LABORATORY MANUAL FOR

CHEMISTRY 68/60

Prepared by:

Department of Physics and Chemistry
Los Angeles Valley College

This Book Belongs To

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Contents

LABORATORY SAFETY AND EQUIPMENT
INTRODUCTION:

The chemistry laboratory, with its equipment, glassware, and chemicals, has the potential for accidents. Everyone doing experimental work must be aware of and follow established safety rules. Following the safety rules for handling chemicals and working with laboratory equipment will provide a safe environment. Everyone working in the chemistry laboratory should follow the policies outlined below.

A. Come to the Laboratory Prepared for Safety

1. At the beginning of the semester, you were given the schedule of laboratory experiments. Read the experiment **BEFORE** you come to class. Make sure you fully understand the experiment before starting the actual work. If you have a question, ask your professor for clarification **BEFORE** starting the procedure.

2. Do only the experiments that have been assigned by your professor. No unauthorized experiments will be allowed.

3. **SAFETY GOGGLES MUST BE WORN AT ALL TIMES IN THE LABORATORY.**

4. **DO NOT WEAR CONTACT LENSES IN THE LABORATORY**

5. Do not wear loose clothing to lab. It is a fire hazard. Wear closed shoes. Tie back long hair so it does not fall into chemicals or into a flame from a Bunsen burner.

6. Learn the location and use of the emergency eye-wash fountain, emergency shower, and fire extinguishers. Memorize their locations in the laboratory. Know the location of the exits in the lab.

7. **NO FOOD OR DRINK IS ALLOWED IN THE LABORATORY**. Never put anything into your mouth while you are in the laboratory. Wash your hands before leaving lab.

8. Behave in a responsible manner while in lab. Be aware of the other students around you.

9. Keep the lab bench clear of all personal items not needed for the experimental work. Store backpacks, purses, and coats in the storage area provided.

B. Handle Chemicals and Equipment in a Safe Manner

1. Double check the label on the container before you remove a chemical. To avoid contamination of the chemical reagents, **NEVER** insert droppers, pipets or spatulas into the reagent bottles.

2. Take only the quantity of chemical needed for the experiment. Pour or transfer a chemical into a small, clean container from your locker. Label the container. Do not take the stock container to your desk.

3. **DO NOT RETURN UNUSED CHEMICALS TO THE ORIGINAL STOCK CONTAINERS.** You risk contamination of the chemicals. Follow your professor's instructions for disposal of unused chemicals.

4. Do not shake laboratory thermometers. Laboratory thermometers respond quickly to the temperature of their environment. Shaking a thermometer is unnecessary and can cause breakage.

5. **Clean up spills.** Spills of chemicals or water in the work area or on the floor should be cleaned up immediately. Small spills of liquid can be cleaned up with a paper towel. Use Sodium Bicarbonate to neutralize any acid spills. **Mercury** spills require special attention. Notify your professor if you break a thermometer so that special methods can be used to remove the mercury.

6. Dispose of broken glass in the special containers provided. **Do <u>not</u>** put broken glass in the wastepaper basket.

7. Heat only heat-resistant glassware (marked Pyrex or Kimax). Other glassware may shatter when heated. Be very careful of hot objects. Iron or glass looks the same when it is hot as it does at room temperature.

8. Be careful of fires. Small fires can be extinguished by covering them with a watch glass. If a larger fire is involved, a fire extinguisher can be used. If clothing or hair catches on fire the rule is **drop and roll** to extinguish the flames.

9. Report any injuries that occur in the laboratory to your professor.

PROCEDURE:

1. Watch the Safety Video

2. Observe your professor's Safety demonstration in the laboratory. Learn the location of the Safety equipment in your lab. Note the location of the exits from the lab.

3. You will be assigned a locker. Go through the equipment list and make sure all of the items listed are present and in good condition. **YOU ARE RESPONSIBLE FOR THIS EQUIPMENT. YOU MUST RETURN THIS EQUIPMENT AT THE END OF THE SEMESTER.**

4. You must learn the names and functions of all the equipment in your drawer.

5. Sign your Safety Pledge. Fill out your locker card **COMPLETELY**. Give the Safety Pledge and locker card to your professor, and print your name on the locker list.

Quantity	Description
1	Beaker, 100 mL
1	Beaker, 150 mL
1	Beaker, 250 mL
1	Beaker, 400 mL
1	Ceramic Tile
1	Clamp, Buret
1	Crucible Cover
1	Crucible, Porcelain, 30 mL
1	Cylinder, Graduated, 50 mL
1	Dish, Porcelain, Evaporating
3	Dropper
1	Flask, Erlenmeyer, 125 mL
1	Flask, Erlenmeyer, 250 mL
1	Holder, Test Tube
1	Rack, Test Tube
1	Rule, Metric
3	Shell Vials
3	Stir Rods, Glass
1	Spot Plate
1	Test Tube Brush
8	Test Tube, 15 mm x 125 mm
1	Test Tube, 25 mm x 200 mm
1	Thermometer, -20°C to 110°C
1	Tongs
1	Triangle, Clay
1	Watch Glass, 100 mm
1	Wire Gauze

LABORATORY SAFETY RULES

Note: Failure to follow safety rules will result in expulsion from this course.

1. Wear approved safety goggles AT ALL TIMES in the laboratory.
2. It is not advisable to wear contact lenses during lab.
3. Do not wear loose clothing to lab. It is a fire hazard.
4. Tie back long hair. It too is a fire hazard.
5. Wear closed shoes to lab.
6. Never put anything into your mouth while in the lab.
7. Immediately wash off any chemicals spilled on your skin or clothes.
8. Keep the lab neat. **Return reagent containers and equipment to proper locations.** Put any belongings not needed for experimental work on the shelves provided.
9. Clean up all chemical spills or broken glass immediately. You should report Mercury spills to your instructor or the stockroom (do NOT attempt to clean up spilt Mercury).
10. Think about how much chemical you will need before you take it from a stock (reagent) bottle. NEVER return unused chemicals to stock bottles.
11. Dispose of waste chemicals only as instructed.
12. Behave in a responsible manner.
13. Be aware of the location and use of laboratory safety equipment.
14. Immediately report accidents and injuries to your professor.
15. Do NOT perform unauthorized experiments
16. Thoroughly wash your hands any time you leave the lab.
17. No smoking on the Los Angeles Valley College campus.

I have carefully read all of the safety precautions summarized above and recognize that it is my responsibility to observe them throughout this course.

Chemistry 68		Printed Name
Date	Section Number	Signature

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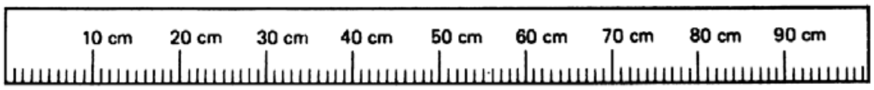
COME TO LAB PREPARED!! CAREFULLY READ THE EXPERIMENT BEFORE COMING TO LAB.

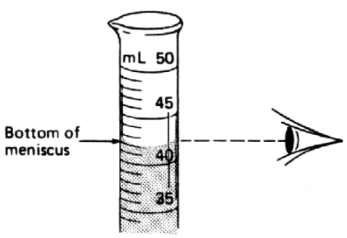
MEASUREMENTS: LENGTH, VOLUME, AND TEMPERATURE
INTRODUCTION:

The world uses a variety of units to measure length. Scientists use the metric system in which the unit of length is the meter (m). Using appropriate prefixes, one can indicate a length that is greater or less than a meter. A meter stick can be divided into 100 centimeters. Each centimeter could be divided into 10 millimeters.



When a meter stick or other measuring device is used, the measurement must be reported as precisely as possible. The number of significant figures that can be included depend on the markings on the device that is used. When a piece of data is recorded, the **next to the last** digit reported is the number represented by the smallest increment marked on the measuring instrument. The **last** digit is an estimate. If the quantity being measured is estimated to fall exactly on a line marked on the measuring device, the last digit of the measured number is a significant zero. If the quantity being measured is seen to fall between two lines, an estimate is made as to the distance the quantity is between the lines. This estimated number becomes the last digit recorded for the measurement.

In this experimental procedure, you will use graduated cylinders to determine the volumes of several substances. To read the volume of liquid properly you must avoid parallax. You should set the cylinder on a level surface and bring your eyes to a level even with the top of the liquid. You will notice that the liquid level is not a straight line, but curves down at the center. This curve, called a *meniscus*, is read at its lowest point (center) to obtain the volume measurement of the liquid. In the graduated cylinder shown, the volume of the liquid can be read as 42.1 mL. (Note: the smallest markings on the cylinder shown are for 1 mL increments. By estimating the volume **between** the 1-mL markings, the volume can be reported to the tenths (0.1) of a milliliter.)



Temperature can be measured by several different methods. One method uses the fact that when most liquids are heated, their volume increases are almost directly proportional to the change in temperature. A mercury-in-glass thermometer, the type that will be used in this experiment, is constructed of a very small diameter capillary tube connected to a relatively large reservoir of mercury in a bulb. As the mercury is warmed it expands into the capillary tube. The change in height of the mercury column can be calibrated to correspond to a temperature scale. The scale most often used by chemists is the Celsius scale in which the normal freezing point temperature of water is 0 °C and the normal boiling point temperature of water is 100 °C.

PROCEDURE:

A. Length Measurements

Use the metric scale on a ruler or meter stick to make the measurements indicated on the report sheet (be sure to record the number of significant figures appropriate for the measuring device(s) that you use). In some cases you will need to use a piece of string to determine the distance (around your wrist, for example).

B. Measuring Volumes of Liquids

1. A display of graduated cylinders containing different volumes of liquids has been set up for you. Review the information in the Introduction concerning parallax, reading a meniscus, and reporting significant figures. Read and record the volume of the liquid in each graduated cylinder using the number of significant figures appropriate for that cylinder.

2. Using a measuring cup (available on the cart), measure out 1 cup of water as exactly as possible. Use a 250-mL or 500-mL graduated cylinder (also available on the cart) to measure the volume of the water in mL. Record your answer. Repeat this procedure two more times. Average your data to obtain an experimental value for the number of mL in a cup. Use English-Metric Unit conversion factors found in your textbook to calculate the true (actual) value for the number of mL in 1 cup. Comment on how close your experimental value is to the true value.

C. Temperature Measurements

The bulb of the thermometer must remain in the liquid while the measurement is being made. Beakers of water at different temperatures are available in the hoods. Measure and record the temperature of each in °C. Convert each temperature to Kelvin.

D. English-Metric Conversions

A selection of product containers is available in the lab. Pick four products and record the volume of the contents of each in both fluid ounces (English system), and Liters (metric system). Based on the volumes given, calculate an English/metric conversion factor in fluid ounces/L from each set of data.

REPORT
MEASUREMENTS: LENGTH, VOLUME,
AND TEMPERATURE

NAME ______________________________

SECTION ___________________________

A. Length measurements

Complete the following equalities:

1 cm = ____________ m   1 km = ______________ m   1 mm = _____________ m

Make each of the following measurements **in centimeters**, and then convert the measurements to millimeters and meters.

	Measured number	Calculated numbers	
	cm	mm	m
Width of little finger nail			
Width of desk top			
Length of shoe			
Your height			
Length of a pencil			

Determine the length and width of the sides of the rectangle shown above **in centimeters.** Have a second person repeat the measurements. Record both sets of data below. Calculate the area of the rectangle each of the sets of data..

	Measurement (first person)	Measurement (second person)	Measurement (Average)
Length, cm			
Width, cm			
Area, cm^2			

Would you expect the sets of data above to be identical? **Explain** why or why not.

REPORT FOR MEASUREMENTS: NAME________________________
LENGTH, VOLUME & TEMP. (cont.)

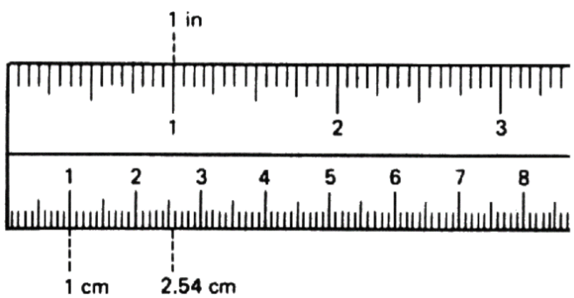
Repeat the measurement of the length and width of the rectangle shown on the previous page using the **inches** scale on a ruler. **Note:** When working in English units it is more difficult to determine how many significant figures to record. Try to estimate the inches to two places after the decimal.

Length = ________________ in. Width = _____________________ in.

Use the **average** length and width measured in centimeters (from the previous page) and length and width in inches (from above) to **calculate** an **experimental** ratio for the number of centimeters in 1 inch.

$$\frac{\_\_\_\_\_\_\_\_\_\_\text{(measured length in cm)}}{\_\_\_\_\_\_\_\_\_\_\text{(measured length in in)}} = \_\_\_\_\_\_\_\_\_\_\_\_ \ {}^{\text{cm}}\!/\!_{\text{in}}$$

$$\frac{\_\_\_\_\_\_\_\_\_\_\text{(measured width in cm)}}{\_\_\_\_\_\_\_\_\_\_\text{(measured width in in)}} = \_\_\_\_\_\_\_\_\_\_\_\_ \ {}^{\text{cm}}\!/\!_{\text{in}}$$



Calculate the percent difference between your experimental ratio and the accepted value of 2.54 cm/in. $\frac{\left(\text{experimental ratio} - 2.54\ {}^{\text{cm}}\!/\!_{\text{in}}\right)}{2.54\ {}^{\text{cm}}\!/\!_{\text{in}}} \times 100$
Show your calculations here.

B. Measuring the Volume of Liquids

Volume in Display Cylinders, (mL)		
Cylinder A	Cylinder B	Cylinder C

REPORT FOR MEASUREMENTS: NAME________________________
LENGTH, VOLUME & TEMP (cont.)

Volume of 1 cup of water, (mL)		
Trial 1	Trial 2	Trial 3
Average Volume of 1 cup, (mL)		

Using conversion factors found in your textbook, calculate the actual number of milliliters in exactly one cup. (One quart is exactly four cups.) Determine the percent difference between your experimental value and the accepted value you just calculated. Show your calculations here.

C. Temperature Measurements

		Temperature in °C	Temperature in Kelvin
1	Ice water		
2	Room temperature water		
3	Warm water		

D. English-Metric Conversions

	Name of Product	Volume in English Units (fluid ounces)	Volume in Metric Units (Liters)	Conversion factor (fluid ounces/L)
1				
2				
3				
4				

What do you notice about the conversion factors obtained from each product?

QUESTIONS FOR MEASUREMENTS: NAME________________________
LENGTH, VOLUME & TEMPERATURE

You may need to look at the English-metric conversion table in your textbook to complete some of these problems. **SHOW YOUR WORK** including all conversion factors.

1. Convert 3.85×10^{-4} kilometers to
   (a) micrometers

   (b) feet

   (c) centimeters

2. A piece of string is found to be 35.9 meters long. How long is the string in inches?

3. A section of lawn that is 25.5 feet by 75.0 feet needs fertilizer. The fertilizer is sold in 5.00 pound boxes and 1.00 pound of fertilizer is needed for 10.0 square yards of lawn. If each box costs $1.65, how much will it cost to fertilize the lawn?

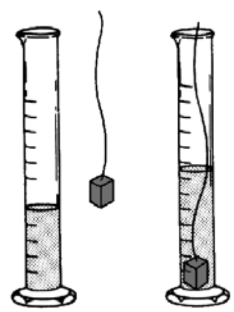
4. If a gallon of gas costs $1.50, how many cents would a liter of gas cost? (Round your answer to the nearest whole number of cents).

5. Convert 75.4 °F to Celsius and then to Kelvin.

MEASURMENTS: VOLUME, MASS AND DENSITY

INTRODUCTION:

Experimental chemistry often involves making a series of measurements. Because measurements are not exact, the data recorded must reflect the precision of the equipment used. The graduated cylinders used to measure volume, and the balances used in this experiment to determine mass, have calibration marks on them. It is the accepted practice to record all the digits of a piece of data that are represented by the calibration marks on the measuring device **plus** one additional digit at the end of the number that is an estimate. Sometimes the same measurement is made more than once and the data are averaged. This practice reduces random errors that result from experimental values that fluctuate around the true value. The number of significant figures in a number that is an average should be the same as the number of significant figures in the least precise piece of data that was used when calculating that average.



Volumes of solids will be determined by two methods. The dimensions of a solid will be determined and its volume calculated in cubic length units. In addition, the volume of a solid will be determined by measuring the volume of liquid displaced. When an object is submerged in a liquid, it displaces its own volume of liquid which causes the liquid level to rise. The difference in the liquid level before and after the object is submerged is due to the volume of the object.

Students will become familiar with the use of the balances in the laboratory and will determine the masses of several objects. Some of these data will be used to calculate densities and percent by mass. Review the material in your textbook for the rules on the correct number of significant figures to use for calculated quantities.

PROCEDURE:

A. Mass Measurements

1. Observe your professor's demonstration of the use of the beam balance.

<u>Note</u>: The smallest increments marked on this balance are the hundredths (0.01) place. Therefore, the thousandths (0.001) place (an estimate) must be recorded.

2. Choose several small objects (pencils, keys, erasers, etc.). Measure and record their mass.

3. Obtain a metal cylinder (with hook) from your professor. **EVERYONE IN THE GROUP IS RESPONSIBLE FOR RETURNING THIS CYLINDER AT THE END OF THE LAB SESSION.**

4. Record the identity of the metal in the cylinder. Determine the mass of the metal cylinder plus hook. **<u>DO</u> <u>NOT</u> REMOVE THE HOOK FROM THE CYLINDER!!** Subtract the mass of the hook (**<u>given</u> <u>on</u> <u>the</u> <u>report</u> <u>sheet</u>**) from the total mass to obtain the mass of the metal.

B. Measuring the Volume of a Solid

1. Obtain a ruler and graduated cylinder from the supply cart. Use a ruler to determine the height and diameter of the metal cylinder in centimeters (cm). Ignore the height of the hook. Calculate the volume of the metal solid in cm^3. Follow the rules for determining the correct number of significant figures for your data.

2. Fill the graduated cylinder about three-fourths full of water. Measure and record the exact volume of the water in the graduated cylinder. Suspend the metal solid from a piece of monofilament line and lower it into the water in the graduated cylinder until the top of the metal solid (but <u>not</u> the hook) is submerged. Read and record the new volume of water. The difference between the two volumes is the volume displaced by the metal solid (in mL). Convert the volume of the metal solid obtained by the displacement method from mL to cm^3. Comment on how close the values obtained by displacement and the dimension measuring methods are to each other. What are some reasons that your two methods might give different numbers?

C. Density of a Solid

1. Determine the volume of the cylinder in cm^3 by averaging the volumes obtained by the displacement and dimension measuring methods.

2. Use the mass of the cylinder and the average volume to calculate the density of the metal in the cylinder.

$$\text{Density of a substance} = \frac{\text{mass of the substance}}{\text{volume of the substance}}$$

D. Density of a Liquid

1. Place about 20 mL of water in a 50-mL graduated cylinder. Record the actual volume of the water to 0.1 mL. <u>Save</u> this water for the next step. **Do not use the markings on beakers to measure volume; they are not accurate.**

2. The mass of a liquid is found by difference. First, measure and record the mass of a small <u>**DRY**</u> beaker. Pour the water from step D.1 into the beaker. Measure and record the combined mass. Calculate the mass of the water.

3. Calculate the density of the water.

E. Percent by Mass

1. Obtain a <u>**DRY**</u> beaker and 2 stoppers from the supply cart. Measure the mass of the beaker alone. Then, place the 2 stoppers into the beaker and measure the combined mass of the beaker and stoppers.

2. What percent of the combined mass is the mass of the stoppers?

$$\text{mass \% of A} = \frac{\text{mass of A}}{\text{total mass}} \times 100$$

F. Thickness of Aluminum Foil

1. Obtain a piece of aluminum foil from your professor. Use a ruler and measure the length and width of the foil in cm (use the correct number of sig. figs.).

2. After the foil has been measured, fold it and determine its mass.

3. From the data given in the table on page 23 of this lab manual, record the density of aluminum.

4. Using the density for aluminum, the mass of your foil, and the area of your foil, determine the thickness of your foil in cm.

DRY THE METAL CYLINDER WITH A PAPER TOWEL AND RETURN IT TO YOUR PROFESSOR. Untie as many knots as possible to the monofilament line and return the line and other equipment to the cart.

REPORT
MEASUREMENTS: VOLUME,
MASS, AND DENSITY

NAME ___________________________

SECTION ________________________

A. Mass Measurements

Masses of Several Objects:

Name of Object	Mass of Object

Identity of the Metal in the Cylinder	
Mass of Metal Cylinder and Hook, (g)	
Mass of Hook, (g) (**DO NOT REMOVE THE HOOK)**	1.250 g
Mass of Metal Cylinder, (g)	

B. Measuring the Volume of a Solid

Metal Cylinder dimensions	radius (cm)	height (cm)
Formula for the volume of a cylinder:		
Volume of Metal Cylinder (Dimension method), (cm^3)		
Volume of Water and Metal Cylinder, (mL)		
Volume of Water alone, (mL)		
Volume of Metal Cylinder (Displacement Method), (mL)		
Volume of Metal Cylinder (Displacement Method), (cm^3)		

REPORT FOR MEASUREMENTS: NAME______________________
VOLUME, MASS & DENSITY (cont.)

C. Density of a Solid

Average Volume of Cylinder used in Part B, (cm^3)	
Experimental Density of the Metal in the Cylinder , (g/cm^3)	

SHOW CALCULATIONS:

D. Density of a Liquid

Volume of Water, (mL)	
Mass of Water and Beaker, (g)	
Mass of Beaker alone, (g)	
Mass of Water, (g)	
Density of Water, (g/cm^3)	

SHOW CALCULATIONS:

E. Percent by Mass

Mass of Beaker and Stoppers, (g)	
Mass of Beaker, (g)	
Mass of Stoppers, (g)	
Mass Percent for Stoppers, (%)	

SHOW CALCULATIONS:

REPORT FOR MEASUREMENTS: NAME___________________________
VOLUME, MASS & DENSITY (cont.)

F. Thickness of Aluminum Foil

Density of Aluminum from table on the next page, (g/cm^3)	
Width of Aluminum Foil, (cm)	
Length of Aluminum Foil, (cm)	
Mass of Aluminum Foil, (g)	
Thickness of Aluminum Foil, (cm)	

SHOW CALCULATIONS:

QUESTIONS FOR MEASUREMENTS: NAME ____________________________
VOLUME, MASS & DENSITY EXP.

1. The table at the right gives the accepted value for the densities of the metals used in Part C of this experiment. Use the experimental value determined for the metal assigned to your group and calculate the percent error:

Metal	Density (g/cm^3)
Copper	8.92
Aluminum	2.702
Tin	7.28
Zinc	7.14
Steel	7.83

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

What are some possible sources of error in doing the density experiment?

2. A sugar solution consists of 35.8 g sugar and 125.35 g of water. What is the percent by mass of the sugar and the percent by mass of the water in the solution?

3. If the solution in problem 2 above has a volume of 132.8 mL, calculate the density of the solution in g/mL.

QUESTIONS FOR MEASUREMENTS: NAME_______________________
VOLUME, MASS & DENSITY EXP. (cont.)

4. In the apothecary system, there are exactly 65 milligrams in one grain. A physician ordered 27 grains of a drug to be administered to a patient. The medication is available in 650.0 mg tablets. How many tablets should be given to the patient? (Round your answer to the nearest whole number of tablets.)

5. The water in a graduated cylinder is measured as having a volume of 25.3 mL. An object with a mass of 9.567 g was submerged in the water of the cylinder and the water rose to a volume equal to 30.1 mL. What was the density of the object?

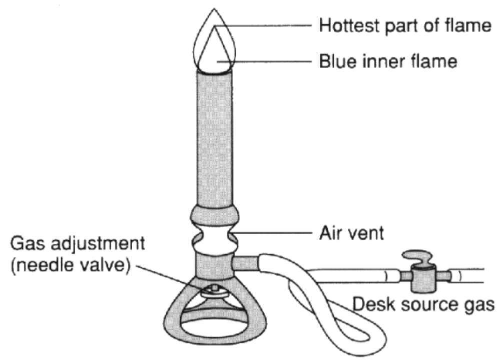
FLAME TESTS

INTRODUCTION:

When elements are exposed to the heat of a flame, electrons in that element can be excited from their ground-state arrangement into higher energy levels. As these excited electrons return to the vacated lower levels excess energy MUST be emitted. If the energy emitted is in the visible region of the electromagnetic radiation spectrum, then a color will be imparted to the flame. Only certain elements have electronic transitions that involve the energy associated with visible light. But, just as the electron arrangement in an element affects its chemistry, the allowed electron transitions impart unique color to the flame. Because the energy of allowed levels for electrons in each element are unique, any color imparted to a flame is distinctive and an element can be identified by its flame test.

PROCEDURE:

A. Bunsen Burner



Observe your professor's demonstration of the correct method of lighting a Bunsen burner.

B. Flame Tests

1. Obtain a flame test wire, cork and two spot plates.

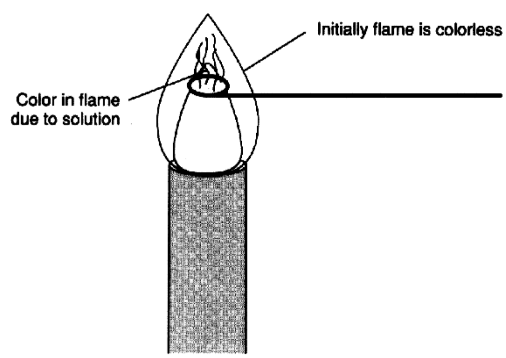
2. Clean the spot-plates but do <u>NOT</u> dry them. After rinsing the clean plates with deionized water, simply shake off the excess water.

3. Put one of the spot plates on a piece of paper and label the paper at the positions of the spot plate wells with LiCl, NaCl, KCl, $CaCl_2$, $BaCl_2$, or $SrCl_2$. Take the spot plate (on the labeled paper) to the stock bottles and **using the droppers that are in the bottles**, place 2-3 drops of each chemical into its labeled well. **DO NOT MIX UP THE DROPPERS. AS YOU USE THEM, RETURN EACH TO THE CORRECT BOTTLE.**

4. Form the end of the flame test wire into a SMALL loop. Insert the straight end of the wire into a cork. The cork will protect your fingers from heat when you place the wire into the flame of the Bunsen burner.

5. <u>Clean</u> the wire by dipping it in a shell vial containing a <u>**few**</u> mL (1/4th full) of 6 <u>M</u> HCl and then placing the wire loop just above the inner blue cone of a burner's flame. Repeat this process until no coloration of the flame is observed. (The wire is clean when no color is imparted to the flame.)

   CLEAN UP ANY ACID SPILLS WITH SODIUM BICARBONATE



<u>NEVER</u> DIP A <u>HOT</u> FLAME-TEST WIRE INTO A KNOWN OR UNKNOWN SOLUTION. THE CATION WILL BE "BAKED" ONTO THE WIRE AND IT WILL BE IMPOSSIBLE TO CLEAN!

6. Dip the clean wire into one of the <u>known</u> solutions. Place the wire into a flame just above the inner blue cone. Observe and record the color imparted to the flame. Sometimes it is easier to see the color when the wire is held just above the visible flame.

7. You MUST clean the flame test wire (see step 4) BEFORE testing a different solution. Repeat the flame test process for the other known solutions. Save your solutions on your spot plate in case you want to recheck your results when you are testing your unknown solutions.

8. Bring your second clean spot plate (on a piece of paper) to your professor who will dispense your unknowns into it. On the paper, label each your unknowns with the numbers given to you by your professor. **Record these identification numbers on your report sheet**.

9. Make sure to clean your wire **WELL** between each test and do a flame test on each of your unknown solutions.

10. Dispose of your excess 6 M HCl by pouring it into a beaker of water and then pour that solution down the sink flushing with additional water. **DISCARD** your flame test wire and return the cork and spot plates.

REPORT
FLAME TESTS

NAME ____________________________

SECTION __________________________

Flame Tests for Known Solutions

Identity of Solution	Flame Color

Identification of Unknowns

Unknown Number	Flame Color	Identity of Unknown Solution

QUESTIONS FOR FLAME TEST EXP. NAME______________________________

Write the **COMPLETE** electron configuration and orbital diagram for each of the following elements:

Na

Kr

N

Use the noble gas short form to write the electron configuration and orbital diagram for the following elements:

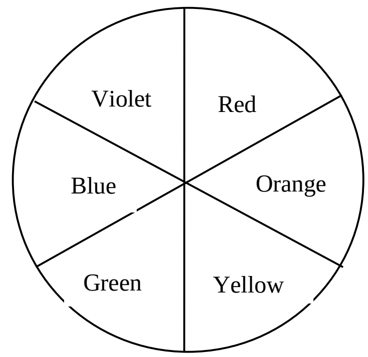
W

Sb

U

LIGHT AND COLOR

Introduction:



The light the human eye can detect is called the visible spectrum. It is a very small part of the entire electromagnetic spectrum. Our eyes are able to distinguish between many different colors. Usually the visible spectrum is broken down into red, orange, yellow, green, blue, and violet. Each color in the spectrum is identified by its wavelength. White light is equal mixture of each of the colors in the visible spectrum. When white light shines on a substance one or more of the colors may be absorbed. The colors we see are those that are reflected or transmitted (the eye sees the colors that are not absorbed). A simplified way to think about this is to look at the color wheel shown on the right. The color(s) our eyes see are the ones that are **opposite** the color(s) being absorbed. For example, if violet light was being absorbed, our eyes would see yellow.

In this experiment, we will use a spectrophotometer to look at the correlation between the wavelength of visible light and its color. Then we will also measure the absorption spectra of two colored solutions. Spectrophotometry is a method that is often used in Chemistry. It has applications ranging from the identification of substances to quantitatively determining the amount of substance present in a solution. There are many types of spectrophotometry using a variety of kinds of light from ultra-violet through visible to infrared. All of these use a portion of the electromagnetic spectrum. The type of spectrophotometer we will use can measure the amount of light absorbed at only one wavelength at a time. Consequently, we must measure the absorption at a variety of wavelengths to obtain the spectrum of a solution. In the spectrophotometers like the ones we will use, there is a source of white light (a light bulb) and a diffraction grating, which works like a prism that separates the light into its component wavelengths. A selected wavelength of light is passed through a sample to a detector and a display shows of the amount of that wavelength of light that is absorbed by the sample.

PROCEDURE:

A. Colors of visible light

1. Turn on the spectrophotometer and let it warm up for a few minutes. Obtain a cuvet. **DO NOT PLACE ANY LIQUID IN THIS CUVET. YOU WILL USE IT EMPTY.**

2. Use the ***nm↑*** and ***↓nm*** buttons to set the wavelength to 525 nm.

3. Open the sample door on the right front side of the machine and leave it open for this part of the experiment. Place the empty cuvet into the holder and adjust it up and down until you see a bar of light on the bottom of the cuvet.

4. Use the nm↑ and nm↓ buttons to set the wavelength to 400 nm. Look down into the tube and see if you can detect a bar of colored light. (Most humans cannot see light with a wavelength of 400 nm. Can you?)

5. Increase the wavelength to 425 nm. Can you see any color now? If so, record the color you see. Continue to increase the wavelength by 25 nm increments (recording any colors that you see) until you reach 700 nm.

B. Spectra of Red and Blue Solutions

1. Part B of this experiment will require three cuvets. Cuvets are made of optically uniform glass that needs to be protected against scratches. **Use only chem. wipes** to clean then and do not allow solutions to stand in them. Rinse them well with deionized water immediately after using them.

2. If necessary, press the ***A/T/C*** button to switch the display so that the ***A*** is shown in the display. The instrument is now set to the display amount light that is being absorbed. Set the wavelength to 400 nm.

3. Fill one cuvet about 1/3 full with deionized water, one about 1/3 full with the red solution and one about 1/3 full will blue solution.

4. In Part B of this experiment **every time** the wavelength is changed the spectrophotometer must be "zeroed" for that wavelength. This means that light absorbed by the solvent and light scattered by the cuvet must be "blanked out" so that is does not register on the display. To "zero," wipe the cuvet containing water (the "blank") with a chem wipe and place it into the cuvet well. **Close the cover** and press the ***0 ABS/100% T*** button. After a short delay, the display should read 0.000 (indicating all of the light is being transmitted, zero is being absorbed).

5. Wipe the tube containing the red solution. Remove the blank tube and put the cuvet containing the red solution into the cuvet well. **Close the cover**. Read and record the absorbance from the display.

6. Repeat step 5 with the blue solution.

7. Increase the wavelength by 25 nm. You must now re-zero the spectrophotometer. Repeat steps 4 through 7. Continue repeating step 4 through step 7 until you reach 700 nm. **NOTE: When you reach 500 nm, you must change the filter wheel to 3 before you continue.**

8. When you have collected your data, rinse the cuvets with deionized water and return them to the cart.

9. Draw a **full page** graph for **each** of the red and blue solutions plotting absorbance (which is unitless) on the **y-axis** and wavelength (in nm) on the **x-axis**. Be sure to follow all the instructions from the graphing lab in regard to choosing intervals, labeling, significant figures, and titles. Note: these will be **"connect the dot"** graphs (**not** straight lines or curves).

REPORT
LIGHT AND COLOR

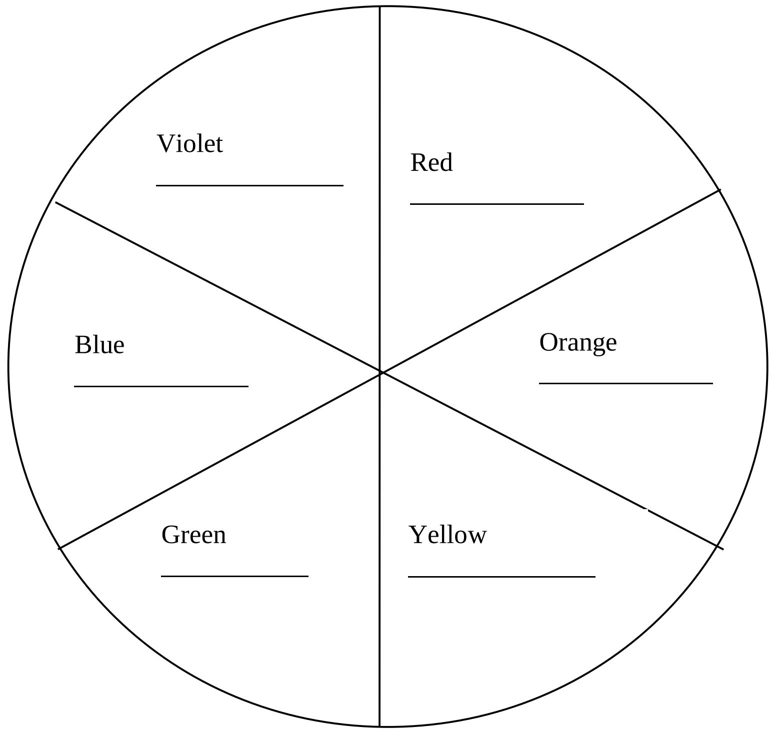
NAME________________________________

SECTION_____________________________

Wavelength(nm)	Color	Red solution ABS	Blue solution ABS
400			
425			
450			
475			
500			
525			
550			
575			
600			
625			
650			
675			
700			

QUESTIONS FOR LIGHT AND COLOR NAME ________________________________

1. Fill in each section of the color wheel with the range of wave lengths that best correspond to the colors you observed. For example, for red you might record 700nm to the lowest wavelength that appeared reddish. The next lowest wavelength would be listed as orange and so on. Your completed wheel should contain all the wavelengths for which you were able to see a color.



2. Based **on your completed graphs**, which wavelength(s) showed the largest absorbance for the

   a. red solution ________________________________

   b. blue solution ________________________________

3. Based on your answers **on the color wheel above**, which color(s) were being most absorbed (had the largest absorbance value) in the

   a. red solution ________________________________

   b. blue solution ________________________________

4. Review the explanation of the color wheel given in the introduction. Based on the color(s) being absorbed, which colors would you **expect** to see for the

   a. red solution ____________________________________

   b. blue solution ____________________________________

5. Do your answers for question number 4 match what you **actually** see for the colors of the two solutions? (If they do not match exactly, are they close?) Explain.

OBSERVATION AND CRITICAL THINKING

INTRODUCTION:

In this experiment, various chemical solutions will be mixed and some of the combinations will result in chemical reactions. In a chemical reaction, changes often can be observed as the original substances (the reactants) form new substances (the products of the reaction). These changes could include the formation of a gas, the formation of a precipitate (a solid), the evolution of heat, color changes, or a combination of changes. When the manner in which certain substances react with each other has been established, this information can be used to determine the identity of those substances when they are used as unknowns.

Critical thinking is a process by which information from knowledge gained in previous experiences and/or data obtained from the present problem is examined. This process involves using logic to look at several possible solutions to the problem. This method of problem solving often uses the process of elimination to discard incorrect answers until only one possible solution (the correct one) remains. The identity of the unknown substance in this experiment will be determined using this method.

PROCEDURE:

BE CAREFUL OF ALL THE CHEMICALS USED IN THIS EXPERIMENT. SOME OF THE SOLUTIONS COULD STAIN OR BURN YOUR SKIN (OR MAKE HOLES IN YOUR CLOTHES)!!!

1. Check out 8 three-inch test tubes from the stockroom. These will be the “experimental” set of test tubes. Clean these tubes.

2. Clean 8 six-inch test tubes (from your lab locker). These will be the “stock” set of test tubes. Rinse both “experimental” tubes and “stock” tubes with deionized water and shake out any residual water. The tubes do not have to be dry for this experiment.

3. Label both the set of 8 “experimental” tubes and the set of 8 “stock” tubes as 1-6, water, and unknown.

4. Using your graduated cylinder, measure out about 6 mL of deionized water and pour it into the “stock” tube labeled “water.”

5. Using a dropper, place 15 drops (which is about ¾ of an mL) of deionized water in the experimental” tube labeled “water.

6. Fill each of the “stock” tubes 1 through 6 with approximately 6 mL of the correspondingly numbered solution from the reagent bottles. Use the “stock” water tube as a visual guide for the volume to take for each of these solutions.

7. Return to your work area with your water and “stock” tubes.

8. Using the “experimental” water tube as a volume guide, pour about ¾ of a mL of your “stock” solution 1 into your experimental tube 1. Pour about ¾ mL of “stock” solution 2 into “experimental tube 2. Continue until ¾ of a mL of each of the “stock” solutions has been transferred to its corresponding “experimental” tube. (Save your unknown “experimental” tube and “stock” tubes for later).

9. Pour the contents of “experimental” tube 1 <u>into</u> the solution in “experimental” tube 2. Mix the contents of “experimental” tube 2 thoroughly and record your observations in the first unshaded square on the top row of the observation table. **Look at the headings at the top of each vertical column. Each horizontal <u>row</u> on the observation table provides space to describe any reaction between the solution in the vertical column and the solution shown in the square at the beginning of the row. You will <u>only</u> do the tests represented**

by squares on the report sheet that are <u>not</u> shaded in. This avoids wasting reagent by doing any test twice or testing a solution with itself.

10. Again measure about ¾ of a mL of **'Stock" solution 1** into "exp." tube 1. Pour the contents of Exp. tube 1 into the solution in Exp. tube 3. Mix well and record your observations on the next square on row one (in the vertical column for Solution 3). Again place ¾ mL of "Stock" solution 1 into Exp. tube 1, mix with Exp. tube 4, and record observations. Continue across the row until you have observed the results of adding solution 1 to solutions 2 through 6.

11. Set Exp. tube 1 aside to be used later (you do not need to wash it). Wash Exp. tubes 2-6 and rinse them with deionized water. Shake the excess water from the tubes. Fill experimental tubes 2-6 with their ¾ mL of their respective solutions.

12. You will now repeat the procedures in steps 9 and 10 **except that** for each test you will be transferring the contents of Exp. tube 2 (containing **solution 2)** into tubes 3, 4, 5, and 6, **refilling Exp. tube 2 with ¾ mL of solution 2** for each transfer and test. Remember to mix well before recording your observations for each test.

13. Set Exp. tube 2 aside to be used later (you do not need to wash it). Wash Exp. tubes 3-6 and rinse them with deionized water. Shake the excess water from the tubes. Fill experimental tubes 3-6 with their ¾ mL of their respective solutions.

14. Continue in the manner outlined above testing all the combinations of the solutions (except the unknown) and filling in the observation table. After the testing on each horizontal row is complete, set aside the Exp. tube that was used to transfer solutions and wash **any tube in which there was a mixture.**

15. Bring your "Stock" unknown tube to your professor to obtain your unknown sample. Record the letter of your unknown.

16. Fill "experimental" tubes 1-6 with about ¾ mL of their corresponding solutions. Fill the "experimental" unknown tube with about ¾ mL of unknown.

17. Pour the contents of the Exp. unknown tube into the solution in Exp. tube 1. Mix well and record your observations.

18. Place another ¾ mL sample of unknown into the Exp. unknown tube. Pour the contents of the Exp. unknown tube into the solution in Exp. tube 2. Mix and record observations. Continue the procedure until the unknown has been tested with all of the Exp. solutions 1-6.

19. Clean all of your test tubes and return the 8 that you checked out to the stockroom.

20. Your unknown solution is the same as one of the 6 known solutions. Use the results of your observation table to determine the identity of your unknown. Do this <u>**using a process of elimination and show your work:**</u>

    a. First, consider the possibility that your unknown might be Solution 1. Look at your observations for the test in which you mixed your unknown with solution 1. Did it react? If so, your unknown could <u>not</u> be solution 1 because in it would not react with itself.

    b. Look at the reaction (if any) for the reactions between solutions 1 and 2, 1 and 3, 1 and 4, 1 and 5, and 1 and 6 (**all on the top row of your observation table)**. Were your observations for the reactions between your unknown and solutions 2, 3, 4, 5, and 6, the same as your observations for the reactions of solution 1 with solutions 2, 3, 4, 5, and 6? **If not, your unknown is not solution 1.**

c. Now consider the possibility that your unknown might be Solution 2. Did your unknown react with solution 2? If so, your unknown is not solution 2. Now look at your observations for any reaction that occurred between solutions 2 and 1 (found on the **top row of your table)** and solutions 2 and 3, 2 and 4, 2 and 5, and 2 and 6 (**2nd row of the observation table).** Were your observations for the reactions between your unknown and solutions 1, 3, 4, 5, and 6, the same as your observations for the reactions of solution 2 with solutions 1, 3, 4, 5, and 6? **If not, your unknown is not solution 2.**

d. Continue this procedure, showing your work, for each of the six possibilities. By the time you have finished you should have eliminated all of the possible choices except the correct one.

REPORT
OBSERVATION AND CRITICAL THINKING

NAME ___________________________

SECTION _________________________

Observations Table

To each tube in the row → add the solution below	Exp. tube 1 which contains Solution 1	Exp. tube 2 which contains Solution 2	Exp. tube 3 which contains Solution 3	Exp. tube 4 which contains Solution 4	Exp. tube 5 which contains Solution 5	Exp tube 6 which contains Solution 6
Solution 1						
Solution 2						
Solution 3						
Solution 4						
Solution 5						
Solution 6						
Unknown						

Unknown Identification Letter ___________ Identity of Unknown __________________

SHOW YOUR WORK on separate sheets of paper. Your work should show your comparison of your observations of the test results for your unknown with the test results for solutions 1-6. For the 5 solutions that your decided were **not** your unknown, explain **for each one** which observations caused you to eliminate it as a possibility. **If you do not show this work you will lose points for this lab.**

DETERMINATION OF CHEMICAL FORMULAS

INTRODUCTION

A chemical formula can be interpreted on two levels. On an individual (microscopic) basis, a chemical formula indicates the number of atoms of each element present in one molecule or formula unit of a substance. The subscripts of the formula can represent the numbers of atoms of the various elements present in one unit of the substance. At a macroscopic level, the subscripts in the chemical formula represent the numbers of moles of atoms of the different elements present in one mole of the substance. In an *empirical* formula, the subscripts represent the smallest whole number ratio of the atoms present in a substance. For most ionic compounds, their chemical formula is an empirical formula.

PROCEDURE

1. Clean an evaporating dish. Place the evaporating dish on a wire screen suspended on a ring approximately 5 cm above a Bunsen burner flame. Heat the dish for at least 5 minutes.

2. Place the evaporating dish in a desiccator and allow the evaporating dish to cool to room temperature.

DO <u>NOT</u> PLACE THE DESICCATOR'S LID ON THE BENCHTOP. THE GREASE WILL BECOME CONTAMINATED AND NO LONGER SEAL THE DESICCATOR PROPERLY.

3. Record the mass of the <u>cooled</u> evaporating dish (to the nearest 0.001 g). (Note that warm objects <u>cannot</u> be weighed accurately due to the convection currents of the atmospheric gases that are established around warm objects.)

4. Place 0.2 to 0.3 grams of manganese metal into the evaporating dish. Determine and record the mass of the evaporating dish and metal.

5. Place approximately 150 mL of water into a 400 mL beaker. Place the beaker on a hot plate in the fume hood and turn the hot plate to "High." Later, when the water begins to boil vigorously, turn the heat down. The water should be kept at a low boil throughout the experiment.

6. Place the evaporating dish containing the manganese on top of the beaker containing the boiling water. **CAREFULLY** add approximately 20 drops of 6 M hydrochloric acid to the manganese metal letting the acid run down the inside surface of the evaporating dish. Gently agitate the evaporating dish to mix the metal and the hydrochloric acid. Allow the reaction to proceed while frequently agitating the dish. When the reaction appears to be slowing, add 2 to 3 more drops of acid. Continue adding acid and agitating the dish until the Manganese is completely dissolved.

7. When the manganese has completely dissolved, stop adding acid and allow the solution in the dish to evaporate completely.

8. When all of the liquid in the dish has evaporated, the salt that remains should be a light pink. There should be no brown spots. When you are <u>**sure**</u> there is no liquid remaining, use beaker tongs to transfer the evaporating dish to the ring and wire screen over a Bunsen burner.

9. Heat the evaporating dish <u>**gently**</u>. If you see any signs of bubbling, immediately remove the Bunsen burner. Continue to apply and remove heat until there is no bubbling.

10. Heat over a very low flame for an additional 10 minutes. DO NOT GET THE DISH TOO HOT. Watch the manganese chloride product. It should stay light pink. If it starts to turn brown, it is getting <u>too</u> <u>hot</u>.

11. Immediately place the evaporating dish in the desiccator to cool to room temperature.

12. Determine the mass of the evaporating dish and manganese chloride product.

13. Calculate the mass of manganese used, the mass of manganese chloride product, and the mass of chlorine that reacted with the manganese.

14. From the mass of manganese and the mass of chlorine, determine the percent by mass of each element in the product.

15. Each group will write their mass percentages on the board. Determine the average mass percent for each element using the class data.

16. Use the average of the mass percentages for each element to determine the empirical formula of the manganese chloride product.

REPORT
DETERMINATION OF CHEMICAL FORMULAS

NAME ______________________

SECTION ____________________

Group data

Mass of manganese and evaporating dish, (g)	
Mass of empty evaporating dish, (g)	
Mass of manganese, (g)	
Mass of manganese chloride product and evaporating dish, (g)	
Mass of empty evaporating dish, (g)	
Mass of manganese chloride product, (g)	
Mass of chlorine, (g)	

Class data

Percent by mass of manganese	Percent by mass of chlorine
Average % Mn	Average % Cl

Empirical Formula of the manganese chloride product: ______________________

SHOW CALCULATIONS (Use additional sheets)

QUESTIONS FOR DETERMINATION OF CHEMICAL FORMULAS

NAME____________________________

1. When 2.50 g of copper metal reacts with molecular oxygen, the copper oxide product of the reaction has a mass of 2.81 g. What is the empirical formula of the copper oxide product?

2. When 10.8 g of silver was reacted with sulfur, 12.4 grams of product was produced (there was only one product). What is the empirical formula of the product?

3. Lysine is a compound composed of carbon, hydrogen, nitrogen and oxygen. When 1.50 g of lysine is burned, 2.72 g of carbon dioxide, 1.29 g of water and 0.287 g of nitrogen gas are produced. What is the empirical formula of lysine? If the molar mass of lysine is 146.19 g/mol, what is the molecular formula?

CHEMICAL REACTIONS

INTRODUCTION

In chemical reactions, the atoms of the substances present at the start of the reaction (the reactants) are rearranged into different combinations to produce other substances (the products of the reaction). However, there is no change in the number of each type of atom (mass and atomic identity is conserved). That means that the total number of atoms of each element in the reactants is equal to the total number of atoms of that element present in the products. This principle is used to balance the chemical equation that represents a chemical reaction. As with chemical formulas, chemical reaction equations can be interpreted on a microscopic level in which the coefficients (the numbers in front of each substance) can represent the number of individual units of that substance present. Reaction equations can also be interpreted on a macroscopic level in which the coefficients in the equation represent the mole-to-mole relationships between the reactant and product substances.

PROCEDURE

A. Metals with Hydrochloric Acid. (Demonstration)

1. Your professor will set up three test tubes, each containing about 3 mL of 6 M HCl. A small piece of three different metals will be added separately to the tubes of acid. Copper, zinc, and magnesium will be used. Record the appearance of each piece of metal before it is placed in the acid.

2. As each piece of metal is placed in the reaction, carefully observe the metal to determine if bubbles of gas are being formed. If the metal does react with the acid, hydrogen gas and a soluble metal chloride compound are formed. For the reaction with magnesium, feel the tube as the reaction occurs. What do you observe in regard to heat generated?

3. Write a balanced chemical equation for any reaction that occurs (or write "no reaction" if there is none). **INCLUDE PHASE LABELS.**

NOTE: For all the experiments listed below, the amount of chemical used can be **estimated.**

B. Zinc and Copper(II) Sulfate

1. Place **approximately** 3 mL of 1 M copper(II) sulfate solution in each of two test tubes.

2. Place a small piece of zinc in one of the tubes. Keep the other tube as a reference. Observe and record the color of the copper(II) sulfate solution.

3. Place the two test tubes in your test tube rack. Stir the solution containing the zinc every 15-20 minutes and observe the tubes periodically for about an hour. Write a balanced chemical equation for the reaction that occurred. (Hint: it is a single replacement reaction). **INCLUDE PHASE LABELS.**

4. Discard the solutions down the drain with lots of water and place any solid remaining in a waste paper basket. **DO NOT PUT THE SOLID IN THE SINK.**

C. More Chemical Reactions

Note: In each of the following group of experiments, two substances will be mixed together. Observe each substance **before** mixing, and observe again after mixing. Look for changes in color, formation of precipitates (solids), the dissolving of solids, and/or the formation of a gas. For each experiment, write a balanced chemical reaction for the reaction that occurs (Hint: these are double replacement reactions). **INCLUDE PHASE LABELS.**

1. Clean 12 of your medium sized test tubes.

Use one of your test tubes and add 20 drops of water. This volume is approximately 1 mL. Observe the level of the water in the tube. Discard the water. For all of the following experiments you will be using approximately 1 mL of each solution. The exact volume IS NOT CRITICAL. Just pour about 1 mL of each solution required into separate labeled tubes.

2. Mix **about** 1 mL of 0.1 M calcium chloride with **about** 1 mL of 0.1 M sodium phosphate. Record observations and write a balanced chemical equation. The solid formed in this reaction is calcium phosphate. What compound remains in solution?

3. Mix about 1 mL of 0.1 M barium chloride with about 1 mL of 0.1 M sodium sulfate. Record your observations and write a balanced chemical equation. The solid formed in this reaction is barium sulfate. What compound remains in solution?

4. Mix about 1 mL of 0.1 M iron(III) chloride with about 1 mL of 0.1 M KSCN Record observations and write a balanced chemical equation. . The red product is $FeSCNCl_2$ which remains in solution. The second product is potassium chloride. (This is a modified double-replacement reaction).

5. Mix about 1 mL of 0.1 M silver nitrate with about 1 mL of 0.1 M sodium chloride. Record your observations and write a balanced chemical equation. The solid formed in this reaction is silver chloride. What compound remains in solution?

6. Mix about 1 mL of 0.1 M lead(II) nitrate with about 1 mL of 0.1 M potassium chromate. Record your observations and write a balanced chemical equation. The solid formed in this reaction is lead(II) chromate. What compound remains in solution?

7. Place about 1 mL of 6 M hydrochloric acid in a test tube. Obtain a small amount of solid sodium carbonate (about the size of a pea) on a scrap of paper. Also obtain a wood splint. Light your Bunsen burner and just **before** you are ready to combine the chemicals, light the wooden splint. Pour the sodium carbonate into the acid, and, keeping the splint **dry**, quickly insert the burning splint into the top of the test tube. What happens to the flame? Record your observations. Write a balanced chemical equation for the reaction that occurred between the sodium carbonate and hydrochloric acid. The products of this reaction are carbon dioxide, water, and a soluble ionic compound.

REPORT
CHEMICAL REACTIONS

NAME______________________________

SECTION __________________________

A. Metals with Hydrochloric Acid

Copper

Observation____

Balanced Chemical Equation

____

Zinc

Observation____

Balanced Chemical Equation

____

Magnesium

Observation____

Balanced Chemical Equation

____

What did you observe concerning the heat generated for this experiment?

B. Zinc and Copper(II) Sulfate

Observation:____

____

Balanced Chemical Equation:

____

REPORT FOR CHEMICAL REACTIONS (cont.) NAME ___________________________

C. More Chemical Reactions

Calcium chloride and sodium phosphate
Observations
Balanced chemical equation
Barium chloride and sodium sulfate
Observations
Balanced chemical equation
Iron(III) chloride and KSCN
Observations
Balanced chemical equation
Silver nitrate and sodium chloride
Observations
Balanced chemical equation
Lead(II) nitrate and potassium chromate
Observations
Balanced chemical equation
Hydrochloric acid and sodium carbonate
Observations
Balanced chemical equation

QUESTIONS FOR CHEMICAL REACTONS NAME ___________________________

Write balanced chemical equations for each of the following reactions:

1. The decomposition of diiodine pentoxide to form iodine and oxygen.

2. Silver nitrate reacting with potassium sulfate in a double replacement reaction.

3. The combination of lithium and nitrogen to form lithium nitride.

4. The decomposition of potassium carbonate to form potassium oxide and carbon dioxide.

5. A single replacement reaction in which sodium metal reacts with aluminum oxide and replaces the aluminum.

6. The combustion of C_7H_{14} to form carbon dioxide and water.

HEAT

INTRODUCTION

Kinetic energy is the energy of motion, and temperature is proportional to the kinetic energy of the particles in a substance. When heat energy is added to a substance, the temperature of the substance increases. Conversely, the temperature decreases when energy is removed. However, substances differ in the amount their temperatures change relative the amount of heat energy they have gained or lost. Change in temperature is related to change in energy by the equation:

$$Q = mC\Delta T$$

where Q is the heat energy gained or lost by the substance, C is the specific heat of the substance, m is the mass of the substance, and ΔT is the change in temperature for the substance. Specific heat is defined as the amount of heat energy needed to change the temperature of exactly 1 g of a substance by exactly 1°C.

Calorimetry is an experimental technique in which temperature changes are measured and related to changes in heat energy. The apparatus with which these studies are done is called a *calorimeter.* The calorimeter that will be used in this experiment consists of a Styrofoam® cup containing a known mass of water, a cardboard lid, and a thermometer. The mass of a piece of metal will be determined and the metal will be heated. The temperature of the heated metal will be measured and then the metal will be placed into the water of the calorimeter (the temperature of which has also been measured). When two substances such as water and metal are in contact, the heat energy from one is transferred to the other. The temperature drops for the substance that loses heat. The temperature increases for the substance that gains heat. Because the amount of heat that is transferred is the same for each,

$$-Q_{\text{substances losing energy}} = Q_{\text{substances gaining energy}}$$

Styrofoam® absorbs very little heat. In this experiment, we will assume that the amount of energy absorbed by the cup and the thermometer is small enough to be neglected and that all the energy lost by the metal is gained by the water in the calorimeter. Therefore,

$$-\left[C_{metal}m_{metal}\left(T_{\text{final metal}} - T_{\text{initial metal}}\right)\right] = C_{water}m_{water}\left(T_{\text{final water}} - T_{\text{initial water}}\right)$$

Using the specific heat for water, and the measured masses and temperature changes for the water and metal, the specific heat of the metal can be determined.

When measuring the heat transferred during a change of state, there is no temperature change. The energy added or removed goes into changing the state (creating or breaking attractive forces between the particles) instead of being used to change the temperature. The heat transferred is still proportional to the amount of substance present, but the constant of proportionality is known as the heat of vaporization (for boiling/ condensing) or the heat of fusion (for freezing/melting).

In this experiment, the heat of fusion of water will be experimentally determined in units of J/g. Heats of fusion also may have units of J/mole. Dimensional analysis here shows that heat can be calculated by multiplying the mass of the substance by the heat of fusion:

$$\text{heat}(J) = \text{mass}(g) \times \text{heat of fusion}\left(\tfrac{J}{g}\right)$$

PROCEDURE

A. Determination of the Specific Heat of a Metal

1. Obtain a Styrofoam® cup, and lid. Obtain a cylinder and record the identity of the metal of your cylinder. **DO <u>NOT</u> REMOVE THE HOOK FROM THE CYLINDER!**
2. Measure and record the mass of the cylinder (this mass will include the mass of the hook).

Note: Because the metal of the hook will absorb and release heat energy during the experiment, its mass must be included in the total mass of the cylinder. However, because the metal used in the hook is probably different than the metal in the cylinder, a small unavoidable error will be present in the experimental determination of the specific heat of the metal in the cylinder.

3. Measure and record the mass of the empty Styrofoam® cup. Place about 50 mL of water in the cup. Measure and record the combined mass of the cup and water. Calculate the mass of water in the cup.

4. On a ring stand, place a ring and wire gauze. The ring should be about 5 cm above the height of your Bunsen burner. Fill your largest Erlenmeyer flask about 2/3 full of water and place it on the wire gauze. Clamp the neck of the flask to a ring stand. Light your Bunsen burner and start heating the water in the flask.

5. Use a second clamp to suspend your largest test tube (it should be clean and **dry**) in the water in the flask. The test tube should not touch the bottom of the flask.

6. Obtain a piece of monofilament line. Using a slipknot, tie the piece of line to the hook on the cylinder. **Gently** lower the cylinder into the test tube leaving the end of the line outside the tube. Continue to heat the water in the flask.

7. When the water starts to boil, measure the temperature of the boiling water (to the nearest 0.1 °C). Be sure the tip of the thermometer is in the water (but not touching the glass) when you read it. This temperature will be recorded as **the initial temperature of the metal.**

8. Measure the temperature of the water in the Styrofoam® calorimeter cup.

9. After the water in the flask has been boiling for at least 10 minutes, it can be assumed that the temperature of the metal cylinder (its initial temperature) is the same temperature as the boiling water. Using the monofilament line, rapidly remove the metal from the tube and gently place it into the water of the Styrofoam® calorimeter cup. Insert your thermometer through the cardboard lid and put the thermometer into the water of the cup. Stir the water gently with the thermometer. **DO NOT HIT THE METAL WITH THE THERMOMETER.** Approximately every 30 seconds, read the temperature of the water. Record the maximum temperature reached. This is the final temperature for **both** the water and metal.

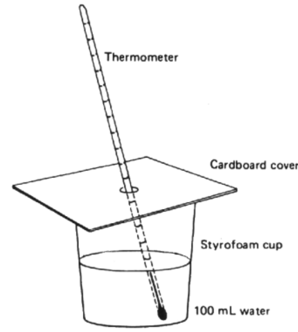
10. Thoroughly dry the piece of metal with a paper towel. Repeat steps 4 through 9 starting with a new 50 mL of water. Determine the combined mass of the water and Styrofoam® cup. You can use the empty mass of the cup and the mass of the cylinder from trial 1 but you must determine the new mass of water used in trial 2.

11. For **EACH** of your trials, calculate the specific heat of the metal in J/g °C (see Introduction). Assume that the initial temperature of the metal was the temperature of the boiling water. Use a specific heat for water of 4.184 J/g °C. Average the results from each of the trials to obtain the experimentally determined specific heat of your metal.

12. Dry your cylinder **AND RETURN IT TO YOUR INSTRUCTOR. EVERYONE IN THE GROUP IS RESPONSIBLE FOR RETURNING THE CYLINDER!!!** Untie as many knots as possible to the monofilament line and return the line to the cart.

B. Determining the heat of fusion of water

1. Obtain a calorimeter (two nested Styrofoam® cups and lid).
2. Measure and record the mass of the empty calorimeter (cups only, not the lid).
3. Add about 100 mL of water to the calorimeter. Determine and record the total mass of the water plus the calorimeter (cups only).
4. Measure and record the initial temperature of the water in the calorimeter. See picture at right.
5. Place 2 or 3 medium-sized ice cubes into the calorimeter and stir gently. When the temperature drops below 5 °C, remove any ice that has not melted.



6. Immediately measure and record the temperature of the water in the calorimeter.
7. Measure and record the mass of the calorimeter (cups only) and its contents. The increase in mass is due to the ice that melted.
8. From the change in temperature of the original water, the initial mass of water present, the mass of ice melted, and the change in temperature of the water from the melted ice, calculate the heat of fusion of water. Remember the heat gained by the ice melting and that water warming is equal to the heat lost by the original water. Also recall that the final temperature of the ice water and the final temperature of the original water will be the same. Assume that the initial temperature of the water obtained from the melting of the ice was 0.0 °C.

$$\left(\Delta H_{fus}\right)\left(m_{ice}\right)+C\left(m_{ice\ \text{water}}\right)\left(T_{final,ice\ \text{water}}-T_{inital,ice\ \text{water}}\right)=-C\left(m_{orig\ water}\right)\left(T_{final,orig\ water}-T_{initial,orig\ water}\right)$$

In the above equation, T is temperature, ΔH_{fus} is the heat of fusion of water, *m* is mass, and *C* is the specific heat of liquid water (4.18J/g °C).

9. Use your textbook to look up the accepted value for the heat of fusion of water and calculate your percent error for this experiment.

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

REPORT
HEAT

NAME ___________________________

SECTION ________________________

A. Determination of the Specific Heat of a Metal

Identity of metal		
Mass of metal, (g)		
	Trial 1	Trial 2
Mass of Styrofoam® cup and water, (g)		
Mass of Styrofoam® cup, (g)		
Mass of water, (g)		
Initial temperature of metal, (°C) (Temperature of the boiling water)		
Initial temperature of water in cup, (°C)		
Final temperature of water and metal (°C)		
Specific heat of metal, (J/g°C)		
Average specific heat of metal, (J/g°C)		

SHOW CALCULATIONS: (Use extra sheets of paper if needed)

REPORT FOR HEAT (cont.) NAME ____________________________

B. Determining the heat of fusion of water

Mass of empty calorimeter, (g)	
Mass of calorimeter + original water, (g)	
Mass of calorimeter + original water + melted ice water, (g)	
Mass of original water, (g)	
Mass of melted ice water, (g)	
Initial temperature of original water, (°C)	
Final temperature of original water + melted ice water, (°C)	
Experimental heat of fusion of water, (J/g)	
Accepted heat of fusion of water, (J/g) (from textbook)	
Percent error, (%)	

SHOW CALCULATIONS: (Use additional sheets if necessary)

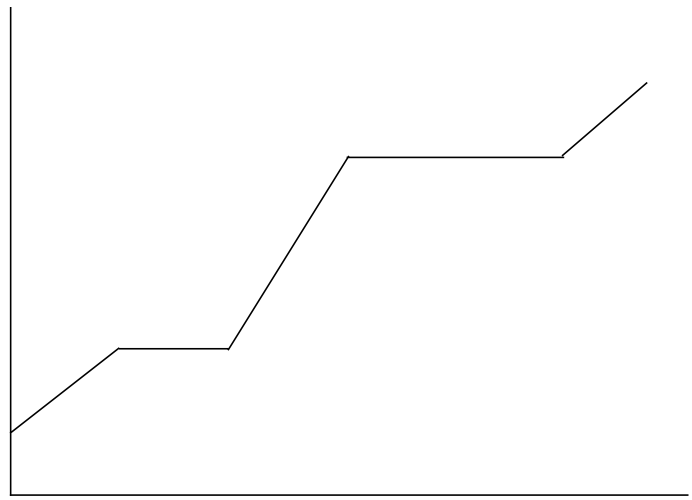
QUESTIONS FOR HEAT EXP. NAME____________________________

1. The specific heat of liquid ethyl alcohol is 2.42 J/g °C and its density is 0.7893 g/mL. A piece of solid silver (solid silver has a specific heat of 0.24 J/g °C) was heated to 95 °C and placed in 25.6 mL of ethyl alcohol that had a temperature of 19.27 °C. The final temperature of the alcohol and Silver was 23.5 °C. How many grams of Silver were used?

2. A 10.35 g piece of copper metal (solid copper has a specific heat of 0.382 J/g °C) was heated to 85.5 °C and then placed into 23.6 grams of water (liquid water has a specific heat of 4.18J/g °C). The initial temperature of the water was 18.3 °C. What was the final temperature of for the water and the copper?

QUESTIONS FOR HEAT EXPERIMENT (cont.) NAME ___________________________

3. Label the heating curve to show which areas on the graph would represent solid, liquid, and gas phases. Label the axes. Identify the phase changes and label the melting and boiling points.



4. How many joules of heat are lost when 150.0 g of steam are cooled from 124 °C to 86 °C? (Use your textbook to obtain any needed constants.)

5. What is the final temperature of the water if 35.1 kJ of energy are added to 100.0 g of ice at -4.5 °C?

CHARLES' LAW

INTRODUCTION

The temperature and volume of a gas are related. If the pressure of a gas and the number of moles of that gas remains constant, the volume will decrease with a decrease in temperature. If the temperature increases, the volume increases. This is the essence of Charles' Law, which is stated mathematically as:

$$V \propto T_K$$

or

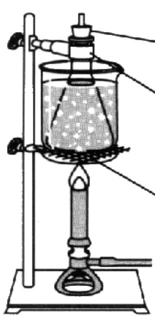
$$\frac{V}{T_K} = C$$

where V is the volume of the gas, T_K is the Kelvin temperature, and C is a constant.

In this experiment you will measure the volume of a gas at various temperatures. The pressure and number of moles of gas will be kept constant. By graphing the temperature of the gas versus its volume, you will be able to determine an estimate of the value in °C for absolute zero on the Kelvin scale.

PROCEDURE:

1. For this lab you will need a **DRY** 125 mL Erlenmeyer flask. DO **NOT** wash the flask (you can not get it dry enough to do the experiment). Look at the flasks that your group members have and choose the cleanest. Take it to the stockroom to be fitted with a one-hole stopper with glass tubing attached.

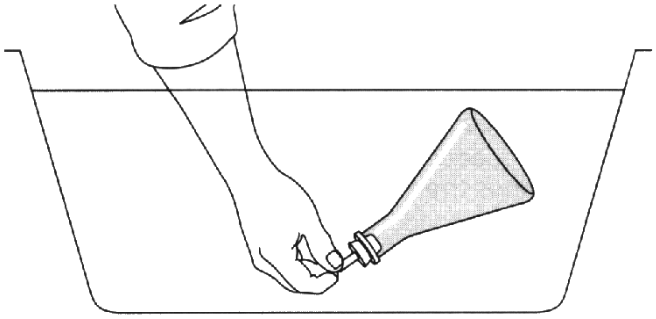


2. Set up a hot water bath as in the figure on the left. Place a 400 mL beaker on your wire gauze on a ring stand. Insert the stopper firmly into the **empty DRY** 125 mL Erlenmeyer flask. Use a clamp to hang the flask in the beaker. The bottom of the flask should be about 0.5 cm above the bottom of the beaker. Fill the **BEAKER** with water up to the neck of the flask. **Make sure you leave enough room for the water to boil without boiling over.**

3. Bring the water to a boil and gently boil for a **full 10** minutes. After the flask has been in the boiling water bath for at least 10 minutes you can assume that the temperature of the air in the flask in that of the boiling water. Measure and record the temperature of the boiling water.

Note: Several water baths at different temperatures have been set up in the hoods. Your professor will assign your group to a certain water bath. The class will combine their data from all the water baths to complete the report for the experiment.

4. Turn off your burner. Place your finger over the hole in the stopper and, using the clamp as a handle, transfer the flask to the water bath **that was assigned to your group**. Keep the stoppered end of the flask pointed down in the bath (see the figure below) and remove your finger from the hole after submerging the flask into the bath. (Do not allow any air to escape from the flask while it is in the bath.)

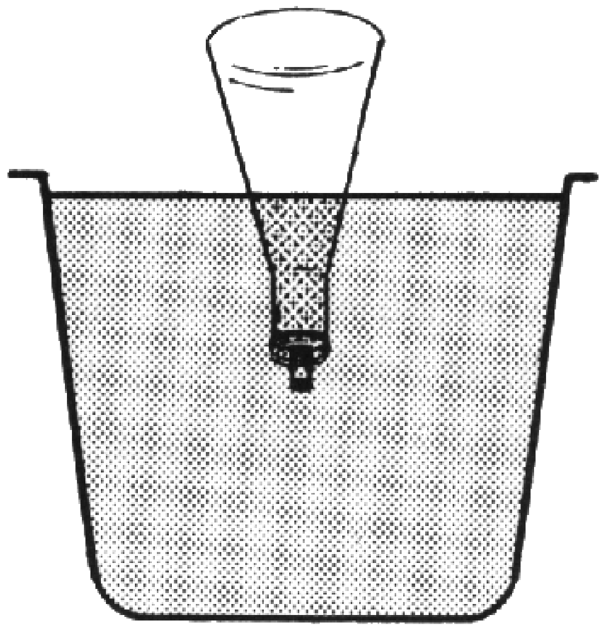


5. After the flask has been in the pan for **a full 10 minutes**, measure and record the temperature of the water in the pan.

6. Keeping the flask inverted, adjust the level of the flask so that the level of water in the flask is the same as the level of the water in the pan (see the figure below). This is the point at which the pressure in the flask is equal to the pressure outside the flask.

7. Place your finger over the hole in the stopper and remove the flask from the pan.

8. Use a graduated cylinder to measure the volume of water which entered the flask. Record this volume.



9. Determine the **<u>total</u>** volume of air space in the Erlenmeyer flask and stopper. To do this, fill the flask to the top with water. Insert the stopper <u>firmly</u> allowing water to go up into the glass tube. If the glass tube does not completely fill with water, use a dropper to fill it to the top. **Make sure there are no trapped air bubbles.** Remove the stopper allowing the water in the glass tube to flow back down into the flask. Use your graduated cylinder to measure all the water in the flask. You will need to fill your cylinder several times and add the volumes. Record the total volume. This volume is equal to the total volume of air present in the flask and tube when no water is present (the volume of air present in the boiling water bath).

10. On the chalkboard, in the first data table prepared by your professor, record your boiling water bath temperature and your boiling water bath gas volume (total flask and stopper volume). On the second table, record your cooler water bath temperature and the volume of water present in the flask after it was in the cooler water bath. Transfer the data for the entire class (including your data) into the data tables of your report.

11. Determine an **<u>average</u>** boiling water bath temperature using the data from all the groups. Determine an **<u>average</u>** total gas volume using the data from all the groups.

12. At each cooler water bath temperature, calculate the volume of gas in each flask. This will be done by subtracting the water left in this flask from the **<u>average</u> total gas volume** determined in step 11. (Using the <u>average</u> volume of all the flasks in this calculation reduces the error introduced by using several flasks, each of which has a slightly different volume. The residual error is compensated for by the fact the data will be graphed and a best-fit straight line determined.)

13. Using the graph paper provided at the back of your manual, draw a graph that will be used to represent the volume-temperature relationship of a gas. Turn your paper **sideways** and use the longest direction for your x-axis where you plot temperature. Your temperature scale (the horizontal axis) should run from -400 °C to 100 °C. Your volume scale (the vertical axis) should run from 0 mL to 150 mL. Review your Learning Laboratory exercise on Graphing Techniques.

14. Plot the experimental data (volume of air at each temperature) on your graph. One of the data points will be the **average volume and temperature** from the boiling water baths.

15. Draw a best-fit straight line through your data points and extend it all the way to the temperature axis. The point at which it crosses the temperature axis is your determination of absolute zero in °C.

16. Choose the air volume data for the coolest water bath temperature. Using the **average** temperature and air volume from the boiling water baths, use gas laws to calculate a **theoretical** volume at the cooler temperature. How well does this compare to the volume actually measured at that temperature? Calculate the % error.

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

REPORT FOR CHARLES' LAW EXP.

NAME ____________________________

SECTION__________________________

Boiling Water Bath

Volume of air in empty flask and stopper, (mL)	
Temperature of boiling water, (°C)	

Cooler Water Bath

Temperature of water bath, (°C)	
Volume of water in flask	

Boiling Water Bath Data

Temperature, °C	Empty Flask and Stopper Volume, mL
Average, (°C)	Average, (mL)

REPORT FOR CHARLES' LAW EXP. (cont.) NAME ___________________________

Cooler Water Bath Data

Temperature, °C	Volume of Water in Flask, mL	Volume of Air in Flask, mL (Subtract mL of water from **average** empty flask and stopper volume)

Estimate of absolute zero in °C (read from graph) ______________________________

Calculate the theoretical volume of air at the coolest water bath temperature.

% Error ____________________________

QUESTIONS FOR CHARLES' LAW EXP. NAME ___________________________

1. A balloon is filled with Helium on a warm summer's day. The balloon has a volume of 15.4 L on the day it is filled when the temperature is 89 °F. The following day a cold front moves through and the temperature drops to 68 °F. What is the volume of the balloon after the cold front moves in? Assume that the pressure is the same on both days and that no Helium leaks out of the balloon.

2. A second balloon is filled with 13.8 L of Helium on a day it is cold (air temperature is 43 °F). The balloon can hold a maximum of 16.00 L. The day after it is filled a warm front moves through the area and the temperature increases to 105 °F. Will the balloon burst? Again assume that the pressure is the same on both days and no Helium is lost.

3. Calculate the temperature at which a piston in a cylinder will be at a volume of 1.4 L if the piston is initially at 1.6 L when the temperature is 51 °C.

GAY-LUSSAC'S LAWS

INTRODUCTION

The temperature and pressure of a gas are related. If the volume and the number of moles of gas remain constant, Gay-Lussac's Law states that the pressure of a gas is directly proportional to its Kelvin temperature. This law can be stated mathematically

$$P \propto T_K$$

or

$$\frac{P}{T_K} = \text{const}$$

where P is the pressure of the gas and T_K is the Kelvin temperature. In this experiment, the pressure of a gas will be measured at several different temperatures and the data from the experiment will be put into graph form.

Gases are often involved as reactants or products in a chemical reaction. It is usually easier to measure the volumes of the gases involved than it is to measure their masses. Gay-Lussac's Law of Combining volumes states that the <u>volumes</u> of the different gases involved in a reaction (when measured at the same temperature and pressure) are in the same ratio as the coefficients for these gases in the balanced chemical equation for the reaction. This concept is also related to Avogadro's Law that states that equal numbers of gas molecules occupy the same volume. It is important to remember that these volume relationships apply only to gases. The volumes of liquids and solids do not show this type of relationship.

Electrolysis is a process by which electrical current (electron flow) can be used to cause a non-spontaneous chemical reaction to occur. During electrolysis, electrons are removed from one reactant (at the anode) and are donated to another reactant (at the cathode). Electrolysis of water causes the water to break down into its elements, oxygen and hydrogen, both of which are diatomic gases at room temperature and normal pressures. Oxygen will be generated at the anode and hydrogen will be formed at the cathode. By identifying the gases and measuring their volumes, the mole-to-mole relationship of oxygen and hydrogen in water can be determined.

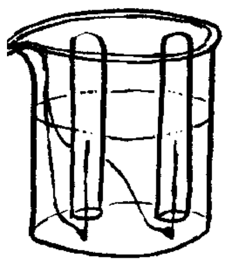
PROCEDURE

A. Electrolysis of Water

1. Each group will check out (from the stockroom) one set of electrodes and a 10 mL graduated cylinder.

2. Use a 400 mL beaker to obtain about 300 mL of 0.5 M sodium sulfate. (DO NOT DISCARD THIS SOLUTION AT THE END OF THE EXPERIMENT. IT WILL BE RECYCLED.) Obtain two of the small rubber bands.

3. Clean your smallest test tubes. Slip one rubber band on to each tube. Pour some of the sodium sulfate Solution into each of the cleaned test tubes. Fill the tubes to the very top. Put your finger tightly over the end of one of the tubes, invert it, and place the tube open end down in the beaker of sodium sulfate. Do not allow any air to enter the tube during this procedure. If you get air into the tube, remove it and try again. Repeat this procedure with the second tube. Wash the salt off your hands.



4. **BE VERY CAREFUL WITH THE ELECTRODES. THEY CAN BE EASILY DAMAGED. DO NOT TWIST OR BEND THEM.** If you do not think they are in the correct orientation, have your professor check them. Without bringing the open end of the tubes above the surface of the solution, insert an electrode into each tube (see diagram).

5. **HAVE YOUR PROFESSOR CHECK YOUR ASSEMBLY.** Plug in the transformer. **DO NOT TOUCH THE ELECTRODES WHILE THE TRANSFORMER IS PLUGGED IN.** After a few moments you should see bubbles forming on both metal electrodes. **Make sure all the bubbles are going into the tubes.** Allow the electrolysis to proceed for about 45 minutes (or for the length of time indicated by your professor). During this time, observe the experiment and record data for Part B: Relationship Between Pressure and Temperature.

6. After the electrolysis has run the allotted time, **UNPLUG THE TRANSFORMER. DO NOT REMOVE THE ELECTRODES FROM THE TUBES UNITL THE TRANSFORMER IS UNPLUGGED**. Without allowing the bottom of the tube to come above the surface of the solution, and without allowing any gas to escape, raise the tubes off the electrodes and remove the electrodes.

7. Leaving the tubes in the solution, raise or lower the tubes until the level of the solution inside each tube is at the same level as the solution outside the tube (you may need to add a little more sodium sulfate solution to your beaker to be able to do this). Keeping the levels of the solution the same on the inside and outside of the tube, roll the rubber band to the level of the solution. Allow the tubes to remain in the beaker.

8. The next step requires two people working together. One person will light a wood splint (light a Bunsen burner and use it to light your splints). A second person will put their finger over the end of the tube **THAT APPEARS TO CONTAIN THE LARGEST AMOUNT OF GAS** (try not to lose any gas and **<u>do</u> <u>not</u> disturb the location of the rubber band**). Keep your finger over the end of the tube to prevent the loss of any gas and turn the tube right side up (it will still contain some solution). As the person with the tube removes their finger, the person with the burning splint will **QUICKLY** insert the splint into the open end of the tube. Do not insert the splint into the solution or allow it to get wet. Record your observations. Label this tube as "A."

9. Remove the second tube (the one **THAT APPEARS TO CONTAIN THE SMALLEST AMOUNT OF GAS**) in the same manner, keeping a finger over the end to prevent the loss of the gas. In this case, light a splint and then put out the flame LEAVING A RED GLOW. Without letting the splint get wet, insert the end of the glowing splint into the end of the tube and record your observations. Label this tube as "B."

Note: When a burning splint is placed into carbon dioxide gas, the flame goes out. When a burning splint is placed into hydrogen gas, a "pop" is heard as the hydrogen burns. A glowing splint (no flame) placed into oxygen flares up into a larger flame. Use this information to determine the probable identity of the gas in each of the tubes.

10. Pour any solution remaining in the tubes back into the beaker. Wash the salt from your hands.

11. Determine the volumes of gas that were in tubes "A" and "B." Put water in the test tube to the level that was occupied by the gas (marked by the rubber band). Pour the water into the 10 mL graduated cylinder and record the volume of the water (watch your sig. figs.). The volume of the water is the volume of gas that was in each tube.

12. Divide the volume for the hydrogen by the volume for the oxygen. What was the ratio?

13. Use the correct formula for water to write a balanced chemical equation for the electrolysis of water to form hydrogen and oxygen gas.

14. Return the Sodium Sulfate solution to the original container (this will be one of the rare times that you are allowed to return chemicals to the original container). Rinse and dry the rubber bands, and return them to their original container. Rinse your metal electrodes 2-3 times with deionized water. Return everything you checked out to the stockroom.

B. Relationship between Pressure and Temperature

1. While you are waiting for the electrolysis experiment, watch your professor's demonstration and record the data.

2. You will be recording pressures in pounds per square inch (psi) and temperatures in Celsius.

3. Convert your pressures to torr (1psi = 51.7 torr) and your temperatures to Kelvin.

4. Use graph paper from the end of the lab manual and plot pressure in torr on the vertical axis and temperature in Kelvin on the horizontal axis. Is the relationship linear?

REPORT FOR GAY-LUSSAC'S LAWS

NAME ______________________________

SECTION____________________________

A. Electrolysis of Water

Volume of Gas in Tube A, (mL)	
Wood Splint Observations:	
Probable Identity of Gas in Tube A	
Volume of Gas in Tube B, (mL)	
Wood Splint Observations:	
Probable Identity of Gas in Tube B	
Hydrogen/Oxygen Ratio (use correct sig. figs.)	
Balanced Equation for the Electrolysis of Water:	

REPORT FOR GAY-LUSSAC'S LAWS (cont.) NAME ___________________________

B. Relationship between Gas Pressure and Temperature

Pressure, psi	Pressure, torr	Temperature, °C	Temperature, K

Graph Pressure (in torr) on the vertical axis and temperature (in Kelvin) on the horizontal axis.
Is the relationship linear?

Show Calculations:

PROBLEMS FOR GAY-LUSSAC'S LAW NAME ___________________________

1. A balloon is inflated to a pressure of 2.55 atm at a temperature of 25 °C. What temperature (**in °C**) is required to maintain the same volume if the pressure decreases to 1.39 atm?

2. The decomposition of ammonia gas (a process that occurs at high temperatures) is

$$2\ NH_3(g) \rightarrow N_2(g) + 3\ H_2(g)$$

If a balloon was filled with 3.5 L of ammonia gas and all of the gas decomposed, what would be the total volume of gas in the container? Assume the pressure and temperature remains constant.

3. What volume of dinitrogen pentoxide gas can be synthesized from 2.54 L of nitrogen gas and 1.83 L of oxygen gas? (Hint: it is a limiting reactant problem.)

SOLUTIONS

INTRODUCTION

A solution is a homogeneous mixture of two or more substances. The substance that is present in the larger molar amount is known as the solvent. The substance that is present in the smaller molar amount is known as the solute. The solute is dissolved into the solvent. Most solutions with which we are familiar are liquids. However, solids, liquids or gases may be either solute or solvent. For example, air is a solution of oxygen and other gases in nitrogen. Oxygen is a solute and nitrogen is the solvent.

The solubility of a substance depends on:

1. The nature of the solute and solvent
2. The amount of solute and solvent
3. The temperature
4. The pressure (for gaseous solutions).

A solution will form between the solute and solvent when the attractive forces between the solute and solvent are large enough to overcome the attractive forces within the solute and within the solvent. For aqueous solutions, temperature plays a role in the amount of solute that will dissolve. Usually, the higher the temperature, the greater the quantity of solute that will dissolve, but this is not always true. Some substances become <u>less</u> soluble in water as temperature increases. In this experiment, you will explore the effect of temperature on solubility by collecting data on the amount of solute (Potassium Nitrate) that will dissolve in a given amount of solvent (water) at a certain temperature. This data combined with the data from the rest of the class will allow you to create a temperature-solubility curve for Potassium Nitrate in water.

The concentration of a solution is calculated from the amount of solute present in a certain amount of solution. The concentration may be expressed using different units for amount of solute and solution. A **mass/mass percent** concentration expresses the grams of solute the grams of solution. The **mass/volume percent** concentration of a solution states the grams of solute present in the milliliters of the solution. The **molarity** of the solution describes the moles of solute per liter of solution.

PROCEDURE

A. Concentrations of a Sodium Chloride Solution

1. Each group should check out a 10 mL graduated cylinder.

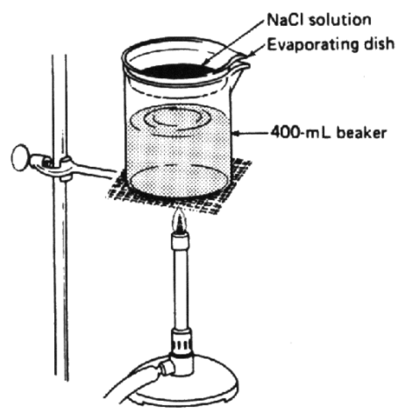
2. Measure and record the mass of a clean dry evaporating dish.

3. Obtain about 6 mL of the unknown sodium chloride solution in the 10 mL graduated cylinder. Read and record the exact volume (watch sig. figs.) of the solution in the cylinder.

4. Pour about 5 mL of the solution into the weighed evaporating dish. Set the cylinder back on the desktop for a few moments, then read and record the volume of solution remaining in the cylinder. Subtract the remaining volume of solution from the original volume to obtain the volume of solution transferred to the dish.

5. Determine the mass of the evaporating dish and solution combined. Calculate the mass of the solution in the dish.

6. Fill a 400 mL beaker about one-half full of water. Set it on an iron ring and wire screen above a Bunsen burner. Place the evaporating dish containing the weighed solution on top of the beaker. Heat the water in the beaker to boiling and keep it boiling as the water in the solution evaporates. Occasionally stir the salt crystals that form as the solution evaporates. Make sure that any salt crystals that adhere to the stir rod are scraped off into the evaporating dish. Replenish the water in the **BEAKER** if needed. Start part B of this experiment while you are evaporating the salt solution.



7. When the sodium chloride salt remaining in the evaporating dish appears completely dry, CAREFULLY remove the evaporating dish and dry the bottom of the dish with a paper towel. Remove the beaker of water and place the dish directly on the wire screen. Gently heat the dish with a low Bunsen burner flame for about 5 minutes.

8. Set the dish on a ceramic tile and allow it to cool for **at least** 10 minutes.

9. Determine the mass of the evaporating dish and the dried sodium chloride Salt.

10. Calculate the % (m/m), % (m/v), molarity, and density of the sodium chloride solution.

B. Preparation of a solubility curve for Potassium Nitrate.

1. Obtain from your professor a sample of potassium nitrate. Each group will be given a different amount of potassium nitrate (3-7 grams).

2. Weigh and record the mass of the potassium nitrate and paper. Place the potassium nitrate into a large test tube. Weigh and record the mass of the paper. Calculate the mass of potassium nitrate used.

3. Use the 10 mL graduated cylinder to obtain **about** 6 mL of deionized water. Read and record the exact volume (watch sig. figs.) of water in the cylinder. Pour **about** 5 mL of the water into the tube containing the Potassium Nitrate. Set the cylinder back on the desktop for a few moments, then read and record the volume of water remaining in the cylinder. Subtract

the remaining volume of water from the original volume to obtain the volume of water added to the potassium nitrate. The KNO_3 should not dissolve completely at room temperature in the amount of water being used.

4. Clamp the test tube to a ring stand and place the test tube into a beaker of water. Use a Bunsen burner to heat the water in the beaker.

5. Stir the mixture in the test tube with a **stir rod** as you begin heating. Continue heating and stirring until all of the KNO_3 dissolves.

6. After all of the KNO_3 dissolves **raise the test tube out of the hot water** and remove the stir rod. Allow the test tube and contents to gradually cool as you **gently** stir the contents with a thermometer.

7. Record the temperature at which the **FIRST** crystals appear. This is the temperature at which the solution is saturated. The amount of KNO_3 represents the solubility at this temperature.

8. Repeat the warming and cooling of the solution **at least** 3 more times. If the temperatures do not agree within 3 °C, do more trials and discard the temperature measurements that do not agree. Calculate an average temperature using data from 4 trials that agree within 3 °C. This is the temperature at which your solution is saturated and at which your volume of water contains the mass of KNO_3 used in your experiment.

9. When you are finished, discard the KNO_3 solution in the sink.

10. Solubility is expressed in grams of solute per 100 mL of water. You now need to convert the solubility obtained earlier to these units.

$$? \text{ grams } KNO_3 = 100.0 \text{ mL } H_2O \times \frac{\text{your mass of Potassium Nitrate}}{\text{your volume of water}}$$

11. Share your results (in g KNO_3/100 mL water) with the rest of the class. Collect the information on solubilities and temperatures of the other solutions.

12. Use the data collected to prepare a graph of the solubility curve for potassium nitrate. Plot the solubility (0.0 to 150.0 g KNO_3/100 mL water) on the vertical axis and the temperature (0.0 °C-100.0 °C) on the horizontal axis. Follow the rules for correctly labeling your graph. Draw a smooth curve (the curve may not pass through all the data points).

REPORT
SOLUTIONS

NAME ____________________________

SECTION __________________________

A. Concentrations of a Sodium Chloride Solution

Initial Volume of NaCl solution in cylinder, (mL)	
Volume of NaCl solution remaining in cylinder, (mL)	
Volume of NaCl solution transferred to dish, (mL)	
Mass of dish, (g)	
Mass of dish and NaCl solution, (g)	
Mass of NaCl solution, (g)	
Mass of dish and dried NaCl salt,(g)	
Mass of dried NaCl salt, (g)	
Percent NaCl (m/m), (%)	
Percent NaCl (m/v), (%)	
Molarity of NaCl, (moles/L)	
Density of solution, (g/mL)	

SHOW CALCULATIONS:

REPORT for SOLUTIONS (cont.) NAME ___________________________

B. Preparation of a solubility curve for Potassium Nitrate

Mass of paper and KNO_3, (g)	
Mass of paper, (g)	
Mass of KNO_3 dissolved, (g)	
Initial volume of water in cylinder, (mL)	
Volume of water remaining in cylinder, (mL)	
Volume of water transferred to tube, (mL)	
Solubility, (g KNO_3/ 100 mL water)	

Temperature of crystal formation, (°C)

trial 1	trial 2	trial 3
trial 4	trial 5	trial 6

Average Temperature, °C ______________________

Class data

Solubility of KNO_3, (g/100 mL)	Temperature, (°C)

QUESTIONS FOR SOLUTIONS NAME ____________________________

1. Why does the solubility of many substances increase with temperature? (Remember what an increase in temperature means on a microscopic scale.)

2. Based on your graph what is the solubility (in g/100mL H_2O) of KNO_3 at 51 °C?

   At 75 °C?

3. Based on your graph, at what temperature would the solubility of potassium nitrate be 95 g/100 mL H_2O?

4. A 3.0% (m/v) potassium chloride solution has a volume of 25.0 mL. How many moles of potassium chloride are in the sample?

5. How many grams of a 25% (m/m) sodium chloride solution contain 0.250 moles of sodium chloride?

QUESTIONS FOR SOLUTIONS (cont.) NAME ____________________________

6. A 15.0 mL sample of a potassium nitrate solution that has a mass of 15.78 g is placed in an evaporating dish and evaporated to dryness. The remaining salt has a mass of 3.26 g. Calculate the following concentrations for the potassium nitrate solution.

   (a) percent (m/m)

   (b) percent (m/v)

   (c) molarity

7. Hydrochloric acid, HCl, reacts with barium hydroxide to produce barium chloride and water. How many mL of a 3.00 M hydrochloric acid solution would be required to react with 25.5 mL of a 4.65 M barium hydroxide solution?

MATH AND THE CALCULATOR

INTRODUCTION

Numbers used in science are usually measurements of some sort. Some examples are the mass of a chemical used in a reaction, the speed of a spacecraft, the distance to some astronomical object, or the mass of a new sub-atomic particle. The number of digits in the measurement is dictated by the instrument used to take the measurement. A meter stick with no markings on it other than the zero meter mark and the one-meter mark should not be used to take an accurate measurement of the thickness of a dime. The number of digits used in a measurement is called the significant figures of that measurement. The significant figures include all of the digits about which we are sure, the digits read directly from the measuring device, and one additional digit (the last one in the number), which is estimated.

Other numbers used in science are exact. They have an infinite number of significant figures. Counted numbers are exact. You do not have 0.995643 cars. You have exactly 1 car. Or maybe you have exactly 0 cars. Conversions within a unit system are also exact. They are defined numbers. There are **exactly** 100 centimeters in 1 meter. There are **exactly** 12 inches in a foot.

Some numbers that are used in chemistry are either too small or too large to write easily. For example, one molecule of water has a mass of 0.00000000000000000000002992 g. This number is rather tedious to write and if it needs to be written many times, there is a large chance for a "typo" error. This number can be expressed more easily in scientific notation. Scientific notation is a way of writing numbers using powers of ten. Negative powers of ten indicate number that are less than 1. Numbers larger than 1 use positive powers of ten. The number above can be re-written as 2.992×10^{-23}. The negative sign in the exponent of the ten tells us that the decimal point in the standard form of this number lies 23 places to the left of where it is now. The nearest star to the sun is 40700000000000 km away. This can be re-written as 4.07×10^{13} km. The positive exponent here tells us that the decimal place in the standard form of the number is located 13 places to the right.

When entering numbers into your calculator in scientific notation, you should follow the rules shown below. As an example, input the number 4.7756×10^{-14} into your calculator.

1. Enter the **coefficient** (the part of the number before the multiplication sign).
2. Press the key marked **EE** or **EXP** (depending on what kind of calculator you have).
3. Enter the power of ten with the appropriate sign. (To change the sign of the power of ten, press the key marked **+/-**.)

DO NOT enter the "×10" into the calculator. The **EE** or **EXP** replaces the "×10" part of the number.

The rules for calculating with significant figures and scientific notation are given in your textbook. For complete instructions on using your calculator, **CONSULT THE MANUAL THAT CAME WITH YOUR CALCULATOR.**

EXERCISES

A. Indicate whether each of the following is a measured or an exact quantity.

5 books ____________________ 12 roses ____________________

5 lb ____________________ 16 ounces in 1 pound ____________________

9.25 g ____________________ 361 miles ____________________

0.035 kg ____________________ 1000 m in 1 km ____________________

B. Write the following numbers in scientific notation.

4,450,000 (4 sig figs) ____________________ 0.00032 ____________________

38,000 (2 sig figs) ____________________ 25.2 ____________________

0.0000000021 ____________________ 0.0505 ____________________

Write the following as standard decimal numbers.

4.09×10^{2} ____________________ 3×10^{-4} ____________________

5.315×10^{1} ____________________ 8.2×10^{-3} ____________________

3.156×10^{3} ____________________ 2.46×10^{-6} ____________________

C. State the number of significant figures in each of the following measurements.

4.5 m ____________________ 204.25 g ____________________

0.0004 L ____________________ 6.25×10^{5} mm ____________________

805 lb ____________________ 34.80 km ____________________

2.50×10^{-3} L ____________________ 8×10^{5} g ____________________

D. Round off each of the following to the number of significant figures indicated. Put the number in scientific notation if needed to avoid writing an ambiguous number.

	Three significant figures	Two significant figures
0.4108		
532,800		
143.63212		
5.448×10^{2}		
0.00858345		

E. Perform the following multiplication and division calculations. Give a final answer with the correct number of significant figures.

4.5×0.28	
$0.1184 \times 8.00 \times 0.0345$	
$(2.5 \times 10^{4}) \times (5.0 \times 10^{-7})$	
$\frac{(42.4)(15.6)}{1.265 \times 10^{2}}$	
$\frac{(35.56)(1.45)}{(4.8 \times 10^{-1})(0.56)}$	

F. Perform the following addition and subtraction calculations. Give a final answer with the correct number of significant figures.

13.45 mL + 0.4552 mL	
145.5 m + 86.58 m + 1045 m	
$1315 + (2.0 \times 10^{2}) + (1.10 \times 10^{3})$	
245.625 g - 80.2 g	
4.62 cm - 0.855 cm	

G. Mixed operations.

(12.6 + 0.57)/3.1415	
$(13.54 \times 1.56) + (22.6 \div 3.5)$	
$33.5 \times \frac{(15.4 + 3.45)}{(44.65 - 23.6)}$	

H. Problems

1. A number that counts something is an exact number. When you say 4 mugs, 5 books, or 2 watches, you are using exact numbers. However, when you use a rule to measure the height of your friend as 155.2 cm, you obtain a measured number. Why are some numbers called exact while other numbers are called measured?

2. In what place (ones, tens, tenths, etc.) is the estimated digit in each of the following measurements.

   1.5 cm ____________ 4.50×10^3 mi ____________

   0.0782 m ____________ 42.50 g ____________

3. Bill and Beverly have measured the sides of a rectangle. Each recorded the length of the rectangle as 6.7 cm and the width as 3.9 cm. When Bill calculates the area (by multiplying the length by the width), he gives an answer of 26.13 cm^2. However, Beverly correctly records her answer for the area as 26 cm^2.

   (a) Why is there a difference between the two answers when both students used the same measurements?

   (b) You are going to tutor Bill. What would you tell him to help him correct his answer?

I. Activity

List at least eight numbers (and units) you might use throughout the day (such as measurements, prices, definitions, cooking quantities, gas purchases, prescription dosages, etc.). Identify each number as exact or measured and explain why it is exact or measured. Did you use a measuring tool? Did you count out something? Did you use a definition?

Number used	**Type of number** (exact or measured)	**Explanation**

Additional Study Problems

1. How many significant figures are in each of the following numbers?

20.03 0.0094 53.000 2.474×10^{-3}

2. Convert each of the following decimal numbers to scientific notation retaining the correct number of significant figures.

300.4 0.0059301 0.0000003

3. Convert each of the following numbers to standard decimal notation retaining the correct number of significant figures.

7.653×10^{2} 8.00×10^{-3} 3.006×10^{-1}

4. Do the following calculations and express your answer using the correct number of significant figures.

$1234.7 + 7 - 3.25 =$

$4.35 \times 10^{3} + 7.966 \times 10^{2} =$

$$\frac{11.45\times 2.3}{2.00\times 0.0345}=$$

$$\frac{2.330-1.2}{3.444}=$$

$(2.45 \times 10^{-3})(5.0 \times 10^{2}) =$

DIMENSIONAL ANALYSIS STUDY QUESTIONS

Use your textbook to up any English-metric conversions you need for these problems.
Some useful English to English conversions are:

1 quart = 2 pints (exact)
1 pint = 2 cups (exact)
1 peck = 8 quarts (exact)
1 gallon = 4 quarts (exact)

1. Suppose there are 3 (counted) peppers in a cup of peppers. If Peter Piper picked a peck of pickled peppers, how many pickled peppers did Peter Piper pick?

2. An alloy is 25.4% silver by mass and has a density of 12.9 g/cm^3.
   (a) What is the volume in m^3 of alloy do you need to have 10.0 grams of Silver?
   (b) What volume in cm^3 of alloy would you need to have 1.00 lb of Silver?

3. A cube has sides that are 2.633 inches long and a density of 7.85g/cm^3. How many pounds does the cube weigh?

4. You are going to build a fence that is 45 feet long. Each meter of fence uses 18 nails. At the hardware store, one pound of nails contains 36 nails, and they sell for 85 cents per kilogram. What will be the cost <u>in</u> <u>dollars</u> to purchase the nails for the fence?

5. A large diamond has a mass of 35.8 carats. If a carat is 200.0 mg, what is the mass of the diamond in mass ounces?

6. A bracelet that has a volume of 15.40 mL is made of an alloy that is 60.0% gold by mass and has a density of 14.3 g/mL. How many ounces (mass ounces) of gold are in the bracelet?

7. The fuel tank in a motorcycle has a capacity of 2.75 gallons. If the density of gasoline is 42.0 lbs/ft^3, what is the mass in grams of the gasoline in a full tank of gas?

8. A type of bronze is 80.0% by weight copper and 20.0% tin. The bronze has a density of 8.74 g/cm^3. A sculptor is preparing to cast a bronze figure that requires 0.350 L of bronze. How many mass ounces of copper must be used to prepare the bronze to use in the figure?

PHYSICAL PROPERTIES OF THE ELEMENTS

Observe the elements in the Periodic Table display and fill in the following table.

Element	Symbol	Atomic Number	Color	Luster	Metal or Nonmetal
aluminum					
carbon					
copper					
iron					
nitrogen					
magnesium					
manganese					
nickel					
oxygen					
phosphorus					
potassium					
silicon					
gold					
silver					
tin					
zinc					
cobalt					
lead					
chlorine					
osmium					
cesium					
antimony					

PHYSICAL PROPERTIES OF ELEMENTS AND COMPOUNDS

Observe the display of chemicals. Record the chemical formula given for each compound. Describe the physical properties of each compound including color, luster, and phase (solid, liquid or gas). Give the full name of each element in the compound. Observe the display of elements and record the physical properties of each element in the compound.

Formula of the compound	Physical properties of the compound	Physical Properties of each element in the compound
Example: **$CuSO_4 \cdot 5H_2O$**	**Deep blue crystalline solid**	**Copper—a yellow-brown, shiny solid** **Sulfur—a yellow, powdery solid** **Oxygen—a colorless gas** **Hydrogen—a colorless gas**

1. When elements combine to form compounds, the physical properties of the compound are the same as those of the elements. True or false? Explain your answer based on your observations.

2. In a compound, there is a variable composition of the elements. True or false? Explain your answer.

GRAPHS

INTRODUCTION

It is often necessary to examine relationships between experimental quantities. This is frequently done by organizing experimental measurements into data tables. However, it is often easier to see the relationship between quantities if the data are represented in the form of a graph. A number of rules must be followed when constructing graphs and this exercise will give the student practice in correct graphing techniques.

Construction of a graph

1. Select a good quality graph paper that is easy to use with the metric scale. Graph paper that has divisions marked in blocks with different shades of lines is easier to use (less counting) than paper that has uniform shading. Choose paper that is divided into 5 x 5 or 10 x 10 small squares within a larger grid. Avoid paper in which the large squares are divided into 4 x 4 or 8 x 8 blocks (this type of graph paper is for drafting classes that use English system units). A few sheets of an appropriate type of graph paper are available at the end of this lab manual.

2. Title each graph. The title should reflect what quantities are being plotted. The title might simply be an equation that has been provided or it might be the description of experimental quantities.

3. Label each axis with name of the quantity represented on that axis.

4. Use a scale for each axis that will spread the data points to be plotted over the full page (or over the space assigned). Do not crowd the data into one corner. However, your scale should result in convenient units (such as 10, 20, 30, etc. or 2, 4, 6, 8, etc.) for each major division on the graph. A compromise may be necessary.

5. Use a constant scale (the same number of divisions/unit) along each axis. Because different quantities are plotted on each axis, you would not expect the scale on the x and y axes necessarily to be the same. If you do not have any data close to a zero value, you need not place "zero" in the lower left-hand corner of the graph. The graph origin can begin at any convenient value (provided it is labeled).

6. It is only necessary to mark (and label) the intervals at 5 to 10 places along an axis (more than that gets cluttered). For example, if you had mass readings ranging from 7 to 68 g, you might mark and label the axis at 0, 20, 40, 60, and 80 g. Do NOT mark your axes at the data points.

7. The precision in the labels for the axes intervals should reflect the precision in the data being plotted. For example, if masses were determined to one place after the decimal (such as 9.1 g, 15.4 g, etc.) the intervals on the graph should be labeled 0.0, 20.0, 40.0 and so on.

NOTE: The precision for measurements plotted on the y-axis may differ from those for the x-axis.

Example

Suppose we had measured the distance and time that a bicycle rider traveled and that we had collected that data in the table shown below. These data also could be represented graphically. On graph paper, a vertical and a horizontal axis is drawn (See Fig. 1). The lines should be set in to leave a margin for numbers and labels, but the graph should cover most of the allotted space. A title should be chosen and placed at the top of the graph. Each axis must be labeled and a scale designed for each axis. The intervals on each axis must be **equally spaced**, and be in convenient divisions. In the sample graph shown below, the distance scale starts at 0 km and goes to 50 km (a distance that includes all the data points given), and every 10 km interval is marked. The time scale is divided into 1-hour intervals and covers the 10-hour time span of the data.

Time and Distance Measurements for a Bicycle Rider

Time(hr)	Distance Covered (km)
1	5
3	14
4	20
6	30
7	36
8	40
9	46
10	49

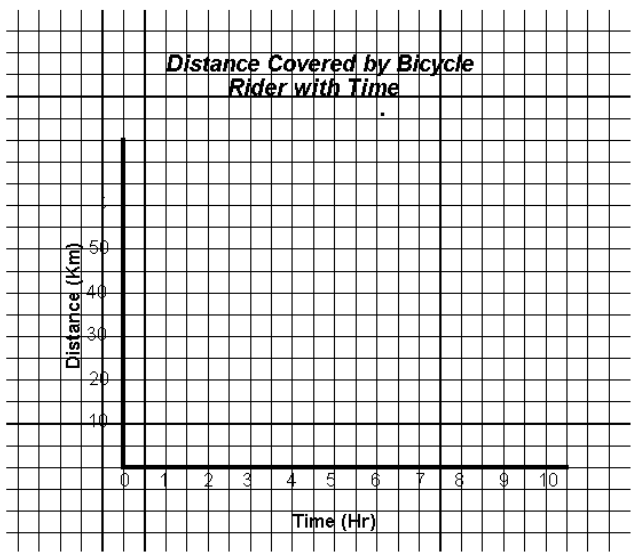


Figure 1

8. Now the data can be plotted. Each set of measurements makes a point on the graph. (See Fig. 2)

9. After the data have been plotted, draw either a straight line or a smooth curve that best represents the data points. Do NOT connect the dots with individual straight lines. Because of experimental error, you should not expect the line to pass directly through every data point. Construct a "best-fit" plot in which the points that do not fall on the line are randomly scattered. The sum of the distances between the line and the points above it should be about the same as the sum of the distances between the line and the points below it. In addition, the line should be drawn so that these distances are minimized.

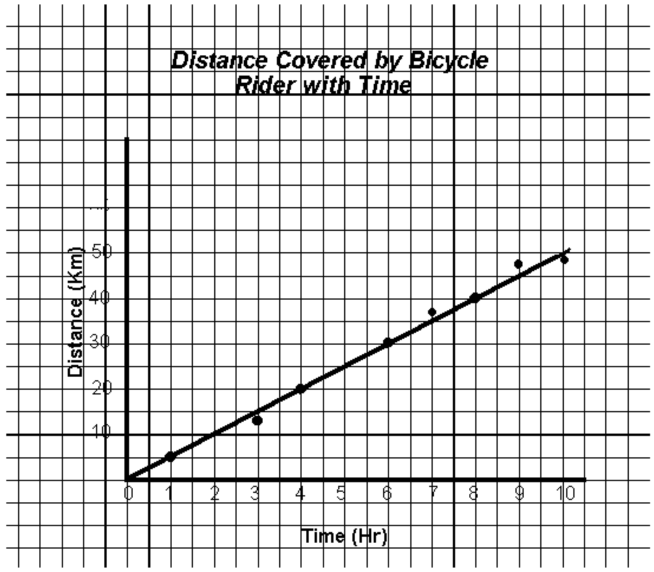

Distance Covered by Bicycle Rider with Time
Distance (Km)
50
40
30
20
10
0
1
2
3
4
5
6
7
8
9
10
Time (Hr)

Figure 2

GRAPHING EXERCISES

Watch your professor's demonstrations and record your data on a piece of your paper.

A. Relationship between the Mass and Volume of a Liquid

In this graphing activity, the relationship between mass and volume of a liquid will be shown. The volume and mass of five different samples will be measured. After the data are collected, the data will be plotted and the density of the samples will be visually represented on a graph.

1. Place a **dry**, clean 50 or 100 mL beaker on a top loading balance and *tare* it (your professor will explain what this means).

2. Obtain approximately 60 mL of water in another small beaker.

3. Using a 10 mL graduated cylinder, carefully measure out a 10.0 mL sample of water and transfer it to the tared beaker. Determine and record the mass of the water.

4. Add a second 10.0 mL sample of water to the first water in the beaker and determine the mass.

5. Continue to increase the total volume of the liquid by adding 10.0 mL increments (determining the combined mass after each addition) until there is a total of 50.0 mL of water in the beaker.

6. Use graph paper from the back of your lab manual and plot the data that has been collected. Put the mass on the vertical axis and the volume on the horizontal axis.

7. Draw a best-fit straight line through the data points. Determine the slope of the line (vertical change divided by horizontal change). This is the density of the water used in this experiment.

B. Relationship between Mass and Volume of a Solid

1. Fill a 50 mL graduated cylinder about one-half full of water. Record its exact volume.

2. Place the graduated cylinder on the top-loading balance and tare it.

3. Add two or three pennies to the graduated cylinder. Record the new volume of the graduated cylinder and the mass of the pennies. Subtract the original volume of the water from the volume of the water plus pennies to determine the volume of the pennies.

4. Add two or three more pennies to the graduated cylinder. Record the combined mass of the pennies. Determine the volume of the pennies plus water and subtract **original** volume of water to obtain the new volume of the pennies. Continue this procedure until you have collected five sets of data.

5. Construct a graph in which you plot the mass of the pennies on the vertical axis and the volume of the pennies on the horizontal axis. Determine the slope of the line (which is the density of the pennies).

SUBATOMIC PARTICLES AND ISOTOPES

Element	Complete Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
fluorine			19			
	$^{31}_{15}P$					
calcium		20			22	
			43	20		
iron			56			
iron					31	
bromine			80			
				44	56	
					57	44
gold					118	
			20			10

Naturally occurring neon has 3 isotopes: 90.48% is neon-20 (19.999 amu), 0.27% is neon-21 (20.993 amu) and the remainder is neon-22 (21.991 amu). Calculate the average atomic mass of neon. How does your calculated number compare with that from the periodic table?

ELECTRON CONFIGURATIONS AND ORBITAL DIAGRAMS

Write the complete electron configurations and orbital diagrams for:

Fe

P

Ne

Mg

Si

Co

Using the noble gas short form, write the electron configurations and orbital diagrams for:

Ir

Y

I

Os

DISTINGUISHING ELECTRON

Give the shell number (n) and subshell letter of the distinguishing electron for each of the following elements. Also, list the period in which each element is located.

Element	Shell and Subshell	Period
titanium		
uranium		
arsenic		
barium		
osmium		
chlorine		
rubidium		
tin		
silver		

CLASSIFICATION OF THE ELEMENTS

Give the names for each of the following elements

Noble gas in period 2	
Halogen in period 3	
First Transition metal in period 4	
Alkali metal in period 5	
Third Inner Transition metal in Period 7	
Halogen in Period 5	
Noble gas in Period 6	

ELECTRON DOT STRUCTURES
INTRODUCTION

Drawing Lewis Electron Dot structures helps visualize the chemical bonds that are present in ionic and molecular compounds, and in polyatomic ions. The formation of a chemical bond is the result of the tendency of atoms to gain or lose electrons until they have obtained an electron configuration that is the same as that of a noble gas. Except for Helium, which has 2 valence electrons, the noble gases all have 8 valence electrons (an octet). The valence electrons are the electrons in the outermost s and p orbitals. Except for Hydrogen (which would prefer to be associated with 2 valence electrons like Helium), the Representative elements (Group A elements) gain, lose, or share electrons in a manner that results in them being associated with 8 valence electrons. Review the rules in Chapter 7 in your textbook for the rules on drawing electron dot structures for atoms, ions, polyatomic ions, ionic compounds, and molecular compounds.

PROBLEMS

Before writing a Lewis dot structure for a compound, one must remember whether the bonding in the compound is ionic or covalent and write the type of structure appropriate for each.

Compound	Type of bonds (ionic or covalent)
LiBr	**Ionic**
$CaCl_2$	
CCl_4	
NH_3	
K_2S	
$FeCl_3$	
MgO	

Write electron **configurations** for the following ions.

1. I^{1-}
2. P^{3-}
3. Mg^{2+}
4. S^{2-}

Write the **orbital diagram** for the following ions.

1. K^+
2. O^{2-}
3. N^{3-}
4. Cl^{1-}

Draw electron dot structures for the following neutral atoms:

1. Ca
2. Se
3. Rb
4. Xe
5. Si
6. P

Draw electron dot structures for the following ions:

1. Br^{1-}
2. Al^{3+}
3. N^{3-}
4. O^{2-}

Draw electron dot structures for the following ionic compounds:

1. $BaCl_2$
2. NaI
3. K_3N
4. Mg_3P_2

Draw electron dot structures for the following molecules:

1. SiO_2

2. F_2CO (carbon is central)

3. PH_3

4. SO_2

Write electron dot structures for the following polyatomic ions.

1. CN^{1-}

2. SO_4^{2-}

3. NO_3^{1-}

Write an electron dot structure for the ionic compound, NH_4NO_2.

MOLECULAR MODELS

INTRODUCTION

The chemical and physical properties of a substance are influenced by the distribution of outer shell (valence) electrons and the three-dimensional arrangement of its nuclei. A variety of experimental methods are employed to map out the relative positions of the nuclei in the molecule or complex ion.

Most molecules and complex ions have a central atom. You will determine the distribution of electrons and bonded atoms about the central atom. In so doing, you will be able to determine the electron pair and molecular geometries, and polarity of the species in question.

Fill in the following table:

CBr_4	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?
N_2O (one of the nitrogens is central)	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?
CO	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?
SO_3^{2-}	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?

CCl_2F_2	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?
HCN	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?
SO_2	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?
NO_2^{1-}	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?

H_2O	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?
F_2CS	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?
NH_3	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?
ClO_2^{1-}	Electron Pair Geometry
	Molecular Geometry
	Polar or Nonpolar?

CHEMICAL NOMENCLATURE

INTRODUCTION

Chemical nomenclature is the system of names used to distinguish compounds from each other and the rules needed to devise these names. The most modern nomenclature system uses rules that follow the recommendations of the International Union of Pure and Applied Chemistry (IUPAC). For nomenclature purposes, it is necessary to determine the nomenclature category to which a compound belongs. In Chemistry 68/60, three nomenclature categories are studied – ionic compounds, binary molecular compounds, and acids. To master chemical nomenclature, it is necessary to memorize the names, formulas and charges of polyatomic ions (there are hundreds of polyatomic ions – you are responsible for the ones assigned by your professor). It is also necessary to memorize the charges on the fixed charge ions. Review the rules on chemical nomenclature in Chapter 8 of your textbook.

EXERCISES

Name the following acids (all compounds are in aqueous solution):

H_2CO_3

$HClO_3$

$HC_2H_3O_2$

H_2S

HNO_2

Write the chemical formulas for the following acids:

Hydrobromic acid

Chloric acid

Phosphoric acid

Sulfuric acid

Nitric acid

Hypochlorous acid

Writing Ionic Formulas

Name	Positive ion	Negative ion	Formula
sodium oxide	Na^{+}	O^{2-}	Na_2O
magnesium chloride			
potassium chloride			
calcium oxide			
aluminum bromide			
lithium phosphide			
aluminum sulfide			
aluminum nitride			
calcium nitride			

Name the following ionic compounds:

Na_2S ______________________________

MgF_2 ______________________________

MgS ______________________________

K_3N ______________________________

Ca_3P_2 ______________________________

Ionic Charges for Transition Metals

Name	Positive ion	Negative ion	Formula
iron(III) sulfide	________	________	____________________
copper(II) chloride	________	________	____________________
copper(I) sulfide	________	________	____________________
copper(II) nitride	________	________	____________________
zinc oxide	________	________	____________________
silver sulfide	________	________	____________________

Name the following ionic compounds:

Cu_3P ____________________________________

Fe_2O_3 ____________________________________

FeI_3 ____________________________________

$CuCl$ ____________________________________

$ZnBr_2$ ____________________________________

Polyatomic Ions

Name	positive ion	negative ion	Formula
sodium nitrate	Na^+	NO_3^-	$NaNO_3$
lithium carbonate			
potassium sulfate			
calcium bicarbonate			
aluminum hydroxide			
lithium sulfite			
sodium phosphate			
iron(II) phosphate			

Name the following ionic compounds:

$CaSO_4$ ____________________

Cu_3PO_4 ____________________

$Al(NO_3)_3$ ____________________

Na_2CO_3 ____________________

$MgSO_3$ ____________________

$Fe(HCO_3)_2$ ____________________

EXCERCISES

Name the following molecular compounds:

1. SeO_2
2. N_2O
3. PCl_3
4. CI_4

Write the chemical formulas for the following molecular compounds:

1. sulfur hexafluoride
2. iodine monochloride
3. tetraphosphorus hexasulfide
4. boron tribromide

Name the following ionic compounds:

1. $MnCl_3$
2. $[NH_4]_2S$
3. K_2O_2
4. $AgBr$
5. $FeCl_3$
6. Au_2SO_4
7. $PbHPO_4$
8. $Sn(OH)_4$

Write the chemical formula for the following ionic compounds:

1. sodium acetate
2. nickel(II) hydrogen sulfate
3. molybdenum(III) permanganate
4. potassium cyanide

Give the name of each of the following compounds:

1. $ZnSO_3$
2. $Na_2Cr_2O_7$
3. $AsCl_3$
4. $Co_2(C_2O_4)_3$
5. HBr(aq)
6. $V_2(CrO_4)_3$
7. N_2O_3
8. PdI_4
9. $Cr(NO_3)_3$

Give the chemical formula for each of the following compounds:

1. ammonium acetate
2. ruthenium(III) sulfide
3. dinitrogen trioxide
4. iron(III) dihydrogen phosphate
5. cadmium chlorite
6. perchloric acid
7. aluminum hydroxide
8. phosphorus pentachloride
9. titanium(III) phosphide
10. selenium dioxide

MOLES

List the number of moles of each kind of element in one mole of each of the following formulas:

Formula	Moles of each kind of atom in one mole of the compound
H_2O	
$CuCl_2$	
Al_2S_3	
$Ba(NO_3)_2$	
C_4H_{10}	

Write the formula of the following compounds from the number of moles given. The elements are listed in the order in which they appear in the compound.

1 mole of carbon and 2 moles of oxygen	
1 mole of nitrogen and 3 moles of hydrogen	
1 mole of carbon and 4 moles of chlorine	
2 moles of iron and 3 moles of oxygen	
1 mole of barium, 1 mole of sulfur, and 4 moles of oxygen	

MOLE DISPLAY

Observe the display of one mole quantities of several substances. For each, write down the name and chemical formula of the substance. For each of the substances, determine its molar mass in grams (to 4 significant figures) and calculate the number of chemical units of each substance in 1.000 gram of that substance. For example, if the compound is water (H_2O) which has a molar mass of 18.015 g, how many molecules of water are in 1.000 g of water?

$$1.000 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.015 \text{ g } H_2O} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 3.343 \times 10^{22} \text{ molecules}$$

Name of Substance	Formula	Molar Mass	Chemical Units/g
Example: Water	H_2O	18.015g/mole	3.343×10^{22} molecules (see calculation above)

SHOW CALCULATIONS:

Problems

1. Calculate the number of grams in each of the following:
   (a) 0.00500 mol of MnO_2

   (b) 1.12 mol of CaH_2

   (c) 0.250 mol of $C_6H_{12}O_6$

   (d) 4.61 mol of $AlCl_3$

2. How many grams of V_2O_3 would you need to have 7.56×10^{22} atoms of oxygen?

3. In a sample of 12.3 grams of carbon tetrachloride, CCl_4:
   (a) what is the number of carbon tetrachloride molecules?

   (b) what is the mass of chlorine present?

   (c) how many chlorine atoms are present?

4. How many grams of N_2O_5 would you have if the sample contains 2.55 g of nitrogen?

5. How many molecules of SiF_4 contain 1.00 gram of fluorine?

6. How many nitrogen atoms are in 3.5 grams of $(NH_4)_3N$?

7. How many grams of $Fe_2(Cr_2O_7)_3$ would you need to have 2.578×10^{22} atoms chromium?

DETERMINATION OF EMPIRICAL FORMULAS

An empirical formula is a chemical formula that denotes the lowest whole number ratio between the individual atoms or ions for a compound. This differs from a molecular formula in that a molecular formula gives the exact number of each atom that is in each molecule of that substance.

Examples:

Chemical formula	Empirical Formula
$NaCl$	$NaCl$
H_2O	H_2O
H_2O_2	HO
NO_2	NO_2
N_2O_4	NO_2

As we can see from these examples a particular molecular formula defines what a particular substance is. However, more than one substance can share the same empirical formula (as is seen in the last two pairs of examples).

How can we go about calculating an empirical formula? This can be done by remembering that the formula (empirical or molecular) gives ratios of moles of atoms present in one mole of the substance. If we already have the masses of the elements in the substance we can calculate the empirical formula. Let's demonstrate the method by using an example. A compound is determined to contain 0.2705 g of sodium, 0.1648 g of nitrogen, and 0.5562 g of oxygen. The first step is to convert the mass of each element present in the compound to moles of that element.

Step 1

$$0.2705 \text{ g Na} \times \frac{\text{mol Na}}{22.99 \text{ g Na}} = 0.01177 \text{ mol Na}$$

$$0.1648 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.01176 \text{ mol N}$$

$$0.5562 \text{ g O} \times \frac{\text{mol O}}{16.00 \text{ g O}} = 0.03476 \text{ mol O}$$

We then determine which element is present in the fewest number of moles, in this case it is Nitrogen with 0.01176 mol. We divide each of the numbers of moles calculated in the first step by the smallest of those numbers. (See Step 2 on the next page.)

<u>Step 2</u>

$$\frac{0.01177 \text{ mol Na}}{0.01176 \text{ mol}} = 1.001 \text{ Na}$$
$$\frac{0.01176 \text{ mol N}}{0.01176 \text{ mol}} = 1.000 \text{ N}$$
$$\frac{0.03476 \text{ mol O}}{0.01176 \text{ mol}} = 2.956 \text{ O}$$

This should give us the ratio of the moles of atoms in the compound. However, looking at the numbers above we find a value of 2.956 for oxygen and we do not see fractional numbers in the formulas. Some things to consider, if the value we get from this last calculation is less than 0.1 away from an integer (3.000-2.956 = 0.044 < 0.1) we just round to the nearest integer, in this case 3. In the example on which we are working we obtain the empirical formula $NaNO_3$.

If the value is more than 0.1 away from an integer we will need to multiply ***all*** of the values by an integer to get values that are close to whole numbers. For example, if one of our values above had been 2.499 (which is approximately 2.5 or 2 ½) we would multiply all of the values by 2. This brings the 2.499 to 4.998 (or 5). For example, we have a compound that contains 0.3685 g of Nitrogen and 0.6315 g of Oxygen. Our calculation is as follows:

<u>Step 1</u>

$$0.3685\text{g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.02630 \text{ mol N}$$
$$0.6315 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.03947 \text{ mol O}$$

<u>Step 2</u>

$$\frac{0.02630 \text{ mol N}}{0.02630 \text{ mol}} = 1.000 \text{ N}$$
$$\frac{0.03947 \text{ mol O}}{0.02630 \text{ mol}} = 1.501 \text{ O}$$

In this example the second number is 1.501 which is the same as 1½ so we would multiply all of the above numbers by 2.

<u>Step 3</u>

1.000 N × 2 = 2 N
1.501 O × 2 = 3 O

This gives us an empirical formula of N_2O_3.

Usually several experiments are required to determine the amounts of each element present in a compound. In that case the data is collected together as percent composition data. Percent composition data can also be used to determine empirical formulas. Remember that percent composition is the same as saying that there are a certain number of grams of a particular element in every 100.00 g of the compound. Therefore, if we assume that there are 100.00 g of the compound we can determine the mass of each element present in that 100.00 g.
For example, we have a compound with the following percent composition, 32.69% sulfur, 2.06% hydrogen, and 65.25% oxygen. If we assume there are 100.00 g of the compound then we have:

100.00 g of compound × (32.96 g S / 100.00 g of compound) = 32.69 g S

100.00 g of compound $\times$ (2.06 g H / 100.00 g of compound) = 2.06 g H
100.00 g of compound $\times$ (65.25 g O / 100.00 g of compound) = 65.25 g O

Now that we have masses of the elements present the process is the same as we used earlier. We can combine these steps into one calculation:

$$100.00\text{ g c'p'd}\times\frac{32.69\text{ g S}}{100.00\text{ g c'p'd}}\times\frac{1\text{ mol S}}{32.06\text{ g S}}=1.020\text{ mol S}/1.020\text{ mol} = 1.000\text{ S} = 1\text{ S}$$

$$100.00\text{ g c'p'd}\times\frac{2.06\text{ g H}}{100.0\text{ g c'p'd}}\times\frac{1\text{ mol H}}{1.008\text{ g H}}=2.044\text{ mol H}/1.020\text{ mol}=2.004\text{ H} = 2\text{ H}$$

$$100.00\text{ g c'p'd}\times\frac{65.25\text{ g O}}{100.00\text{ g c'p'd}}\times\frac{1\text{ mol O}}{16.00\text{ g O}}=4.078\text{ mol O}/1.020\text{ mol}= 3.998\text{ O} = 4\text{ O}$$

This gives an empirical formula of H_2SO_4.

To obtain the chemical (molecular) formula from the empirical formula, the molar mass of the compound is required. The first step is to calculate the empirical formula. Next, calculate the molar mass of an empirical formula unit. To find the number of empirical formula units in one chemical formula unit, divide the molecule's molar mass by the empirical mass (you should get very close to a whole number). Multiply the subscripts of the empirical formula by the number of empirical formula units / chemical formula. For example if we know that the empirical formula is CH_2O and the molecule has a molar mass of 180.15 g/mol, what is the molecular formula?

Find the empirical mass.
In this case the empirical mass is 30.03 g/mol.

Divide the molar mass by the empirical mass.

$$\frac{180.15}{30.03}=5.999=6$$

We must now multiply all the subscripts in the empirical formula by 6 which gives us a molecular formula of $C_6H_{12}O_6$.

Empirical Formula Problems

1. A compound contains 43.6% phosphorus by mass. The remainder is oxygen. What is the empirical formula of the compound?

2. What is the empirical formula of a compound that is (by mass) 37.70% sodium, 22.95% silicon and the remainder is oxygen?

3. What is the empirical formula of a compound that is (by mass) 70.57% carbon, 5.93% hydrogen and the rest is oxygen.

4. A compound has the following percent composition by mass: 40.69% K, 16.12% P, 41.62% O and the remainder is H. What is the empirical formula? If the molar mass is 384.38 g/mol, what is the molecular formula of the compound?

5. A 3.054 gram sample of vanadium was reacted with oxygen to form 5.454 grams of product. What is the empirical formula of the product?

6. When 2.500 grams of a compound that contained carbon, hydrogen and oxygen was burned (reacted with molecular oxygen) the products of the reaction was 3.545 grams of carbon dioxide and 2.176 grams of water. The molar mass of the compound is 62.08 g/mole. What is the molecular formula of the compound?

CHEMICAL EQUATIONS STUDY QUESTIONS

Balance the following chemical reactions:

$B_2O_3 + C + Cl_2 \rightarrow BCl_3 + CO$

$BaF_2 + H_3PO_4 \rightarrow Ba_3(PO_4)_2 + HF$

$NH_3 + O_2 \rightarrow N_2 + H_2O$

$KNO_3 + K \rightarrow K_2O + N_2$

$BF_3 + H_2O \rightarrow H_3BO_3 + HBF_4$

$LiOH + CO_2 \rightarrow Li_2CO_3 + H_2O$

Balance the following chemical reactions:

aluminum + iron(III) oxide $\rightarrow$ aluminum oxide and iron
(what would be the classification for this reaction?)

ammonium nitrate $\rightarrow$ dinitrogen monoxide and water
(what would be the classification for this reaction)

In a synthesis reaction, aluminum metal and chlorine gas react to form aluminum chloride.

CHEMICAL REACTION PROBLEMS

1. How many grams of hydrogen gas are required to completely convert 25 g of hot magnetic iron oxide (Fe_3O_4) into elemental iron? Water vapor is the other product of the reaction.

2. Air is 23% oxygen by mass and has a density of 0.0012 g/L. How many liters of air are required to complete the combustion of 93 g of phosphorus to form diphosphorus pentoxide (the only product of the reaction)?

3. Calcium hydroxide reacts with hydrochloric acid, HCl, to give calcium chloride and water. If the reaction starts with 1.00 gram each of calcium hydroxide and hydrochloric acid, how many grams of calcium chloride will be formed in the reaction? What will be the mass of the reactant in excess that is left at the end of the reaction?

4. Ammonia gas, $NH_{3(g)}$, can be prepared by reacting hydrogen gas with nitrogen gas. Using 1.00 g of hydrogen and excess nitrogen, how many grams of ammonia will be produced if the percent yield for the reaction is 64%?

5. A mixture was created using table salt, sodium chloride, and sugar, $C_6H_{12}O_6$. A 5.00 g sample of the mixture was burned (only the sugar will burn). Water and 2.20 g of carbon dioxide were produced. What is the percent by mass of salt in the **original** mixture?

6. Balance the following reaction:

   $$P_4 + NaOH + H_2O \rightarrow PH_3 + NaH_2PO_2$$

   How many grams of phosphine, PH_3, can be produced from 20.0 grams of phosphorus, 50.0 grams of sodium hydroxide, and 15.0 grams of water?

KINETIC ENERGY PROBLEMS

$$\text{Kinetic Energy} = \tfrac{1}{2} \times \text{mass} \times \text{velocity}^2$$

1. A car with a mass of 1.5×10^3 lbs is traveling down the 405 Freeway at a speed of 81 miles per hour. What is the kinetic energy of the car in Joules? (1 lb. = 454 g, 1 mile = 1.61 km)

2. The Road Runner dropped a 75 lb. anvil on Wile E. Coyote's head and it hits with a kinetic energy of 68 kJ. What was the speed of the anvil (in miles per hour) as it hit Mr. Coyote's head?

3. An asteroid hits Los Angeles with a velocity of 1.5×10^3 km/s. It releases 2.25×10^{13} kJ of energy. What is the mass of the asteroid in pounds?

SPECIFIC HEAT AND PHASE CHANGE PROBLEMS

1. Solid gold has a specific heat of 0.13 J/g °C. If 12.5 calories of energy are added to 3.75 g of gold with a starting temperature of 19.1 °C, what will be the final temperature of the gold?

2. Solid iron has a specific heat of 0.444 J/g °C. If 75 calories of energy were removed from a piece of iron, and the temperature of the iron went from 21.4 °C to –25.2 °C, what was the mass of the iron?

3. A certain substance has a specific heat of 0.785 J/g °C as a solid and a specific heat of 1.56 J/g °C as a liquid. Its melting point is -15.5 °C. Its Heat of Fusion is 600.0 cal/g. How many joules of energy must be removed from 15.0 grams of substance to change it from liquid at 35.0 °C to solid at -25.3 °C.

4. The melting point temperature of Copper is 1083 °C. Copper's Heat of Fusion is 13.0 kJ/mole. The specific heat of solid Copper is 0.382J/g °C. How many kJ of energy must be removed from 7.65 grams of liquid copper at its melting point temperature to convert it to solid Copper at 20.5 °C?

5. For water:

Melting point temperature = 0.0 °C	<u>Specific Heats (J/g °C)</u>
Boiling point temperature = 100.0 °C	gas = 2.0
Heat of Fusion = 6.01 kJ/mole	liquid = 4.18
Heat of Vaporization = 40.7 kJ/mole	solid = 2.1

Calculate the kJ of energy needed to convert 10.0 grams of ice at -15.3 °C to steam at 125.6 °C.

6. A sample of metal that has a mass of 45.0 grams was heated to 85.3 °C and dropped into 65.0 grams of water (specific heat = 4.18 J/g °C) having an initial temperature of 25.0 °C. The final temperature of the metal and water was 31.5 °C. What is the specific heat of the metal?

7. A 3.62 kg sample of a ceramic material is being tested for use in storing heat in a new type of solar heating system. After the sample is heated to 92.8 °C, it is dropped in 8555 grams of water a 22.5 °C. Both water and the ceramic material reach a final temperature of 34.0 °C. Calculate the specific heat of the ceramic. The specific heat of liquid water is 4.18 J/g °C.

Intermolecular Forces:

1. What intermolecular forces will be present between molecules of each of the following substances?

   a. H_2S

   b. CF_4

   c. NH_3

   d. CH_3OH

2. In each of the following pairs of molecules, predict which one would have the higher boiling point.

   a. H_2O or H_2S

   b. $CH_3CH_2CH_3$ or $CH_3CH_2CH_2CH_3$

   c. $HOCH_2CH_2OH$ or CH_3CH_2OH

   d. CO or N_2

3. Which out of the following pairs would have the highest vapor pressure?

   a. H_3COCH_3 or CH_3CH_2OH

   b. Cl_2 or Br_2

   c. NH_3 or H_2O

   d. CO or N_2

BOYLE'S LAW

INTRODUCTION

The pressure of a gas is inversely proportional to volume of a gas if the number of moles of gas and the temperature remain constant. This relationship can be expressed as:

$$P \propto \frac{1}{V}$$

or

$$PV = C$$

This relationship is known as Boyle's Law. It states that when the number of moles of a gas and the temperature of a gas is constant, the product of the pressure and the volume is also a constant.

In this experiment, you will measure the pressure of a sample of air at several different volumes. You will also graph the volume versus the reciprocal of the pressure (1/P) to show that the pressure is inversely proportional to the volume of a gas.

PROCEDURE

Observe your professor's demonstration of Boyle's Law using an apparatus consisting of a pressure gauge connected to a calibrated glass syringe.

1. **Before** connecting the pressure gauge to the syringe, pull the plunger about half way out. Connect the syringe to the gauge. Measure and record the syringe volume (to 0.1 mL) and the pressure (to 0.1 pounds per square inch).

2. Set the syringe two seven other volumes, both larger and smaller than the first volume. Measure and record each volume and each pressure at that volume.

3. To accurately demonstrate Boyle's Law, the volume of air must reflect the total volume in the system. In addition to the air in the syringe, there is an additional 5.0 mL of air in the tubing, gauge and fittings. Add 5.0 mL to each measured syringe volume and record the total volume of air for each reading.

4. Convert the pressures for each reading to mmHg (1 psi = 51.7 mmHg).

5. Calculate the reciprocal pressure (1/mmHg). Put these answers in scientific notation **with all** the numbers having an exponent of -3.

6. Draw a graph in which you plot the **total** volume of air on the y-axis (vertical axis) and reciprocal pressure (in scientific notation) on the x-axis (horizontal axis) **Note: your graph increments do not need to start at zero for either axis.** Draw a best-fit straight line through your data points.

Reading	Pressure of air in psi	Pressure in mmHg	1/Pressure (in sci. note. exp = 10^{-3})	Measured Syringe volume	Total Volume of air (add 5.0 mL)
1					
2					
3					
4					
5					
6					
7					
8					

Did your plot of volume against 1/P fall in a straight line?

What do these data tell you about the relationship between pressure and volume of a gas sample?

GAS LAW PROBLEMS

1. A proposed space missions is to send an atmospheric balloon to Mars to carry an instrument array across the surface of the "Red Planet." The temperature on Mars is about the same as it is in Antarctica. Assume the balloon is filled to a volume of 1.5 L in Antarctica. The atmospheric pressure there is 725 mmHg. The atmospheric pressure on Mars is about 70 mmHg. What will be the volume of the balloon when it is released on Mars?

2. A piston is placed in a cylinder when the external pressure is 747 mmHg. The volume enclosed by the piston and cylinder is 14.7 L. A high-pressure ridge moves through the area and the volume in the cylinder decreases to 14.3 L. What is the pressure after the ridge moves through?

3. A gas sample has a pressure of 583 mmHg and a volume of 3.75 L. If it is compressed to 1.05 L, what will be its pressure in atm?

4. A gas sample has a volume of 2.50 L at 34.5 °C. What will be its volume in m^3 at 525 °C?

5. A gas sample has a pressure of 378 torr at 27.6 °C. What will be its temperature in °C if its pressure becomes 2.45 atm? Assume the volume remains constant.

6. What is the new volume in milliliters, of a 4.00 mL sample of air at 0.875 atm and 250.5 °C that is compressed and cooled to 305 torr and 185 °C?

7. A sample of neon gas in a cylinder had a volume of 1.05 L, a pressure of 1.90 atm, and a temperature of 55.2 °C. The volume of the cylinder was increased to 1.50 L and the pressure of the gas became 735 torr. What was the new temperature of the gas in °C?

8. When 150.4 grams of magnesium carbonate decomposes to magnesium oxide and carbon dioxide gas, how many liters of carbon dioxide are produced at 28 °C and 745 torr?

9. A 2.50 liter flask at a temperature of 28.5 °C was used to collect gaseous propane (C_3H_8). After the collection was complete, the pressure in the flask was 1.35 atm. How many grams of propane were in the flask?

10. 3.455 grams of gas are placed into a 150.0 mL container at 25 °C. The pressure is measured at 867 mmHg. What is the molar mass of the gas?

11 How many liters of oxygen gas will be required for the complete combustion of 5.00 grams of ethane (C_2H_6)

(a) at STP?

(b) at 695 torr and 45 °C?

12. Suppose 435 mL of Ne gas at 21 °C and 1.09 atm, and 456 mL of SF_6 at 25 °C and 0.89 atm are put into a 325 mL flask at 30.2 °C.

(a) What will be the total pressure in the flask?

(b) What is the mole fraction of for each of the gases in the flask?

(c) How many grams of each gas are in the flask?

13. A 3.75 L flask contains 3.56 grams of neon, 4.35 grams of nitrogen and 8.24 grams of carbon dioxide gases at a temperature of 35.2 °C.
    (a) What is the total pressure inside the flask?

    (b) What is the mole fraction of each gas in the flask?

14. A gas sample contains argon and oxygen gases and has a total pressure of 2.35 atm. If the partial pressure of the argon is 865 torr, what is the mole fraction for each of the gases in the mixture?

15. H_2S gas burns in oxygen gas to produce sulfur dioxide gas and water. How many grams of water could be produced from 2.5 liters of H_2S gas and 5.35 liters of oxygen gas? Both the H_2S gas and oxygen gas were at 745 torr and 65 °C.

16. Calcium nitride can be synthesized from calcium metal and nitrogen gas. How many grams of calcium nitride can be made from 5.00 grams of calcium and 5.00 liters of nitrogen gas at 1.15 atm and 25 °C?

SOLUTIONS

INTRODUCTION

A solution is a homogeneous mixture of two or more substances. The substance that is present in the larger amount is known as the solvent. The substance that is present in the smaller amount is known as the solute. The solute is dissolved into the solvent. Most solutions with which we are familiar are liquids. However, solids, liquids or gases may be either solute or solvent. For example, air is a solution of oxygen and other gases in nitrogen. Oxygen is the solute and nitrogen is the solvent.

The solubility of a substance depends on:

1. The nature of the solute and solvent
2. The amount of solute and solvent
3. The temperature
4. The pressure (for gaseous solutions).

A solution will form between the solute and solvent when the attractive forces between the solute and solvent are large enough to overcome the attractive forces within the solute and within the solvent. Therefore, a polar solute will dissolve in a polar solvent and a non-polar solute will dissolve in a non-polar solvent. For this reason we say that "like dissolves like." For aqueous solutions temperature plays a role in the amount of solute which will dissolve. Usually, the higher the temperature, the more solute will dissolve.

PROCEDURE

(Professor Demonstration)

A. Paper discs.

A jar is filled with two liquids, water and carbon tetrachloride. Some paper discs with graphite on one side are added to the jar. Observe the orientation of the discs in the jar. Shake the jar and observe the orientation of the disks after the contents have settled.

Questions:

1. What is the polarity of the two liquids in the jar? What does that tell you about the polarity of the fibers in the paper? About the polarity of graphite?

2. Carbon tetrachloride is denser than water so it is on the bottom. Cyclohexane is less dense than water and has the same polarity as carbon tetrachloride. What would be the orientation of the disks, if cyclohexane replaced carbon tetrachloride in the jar?

B. Three Layers

A large test tube is filled with cyclohexane, water, and carbon tetrachloride. They form three layers with the cyclohexane on the top and the carbon tetrachloride on the bottom. A small amount of iodine is added.

Questions:

1. What happens to the appearance of the layers in the test tube after the iodine is added? What does this tell you about the polarity of iodine?

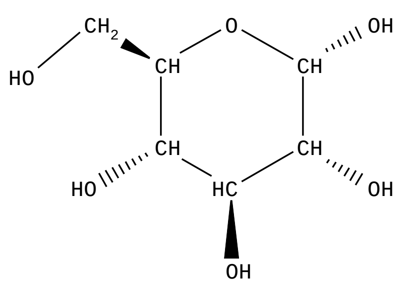
2. What would happen to the appearance of the contents of the test tube if it were stoppered and the contents mixed? (Remember "like dissolves like.")

Questions:

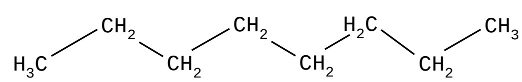
1. NaCl is soluble in water but I_2 is not. Explain. What type(s) of intermolecular forces are involved?

2. Would you expect sugar to dissolve in gasoline? Explain. What type(s) of intermolecular forces are involved?

Sugar



Gasoline



3. Would you expect carbon monoxide to be soluble in water? Why or why not?

4. Does carbon tetrahydride (methane) dissolve in water? Explain.

CONCENTRATION STUDY PROBLEMS

1. A solute concentration is 3.74%(m/m). What would this concentration be in μg solute per mg of solution?

2. A solution was created by dissolving 45.6 grams of sodium sulfate in enough water to make 435 mL of solution. The density of the solution that is created is 1.04 g/mL. For the solution:
   (a) Calculate the %(m/m)

   (b) Calculate the %(m/v)

   (c) Calculate the molarity

3. A solution contains 2.46 grams of calcium citrate in 125 mL of solution. The density of the solution is 1.0 g/mL.
   (a) What is the %(m/v) of calcium nitrate in this solution?

   (b) What is the percent by mass of the calcium nitrate in the solution?

4. A solution is 8.25% by mass ammonium acetate. How many grams of ammonium acetate would be in 235 mL of solution that has a density of 1.18 g/mL?

5. The density of water at 25°C is 0.9975 g/mL and the density of ethanol, C_2H_5OH, is 0.786 g/mL. What is the molarity of a solution made from 17.6 grams of ethanol and 146.7 grams of water? Assume the volumes are additive.

6. A quantity of solid sodium peroxide is reacted with excess water. The products of the reaction are aqueous sodium hydroxide and oxygen gas.
   (a) If 65.0 mL of oxygen gas at STP and 150.0 mL of sodium hydroxide solution resulted from the reaction, what is the molarity of the sodium hydroxide?

   (b) How many grams of sodium peroxide were used in problem 6(a)?

   (c) What would be the answer to question 6(a) if 84.5 mL of oxygen gas was collected at 35°C and 745 torr?

7. Iron(II) iodide reacts with chlorine gas to produce iron(III) chloride and iodine. How many grams of iodine could be made from 35.5 mL of 6.45 M iron(II) iodide and excess chlorine? What type (classification) is this reaction?

8. How many grams of nickel(II) phosphate could be made from 8.7 mL of 1.50 M nickel(II) nitrate and 5.00 mL of 1.25 M ammonium phosphate in a double replacement reaction?

9. How many mL of 4.50 M calcium hydroxide solution would be required to completely react with 6.36 mL of 2.5 M hydrochloric acid in a double replacement reaction?

10. When 132.8 mL of a barium chloride solution reacted with excess potassium chromate solution in a double replacement reaction, 1.05 grams of barium chromate was formed. What was the molarity of the barium chloride solution?

11. 10.0 grams of aluminum was added to 155 mL of 0.950 M hydrochloric acid. The products of the reaction were a solution of aluminum chloride and hydrogen gas. Assume that the volume of the aluminum chloride solution is 155 mL.
    (a) What is the molarity of the aluminum chloride Solution?

    (b) How many liters of hydrogen gas were produced if the pressure was 745 torr and the temperature was 32.5°C?

12. A 6.35 gram sample of a mixture containing sodium carbonate is found to react completely with 75.0 mL of 1.25 M hydrochloric acid. The other components of the mixture did not react. The products of the reaction between the sodium carbonate and hydrochloric acid were sodium chloride, carbon dioxide, and water.
    (a) What was the percent by mass of sodium carbonate in the original mixture?

    (b) How many liters of carbon dioxide would have been generated if it had been collected at STP?

13. Hydrochloric acid reacts with strontium hydroxide to produce strontium chloride and water. If 35.3 mL of a hydrochloric acid solution completely neutralized 85.5 mL of a 3.85 M strontium hydroxide solution, what was the molarity of the hydrochloric acid?

14. Sulfuric acid reacts with sodium hydroxide to produce sodium sulfate and water. How many mL of a 4.85 M sulfuric acid solution would it take to completely neutralize 15.85 grams of sodium hydroxide?

IONIC AND NET IONIC EQUATIONS

Some Solubility Rules:

(a) Group 1A, ammonium, acetate, and nitrate compounds are soluble.
(b) Chloride, bromide, and iodide compounds are soluble except silver, lead(II) and mercury(I).
(c) Carbonates and phosphates are insoluble except for Group IA and ammonium compounds.
(d) Hydroxides are usually insoluble except for Group IA, strontium, barium and ammonium compounds.
(e) Sulfides are usually insoluble. Some exceptions are Group IA and ammonium compounds

Use the solubility rules given above and write chemical equations, ionic equations **and** net ionic equations for the double replacement reactions that occur between the following chemicals (if no reaction occurs write "no reaction" for the net ionic equation):

1. Ammonium carbonate and vanadium(III) bromide

2. Hydriodic acid and potassium hydroxide

3. Sodium chloride and silver nitrate

4. Lead(II) acetate and rubidium chloride

5. Barium hydroxide and phosphoric acid (complete neutralization)

6. Potassium chloride and ammonium acetate

7. Iron(III) nitrate and sodium sulfide

FINAL REVIEW PROBLEMS

1. $$\frac{\left(5.117\times10^{-2}+0.09327\right)\left(2.11\times10^{2}\right)}{\left(1.044-0.556\right)}$$

2. $$3.63+\frac{2677\times10^{1}}{\left(1.4+8.2\right)}$$

3. Suppose that in a different universe there is an element, Ω, that has an average atomic mass of 415.5 amu. Ω has two isotopes, one with a mass of 412.2 amu and one with a mass of 416.6 amu. What is the percent abundance for each of the isotopes of Ω?

4. Write the electron configuration and orbital diagrams of each of the following elements using the noble gas abbreviation form:

   (a) Phosphorus

   (b) Zirconium

   (c) Lead

   (d) Uranium

   (e) Iodine

5. For each of the following, give the electron dot structure. For any **molecular** compound or polyatomic ion, also give the molecular and electron pair geometries, and indicate if the molecule is polar.

   (a) Phosphorus trichloride

   (b) Carbonate ion

   (c) Magnesium bromide

   (d) H_2CS (carbon central)

   (e) Calcium nitride

   (f) Sulfur dioxide

6. For the following reaction,

$$3\ Na(s) + P(g) \rightarrow Na_3P(s)$$

50.00 g of sodium and 60.00 L of phosphorus gas at 755 torr and 645°C were combined. How many grams of sodium phosphide were produced?

7. A certain substance has a specific heat of 0.355 J/g °C for the solid, a specific heat of 0.552 J/g °C for the liquid, and a melting point of 843 °C. A 14.3 gram sample of the substance required 4.35 kJ of energy to change its temperature from 825 °C to 853 °C.
(a) What was the heat of fusion for the substance in cal/g.

(b) If energy was being added to the substance at a rate of 15 J/s, how many minutes would it take for 6.25 g of the substance to melt?

8. A certain molecular compound is 30.45% by mass nitrogen and the remainder is oxygen. At 25.5 °C and a pressure of 875 mmHg, a gaseous sample of this compound has a density of 4.329 g/L. What is the molecular formula of the compound?

9. An aqueous solution of ammonium hydrogen phosphate has a concentration of 1.350 M and a density of 1.045 g/mL.
   (a) What is the percent by mass of the ammonium hydrogen phosphate in the solution?

   (b) How many atoms of hydrogen would be in the ammonium hydrogen phosphate dissolved in 2.55 mL of the solution?

10. Potassium carbonate reacts with nitric acid to form potassium nitrate, water and carbon dioxide gas. When 5.00 g of potassium carbonate reacts with 35.5 mL of 1.75 M nitric acid solution, 0.695 L of carbon dioxide gas is formed at 25 °C and 745 torr. What is the percent yield of the reaction?

11. Potassium and nitrate containing compounds are soluble in water. Most sulfates are soluble but one of the exceptions is lead(II). When 25.5 mL of aqueous lead(II) nitrate was reacted with excess potassium sulfate solution, 5.377 grams of lead(II) sulfate were produced.
    (a) Write a chemical equation, ionic equation, and net ionic equation for the reaction. Include phase labels.

    (b) Calculate the molarity of the lead(II) nitrate solution used.