7 min.&22 sec)= 442 sec	Bubbling quicker than other rocks.
$CaCO_3$ Rock #5	Heated, 4M HCl	0.308 g	(1 min.&19 sec)= 79 sec	Reaction continued fastest of rocks.

Calculations

Calculate the rate of each of the reactions in g/sec.

- **Reaction 1** (Rock, iced, 4M HCl):

  (0.320 g/ 1440 sec) = 2.22 x 10 $^{-4}$ g/sec

- **Reaction 2** (Rock, room temperature, 2M HCl):

  (0.313 g/ 780 sec) = 4.01 x 10 $^{-4}$ g/sec

- **Reaction 3** (Powder, room temperature, 2M HCl):

  (0.303 g/ 12.73 sec) = 2.38 x 10 $^{-2}$ g/sec

- **Reaction 4** (Rock, room temperature, 4M HCl):

  (0.324 g/ 442 sec) = 7.33 x 10 $^{-4}$ g/sec

- **Reaction 5** (Rock, heated, 4M HCl):

  (0.308 g/ 79 sec) = 3.90 x 10 $^{-3}$ g/sec

Post Lab Questions

1. All the reactions that you performed were chemically the same. You just varied several factors. What were the factors that were varied?

The three altered factors among the reactions were the surface area of the $CaCO_3$, the concentration of the HCl, and the temperature at which the reaction occurred.

2. Which factor do you think made the biggest influence on the reactions? Why?

The surface area of the $CaCO_3$ appeared to make the greatest impact on the rate of reaction. The powder sample had a much faster rate than the rock sample tested at the same conditions.

3. Out of the 5 different reactions, which reaction was the slowest? Was this what you expected? Why?

The $CaCO_3$ rock reacted with 4M HCl in an ice bath had the slowest rate of reaction. This was expected because the reaction was conducted at a lower temperature. Lower temperatures will lower the number of molecular collisions that occur. Also, a rock has a lower amount of surface area than a powder. This will also slow the reaction rate.

4. Why do you think marble statues require long periods of time to degrade in areas that have acid rain?

The concentration of the acid found in acid rain is much less than the acid we used. Also marble statues have a much lower amount of exposed surface area than the rock pieces or powder used in this lab.

Lab

21

Acceleration of Reactions by a Catalyst

*TN Standard **2.2:** The student will explore the interactions of matter and energy.*

Why do bubbles form when you put hydrogen peroxide on a wound?

Ouch! You just scraped your knee. After you hobble to a sink to wash it off, you put some hydrogen peroxide on it. Small bubbles start forming almost immediately. Why? The reason is because blood and tissue contain a certain enzyme that accelerates the decomposition reaction of hydrogen peroxide forming oxygen gas and water. The enzyme is a biological catalyst. When there is a need for speed in a reaction, using a catalyst is often the best method.

Introduction

Many of the chemicals the human body needs are made within the cells. This means the human body has a real need for super fast chemical reactions. A high reaction temperature or a large concentration of the reactants will often sufficiently speed up a reaction in the laboratory. But our cells can't rapidly increase their temperature or suddenly add a lot more of certain chemicals. Instead, the human body uses catalysts. A catalyst is a substance that speeds up a reaction, but is not used up during the reaction. Biological catalysts are called enzymes. There are many different types of enzymes. Each type speeds up a certain reaction your body needs to have happen right then! Without catalysts your body could not do even the simplest task.

Reactions have a minimum amount of energy they need in order to occur. This is called the a**ctivation energy**. A catalyst will lower the activation energy. It does this by making a way for the reaction to take place that requires less energy.

Chemists cannot always speed up a reaction by changing the usual variables, and reactions that take a long time are seldom very useful. For this reason, chemists often use catalysts to speed up reactions.

After a chemist decides to use a catalyst there are several things that have to be studied. One of the biggest challenges is finding just the right catalyst. Frequently several catalysts are found that will work, and they have to be compared to each other to determine which is the best one. Some of the factors that a chemist will take into consideration in choosing a catalyst include the desired speed of the reaction, the cost of the catalyst, how long the catalyst will work, and if it is toxic or harmful to the environment. In addition to the catalyst in blood and tissue, there are several other catalysts that can be used to catalyze the reaction to decompose hydrogen peroxide into water and oxygen gas. For example, manganese dioxide, many fruits and vegetables, household bleach, and even dirt can all be used to catalyze this reaction.

Another thing to be decided is whether to use a homogeneous or a heterogeneous catalyst. A **heterogeneous catalyst** is in a different phase as the reactants and a **homogeneous catalyst** is in the same phase. A piece of a carrot put into a solution of hydrogen peroxide is an example of a heterogeneous catalyst. This is because the carrot is in a solid phase and the hydrogen peroxide is in a liquid phase. If instead the carrot is made into a juice and added to the hydrogen peroxide solution, it is a homogeneous catalyst since it is then in the liquid phase. Generally homogeneous catalysts will react faster, but heterogeneous catalysts are easier to remove from the products.

Chemists also have to determine the best amount of a catalyst to use. If a catalyst is very expensive, toxic, or hard to remove from the product, they may use the least amount of catalyst that will work. If instead there is a need to have the reaction happen more rapidly, a chemist may choose to add more catalyst. There is a point however, where adding more catalyst will not increase the reaction rate. This is because there is as much or more catalyst than the limiting reactant.

In this laboratory exercise you will evaluate carrots, tomatoes, yeast, and dirt as catalysts to decompose hydrogen peroxide. You will also observe the differences between using a piece of carrot or carrot juice as a catalyst for this reaction. Finally you will observe the effects of adding different amounts of carrot juice on the reaction rate.

Objectives

- Evaluate different catalysts to determine which one is the best choice.
- Illustrate the differences between using heterogeneous and homogenous catalysts.
- Demonstrate how varying quantities of catalyst affect the reaction.

Date:________ Name:_______________________ Period:_________

Pre-lab Questions

1. What is a catalyst?

A catalyst is a substance used to speed up a reaction, but is not consumed during the reaction.

2. If you continue to add more catalyst will the speed of a reaction always continue to increase? Explain your answer.

No. There is a maximum point that can be reached where additional catalyst will no longer increase the rate of a reaction.

3. In this lab you will produce oxygen and water from hydrogen peroxide (H_2O_2). Write a balanced reaction equation for this reaction.

$$2H_2O_2 \rightarrow O_2(g) + 2H_2O$$

4. What causes the bubbles to form in this reaction?

The bubbles are caused by the oxygen gas that is formed.

5. The exhaust gas from car engines pass through catalytic converters that contain very small amounts of solid platinum, palladium, and rhodium catalysts. Are these metals homogeneous or heterogeneous catalysts?

Heterogeneous

Preparation Materials

Teaching Preparation: The carrot juice MUST be fresh! It should be prepared just prior to the procedure. The carrot/water ratio can be adjusted to accommodate class size. A more concentrated carrot juice works faster.

General Supplies	**Laboratory Equipment**
Carrots	11 medium test tubes (18 x 150 mm)
Fresh carrot juice: For preparation add 1 medium carrot and 100 mL distilled water into a blender. Liquefy and then pour through a strainer into a beaker.	10 mL graduated cylinder
3% Hydrogen peroxide (H_2O_2)	Test tube rack
Yeast	3 Droppers
Tomatoes	Blender
Dirt	Strainer
Sharpie marker	

Procedure

SAFETY FEATURES

Safety Goggles

MATERIALS NEEDED

- Fresh carrot juice
- 11 Medium test tubes
- 10 mL Graduated cylinder
- Small piece of a carrot
- Small piece of a tomato
- Dirt
- Hydrogen peroxide (H_2O_2)
- Yeast
- 3 Droppers
- Test tube rack
- Sharpie marker

Helpful HINT: ***Hydrogen Peroxide should not be left open or exposed to the air. It will degrade with time.***

Teaching Alternative: ***The solid piece of carrot can be substituted with a potato or turnip. Also, Liver (turkey, chicken, or deer) can be purchased at many grocery stores and works very well in this reaction.***

Part 1. Comparison of Different Catalysts

1. Place 11 medium sized test tubes in a test tube rack.

2. Use a Sharpie or wax pencil to label 5 of the test tubes **C**, **T**, **D**, **CJ**, and **Y**, symbolizing carrot, tomato, dirt, carrot juice, and yeast. **HINT**: It's best to clearly label glassware to prevent cross contamination.

3. Use a 10 mL graduated cylinder and a dropper to add 3 mL of hydrogen peroxide to EACH of the 5 **labeled** test tubes.

4. Carefully add a small piece (~1 cm^2) of carrot to the test tube labeled **C**, and a small piece of tomato to the test tube labeled **T**. Add a similar small amount of dirt to the test tube labeled **D**.

5. Measure 3 mL of carrot juice into a clean 10 mL graduated cylinder. Slowly add the carrot juice to the test tube labeled **CJ**. Add a small amount of yeast to test tube **Y** and swirl until all bubbling and foaming stops. This will indicate the completion of the reaction.

6. Record observations for each of the reactions in the Initial Observations column in Table 1 in the data section. Let the reactions continue until the end of the next part of the procedure.

Part 2. Catalyst Quantity Comparison

7. Use a Sharpie or wax pencil to label the 6 remaining test tubes, **1**, **5**, **10**, **1A**, **5A**, and **10A**.

8. Use a 10 mL graduated cylinder and dropper to add 1 mL of hydrogen peroxide to the test tubes labeled **1**, **5**, and **10**.

9. Use a **clean** 10 mL graduated cylinder to add 1 mL of fresh carrot juice to test tube **1A**, 5 mL of carrot juice to test tube **5A,** and 10 mL of carrot juice to test tube **10A**.

10. **Simultaneously** (or as close to the same time as possible) pour the carrot juice from test tubes **1A**, **5A**, and **10A** into the corresponding test tubes **1**, **5**, and **10** containing the hydrogen peroxide.

11. Record your observations of the 3 reactions in Table 2 of the data section. **HINT**: Be sure to note observations both of similarities and differences in the bubbling and foaming among the 3 reactions.

12. Return to the test tubes in Part 1 and make final observations. Record your observations in the Final Observations Column in Table 1 of the data section.

13. To clean-up decant the liquid from the heterogeneous catalysts and wash all solutions down the sink with plenty of water. Dispose of the carrot, tomato, and dirt in the trash.

Date:________Name: __________________________ Period:_________

Lab Partner: __________________________________

Data

Table 1: Comparison of Different Catalysts

Type of Catalyst	Initial Observations	Final Observations
Carrot	*Only a very small amount of bubbles were produced from the carrot in the beginning.*	*The reaction was still going since bubbles continued to be produced.*
Tomato	*Bubbles and slight foaming were produced from the tomato.*	*The reaction was still going since bubbles continued to be produced.*
Dirt	*Bubbles were produced from the dirt.*	*The reaction was still going since bubbles continued to be produced.*
Carrot Juice	*Bubbles were produced at a fairly fast rate.*	*No more reaction was occurring.*

Yeast	*Bubbling at a rapid, violent rate and foaming.*	*No more reaction was occurring and the yeast was dissolved.*

Table 2: Catalyst Quantity Comparison

Amount of Carrot Juice	Observations
1 mL	*Slow bubbling.*
5 mL	*Faster bubbling and slight foaming.*
10 mL	*Bubbling and foaming much like the reaction using 5 mL of carrot juice.*

Post Lab Questions

1. Classify each catalyst you used as homogeneous or heterogeneous.

   Homogeneous – ***Yeast, Carrot Juice***

   Heterogeneous – ***Carrot piece, Tomato, Dirt***

2. Which catalyst made the reaction go the fastest? Is it a homogeneous or heterogeneous catalyst?

The yeast was the fastest. It is a homogeneous catalyst.

3. Which catalyst would be the easiest to remove from the water that was formed? Is it a homogeneous or heterogeneous catalyst?

The piece of carrot would be the easiest to remove. It is a heterogeneous catalyst.

4. Which amount of carrot juice you tested would be the best to use? Explain your answer.

Usually the answer will be 5 mL since the reaction was faster than the 1 mL and not much difference from the 10 mL. However, this answer can vary depending on the freshness and dilution of the carrot juice.

Lab

22

Properties of Acids and Bases

TN Standard ***4.2****: The student will investigate the characteristics of acids and bases.*

Have you ever brushed your teeth and then drank a glass of orange juice?

What do you taste when you brush your teeth and drink orange juice afterwards. Yuck! It leaves a really bad taste in your mouth, but why? Orange juice and toothpaste by themselves taste good. But the terrible taste results because an acid/base reaction is going on in your mouth. Orange juice is a weak acid and the toothpaste is a weak base. When they are placed together they neutralize each other and produce a product that is unpleasant to taste. How do you determine what is an acid and what is a base? In this lab we will discover how to distinguish between acids and bases.

Introduction

Two very important classes of compounds are acids and bases. But what exactly makes them different? There are differences in definition, physical differences, and reaction differences. According to the Arrhenius definition, acids ionize in water to produce a hydronium ion (H_3O^+), and bases dissociate in water to produce hydroxide ion (OH^-).

Physical differences can be detected by the senses, including taste and touch. Acids have a sour or tart taste and can produce a stinging sensation to broken skin. For example, if you have ever tasted a lemon, it can often result in a sour face. Bases have a bitter taste and a slippery feel. Soap and many cleaning products are bases. Have you accidentally tasted soap or had it slip out of your hands?

Reactions with acids and bases vary depending on the substances being reacted. Acids and bases react differently. For example, bases do not react with most metals, but an acid will react readily with certain metals to produce hydrogen gas and an ionic compound. An ionic compound is often referred to as a **salt**. An example of this type of reaction occurs when magnesium metal reacts with hydrochloric acid. In this reaction magnesium chloride (a salt) and hydrogen gas are formed.

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2(g)}$$

metal + acid → a salt + hydrogen gas

Acids may also react with a carbonate or bicarbonate to form carbon dioxide gas and water. The general reaction equation for a reaction between an acid and a carbonate can be represented in this manner:

$$CO_3{}^{2-}{}_{(aq)} + 2H_3O^+{}_{(aq)} \rightarrow CO_{2(g)} + 3H_2O_{(l)}$$

carbonate + acid → carbon dioxide + water

The general reaction equation for a reaction between an acid and a bicarbonate is similar and can be represented in this manner:

$$HCO_3{}^-{}_{(aq)} + H_3O^+{}_{(aq)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$$

bicarbonate + acid → carbon dioxide + water

Acids and bases can also react with each other. When the two opposites react with each other they cancel each other out so that the product formed has neither the acid nor the base properties. This type of reaction is called a **neutralization** reaction. The general reaction equation for the reaction between an acid and a base is represented in this manner:

$$H_3O^+ + OH^- \rightarrow 2H_2O$$

Acid + Base → Water

An example of a neutralization reaction is when an aqueous solution of HCl, a strong acid, is mixed with an aqueous solution of NaOH, a strong base. HCl when it is in water forms H_3O^+ and Cl^-. NaOH in water forms Na^+ and OH^-. When the two solutions are mixed together the products are water and common table salt. Neither water nor table salt has acid or base properties. Generally this reaction is written without the water solvent shown as a reactant.

$$HCl + NaOH \rightarrow H_2O + NaCl$$

An aqueous (water) solution that has a lot of hydronium ions but very few hydroxide ions is considered to be very acidic. If instead an aqueous solution has a lot of hydroxide ions but very few hydronium ions, it is considered to be very basic. Acids and bases are measured on a scale called pH. The pH is minus the log of the hydronium ion concentration.

$$pH = -\log[H_3O^+]$$

pH ranges from less than 1 to 14. It lets us quickly tell if something is very acidic, a little acidic, neutral (neither acidic nor basic), a little basic or very basic. A pH of 1 is highly acidic, a pH of 14 is highly basic, and a pH of 7 is neutral.

pH indicators, litmus paper, and pH paper can be used to determine whether something is an acid or a base and the strength of its acidity or basicity. An indicator is a substance that turns a different color at a certain pH. Litmus paper is a form of an indicator. It is made by coating paper with the indicator litmus. Litmus is known to change color at a pH of about 7. Either red or blue litmus paper can be purchased. Blue litmus paper remains blue when dipped in a base, but it turns red when an acid touches it. Red litmus paper stays red when dipped in an acid, but turns blue when a base touches it.

Another way to more specifically determine an acid or base is through the use of pH paper. pH paper allows us to determine to what degree a substance is acidic or basic. When a substance is placed on pH paper a color appears. The color is compared to a color chart showing the color the pH paper will turn at different pH values.

In this experiment, we will observe the neutralization of acids and bases using grape juice as an indicator, and an acid-base reaction of a bicarbonate. We will also test common household products for their acidity or basicity.

Objectives

- To understand the properties and reactions of acids and bases.
- To relate these properties to common household products.

Date:________ Name:______________________ Period:_________

Pre-lab Questions

1. What is a neutralization reaction?

   A neutralization reaction is a reaction between an acid and a base in which the product is neither acidic nor basic.

2. Hydrochloric acid (HCl) is a strong acid. About what pH would you expect it to be?

   pH of 1

3. Sodium hydroxide (NaOH) is a strong base. About what pH would you expect it to be?

   pH of 14

Preparation Materials

General Supplies	Laboratory Equipment	Chemicals
Concentrated grape juice (diluted ~10% of juice in distilled water) (Welch's works well)	3 Medium test tubes (18 x 150 mm) 7-10 additional test tubes can be helpful to tape to sample container to hold a dropper for each sample.	1 M HCl Prep: 8.3 mL concentrated HCl per 100 mL H_2O
Vinegar	2 Droppers 7-10 additional droppers could be helpful to have a clean dropper for specific for each sample.	1M NaOH Prep: 4.0g per 100 mL
Tomato juice	pH paper	Universal indicator
Household ammonia	Red and/or blue litmus paper	
Lemon juice	Spot plate	
Milk	10 mL Graduated cylinder	
Dilute dishwashing liquid	100 mL Beaker	
Distilled water	Balance	
Baking soda (sodium bicarbonate)	Stirring rod	

Procedure

SAFETY FEATURES

- Safety Gloves
- Safety Goggles

MATERIALS NEEDED

- ~ 10% Welch's Grape Juice solution
- 3 Medium size test tubes
- 1 M HCl
- 1 M NaOH
- 2 droppers
- Stirring rod
- 10 mL graduated cylinder
- 100 mL Beaker
- Balance
- pH paper
- Red and blue litmus paper
- Sodium bicarbonate (baking soda)
- Distilled water
- Universal indicator
- Vinegar
- Tomato juice

Helpful HINT:* *If the grape juice is not dilute enough or the base is not as strong as needed, students may continue adding drops of base.

Part 1: Acid-Base Neutralization

1. Label 3 medium sized test tubes **1**, **2**, and **Standard**.

2. Pour 10 mL of the dilute grape juice solution into each test tube.

3. Note the color of the juice in the test tube labeled **standard** in Table 1.

4. Add 10 drops of 1M HCl into test tube **1.** Record your observations concerning the color change in Table 1 of the data section. Use the juice in the test tube labeled **standard** for comparison.

5. Add 10 drops of 1M NaOH into test tube **2**. Record your observations concerning the color change in Table 1 of the data section. Use the juice in the test tube labeled **standard** for comparison.

6. Use pH paper to determine the pH of the solution in each of the 3 test tubes. Record the pH values in Table 1.

7. Add drops of 1M NaOH to test tube **1** until it returns to its original color. Record your observations in Table 2.

8. Add drops of 1M HCl to test tube **2** until it returns to its original color. Record your observations in Table 2.

9. Use pH paper to test the pH of the 3 solutions. Record the pH values in Table 2.

Part 2: Acid / Bicarbonate Reaction

1. Weigh out 0.5 grams of sodium bicarbonate into a 100 mL beaker.

2. Add 3 mL of distilled water to the beaker and stir until most of the sodium bicarbonate is dissolved.

3. Add ~3 drops of universal indicator to the beaker. Compare the color of the solution to the pH paper indicator chart. Record the color and corresponding pH in Table 3.

4. Measure 9 mL of vinegar into a graduated cylinder.

5. Add ~3 drops of the universal indicator to the vinegar. Compare the color of the solution to the pH paper indicator chart. Record the color and corresponding pH in Table 3.

6. Use a dropper to add 3 mL of the vinegar in the 10 mL graduated cylinder into the sodium bicarbonate solution. Note the color change and pH in data Table 3 under mixture.

7. Add 3 mL increments of the vinegar into the sodium bicarbonate and vinegar mixture two additional times. Note the color and pH in Table 3 under "additional vinegar" each time.

Teaching Alternative*: *If laboratory periods are short some test items can be eliminated. However, to maintain a range of pH values, only the following items should be eliminated: tomato juice, milk, detergent, and/or water.

Part 3: Testing acidity and basicity of common household products

1. Use the droppers provided to place into different wells of your spot plate a couple of drops of each of the following items: tomato juice, household ammonia, milk, vinegar, lemon juice, and dilute dishwashing liquid. Be sure to label or write down where each item is located in the spot plate. **CAUTION: Do not contaminate the items being tested. Be sure to use only the designated dropper for each item.**

Helpful HINT: Tape a small test tube on each item to be tested in Part 3. Place a dropper for the students to use to obtain their samples in each of the test tubes.

2. Guess the pH of each of the items before you find the experimental value and record your guess in Table 4.

3. Test each item with litmus paper and pH paper. Record your results in Table 4.

4. To clean up rinse all chemicals into a waste beaker. Neutralize the waste to a pH between 4 – 8 using either baking soda or vinegar. Wash the waste solution down the drain.

Date:________ Name: ___________________________ Period:_________

Lab Partner: ____________________________________

Data

Table 1: Acid-Base Neutralization

	Test tube 1	Test tube 2	Standard
Step 1	Add acid	Add base	Neutral
Color	pink	green	Pink
pH paper	3	8	4

Table 2: Acid-Base Neutralization

	Test tube 1	Test tube 2	Standard
Step 1	Add base	Add acid	Neutral
Color	Pink	Pink	Pink
pH paper	3	4	4

Table 3: Acid Base Reactions

Substance	Color	pH
Sodium bicarbonate	Blue green	10
Vinegar	Orange	3
Mixture	Light green	8
First Additional vinegar	Yellow	5

Second Additional vinegar	*Orange*	*3*
Third Additional vinegar	*Orange*	*3*

Table 4: Testing for acidity and basicity in common household products

Product	Hypothesized pH	Color of Litmus Paper	Color of pH Paper	Actual pH
Ammonia	*11*	*blue*	*dark blue*	*11*
Tomato juice	*4*	*Light orange*	*Orange*	*3*
Milk	*9*	*purple*	*light green*	*7*
Lemon juice	*2*	*pink*	*Orange*	*2*
Vinegar	*4*	*pink*	*Light orange*	*3*
Dish detergent	*8*	*blue*	*Green*	*8*
Water	*7*	*light pink*	*Yellow-green*	*6*

Post Lab Questions

1. Why did the grape juice change color when an acid or base was added?

   The color changed due to a change of pH.

2. You added a base, NaOH, to test tube **1** that contained HCl and an acid to test tube **2** that contained base. Why did the grape juice return to its original color?

The original color of the solutions returned, because a neutralization reaction occurred in both of them.

3. In part 2 of the procedure universal indicator was added in the reaction. Why was it added?

The universal indicator was added to indicate how the pH of the solution was affected by the neutralization of the reaction.

4. Name 2 acids and 2 bases you often use.

Some examples of acids are batteries, sour candies, orange juice, and carbonated drinks. Some examples of bases are milk, soap, and medications for stomach acid (TUMS®).

Lab

23

Titration of Acidic Candy

*TN Standard **4.2**: The student will investigate the characteristics of acids and bases.*

When do you use something known to find an unknown?

Every day we face things that are unknown to us, and every day we use what we do know to determine the unknown. When you wake up in the morning and see that it is cloudy, you speculate that you might need an umbrella because there is a possibility of rain. You do not know that it is going to rain, but you do know that it is cloudy and many cloudy days often lead to rainy days. You use what you know to determine what you do not know. In the same way, in an acid/base titration, we determine the unknown concentration of an acid or base with a known concentration of the opposite acid or base.

Introduction

For this laboratory exercise candy will be used, not to eat, but to determine the concentration of an acid in a candy. In general, to find an unknown one must know how it relates to a known. Remember when acids and bases are combined, a neutralization reaction occurs. To determine the concentration of acid in an acidic candy solution, a base of known concentration, called a **standard solution**, is added in small quantities until complete neutralization occurs. When neutralization occurs, the **equivalence point** of the titration has been reached. This procedure is called an **acid-base titration**.

When the standard solution is reacted with an acid or base how will one know when the neutralization reaction is complete? An additional substance is required for a titration - an acid-base indicator. An indicator provides a sudden color change when a certain pH is reached. The pH at which the indicator changes color, known as the **end point** of the titration, depends on the chemical nature of the indicator. An indicator for a titration should be chosen that changes color at a pH near the equivalence point. At the equivalence point, one can assume that the total number of H^+ ions donated by the acid equals the total number of H^+ ions accepted by the base.

But even more information can be learned from the titrations. You can even calculate how many moles of acid there is per gram of candy if the type of acid in the candy is known. To do this the following equation is used:

$$moles_{acid} = \frac{a \times M_b \times V_b}{b}$$

a – reaction coefficient of the acid

M_b — Molarity of the basic solution

V_b – Volume of base used in titration converted to Liters

b — reaction coefficient of the base

To determine the reaction coefficients of the acid and base you must first balance the reaction equation. The reaction coefficients are just the number in front of the acid or the base in the balanced equation. In this experiment the tart candies you will analyze are Sweet Tarts© which contains malic acid, and Smarties© which contains citric acid. The reactions are shown below.

Sweet Tarts: Malic Acid

(acid ↓) (base ↓) (salt↓)

$HO_2CCH_2CH(OH)CO_2H + 2\ NaOH \rightarrow NaO_2CCH_2CH(OH)CO_2Na + 2H_2O$

Notice that for the Sweet Tarts, 2 molecules of NaOH are required for every one molecule of malic acid. This is because malic acid has 2 acid groups. For this reaction **a** is 1 and **b** is 2.

Smarties: Citric Acid

(acid ↓) (base ↓) (salt↓)

$HOC(CO_2H)(CH_2CO_2H)_2 + 3\ NaOH \rightarrow HOC(CO_2Na)(CH_2CO_2Na)_2 + 3H_2O$

For Smarties 3 NaOH molecules are required for every one citric acid molecule. This is because citric acid has 3 acid groups.

After the reaction coefficients are determined for a titration, you then calculate how many moles of acid there are in each gram of candy. Here is an example.

Example 1: Calculating the Moles of Acid in Sweet Tarts
If 0.455 g of Sweet Tarts required 11.51 mL of 0.100 M NaOH to reach the endpoint, how many moles of acid is there in each gram of the candy?

Answer: Since 1 molecule of malic acid requires 2 molecules of NaOH, **a = 1** and **b = 2.** Next 11.51 mL of NaOH must be converted to liters of NaOH.

$$\mathbf{V_b} = 11.51 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.01151 \text{ L}$$

$\mathbf{M_b}$ is known to be 0.100 mol/L NaOH.

Solving for moles acid, we get:

$$\mathbf{Moles_{acid}} = \frac{a \times M_b \times V_b}{b} = \frac{(1) \times (0.100 \text{ mol/L}) \times (0.0115 \text{ L})}{2}$$

$$\mathbf{Moles_{acid}} = 5.75 \times 10^{-4} \text{ moles}$$

$$\mathbf{Moles_{acid}/g\ candy} = \frac{5.75 \times 10^{-4} \text{ moles}}{0.455 \text{ g}}$$

$$\mathbf{Moles_{acid}/g\ candy} = 1.26 \times 10^{-3} \text{ moles acid/g Sweet Tarts}$$

In this experiment you will dissolve Sweet Tarts© and Smarties© in water and titrate the solutions with 0.1 M NaOH. The end point will be determined with phenolphthalein indicator. It changes color at approximately pH 8. You will then determine whether Sweet Tarts or Smarties required more NaOH per gram of candy to reach the endpoint. You will also calculate the moles of acid per gram of candy for both Sweet Tarts© and Smarties©.

Objectives

- To understand the process of titration.
- To use titration with a standard NaOH solution to determine whether Sweet Tarts© or Smarties© candies requires more standard base per gram of candy to reach the endpoint.
- To use titration with a standard NaOH solution to determine whether Sweet Tarts© or Smarties© candies has more moles of acid per gram of candy.

Date:________ Name:______________________ Period:_________

Pre-lab Questions

1. What is the difference between equivalence points and end points?

   The equivalence point of the titration is when the neutralization occurs. The end point is the pH at which the indicator changes color.

2. What would happen if you forgot to put the indicator in?

   There would never be a color change so you would not know how many mL of NaOH was needed to neutralize the acid.

3. If 1.578 g of Smarties required 2.81 mL of 0.100 M NaOH to reach the endpoint, how many moles of acid is there in each gram of the candy?

$$Moles_{acid} = \frac{a \times M_b \times V_b}{b} = \frac{(1)\times(0.100\ mol/L)\times(0.00281\ L)}{3}$$

$$Moles_{acid} = 9.36 \times 10^{-5} \text{ moles acid}$$

$$Moles_{acid}/g\ candy = \frac{9.36 \times 10^{-5} \text{ moles}}{1.578 \text{ g Smarties}}$$

$$Moles_{acid}/g\ candy = 5.93 \times 10^{-5} \text{ moles acid/g Smarties}$$

Preparation Materials

General Supplies	Laboratory Equipment	Chemicals
Smarties© (8 per group) (Light colored candies work best)	Balance	0.1 M NaOH Prep: 0.40 g NaOH per 100 mL H_2O
Sweet Tarts© (6 per group) (Light colored candies work best) (2.0 g of Nerds© or one Sweet Tart Shocker© per trial may be substituted)	2 - 250 mL Erlenmeyer flask	Phenolphthalein indicator solution with dropper
	25 mL buret	
	Buret clamp	
	Ring stand	
	Stirring rod	
	2 – 150 mL Beakers	
	Hot plate or Coffee Maker	
	Mortar and Pestle (optional)	

Procedure

> ***Teaching Alternatives:***
> ***If mortar and pestles are available, students may crush the candies first and then weigh and place into the Erlenmeyer flask .***

> ***Helpful HINT:***
> ***To save time, water can be heated before the allotted lab time. A clean empty coffee maker works well. It makes for quick heating and easy pouring directly into Erlenmeyer flask.***

SAFETY FEATURES

- Measure Accurately
- Safety Goggles

MATERIALS NEEDED

- 0.1M NaOH
- 25 mL buret
- Lightly colored Smarties© candies
- Lightly colored Sweet Tart© candies.
- 2 – 150 mL Beakers
- 2 – 250 mL Erlenmeyer Flasks
- Balance
- Hot plate
- Phenolphthalein
- Stirring rod
- Ring stand
- Buret clamp

Part 1: Titration of Sweet Tarts

1. Weigh 3 lightly colored Sweet tarts© and record their mass in the Data Table. Place them into a 250 mL Erlenmeyer flask. **HINT**: Candies that are of a lighter color work much better.

2. Add ~50 mL of deionized water to the flask with the candies. Heat the mixture on a hot plate while crushing and stirring the candy with a stirring rod until the candy is completely dissolved. Allow the solution to cool. Add about 5 drops of phenolphthalein to the Erlenmeyer flask.

3. Attach a buret clamp to a ring stand. Clamp a 25 mL buret in the buret clamp.

4. Record the molarity of the NaOH you are using in the data table under Concentration of NaOH $\mathbf{M_b}$.

5. Pour approximately 50 mL of 0.100 M NaOH into a 150 mL beaker. Place a funnel on top of the buret and a waste beaker below it. Carefully rinse through the buret a few mL of the 0.100 M NaOH. Rinse the buret at least two times.

6. Close the buret stopcock, and fill it to where the bottom of the meniscus or curve touches the 0.00 mL mark. Record this volume as the initial buret reading in the data table. **HINT**: Fill the buret slightly over the 0.00 mL mark, and then use a waste beaker to drain the buret until it is right on the 0.00 mL mark. Make sure there are no air bubbles in the buret tip.

7. Place the Erlenmeyer flask containing the dissolved candy directly under the buret.

8. While swirling the flask, slowly and carefully begin to add 0.100 M NaOH solution to it. The moment that the solution turns pink (or until the solution undergoes a **definite** color change), stop adding the base and swirl the flask. **HINT**: The color may vary depending on the original color of the solution.

9. Continue the titration by adding a drop at a time while swirling the flask. When the whole solution stays pink or the color change persists after swirling, stop the titration.

10. Determine the final buret reading to the 0.00 mL place, and record this in the Data Table under the final buret reading. **HINT**: Remember to take the final buret reading where the bottom of the meniscus is seen. Record this end point value in the Data Table under Trial 1 for the Sweet tarts©.

11. Repeat the procedure for the Sweet tarts©. Record the data in the Data Table under Trial 2 for the Sweet tarts©.

12. Do two trials for the Smarties© candies. Follow the same procedure only use 4 Smarties© candies instead of 3.

Date:________Name:__________________________ Period:_________

Lab Partner: ___________________________________

Data

Table 1: Titration data (Use calculations section to help complete table)

	Sweet Tarts Trial 1	Sweet Tarts Trial 2	Smarties Trial 1	Smarties Trial 2
Mass of candy (g)	2.88 g	2.88 g	1.52 g	1.51 g
Concentration of NaOH M_b (M) (on the bottle)	0.100 M	0.100 M	0.100 M	0.100 M
Initial buret reading (mL)	0.00 mL	11.00 mL	12.00 mL	15.00 mL
Final buret reading (mL)	11.51 mL	23.50 mL	14.64 mL	17.98 mL
Volume NaOH solution used (mL)	11.51 mL	12.50 mL	2.64 mL	2.98 mL

Calculations

1. Determine the amount of NaOH used per each trial.
HINT: Final – Initial buret reading.

Sweet Tarts Trial 1: ***11.51 – 0.00 = 11.51 mL***

Sweet Tarts Trial 2: ***23.50 – 11.00 = 12.50 mL***

Smarties Trial 1: ***14.64 – 12.00 = 2.64 mL***

Smarties Trial 2: ***17.98 – 15.00 = 2.98 mL***

2. Determine the amount of base (NaOH) per gram of candy.
HINT: Divide the amount of NaOH used by the mass of the sample of candies.

Sweet Tarts Trial 1:

(11.51 mL)/(2.88 g)= 4.00 mL NaOH/g candy

Sweet Tarts Trial 2:

(12.50 mL)/(2.88 g)= 4.34 mL NaOH/g candy

Smarties Trial 1:

(2.64 mL)/(1.52 g)= 1.74 mL NaOH/g candy

Smarties Trial 2:

(2.98 mL)/(1.51 g)= 1.97 mL NaOH/g candy

3. Determine the average of the two trials for the amount of base (NaOH) used per gram of candy.

Sweet Tarts Average:

(4.00 mL + 4.34 mL)/2 = 4.17 mL NaOH/g candy

Smarties Average:

(1.74 mL + 1.97 mL)/2 = 1.86 mL NaOH/g candy

4. Determine the $moles_{acid}$ per gram of candy. **HINT:** This is found by dividing the moles of acid found for each type of candy by the mass of the candy initially weighed out and recorded in the data table. Complete in the calculation in a stepwise process.

- **Moles of acid Sweet Tarts Trial 1:**

  A. Write the balanced equation.

$$HO_2CCH_2CH(OH)CO_2H + 2NaOH \rightarrow NaO_2CCH_2CH(OH)CO_2Na + 2H_2O$$

B. Determine the mol to mol ratio of the acid to base.

1:2 mole ratio

C. Apply the equation. (**HINT:** $moles_{acid} = a \times M_b \times V_b / b$. ***Be sure to convert mL to L.)***

$$Moles_{acid} = \frac{(1)x(0.1\ mol/L)x(11.51 \times 10^{-3})}{2} = 5.76 \times 10^{-4}\ mol$$

D. Divide the number of moles by the mass in grams of the candy.

$$Moles_{acid}/g = \frac{5.76 \times 10^{-4}\ mol\ acid}{2.88\ g} = 2.00 \times 10^{-4}\ mol/g$$

- **Moles of acid in Sweet Tarts Trial 2:**

A. Write the balanced equation.

$$HO_2CCH_2CH(OH)CO_2H + 2NaOH \rightarrow NaO_2CCH_2CH(OH)CO_2Na + 2H_2O$$

B. Determine the mol to mol ratio of the acid to base.

1:2 mole ratio

C. Apply the equation.

$$\text{Moles}_{acid} = \frac{(1) \times (0.1\ mol/L) \times (12.50 \times 10^{-3})}{2} = 6.25 \times 10^{-4}\ mol$$

D. Divide the number of moles by the mass in grams of the candy.

$$\text{Moles}_{acid}/g = \frac{6.25 \times 10^{-4}\ mol\ acid}{2.88\ g} = 2.17 \times 10^{-4}\ mol/g$$

- **Moles of acid in Smarties Trial 1**:

A. Write the balanced equation.

$$HOC(CO_2H)(CH_2CO_2H)_2 + 3NaOH \rightarrow HOC(CO_2Na)(CH_2CO_2Na)_2 + 3H_2O$$

B. Determine the mol to mol ratio of the acid to base.

1:3 mole ratio

C. Apply the equation.

$$\text{Moles}_{acid} = \frac{(1)\times(0.1\ mol/L)\times(2.64 \times 10^{-3})}{3} = 8.80 \times 10^{-5}\ mol$$

D. Divide the number of moles by the mass in grams of the candy.

$$\text{Moles}_{acid}/g = \frac{8.80 \times 10^{-5}\ mol\ acid}{1.52\ g} = 5.79 \times 10^{-5}\ mol/g$$

- **Moles of acid in Smarties Trial 2:**

A. Write the balanced equation.

$$HOC(CO_2H)(CH_2CO_2H)_2 + 3NaOH \rightarrow HOC(CO_2Na)(CH_2CO_2Na)_2 + 3H_2O$$

B. Determine the mol to mol ratio of the acid to base.

1:3 mole ratio

C. Apply the equation.

$$\text{Moles}_{acid} = \frac{(1)\times(0.1\ mol/L)\times(2.98 \times 10^{-3})}{3} = 9.93 \times 10^{-5}\ mol$$

D. Divide the number of moles by the mass in grams of the candy.

$$\text{Moles}_{acid}/g = \frac{9.93 \times 10^{-5}\ mol\ acid}{1.51\ g} = 6.58 \times 10^{-5}\ mol/g$$

Post Lab Questions

1. Which candy required more base per gram of candy? Is this what you expected? Explain your answer.

The Sweet tarts required much more base/gram of candy than the Smarties. Yes, this was expected, because they taste sourer.

2. Which candy contained more moles of acid per gram of candy? Is this what you expected? Explain your answer.

The Sweet tarts contained more moles of acid/gram of candy than the Smarties. Yes, this was expected, because they taste sourer.

3. How would your NaOH/g of candy reported change if there was a large air bubble in the buret tip when you started your titration?

The NaOH/g of candy reported would be higher than expected.

Lab

24

Who Done It?

*TN Standard **2.1**: The student will investigate the characteristics of matter.*

*TN Standard **3.1:** The student will investigate chemical bonding.*

*TN Standard **4.2**: The student will investigate the characteristics of acids and bases.*

A crime has been committed in your chemistry classroom. Can you solve the case?

Someone stole your teacher's most prized coffee mug. It was sitting on your teacher's desk. A list of the suspects is written on the board. You now need to get with the other crime scene investigators in your group to strategize. After your group is together, you need to decide who is going to be in charge. Next, you need to thoroughly analyze the crime scene and brainstorm to determine what needs to be done to solve this crime. Once you decide what to do, assign each task. Be sure to work accurately, and to carefully record all data, information, and observations. After all, it may have to stand up in court. Your group will need to meet periodically to evaluate the information obtained, and to determine what additional information needs to be gathered or additional tests need to be run. When all of the information is collected, your group will need to meet and put the clues together to solve the case.

Note to the Teacher

In this lab, students are to apply what they have learned to solve a crime. You will need to come up with a coffee mug that is stolen off of your desk. Be sure to refer to the coffee mug several times earlier in the semester. You will also have to find several other people who are willing to be suspects, and are available for questioning. This can be other teachers, aides, student assistants, custodians, the school secretary, or even a principal. Write the suspects names on the board. You may want to rope off the crime scene to make it more realistic. Have the students work in groups of around four.

Design into the crime scene several clues that require a chemical test related to ones the students have previously performed. Be sure the needed equipment is readily available. Examples of tests include density, polarity, type of bonds, and pH. You may also want have a microscope available and leave a couple of clues requiring its use.

In one possible crime scenario the thief is not very careful. He/she leaves a Styrofoam cup, a Styrofoam or other type of plastic plate, and a crumpled up note in the trash can next to your desk. Upon investigation the students should find a small hole or crack in the cup. Students may conclude the thief needed a cup to quickly pour their drink into since their cup had sprung a leak. Some food scraps from a sandwich and a liquid (pickle juice, etc.) is present on the plate. The thief leaves a spilled white granular substance on the desk. This could be sugar that was used to sweeten their drink or salt used to season their food. Remind students that determining a substance's identity by tasting it is not considered admissible evidence. A small metal object that looks like a fishing weight is also left on the floor close to the desk as if it fell out of a hole in the thief's pocket or their hand. You may want to leave a hair or thread on the desk if a microscope is available.

The students should be able to question the suspects to find out what each of the suspects typically drinks, if they normally salt their food, if they eat pickles with or on their sandwiches, what type of ink pen they use, and if they have recently been fishing. Possible answers are given that you can match up to your suspects. The number of suspects will depend on your situation. Be sure the suspects know how they are to answer student's questions and that only one of your suspect's answers matches all of the evidence.

Suspect	**Typical Drink**	**Uses salt?**	**Eats pickles?**	**Polar or nonpolar ink?**	**Fishes /type of weight used**
Suspect 1	Black coffee or unsweetened tea	Yes	Yes	Polar	Yes / Lead
Suspect 2	Coke or juice	Yes	Yes	Nonpolar	Yes / Lead
Suspect 3	Coffee or tea with sugar	No	Yes	Polar	No
Suspect 4	Coffee or tea with sugar	Yes	No	Nonpolar	No
Suspect 5	Coffee with cream no sugar	Yes	Yes	Polar	Yes / Non-lead

An example is that students would find the thief uses salt, eats pickles, uses polar ink, and fishes using non-lead weights. From this they would conclude the thief is Suspect 5.

Crime Investigation Information

Supervisor ____

Investigators____

____

____

____

Description of the Crime and the Crime Scene:

Notes from Interviews:

Task 1

Description

Assigned to ____

Task 2

Description

Assigned to ____

Task 3

Description

Assigned to ____

Task 4

Description

Assigned to ____

Task 5

Description

Assigned to ____

Task 6

Description

Assigned to ____

Test Results:

Task 1 -

Task 2 -

Task 3 -

Task 4 -

Task 5 -

Task 6 -

Thief's Profile:

Thief: ____