Solano Community College

CHEM010 Lab Manual

Intermediate Chemistry

(The original work is from Santa Monica Community College)

Notes:

Experiment Number	Notes
1	Do not do the melting point part of this experiment as it seems very dangerous for introductory chemistry students. Or an option would be to use a hot plate and stir instead
2	
3	
4	
5	
6	
7	
8	
9	
10	Clean out the eudiometers before using because soap creates excess bubbles. Procedure #2 - "Note that this mass should be less than **0.080** grams." A 18 cm copper wire should be enough to wrap the metal strip.
11	
12	We are not doing this one because equilibrium is not covered in this course. Also, some of the reagents aren't very safe for students to handle.

Table of Contents

Name: _____ Chem 10, Section: _____

Prelab Assignment: Introducing Measurements in the Laboratory

1. In Part A of this lab, you will measure the dimensions (length, width, diameter) of several geometric shapes.

a. Using a ruler, you measure the length of a rectangle to be 12.75 cm and the width to be 3.64 cm. Calculate the area of this rectangle (show work), reporting your answer to the correct number of significant figures.

b. What is the formula for the area of a circle? _____

2. In Part B of this lab, you will measure the volume of a sample of water, in milliliters (mL).

a. What <u>two</u> measuring instruments will you use to measure the water volume?

b. Consider the following two volume measurements: 57.7 mL and 57.68 mL. Which of these is the more precise measurement, and why?

3. In Part C of this lab, you will measure the mass of several different items, in grams (g).

a. What are the <u>two</u> types of balances you will use to measure mass?

b. An empty beaker has a measured mass of 29.456 g. When some salt is added to the beaker, the combined mass is 36.176 grams. Calculate the mass of the salt only (show work), reporting your answer to the correct number of significant figures.

4. In Part D of this lab, you will measure the melting point of an unknown solid, in degrees Celsius (°C).

a. Define “melting point".

b. Is melting point a physical or chemical property of matter? _____

c. A student measures the melting point of an unknown compound to be 53.5 °C. She later discovers that the compound is chlorothymol, with a true melting point of 58.8 °C. Calculate her percent error (show work) to the correct number of significant figures. The required formula is on page 3 of the Procedure document.

d. You will use a variety of equipment to measure the melting point of the solid. Sketch a set-up of the equipment on the back of this page, and label all items in your sketch.

Introducing Measurements in the Laboratory

Objectives

The objectives of this laboratory are:

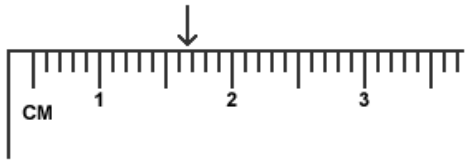
a) To use a metric ruler to measure the dimensions of regular geometric shapes, and to use these measurements to determine the areas of the shapes.
b) To measure the volume of a sample of water using a graduated cylinder and a beaker in order to compare their precision.
c) To measure the mass of an item using a triple-beam balance and an analytical (electronic) balance in order to compare their precision; also, to determine the mass of a powder by weighing by difference.
d) To measure the melting point of an unknown solid and identify it using this measured value.

Background

Our knowledge of chemistry and chemical processes largely depends on our ability to obtain correct information about matter. Often this information is quantitative, in the form of measurements. In this lab, students will be introduced to some common measuring instruments so that they can practice making measurements, and to learn about instrument precision. In Part A of this lab, a metric ruler will be used to measure length in centimeters (cm). In Part B, a beaker and a graduated cylinder will be used to measure liquid volume in milliliters (mL). In Part C, an electronic balance and a triple-beam balance will be to measure mass in grams (g). In Part D, a thermometer will be used to measure temperature in degrees Celsius (°C).

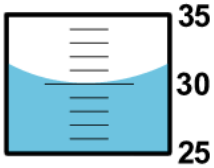
Since all measuring devices are subject to some error, it is impossible to make exact measurements. Scientists record all the digits of a measurement that are known exactly, plus the first one that is uncertain. These digits are collectively referred to as significant digits. Digital instruments, such as an electronic balance, are designed to limit themselves to the correct number of significant digits, and their readings are properly recorded as given. However, when using analog instruments such as rulers and thermometers, the experimentalist is responsible for determining the correct number of significant figures. These instruments are properly read to one place beyond the graduations of the scale.

Example 1: Measuring Length



The ruler markings are every 0.1-centimeter. The correct reading is **1.67 cm**. The first 2 digits **1.6**7 are known exactly. The last digit 1.6**7** is uncertain. You may have instead estimated it as 1.68 cm.

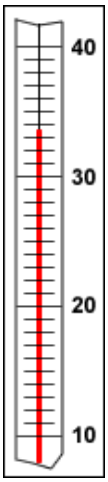
Example 2: Measuring the Volume of a Liquid



When measuring liquid volumes, the graduated scale must be read from the lowest point of the curved surface of the liquid – the liquid meniscus.

The graduated cylinder markings are every 1-milliliter. The correct reading is **30.0 mL**. The first 2 digits **30**.0 are known exactly. The last digit 30.**0** is uncertain. Even though it is a zero, it is significant and must be recorded.

Example 3: Measuring Temperature



Here, the thermometer markings are every 1-degree. The correct reading is **33.6 °C**. The first 2 digits **33**.6 are known exactly. The last digit 33.**6** is uncertain. You may have instead estimated it as 33.5 °C.

Note that the measuring devices used in this lab may have different scale graduations than the ones shown in these examples. Thus, be sure to make it a regular habit to check the scales on all equipment.

When making measurements, it is important to be as accurate and precise as possible. Accuracy is a measure of how close an experimental measurement is to the true, accepted value. Precision refers to the degree of uncertainty in a measurement. For example, a mass measurement of 48.26 g has an uncertainty of ±0.01 g, while a measurement of 48.3 g has an uncertainty of ±0.1 g. Since the measurement of 48.26 g has less uncertainty, it is the more precise measurement. In general, the more decimal places provided by a device, the more precise the measurement will be.

Since measurements are often used in calculations to obtain other values of interest, it is important to consider the number of significant figures that should be recorded for the results of such calculations. If multiplying or dividing measured values, the result should be reported with the lowest number of *significant figures* used in the calculation. If adding or subtracting measured values, the result should be reported with the lowest number of *decimal places* used in the calculation.

Example 4: Significant Figures in Calculated Values

(a) A student runs 18.752 meters in 54.2 seconds. Calculate his average velocity (or speed).

velocity = distance/time
= 18.752 m / 54.2 s
= 0.345978 m/s from calculator
= **0.346** m/s to 3 significant figures

(b) The mass of a glass is measured to be 12.466 grams. If 10.33 grams of water are added to this glass, what is the total combined mass?

total mass = 12.466 g + 10.33 g
= 22.796 g from calculator
= **22.80 g** to 2 decimal places

The temperature that will be measured in this lab is the melting point of an unknown solid. Melting point is a physical property. When a solid is heated continuously, a point will eventually be reached where it undergoes a physical change and becomes a liquid. The temperature at which liquid first appears is defined as the melting point of that substance. Since all pure substances have unique melting points, a measured melting point can be used to identify an unknown substance by comparing it with a list of known substances and their accepted, true melting points.

The accuracy of a measured value, such as a melting point, may be evaluated by a calculation of percent error. Percent error is a common way of reporting how close a measured experimental value (*EV*) is to the true value (*TV*):

$$Percent\ Error = \frac{|EV - TV|}{TV} \times 100$$

Accurate measurements will typically have low percent errors of <5%.

Procedure

Safety

In Part D you will be heating a solid powder and several pieces of equipment with an open Bunsen burner flame. Exercise extra caution while using the Bunsen burner, and please remember that the heated items will be very hot to the touch.

Materials and Equipment

Metric ruler*, shape sheet, electronic balance, large test tube, 100-mL beaker, 100-mL graduated cylinder, triple-beam balance, 250-mL Erlenmeyer flask, electronic balance, sugar, Bunsen burner, thermometer, 400-mL beaker, stand and ring clamp, small watch glass, wire gauze, capillary tube, latex tubing, scoopula and unknown solids.

Part A: Measuring the Dimensions of Regular Geometric Shapes

1. Check out a ruler from the stockroom.
2. Obtain a "shape sheet" from your instructor, and then use the ruler to measure the dimensions of the two geometric shapes on it. Measure the length and width of the rectangle, and the diameter of the circle. Record these measurements on your report form.
3. When finished, return the ruler to the stockroom.
4. Use your measurements to calculate the areas of the assigned geometrical shapes.
   - Area of a rectangle = $l \times w$
   - Area of a circle = πr^2
     (where r = radius = ½ the diameter)

Part B: Measuring the Volume of a Sample of Water

1. Obtain a large test tube from your instructor. Fill this test-tube to the brim with tap water, then carefully transfer it to a 100-mL beaker (obtain from your locker). Note that if your 100-mL beaker has no scale markings on it, you will need to take it to the stockroom and swap it for one that does. Measure and record the volume of water in the beaker.
2. Again, fill the same test-tube to the brim with tap water, then carefully transfer it to a 100-mL graduated cylinder (obtain from your locker). Measure and record the volume of water in the graduated cylinder. Do these measured volumes have the same number of significant figures?

Part C: Measuring the Mass of Solids

Comparing the Precision of two types of Balances

1. Use a triple-beam balance to obtain the mass of a 250-mL Erlenmeyer flask (obtain from your locker).
2. Now use an electronic balance to obtain the mass of the same Erlenmeyer flask. Do these measured masses have the same number of significant figures? Be sure to record your measured masses on your report form.

Weighing by Difference

3. Using the electronic balance again, obtain the mass of a 100-mL beaker. If you already used this same beaker in Part B, make sure that you carefully dry it before weighing it.
4. Add two spoonfuls of sugar to this beaker, using your scoopula. **Do not do this over the balance!** Then obtain the new combined mass of both the beaker and the sugar. Be sure to use the same electronic balance as before.
5. When finished, dispose of the sugar used in the sink.
6. Use your two measurements to determine the mass of sugar (only) weighed out.

Part D: Measuring the Melting Point of an Unknown Solid

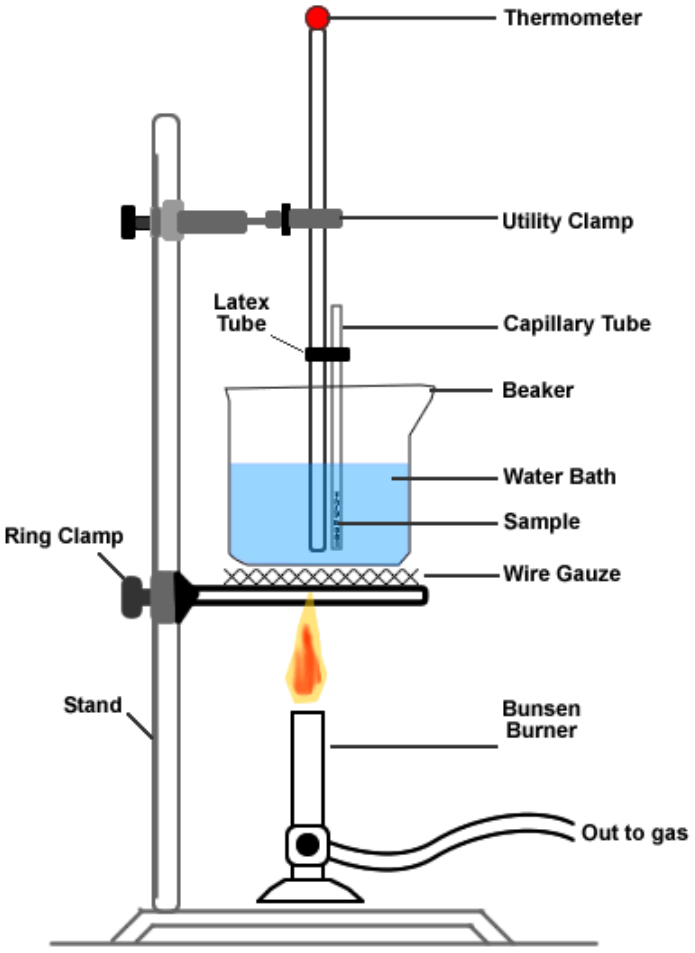
1. Record the ID code of the unknown solid assigned to you and your partner. Powdered samples of the unknown solids are located at the front of lab on the instructor's lab bench.
2. Obtain a capillary tube from your instructor. Press the open end of the capillary tube into the powder. Then turn the tube over and tap the tube lightly against the lab bench to allow the powder to fall into the sealed end. Repeat until you have a depth of about 2-mm of solid in the tube.
3. Assemble your equipment as shown in the diagram on page 5.
    - Use a large 400-mL beaker half-filled with tap water for the hot water bath.
    - Use a small piece of latex tubing like a rubber band to attach the capillary tube to your thermometer. The sealed end should be close to the bulb of the thermometer.
    - Place a slotted stopper around the thermometer, and using a clamp, suspend it in the water bath.
4. Heat the water bath slowly with your Bunsen burner. The flame should be adjusted to a moderate temperature, with its tip touching the bottom of the beaker. Stir the bath continuously and watch your sample carefully.

5. The melting point is the temperature at which liquid first appears. Record your melting point.
6. Share your measured value with all other groups who were assigned the same unknown solid as you. You will also need to obtain and record the melting points that they have measured.
7. Your unknown solid is one of the substances listed in the table below. Identify your solid by comparing your experimental melting point with the true melting points supplied. Then evaluate the accuracy in your measurement by calculating your percent error.

Substance	**Melting Point (in °C)***
L-Menthol	41.6
Benzophenone	48.2
Myristic Acid	54.0
Palmitic Acid	62.0
Stearic Acid	68.0
Vanillin	82.1

* This melting point data was obtained from the NIST Standard Reference Database Number 69 (http://webbook.nist.gov/chemistry/). Please note that organic solids actually melt over a range of temperatures. The melting points given in the table represent the lowest temperature in that range, where liquid formation is first observed. Also note that melting points depend on the purity of the solid.

8. The unknown sample and the capillary tube (together) should be disposed of in the labeled waste container provided when you are finished.



Name: _____ Chem 10, Section: _____

Lab Partner: _____ Experiment Date: _____

Lab Report: Introducing Measurements in the Laboratory

Part A: Measuring the Dimensions of Regular Geometric Shapes

Experimental Data

Shape	Dimensions	Measurement	# Significant Figures
Rectangle	Length		
	Width		
Circle	Diameter		

Data Analysis

1) Perform the conversions indicated below. Show your work, and report your answers in scientific notation.

a. Convert the measured rectangle length to pm.

b. Convert the measured circle diameter to km.

2) Calculate the areas of your rectangle and circle in cm^2. Show your work, and report your answers to the correct number of significant figures.

a. Area of rectangle

b. Area of circle

3) Convert the area of your circle to μm^2. Show your work, and report your answer in scientific notation.

Part B: Measuring the Volume of a Sample of Water

Experimental Data

Measuring Device	Volume Measurement	# Significant Figures
100-mL Beaker		
100-mL Graduated Cylinder		

Data Analysis

1) Compare your volume measurements in the table above. Which instrument, the beaker or the graduated cylinder, provides the more precise measurement? Explain.

2) Convert the volume of water obtained using the graduated cylinder to hm^3. Show your work, and report your answer in scientific notation.

Part C: Measuring the Mass of Solids

Experimental Data

Table 1 – Mass of an Erlenmeyer Flask

Measuring Device	Mass Measurement	# Significant Figures
Triple-Beam Balance		
Electronic Balance		

Table 2 – Weighing by Difference

	Mass Measurement	# Significant Figures
Mass of Empty Beaker		
Mass of Beaker + Sugar		

Data Analysis

1) Compare your mass measurements obtained for the Erlenmeyer flask in Table 1. Which balance, triple-beam or electronic, provides the more precise measurement? Explain.

2) Consider the data you obtained in Table 2.

a. Calculate the mass of sugar weighed out. Show your work.

b. Circle one: When performing the above calculation, significant figures / decimal places are the primary consideration.

3) Perform the conversions indicated below. Show your work, and report your answers in scientific notation.

a. Convert the mass of the sugar weighed out to fg.

b. Convert the mass of the sugar weighed out to Gg.

Part D: Measuring the Melting Point of an Unknown Solid

Experimental Data

Unknown Compound ID Code: _____

Measured by	Melting Point	# Significant Figures
Group 1: You and your partner		
Group 2		
Group 3		
Group 4		
Average Value		

Data Analysis

1) Using the average value above, identify your unknown compound (see Procedure, Part D, #7).

   Name of Compound: _____ True Melting Point: _____

2) Which of the measured melting points recorded in the table was the most accurate? Explain.

3) Calculate the percent error between the experimental melting point that *you and your partner* measured and the substance's true melting point. Report your answer to the correct number of significant figures.

4) Perform the temperature conversions indicated below. Show your work, and report your answers to the correct number of significant figures.

a. Convert the true melting point of your compound to K.

b. Convert the true melting point of your compound to °F.

Name: _____ Chem 10, Section: _____

Prelab Assignment: The Density of Solids and Liquids

1. Circle the correct responses in the following statement:
   Density is a physical / chemical property of matter and an intensive /extensive property of matter.

2. What devices will you use to measure the *mass* and the *volume* of water in Part A of this lab?

3. In Part B of this lab you will perform several measurements in order to determine the density of a metal.

a. Name this metal. _____

b. Describe the technique you will use to measure the volume of this metal.

4. Consider the tabulated data collected by a student for an unknown metal sample. Use this data to calculate the density of the metal (in g/cm^3). Show your work clearly.

Mass of Empty Beaker	44.656 g
Mass of Beaker and Metal sample	124.400 g
Initial volume of water in cylinder	12.7 mL
Final volume of water and Metal sample	21.6 mL

5. In Part C of this lab, you will measure the mass, height and diameter of four cylinders composed of some unknown material.

a. Calculate the volume (in cm^3) of a cylinder with a measured height of 11.76 cm and a diameter of 7.22 cm. Show your work clearly.

b. Each pair of mass and volume values (for each cylinder) will be plotted on a scatter plot, with mass on the y-axis and volume on the x-axis. A best-fit line will then be applied to the plotted data.

- How will you calculate the value of the slope of this best-fit line?

- How will the value of the slope help you identify the unknown material that the cylinders are made of?

The Density of Liquids and Solids

Objectives

The objectives of this laboratory are:

a) To determine the density of pure water;
b) To determine the density of aluminum (applying the technique of water displacement) and to use this value to determine the thickness of a piece of aluminum foil;
c) To measure the mass and volume (via measured dimensions) of several cylinders of an unknown material, and to determine the density of this material via graphical analysis of the collected data.

Background

Density

Density is defined as the mass per unit volume of a substance, and it is a physical property of matter. A physical property can be measured without changing the chemical identity of the substance. Since pure substances have unique density values, measuring the density of a substance can help identify that substance. Density is determined by dividing the mass of a substance by its volume:

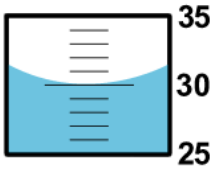
$$Density = \frac{Mass}{Volume}$$

The units of density are commonly expressed as g/cm^3 for solids, g/mL for liquids, and g/L for gases.

Density is also an intensive property of matter. This means that the value of density is independent of the quantity of matter present. For example, the density of a gold coin and a gold statue are the same, even though the gold statue consists of the greater quantity of gold. This is in contrast to extensive properties, like volume (the amount of space occupied by matter), which depend of the quantity of mater present. The more matter present, the larger the volume.

In Part A of this lab, the mass and volume of distilled water will be measured in order to determine the density of water. Measurements will be performed on three samples of water to improve precision and accuracy. Mass will be measured with an electronic balance, in grams (g), and volume will be measured directly with a graduated cylinder, in milliliters (mL). Recall that when measuring liquid volumes, the graduated scale must be read from the lowest point of the curved surface of the liquid (the meniscus).

Measuring the Volume of a Liquid

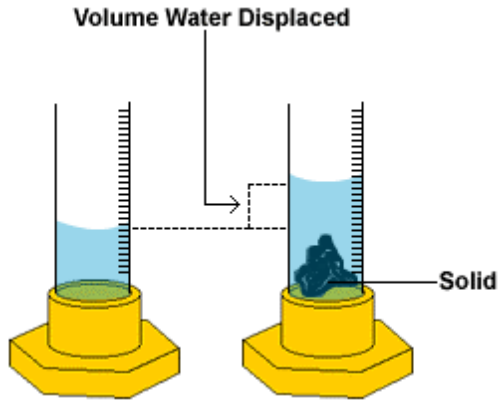


The graduated cylinder markings are every 1-milliliter. When read from the lowest point of the meniscus, the correct reading is **30.0 mL**. The first 2 digits **30**.0 are known exactly. The last digit 30.**0** is uncertain. Even though it is a zero, it is significant and must be recorded.

The accuracy of the experimentally determined density of water will then be evaluated by comparison to the true, accepted density of water.

In Part B of this lab, the density of aluminum will be determined using aluminum pellets. Again, mass will be measured using an electronic balance, in grams (g). However, since the pellets have irregular shapes, their volume must be measured indirectly using the technique of water displacement (also known as Archimedes Principle). This is because the volume of water that the solid displaces when it is immersed in the water is the same as the volume of the solid itself. The accuracy of this experimentally determined density will also be evaluated by comparison to the true, accepted density of aluminum.

Measuring the Volume of an Irregularly Shaped Solid



Volume water displaced = Final volume – Initial volume
Volume water displaced = Volume of solid

Note that 1 mL = 1 cm^3.

The density of aluminum will then be used in an applied problem to determine the thickness of a piece of aluminum foil. The piece of foil used can be considered to be a very flat rectangular box, where

$$Volume\ of\ foil = length \times width \times thickness$$

The foil volume can be obtained from the measured mass of the foil and the density of aluminum. Thus, if the length and width of the foil rectangle are measured, then the foil's thickness may be calculated.

Density and Graphical Analysis

Laboratory investigations involve collecting data, which is often numeric. One common method of interpreting data is graphical analysis.

In Part C of this lab, the mass and volume of several cylindrical pieces of an unknown solid material will be measured. Once again mass will be obtained using an electronic balance, in grams (g). But since the cylinders are regularly-shaped solids, their volumes (in cubic centimeters, cm^3) will be calculated from their measured dimensions by using the appropriate volume formula:

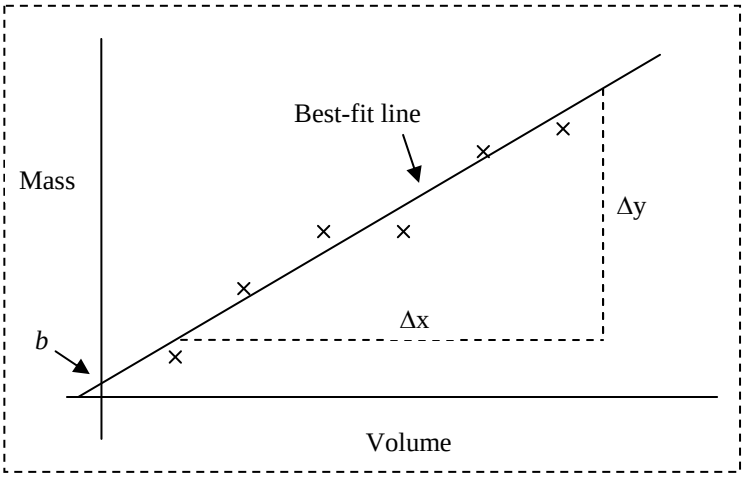
$$Volume\ of\ a\ cylinder = \pi r^2 h$$

h = cylinder height or length

r = cylinder radius = ½ the diameter

Each pair of mass and volume values will then be plotted on graph paper as a scatter plot, with mass plotted on the y-axis and volume plotted on the x-axis. Since the plotted data generate (or at least approximate) a straight line, a "best-fit line" can be added to the graph. A best-fit line is a single line that comes as close as possible to all the plotted points.

The equation of this best-fit line will have the familiar form $y = mx + b$, where m represents the slope of the line, and b represents the y-intercept. This is illustrated in the figure below.



Best-fit line equation: y = *m*x + *b*

b = y-intercept

m = slope

The y-intercept (*b*) is the point on the y-axis where the line crosses the axis. In this experiment, the value of *b* should be equal to zero. This is because if there is no mass, the volume must also be zero. However, note that your best-fit line might not pass exactly through the origin (0,0) due to experimental error – but it should be quite close.

The slope of the line (*m*) is the change in the y-axis values divided by the change in x-axis values (or, rise over run):

$$m = \frac{\Delta y}{\Delta x} = \frac{y_1 - y_2}{x_1 - x_2}$$

Since Δy is really the change in mass (Δmass), and Δx is really the change in volume (Δvolume), this means that the slope of the best-fit line yields the density of the unknown material:

$$m = \frac{\Delta y}{\Delta x} = \frac{\Delta mass}{\Delta volume} = density$$

Once the density is determined in this manner, it will be used to identify the unknown material analyzed.

Procedure

Safety

Be especially careful when adding the aluminum to your graduated cylinder, as the glass could break. Tilt the graduated cylinder and allow the pellets to gently slide to the bottom.

Materials and Equipment

100-mL graduated cylinder, metric ruler*, aluminum pellets, small beaker, aluminum foil, thermometer, electronic balance, distilled water, tube of unknown solid cylinders* and graph paper.

Part A: The Density of Water

1. Using the electronic balance, obtain the mass of your 100-mL graduated cylinder. Make sure it is dry before you weigh it.
2. Add 20-25 mL of distilled water to the graduated cylinder. Precisely measure this volume of water. Then measure the combined mass using the electronic balance.
3. Add another 20-25 mL of distilled water to the graduated cylinder. Again, precisely measure this volume of water, and then measure the combined mass using the electronic balance.
4. Repeat Step 3 to obtain a third set of mass and volume measurements.
5. Use your thermometer to record the temperature of the water in your graduated cylinder.
6. Analysis: Subtract the mass of the empty cylinder from each combined mass measurement to obtain three mass measurements of water. Use the three sets of mass and volume measurements to calculate three density values for water. Then take the average of these three density values. Finally, look up the true density of water at the temperature used, and evaluate the accuracy of your average density value by calculating your percent error.

Part B: The Density of Aluminum and the Thickness of Foil

The Density of Aluminum

1. Using the electronic balance, obtain the mass of a clean, dry small beaker.
2. Obtain a sample of aluminum from your instructor. Transfer all the pellets to the beaker, and measure the mass of the beaker and pellets.
3. Pour 30-35 mL of water into your 100-mL graduated cylinder. Precisely measure this volume.
4. Carefully add all the aluminum pellets to the water, making sure not to lose any water to splashing. Also make sure that the pellets are all completely immersed in the water. Measure the new volume of the water plus the pellets.
5. When finished, retrieve and dry the aluminum pellets and return them to your instructor.
6. Analysis: Use your measured mass and volume (obtained via water displacement) of the aluminum pellets to calculate the density of aluminum. Then look up the true density of aluminum and evaluate your accuracy by calculating your percent error.

The Thickness of Aluminum Foil

7. Now obtain a rectangular piece of aluminum foil from your instructor. Use the ruler to measure the length and width of the piece of foil.
8. Measure the mass of the foil using the electronic balance.
9. When finished, return the foil to your instructor and the ruler to the stockroom.
10. Analysis: Use these measurements along with the density of aluminum to calculate the thickness of the foil.

Part C: Graphical Analysis of Mass and Volume Data of an Unknown Solid

1. Check out a ruler from the stockroom and obtain a tube containing cylindrical pieces of an unknown solid material from your instructor. Record the ID Code of the unknown solid on your report form.
2. Using the ruler, measure the dimensions (diameter and height) of each cylindrical object. Start with the smallest object first and progress in order of increasing object size.
3. Measure the mass of each cylindrical object using an electronic balance. Again, start with the smallest object first and progress in order of increasing object size.
4. Replace all the objects in the tube and return the tube to your instructor.
5. Analysis: Use the measured dimensions to calculate the volume of each solid object. Then, on the graph paper supplied, plot the mass (Y) versus the volume (X) of each measured object. Add a best-fit line to this plot. Calculate the slope of this line, which is the density of the unknown solid. Then use this density to identify the unknown material analyzed. Your unknown material is one of the substances listed in the table below.

Substance	**Density (g/cm^3)**
Polyvinylchloride (PVC)	1.35
Maple	0.77
Acrylic	1.16
Polytetrafluoroethylene (Teflon)	2.20
Polypropylene	0.90
Aluminum	2.71
Polyurethane	1.23

Name: _____ Chem 10, Section: _____

Lab Partner: _____ Experiment Date: _____

Lab Report: The Density of Liquids and Solids

Part A: The Density of Water

Experimental Data

	1st Water Addition	2nd Water Addition	3rd Water Addition
Mass of Empty Cylinder			
Mass of Cylinder + Water			
Mass of Water only			
Volume of Water			
Density of Water			
Average Density of Water			

Temperature of Water: _____

Data Analysis

1) Look up the true density of water at the temperature recorded: _____
Obtain this value from http://jupiter.plymouth.edu/~jsduncan/courses/2012_Spring/Techniques/Exams/DensityOfWater-vs-Temp.pdf. Then use this to calculate the percent error in your average density of water. Show your work.

Part B: The Density of Aluminum and the Thickness of Foil

Experimental Data

Table 1 – The Density of Aluminum

Mass of Empty Beaker	
Mass of Beaker and Al pellets	
Mass of Al pellets	
Initial volume of water in cylinder	
Final volume of water and Al pellets	
Volume of Al pellets	

Table 2 – The Thickness of Aluminum Foil

Mass of Al Foil	
Length of Al Foil	
Width of Al Foil	

Data Analysis

1) Use your measured mass and volume of the pellets (in Table 1) to calculate the density of aluminum, in g/cm^3. Show your work, and report your answer to the correct number of significant figures.

2) Look up the true density of aluminum at http://www.chemicool.com: _____
Use this to calculate the percent error in your experimentally determined density value. Show your work.

3) Use your measurements for the aluminum foil (in Table 2) along with the true density of aluminum to calculate the foil thickness, in cm. Show your work, and report your answer in scientific notation. Consider the foil to be a very flat rectangular box, where: Volume of foil $= length \times width \times thickness$

Part C: Graphical Analysis of Mass and Volume Data of an Unknown Solid

Experimental Data

ID Code of Unknown Solid: _____

	Small Cylinder	Medium Cylinder	Large Cylinder	EX Large Cylinder
Mass				
Length				
Diameter				
Calculated Volume				

Show a sample calculation for volume using your measured dimensions for the small cylinder below.

Data Analysis

1) On the graph paper supplied, plot "Mass (Y) versus Volume (X)" for all four cylinders measured. Staple your graph to this report form.

Instructions for Graphing Data
- Use a sharpened pencil. - Use a ruler to draw your axes. - Choose axis scales that make use of the entire sheet of graph paper. - Clearly number and label your axes. - Use "X" symbols for each plotted point. - Draw a best-fit straight line through your data points using a ruler. - Give your graph an appropriate title.

2) Choose two points on your best-fit line separated far from each other. The points chosen cannot be any of your plotted data points. Circle the two points selected on your graph, and complete the table below.

	X Value	Y Value
Point 1		
Point 2		

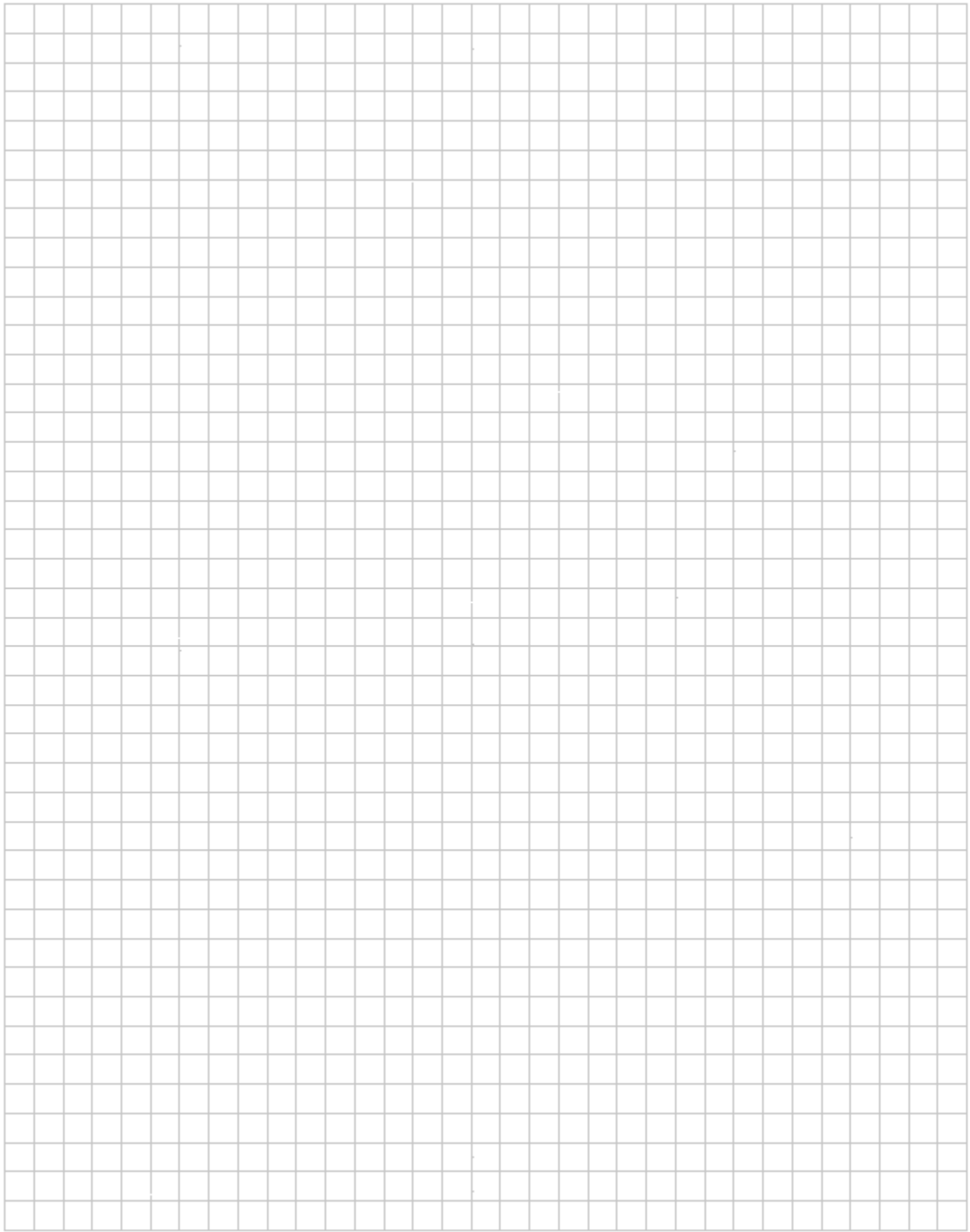
Now calculate the slope (m) of your best-fit line using the equation: $m = y_2\text{-}y_1/x_2\text{-}x_1$. Show your work, and report your result to the correct number of significant figures.

3) The value of the slope obtained above in #2 yields the density of your unknown solid, in units of g/cm^3. Using this value, identify your unknown solid (see Procedure, Part C, #5).

ID Code of Unknown: _____

Name of Unknown: _____ True Density: _____

4) You are supplied with another cylinder made of the same material. If the cylinder length is 1.83 feet, calculate the mass of this cylinder, in grams. Use the true density of the solid in this calculation, and assume that the cylinder diameter is the same as all the other cylinders you measured. Show your work.



Nomenclature of Ionic Compounds

Ionic compounds are composed of ions. An ion is an atom or molecule with an electrical charge. Monatomic ions are formed from single atoms that have gained or lost electrons. Polyatomic ions are formed from molecules (groups of atoms bonded together) that have gained or lost electrons.

Negative ions are called anions, and are formed when an atom or molecule gains electrons. All non-metals form negatively charged ions. Positive ions are called cations, and are formed when an atom or molecule loses electrons. All metals form positively charged cations. Ions with opposite charges (positive metal cations and negative non-metal anions) will experience a strong electrostatic attraction and form an ionic bond, which leads to the formation of the ionic compound.

Non-metal Anions

Non-metals will form anions with only one possible negative charge. The following Periodic Table shows the charges for non-metal anions commonly found in ionic compounds:

1A	2A											3A	4A	5A	6A	7A	8A
																H^{-1}	
												B	C^{-4}	N^{-3}	O^{-2}	F^{-1}	
													Si	P^{-3}	S^{-2}	Cl^{-1}	
														As	Se^{-2}	Br^{-1}	
															Te	I^{-1}	
																At	

Note that

a) The magnitude of the negative charges on these anions is equal to 8 minus their Group Number.

b) The names of these anions are based on the element names, but the endings are all changed to –ide.

Example 1: Sulfur in Group 6A forms anions with a [8–6] = 2, or, **–2 charge**.

Example 2: The Cl^{-1} anion is called the chlor**ide** ion.

Metal Cations

Most (but not all) main group metals will form cations with only one possible charge. Most (but not all) transition metals will form cations with more than one possible charge. The following Periodic Table shows the charges for metal cations commonly found in ionic compounds:

1A	2A	Transition Elements (B)										3A	4A	5A
H^{+1}														
Li^{+1}	Be^{+2}													
Na^{+1}	Mg^{+2}											Al^{+3}		
K^{+1}	Ca^{+2}		Ti^{+2}, Ti^{+4}		Cr^{+2} Cr^{+3} Cr^{+6}	Mn^{+2} Mn^{+3} Mn^{+4}	Fe^{+2} Fe^{+3}	Co^{+2} Co^{+3}	Ni^{+2} Ni^{+3}	Cu^{+1} Cu^{+2}	Zn^{+2}	Ga^{+3}	Ge	
Rb^{+1}	Sr^{+2}									Ag^{+1}	Cd^{+2}	In^{+1} In^{+3}	Sn^{+2} Sn^{+4}	Sb
Cs^{+1}	Ba^{+2}									Au^{+1} Au^{+3}	Hg_2^{+2} Hg^{+2}		Pb^{+2} Pb^{+4}	Bi^{+3} Bi^{+5}

Note that

a) The magnitude of the positive charge on the main group metal cations is generally equal to their Group Number.

b) The names of metal cations with only one possible charge are the same as the names of the metals themselves.

c) For metal cations with more than one possible charge, the ion charge must be indicated in the ion name. In the IUPAC system, the ion charge is indicated in the name as Roman numerals in brackets.

Example 3: Magnesium in Group 2A forms cations with a **+2 charge**.

Example 4: Al^{+3} is called the **aluminum** cation.
Ag^{+1} is called the **silver** cation.

Example 5: Pb^{+2} is called the **lead(II)** cation.
Pb^{+4} is called the **lead(IV)** cation.

Polyatomic Ions

Polyatomic ions are formed from molecules (groups of atoms bonded together) that have gained or lost electrons. The table below includes a list of common polyatomic ions that must be memorized.

OH^{-1}	Hydroxide	O_2^{-2}	Peroxide
CN^{-1}	Cyanide	CO_3^{-2}	Carbonate
SCN^{-1}	Thiocyanate	SO_3^{-2}	Sulfite
HCO_3^{-1}	Bicarbonate (Hydrogen Carbonate)	SO_4^{-2}	Sulfate
HSO_3^{-1}	Bisulfite (Hydrogen Sulfite)	$S_2O_3^{-2}$	Thiosulfate
HSO_4^{-1}	Bisulfate (Hydrogen Sulfate)	$C_2O_4^{-2}$	Oxalate
$C_2H_3O_2^{-1}$	Acetate	CrO_4^{-2}	Chromate
NO_2^{-1}	Nitrite	$Cr_2O_7^{-2}$	Dichromate
NO_3^{-1}	Nitrate		
MnO_4^{-1}	Permanganate	PO_3^{-3}	Phosphite
ClO^{-1}	Hypochlorite	PO_4^{-3}	Phosphate
ClO_2^{-1}	Chlorite		
ClO_3^{-1}	Chlorate	NH_4^{+1}	Ammonium
ClO_4^{-1}	Perchlorate	Hg_2^{+2}	Mercury (I)

Note that

a) Almost all the polyatomic ions are negatively charged anions.

b) Most of the names of polyatomic anions end in either –ate or –ite. The –ate's always have one more oxygen than the -ite's.

Example 6: Sulf**ate**, SO_4^{-2}, has one more oxygen atom present than sulf**ite**, SO_3^{-2}.

Formulas and Names of Ionic Compounds

Ionic compounds are formed when positive cations and negative anions are attracted to each other via strong electrostatic forces. This attraction is called an ionic bond. The following are the basic rules for writing the formulas and names of ionic compounds:

Writing Formulas

1. Determine the formulas and charges on the cation and anion involved in the compound.
2. Combine the ions in a ratio that results in the formation of a neutral ionic compound. In other words, the total charge of all the positive cations must equal the total charge of all the negative anions in the compound. The numbers of each element present in the compound are shown as subscripts after the element symbol.

Writing Names

1. Both the cation and anion must be named.
2. Always name the cation first, then the anion.

Example 7: Write the formula and name for the compound formed between calcium and fluorine.

Ca (metal) forms a +2 cation Ca^{+2} the **calcium** cation.
F (non-metal) forms a -1 anion F^{-1} the **fluoride** anion.

To obtain a neutral compound, **1** Ca^{+2} is needed for every **2** F^{-1}

The formula of the compound is **CaF_2**

The name of the compound is **Calcium Fluoride**

Example 8: Write the formula for iron(III) chloride.

First identify the cation and the anion in this compound.

Cation = iron(III) = Fe^{+3} (transition metal cation)
Anion = chloride = Cl^{-1} (non-metal anion)

To obtain a neutral compound, **1** Fe^{+3} is needed for every **3** Cl^{-1}

The formula of the compound is **$FeCl_3$**

Example 9: Write the formula for magnesium phosphate.

First identify the cation and anion in this compound.

Cation = magnesium = Mg^{+2} (metal cation)
Anion = phosphate = PO_4^{-3} (polyatomic anion)

To obtain a neutral compound, **3** Mg^{+2} are needed for every **2** PO_4^{-3}

The formula of the compound is **$Mg_3(PO_4)_2$**

Note in the above Example 9 that parentheses are placed around the polyatomic portion of compound, to indicate that it must be treated as a complete and whole unit.

Example 10: Name the ionic compound $Al(NO_3)_3$.

First identify the cation and anion in this compound.

Cation = Al^{+3} = the aluminum cation (metal cation)
Anion = NO_3^{-1} = the nitrate anion (polyatomic anion)

The name of this compound is **Aluminum Nitrate**

Example 11: Name the compound TiO_2.

First identify the cation and anion in this compound.

Cation = Ti^{+4} = the titanium(IV) cation (transition metal cation)
Anion = O^{-2} = the oxide anion (non-metal anion)

The name of the compound is **Titanium(IV) Oxide**

Nomenclature of Simple Covalent Compounds

Covalent compounds are compounds formed between non-metals only. Simple binary covalent compounds contain just two different types of non-metal elements. When non-metals combine they can form several different covalent compounds. These compounds must therefore be identified with unique names and formulas.

Example 12: Carbon and oxygen combine to form two common covalent compounds CO_2 and CO.

Formulas and Names of Simple Covalent Compounds

1. Always write/name the element with more metallic character first. Metallic character increases going from right to left, and top to bottom on the Periodic Table.
2. Then write/name the second (less metallic) element, changing the ending of its name to -ide.
3. Since nonmetals often combine in different proportions to form a number of different compounds, prefixes must be included in the names to indicate the numbers of each kind of atom present. Prefixes for 1-10 atoms are given in the following table.

Number	**Prefix**	**Number**	**Prefix**
1	Mono	6	Hexa
2	Di	7	Hepta
3	Tri	8	Octa
4	Tetra	9	Nona
5	Penta	10	Deca

Example 13: A compound contains 3 atoms of sulfur and 4 atoms of phosphorus. Write its name and formula.

Since **phosphorus is the more metallic element** (left of sulfur), it must be named first. Sulfur being the less metallic element is named second with an -ide ending = **sulfide**.

The prefix for **three** is **tri**, and the prefix for **four** is **tetra**.

The name of this compound is **tetraphosphorus trisulfide**

The formula of this compound is $\mathbf{P_4S_3}$

Example 14: Write the name of the compound N_2O.

The **two N atoms** require the prefix **di** in the name.

The **one O atom** requires the prefix **mono** in the name. The ending of oxygen must be changed to –ide = **oxide**.

The name of this compound is **dinitrogen monoxide**

There are two important exceptions to the naming rules outlined so far:

- Never use the prefix "mono" for the first element, even if just one atom is present.
- Never use any prefixes at all for simple covalent compounds containing Hydrogen.

Example 15: Write the name of the compound BCl_3.

Although just one B atom is present, the prefix **mono is not used** since it is the first element in this compound.

The **three Cl atoms** require the prefix **tri** in the name. The ending of chlorine must be changed to –ide = **chloride**.

The name of this compound is **boron trichloride**

Example 16: HF is **hydrogen fluoride** not hydrogen monofluoride or monohydrogen monofluoride.

Please note that many simple covalent compounds have common, rather than systematic names. Please memorize the common names of the following three compounds:

- H_2O water
- NH_3 ammonia
- CH_4 methane

Covalent compounds containing more than two non-metal elements become increasingly more difficult to name, and common names for these compounds are more extensively used. You will not have to learn these yet.

Nomenclature of Acids

Acids are compounds that release hydrogen cations (H^{+1}) when dissolved in water. They are all found in the aqueous state (*aq*).

$HCl\ (aq)$	$\rightarrow$	$H^{+1}\ (aq)$	+	$Cl^{-1}\ (aq)$
hydrochloric acid		hydrogen ions		chloride ions
$HNO_3\ (aq)$	$\rightarrow$	$H^{+1}\ (aq)$	+	$NO_3^{-1}\ (aq)$
nitric acid		hydrogen ions		nitrate ions
$H_2SO_3\ (aq)$	$\leftrightarrows$	$2\ H^{+1}\ (aq)$	+	$SO_3^{-2}\ (aq)$
sulfurous acid		hydrogen ions		sulfite ions

In acids the element hydrogen actually behaves like a Group 1A metal cation. Since it behaves like a +1 cation, hydrogen is always written first in the formulas of all acids. The anion in the acid can be either monatomic or polyatomic, and affects how the acid is named.

Acids containing Non-Metal Anions

These acids contain the H^{+1} cation and a monatomic non-metal anion.

Example 17: Hydrochloric acid, HCl (*aq*), contains **H^{+1}** and the monatomic anion, **chloride, Cl^{-1}**.

Acids containing monatomic anions are named using the prefix hydro + the name of the anion with the suffix -ic + the word acid.

Example 18: Name the acid HBr (*aq*).

This acid contains **H^{+1}** and the monatomic anion, **bromide, Br^{-1}**.

The name of this acid is **hydro** + brom-**ic** + **acid** = **hydrobromic acid**

The formulas of these acids are obtained in an identical fashion to regular ionic compounds. The H^{+1} cation and the monatomic anion are combined in a ratio to yield a neutral compound.

Example 19: Write the formula and name for the acid containing the phosphide anion.

This acid contains **H^{+1}** and the **P^{-3}** anion (monatomic). Just combine them together like any other cation and anion: **3** H^{+1} : **1** P^{-3}

The formula of this acid is **H_3P (*aq*)**

H_3P (*aq*) is named **hydro** + phosphor-**ic** + **acid** = **hydrophosphoric acid**

Acids containing Polyatomic Anions

These acids contain the H^{+1} cation and a polyatomic anion.

Example 20: Nitric acid, HNO_3 (*aq*), contains **H^{+1}** and the polyatomic anion, **nitrate, NO_3^{-1}**.

These acids have names that are based on the name of the polyatomic ion in the acid. If the polyatomic ion has the ending -ate, in the acid the ending is changed to -ic + acid. If the polyatomic ion has the ending -ite, in the acid the ending is changed to -ous + acid.

Example 21: Name the acid $HClO_3$ (*aq*).

This acid contains H^{+1} and the polyatomic anion **ClO_3^{-1} = chlorate**.

To name this acid, the ending **-ate** is switched to **–ic**.

The acid $HClO_3$ (*aq*) is thus called **chlor<u>ic</u> acid**

Example 22: Name the acid H_2SO_3 (*aq*).

This acid contains H^{+1} and the polyatomic anion **SO_3^{-2} = sulfite**.

To name this acid, the ending **–ite** is switched to **–ous**.

The acid H_2SO_3 (*aq*) is thus named **sulfur<u>ous</u> acid**

Again, the formulas of these acids are obtained in an identical fashion to regular ionic compounds. The H^{+1} cation and the polyatomic anion are combined in a ratio to yield a neutral compound.

Example 23: Write the formula for oxalic acid.

Oxalic acid must contain (by reverse logic) the polyatomic anion **oxalate = $C_2O_4^{-2}$**, as well as H^{+1} (the cation in all acids).

Once the ions have been identified, just combine them together like any other cation and anion: **2** H^{+1} : **1** $C_2O_4^{-2}$

The formula of oxalic acid is **$H_2C_2O_4$ (*aq*)**

Nomenclature of Hydrates

A hydrate is typically an ionic compound with a certain number of water molecules loosely bound to it.

The general formula of a hydrate is $MX \bullet nH_2O\ (s)$, where M is the cation in the ionic compound, X is the anion in the ionic compound and nH_2O are the water molecules loosely bound to the ionic compound.

Hydrates are named by writing the name of the ionic compound first, followed by the word "hydrate". To indicate the number of water molecules present, prefixes must be used.

Example 24: Name the hydrate $MgSO_4 \bullet 7H_2O$.

$MgSO_4$ is the ionic compound **magnesium sulfate**.
Since there are **seven water molecules** present, the correct prefix to use is **hepta**.

The name of this hydrate is **magnesium sulfate heptahydrate**

Example 25: Write the formula for copper(II) chloride dihydrate.

Copper(II) chloride has the formula **$CuCl_2$**
The prefix **di** indicates that there are **two** water molecules present.

The formula of this hydrate is **$CuCl_2 \bullet 2H_2O$**

The water molecules in a hydrate can be removed with relative ease by heating the hydrate. The ionic compound that remains after heating is called an anhydrous salt.

$$\underset{\text{Hydrate}}{MX \bullet nH_2O\ (s)} \rightarrow \underset{\text{Anhydrous salt}}{MX\ (s)} + \underset{\text{Free Water}}{n\ H_2O}$$

Often the anhydrous salt has a completely different color and texture from the hydrate.

Example 26: Copper(II) sulfate pentahydrate is blue and crystalline, whereas anhydrous copper(II) sulfate is white and powdery.

Name: _____ Chem 10, Section: _____

[Individual exercise] Exercise Date: _____

Chemical Nomenclature

Write the names and formulas for the following inorganic compounds in the spaces provided.

Part 1: Ions and Ionic Compounds

Write formulas/charges or names as appropriate for each of the following monatomic ions.

1. Calcium ion _____
2. Phosphide ion _____
3. Iodide ion _____
4. Gallium ion _____
5. Titanium(IV) ion _____
6. C^{-4} _____
7. Rb^{+1} _____
8. Pb^{+4} _____
9. S^{-2} _____
10. Cr^{+2} _____

Write formulas or names as appropriate for each of the following ionic compounds.

1. Magnesium nitride _____
2. Lithium oxide _____
3. Aluminum sulfite _____
4. Copper(II) bicarbonate _____
5. Sodium nitrate _____
6. SrI_2 _____
7. $Ba_3(PO_4)_2$ _____
8. $(NH_4)_2O$ _____
9. $Fe(ClO)_3$ _____
10. $ZnCrO_4$ _____

Part 2: Covalent Compounds

Write formulas or names as appropriate for each of the following covalent compounds.

1. Dichlorine monoxide _____
2. Disulfur dichloride _____
3. Carbon tetrafluoride _____
4. Phosphorus pentachloride _____
5. Nitrogen tribromide _____
6. AsI_3 _____
7. P_4O_{10} _____
8. Cl_2O_7 _____
9. $SeCl_6$ _____
10. NO _____

Part 3: Acids

Write formulas or names as appropriate for each of the following acids.

1. Hydroiodic acid _____
2. Carbonic acid _____
3. Chlorous acid _____
4. Sulfuric acid _____
5. Phosphorous acid _____
6. HCN *(aq)* _____
7. $H_2C_2O_4$ *(aq)* _____
8. HNO_2 *(aq)* _____
9. $H_2Cr_2O_7$ *(aq)* _____
10. $HMnO_4$ *(aq)* _____

Part 4: Hydrates (optional – check with your instructor to see if you are responsible for this section)

Write formulas or names as appropriate for each of the following hydrates.

1. Magnesium sulfate heptahydrate _____
2. Copper(I) sulfate pentahydrate _____
3. Potassium phosphate decahydrate _____
4. Calcium chloride hexahydrate _____
5. Iron(III) nitrate nonahydrate _____
6. $CoSO_4 \cdot H_2O$ _____
7. $Na_2CrO_4 \cdot 4H_2O$ _____
8. $CuF_2 \cdot 2H_2O$ _____
9. $Sr(NO_3)_2 \cdot 6H_2O$ _____
10. $ZnSO_4 \cdot 7H_2O$ _____

Part 5: Nomenclature of Ionic Compounds, Covalent Compounds and Acids

	Classification	Name or Formula
1. C_3O_2		
2. IF_7		
3. Rb_2CO_3		
4. SnS_2		
5. $Au(CN)_3$		
6. H_2CrO_4 (*aq*)		
7. H_3P (*aq*)		
8. Li_3PO_4		
9. Mg_3N_2		
10. $Ti(C_2H_3O_2)_4$		
11. Fe_2O_3		
12. NaH		
13. Br_3O_8		
14. MnS_2O_3		
15. NH_4NO_2		
16. $Cd(ClO_2)_2$		
17. $Ba(HSO_3)_2$		
18. Cu_2O		
19. $NiBr_3$		
20. $Sr(OH)_2$		
21. Perchloric acid		
22. Potassium permanganate		
23. Calcium hydride		
24. Vanadium(II) bicarbonate		
25. Bismuth(V) nitrate		
26. Rubidium peroxide		

27. Strontium hydrogen phosphite		
28. Hydrofluoric acid		
29. Chromium(III) thiocyanate		
30. Acetic acid		
31. Molybdenum(IV) carbonate		
32. Tetraiodine nonaoxide		
33. Diphosphorus tetrafluoride		
34. Aluminum sulfate		
35. Ammonium hydroxide		
36. Sodium dichromate		
37. Carbon disulfide		
38. Nickel(II) oxalate		
39. Barium selenide		
40. Silver bisulfate		

Questions

1. How are the following types of compounds recognized from their *formulas*?

   Ionic _____

   Covalent _____

   Acid _____

2. When do *parentheses* appear in the formulas of ionic compounds?

3. Do *Roman Numerals* appear in the names of ionic or covalent compounds? Explain why they are used.

4. Do *Greek Prefixes* appear in the names of ionic or covalent compounds? Explain why they are used.

5. What is the relationship between the ***number*** *of hydrogens* in an acid and the ***charge*** *on the anion* that they are combined with?

Name: _____ Chem 10, Section: _____

Prelab Assignment: The Properties of Oxygen Gas

1. Oxygen gas will produced via a decomposition reaction of a certain substance.

a. Name the substance that will be decomposed. _____

b. Name the two products generated by this reaction. _____

2. A catalyst called manganese(IV) oxide, MnO_2, will be used to facilitate the production of oxygen gas. Exactly what does the catalyst do?

3. Carefully read the procedure for producing oxygen gas (Part A) and examine the accompanying figure of the equipment set-up.

a. What type of flask does the decomposition reaction occur in? _____

b. What chemical is added to this flask through the thistle tube? _____

c. What chemical(s) are already in the flask? _____

d. What type of bottles is the oxygen gas collected in? _____

e. After the oxygen is collected, do you store it in these bottles right-side up or upside down? (circle one) Explain why.

4. After generating and collecting the oxygen, you will then investigate its role in combustion reactions

a. Is oxygen a reactant or product in a combustion reaction? _____

b. Are combustion reactions exothermic or endothermic? _____

5. In Part B you will burn a variety of substances in the oxygen gas collected from Part A.

a. Which *one* of these substances must be burned in the <u>hood</u>? _____

b. Which *two* of these substances must be burned in <u>air</u> only? _____

c. Which substance (one only) will be burned by the instructor? _____

The Properties of Oxygen Gas

Objectives

The objectives of this laboratory are:
a) To generate (and collect) oxygen gas via the decomposition of hydrogen peroxide.
b) To investigate the properties of oxygen, particularly as an agent of combustion.

Background

Oxygen is one of the most abundant elements on this planet. Our atmosphere is 21% free elemental oxygen. Oxygen is also extensively combined in compounds in the earths crust, such as water (89%) and in mineral oxides. Even the human body is 65% oxygen by mass.

Free elemental oxygen occurs naturally as a gas in the form of diatomic molecules, O_2 (*g*). Oxygen exhibits many unique physical and chemical properties. For example, oxygen is a colorless and odorless gas, with a density greater than that of air, and a very low solubility in water. In fact, the latter two properties greatly facilitate the collection of oxygen in this lab. Among the unique chemical properties of oxygen are its ability to support respiration in plants and animals, and its ability to support combustion.

In this lab, oxygen will be generated as a product of the decomposition of hydrogen peroxide. A catalyst is used to speed up the rate of the decomposition reaction, which would otherwise be too slow to use as a source of oxygen. The catalyst does not get consumed by the reaction, and can be collected for re-use once the reaction is complete. The particular catalyst used in this lab is manganese(IV) oxide.

Generating Oxygen Gas:

$$\underset{\text{hydrogen peroxide}}{2\ H_2O_2\ (aq)} \xrightarrow{\text{catalyst}} \underset{\text{water}}{2\ H_2O\ (l)} + \underset{\text{oxygen}}{O_2\ (g)}$$

The oxygen gas produced will be collected in bottles by a method known as the downward displacement of water (see figure on page 3). Once collected, several tests will be performed in order to investigate the role of oxygen in several combustion reactions.

A combustion reaction is commonly referred to as "burning". During a combustion reaction, oxygen reacts chemically with the substance being burned. Note that since our atmosphere is roughly 21% oxygen, many substances readily burn in air. Both oxygen and the substance being burned (the reactants) are consumed during the combustion reaction, while new substances (the products) and heat energy are generated. Since heat is produced, this is an exothermic reaction.

Combustion Reactions:

$$\text{Substance being burned} + \text{Oxygen} \rightarrow \text{Products} + \text{Heat}$$

The actual products of a combustion reaction depend on what substance is burned and how much oxygen is present. In general, however, when a pure element burns in oxygen the product is called an oxide. An oxide is a compound containing both the element and oxygen chemically combined together.

Some examples of element combustion are shown below. Several such reactions will be performed using the oxygen gas collected in this lab.

Combustion of an Element:

$$\text{Element} + \text{Oxygen} \rightarrow \text{Oxide of Element} + \text{Heat}$$

$$\underset{\text{carbon}}{\text{C}\,(s)} + \underset{\text{oxygen}}{\text{O}_2\,(g)} \rightarrow \underset{\text{carbon dioxide}}{\text{CO}_2\,(g)} + \text{Heat}$$

$$\underset{\text{mercury}}{2\,\text{Hg}\,(l)} + \underset{\text{oxygen}}{\text{O}_2\,(g)} \rightarrow \underset{\text{mercury(II) oxide}}{2\,\text{HgO}\,(s)} + \text{Heat}$$

Procedure

Safety

First, be sure to exercise caution when using the hydrogen peroxide (H_2O_2) and the hydrochloric acid (HCl) as they can cause chemical burns and skin irritation. If either of these chemicals comes into contact with your skin, immediately rinse with water for a minimum of fifteen minutes and notify your instructor. Second, do not look directly at the burning magnesium. In addition to being very bright, it emits harmful UV radiation that could cause damage to the retina of your eyes.

Materials and Equipment

Materials: 9% Hydrogen peroxide solution, manganese(IV) oxide, wooden splints, candle, sulfur, steel wool, magnesium ribbon, zinc metal and 6M hydrochloric acid

Equipment: 250-mL Erlenmeyer flask, five wide-mouth bottles, four glass 'cover' plates, pneumatic trough, "stopper + thistle tube + tubing" apparatus*, utility clamp, stand, deflagration spoon, crucible tongs, small beaker, medium beaker and a large test tube.

Part A: Generating and Collecting Oxygen Gas

1. Obtain the following equipment:
   - A 250-mL Erlenmeyer flask (locker)
   - The "two-hole stopper + thistle tube + glass tubing + rubber tubing" apparatus (stockroom)
   - Five wide mouth 'gas-collecting' bottles (under sink)
   - Four glass 'cover' plates (front desk)
   - A pneumatic trough (under sink) filled with water to ½ inch above the metal shelf
2. Fill four of the five wide-mouth bottles ***to the brim*** with water (the fifth will be used later). Then gently slide a glass plate over the mouth of each bottle. Make sure that there are no air bubbles at the top of the glass plate.
3. While holding the glass plate with your fore and middle finger, gently invert a bottle and lower it into the water in the pneumatic trough. Remove the glass plate when the mouth of the bottle is below the water level in the pneumatic trough. Repeat this for all four bottles. Place the glass plates aside on a paper towel, as they will be used later.
4. Place one gas-collecting bottle on the metal shelf. Make sure that the mouth of the bottle does not come out of the water.

5. Now focus on your reaction vessel, the Erlenmeyer flask. Add a pea-sized amount of manganese(IV) oxide (the catalyst) to the flask, followed by about 50-mL of tap water.

6. Finally, assemble all your equipment together as demonstrated by your instructor, or as shown in the figure below. Make sure that
   - the end of the thistle tube is completely covered with water at the bottom of the flask,
   - the end of the glass tubing running from the Erlenmeyer flask is inserted under the opening in the bottom of the metal shelf into the gas-collecting bottle (which is full of water),
   - the Erlenmeyer flask is stabilized with a utility clamp.

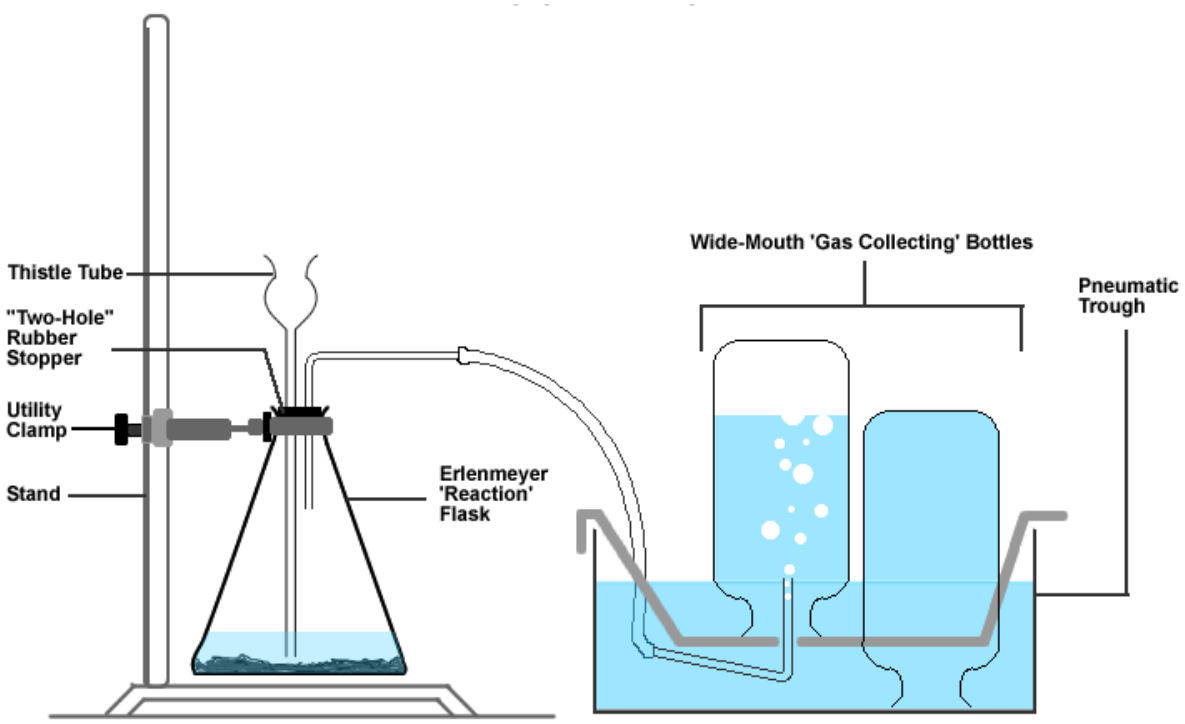
7. Obtain about 30-mL of 9% aqueous hydrogen peroxide (H_2O_2) in your smallest beaker. Then carefully add about 10-mL of this H_2O_2 through the thistle tube. The generation of oxygen gas should begin immediately. If at any time the rate of the reaction in the Erlemeyer flask appears to slow down, add another 10-mL portion of H_2O_2.

8. The oxygen produced will fill the inverted bottle by displacing the water in it. This is because oxygen does not dissolve in water, due to its low solubility. When the first bottle is completely filled with gas, place the second bottle on the metal rack in its place and allow it to fill in a like manner. Repeat this for the third and fourth bottles.

9. As soon as each bottle is completely filled, remove it by placing a glass plate under the bottle's mouth while under water, then lifting the bottle and plate from the pneumatic trough. Place the bottle on the lab bench ***mouth up*** and ***do not remove the glass plate***. Since oxygen is denser than air, it sinks to the bottom of the flask and will not readily leak out the top.

10. Using masking tape, label each bottle of gas in the order they are collected: Bottle #1, Bottle #2, Bottle #3 and Bottle #4. Label the fifth unused empty bottle "Air Bottle".

11. Once all four bottles are filled with oxygen, do not add any more H_2O_2 to the Erlenmeyer flask. Set it aside and allow the reaction to go to completion. At the end of the lab, the chemicals remaining in the reaction flask and any unused H_2O_2 must be disposed of in the labeled waste container in the hood. In the meantime, proceed to Part B.



Part B: The Properties of Oxygen Gas

Dispose of all chemicals used in these tests as indicated by your instructor

Test 1: Combustion of wood

Light a wooden split, and then blow it out. While it is still glowing red, quickly insert the splint into Bottle #1 (oxygen-filled). How many times can you repeat this? Record your observations. Now re-light the same wooden split, and again blow it out. Place it in the empty bottle (air-filled) while it is still glowing. Record your observations.

Test 2: Combustion of candle wax

Place a small candle on a glass plate and light it. Then uncover and carefully lower Bottle #2 (oxygen-filled) over the candle. Measure and record the number of seconds that the candle continues to burn. Then re-light the candle lower the empty bottle (air-filled) over it. Again, measure and record the number of seconds that the candle continues to burn. Also be sure to record any other relevant observations.

Test 3: Combustion of sulfur

This test must be performed in the hood under instructor supervision. Take your Bottle #3 (oxygen-filled) and your empty bottle (air-filled) to the hood your instructor directs you to. Place a small lump of sulfur in a deflagrating spoon (located in the hood). Light the Bunsen burner in the hood, and heat the sulfur in the spoon. The sulfur will first melt, then burn with an almost invisible blue flame. Insert the spoon with the burning sulfur in Bottle #3 and record your observations. Then insert it in the empty bottle, and again record your observations. When finished, extinguish the burning sulfur in the beaker of water provided in the hood.

Test 4: Combustion of iron

Pour about 20-mL of tap water into Bottle #4 (oxygen-filled) and replace the glass plate quickly. Take a loose, frayed out 2-3 centimeter piece of steel wool and hold it in a Bunsen burner flame for a very brief instant with your crucible tongs (it will glow red). Then immediately lower the steel wool into Bottle #4. Record your observations. Repeat with the empty bottle (air-filled) and record your observations.

Test 5: Combustion of hydrogen

This test must be performed in air only. (Note: The hydrogen burned in this test must be first generated by a reaction between zinc and hydrochloric acid.) To a large test tube, add 1-2 pieces of zinc metal followed by about 3-mL of hydrochloric acid. Rapid bubbling should begin immediately as hydrogen gas is produced, and the bottom of the test tube will get quite hot. Place the test tube in the medium beaker. After 60 seconds have elapsed, light a wooden splint. Do not blow it out. Hold the burning splint to the mouth of test tube (where the hydrogen gas is being evolved) and record your observations.

Test 6: Combustion of magnesium

This test is an instructor demonstration. ***It must be performed in air only***. Hold a 1-inch piece of magnesium metal in a Bunsen burner flame with your crucible tongs until it ignites (in air). Record your observations, remembering not to look directly at the burning magnesium!

Name: _____ Chem 10, Section: _____

Lab Partner: _____ Experiment Date: _____

The Properties of Oxygen Gas

Part A: Generating and Collecting Oxygen Gas

1) Write the equation for the reaction used to generate oxygen gas.

   Word Equation: _____

   Formula Equation: _____

2) What is the *name* and *formula* of the catalyst used in this reaction?

   _____

   What is the purpose of this catalyst?

3) In addition to oxygen, what other substance is produced by this reaction? Where is this substance collected?

4) Two notable physical properties of oxygen are its low solubility in water and a density greater than air.

a. Which one of these properties allows the oxygen gas collected to be stored in the bottles *mouth up*? Explain.

b. Which one of these properties allows the oxygen gas to be collected via the *displacement of water*? Explain.

Part B: The Properties of Oxygen Gas

Test 1	Observations
Glowing splint in Bottle #1	
Glowing splint in air bottle	

Test 2	Observations
Burning candle in Bottle #2	Candle burned for _____ seconds.
Burning candle in air bottle	Candle burned for _____ seconds.
Test 3	Observations
Burning sulfur in Bottle #3	
Burning sulfur in air bottle	
Test 4	Observations
Glowing steel in Bottle #4	
Glowing steel in air bottle	
Test 5	Observations
Burning hydrogen in air	
Test 6	Observations
Burning magnesium in air	

Analysis of Combustion Results

1) Consider your results for the first four tests you performed. In which bottles, air-filled or oxygen-filled, did the combustion reactions occur more vigorously? Why?

2) Are the combustion reactions of oxygen exothermic or endothermic? Support your answer with one or more <u>specific</u> observations from the tests you performed.

3) Consider your Test 2 results. Although the candle burns for a longer period of time in one bottle, it eventually goes out in both the empty bottle and Bottle #2. Why does it go out?

4) When an element burns in oxygen gas, the product is called an oxide.

a. The wood in the splint consists mostly of carbon. The combustion of carbon produces carbon dioxide, CO_2. Write the equation for the combustion of wood (carbon).

Word Equation: _____

Balanced Formula Equation: _____

b. The combustion of sulfur produces sulfur dioxide, SO_2. Write the equation for the combustion of sulfur.

Word Equation: _____

Balanced Formula Equation: _____

c. Steel wool consists mostly of iron. The combustion of iron produces iron(III) oxide, Fe_2O_3. Write the equation for the combustion of steel wool (iron).

Word Equation: _____

Balanced Formula Equation: _____

d. The combustion of hydrogen produces water, H_2O. Write the equation for the combustion of hydrogen.

Word Equation: _____

Balanced Formula Equation: _____

e. The combustion of magnesium produces magnesium oxide, MgO. Write the equation for the combustion of magnesium.

Word Equation: _____

Balanced Formula Equation: _____

5) Do you expect the product formed during the combustion of magnesium in Test 6 (the ashy magnesium oxide) to weigh more than, less than, or the same as the original piece of magnesium? Explain.

Name: _____ Chem 10, Section: _____

Prelab Assignment: The Composition of Potassium Chlorate

1. In Part A of this lab, you will analyze a sample of potassium chlorate to determine the mass percent of oxygen present in it. To perform the analysis, you will decompose the potassium chlorate by heating it. Write the word equation and the balanced formula equation for this decomposition reaction.

   Word Equation: _____

   Formula Equation: _____

2. The potassium chlorate sample will be heated in a specialized "container".

a. What is this container called? _____

b. Will this container be covered or uncovered while heating? _____

3. You will have to heat your sample of potassium chlorate at least twice.

a. How long must the sample be heated the first time (total)? _____

b. How long must the sample be heated the second time? _____

4. A residue of potassium chloride will be left in the "container" after the heating is completed. Do you expect it weigh more than, less than or the same as the original potassium chlorate sample? Why?

5. In Part A you will be performing several mass measurements. What are <u>two</u> precautions you must observe when using the electronic balance?

6. In Part B of this lab, you will analyze the residue in left the "container" in order to experimentally verify its identity. To do this, you will need three test tubes. Potassium chlorate is added to tube #1, potassium chloride to tube #2, and the residue to tube #3. These solids are all dissolved in distilled water.

a. What <u>two</u> chemicals will then be added to each of these substances to test them?

b. What will you observe if you obtain a positive test for chloride ions?

The Composition of Potassium Chlorate

Objectives

The objectives of this laboratory are:

a) To experimentally determine the mass percent of oxygen in the compound potassium chlorate ($KClO_3$) via the thermal decomposition of a sample of potassium chlorate.
b) To qualitatively demonstrate that the residue resulting from the decomposition of potassium chlorate is potassium chloride.

Background

All compounds consist of elements chemically combined in fixed proportions – they obey the Law of Constant Composition. One way to express the proportion each of element in a compound is as a percentage by mass, or mass percent.

In Part A of this lab, a sample of potassium chlorate will be experimentally analyzed in order to determine the mass percent of elemental oxygen present in it. To do this, the potassium chlorate must be heated to temperatures greater 400 °C, causing it to thermally decompose into potassium chloride and free oxygen:

$$\underset{\text{Potassium Chlorate}}{2\ KClO_3\ (s)} \xrightarrow{\text{heat}} \underset{\text{Potassium Chloride}}{2\ KCl\ (s)} + \underset{\text{Oxygen}}{3\ O_2\ (g)}$$

Students will perform a quantitative analysis of the reactants and products of this reaction, measuring the initial mass of solid potassium chlorate used (before heating), and the mass of the solid potassium chloride product, or residue, remaining after heating. Applying the Law of Mass Conservation, the difference in these measured masses is the mass of oxygen released (from the original potassium chlorate sample). From this data, the *experimental* mass percent of oxygen in potassium chlorate will be determined:

$$\text{Mass Percent of Oxygen (experimental)} = \frac{\text{Mass of Oxygen Released}}{\text{Mass of Potassium Chlorate Used}} \times 100$$

Mass percentages of elements in compounds can also be theoretically calculated using molar masses, along with the known chemical formula of the compound. Thus, the *theoretical* mass percent of oxygen in potassium chlorate would be calculated using the expression:

$$\text{Mass Percent of Oxygen (theoretical)} = \frac{3\ (\text{Molar Mass of O})}{\text{Molar Mass of } KClO_3} \times 100$$

Students can therefore evaluate their accuracy in this experiment by comparing their experimental results to the true theoretical value, and by calculating their percent error.

In Part B of this lab, the residue left after heating will be qualitatively analyzed in order to demonstrate that it is chemically different from the initial potassium chlorate sample. Specifically, the residue will be tested for the presence of *chloride ions* by the addition of nitric acid and aqueous silver nitrate. A positive test is indicated by the formation of a *white precipitate*. The actual identity of the residue will then be conclusively verified by comparing this result to those obtained for identical tests on known samples of potassium chlorate and potassium chloride.

Procedure

Safety

Be especially careful when using the Bunsen burner and handling hot equipment. Remember that most items look exactly the same whether they are hot or cold. Heat the potassium chlorate sample slowly to avoid any splattering. Be aware that silver nitrate may stain the skin and *nitric acid may burn the skin*. If a spill of either chemical occurs, rinse under running water and report the accident to your instructor. Nitric acid spills may also be neutralized using the sodium bicarbonate solution by the sinks.

Materials and Equipment

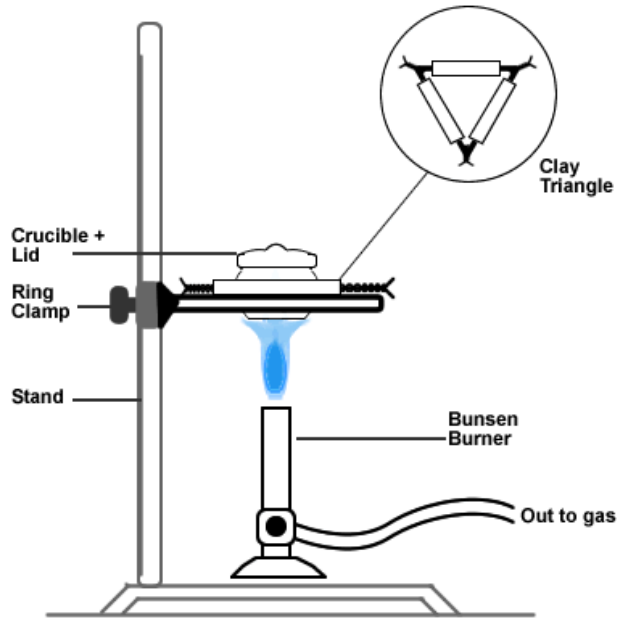
Solid potassium chlorate ($KClO_3$), solid potassium chloride (KCl), 6M nitric acid (HNO_3), 0.1M silver nitrate ($AgNO_3$), two crucibles with lids, stand and ring clamp, clay triangle, crucible tongs, Bunsen burner, three medium-sized test tubes, test tube rack, stirring rod, and an electronic balance.

Part A: Mass Percent of Oxygen in Potassium Chlorate

The following steps should be carried out for two separate samples of potassium chlorate.

1. Clean both crucibles and their lids (obtained from the stockroom) by thoroughly rinsing with distilled water then drying as completely as possible with a paper towel.
2. Weigh the first crucible and lid on an electronic balance and record this mass on your report form.
3. Add approximately 1 gram of potassium chlorate to the crucible. ***Do not do this over the balance!*** Then weigh and record the mass of the crucible, lid and potassium chlorate sample.
4. Fetch a stand and ring clamp from the back of the lab. As shown in the figure and photo on the following page, place your clay triangle on the ring, and then place the crucible containing the sample onto the triangle. Cover the crucible with the lid.
5. Using a Bunsen burner, heat the crucible and sample for a total of 12 minutes. Be sure that the crucible is covered, and that that the top of the flame is touching the bottom of the crucible.
   - For the first 6 minutes, the sample should be gently heated by adjusting the Bunsen burner flame to a low-moderate temperature. Note that heating the sample too intensely at this point could cause loss of the sample via splattering, and could crack the crucible.
   - For the last 6 minutes, the sample should be strongly heated by adjusting the Bunsen burner flame to a high temperature.
6. Allow the crucible to cool to room temperature. Then weigh and record the mass of the crucible, lid, plus the residue that remains. Note that the weight of your sample is expected to decrease by at least 30 % of its original mass (~ 0.3 g).

7. Now heat the sample a second time for an additional 6 minutes using a high temperature flame. Then, once again, allow it to cool to room temperature. Weigh the cooled crucible, lid and sample after this second heating and record the mass. If this mass is within 0.050 grams of your mass measurement after the first heating (see step 6), no further heating is necessary and you may begin Part B. ***Do not dispose of the residue as you will need it for Part B.***
8. If the sample from step 7 is not within 0.050 grams of the mass from step 6, heat again for a third time, cool and record the mass.
9. Repeat all steps for your second crucible and second sample of potassium chlorate.
10. Analysis: For each sample analyzed, obtain the mass of potassium chlorate before heating, the mass of KCl residue after heating, and the mass of oxygen released. Use the residue mass after the final heating for these calculations. Then using these values, determine the experimental mass percent of oxygen in potassium chlorate (see Theory Section for equation required).



Part B: Qualitative Analysis of Residue

1. Place three medium-sized test tubes in the test tube rack. The test tubes should be thoroughly cleaned and rinsed with distilled water. Label them tube #1, tube #2 and tube # 3. Continue to use only distilled water for the rest of Part B.
2. To tube #1 add a pea-sized amount of *potassium chlorate*. Then fill tube #1 half-way with distilled water and mix with a stirring rod until the solid dissolves. If the solid does not dissolve completely, decant the clear part of the solution into a clean test tube, labeling this #1 instead.
3. To tube #2 add a pea-sized amount of *potassium chloride*. Then fill tube #2 half-way with distilled water and mix until the solid dissolves. Be sure to rinse your stirring rod with distilled water first or you will contaminate your sample.
4. Add some distilled water to your crucible and *residue*. The residue should dissolve. Then pour the resulting solution into tube #3. Repeat the process until tube #3 is half-filled (like tubes 1 and 2).

5. To each of the tubes #1-3, add ***6 drops of nitric acid*** immediately followed by ***6 drops of silver nitrate*** solution. Then record your observations on the report form. Discard all test tube waste in the container provided.

6. Analysis: Based on your observations for these tests (and any others observations made), what evidence do you have that the residue in your crucible is really potassium chloride?

Name: _____ Chem 10, Section: _____

Lab Partner: _____ Experiment Date: _____

The Composition of Potassium Chlorate

Part A: Mass Percent of Oxygen in Potassium Chlorate

Experimental Data

	Sample 1	Sample 2
(a) Mass of crucible + lid		
(b) Mass of crucible, lid + $KClO_3$		
(c) Mass of crucible, lid + residue after 1st heating		
(d) Mass of crucible, lid + residue after 2nd heating		
(e) Mass of crucible, lid + residue after 3rd heating		

Data Analysis

- Use your data to determine the experimental mass percent of oxygen in $KClO_3$. Show your work clearly for each step in the table below.

	Sample 1	Sample 2
Mass of original $KClO_3$ sample		
Mass of KCl residue		
Mass of Oxygen released		
Mass Percent of Oxygen in $KClO_3$		
Average Mass Percent Oxygen		

- Using molar masses along with the known formula of potassium chlorate, calculate the theoretical mass percent of oxygen in $KClO_3$. Show your work clearly.

- Calculate the percent error between your average experimental value and theoretical value for the mass percent of oxygen in $KClO_3$. Show your work clearly.

Part B: Qualitative Examination of Residue

Observations and Analysis

Tube	Observations (after the addition of <u>both</u> nitric acid and silver nitrate)
#1: Potassium Chlorate	
#2: Potassium Chloride	
#3: Residue from Crucible	

- Explain how your observations in the table above verify that the residue in your crucible after heating is potassium chloride.

- Are there any <u>other observations</u> that you have made during this experiment (<u>not</u> those in the table above) that would suggest that the potassium chlorate was converted to a new substance upon heating?

Questions

1) Was your average experimental mass percent of oxygen in potassium chlorate higher or lower than the theoretical value (circle one)? Higher Lower

   Which of the following sources of error **could** be used to explain this discrepancy (circle one)?
   A. The potassium chlorate sample was not heated strongly or long enough.
   B. Some of the potassium chloride product splattered out of the crucible during the heating process.

   Explain your choice. Your response should include an analysis of the calculations you performed with your raw data to obtain your experimental % of oxygen.

2) Suppose the stockroom made a mistake and gave you a mixture of potassium chlorate and potassium chlorite. Upon analysis of this mixture, would you obtain a larger or smaller mass percent of oxygen than you would for an equal mass of pure sample of potassium chlorate (circle one)? Larger Smaller

   Explain your choice. Your response should include an analysis of the formulas of the compounds involved.

3) Show your calculations clearly. Suppose you are provided with a 36.55 g sample of potassium chlorate.

a. What mass of oxygen should theoretically be released upon heating?

b. What mass of potassium chloride residue should theoretically be left over after heating?

Name: _____ Chem 10, Section: _____

Prelab Assignment: Single and Double Displacement Reactions

1. In this lab you will perform a variety of single and double displacement reactions. What are <u>three</u> observable signs that a chemical reaction has occurred?

2. What is the general equation of a single displacement reaction?

3. For each of the following sets of reactants, write the balanced equation for the single displacement reaction that occurs. If you determine that a reaction will not occur, write "NR", and provide a brief explanation.

a. Aluminum metal + aqueous nickel(II) nitrate

b. Gold metal + hydrobromic acid

4. What is the general equation of a double displacement reaction?

5. For each of the following sets of reactants, write the balanced equation for the double displacement reaction that occurs. If you determine that a reaction will not occur, write "NR", and provide a brief explanation.

a. Aqueous zinc chloride + aqueous sodium chromate

b. Aqueous lithium hydroxide + phosphoric acid

6. The equipment required for this lab is fairly simple - just 8 small test tubes and 6 large test tubes.

a. Using the <u>small</u> test tubes you will mix two aqueous solutions together and observe whether or not a reaction occurs. What quantity of each solution will you use? How will you estimate this quantity?

b. What do the reactions studied in the <u>large</u> test tubes all have in common?

c. In the reactions involving both a solid and a solution as reactants, which do you place in the test tube first?

Single and Double Displacement Reactions

Objectives

The objectives of this lab are:
a) To perform and observe the results of a variety of single and double displacement reactions,
b) To become familiar with some of the observable signs of these reactions,
c) To identify the products formed in each of these reactions,
d) To write balanced chemical equations for each single and double displacement reaction studied.

Background

During a chemical reaction both the form and composition of matter are changed. Old substances are converted to new substances, which have unique physical and chemical properties of their own. Some of the observable signs that a chemical reaction has occurred include the following:

- A metallic deposit appears
- Bubbles appear
- A temperature change occurs
- A color change occurs
- A precipitate (cloudy, tiny particles) appears

Note that there are many other observable signs for chemical reactions, but these are the ones most likely to be encountered in this lab.

Single Displacement Reactions

All single displacement reactions have the general form: **A + BC → B + AC**

Here, A is an element and BC is usually an aqueous ionic compound or an acid (consisting of B^+ and C^- aqueous ions). A displaces B in BC, resulting in the formation of a new element B and a new ionic compound or acid, AC. If the new element B is a metal, it will appear as a metallic deposit. If it is a gas, it will appear as bubbles.

An *Activity Series* of elements is often used to determine if A will displace B in a single displacement reaction. An *Activity Series* is provided at the end of the Background section. As a rule, if A has a higher activity that B, a single displacement reaction will occur. However, if A has lower activity than B, a single displacement reaction will not occur.

Example 1:	magnesium metal + aqueous aluminum chloride Since Mg is more active than Al, a single displacement reaction will occur. The predicted products are aluminum metal and aqueous magnesium chloride Reaction Equation: $3\ Mg\ (s) + 2\ AlCl_3\ (aq) \rightarrow 2\ Al\ (s) + 3\ MgCl_2\ (aq)$

Double Displacement Reactions

All double displacement reactions have the general form: **AB + CD → AD + CB**

Reactions that can be classified as double displacements include precipitation reactions, neutralization reactions and gas forming reactions.

Precipitation Reactions

Here AB and CD are usually aqueous ionic compounds (or acids) consisting of aqueous ions (A^+ and B^-, C^+ and D^-). When a double displacement reaction occurs, the cations and anions switch partners, resulting in the formation of two new ionic compounds AD and CB, one of which is in the solid state. This solid product is an insoluble ionic compound called a precipitate. To determine whether a product ionic compound will be soluble or insoluble, consult the *Solubility Rules* provided at the end of the Background section. Note that if both of the predicted products are soluble, a precipitation reaction will not occur.

Example 2:	aqueous lead(II) nitrate + aqueous potassium chloride The predicted products are lead(II) chloride (insoluble) and potassium nitrate (soluble). Since one of the predicted products is insoluble, a precipitation reaction is will occur. Reaction Equation: $Pb(NO_3)_2\ (aq) + 2\ KCl\ (aq) \rightarrow 2\ KNO_3\ (aq) + \mathbf{PbCl_2\ (s)}$

Neutralization Reactions

Here AB is an acid (consisting of H^+ and X^- aqueous ions) and BC is a base (consisting of M^+ and OH^- ions). When a double displacement reaction occurs, the cations and anions switch partners, resulting in the formation of water and a new ionic compound (or salt), which is usually soluble. Neutralization reactions are exothermic, and are generally accompanied by a noticeable release of heat.

Example 3:	sulfuric acid + aqueous lithium hydroxide The predicted products are water and lithium sulfate. Reaction Equation: $H_2SO_4\ (aq) + 2\ LiOH\ (aq) \rightarrow Li_2SO_4\ (aq) + 2\ H_2O\ (l)$

Gas Forming Reactions

In these reactions one of the products (AD or CB) after the double displacement is in the gaseous state. One such example is hydrogen sulfide (H_2S). However, one of the products could also be carbonic acid (H_2CO_3) or sulfurous acid (H_2SO_3). Both carbonic acid and sulfurous acid are unstable and will decompose to form carbon dioxide and sulfur dioxide gases, respectively:

Carbonic acid

$H_2CO_3\ (aq) \rightarrow H_2O\ (l) + CO_2\ (g)$

Sulfurous Acid

$H_2SO_3\ (aq) \rightarrow H_2O\ (l) + SO_2\ (g)$

Example 4: nitric acid + aqueous sodium sulfite

The predicted products are sulfurous acid and sodium nitrate. However sulfurous acid decomposes to sulfur dioxide and water:

Reaction Equation: $2\ HNO_3\ (aq) + Na_2SO_3\ (aq) \rightarrow 2\ NaNO_3\ (aq) + \mathbf{H_2SO_3\ (aq)}$ **decomposes**

Final Equation: $2\ HNO_3\ (aq) + Na_2SO_3\ (aq) \rightarrow 2\ NaNO_3\ (aq) + \mathbf{H_2O\ (l) + SO_2\ (g)}$

Writing Equations for Reactions

- Write the correct formulas for each reactant and place a yield arrow (→) after the last reactant.
- Identify the reaction type – single or double displacement, using the guidelines outlined thus far.
- If you determine that a reaction will occur, write the correct formula(s) of the products after the arrow. If you determine that a reaction will not occur, simply write "no reaction" after the arrow.
- Balance the equation (to ensure mass conservation).
- Be sure to include the physical states of all reactants and products in your final equation.

Solubility Rules and Activity Series

SOLUBILITY RULES

1. Alkali metal compounds, acetates, nitrates, and ammonium compounds are all soluble.
2. Hydroxides of alkali metals and NH_4^{+1}, Ca^{+2}, Sr^{+2}, and Ba^{+2} are soluble. All others are **in**soluble.
3. All halides (chlorides etc.) are soluble except for those containing Ag^{+1}, Pb^{+2}, and Hg_2^{+2}.
4. Most sulfates are soluble, except for $BaSO_4$, $SrSO_4$, Ag_2SO_4, $PbSO_4$, and $CaSO_4$.
5. Most phosphates, carbonates, chromates and sulfides are **in**soluble (except those of the alkali metals and ammonium).
6. In addition, all acids are soluble!

ACTIVITY SERIES

highest activity	Li
	K
	Ca
	Na
	Mg
	Al
	Zn
	$Cr \rightarrow Cr^{+3}$
	$Fe \rightarrow Fe^{+2}$
	Cd
	$Ni \rightarrow Ni^{+2}$
	$Sn \rightarrow Sn^{+2}$
	$Pb \rightarrow Pb^{+2}$
	H_2
	$Cu \rightarrow Cu^{+2}$
	Ag
	$Hg \rightarrow Hg^{+2}$
lowest activity	$Au \rightarrow Au^{+3}$

Procedure

Safety

Be especially cautious when using the 6M HCl and 6M NaOH as they can burn your skin. Also be aware that skin discoloration will result from contact with $AgNO_3$. If you feel any tingling sensations or see any color changes on your skin, flush with water immediately for a minimum of 15 minutes. Inform your instructor of any chemical contact as soon as possible.

Materials and Equipment

Solids: Copper metal, zinc metal, magnesium metal, solid sodium bicarbonate
Solutions: 6M sodium hydroxide, 6M hydrochloric acid, 6M ammonium hydroxide, 5% acetic acid; all other solutions are 0.1M and include silver nitrate, barium chloride, sodium sulfate, potassium chloride, lead(II) nitrate, iron(III) chloride, sodium carbonate, cobalt(II) nitrate, sodium phosphate, zinc nitrate, copper(II) sulfate, sodium chloride, potassium nitrate, nickel(II) nitrate.
Equipment: 6 large test tubes, 8 small test tubes, plastic test tube rack (or large beaker)

Instructions for Performing Reactions

- For the reactions involving solid reactants (#2, 4, 5, 7, 11, 13), use the large test tubes. For reactions involving solutions only, use small test tubes. Always use clean test tubes that have been rinsed with ***distilled water***. The test tubes do not have to be dry.
- Use approximately 3-mL quantities of all solutions. A good estimate is to use two full dropper squirts of each chemical.
- For reactions involving metals, use just 1-2 pieces of each metal. Place the metal in the test tube first, and then add the solution. The metal should be completely immersed in the solution used.
- Perform the following reactions and record your observations for each on the report form. Note that some reactions take longer than others. Thus, if results are not obtained immediately, give the reaction some time. ***All waste is to be disposed of in the plastic container in the hood*!**

    1. Aqueous barium chloride + aqueous sodium sulfate
    2. Zinc metal + hydrochloric acid
    3. Aqueous sodium phosphate + aqueous copper(II) sulfate
    4. Copper metal + aqueous silver nitrate
    5. *Solid* sodium bicarbonate + acetic acid
    6. Aqueous nickel(II) nitrate + aqueous sodium hydroxide
    7. Copper metal + aqueous zinc nitrate
    8. Aqueous potassium chloride + aqueous silver nitrate
    9. Hydrochloric acid + aqueous sodium hydroxide
    10. Aqueous sodium carbonate + aqueous cobalt(II) nitrate
    11. Zinc metal + aqueous lead(II) nitrate
    12. Aqueous sodium chloride + aqueous potassium nitrate
    13. Magnesium metal + acetic acid
    14. Aqueous iron(III) chloride + aqueous ammonium hydroxide

- When finished, complete your lab report by writing the balanced equations for each reaction studied.

Name: _____ Chem 10, Section: _____

Lab Partner: _____ Experiment Date: _____

Single and Double Displacement Reactions

For each of the reactions performed,
-- predict the reaction type (single or double displacement)
-- record your observations
-- predict the names and states of the products formed
-- write the balanced "molecular" equation, including all physical states.

1. Aqueous barium chloride + aqueous sodium sulfate

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

2. Zinc metal + hydrochloric acid

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

3. Aqueous sodium phosphate + aqueous copper(II) sulfate

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

4. Copper metal + aqueous silver nitrate

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

5. Solid sodium bicarbonate + acetic acid

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

6. Aqueous nickel(II) nitrate + aqueous sodium hydroxide

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

7. Copper metal + aqueous zinc nitrate

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

8. Aqueous potassium chloride + aqueous silver nitrate

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

9. Hydrochloric acid + aqueous sodium hydroxide

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

10. Aqueous sodium carbonate + cobalt(II) nitrate

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

11. Zinc metal + aqueous lead(II) nitrate

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

12. Aqueous sodium chloride + aqueous potassium nitrate

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

13. Magnesium metal + acetic acid

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

14. Aqueous iron(III) chloride + aqueous ammonium hydroxide

Reaction Type:	
Observations:	Product Names & States (if none, why not?):
Balanced Equation:	

Questions

1. Consider Reactions 3 and 14 studied in this lab. Write the **balanced molecular equation** (identical to what you completed in the previous section), the **complete ionic equation** and the **net ionic equation** for these reactions. Include all physical states, and circle the spectator ions in the complete ionic equations.

Reaction 3: *Aqueous sodium phosphate + aqueous copper(II) sulfate*

Balanced Molecular Equation (from page 1):

Complete Ionic Equation:

Net Ionic Equation:

Reaction 14: *Aqueous iron(III) chloride + aqueous ammonium hydroxide*

Balanced Molecular Equation (from page 3):

Complete Ionic Equation:

Net Ionic Equation:

2. Predict the products for the following single and double displacement reactions, and write balanced molecular equations (including physical states) for each of them. If you predict that no reaction will occur, write "NR", followed by a brief explanation.

a. *Aluminum metal + aqueous silver acetate*

b. *Aqueous zinc nitrate + aqueous lithium chloride*

c. *Hydrobromic acid + solid magnesium sulfite*

d. *Aqueous rubidium hydroxide + perchloric acid*

e. *Tin metal + phosphoric acid*

f. *Aqueous lithium chromate + aqueous gold(III) iodide*

Name: _____ Chem 10, Section: _____

Prelab Assignment: Mole Ratios and Reaction Stoichiometry

1. Write balanced equations for the two reactions you will perform in this lab.

   Reaction A: _____

   Reaction B: _____

2. Your goal in this lab is to experimentally verify the mole-to-mole ratios between a certain reactant and a certain product in both reactions.

a. Identify the two substances in Reaction B. _____

b. What is this theoretical mole-to-mole ratio in Reaction B? _____

3. In Reaction A, you will react a pre-weighed sample of sodium bicarbonate with acid. In Reaction B, you will use sodium carbonate instead of the bicarbonate.

a. Name the "container" that you will perform both reactions in. _____

b. What is the purpose of the watch glass?

4. How will you know when enough acid has been added to the sodium bicarbonate (in Reaction A) or sodium carbonate (in Reaction B), and that the reactions are complete?

5. After mixing the reactants together, you will then use a Bunsen burner to heat the contents of the reaction "container". When should you stop heating?

6. Once heating is complete, you will then weigh the substance that remains in the reaction "container".

a. What is the name of this substance? _____

b. The mass of this substance could be described as your (circle one): experimental yield / theoretical yield / percent yield

Mole Ratios and Reaction Stoichiometry

Objectives and Background

The objectives of this laboratory are to experimentally determine the mole-to-mole ratios between the underlined reactants and products in the following two double displacement "gas forming" reactions:

(A) sodium bicarbonate + hydrochloric acid $\rightarrow$ sodium chloride + carbon dioxide + water

(B) sodium carbonate + hydrochloric acid $\rightarrow$ sodium chloride + carbon dioxide + water

The easiest way to obtain the mole-to-mole ratios would be to simply balance the chemical equations for these reactions. This would be considered a theoretical approach to the problem. The balanced equations for reactions A and B are:

(A) $\mathbf{1}\ NaHCO_3\ (s) + 1\ HCl\ (aq) \rightarrow \mathbf{1}\ NaCl\ (aq) + 1\ CO_2\ (g) + 1\ H_2O\ (l)$

(B) $\mathbf{1}\ Na_2CO_3\ (s) + 2\ HCl\ (aq) \rightarrow \mathbf{2}\ NaCl\ (aq) + 1\ CO_2\ (g) + 1\ H_2O\ (l)$

In reaction A, the balancing coefficients indicate that there is a 1:1 mole ratio between reactant $NaHCO_3$ and product NaCl. This means that for every 1 mole of sodium bicarbonate that reacts, 1 mole of sodium chloride should be produced. However in reaction B, the balancing coefficients indicate that there is a 1:2 mole ratio between reactant Na_2CO_3 and product NaCl. In this case, for every 1 mole of sodium carbonate that reacts, 2 moles of sodium chloride should be produced.

To determine these mole-to-mole ratios experimentally, a quantitative analysis of both reactions is required. Specifically, a pre-weighed mass of sodium bicarbonate/carbonate will be allowed to react with a slight excess of hydrochloric acid. The sodium chloride product will then be carefully retrieved, dried and weighed at the end of the reaction. This mass of collected sodium chloride is called an *experimental yield*. Both the mass of sodium bicarbonate/carbonate reactant used and the mass of sodium chloride product can be converted to mole quantities via their respective molar masses. Finally, dividing both the reactant and product moles by the lower of the two values should yield the simplest whole number mole-to-mole ratio of reactant and product.

While an experimental product yield is obtained by actually performing a reaction in lab, a *theoretical yield* is the maximum mass of product that could be obtained from a reaction provided that no errors occur. Theoretical product yields can only be determined by performing a series of stoichiometric calculations. Thus, using this method, theoretical yields of sodium chloride will be calculated for reactions A and B. Once obtained, the *percent yield* of sodium chloride can be determined for both reactions, where

$$\textit{Percent Yield} = \frac{\textit{Experimental Yield}}{\textit{Theoretical Yield}} \times 100$$

Good experimental practices in the lab (with minimum error) generally result in a high percent yield, where the experimental yield closely approaches the theoretical yield.

Procedure

Safety

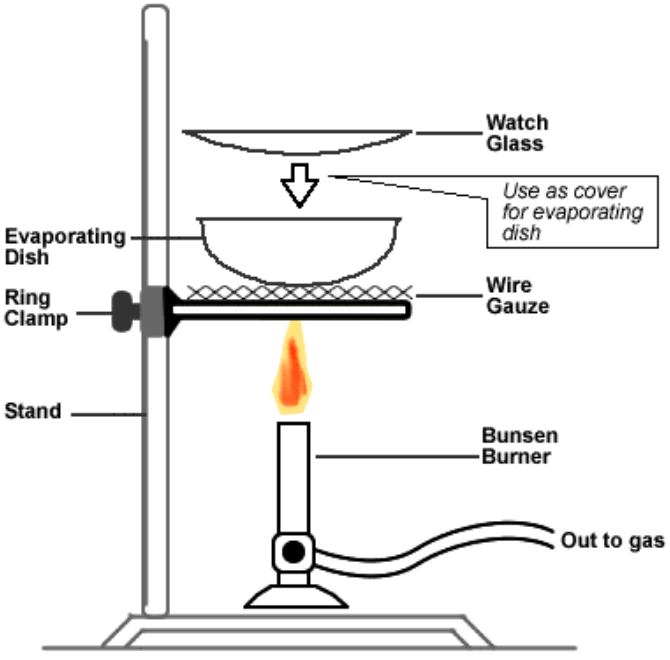
Be especially careful when handling the 6M HCl (*aq*), as it can cause chemical burns to the skin. If any acid spills on you, rinse immediately under running water for up to 15 minutes and report the accident to your instructor. Acid spills may also be neutralized using the sodium bicarbonate solution by the sinks. Also, be sure to exercise appropriate caution when using the Bunsen burner and handling hot equipment.

Materials and Equipment

Solid sodium bicarbonate ($NaHCO_3$), solid sodium carbonate (Na_2CO_3), 6M hydrochloric acid (HCl), electronic balance, evaporating dish, watch glass (to fit as a cover for the evaporating dish), stand and ring clamp, wire gauze, dropper pipette, small beaker, and Bunsen burner.

Experimental Procedure

1. Measure and record the mass of your clean dry evaporating dish + watch glass (assembled together with the watch glass acting as a cover on top of the evaporating dish).
2. Carefully add 0.3 – 0.4 g of solid sodium bicarbonate ($NaHCO_3$) to the evaporating dish. ***Do not do this over the balance!*** Then measure and record the mass of the evaporating dish + watch glass + $NaHCO_3$.
3. Obtain about a 5-mL quantity of hydrochloric acid (HCl) in your small beaker. Then using your dropper pipette, add the HCl ***drop by drop*** to the sodium bicarbonate in the evaporating dish. The reaction will be evident by the bubbling that takes place. Gently mix the reactants after every 4-5 drops of HCl. Continue adding HCl until the bubbling stops and all of the $NaHCO_3$ is dissolved – this indicates that the reaction is complete.
4. Assemble the stand, ring clamp and wire gauze apparatus for heating as shown in the figure below. Cover the evaporating dish with the watch glass and place it on the wire gauze.



5. ***Gently heat*** the solution in the covered evaporating dish with a Bunsen burner flame in order to remove the water generated in the reaction (as well as any excess HCl present). The flame should be adjusted to a lower temperature and wafted under the evaporating dish constantly. Continue heating until the contents are ***completely dry***. Note that the watch glass cover should also be dry!
6. After allowing the evaporating dish to cool to room temperature, measure and record the mass of the evaporating dish + watch glass + residue (NaCl).
7. Repeat steps 1 to 6 with a 0.3 – 0.4 g sample of sodium carbonate (Na_2CO_3).
8. The waste from this experiment may be disposed of in the sink.
9. Analysis: Experimental Mole-to-Mole Ratios – Convert the initial mass of sodium bicarbonate (or carbonate) reactant to moles (via its molar mass). In a similar manner, convert the final mass of collected sodium chloride product to moles (via its molar mass). Finally, obtain the simplest whole number mole-to-mole ratio by dividing both the reactant and product moles by the lower of the two values.
10. Analysis: Percent Yields – Calculate the theoretical yield of NaCl for both reactions A and B via standard mass-to-mass stoichiometry. Use your masses of sodium bicarbonate/carbonate reactants weighed out in lab as the starting point and the mole ratios from the balanced equations for these calculations. Then determine your percent yield for each reaction using the calculated theoretical yields along with your experimental yields of NaCl, obtained in lab. See the Theory Section for the equation, if necessary.

Name: _____ Chem 10, Section: _____

Lab Partner: _____ Experiment Date: _____

Mole Ratios and Reaction Stoichiometry

Reaction A: Sodium Bicarbonate and Hydrochloric Acid

Experimental Data

(a) Mass of evaporating dish + watch glass	
(b) Mass of evaporating dish + watch glass + sodium bicarbonate	
(c) Mass of sodium bicarbonate used	
(d) Mass of evaporating dish + watch glass + sodium chloride	
(e) Mass of sodium chloride collected (experimental yield)	

Data Analysis

1) Use your data to determine the experimental mole-to-mole ratio between sodium bicarbonate and sodium chloride. Show your work for each step.

- Convert the mass of sodium bicarbonate used to moles.

- Convert the mass of sodium chloride collected to moles.

- Divide both of your results from the preceding two steps by the lower mole value to determine the simplest mole-to-mole ratio between sodium bicarbonate and sodium chloride.

Simplest mole ratio *before* rounding

_____ moles $NaHCO_3$ **:** _____ moles NaCl

Simplest whole number mole ratio *after* rounding

_____ moles $NaHCO_3$ **:** _____ moles NaCl

2) Determine your percent yield of sodium chloride in reaction A. Show your work for each step.

- Write the balanced equation for reaction A – the reaction between sodium bicarbonate and hydrochloric acid.

- Using mass-to-mass stoichiometry, calculate the theoretical yield of NaCl for reaction A. Use your initial mass of sodium bicarbonate reactant as a starting point, along with the relevant mole ratio from the balanced equation to perform this calculation.

- Calculate your percent yield of sodium chloride product.

Reaction B: Sodium Carbonate and Hydrochloric Acid

Experimental Data

(a) Mass of evaporating dish + watch glass	
(b) Mass of evaporating dish + watch glass + sodium carbonate	
(c) Mass of sodium carbonate used	
(d) Mass of evaporating dish + watch glass + sodium chloride	
(e) Mass of sodium chloride collected (experimental yield)	

Data Analysis

1) Use your data to determine the experimental mole-to-mole ratio between sodium carbonate and sodium chloride. Show your work for each step.

- Convert the mass of sodium carbonate used to moles.

- Convert the mass of sodium chloride collected to moles.

- Divide both of your results from the preceding two steps by the lower mole value to determine the simplest mole-to-mole ratio between sodium carbonate and sodium chloride.

Simplest mole ratio *before* rounding

_____ moles Na_2CO_3 **:** _____ moles NaCl

Simplest whole number mole ratio *after* rounding

_____ moles Na_2CO_3 **:** _____ moles NaCl

2) Determine your percent yield of sodium chloride in reaction B. Show your work for each step.

- Write the balanced equation for reaction B – the reaction between sodium carbonate and hydrochloric acid.

- Using mass-to-mass stoichiometry, calculate the theoretical yield of NaCl for reaction B. Use your initial mass of sodium carbonate reactant as a starting point, along with the relevant mole ratio from the balanced equation to perform this calculation.

- Calculate your percent yield of sodium chloride product.

3) Is your percent yield here for reaction B greater than or less than 100%? Give one possible source of error that could explain the percent yield you obtained.

Name: _____ Chem 10, Section: _____

Prelab Assignment: Flame Tests of Metal Cations

1. In this lab, you will perform flame tests of several different metal cations. The characteristic colors observed are due to emitted electromagnetic radiation from the excited metal cations.

a. In this lab, how do the metal cations become "excited"?

b. Circle the correct responses to complete the following statement:

EM radiation is emitted when electrons make transitions from low / high to low / high energy levels.

2. In a flame test, the element Boron emits EM radiation that is predominantly green in color.

a. Use the Table in the Procedure to obtain the wavelength of this emitted radiation (in nm). Then convert this wavelength from nm to m.

b. Calculate the frequency of this emitted green radiation, in s^{-1}. Show your work.

c. Calculate the energy of this emitted green radiation, in J. Show your work.

3. Circle the type of EM radiation that has the property indicated:

a. Higher energy: UV or Visible light?

b. Longer wavelength: Orange or Blue light?

c. Higher frequency: IR or Yellow light?

Flame Tests of Metal Cations

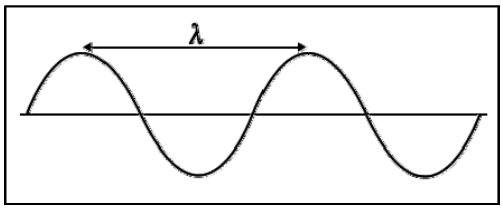
Objectives

The objectives of this lab are to:

a) Perform flame tests of metal cations in order to observe their characteristic colors,
b) Match the flame colors observed to an appropriate wavelength of visible light, and then perform calculations to determine the frequency and energy of the emitted photons,
c) Relate these results to the types of electronic transitions occurring in these elements,
d) Practice writing electron configurations for these (and other) elements.

Background

Electromagnetic radiation is composed of perpendicular waves oscillating in the electric and magnetic fields (through space or matter). These waves are characterized by their wavelength (λ) and frequency (υ). Wavelength is defined as the distance between successive crests (or troughs) on a wave, and is measured in meters. Frequency is defined as the number of waves that pass a given point every second, and is measured in 1/seconds, or Hertz (Hz).



All electromagnetic waves travel at the speed of light (c), or 2.998 x 10^8 m/s. The relationship between the wavelength, frequency and speed of an electromagnetic wave is given by the equation:

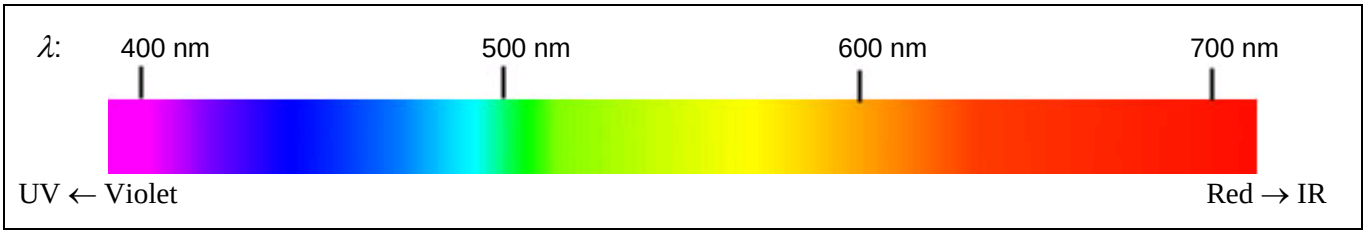
$$c = \lambda \times \upsilon$$

Electromagnetic radiation also occurs as discreet packets of energy (or quanta) called photons. The energy per photon (in Joules) is given by the equation:

$$E_{photon} = h \times \upsilon$$

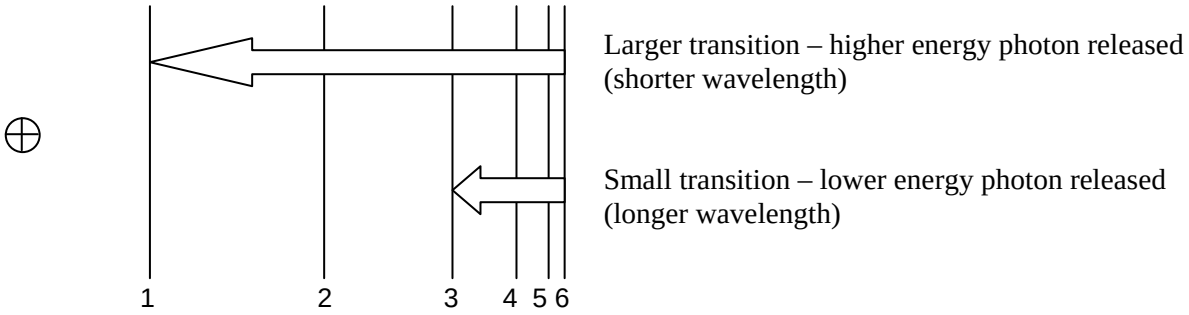
Here, h is Planck's constant, which has a value of 6.626 x 10^{-34} J•s.

Visible light is the most familiar example of electromagnetic radiation. Differences in the wavelengths of visible light are manifested as different colors, shown in the Color Spectrum below (colors can be seen in the PDF document on-line). Other examples of electromagnetic radiation include X-rays, ultraviolet light, infrared light, microwaves and radio waves.

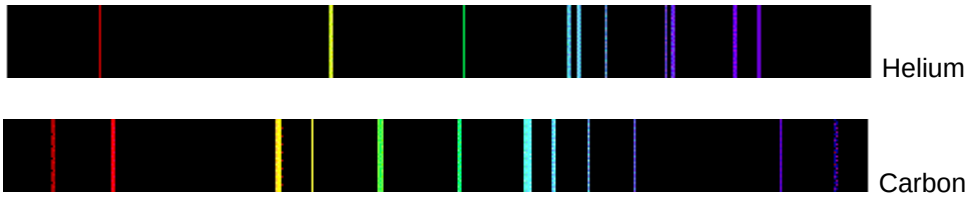


So, how does electromagnetic radiation relate to flame tests? Well, when an atom (or ion) absorbs energy, its electrons can make transitions from lower energy levels to higher energy levels. The energy absorbed could be in the form of heat (as in flame tests), or electrical energy, or electromagnetic radiation. However, when electrons subsequently return from higher energy levels to lower energy levels, energy is released predominantly in the form of *electromagnetic radiation*.

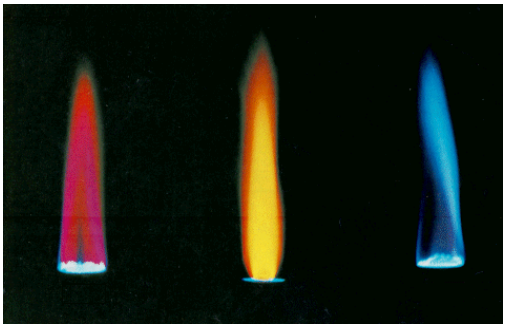
The spacing between energy levels in an atom determines the sizes of the transitions that occur, and thus the energy and wavelengths of the collection of photons emitted:



If emitted photons are in the visible region of the spectrum, they may be perceived as lines of different colors (note that photons outside the visible spectrum may also be emitted, but cannot be seen by eye). The result is called a *line emission spectrum*, and can serve as a 'fingerprint' of the element to which the atoms belong. For example, the line spectra shown below for the elements helium and carbon are clearly quite different.



Unfortunately, techniques more sophisticated than those used in this lab are required to obtain such line spectra. To the naked eye, when an element is vaporized in a flame (or an electrical discharge) the emission spectrum will appear to be just *one color*. For example, helium gas when excited by an electrical discharge emits light that appears an orange-peach color. This one color results from a *combination of all lines* of the emission spectrum, in proportion to their intensities. As many elements will still produce distinctive colors under such conditions, simple flame tests can be used to identify these elements. In fact, flame tests were used to identify elements long before the invention of modern techniques, such as emission spectroscopy.



Procedure

Safety

Exercise appropriate caution when using the Bunsen burner.

Materials and Equipment

Looped platinum or nichrome wires, wash bottle with distilled water, Bunsen burner, and the following solutions: LiCl (*aq*), NaCl (*aq*), KCl (*aq*), $CuCl_2$ (*aq*), $BaCl_2$ (*aq*), $CaCl_2$ (*aq*).

Experimental Procedure

This experiment will be performed as an instructor demonstration only.

Your instructor will dip a looped wire into one of the solutions supplied, and then hold it in the Bunsen burner flame. Students will record the dominant flame color observed. The table below contains a list of appropriate colors to choose from. Your instructor will then repeat this for the remaining five solutions, using a fresh looped wire each time.

Analysis: For each metal cation tested, obtain the wavelength of light corresponding to the observed flame color from the table below. Note that the wavelengths supplied here are in nanometers. Using these wavelengths, calculate the frequency and energy of the photons emitted during the flame tests. Finally, answer the questions and perform the exercises as indicated on your Report form.

Dominant Color	Approximate Wavelength (in nm)*
Red	701
Red-Orange	622
Orange	609
Orange-Yellow	597
Yellow	587
Yellow-Green	577
Green	535
Green-Blue	492
Blue	474
Blue-Violet	455
Violet	423

*****Wavelength values here are given for the mid-range of the color indicated.

Name: _____ Chem 10, Section: _____

Lab Partner: _____ Experiment Date: _____

Flame Tests of Metal Cations

Experimental Data and Observations

Solution	Dominant Flame Color	Wavelength (nm)
LiCl		
NaCl		
KCl		
$CuCl_2$		
$BaCl_2$		
$CaCl_2$		

Data Analysis

Using the wavelengths recorded above, calculate the corresponding frequencies and photon energies for the emitted radiation observed for each compound tested. Record the results of your calculations in the table below.

Solution	Wavelength (m)	Frequency (s^{-1})	Photon Energy (J)
LiCl			
NaCl			
KCl			
$CuCl_2$			
$BaCl_2$			
$CaCl_2$			

Show a set of sample calculations for LiCl only below. Clearly show any equations you have used.

- Wavelength (in m):

- Frequency (in s^{-1}):

- Photon Energy (in J):

Questions

1) Complete the following paragraph by circling the correct responses:

In this experiment, the metal cations in the solutions were initially in the (ground, excited) state. When placed in the flame, the metals then (absorbed, emitted) energy as (electricity, heat, EM radiation). When this occurred, electrons made transitions from (low, high) energy levels to (low, high) energy levels. The metals were then in the (ground, excited) state. The electrons in these metals then made transitions from (low, high) energy levels to (low, high) energy levels, resulting in the (absorption, emission) of energy as (electricity, heat, EM radiation).

2) What evidence is there that the colors observed in the flame tests are due to the metals, and not the nonmetals in the compounds tested?

3) Which metal cation was observed to emit radiation with the *shortest* wavelength? _____

Compared to the other metals studied, did the radiation emitted by this metal cation (identified above) have

- the highest or lowest frequency? _____
- the highest or lowest photon energy? _____

From this, would you conclude that the relationships between the following are direct or inverse?

- wavelength and frequency _____
- frequency and photon energy _____
- wavelength and photon energy _____

4) The energy, wavelength and frequency of an emitted photon are all related to the size of the electronic transition (high $\rightarrow$ low energy levels) occurring in the metal cation. Based on your observations, in which metal did the *smallest* electronic transitions occur? Briefly explain your response.

5) When heated in a flame, the element Indium emits electromagnetic radiation with a distinctive indigo blue color (the name indium is derived from the word indigo). The emitted photons that give rise to this color have energies of 4.405×10^{-19} J. Calculate the wavelength of this radiation in *nanometers*.

Exercise

In the Bohr Model of the atom, electrons occupy fixed orbits around the nucleus called energy levels. However in the Quantum Mechanical Model of the atom, electrons occupy orbitals. Orbitals are grouped by size and shape into shells and subshells (or, levels, and sublevels). Electron configurations and orbital diagrams are used to show the arrangement of electrons in shells (levels), subshells (sublevels) and orbitals for specific atoms.

Write complete electron configurations and abbreviated orbital diagrams for each of the elements given below. Circle the valence electrons in your complete electron configurations.

Chlorine

complete configuration: _____

abbreviated orbital diagram:

Tin

complete configuration: _____

abbreviated orbital diagram:

Selenium

complete configuration: _____

abbreviated orbital diagram:

Cobalt

complete configuration: _____

abbreviated orbital diagram:

Boron

complete configuration: _____

abbreviated orbital diagram:

Bismuth

complete configuration: _____

abbreviated orbital diagram:

Magnesium

complete configuration: _____

abbreviated orbital diagram:

Name: _____ Chem 10, Section: _____

Prelab Assignment: Lewis Structures and Molecular Shapes

1. Lewis Structures are used to represent covalently bonded molecules and polyatomic ions. Draw the Lewis Structure of the OF_2 molecule. A copy of the "Rules for Drawing Lewis Structures" may be found on page 4 of the Procedure Handout.

2. Consider the molecule carbon monoxide, shown below.

a. Are the electrons in the triple bond pulled closer to the C or the O, or, are they equally shared? Explain your response, referring to the concept of electronegativity.

b. Is this triple bond polar or non-polar? _____

3. In this lab you will draw Lewis Structures for a number of molecules, and then you will build each molecule with the Model Kit provided. The kits contain three items: colored balls, short sticks and long flexible sticks.

a. The colored balls correspond to different atoms. How will you know which color to use for specific atoms?

b. What type of bond is a short sticks used for?

c. Suppose you want to construct a triple bond. What type of stick should you use, and how many?

4. One of the goals of this lab is to become familiar with different shapes of simple molecules.

a. What is the name of the theory used to predict molecular shapes?

b. Suppose a molecule consists of a central atom (X) bonded to 2 outer atoms (both Y). There is one lone pair on the central atom.

What is the name of the shape of this molecule? _____

What are all the bond angles in this molecule? _____

c. Suppose a molecule consists of a central atom (X) bonded to 4 outer atoms (all Y). There are no lone pairs on the central atom.

What is the name of the shape of this molecule? _____

What are all the bond angles in this molecule? _____

Lewis Structures and Molecular Shapes

Objectives

The objectives of this laboratory are:
a) To practice drawing Lewis Structures for various covalently bonded molecules and polyatomic ions.
b) To use model kits to construct these molecules/ions in order to explore their structure and shapes.
c) To practice predicting molecular shapes (using VSEPR theory) and molecular polarity.

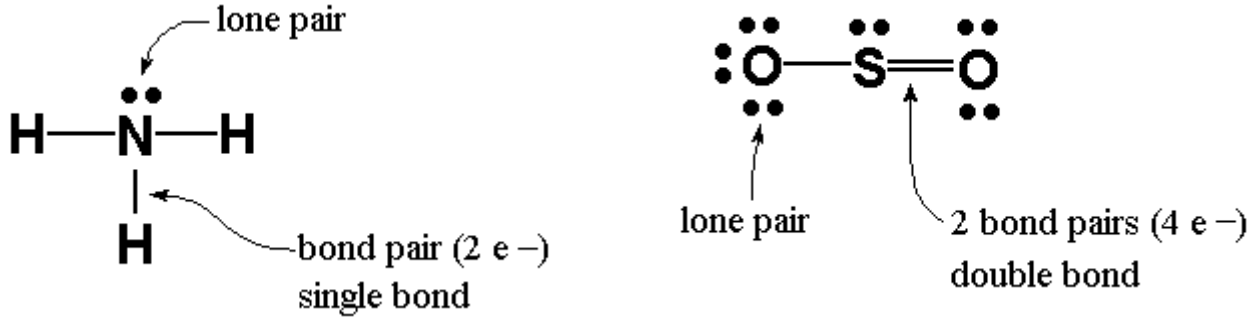
Background

Non-metal atoms bond covalently, resulting in the formation of either neutral molecules or polyatomic ions. A covalent bond is formed when non-metal atoms *share* their valence electrons, which they do in order to achieve filled valence orbitals like their nearest noble gas neighbor. This means that most bonded non-metal atoms will acquire a total of eight valence electrons via the sharing process – often referred to as the octet rule. A notable exception is hydrogen, which only needs to acquire two electrons to be like its nearest noble gas neighbor, helium.

Lewis Structures

A Lewis Structure is a representation of covalent molecules (or polyatomic ions) where all the valence electrons are shown distributed about the bonded atoms as either shared electron pairs (bond pairs) or unshared electron pairs (lone pairs). A shared pair of electrons is represented as a short line (a single bond). Sometimes atoms can share two pairs of electrons, represented by two short lines (a double bond). Atoms can even share three pairs of electrons, represented by three short lines (a triple bond). Pairs of dots are used to represent lone pair electrons.

Examples:



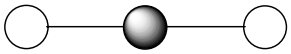
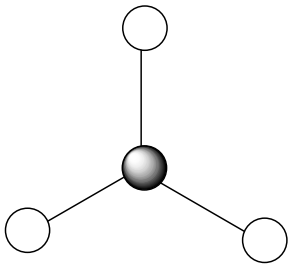
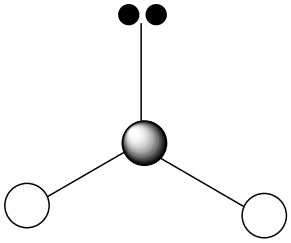
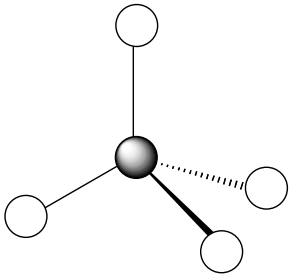
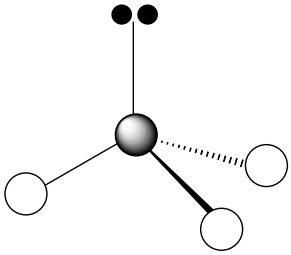
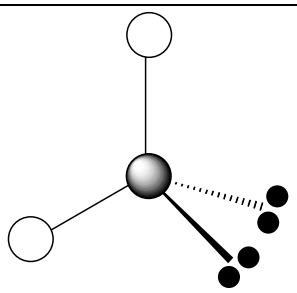
The rules for drawing Lewis structures can be found in the Procedure Section of this handout.

Molecular Shapes

The shape of a molecule depends on the distribution of atoms in space about the central atom, and their bond angles. Bond pair electrons and lone pair electrons repel one another, thus they will be arranged around a central atom as far apart as possible in order to minimize repulsions. This is known as:

Valence **S**hell **E**lectron **P**air **R**epulsion theory, or **VSEPR** theory.

The following VSEPR table supplies the names, sketches and descriptions of the most common types of molecular shapes that you will encounter. Note that several other molecular geometries do exist, however, they are beyond the scope of this course.

Shape Name	Sketch	Description
Linear		Only two outer atoms surround the central atom. There are no lone pairs on the central atom. Outer atoms are arranged opposite to each other. The bond angles are exactly 180°.
Trigonal Planar		Three outer atoms surround the central atom. There are no lone pairs on the central atom. The central and outer atoms all lie in the same plane (molecule is flat). Bond angles are exactly 120°.
Bent		Two outer atoms and one lone pair surround the central atom. Bond angles are slightly less than 120°.
Tetrahedral		Four outer atoms surround the central atom. There are no lone pairs on the central atom. The four outer atoms are evenly arranged in 3D around the central atom as if at the corners of a regular tetrahedron. The bond angles are exactly 109.5°.
Trigonal Pyramidial		Three outer atoms and one lone pair surround the central atom. Here the central atom is located slightly above the three outer atoms, like a tripod. The bond angles are slightly less than 109.5°.
Bent		Two outer atoms and two lone pairs surround the central atom. Bond angles are slightly less than 109.5°.

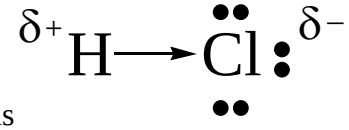
Electronegativity and Bond Polarity

Some atoms in molecules have the ability to pull shared electrons closer to themselves than other atoms, an ability referred to as electronegativity. Electronegativity (χ) increases going across a period and decreases going down a group.

If two bonded atoms have different electronegativities, then the bond pair electrons will be *unequally shared*. The atom with the greater electronegativity will pull the bond electrons closer towards itself, causing it to obtain a very slight negative charge (δ-). The atom with the lower electronegativity will have bond electrons pulled further away from it, causing it to obtain a very slight positive charge (δ+). The result is a *polar covalent bond*. However, if two bonded atoms have the same electronegativity, then the bond pair electrons will be *equally shared*. The result is a *non-polar covalent bond*.

Example:

is better represented as



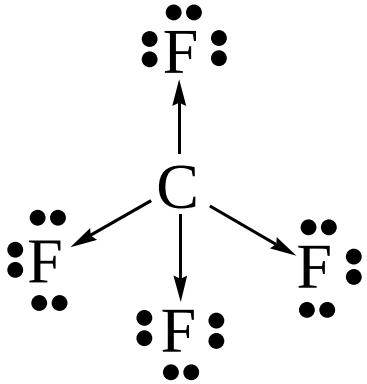
because the bond electrons are pulled closer to Cl due to its greater electronegativity. Thus, HCl contains a polar covalent bond.

Molecular Polarity

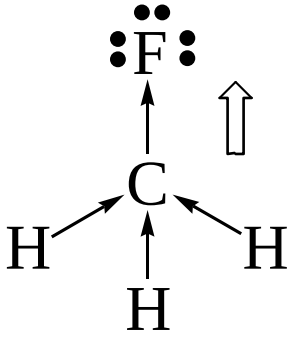
Molecular polarity results when the *entire molecule* (not just a bond in the molecule) ends up with an unequal distribution of electrons. In general, a molecule will be *polar* if it contains polar bonds that are distributed in a *non-symmetrical* arrangement around the central atom. A polar molecule is said to have a net dipole moment. A non-symmetrical arrangement typically results when there are lone pairs on the central atom, or, when different outer atoms surround the central atom.

Not surprisingly, a molecule will be *non-polar* if it contains all non-polar bonds. It will also be non-polar if it contains polar bonds distributed in a *symmetrical* arrangement around the central atom. The symmetry causes the individual bond polarities to cancel out, resulting in a net non-polar molecule. A symmetrical arrangement typically results when there are no lone pairs on the central atom, and if all the outer atoms are identical.

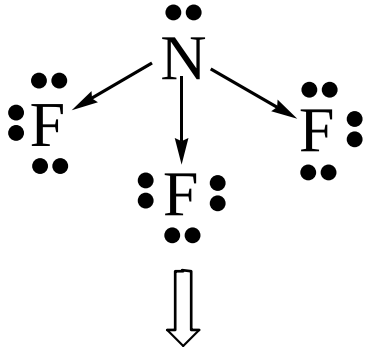
Examples:



Non-Polar



Polar



Polar

Procedure

Drawing Lewis Structures

1. Draw Lewis structures for each of the molecules and polyatomic ions given on your Report Form. Clearly show all bond pair electrons as lines and lone pair electrons as pairs of dots.

Rules for Drawing Lewis Structures

1. Total the number of valence electrons that each atom contributes to the molecule/polyatomic ion.
   - The quickest way is to find the group number for each atom.
   - For polyatomic anions, you must add electrons (equal to the negative charge) to the total number of valence electrons. For polyatomic cations, you must subtract electrons (equal to the positive charge) from the total number of valence electrons.
2. Draw a stick structure for the molecule.
   - Most molecules/polyatomic ions consist of one central atom bonded to 2, 3 or 4 other atoms.
   - The least electronegative atom is the central atom. Hydrogen is the only exception to this, as it forms only 1 bond. The central atom will usually need to form multiple bonds.
   - The other atoms are arranged around the central atom, and are attached to it via single bonds.
3. The octet rule must be obeyed for all elements except hydrogen (follows a "duet" rule).
   - Each bond in the stick structure contains two electrons, which must be subtracted from the total number of valence electrons.
   - Starting with the outside atoms, place the remaining electrons around each atom until it has a total of 8 electrons (except hydrogen – it only requires 2 electrons).
   - If there are not enough electrons available to obey the octet rule using single bonds, this indicates that double or triple bonds between two atoms are required in your structure. If short by two electrons, try a double bond, and if short by four electrons, try a triple bond or two double bonds.

Constructing Models, Determining Molecular Shapes and Molecular Polarity

1. Use your molecular model kit to construct a three-dimensional model of each of these molecules and polyatomic ions. Sketch a reasonably detailed picture of this model on your Report Form.

Rules for Constructing Molecules with the Model Kit

1. Each colored ball in your kit corresponds to different atom or different group of atoms. Refer to the inside of the container lid for the correct color key.
2. Use the *short sticks* for *single bonds*.
3. Use *two long flexible sticks* for a *double bond* and *three long flexible sticks* for a *triple bond*.

2. Compare each model constructed to the molecular shapes described in the Theory section. Then identify and record its correct shape name and its bond angles. If the molecule has no definite central atom, then only record the bond angles.

3. Evaluate the polarity of the bonds in each molecule as well as its overall symmetry in order to determine whether it is polar or non-polar.

Name: _____ Chem 10, Section: _____

Lab Partner: _____ Exercise Date: _____

Lewis Structures and Molecular Shapes

1. **H_2S**

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

2. **$COCl_2$**

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

3. **SiI_4** – use a black ball for silicon when constructing this model

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

4. **HCN**

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

5. **NO_2^{-1}**

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

6. **H_3O^{+1}** – use a yellow ball for oxygen (not red) when constructing this model

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

7. **$CHBr_3$**

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

8. **OF_2**

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

9. **CO_3^{-2}**

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

10. **CS_2**

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

11. **PF_3** – use a blue ball for phosphorus when constructing this model

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

12. **NOCl**

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

13. _____

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

14. _____

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

15. _____

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

16. _____

Total # of Valence Electrons:	3-D Model Sketch	Bond Angles
Lewis Structure (show all resonance structures if applicable)		
	Molecular Shape:	
Any polar bonds in the molecule? Yes No	Molecular Polarity: Polar Non-Polar	

Name: _____ Chem 10, Section: _____

Prelab Assignment: Experimental Determination of the Gas Constant

1. What is the name of the gas that will be collected and studied in this lab? _____

   Write the balanced equation for the reaction used to generate this gas.

   _____

2. You will perform several measurements on your collected gas sample in order to experimentally determine the value of the Gas Constant (R).

a. What is the theoretical value of R, and what are its units? _____

b. The Gas Constant is found in the Ideal Gas Law. Write the equation for this law.

3. The magnesium ribbon used in this reaction must be carefully handled.

a. What must you avoid doing with the magnesium ribbon at the lab bench (hint, see Procedure #2)?

b. What mass of the magnesium ribbon should be used? _____

4. What is the name of the specialized "tube" that your gas is collected in? _____

   This tube not only collects the gas, it also allows you to directly measure the gas (circle one):

   pressure / temperature / volume

5. Consider some of the other equipment you will use in this lab.

a. What device will you use to measure atmospheric pressure? _____

b. What TWO temperatures will you measure with your thermometer?

6. Part of the procedure for this experiment involves ensuring that the total pressure of gases collected inside the specialized tube is equal to atmospheric pressure. How is this achieved (hint, see Procedure #8)?

Experimental Determination of the Gas Constant

Objectives

The objectives of this lab are to experimentally determine the value of the Gas Constant, *R*, and to practice using the Gas Laws to solve a variety of problems.

Background

A gas is the state of matter that is characterized by having neither a fixed shape nor a fixed volume. Gases exert pressure, are compressible, have low densities and diffuse rapidly when mixed with other gases. On a microscopic level, the molecules (or atoms) in a gas are separated by large distances and are in constant, random motion.

Four measurable properties can be used to describe a gas quantitatively: pressure (*P*), volume (*V*), temperature (*T*) and mole quantity (*n*). The relationships among these properties are summarized by the Gas Laws, as shown in the table below.

Charles's Law:	$V \propto T$ [*P* and *n* are held constant] As gas temperature increases, gas volume increases.	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Boyle's Law:	$V \propto 1/P$ [*T* and *n* are held constant] As gas pressure increases, gas volume decreases.	$P_1V_1 = P_2V_2$
Avogadro's Law:	$V \propto n$ [*P* and *T* are held constant] As the number of moles of gas increase, gas volume increases.	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$
Combined Law:	$V \propto T/P$ [*n* is held constant] Obtained by combining Boyle's Law and Charles's Law.	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

A closer look at the Combined Law reveals that the volume of a gas depends on both the pressure and temperature. Thus, if the volumes of two gases are to be compared, they must be under the same *P* and *T*. A commonly used set of *P* and *T* reference conditions is known as Standard Temperature and Pressure, or STP. Standard temperature is defined as exactly 0 °C (273 K) and standard pressure is defined as exactly1 atm (760 mm Hg).

The Ideal Gas Law is obtained by combining Boyle's Law, Charles's Law and Avogadro's Law together:

$$PV = nRT$$

Here, *P* represents as the gas pressure (in atmospheres); *V* is the gas volume (in Liters); *n* is the number of moles of gas in the sample; *T* is the gas temperature (in Kelvins). *R* is a proportionality constant called the Gas Constant, and has a theoretical value of 0.08206 L•atm/K•mol. Note that the units of *R* will allow the units of *P*, *V*, *n* and *T* in the Ideal Gas Law to cancel correctly.

In this lab, students will measure various properties of a sample of hydrogen gas in order to experimentally determine the value of the Gas Constant, *R*. The single displacement reaction between magnesium metal and hydrochloric acid will be used to generate the hydrogen gas:

$$Mg\,(s) + 2\,HCl\,(aq) \rightarrow MgCl_2\,(aq) + H_2\,(g)$$

The hydrogen gas will be collected in a eudiometer, a tube closed at one end and marked in milliliter volume units. The gas will be collected in the closed end of the tube over a water bath via the technique of water displacement (see figures on page 4).

Students will then obtain the following values for the collected sample of hydrogen gas: (1) Volume, (2) Temperature, (3) Moles, and (4) Pressure. The hydrogen volume will be directly measured from the eudiometer scale. The hydrogen temperature will also be directly measured using a thermometer. However, the mole quantity and pressure of the hydrogen gas must be determined indirectly. The mole quantity of the collected hydrogen can be easily calculated from the measured mass of the magnesium reactant using stoichiometry. But the hydrogen pressure is a little more difficult to obtain. Since hydrogen is collected over a water bath, a small amount of water vapor is mixed with the hydrogen in the eudiometer. The combined pressure of the H_2 and H_2O gases will be equal (after adjustments) to the external atmospheric pressure:

$$P_{atm} = P_{hydrogen} + P_{water\ vapor}$$

P_{atm} (atmospheric pressure) will be measured using a barometer. $P_{water\ vapor}$ (the partial pressure of water vapor) depends on the temperature of the water bath, and can be obtained from the table supplied below. By substituting these values in the above equation, the pressure of hydrogen ($P_{hydrogen}$) will be determined.

Temperature (°C)	$P_{\text{water vapor}}$ (mm Hg)
16	13.5
17	14.5
18	15.5
19	16.5
20	17.5
21	18.7
22	19.3
23	21.1
24	22.4
25	23.8
26	25.2
27	26.7
28	28.3
29	30.0

Finally, to determine the value of the Gas Constant (*R*), the quantities *V*, *T*, *n* and *P* obtained for the hydrogen gas must simply be substituted into the Ideal Gas Equation. Students can then evaluate their accuracy in this experiment by comparing their experimental result to the true theoretical value of *R*, and by calculating their percent error.

Procedure

Safety

Concentrated HCl is dangerous! Handle it with extreme care as demonstrated by your instructor. If any spills occur, inform your instructor immediately. Wash under running water (sink or shower) and use the neutralizing sodium bicarbonate solution supplied at the sinks if necessary. Also note that hydrogen gas is flammable, so be sure to have no open flames nearby when you perform this experiment.

Materials and Equipment

4.0-cm ribbon of magnesium, length of copper wire (reusable), 6M HCl (*aq*), 50-mL eudiometer*, eudiometer stopper with hole(s)*, burette stand, large beaker, thermometer, small funnel, small graduated cylinder, barometer, large tub of water, electronic balance and sandpaper.

Experimental Procedure

Magnesium Ribbon

1. Obtain a 4.0-cm ribbon of magnesium (Mg), a piece of sandpaper, and a length of copper wire.
2. Carefully sand the outside of the Mg ribbon to remove any oxide coating. **Do not sand on the bench top!** Place the Mg ribbon on a paper towel while sanding. Weigh the cleaned Mg ribbon and record this mass on your report form. Note that this mass should be less than 0.040 grams. If it is heavier, your Mg ribbon will have to be "trimmed" by your instructor.
3. Wrap the Mg around the end of the copper wire. Do this in a tight ball with only a small gap between layers. Then wrap the copper wire to form a cage around the Mg ball. The cage must be tight enough to keep the Mg inside, but loose enough to allow water to easily flow around the wire. Roughly 3-cm of copper wire should be left over as a "handle" (see Figure 1).

Eudiometer Set Up and Reaction

4. Obtain a eudiometer tube and stopper (with holes) from the stockroom. Use the burette clamp to hold it in place, open end up.
5. Add ~10-mL of 6M HCl (*aq*) to the eudiometer tube using a small funnel. Then add tap water to the eudiometer carefully until it is filled to the brim (see Figure 1).
6. Hang the Mg ball inside the open end of the eudiometer, ~2-cm down from the top. Then insert the stopper into this end, and, while holding it in place, quickly invert the entire tube into your largest beaker ¾ filled with water. Clamp the tube in the water in the upside down position (see Figure 2).
7. The reaction will occur as soon as the acid diffuses down the tube and reaches the Mg ribbon. As hydrogen gas is generated it will fill the eudiometer by forcing the water out of the tube and into the beaker via water displacement (see Figure 2). Allow the reaction to proceed until no Mg is left and no further gas is formed. This should take 3-5 minutes.

Pressure Equalization

8. To ensure that the pressure of hydrogen (and water vapor) in the eudiometer is equal to atmospheric pressure, the level of the water inside the tube must be the same as the level of water outside the tube. To achieve this, transfer both the tube and the beaker of water into the large bucket of water in the sink. Then raise or lower the tube until the internal and external water levels are equal.

Measurements

9. After equalizing the water levels, record the following measurements:
   - The volume of hydrogen gas collected (read directly from the eudiometer scale), in mL
   - The temperature of the hydrogen gas collected, in °C. This can be measured by first removing the stopper then placing the thermometer directly in the eudiometer (keep the tube inverted so the gas does not readily escape). It is also acceptable to assume that the temperature of the hydrogen gas is the same as the temperature of the water bath, especially if you wait a while before making your measurements.
   - The atmospheric pressure (use the lab barometer), in mm Hg
   - The temperature of the water in the plastic tub (use the thermometer), in °C
   - The vapor pressure of water at the above temperature (obtain from Table on page 2), in mm Hg
10. When finished, repeat this entire procedure a second time with a fresh piece of magnesium ribbon.

Figure 1
Before Inverting

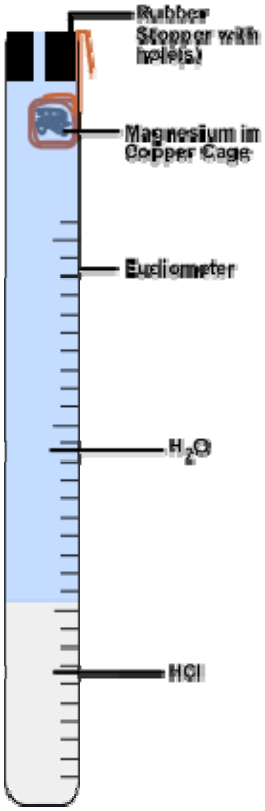
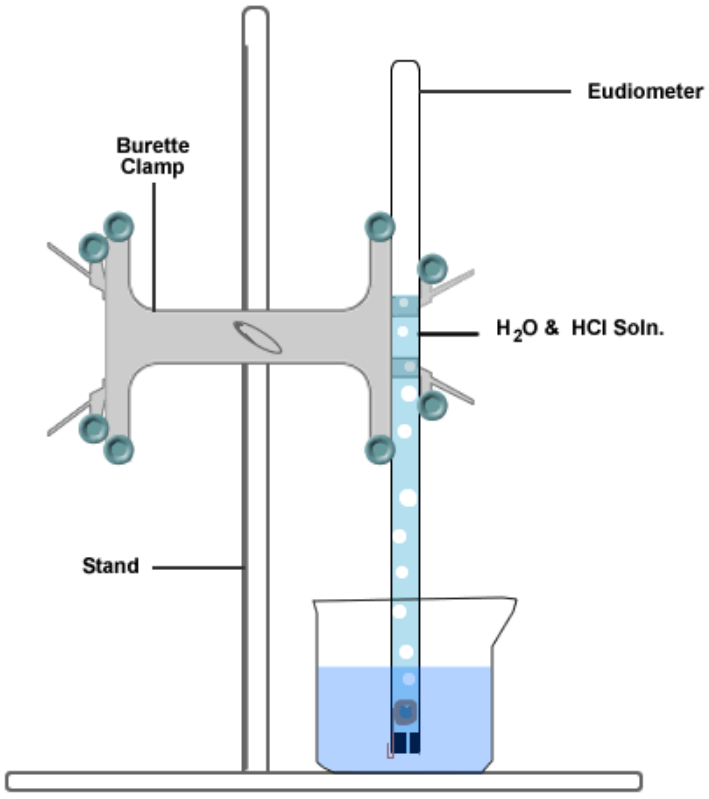


Figure 2
After Inverting



Name: _____ Chem 10, Section: _____

Lab Partner: _____ Experiment Date: _____

Experimental Determination of the Gas Constant

Experimental Data

	Trial 1	Trial 2
(a) Mass of Magnesium metal used		
(b) Volume of H_2 gas collected		
(c) Temperature of H_2 gas collected		
(d) Atmospheric Pressure		
(e) Temperature of Water in bath (bucket)		
(f) Vapor Pressure of Water at above temperature		

Data Analysis

- Using your experimental data, determine the value of *R*, the gas constant. Show all your conversions and calculations for each step clearly in the table below. Pay attention to units and significant figures.

Trial 1	Trial 2
Volume of H_2 gas (in L)	
Temperature of H_2 gas (in K)	
Moles of H_2 gas	
Pressure of H_2 gas (in atm)	

Experimental value of *R* (include units)	

- Average value of *R* (include units): _____
- Percent Error between your average value and the theoretical value of *R* (show work):

Questions

1) The hydrogen generated in this lab was a product of the reaction between magnesium and hydrochloric acid. Which of these reactants was the limiting reactant? Provide experimental evidence to support your choice.

2) Suppose when you inverted the eudiometer, a bubble of air became trapped inside it. Would this make your experimental value of *R* larger, smaller, or have no effect? Briefly explain your response.

3) In Santa Monica, a sample of dry hydrogen gas inflates a balloon to 43.0 mL at 761 torr (sea-level). If the temperature remains unchanged, what is the balloon's volume (in mL) in Denver, where the pressure is 12.2 psi (5000 ft elevation)? Assume that no gas has been added or removed.

4) Another balloon is inflated to a volume of 1.250 L with dry hydrogen gas, at 28.0 °C. The balloon is then cooled, and its volume drops to 964 mL. If the pressure is unchanged, what is the hydrogen gas temperature (in °C) in the cooled balloon? Assume that no gas has been added or removed.

5) Yet another balloon is inflated to a volume of 434 cm^3 using 0.141 moles of dry hydrogen gas. An additional 0.129 grams of hydrogen is then injected into the balloon at constant pressure and temperature. Calculate the new volume of the balloon (in cm^3).

6) Hydrogen gas can be generated from the reaction between *aluminum* metal and hydrochloric acid:

$$2\ Al\ (s)\ +\ 6\ HCl\ (aq)\ \rightarrow\ 2\ AlCl_3\ (aq)\ +\ 3\ H_2\ (g)$$

a. Suppose that 3.00 grams of Al are mixed with excess acid. If the hydrogen gas produced is directly collected into a 850. mL glass flask at 24.0 °C, what is the pressure inside the flask (in atm)?

b. This hydrogen gas is then completely transferred from the flask to a balloon. To what volume (in L) will the balloon inflate under STP conditions?

c. Suppose the balloon is released and rises up to an altitude where the temperature is 11.2 °C and the pressure is 438 mm Hg. What is the new volume of the balloon (in L)?

Name: _____ Chem 10, Section: _____

Prelab Assignment: Titration of Vinegar

1. In this lab, you will perform a titration using sodium hydroxide and acetic acid (in vinegar). Write the balanced neutralization reaction that occurs between sodium hydroxide and acetic acid.

_____

2. Specialized equipment is needed to perform a titration.

a. Consider the sodium hydroxide reactant.

-- Name the specialized device the sodium hydroxide is placed in. _____

-- Is the concentration of the sodium hydroxide known or unknown? _____

-- Is sodium hydroxide the analyte or the titrant? _____

b. Consider the acetic acid reactant.

-- What type of flask is the acetic acid placed in? _____

-- What volume of acetic acid is used? _____

-- What specialized device is used to obtain this precise volume? _____

-- Is the acetic acid the analyte or the titrant? _____

3. You will add sodium hydroxide to the acetic acid until all the acetic acid is consumed. This is a special point in the titration called the _____ point.

4. An indicator solution is used to indicate when all the acetic acid has been consumed and that the reaction in complete.

a. What is the name of the indicator solution? _____

b. Is this indicator mixed with sodium hydroxide or acetic acid? _____

c. How exactly does the indicator let you know when the reaction is complete?

Titration of Vinegar

Objectives

The objectives of this laboratory are to determine the molarity and percent by mass of acetic acid in vinegar.

Background

Vinegar is essentially a solution of acetic acid ($HC_2H_3O_2$) in water. The concentration of acetic acid in vinegar may be expressed as a molarity (in mol/L) or as a mass percent, where:

Molarity	=	Moles of Acetic Acid / Volume of Vinegar (in L)
Mass %	=	[Mass of Acetic Acid / Mass of Vinegar] x 100

In this experiment, a technique known as a titration will be used to determine the concentration of acetic acid in vinegar. A titration involves performing a controlled reaction between a solution of known concentration (the titrant) and a solution of unknown concentration (the analyte). Here, the titrant is an aqueous solution of ~ 0.1 M sodium hydroxide (NaOH) and the analyte is vinegar. When mixed, a neutralization reaction occurs between sodium hydroxide and the acetic acid in vinegar:

$$NaOH\,(aq) + HC_2H_3O_2\,(aq) \rightarrow NaC_2H_3O_2\,(aq) + H_2O\,(l)$$

The sodium hydroxide will be gradually added to the vinegar in small amounts from a burette. A burette is a device that allows the precise delivery of a specific volume of a solution. The NaOH will be added to the vinegar sample until all the acetic acid in the vinegar has been exactly consumed (reacted away). At this point the reaction is completed, and no more NaOH is required. This is called the *equivalence point* of the titration.

In order to know when the equivalence point is reached, an *indicator solution* called phenolphthalein is added to the vinegar at the beginning of the titration. Phenolphthalein is a pH sensitive organic dye. Phenolphthalein is colorless in acidic solutions like vinegar, and deep pink in basic solutions like sodium hydroxide. At the equivalence point of the titration, just one drop of NaOH will cause the entire solution in the Erlenmeyer flask to change from colorless to a very pale pink.


In acidic solution


In basic solution


equivalence point

As the titration is performed, the following data will be collected: (1) the molarity of NaOH (*aq*) used, (2) the volume of NaOH (*aq*) used to neutralize the vinegar, and (3) the volume of vinegar used. Using this data, the molarity and mass percent of acetic acid in vinegar can be determined by performing a series of solution stoichiometry calculations (see Calculations Section).

Procedure

Safety

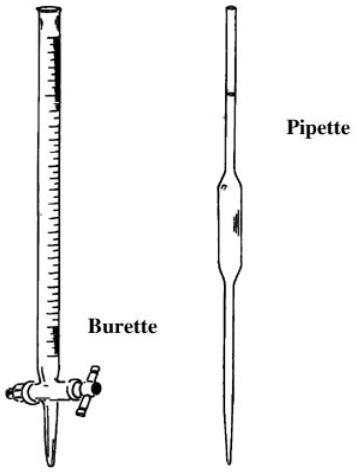
Be especially careful when handling the sodium hydroxide base (NaOH), as it is corrosive and can cause chemical burns to the skin. If any NaOH spills on you, rinse immediately under running water for up to 15 minutes and report the accident to your instructor.

Materials and Equipment

50-mL burette*, 5-mL volumetric pipette*, pipette bulb*, ~ 0.1 M NaOH (*aq*), vinegar, phenolphthalein, burette stand, two 250-mL (or 125 mL) Erlenmeyer flasks, wash bottle with distilled water, funnel

Titration Procedure

Your instructor will demonstrate the correct use of the volumetric pipette and burette at the beginning of the lab session. Detailed instructions on how to use a pipette are also found on the last page of this handout. Note that <u>three</u> titrations must be performed.



1. Obtain a 50-mL burette, 5-mL volumetric pipette and a pipette bulb from the stockroom.

Setting up the burette and preparing the NaOH

2. Rinse the inside of the burette with distilled water. Allow the distilled water to drain out through the tip in order to ensure that the tip is also rinsed.

3. Now rinse the burette with a small amount of NaOH (*aq*). To do this, add about 5-mL of NaOH (*aq*) to the burette, then twirl the burette on its side (over the sink) to rinse its entire inner surface. Then allow the NaOH (*aq*) to drain out through the tip.

4. Fill the burette with NaOH (*aq*) up to the top, between 0-mL and 5-mL. Use a funnel to do this carefully, below eye-level, and preferably over the sink. After this you will need to flush the tip of the burette – your instructor will show you how to do this. Now measure the volume at the level of the NaOH precisely, and record it as the "Initial Burette Reading" on your report. Also record the exact molarity of the NaOH (*aq*), which is labeled on the stock bottle.

Preparing the vinegar sample

5. The volumetric pipette used in this lab is designed to measure and transfer exactly 5.00 mL of solution. First, rinse the inside of the volumetric pipette with distilled water. Using the pipette bulb, draw the water into the pipette up above the 5-mL mark, then allow it to drain out through the tip. You may want to do this several times for practice. Then perform a final rinse, but this time use vinegar.

6. Now use the volumetric pipette to transfer 5.00-mL of vinegar into a clean 250-mL Erlenmeyer flask (see instructions on page 4). Record this volume of vinegar (precise to two decimal places) on your report. Then add about 20-mL of distilled water and 5 drops of phenolphthalein to this Erlenmeyer flask.

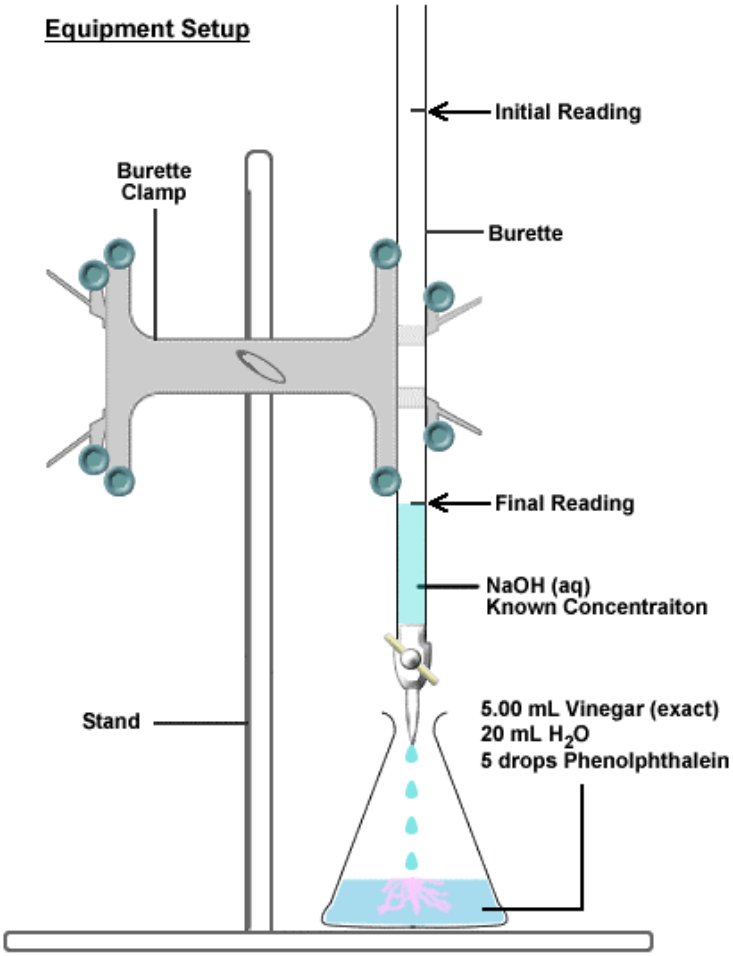
Performing the titration

7. Begin the titration by slowly adding NaOH (*aq*) from the burette to the vinegar in the Erlenmeyer flask. Swirl Erlenmeyer flask as you add the base in order to efficiently mix the chemicals. Some pinkness may appear briefly in the flask as the base is added, but it will quickly disappear as the flask is swirled.

8. As the equivalence point is approached, the pink color will become more pervasive and will take longer to disappear. When this occurs, start to add the NaOH (*aq*) ***drop by drop***. Eventually the addition of just one drop of NaOH (*aq*) will turn the solution in the Erlenmeyer flask a pale pink color that does not disappear when swirled. This indicates that the equivalence point has been reached. ***Do not add any more NaOH (aq) at this point***. Measure this volume of NaOH (*aq*) precisely, and record it as the "Final Burette Reading" on your report. Then show the resulting solution in the flask to your instructor so s/he can record the final color on your report form.

9. Refill your burette with NaOH (*aq*), and then repeat this procedure for a second sample of vinegar, and then a third sample of vinegar. You do not need to flush the tip of the burette again. Note that if you use less than 25-mL of NaOH (*aq*) for the second titration, you do not need to refill the burette for the third titration; also that you will need to clean out and re-use one of your Erlenmeyer flasks for the third titration. You and your partner should take turns performing these titrations.

10. When finished, dispose of your chemical waste as instructed.



Pipetting Instructions

a. Get the appropriate amount of the solution you wish to pipette in a clean, dry beaker. Never pipette directly out of the stock bottles of solution. This creates a contamination risk.

b. Insert the tip of the pipette into the beaker of solution so that it is about a quarter inch from the bottom. Be sure not to press the tip against the bottom of the container.

c. If you are right handed, hold the pipette in your right hand, leaving your index finger free to place over the top of the pipette. With your left hand, squeeze the pipette bulb. Press it firmly over the top of the pipette, but DO NOT INSERT THE PIPET DEEP INTO THE BULB!

d. Release the pressure on the bulb and allow the solution to be drawn up into the pipette until it is above the volume mark. Do not allow the solution to be sucked into the bulb itself.

e. Quickly remove the bulb and place your index finger firmly over the top of the pipette. Then remove the pipette tip from the beaker of solution.

f. Slowly roll your finger to one side and allow the liquid to drain until the bottom of the meniscus is aligned with the volume mark. With practice you will be able to lower the liquid very, very slowly.

g. When the bottom of the meniscus is even with the volume mark, press your index finger firmly on the top of the pipette so no liquid leaks out. Touch the tip once to the side of the beaker to remove any hanging drops.

h. To transfer the solution, place the tip of the pipette against the wall of the receiving container at a slight angle. Then allow the liquid to drain from the pipette.

i. When the solution stops flowing, touch the pipette once to the side of the receiving container to remove any hanging drops. DO NOT blow out the remaining solution. The pipette has been calibrated to deliver the appropriate amount of solution with some remaining in the tip.

Calculations

Molarity of Acetic Acid in Vinegar

- First, using the known molarity of the NaOH (*aq*) and the volume of NaOH (*aq*) required to reach the equivalence point, calculate the moles of NaOH used in the titration.
- From this mole value (of NaOH), obtain the moles of $HC_2H_3O_2$ in the vinegar sample, using the mole-to-mole ratio in the balanced equation.
- Finally, calculate the molarity of acetic acid in vinegar from the moles of $HC_2H_3O_2$ and the volume of the vinegar sample used.

Mass Percent of Acetic Acid in Vinegar

- First, convert the moles of $HC_2H_3O_2$ in the vinegar sample (previously calculated) to a mass of $HC_2H_3O_2$, via its molar mass.
- Then determine the total mass of the vinegar sample from the vinegar volume and the vinegar density. Assume that the vinegar density is 1.000 g/mL (= to the density of water).
- Finally, calculate the mass percent of acetic acid in vinegar from the mass of $HC_2H_3O_2$ and the mass of vinegar.

Name: _____ Chem 10, Section: _____

Lab Partner: _____ Experiment Date: _____

Titration of Vinegar

Experimental Data

	Trial 1	Trial 2	Trial 3
(a) Initial Buret Reading			
(b) Final Buret Reading			
(c) Volume of NaOH (*aq*) used			
(d) Molarity of NaOH (*aq*) used			
(e) Volume of Vinegar used			
Color at equivalence point – to be recorded by *your instructor*			

Data Analysis

Write the balanced equation for the neutralization reaction between aqueous sodium hydroxide and acetic acid.

_____

The Molarity of Acetic Acid in Vinegar

Use your two best sets of results (with the palest pink equivalence points) along with the balanced equation to determine the molarity of acetic acid in vinegar. Show all work for each step in the spaces provided.

Data used ⇒	Trial _____	Trial _____
Moles of NaOH used in titration		
Moles of $HC_2H_3O_2$ neutralized in vinegar sample		
Molarity of $HC_2H_3O_2$ in vinegar		
Average Molarity		

The Mass Percent of Acetic Acid in Vinegar

Use your two best sets of results along with calculated values in the previous table to determine the mass percent of acetic acid in vinegar. Show all work for each step in the spaces provided.

Data used ⇒	Trial _____	Trial _____
Mass of $HC_2H_3O_2$ in vinegar sample		
Mass of vinegar sample (assume density = 1.00 g/mL)		
Mass Percent of $HC_2H_3O_2$ in vinegar		
Average Mass Percent		

Questions

1) What was the purpose of the phenolphthalein indicator in this experiment? Be specific.

2) Suppose you added 40 mL of water to your vinegar sample instead of 20 mL. Would the titration have required more, less or the same amount of NaOH (*aq*) for a complete reaction? Explain.

3) Consider a 0.586 M aqueous solution of barium hydroxide, $Ba(OH)_2$ (*aq*).

a. How many grams of $Ba(OH)_2$ are dissolved in 0.191 dL of 0.586 M $Ba(OH)_2$ (*aq*)?

b. How many *individual* hydroxide ions (OH^{-1}) are found in 13.4 mL of 0.586 M $Ba(OH)_2$ (*aq*)?

c. What volume (in L) of 0.586 M $Ba(OH)_2$ (*aq*) contains 0.466 ounces of $Ba(OH)_2$ dissolved in it?

d. If 16.0 mL of water are added to 31.5 mL of 0.586 M $Ba(OH)_2$ (*aq*), what is the new solution molarity?

e. Suppose you had titrated your vinegar sample with barium hydroxide instead of sodium hydroxide:

$$Ba(OH)_2\ (aq)\ +\ 2\ HC_2H_3O_2\ (aq)\ \rightarrow\ Ba(C_2H_3O_2)_2\ (aq)\ +\ 2\ H_2O\ (l)$$

What volume (in mL) of 0.586 M $Ba(OH)_2$ (*aq*) must be added to a 5.00 mL sample of vinegar to reach the equivalence point? Use your average vinegar molarity (see page 1) in this calculation.

Name: _____ Chem 10, Section: _____

Prelab Assignment: Chemical Equilibrium and Le Chatelier's Principle

1. Consider the reversible reaction: $A + B \leftrightarrows C + D$

a. What happens to the forward and reverse reaction *rates* when equilibrium is achieved?

b. What happens to the reactant (A and B) and product (C and D) *concentrations* when equilibrium is achieved?

2. Le Chatelier's Principle states that if a stress is applied to a reversible reaction at equilibrium, the reaction will undergo a shift in order to re-establish its equilibrium. Consider the following *exothermic* reversible reaction at equilibrium:

$$2\,A \leftrightarrows B + C$$

In which direction (left or right) would the following stresses cause the system to shift?

a. decrease the concentration of A _____

b. increase the concentration of B _____

c. lower the temperature _____

3. In this lab you will explore the effect of Le Chatelier's Principle on several chemical systems at equilibrium. These are supplied on page 2 of the Theory Section. Consider the *third* system you will study: the Aqueous Ammonia Solution.

a. Write the balanced equation for this reversible reaction.

_____

b. Suppose you added some excess ammonium ions to this system at equilibrium.

-- In which direction would a shift occur? _____

-- What color change might you expect to observe? _____

4. List all the equipment you will use in this lab.

Chemical Equilibrium and Le Chatelier's Principle

Objectives

The objective of this lab is to observe the effect of an applied stress on chemical systems at equilibrium.

Background

A reversible reaction is a reaction in which both the conversion of reactants to products (forward reaction) and the re-conversion of products to reactants (backward reaction) occur simultaneously:

forward reaction $\underset{\text{Reactants}}{A + B} \rightarrow \underset{\text{Products}}{C + D}$

backward reaction $\underset{\text{Products}}{C + D} \rightarrow \underset{\text{Reactants}}{A + B}$

reversible reaction $A + B \leftrightarrows C + D$

Consider the case of a reversible reaction in which a concentrated mixture of only A and B is supplied. Initially the forward reaction rate ($A + B \rightarrow C + D$) is fast since the reactant concentration is high. However as the reaction proceeds, the concentrations of A and B will decrease. Thus over time the forward reaction slows down. On the other hand, as the reaction proceeds, the concentrations of C and D are increasing. Thus although initially slow, the backward reaction rate ($C + D \rightarrow A + B$) will speed up over time. Eventually a point will be reached where the rate of the forward reaction will be equal to the rate of the backward reaction. When this occurs, a state of chemical equilibrium is said to exist. Chemical equilibrium is a dynamic state. At equilibrium both the forward and backward reactions are still occurring, but the concentrations of A, B C and D remain constant.

A reversible reaction at equilibrium can be disturbed if a stress is applied to it. Examples of stresses include increasing or decreasing chemical concentrations, or temperature changes. If such a stress is applied, the reversible reaction will undergo a shift in order to re-establish its equilibrium. This is known as **Le Chatelier's Principle**.

Consider a hypothetical reversible reaction already at equilibrium: $A + B \leftrightarrows C + D$. If, for example, the concentration of A is increased, the system would no longer be at equilibrium. The rate of the forward reaction ($A + B \rightarrow C + D$) would briefly increase in order to reduce the amount of A present and would cause the system to undergo a net shift to the right. Eventually the forward reaction would slow down and the forward and backward reaction rates become equal again as the system returns to a state of equilibrium. Using similar logic, the following changes in concentration are expected to cause the following shifts:

Increasing the concentration of A or B causes a shift to the right.

Increasing the concentration of C or D causes a shift to the left.

Decreasing the concentration of A or B causes a shift to the left.

Decreasing the concentration of C or D causes a shift to the right.

In other words, if a chemical is added to a reversible reaction at equilibrium, a shift away from the added chemical occurs. When a chemical is removed from a reversible reaction at equilibrium, a shift towards the removed chemical occurs.

A change in temperature will also cause a reversible reaction at equilibrium to undergo a shift. The direction of the shift largely depends on whether the reaction is exothermic or endothermic. In exothermic reactions, heat energy is released and can thus be considered a product. In endothermic reactions, heat energy is absorbed and thus can be considered a reactant.

exothermic $A + B \leftrightarrows C + D + \text{heat}$

endothermic $A + B + \text{heat} \leftrightarrows C + D$

As a general rule, if the temperature is increased, a shift away from the side of the equation with "heat" occurs. If the temperature is decreased, a shift towards the side of the equation with "heat" occurs.

In this lab, the effect of applying stresses to a variety of chemical systems at equilibrium will be explored. The equilibrium systems to be studied are given below:

1) Saturated Sodium Chloride Solution

$NaCl\,(s) \leftrightarrows Na^{+1}\,(aq) + Cl^{-1}\,(aq)$

2) Acidified Chromate Solution

$2\,CrO_4^{-2}\,(aq) + 2\,H^{+1}\,(aq) \leftrightarrows Cr_2O_7^{-2}\,(aq) + H_2O\,(l)$
yellow orange

3) Aqueous Ammonia Solution (with phenolphthalein)

$NH_3\,(aq) + H_2O\,(l) \leftrightarrows NH_4^{+1}\,(aq) + OH^{-1}\,(aq)$
clear pink

4) Cobalt(II) Chloride Solution

$Co(H_2O)_6^{+2}\,(aq) + 4\,Cl^{-1}\,(aq) \leftrightarrows CoCl_4^{-2}\,(aq) + 6\,H_2O\,(l)$
pink blue

5) Iron(III) Thiocyanate Solution

$Fe^{+3}\,(aq) + SCN^{-1}\,(aq) \leftrightarrows Fe(SCN)^{+2}\,(aq)$
pale yellow colorless deep red

By observing the changes that occur (color changes, precipitate formation, etc.) the direction of a particular shift may be determined. Such shifts may then be explained by carefully examining the effect of the applied stress as dictated by Le Chatelier's Principle.

Procedure

Safety

All of the acids and bases used in this experiment (NH_3, HCl, HNO_3 and NaOH) can cause chemical burns. In particular, concentrated 12M HCl is extremely dangerous! If any of these chemicals spill on you, immediately rinse the affected area under running water and notify your instructor. Also note that direct contact with silver nitrate ($AgNO_3$) will cause dark discolorations to appear on your skin. These spots will eventually fade after repeated rinses in water. Finally, in Part 4 you will be heating a solution in a test tube directly in a Bunsen burner flame. If the solution is overheated it will splatter out of the tube, so be careful not to point the tube towards anyone while heating.

Materials and Equipment

Equipment: 10 small test tubes, test tube rack, test tube holder, Bunsen burner, 2 medium-sized beakers (for stock solutions), 10-mL graduated cylinder, wash bottle, stirring rod, and scoopula.

Chemicals: solid NH_4Cl (s), saturated NaCl (*aq*), concentrated 12M HCl (*aq*), 0.1M $FeCl_3$ (*aq*), 0.1M KSCN (*aq*), 0.1M $AgNO_3$ (*aq*), 0.1M $CoCl_2$ (*aq*), concentrated 15M NH_3 (*aq*), phenolphthalein, 0.1M K_2CrO_4 (*aq*), 6M HNO_3 (*aq*), and 10% NaOH (*aq*).

Experimental Procedure

Record all observations on your report form. These should include, but not be limited to, color changes and precipitates. Note that solution volumes are approximate for all reactions below. Dispose of all chemical waste in the plastic container in the hood.

<u>Part 1: Saturated Sodium Chloride Solution</u>

a. Place 3-mL of saturated NaCl (*aq*) into a small test tube.
b. ***Carefully*** add concentrated 12M HCl (*aq*) drop-wise to the solution in the test tube until a distinct change occurs. Record your observations.

<u>Part 2: Acidified Chromate Solution</u>

a. Place 3-mL of 0.1M K_2CrO_4 (*aq*) into a small test tube.
b. Add an equal amount of 6M HNO_3 (*aq*) to this solution. Record your observations.
c. Now add 10% NaOH (*aq*) drop-wise until the original color is returned. Record your observations. Here the added sodium hydroxide is effectively removing acidic hydrogen ions from the equilibrium system via a neutralization reaction: H^{+1} (*aq*) + OH^{-1} (*aq*) $\rightarrow$ H_2O (*l*).

<u>Part 3: Aqueous Ammonia Solution</u>

Instructor Prep: At the beginning of lab prepare a stock solution of aqueous ammonia. Add 4 drops of concentrated 15M NH_3 (*aq*) and 3 drops of phenolphthalein to a 150-mL (medium) beaker, top it up with 100-mL of distilled water, and mix with a stirring rod. Label the beaker and place it on the front desk. The entire class will then use this stock solution in Part 3.

a. Place 3-mL of the prepared stock solution into a small test tube.
b. Add a medium scoop of NH_4Cl powder to the solution in this test tube. Record your observations.

Part 4: Cobalt(II) Chloride Solution

a. Place 3-mL of 0.1M $CoCl_2$ (*aq*) into 3 small test tubes. Label these test tubes 1-3.
b. The solution in test tube #1 remains untouched. It is a control for comparison with other tubes.
c. To the solution in test tube #2, carefully add concentrated 12M HCl (*aq*) drop-wise until a distinct color change occurs. Record your observations.
d. To the solution in test tube #3, first add a medium scoop of solid NH_4Cl. Then heat this solution directly in your Bunsen burner flame (moderate temperature). Firmly hold test tube #3 with your test tube holder, and waft it back and forth through the flame (to prevent overheating and "bumping") for about 30 seconds, or, until a distinct change occurs. Record your observations. Then cool the solution in test tube #3 back to room temperature by holding it under running tap water, and again record your observations.

Part 5: Iron(III) Thiocyanate Solution

Instructor Prep: At the beginning of lab prepare a stock solution of iron(III) thiocyanate. Add 1-mL of 0.1M $FeCl_3$ (*aq*) and 1-mL of 0.1M KSCN (*aq*) to a 150-mL (medium) beaker, top it up with 100-mL of distilled water, and mix with a stirring rod. Label the beaker and place it on the front desk. The entire class will then use this stock solution in Part 5.

a. Place 3-mL of the prepared stock solution into 4 small test tubes. Label these test tubes 1-4.
b. The solution in test tube #1 remains untouched. It is a control for comparison with other tubes.
c. To the solution in test tube #2, add 1-mL of 0.1M $FeCl_3$ (*aq*). Record your observations.
d. To the solution in test tube #3, add 1-mL of 0.1M KSCN (*aq*). Record your observations.
e. To the solution in test tube #4, add 0.1M $AgNO_3$ (*aq*) drop-wise until all the color disappears. A light precipitate may also appear. Record your observations. Here the added silver nitrate is effectively removing thiocyanate ions from the equilibrium system via a precipitation reaction:
Ag^{+1} (*aq*) + SCN^{-1} (*aq*) $\rightarrow$ AgSCN (*s*).

Name: _____ Chem 10, Section: _____

Lab Partner: _____ Experiment Date: _____

Chemical Equilibrium and Le Chatelier's Principle

1) Saturated Sodium Chloride Solution

Equilibrium System: _____

Observations upon addition of HCl:

_____

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? _____
Explain.

2) Acidified Chromate Solution

Equilibrium System: _____

Observations upon addition of HNO_3:

_____

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? _____
Explain.

Observations upon addition of NaOH:

_____

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? _____
Explain.

3) Aqueous Ammonia Solution

Equilibrium System: _____

Observations upon addition of NH_4Cl:

_____

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? _____
Explain.

4) Cobalt(II) Chloride Solution

Equilibrium System: _____

Observations upon addition of HCl:

_____

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? _____
Explain.

Observations upon heating:

_____

In which direction did heating cause the equilibrium system to shift? Left or Right

Based on these results, is this reaction (as written) exothermic or endothermic: _____
Explain:

5) Iron(III) Thiocyanate Solution

Equilibrium System: _____

Observations upon addition of $FeCl_3$:

_____

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? _____
Explain.

Observations upon addition of KSCN:

_____

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? _____
Explain.

Observations upon addition of $AgNO_3$:

_____

In which direction did this stress cause the equilibrium system to shift? Left or Right

Which ion caused the shift? _____
Explain.