Welcome to Chemistry 51. Regardless your instructor, the following information pertains to all students in Chemistry 51 at LA Valley College.

All courses at Valley College have certain objectives called 'Student Learning Outcomes' or 'SLO'.

Student Learning Outcomes for Chemistry 51 (SLO)

Apply chemical prinicples to the health science fields

The college also has standard policies concerning academic dishonesty. Please read the department's statement on academic dishonesty and the college's statement on plagiarism.

All forms of academic dishonesty (cheating) will not be tolerated in this class. Any student representing another person's work (on exams, quizzes or in lab reports) as their own is guilty of plagiarism. Plagiarism is the use of others' words and/or ideas without clearly acknowledging the source from which they were obtained. When you incorporate those words and ideas into your own work, you must give credit where credit is due. Plagiarism, intentional or unintentional, is considered academic dishonesty and is not tolerated. Plagiarism or cheating (for example, submitting altered papers for regrading, bringing unauthorized notes into exams) will receive a zero or 'F' for that assignment. A student found to be cheating may also be reported to the Vice-President's Office for inclusion in the student's permanent college record. You may refer to the standards of student conduct and disciplinary actions in the current catalog or schedule of classes. For further information on plagiarism, go to the Writing Center website (www.lavc.edu/WCweb/plagiarism.html) and refer to the STANDARDS OF STUDENT CONDUCT AND DISCIPLINARY ACTION in the current Schedule of Classes and Catalog.

Should you require any accommodation for a disability, please read the following:

If you are a student with a disability requiring classroom accommodations, and have not yet contacted the office of Services for Students with Disabilities (SSD), please do so in a timely manner. The SSD office is located in the Student Services Annex, Room 175 or you may call SSD personnel at (818) 947-2681 to make arrangements to meet with an SSD counselor. If the SSD office has already sent a memo confirming the specific accommodations required by an SSD registered student for this class, please meet with your instructor to discuss the necessary arrangements.

CHEMISTRY 51 LAB MANUAL

EXPERIMENT ONE

LABORATORY SAFETY AND EQUIPMENT

The chemistry laboratory, with its equipment, glassware and chemicals, has the potential for accidents. Everyone completing 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 is expected to 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 actual work. If you have a question, ask your professor for clarification **BEFORE** starting the procedure.

2. Perform 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 wand and 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 areas 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 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 paper towel. Use Sodium Bicarbonate to neutralize any acid spills prior to wiping up the spill.

6. Dispose of broken glass in the special container provided. **Do not** 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, use a fire blanket—wrap yourself in the blanket and **stop, drop and roll** to extinguish the flames. Alternatively, (particularly if you find yourself closer to the safety shower), stand under running water to douse 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 to share with two other students. 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. The equipment list and pictures of the equipment is provided in Appendix 1 of this laboratory manual.

5. Print and sign the Safety Pledge provided on the next page and turn it in to your instructor.

SAFETY PLEDGE

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.
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 LAVC 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 51	Printed Name
Date Section Number	Signature

EXPERIMENT TWO
MATH AND THE CALCULATOR

Numbers used in science are usually measurements of some sort. Some examples are the volume of a solution 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 precision of 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 obtained 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 or $\mathbf{x10^x}$** (depending on what kind of calculator you have). Note that the $x10^x$ key is different from the 10^x key—do not use the 10^x key.
3. Enter the power of ten with the appropriate sign. (To change the sign of the power of ten, press the key marked (**+/-**) or (-), depending on your calculator.

<u>DO NOT</u> enter the "times 10" into the calculator. The **EE** , **EXP, or $\mathbf{x10^x}$** enters a scientific number in the calculator—you do not enter a number like 3×10^4 as '3 x 10 EE 4'. If you do, you have multiplied your number by a factor of ten and have actually entered 3×10^5. 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.**

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N A M E

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 _____________________________

61 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 _____________________________

C. Write the following as standard decimal numbers.

4.09×10^{2} ____________________________

3×10^{-4} _____________________________

5.315×10^{1} __________________________

C. (continued)

8.2×10^{-3} ______________________________

3.156×10^{3} ______________________________

D. 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 ______________________________

E. 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		

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

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

G. Perform the following addition and subtraction calculations. Give a final answer with the correct number of significant figures and with correct units where appropriate.

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	

H. Perform the following mixed operations.

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

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 meter stick 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?

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

20.03 __________ 0.0094 __________

53.000 __________ 2.474×10^{-3} __________

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

300.4 ___________________ 0.0059301 ________________

0.0000003 ________________ 1,321,143.00 _______________

6. 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} ___________ 8.54000×10^{5} ___________

7. Perform the following calculations and express your answer using the correct number of significant figures.

$1234.7 + 7 - 3.25 =$ ______________________

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

$\frac{(11.45)(2.3)}{(2.00)(0.0345)} =$ ___________________________

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

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

EXPERIMENT THREE
MEASUREMENT OF LENGTH, VOLUME, and MASS

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N A M E

In this laboratory exercise we will use rulers and meter sticks to make measurements of length and graduated cylinders to make measurements of volumes. We shall also measure the dimensions of a regularly shaped solid object to calculate the volume of this object.

The most important thing about recording your measurements is to indicate the proper degree of precision with which you have made the measurement. The proper degree of precision depends upon the markings on your instrument. For example, a meter stick marked in 1 centimeter increments is always read to the nearest 0.1 cm. On the other hand, you must record all of your measurements to the nearest 0.01 cm if the meter stick is marked in 0.1 centimeter increments. Regardless of the size of your measurement, the degree of precision you record is always the same for any given instrument. This idea is the most important concept involved in this laboratory exercise. Your lab report will be graded primarily on how properly you record the degree of precision of your results.

A. Examine a meter stick and answer the following questions

   a. The smallest marking (expressed in cm) on this meter stick is to the_________cm

   b. The degree of precision with which this instrument is read is to the nearest _________cm

B. Examine the three graduated cylinders in the lab.

   a. The 1 L cylinder is marked in_________mL increments and must be read to the nearest______mL

   b. The 100 mL cylinder is marked in_____mL increments and must be read to the nearest______mL.

   c. The 10 mL cylinder is marked in_____mL increments and must be read to the nearest_____mL.

C. Using a ruler or meter stick, record your measurements of the following objects. Convert your measurements from centimeters to mm and m, making sure to retain the correct degree of precision.

OBJECT	cm	mm	m
Your height	___________	___________	___________
Your little fingernail	___________	___________	___________
Your little finger	___________	___________	___________
Your pencil	___________	___________	___________
Length of your shoe	___________	___________	___________

D. Obtain a rectangular solid object and measure its height, length and width. Report your measurements as centimeters. Calculate its volume using the equation. Show your units for both your measurements and your calculated volume.

$$V = (\text{length}) \times (\text{height}) \times (\text{width}).$$

Length ____________________

Width ____________________

Height ____________________

Calculated volume __________________

Show your calculations below and explain how you determined the precision in your answer.

E. Record the volumes of liquids found in the display of three graduated cylinders.

Cylinder A ____________

Cylinder B ____________

Cylinder C ____________

Add the three volumes together and record your answer with the correct degree of precision.

___________+____________+_____________=______________mL

F. Use the triple beam balances to determine the mass of the following objects. Convert your measurements to mg and kg, keeping the correct degree of precision.

OBJECT	g	mg	kg
Your pencil	_________	__________	___________
Your keys	_________	__________	___________

G. Add the weight of your keys to that of the weight of the pencil.

__________g keys +___________g pencil =____________g

Calculate the percent by mass contributed by your keys to the total mass of the two objects. (Careful with sig figs and show your work)

Calculate the percent by mass contributed by your pencil to the total mass of the two objects. (Careful with sig figs and show your work).

Now take your keys and pencil and weigh them together on your balance.

________________g is the measured mass of the two objects together.

Compare this measured mass with the mass you calculated above—explain any differences in the degree of precision you show and the values of your two measurements.

EXPERIMENT FOUR
EXPERIMENTAL DETERMINATION OF DENSITY

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N A M E

In this experiment you will measure the density of a liquid and a solid object. Be sure to record all of your measurements with the correct degree of precision and perform your calculations properly, according to the rules of arithmetic for measured values.

A. Measurement of the density of water.

Weigh a small beaker (degree of precision?). Place about 25 mL of water in a 50 mL graduated cylinder and record its volume (degree of precision?). Pour the water into the beaker and weigh the beaker and water together. From these data, calculate the density of the water, giving your answer with the proper number of significant figures.

Volume of water in cylinder ______________________mL

Mass of empty beaker ______________________g

Mass of beaker with water ______________________g

Mass of water in beaker ______________________g

Density of water ______________________g/mL

Show calculations

After pouring out the water from your beaker, add a second 25 mL of water to the graduated cylinder. As before, record the volume of water in your graduated cylinder. Pour the water into the beaker and weigh the beaker after adding the water. From these data, calculate the density of the water a second time. You need not dry the beaker before performing this second experiment. Why?

Volume of water in cylinder ______________________mL

Mass of empty beaker ______________________g

Mass of beaker with water ______________________g

Mass of water in beaker ______________________g

Density of water ______________________g/mL

Show calculations

What do you notice about the two values of density obtained in these two experiments? How close are your two values? If there is variation, in what decimal place to you see the variation? Why should you expect variation in the last decimal place of your recorded values?

Calculate the average density you determined for this unknown liquid and record below.

[___________g/mL +________g/mL] / 2 =_______________g/mL

B. Measurement of the density of a solid object.

1. Determination of volume by the displacement of water.

   You can determine the volume of any object (including a regularly shaped object such as your cylindrical piece of metal) by determining how much water it displaces. The volume of liquid it displaces expressed as mL equals its volume expressed as cm^3. Why?

   Tie a piece of fish line or secure a bent piece of copper wire to the hook at the top of your cylindrical object. Place roughly 25 mL of water in your 50 mL graduated cylinder and record the volume (degree of precision). Gently submerge the object to just the level of the bottom of the hook. Record the new volume reading. From your two volume readings, calculate the volume of water displaced by the metal object.

   Volume of water in cylinder before submerging object_____________mL

   Volume of water in cylinder after submerging object _____________mL

   Volume of object _____________mL

   Volume of object expressed as cm^3 _____________cm^3

   Show calculation

2. Calculation of density of cylindrical object

Determine the mass of the object and subtract the mass of the hook (1.250 g) in order to determine the mass of the cylindrical object without the hook. Calculate the density of the metal object from these data. Be sure to record your density with the correct number of significant figures.

Compare your experimentally determined density with the given density of the metal out of which your cylindrical object was made. Calculate a 'percent error' for your experimental determination of density by subtracting your measured, experimental value from the accepted or 'theoretical' value (Table 1 on following page). Divide the difference by the theoretical value and multiply by 100.

$$\left(\frac{(\text{experimental value} - \text{theoretical value}}{\text{theoretical value}}\right) x 100 = \%\,error$$

Mass of cylindrical object with hook	________________g
Mass of hook	1.250 g
Mass of cylinder without hook	_______________g
Volume of cylinder without hook	_____________cm^3
Density of your object (without hook)	____________g/cm^3
Density of metal given in Table 1	____________g/cm^3
Percent error of your density determination	_____________

Show calculations below. Watch significant figures!

Table 1. Densities of several substances

Symbol on bottom of cylinder	**Metal**	**Density (g/cm^3)**
C	Copper	8.92
A	Aluminum	2.70
T	Tin	7.28
Z	Zinc	7.14
S	Steel	7.83

Problems

1 Explain why you showed the number of significant figures that you did in your calculation of the percent error of the density of your cylindrical object.

2. What is the density of a piece of zinc that is 2.00 m x 4.00 m x 10.00 m?

3. What is the mass of a 4.56 cm^3 piece of zinc?

4. What is the mass of a 8.45 m^3 piece of copper?

5. What is the volume occupied by 6.56 kg of Al?

6. Which of the substances listed in Table 1 is not an element? Why?

7. 6 small glass beads are found to weigh 9.34 g. The beads are placed in a graduated cylinder filled with ethyl alcohol (density 0.78 g/mL). The reading of the liquid in the cylinder before the beads were added is 21.5 mL. After the beads sink to the bottom of the alcohol the reading is 24.5 mL. Calculate the density of the beads. Be sure to record your answer with the correct number of significant figures. Did it matter than alcohol rather than water is used in this experiment? Why or why not? Does it matter that this experiment was done with a collection of glass beads rather than a single solid object? Why or why not?

EXPERIMENT FIVE
GRAPHING; PHYSICAL PROPERTIES OF ELEMENTS AND COMPOUNDS

In the first exercise of this experiment you will use the ideas you learned last week about density to prepare a graph of mass vs. volume whose slope will be the density. You will observe you professor measure the mass of several pennies while simultaneously determining their volume by the displacement method we used last week to determine the volume of the metal cylinder. You will then graph your results and, from the slope of your curve, obtain the density of the pennies.

In the second exercises you will observe the physical properties of a number of elements and learn to distinguish the appearance of metals from nonmetals. Lastly, you will observe the physical appearances of several compounds and compare their appearances with those of the elements of which they are composed.

A. 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 total volume in 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). From each added mass of pennies calculate the density of pennies in that experiment. Compare the slope of your graph to each of the calculated densities.

Volume reading before adding pennies (mL)	Mass of pennies added (g) **(y axis)**	Volume reading after adding pennies (mL	Volume of added pennies (mL or cm^3) **(x axis)**	Density calculation (g/cm^3)
	0.00	No reading	0.0	

MASS OF PENNIES vs VOLUME OF PENNIES

B **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						

C. 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.

EXPERIMENT SIX
ATOMIC STRUCTURE AND FLAME TESTS

The object of this experiment is to show how flame tests are used to identify the presence of a metal ion in a solution containing an ionic compound. You will also review the concepts of atomic structure and isotopes that you are discussing in lecture.

When metal ions are exposed to the heat of a flame, electrons in the ion are 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. Just as the electron arrangement in an element affects its chemistry, the allowed electron transitions impart unique color to the flame. Because the energies 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.

A. Bunsen Burner

Observe your professor's demonstration of the correct method of lighting a Bunsen burner and performing flame tests.

B. Flame Tests

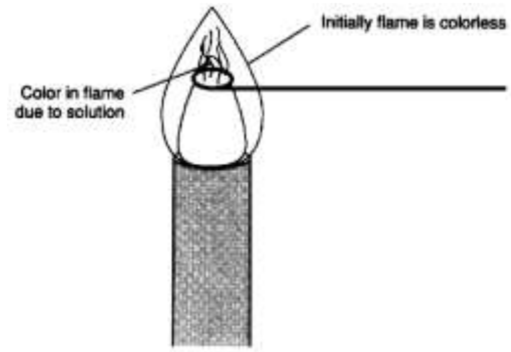
1. Obtain a flame test wire and a cork.

2. Clean the spot-plate from your locker but do <u>NOT</u> dry it. After rinsing the clean plate with deionized water, simply shake off the excess water and place it on a dry paper towel.

3. 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.**

1. If the provided wire has no loop 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. Clean the wire by dipping it in a shell vial containing a **few** mL (1/4th full) of 6-M 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

NEVER DIP A HOT 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 cooled clean wire into one of the known 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 5) BEFORE testing a different solution. Repeat the flame test process for the other known solutions.

8. Clean your spot plate and bring it (on a piece of paper towel) 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 carefully clean your wire between tests of the different solutions and perform a flame test on each of your unknown solutions.

10. If two of your known substances have very similar colors to one of your unknown samples, examine your unknown and the two possible known samples immediately one after the other. You might observe an 'afterflash' of color that is unique to one of the known samples displaying similar colors. Perhaps your unknown sample will or will not have a similar 'afterflash'. Would not that observation help you decide which is your unknown?

11. Dispose of your excess 6-M HCl by pouring down the sink **with the water running into the sink**. Be sure to flush the sink well with water. **DISCARD** your flame test wire in the trash can and return the cork.

____

N A M E

C. Report of Flame Tests for Known Solutions and Determination of Unknowns

Identity of metal ion in the solution	Flame Color

Identification of Unknowns

Unknown Number	Flame Color	Identity of metal ion in the unknown solution

C. Subatomic particles in different isotopes.

Complete the following table for neutral atoms by listing the numbers of protons, neutrons and electrons in each isotope. Recall that the 'isotope number' is the mass number and is the sum of the protons and neutrons in the nucleus, in other words the number of nucleons in the nucleus.

ISOTOPE	PROTONS	NEUTRONS	ELECTRONS
C-14	__________	__________	__________
C-12	__________	__________	__________
__________	11	12	__________
I -131	__________	__________	__________
__________	26	30	__________

Complete the following table for these ions, listing the numbers of protons, neutrons and electrons in each isotopic species.

ISOTOPE	PROTONS	NEUTRONS	ELECTRONS
$Br\text{-}81^{1-}$			
$Ca\text{-}43^{2+}$			
__________	29	23	27
$S\text{-}32^{2-}$		__________	__________

D. Complete the following table of the electron arrangements for several atoms. We are only showing the number of electrons in each principal shell. The pattern is 2 electrons in the first shell, 8 electrons in the second, 8 electrons in the third and up to 18 electrons in the fourth.

ATOM	**NUMBER OF ELECTRONS IN SHELL**			
	1^{ST} SHELL	2^{ND} SHELL	3^{RD} SHELL	4^{TH} SHELL
Na	______	______	______	______
K	______	______	______	______
O	______	______	______	______
S	______	______	______	______
P	______	______	______	______
N	______	______	______	______

When electrons are gained or lost by the 'Representative Elements' the electron arrangement of the ion becomes 'isoelectronic' with the noble gas nearest to it on the periodic table. Complete the table below for ions of the representative elements listed. Use your periodic table to determine with which noble gas the ion's electron arrangement is similar, in other words, isoelectronic.

ION	**NUMBER OF ELECTRONS IN SHELL**				**NOBLE GAS**
	1^{ST} SHELL	2^{ND} SHELL	3^{RD} SHELL	4^{TH} SHELL	
Na^{+}					______
K^{+}					______
O^{2-}					______
S^{2-}					______
P^{3-}					______
N^{3-}					______

Problems

1. How many neutrons are in the following isotopes?

    a. U-235 ______________

    b. U-238 ______________

    c. Fe-60 ______________

    d. Pb-207 ______________

    e. Zn-68 ______________

2. Write nuclide symbols for the following isotopes. As an example recall that you would write P-31 as $^{31}_{15}P$.

    a. U-235 ______________

    b. U-238 ______________

    c. Fe-60 ______________

    d. Pb-207 ______________

    e. Zn-68 ______________

3. Name the noble gas with which the following elements as ions would be isoelectronic.

    a. Ca ______________

    b. Sr ______________

    c. Se ______________

    d. N ______________

    e. Rb ______________

4. Which elements have the given electron arrangements?

a. 2 – 4 ______________

b. 2– 8 – 1 ______________

c. 2 – 8 – 8 – 2 ______________

d. 2 – 8 – 8 – 1 ______________

e. 2 – 3 ______________

5. Complete the following table. Recall that mass number = number of protons + number of neutrons.

Element	Isotopic Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
Fluorine			19			
	${}_{15}^{32}P$					
Calcium		20			22	
			43	20		
Iron			56			
Iron					31	

EXPERIMENT SEVEN

COMPOUNDS—COMPOSITION AND NAMING

____
N A M E

Modern chemists name compounds according to the rules devised by the International Union of Pure and Applied Chemistry (IUPAC). In Chem 51 we learn to name several distinct categories of compounds: (1) binary ionic compounds composed of metals with fixed charges (2) binary ionic compounds composed of metals with variable charges (3) ionic compounds containing polyatomic ions and (4) binary covalent compounds. In this laboratory exercise we will systematically name compounds from each category and then practice naming a set of compounds for which you must decide the correct naming rules.

A. Writing Ionic Formulas For Binary Compounds with Metals of Fixed Charges

Metals of Groups IA, IIA and IIIB always have one kind of charge: Group IA metals are always +1, those of Group IIA are +2. The other metals that have fixed charges that are not in these groups are Al, Zn, Cd and Ag. Most of the other metals on the periodic chart have at least two different possible charges.

Name	Positive ion	Negative ion	Formula
Sodium Oxide	Na^+	O^{2-}	Na_2O
Magnesium Chloride			
Potassium Iodide			
Calcium Oxide			
Strontium Bromide			
Lithium Phosphide			
Potassium Sulfide			
Barium Nitride			

A. (cont)

Calcium Arsenide ______________ ____________ __________________________

Sodium Sulfide ______________ ____________ __________________________

Lithium Phosphide ______________ ____________ __________________________

Name the following binary ionic compounds:

Na_2O ____

MgF_2 ____

BaS ____

K_3N ____

Ca_3P_2 ____

Na_3PO_4 ____

CdI_2 ____

$AlCl_3$ ____

Zn_3N_2 ____

AgCl ____

CaS ____

B. Binary Ionic Compounds with Metals of Variable Charges

Almost all metals other than those of Group IA and Group IIA have several different possible charges. You are not expected to memorize these charges—you must, however, know how to indicate the charge of the metal ion in a compound's name using a Roman numeral. You must also remember that several metals in the transition metal block do not require a Roman numeral because they can only have one charge. Those metals are Al (always Al^{3+}), Zn (always Zn^{2+}), Cd (always Cd^{2+}), and Ag (always Ag^{+}).

Name	Positive ion	Negative ion	Formula
Iron (III) Sulfide	___________	___________	______________________________
Copper (II) Chloride	___________	___________	______________________________
Mercury (I) Oxide	___________	___________	______________________________
Cobalt (II) Nitride	___________	___________	______________________________
Zinc Fluoride	___________	___________	______________________________
Silver Phosphide	___________	___________	______________________________
Cobalt (II) Oxide	___________	___________	______________________________
Gold (I) Sulfide	___________	___________	______________________________
Aluminum Iodide	___________	___________	______________________________
Cadmium Bromide	___________	___________	______________________________
Manganese (II) Nitride	___________	___________	______________________________

Name the following ionic compounds:

Cr_3P_2 ________________________________

FeO ________________________________

MnI_3 ________________________________

CuCl ________________________________

$ZnBr_2$ ________________________________

C. Ionic Compounds Containing Polyatomic Ions

You cannot avoid learning the formulas, names and charges of some of the polyatomic ions. Appendix 2 lists the polyatomic ions you must memorize. This exercise gives examples of the most common of these polyatomic ions.

Name	positive ion	negative ion	Formula
Sodium Nitrate	Na^+	NO_3^-	$NaNO_3$
Lithium Carbonate			
Potassium Sulfate			
Calcium Bicarbonate			
Aluminum Hydroxide			
Cobalt (II) Sulfite			

C. (cont)

Magnesium Cyanide ____________ ____________ ________________________________

Tin (II) Phosphate ____________ ____________ ________________________________

Barium Nitrite ____________ ____________ ________________________________

Lead (II) Carbonate ____________ ____________ ________________________________

Copper (I) Sulfate ____________ ____________ ________________________________

Name the following ionic compounds:

$CaSO_4$ ____

Cu_3PO_4 ____

$Al(NO_3)_3$ ____

Na_2CO_3 ____

$MgSO_3$ ____

$Fe(HCO_3)_2$ ____

$Pb(SO_4)_2$ ____

Al_2S_3 ____

AgBr ____

D. Naming Covalent Compounds.

Name the following molecular compounds. Recall that you must use di, tri prefixes to indicate the numbers of each element in the binary covalent compound.

1. CS_2 ____
2. N_2O ____
3. PCl_3 ____
4. BrF ____

Write the chemical formulas for the following molecular compounds:

1. Sulfur Hexafluoride ____________________________________
2. Iodine Monochloride ____________________________________
3. Tetraphosphorus Hexasulfide ____________________________________
4. Boron Tribromide ____________________________________

E. Fixed and variable charges on metal ions.

Name the following ionic compounds:

1. $MnCl_3$ ____
2. $(NH_4)_2S$ ____
3. K_2O ____
4. AgBr ____
5. $FeCl_3$ ____
6. Au_2SO_4 ____
7. $Sn(OH)_4$ ____

Write formulas of the following ionic compounds

1. Zinc Chloride ______________________________
2. Lead (IV) Sulfide ______________________________
3. Sodium Sulfate ______________________________
4. Ammonium Phosphate ______________________________
5. Calcium Hydroxide ______________________________
6. Potassium Nitride ______________________________
7. Magnesium Phosphide ______________________________
8. Silver Iodide ______________________________
9. Lithium Oxide ______________________________
10. Beryllium Bisulfate ______________________________

Write chemical formulas for the following compounds:

1. Magnesium Sulfide ______________________________
2. Barium Sulfate ______________________________
3. Gold (III) Phosphate ______________________________
4. Dinitrogen Monoxide ______________________________
5. Lithium Nitrite ______________________________
6. Phosphorus Pentachloride ______________________________
7. Aluminum Bisulfate ______________________________
8. Calcium Hydrogen Phosphate ______________________________
9. Potassium Sulfite ______________________________
10. Manganese (III) Carbonate ______________________________

Name the following compounds:

1. N_2O_4 ____________________
2. $Pb(CO_3)_2$ ____________________
3. $(NH_4)_2SO_4$ ____________________
4. Li_3PO_4 ____________________
5. $Zn(OH)_2$ ____________________
6. $SnCl_4$ ____________________
7. $AsBr_3$ ____________________
8. CO ____________________
9. $FeSO_4$ ____________________
10. CrS ____________________
11. $Sr(NO_3)_2$ ____________________
12. $Ba(NO_2)_2$ ____________________
13. N_2O ____________________
14. $CaCl_2$ ____________________
15. SCl_2 ____________________
16. $SnCl_2$ ____________________
17. $ZnCl_2$ ____________________
18. $Fe(CN)_3$ ____________________
19. $SnSO_4$ ____________________
20. PCl_5 ____________________
21. $LiHCO_3$ ____________________
22. $NaHSO_4$ ____________________

EXPERIMENT EIGHT

OBSERVATIONS

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.

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)!!!

PROCEDURE:

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 "experimental" tube 1. Pour the contents of "experimental" tube 1 into the solution in "experimental" 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 "experimental" tube 1, mix with "experimental" 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 "experimental" tube 1 aside to be used later (you do not need to wash it). Wash "experimental" 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 "experimental" tube 2 (containing **solution 2)** into tubes 3, 4, 5, and 6, **refilling "experimental" 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 "experimental" tube 2 aside to be used later (you do not need to wash it). Wash "experimental" 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 "experimental" 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 "experimental" unknown tube into the solution in "experimental" tube 1. Mix well and record your observations.

18. Place another ¾ mL sample of unknown into the "experimental" unknown tube. Pour the contents of the "experimental" unknown tube into the solution in "experimental" tube 2. Mix and record observations. Continue the procedure until the unknown has been tested with all of the experimental 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 <u>not</u> 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 <u>not</u> solution 2. A solution mixed with itself will not produce any reaction. 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 <u>not</u> solution 2.**

    d. Continue this procedure, <u>showing your work</u>, 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.

____
N A M E

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 ___________________

EXPERIMENT NINE
DETERMINATION OF A CHEMICAL FORMULA

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. 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, the chemical formula is the same as its empirical formula.

A. Determination of the formula of a chloride of manganese

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

2. 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 convection currents of the atmospheric gases that are established around warm objects.)

3. 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.

4. Place approximately 250 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.

5. Place the evaporating dish containing the Manganese on top of the beaker containing the boiling water. **CAREFULLY** add approximately 20 drops of 6 <u>M</u> 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.

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

7. 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 a ring and wire screen over a Bunsen burner.

8. 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.

9. 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>.

10. Determine the mass of the evaporating dish and Manganese Chloride product.

11. Calculate the mass of Manganese used, the mass of Manganese Chloride product, and the mass of Chloride that reacted with the Manganese.

12. From the mass of Manganese and the mass of Chloride, determine the percent by mass of each element in the product.

13. Determine the ratio of mol Cl to mol Mn as shown below.

B. Calculation of molar ratio of Cl to Mn in the product

1. From your mass percent data you can calculate the molar ratio of the elements in the compound. First assume that you have 'exactly' 100 g of the compound. The mass percent information tells you how many grams of each element would be in this compound.
2. Convert the mass of each element in the 100 gram product to the equivalent number of moles of that element you would have in 100 grams of the product.
3. Pick the element with the smaller number of moles. Divide that number of moles into the number of moles of the other element.
4. Round your answer to the nearest integer. We assume in our experiment that the number of moles of Mn is 1 and the number of moles of Cl, n, is equal to or larger than one.

   EXAMPLE: You have a compound composed of calcium and chloride that is 36.1 % calcium.

   a. This compound must be 100.0% - 36.1% = 63.9 % Cl

   b. Moles Ca = (36.1 g Ca) x $\left(\frac{1\, mol\, Ca}{40.1\, g\, Ca}\right)$ = 0.900 mol Ca

   c. Moles Cl = (63.9 g Cl) x $\frac{1\, mol\ Cl}{35.5\ g\, Cl}$ = = 1.80 mol Cl

   d. Mol Cl/mol Ca = $\frac{1.80\ mol\ Cl}{0.900\, mol\, Ca} = \frac{2\ mol\ Cl}{1\ mol\, Ca}$

   e. Formula is $CaCl_2$

N A M E

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 Chloride, (g)	

RESULTS

Mass percent Cl in your product:

Mass percent Mn in your product:

Moles Cl in 100 g of your product

Moles Mn in 100 g of your product

Molar ratio of Cl to Mn

Problems

1. Complete the following table:

Formula	Moles of each kind of element
H_2O	
$CuCl_2$	
Al_2S_3	
$Ba(NO_3)_2$	
C_4H_{10}	

2. 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	

3. 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?

4. 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?

EXPERIMENT TEN
CHEMICAL REACTIONS AND EQUATIONS

____

N A M E

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.**

<u>Copper</u>

Observation____

Balanced Chemical Equation

____

<u>Zinc</u>

Observation____

Balanced Chemical Equation

____

<u>Magnesium</u>

Observation____

Balanced Chemical Equation

____

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

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

B. Zinc and Copper(II) Sulfate

1. Place **<u>approximately</u>** 1 mL (20 drops) of 0.5 <u>M</u> Copper(II) Sulfate solution in a test tube.

2. Place a small piece of Zinc into the solution. After about ten minutes compare your test tube with a reference test tube provided in the lab containing the original Copper (II) Sulfate solution without any Zinc. Record the color of the Copper(II) Sulfate solution in your tube with the Zinc and compare it with the reference test tube.

3. Stir the solution containing the Zinc every 15-20 minutes and compare your tube with the lab standard tube every 15 minutes for 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.**

Zinc and Copper(II) Sulfate

Observations:

10 minutes ____

30 minutes ____

45 minutes ____

Balanced Chemical Equation:

____

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 your medium sized test tubes by rinsing in deionized water.

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 other compound has been formed?

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 other compound has been formed?

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 other compound has been formed?

6. Mix about 1 mL of 0.1-M Lead(II) Nitrate with about 1 mL of 0.1-M Potassium Chromate, K_2CrO_4. Record your observations and write a balanced chemical equation. The solid formed in this reaction is Lead(II) Chromate. What other compound has been formed?

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.

For each of the reactions you have performed, give a short description of the reaction mixture before and after the reaction. Then write a balanced chemical equation describing the reaction.

Calcium Chloride and Sodium Phosphate

Appearance of Reactants before reaction:____

____

Appearance after reaction: ____

____

Balanced Chemical Equation:

____

Barium Chloride and Sodium Sulfate

Appearance of Reactants before reaction:____

____

Appearance after reaction: ____

____

Balanced Chemical Equation:

____

Iron(III) Chloride and Potassium thiocyanate, KSCN

Appearance of Reactants before reaction:____

____

Appearance after reaction: ____

____

Balanced Chemical Equation:

____

Silver Nitrate and Sodium Chloride

Appearance of Reactants before reaction:____

____

Appearance after reaction: ____

____

Balanced Chemical Equation:

____

Lead(II) Nitrate and Potassium Chromate, K_2CrO_4

Appearance of Reactants before reaction:____

____

Appearance after reaction: ____

____

Balanced Chemical Equation:

____

Hydrochloric Acid, HCl, and solid Sodium Carbonate

Appearance of Reactants before reaction:____

____

Appearance after reaction: ____

____

Balanced Chemical Equation:

____

Problems

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.

EXPERIMENT ELEVEN

MEASUREMENT OF THE SPECIFIC HEAT OF A METAL

____
N A M E

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 that does not undergo a phase change, 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:

$$Q = C \times m \times \Delta T$$

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.
The specific heat, C, 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 performed is called a *calorimeter.* The calorimeter that will be used in this experiment consists of a Styrofoam® cup containing a measured 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 in 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, lid 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,

$$Q_{water} = C_{water} \times m_{water} \times \Delta T_{water} = -Q_{metal}$$

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.

$$C_{metal} = -\frac{Q_{water}}{m_{metal}\Delta T_{metal}} \quad \text{since } Q_{metal} = C_{metal} \times m_{metal} \times \Delta T_{metal}$$

1. Obtain a Styrofoam® cup, and lid. Obtain a cylinder and record the identity of the metal of your cylinder. **DO NOT 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. Measure the initial temperature of the water in the Styrofoam® calorimeter cup with the same thermometer.

5. Fill your largest Erlenmeyer flask about 2/3 full of water and place it on a hot plate. Turn the hot plate heat setting to about ½ full heating.

6. Use a 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.

7. 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.

8. After the water has been boiling for about five minutes, 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.**

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 in 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 cal/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 1.00 cal/g □C. Average the results from each of the trials to obtain the experimentally determined specific heat of your metal.

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

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 (Temperature of the boiling water)		
Initial temperature of water in cup,		
Final temperature of water and metal		
Specific heat of metal, (cal/g^{o} C)		
Average specific heat of metal, (cal/g^{o} C)		

Calculations:

$\Delta T_{water} = T_f - T_i$ water = ____ºC

$\Delta T_{metal} = T_f - T_i$ metal = _____^{o}C

Q_{water} = (1.00 cal/g^{o}C) x (_____g H_2O) x (______^{o}C) =__________cal

$Q_{metal} = - Q_{water}$

$C_{metal} = Q_{metal} / m_{metal} \times \Delta T_{metal}$ = – _______cal / [(______g) x (_____ºC)] = ________cal/gºC

Show your calculation for the second trial below.

Average specific heat:

[____________cal/g ^{o}C + ___________cal/g ^{o}C] / 2 =___________cal/g ^{o}C

Problems

1. Calculate the heat (expressed in calories) required to heat 115 g of water from 15.4 ^{o}C to 91.4 ^{o}C.

2. Calculate the heat (in calories) lost by 115 g of water as it cools from 91.4 ^{o}C to 15.4 ^{o}C.

3.Calculate the temperature change caused by the absorption of 3.85 kcal heat by 75.4 g water.

4. Calculate the final temperature of 75.4 g water originally at 12.6 ^{o}C after it absorbs 3.85 kcal of heat.

5. A 23.9 g piece of metal heated to 97.8 ^{o}C is placed in 52.4 g water at 21.9 ^{o}C. After the metal is added, the temperature of the water rises to 29.9 ^{o}C. Calculate the specific heat of the metal. Express your answer in the units of cal/g^{o}C.

EXPERIMENT TWELVE

THE STRUCTURE OF MOLECULAR COMPOUNDS

A. LEWIS STRUCTURES OF COVALENT COMPOUNDS

Drawing Lewis Electron Dot structures helps visualize the chemical bonds that are present in molecular compounds. The formation of a chemical bond is the result of the tendency of atoms to gain or lose valence electrons until they have obtained an electron configuration that is the same as that of a noble gas. The valence electrons are the electrons in the outermost electron shell (Table 1). Except for Helium, which has 2 valence electrons, the noble gases all have 8 valence electrons (an octet). Except for Hydrogen (which would prefer to be associated with 2 valence electrons like Helium), many nonmetal atoms share electrons in a manner that results in their being associated with 8 valence electrons (the 'octet rule'). We shall only examine molecules that obey the octet rule in this exercise.

TABLE 1

Numbers of bonds and nonbonding pairs for nonmetals obeying the octet rule and having no formal charge.

	Group IVA	Group VA	Group VIA	Group VIIA
	C, Si	N, P, As	O, S, Se, Te	F, Cl, Br, I
Valence electrons	4	5	6	7
Bonds	4	3	2	1
Nonbonding pairs	0	1	2	3

Lewis dot structures of a single atom show the valence electrons of that atom as dots. Generally only two dots are shown on any one side of the atom. Atoms forming covalent bonds in a molecule share their dots to form a bond—a single bond is formed from one dot (electron) from each atom and a double bond is formed when two atoms each share two of their dots.

We draw 'Lewis Structures' showing all the valence electrons in a molecule. One can approach drawing Lewis structures as a puzzle—arrange the total number of valence electrons contributed by each atom of the molecule in such a way that each atom has eight electrons around it. The electrons around each atom include both the bonding electrons (6 electrons for a triple bond, 4 for a double and, 2 for a single bond) and any nonbonding pairs. Many students prefer to remember the information provided in Table 1—for example Group IV A elements form four bonds and have no nonbonding pairs, Group V A form 3 bonds and one nonbonding pair and so forth. The numbers of bonds formed and nonbonding pairs present always obeys the rules given in Table 1 provided the atoms have no charges.

1. Determine the number of valence electrons in each of the atoms listed below.

    a. C

    b. O

    c. N

    d. F

2. Draw Lewis dot structures for the following atoms.

    a. Si

    b. C

    c. N

    d. As

    e. S

    f. Br

3. Draw electron dot structures for the following molecules:

   a. SiO_2

   b. F_2CO (carbon is central)

   c. PH_3

   d. CO_2

   e. NH_3

f. CH_4

g. SCl_2

B. SHAPES OF MOLECULES

Most simple covalent molecules have a central atom. In this exercise you will determine the distribution of (a) **nonbonding electrons pairs** and (b) **groups** of electrons involved in covalent bonds about the central atom. The total number of nonbonding pairs and bonding groups determine the electron pair geometry (how the electron pairs are arranged in space). Note that a single bond, double bond or triple bond count as a SINGLE **group** of bonding electrons . Also note that nonbonding pairs count as a distinct pair of electrons when determining the electron geometry.

Once the electron geometry has been determined, we can predict the shape of the molecule. The VSEPR model predicts that the total number of electron groups around a central atom (both bonding groups and nonbonding pairs) determine the shape of the molecule. Table 2 summarizes the VSEPR rules.

TABLE 2

Electron geometries and molecular shapes for octet obeying central atoms

Number of electron groups	Electron geometry	Molecular shape
2 groups	linear	linear
3 groups	trigonal planar	trigonal planar
4 groups	tetrahedral	tetrahedral
3 groups + 1 non bonding pair	tetrahedral	pyramidal
2 groups + 2 non bonding pairs	tetrahedral	bent

After drawing the Lewis structure in the large box, complete the table:

CBr_4	Number of nonbonding pairs:
	Number of bonding groups:
	Total number of electron groups:
	Electron Geometry
	Molecular shape
NBr_3	Number of nonbonding pairs:
	Number of bonding groups :
	Total number of electron groups:
	Electron Geometry
	Molecular shape
H_2O	Number of nonbonding pairs:
	Number of bonding groups:
	Total number of electron groups:
	Electron Geometry
	Molecular shape
CO_2	Number of nonbonding pairs:
	Number of bonding groups:
	Total number of electron groups:
	Electron Geometry
	Molecular shape

CH_2O (C is central atom)	Number of nonbonding pairs:
	Number of bonding groups:
	Total number of electron groups:
	Electron Geometry
	Molecular shape
HCN (C is central atom)	Number of nonbonding pairs:
	Number of bonding groups:
	Total number of electron groups:
	Electron Geometry
	Molecular shape
CO_2	Number of nonbonding pairs:
	Number of bonding groups:
	Total number of electron groups:
	Electron Geometry
	Molecular shape
PCl_3^-	Number of nonbonding pairs:
	Number of bonding groups:
	Total number of electron groups:
	Electron Geometry
	Molecular shape

EXPERIMENT THIRTEEN

GAY-LUSSAC'S LAW

Gases are often involved as reactants or products in chemical reactions. 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 volumes 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 at the same temperature and pressure. It is important to remember that these volume relationships apply only to **gases** at the same temperature and pressure. The volumes of liquids and solids do not show this type of relationship.

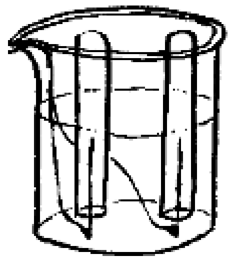
Electrolysis is a process by which electrical current (electron flow) can be used to cause a non-spontaneous chemical reaction to occur. During electrolysis, electrons are removed from one reactant (at the anode) and are donated to another reactant (at the cathode). Electrolysis of water causes the water to break down into its elements, oxygen and hydrogen, both of which are diatomic gases at room temperature and normal pressure. 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.

A. Electrolysis of Water (This experiment may be performed as a demonstration by your professor)

1. Each group will check out one set of electrodes from the instructor.

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 two of your smallest test tubes. Slip one rubber band onto 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 solution off your hands.



4. **BE VERY CAREFUL WITH THE ELECTRODES. THEY CAN BE EASILY DAMAGED. DO <u>NOT</u> 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).

6. After the electrolysis has run the allotted time, **UNPLUG THE TRANSFORMER. <u>DO NOT</u> 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 the splint). A second person will put their finger over the end of the tube **THAT APPEARS TO CONTAIN THE LARGER AMOUNT OF GAS** (try not to lose any gas and **<u>do 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” which contains gas “A”.

9. Remove the second tube (the one **THAT APPEARS TO CONTAIN THE SMALLER AMOUNT OF GAS** in the same manner, keeping a finger over the end to prevent the loss of the gas. For this tube, 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” which contains gas “B”.

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

11. Determine the volumes of gases 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 larger volume (Gas A) by the smaller volume (Gas B) to obtain a ratio of gas volumes.

13. Use the correct formula for water to write a balanced chemical equation for the electrolysis of water to form Hydrogen and Oxygen gases.

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 your instructor.

N A M E

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	
Gas A/Gas B Ratio (use correct sig. figs.)	
Balanced Equation for the Electrolysis of Water:	

Problems

1. 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 remain constant.

2. Calculate what will happen to a 10.0 L balloon at 1 atm pressure and 25^oC if the following changes are made. If you cannot predict a change, briefly explain.

    a. The temperature is doubled from 25 ^oC to 50^oC with no change in pressure

    b. The pressure is doubled from 1 atm to 2 atm but the temperature remains at 25^oC

    c. The pressure and temperature remain at 1 atm and 25^oC but the number of moles of gas in the balloon is doubled.

    d. The pressure and temperature remain at 1 atm and 25^oC but the mass of gas in the balloon is doubled.

EXPERIMENT FOURTEEN

DETERMINATION OF THE CONCENTRATION OF A SOLUTION

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 number of mass units of solute per 100 mass units of solution. The mass units we use in chemistry are typically grams. The **mass/volume percent** concentration of a solution states the number of grams of solute present in 100 milliliters of the solution. The **molarity** of the solution describes the moles of solute per liter of solution.

A. Concentration of a Sodium Chloride solution in expressed as M, %(m/v), and (%m/m).

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

2. Obtain about 6 mL of the unknown Sodium Chloride solution in the 10 mL graduated cylinder. Read and record the volume (watch sig. figs.) of the solution in the cylinder.

3. 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.

4. Determine the mass of the evaporating dish and solution combined by weighing it. Calculate the mass of the solution in the dish.

5. Fill a 400 mL beaker about two thirds full of water and add a boiling chip or piece of broken crucible. Set the beaker on a hot plate. 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. The boiling chip or broken crucible prevents the water from boiling over or ‘bumping’. Replenish the water in the **BEAKER** if needed.

6. 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 hot plate. Gently heat the dish with a low setting on the hot plate for about 5 minutes.

7. Allow the dish to cool for **at least** 10 minutes on the benchtop.

8. Determine the mass of the evaporating dish and the dried Sodium Chloride Salt.

9. Calculate the % (m/m), % (m/v), molarity, and density of the Sodium Chloride solution.

10. Be careful to include all units when performing the calculations. For example, when you calculate the density of a solution, your ratio is the mass of s**olution** divided by the volume of solution—**NOT** the mass of solute divided by volume of solution.

____
N A M E

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 , (g)	
Mass of dried NaCl, (g)	
Density of solution, (g/mL)	
Percent NaCl (m/m), (%)	
Percent NaCl (m/v), (%)	
Molarity of NaCl, (moles/L)	

Calculations:

Density of solution: $\dfrac{\underline{\qquad\qquad}\ \textit{g solution}}{\underline{\qquad\qquad}\ \textit{mL solution}}$ = _______g/ mL

% (m/m) = ($\dfrac{\underline{\qquad\qquad}\ \textit{g solute}}{\underline{\qquad\qquad}\ \textit{g solution}}$) x 100 =___________% (m/m)

%(m/v) = ($\dfrac{\underline{\qquad\qquad}\ \textit{g solute}}{\underline{\qquad\qquad}\ \textit{mL solution}}$) x 100 =__________% (m/v)

Molarity. This is a two step calculation. First you must find the number of moles of solute in your solution. Then divide the number of moles solute by the volume of solution **expressed in L**.

$$(\_\_\_\_\_\_\_\_\text{g solute}) \times \left(\frac{1\ mol\ NaCl}{\_\_\_\_\_\_\_\_\ g\ NaCl}\right) = \_\_\_\_\_\_\text{mol NaCl}$$

$$M = \left(\frac{\_\_\_\_\_\_\_\_\ mol\ NaCl}{\_\_\_\_\_\_\_\_\ L\ solution}\right) = \_\_\_\_\_\_\_\_\ M$$

An alternate solution allows you to perform the calculation of molarity in a single step by starting with the ratio of grams of solute to the volume of solution. You convert this experimentally determined ratio to the number of moles NaCl per L solution.

$$\frac{\_\_\_\_\_\_\_\_\ g\ solute}{\_\_\_\_\_\_\_\_\ mL\ solution} \quad \frac{1\ mol\ NaCl}{\_\_\_\_\_\_\_\_\ g\ NaCl} \times \frac{1000\ mL}{1\ L}) = \frac{\_\_\_\_\_\_\ mol\ NaCl}{1\ L\ solution} =$$

____________M

Problems

1. Use the density of the NaCl solution that you determined to convert the % (m/m) concentration directly into an equivalent % (m/v) concentration.

2. 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 Potassium Nitrate that remains 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

APPENDIX 1

Names and Symbols of Some Commonly Encountered Elements

Aluminum	Al	Fluorine	F	Oxygen	O
Argon	Ar	Gold	Au	Phosphorus	P
Arsenic	As	Helium	He	Platinum	Pt
Barium	Ba	Hydrogen	H	Potassium	K
Beryllium	Be	Iodine	I	Radium	Ra
Boron	B	Iron	Fe	Silicon	Si
Bromine	Br	Lead	Pb	Silver	Ag
Cadmium	Cd	Lithium	Li	Sodium	Na
Calcium	Ca	Magnesium	Mg	Strontium	Sr
Carbon	C	Manganese	Mn	Sulfur	S
Chlorine	Cl	Mercury	Hg	Tin	Sn
Chromium	Cr	Neon	Ne	Titanium	Ti
Cobalt	Co	Nickel	Ni	Uranium	U
Copper	Cu	Nitrogen	N	Zinc	Zn

Elements which are always written as diatomic molecules:

H_2 N_2 O_2 and the halogens (F_2 Cl_2 Br_2 and I_2)

Metals which always have a fixed charge despite not being in Groups IA or IIA

Al^{3+} Zn^{2+} Cd^{2+} Ag^{+}

Names and Formulas of Some Common Polyatomic Ions

Acetate	$C_2H_3O_2^{1-}$	(also designated as CH_3COO^{1-})	
Ammonium	NH_4^{1+}		
Carbonate	CO_3^{2-}	Hydrogen Carbonate (aka Bicarbonate)	HCO_3^{1-}
Chlorate	ClO_3^{1-}	Perchlorate	ClO_4^{1-}
Chlorite	ClO_2^{1-}	Hypochlorite	ClO^{1-}
Cyanide	CN^{1-}		
Hydroxide	OH^{1-}		
Nitrate	NO_3^{1-}		
Nitrite	NO_2^{1-}		
Phosphate	PO_4^{3-}	Hydrogen Phosphate	HPO_4^{2-}
		Dihydrogen Phosphate	$H_2PO_4^{1-}$
Phosphite	PO_3^{3-}	Hydrogen Phosphite	HPO_3^{2-}
		Dihydrogen Phosphite	$H_2PO_3^{1-}$
Sulfate	SO_4^{2-}	Hydrogen Sulfate (aka Bisulfate)	HSO_4^{1-}
Sulfite	SO_3^{2-}	Hydrogen Sulfite (aka Bisulfite)	HSO_3^{1-}