CHEM30A

Inorganic Chemistry for Health Occupations

4 units
3 hours Lecture
3 hours Laboratory

Laboratory Manual

Table of Contents

Activity 1 - Math and the Calculator

Goals

- ❑ State the correct number of significant figures in a measurement.
- ❑ Round off a calculated answer to the correct number of significant figures.
- ❑ Give final answers with the correct number of significant figures.
- ❑ Write a number in scientific notation.

Pre-lab Questions *(answer these on a separate sheet using complete sentences)*

1. How does a measured number differ from an exact number?
2. Why does a measurement have a limited number of significant figures?
3. How do you determine the number of significant figures for a calculated answer?
4. What are the rules for rounding off numbers?

Concepts to Review

Significant figures
Exact numbers
Mathematical operations with significant figures
Scientific notation
Calculator use

Introduction

In the sciences, matter is characterized using a wide variety of numbers. Measured numbers are different from counted or defined numbers. Any budding chemist must learn how to distinguish measurements from exact numbers. Measurements may be used in calculations and the calculated numbers must accurately reflect the quality of the measured number used. The following information and exercises will help prepare you for the mathematics involved in learning chemistry.

Required Materials

Scientific calculator

A. Measured and Exact Numbers

Suppose you used a bathroom scale this morning to determine that your weight is 145 lb. The scale is a measuring device and the number in 145 lb is called a *measured number*. When numbers are obtained by counting objects or using a definition, they are called *exact numbers*. Suppose you counted 22 people in your lab. The number 22 is an exact number because you did not need to use a measuring device. The relationships between units within the metric system (S.I.) or within the American system are defined numbers, which makes them exact numbers as well. For example, the numbers in definitions such as 100 cm in 1 meter and 12 inches in 1 foot are exact. See Example 1.

Example 1.

Describe each of the following as a measured or exact number.

a. 14 inches
b. 14 pencils
c. 60 minutes in 1 hour
d. 7.5 kg

Solution:

a. measured
b. exact (counted number)
c. exact (definition)
d. measured

B. Scientific Notation

In scientific work, small numbers such as 0.000000025 m and large numbers such as 4,000,000 g are often expressed using powers of 10; the above numbers are 2.5×10^{-8} m and 4×10^{6} g. The values 2.5 and 4 are *coefficients*; the values 10^{-8} and 10^{6} are powers of ten (See Table 1 and Example 2). The rules for converting standard numbers to scientific notation are given below.

Table 1. Examples of exponential notation

Standard number		Power of ten	Standard number		Power of ten
10,000	=	10^4	0.1	=	10^{-1}
1000	=	10^3	0.01	=	10^{-2}
100	=	10^2	0.001	=	10^{-3}
10	=	10^1	0.0001	=	10^{-4}

Rules for Writing Numbers in Scientific Notation

For numbers larger than 10:

a. Move the decimal point to the left until it follows the first digit in the number.

b. Write a power of ten that is equal to the number of places the decimal point was moved to the left.

For numbers smaller than 1:

a. Move the decimal point to the right until it is located after the first digit in the number.

b. Write a negative power of ten that is equal to the number of places the decimal point was moved to the right.

Example 2.

Write the following standard numbers in scientific notation:

a. 35,000 b. 608 c. 0.0000815

Solution:

					coefficient		power of ten
a.	35,000	=	3 <u>5 0 0 0 .</u>	=	3.5	×	10^4
b.	608	=	6 <u>0 8 .</u>	=	6.08	×	10^2
c.	0.0000815	=	0 <u>. 0 0 0 0 8</u> 1 5	=	8.15	×	10^{-5}

C. Significant Figures

In measured numbers, all the reported figures are called *significant figures*. The first significant figure is the first nonzero digit. The last significant figure is always the estimated digit. Zeros between other digits or at the end of the decimal part of a number (i.e., trailing zeros in a number with a decimal part) are counted as significant figures. Leading zeros after the decimal point are *not significant;* they are placeholders. Trailing zeros are not *significant* in numbers equal to or greater than 10 if there is no decimal point present; they are placeholders needed to express the magnitude of the number.

In a number written in scientific notation, *all* the figures in the coefficient *are significant*. Examples of counting significant figures in measured numbers are given in Table 3.2 and Example 3.3.

Table 2. Examples of Counting Significant Figures.

Measurement	Significant Figures	Reason
455.2 cm	4	All nonzero digits are significant
0.800 m	3	The trailing zeros in the decimal part are significant
50.2 L	3	A zero between nonzero digits is significant
0.0005 lb	1	Leading zeros are not significant
25,000 ft	2	Placeholder zeros are not significant
3.20×10^4 g	3	All the digits in a coefficient are significant

Example 3.

State the number of significant figures in each of the following measured numbers:

a. 0.00580 m b. 132.08 g c. 1.5×10^3 mL

Solution:

a. Three significant figures. The zeros immediately after the decimal point are placeholder zeros; the last zero is a trailing zero in the decimal part.

b. Five significant figures. The zero between nonzero digits is significant.

c. Two significant figures. All the digits in the *coefficient* of a number in scientific notation are significant.

D. Rounding Off

Often, you will use a measurement in a mathematical operation such as multiplication, division, addition or subtraction. When the calculator display shows more numbers than the measurements support, it is necessary to *round off* the calculated answer. If the numbers to be dropped begin with a number **less than 5,** they are simply dropped. However, if the numbers dropped begin with **a number greater than 5,** the value of the last *retained digit* is increased by 1. If the number to be rounded is **exactly 5** (all figures following 5, if any, are 0)**,** round an odd number to the next highest even number and leave an even number unrounded. For a calculator display already in scientific notation, round off the coefficient to report the correct number of significant figures (sig figs).

On your calculator, an answer may appear in scientific notation, which means that a coefficient and a power of ten are shown. (See Table 3 and Example 4.) In scientific notation, the correct number of sig figs is shown in the coefficient. Since calculators know nothing of sig figs, the numbers shown in a calculator display before the power of ten must be rounded to the correct number of sig figs. Be *sure* to keep the power of ten when recording your final answers!

Please be advised that Table 3 may not represent the exact output of your calculator. Now is the time to make yourself familiar with your calculator, especially regarding how to perform mathematical operations with exponential numbers. *See your instructor before your first exam if you are not successful in mastering this task.*

The exponential function key is usually represented by an “EE” or “EXP” key and acts as “× 10”. Do not use the exponent key **and then** multiply by 10. Only the exponent key is necessary. The consequence of making this mistake is that your results will be consistently too high by a power of ten.

Table 3. Examples of Writing Calculator Results in Scientific Notation.

Calculator display	Number of significant figures to be shown in coefficient	Rounded and written in scientific notation
2.512^{05} or 2.152E5	2	2.5×10^{5}
4.1585^{12} or 4.1585E12	3	4.16×10^{12}
8.775^{-08} or 8.775E-8	3	8.78×10^{-8}

Example 4

Round off each of the following calculator displays to report answers with three significant figures as well as two significant figures.

a. 75.6243 b. 0.528392 c. 387,600 d. 8.027^{-04} (displayed on some as 8.027E-4)

Solution:

	Three significant figures	Two significant figures
75.6243	75.6	76
0.528392	0.528	0.53
387,600	388,000	390,000
8.027^{-04}	8.03×10^{-4}	8.0×10^{-4}

E. Multiplication and Division of Measured Numbers

If a calculated answer is obtained from multiplication and/or division, it is rounded off to the same number of significant figures as the measured number with the *fewest* significant figures. See Examples 5 and 6.

Example 5

Solve: $\dfrac{(0.025\text{ m})(4.62\text{ g})}{3.44\text{ s}} =$

Solution: On the calculator, the steps are:

Enter keys	Display reads	
0.025	0.025	two significant figures
×	0.025	
4.62	4.62	three significant figures
=	0.1155	
	0.1155	(“Enter keys” left blank since no division sign on computer keyboard)
3.44	3.44	three significant figures
=	0.033575581	Calculator display to 9 decimal places
	$0.034 \dfrac{\text{m} \times \text{g}}{\text{s}}$	final answer rounded to **two** significant figures with the correct units.

Example 6

Solve: $\dfrac{3.4 \times 10^{-4}\text{ g}}{2.75 \times 10^{8}\text{ mL}} =$

Solution: On the calculator, the steps are:

Enter keys	Display reads	
3.4	3.4	two significant figures
EXP (or EE)	3.4^{00} or 3.4E	
4	3.4^{04} or 3.4E4	
(-) or (±)	3.4^{-04} or 3.4E-4	
	3.4^{-04} or 3.4E-4	“Enter keys” left blank as above
2.75	2.75	
EXP (or EE)	2.75^{00} or 2.75E	
8	2.75^{08} or 2.75E8	
=	1.2363636^{-12} or 1.2363636E-12	calculator display to 7 decimal places plus two digit exponent
	$1.2 \times 10^{-12} \dfrac{\text{g}}{\text{mL}}$	Coefficient in **final answer** rounded to *two* significant figures; units must be included in your written answer.

F. Addition and Subtraction of Measured Numbers

After you have added or subtracted measured numbers, you may need to round off the result. An answer from addition or subtraction has the last significant figure in the column where **all** of the numbers added or subtracted **also** have significant figures. See Examples 7, 8 and 9.

Example 7

Addition: 42.11 cm + 4.056 cm + 30.1 cm =

Solution:

	42.11	
+	4.056	
+	30.1	digits in 30.1 end at the tenth's place; all other numbers go further than this.
	76.266	calculator display
	76.3 cm	correct answer rounded to give digit in tenth's place accompanied by proper units.

Example 8

Subtraction: 14.621 g - 3.39 g =

Solution:

	14.621	
-	3.39	digits in 3.39 end at the hundredth's place; all others go further.
	11.231	calculator display
	11.23 g	correct answer rounded to give final significant digit in the hundredth's place, with units included.

Example 9

Addition: 1200 m + 14 m + 1.11 m =

Solution:

	1200	digits in 1200 end at the hundred's place; all others go further
+	14	
+	1.11	
	1215.11	calculator display
	1200 m	correct answer rounded to give the last significant digit in the hundred's place, with units included.

Activity 1 - Math and the Calculator Worksheet

Name____________________________

Section____________ **Date**_______

Exercise A. Measured and Exact Numbers

Circle "M" or "E" to indicate whether each of the following numbers is measured or exact:

5 books	M E	12 roses	M E
5 lb	M E	12 inches in 1 foot	M E
9.25 g	M E	361 miles	M E
0.035 kg	M E	100 cm in 1 m	M E

Exercise B. Scientific Notation

Write the following numbers in scientific notation:

4,450,000	___________________	0.00032	_____________
38,000	___________________	25.2	_____________
0.0000000021	___________________	0.0505	_____________

Write the following as standard numbers:

4×10^{2}	____________________	3×10^{-4}	_____________
5×10^{3}	____________________	8.2×10^{-3}	_____________
3.15×10^{5}	____________________	2.46×10^{-6}	_____________

Exercise C. Significant Figures

State the number of significant figures in each of the following measured quantities:

4.5 m	____________________	204.52 g	_____________
0.0004 L	____________________	625,000 mm	_____________
850 lb	____________________	34.80 km	_____________
2.50×10^{-3} L	____________________	8×10^{5} g	_____________

Exercise D. Rounding Off

Round off each of the following to the number of significant figures indicated. Don't forget placeholder zeros when necessary!

	Three significant Figures	Two Significant Figures
e.g.: 0.4108 g	0.411 g	0.41 g
143.63212 mi	____________________	__________________
532,800 ft	____________________	__________________
5.448×10^2 yrs	____________________	__________________
0.00858345 mm	____________________	__________________

Exercise E. Multiplication and Division

Do the following multiplication and division calculations. **Give a final answer with the correct number of significant figures and the correct units. Units can cancel or multiply just like number factors**.

e.g.: $4.5 \text{ ergs} \times 0.281 \text{ in} =$ 1.3 erg·in

$0.1184 \text{ cm} \times 8.00 \text{ cm} \times 0.0345 \text{ cm} =$ __________________

$(2.5 \times 10^4 \text{ m/s}) \times (5.0 \times 10^{-7} \text{ s}) =$ __________________

$$\frac{(42.2 \text{ L})(1.45 \text{ atm})}{(4.8 \text{ mol})(0.0821 \ \frac{\text{L atm}}{\text{mol K}})} =$$ __________________

Exercise F. Addition and Subtraction

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

13.45 mL + 0.4552 mL = ___________________

145.5 m + 86.58 m + 1045 m = ___________________

1315 mi + 200 mi + 1100 mi = ___________________

245.625 g - 80.2 g = ___________________

4.62 cm - 0.885 cm = ___________________

Questions and Problems

1. How can you distinguish an exact number from a measured number?

2. In the scientific community, the last digit in a measured number that is still significant is sometimes called the estimated digit. Circle or underline the estimated digit in each of the following measurements:

   1.5 cm   4500 mi.   0.0782 m   42.50 g

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

   a. Give the most likely explanation for the difference between the two calculated answers despite the fact that both students used the same measurements.

   b. You are going to tutor Bill. Using complete sentences, describe how you would help him to understand why his answer is wrong and Beverly's is right.

Measurements in Your Daily Life

4. Throughout a typical day, list at least eight numbers (and units) you might use such as measurements, prices, definitions, cooking quantities, gasoline purchases, prescription dosages, etc. Identify each number as exact or measured. Explain your choice. (Did you use a measuring tool, or did you count out something, or use a definition?)

Number used	Type of Number	Explanation
________________	___________________	____________________________________
________________	___________________	____________________________________
________________	___________________	____________________________________
________________	___________________	____________________________________
________________	___________________	____________________________________
________________	___________________	____________________________________
________________	___________________	____________________________________
________________	___________________	____________________________________

5. In the list above, were the numbers you used mostly measured numbers, or mostly exact numbers?

6. List the names and abbreviations of five units (metric or American) of the measurement you used. Give the property measured (weight, mass, volume, distance, etc.)

Unit of Measurement	Abbreviation	Property Measured
__________________	___________	________________
__________________	___________	________________
__________________	___________	________________
__________________	___________	________________
__________________	___________	________________

Activity 2 – Dimensional Analysis

Goals

- Develop conversion factors from common equalities.
- Use conversion factors to convert between different units of measure.
- Apply the concept of dimensional analysis to string together conversion factors forming mathematical expressions.

Concepts to Review

Conversion factors
Dimensional analysis
Units of measure
Mathematical operations with significant figures
Scientific notation
Calculator use

Introduction

The following problems involve converting from one unit to another. If the unit of measure changes, the number in front of the unit will also change. All of these calculations require two things: defined relationships between different units and some method to convert one into another. *Dimensional analysis* involves setting up the problem in such a way as to cancel the unwanted units with only the desired units remaining. A consistent method of problem solving including unit cancellation is invaluable in succeeding in this course.

Concepts to Review: English units, Metric units, Conversion factor, temperature (scales, conversions between scales), Exact vs. Measured Numbers, Significant figures in Calculations involving Measurements, Density vs. Specific Gravity.

Use your text to find other common equalities (Table 1.3) and the relationships needed to convert temperatures. Also look for similar problems to help you learn how to successfully set up these problems.

Required Materials

Scientific calculator, selection of commercial products

Activity 2 – Dimensional Analysis Worksheet

Name____________________________

Section____________ **Date**______

Exercise – Measured and Exact Numbers

Look at the commercial products in lab or at your home and create conversion factors (ratios) for fluid ounces to liters or milliliters. To do this, read the amounts present in both English units (fluid ounces, pints, quarts, pounds, ounces, etc.) and metric units (milliliters, liters, grams, kilograms, etc.) and then divide one value by the other, keeping the units.

Product	*English Measurement*	*Metric Measurement*	*Conversion Factors*
Coca Cola	12 fl oz	355 mL	$\frac{355\ \text{mL}}{12\ \text{fl oz}}$ or $\frac{12\ \text{fl oz}}{355\ \text{mL}}$

Questions and Problems – Dimensional Analysis

For all of the following calculations, show all your work. If you need more space than is provided, you may do your work on separate pages but be sure to attach these when you are done. Write your correct answers in the space provided. If you only provide the answers without showing your work, you will not be given full credit. All answers should be clearly identified (boxed off), written in scientific notation if the values are less than 1 or greater than 1000, and all should have the correct number of significant figures.

1. How many seconds are in exactly one day to five significant figures?

2. Seventeen apples weigh 3.25 pounds. They cost 59 cents for 1 pound. What is the cost of 8.95×10^3 apples?

3. The distance from Santa Cruz to Santa Barbara is about 280 miles. If a car gets 23.6 miles per gallon, and the price of gas is 32 cents per liter, how much will it cost for gas to drive from Santa Cruz to Santa Barbara?

4. It is found that 5 pears weigh an average of 1.9 pounds. A box of pears cost $7.94. The price per pound is 55 cents. How many pears are in exactly 7 boxes?

5. During surgery a patient receives 5.0-pts of plasma. (1 quart = 2 pints, 1 liter = 1.057 qt)

    a. How many milliliters of plasma were given?

    b. How many dL were given?

6. How many cubic meters of soil are needed to fill a flower box that is 3.5 feet long, 8 inches wide and 1 foot deep?

7. Body temperatures above 41.1^{o}C can lead to convulsions, especially in children.

   a. What is this temperature in ^{o}F?

   b. What is this temperature in K?

8. The daily dose of ampicillin for the treatment of an ear infection is 115 mg ampicillin per kg of body weight. The pill is dispensed in 500. mg tablets. How many tablets should be given daily for a 75 pound child? An IV pump delivers medication at a constant rate of 24 mg/hr. How long does it take to deliver 9.0×10^{1} mg?

9. The volume of blood plasma in adults is 3.1 L. The density of blood plasma is 1.03 g/cc. How many pounds of blood plasma are there in the average adult body? (Hint: You can use the density as a conversion factor.)

10. Which is the higher temperature, 18 ^{o}F or -1.0 ^{o}C?

11. A bottle of Cabernet Sauvignon is labeled as having an alcohol content of 12.5% by volume.

    a. Write the percentage of the alcohol in the wine as a conversion factor.

    b. If an individual were to consume 320. mL of the wine, how many fluid ounces of pure alcohol would the individual have ingested? (1 pint = 16 ounces; 8 pints = 1 gal)

12. Urine is a water-based solution containing a variety of dissolved solids. The specific gravity of a urine sample of a young wrestler is 1.045, which is outside the normal range of 1.003 – 1.030. (The specific gravity of a substance is its density divided by the density of water at $4^{O}C$, at which the assumption stated below is accurate.)

    a. What is the density (d) of the urine sample? (Assume that $d(H_2O) = 1.00$ g/mL)

    b. Is it more likely that the wrestler is dehydrated or that he recently drank a large amount of water? *You will use words for this answer, no calculations necessary*. (Hint: Review the definition of density)

Activity 3 - Atoms and Elements

Goals

- Write the correct symbols or names of some elements.
- Describe some physical properties of the elements you observe.
- Classify an element as a metal or nonmetal from its physical properties or location on the periodic table.
- Write the complete symbol of a nuclide including its mass number and atomic number.
- Use the complete symbol of a nuclide to determine mass number, atomic number, and the number of protons, neutrons and electrons.

Pre-lab Questions *(answer these on a separate sheet using complete sentences)*

1. What is the periodic table?
2. Where are the alkali metals located on the periodic table? Halogens? Noble gases?
3. Describe the three subatomic particles.
4. How are isotopes alike? Different?

Concepts to Review

Names and symbols of the elements
Properties of metals and nonmetals
Periodic table
Atoms
Subatomic particles
Isotopes

Introduction

Primary substances, called elements, build all the materials about you. There are 117 elements currently known (see http://www.webelements.com/). Some look very different from each other, while others look similar. In this experiment, you will describe the physical properties of some elements in a laboratory display. Then you will show the location of the elements on a blank periodic table.

Metals are elements that are usually shiny, or have a metallic luster. They are usually good conductors of heat and electricity, ductile (can be drawn into a wire), and malleable (can be beaten into sheets). Some of the metals such as sodium or calcium may have a white coating of oxide formed by reacting with oxygen in the air; if the metal were cut, you could see the fresh shiny surface underneath. In contrast, nonmetals are not good conductors of heat and electricity, are brittle (not ductile or malleable), and appear dull, not shiny.

Atoms

Atoms are made up of smaller bits of matter called **subatomic particles**. Of these, we will consider the protons, neutrons, and electrons. **Protons** are positively charged particles, **electrons** are negatively charged, and **neutrons** are neutral (have no charge). In an atom, the protons and neutrons are tightly packed in a tiny central body called the **nucleus**. Most of the atom is empty space, which contains fast-moving electrons. Electrons are so small that their mass is nearly negligible compared to the mass of the proton or neutron.

Every atom is identified by its atomic number and mass number. The **atomic number** is equal to the number of protons in the nucleus, which is the same as the nuclear charge. The **mass number** of an atom is the sum of the number of protons and neutrons, and is usually not equal to the exact measured mass of the atom.

atomic number = number of protons (p^+)
mass number = sum of the number of protons and neutrons ($p^+ + n^0$)

In a **neutral atom** (as opposed to an **ion**), the number of electrons is equal to the number of protons. Protons attract electrons because they have opposite charges.

number of protons (#p$^+$) = number of electrons (#e$^-$)

Isotopes

There are different versions of atoms for each of the elements. **Isotopes** are types of atoms of the same element that differ in the number of neutrons they contain. This means that isotopes of an element have the same number of protons, but different mass numbers. The following example represents the symbol of a sulfur nuclide that has 16 protons and 18 neutrons.

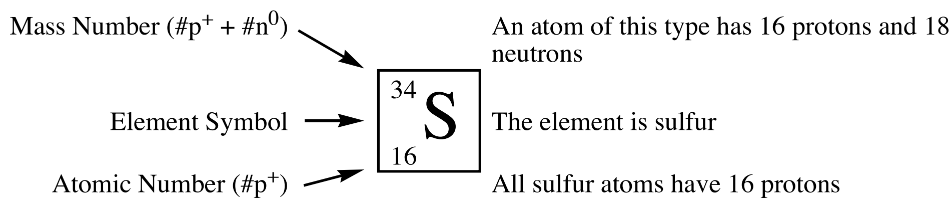


Figure 1. Complete symbol of a nuclide and the meaning of each component.

Periodic Table

The periodic table shown in Figure 1 on the next page contains information about each of the elements. The horizontal rows of the table are called **periods**, and the vertical columns are called **groups**. Each group contains elements that have similar physical and chemical properties. The groups are numbered across the top of the chart. Elements in Group 1A are the **alkali metals**, elements in Group 2A are the **alkaline earths,** and Group 7A contains the **halogens**. Group 8A contains the **noble gases**, which are elements that are generally unreactive with other elements. A double zigzag line that looks like a staircase separates the metals on the left side from the nonmetals on the right side.

1 1A																	18 8A
1 H	2 2A											13 3A	14 4A	15 5A	16 6A	17 7A	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe

Figure 2. The atomic numbers and symbols of some elements on the periodic table.

Materials

A display of samples of elements (metals and nonmetals), colored pencils.

Experimental Procedure

A. Physical Properties of Elements

Observe the elements in the laboratory display. In the worksheet provided, write the symbol and atomic number for each element listed. Describe a selection of physical properties such as color and luster. From your observations, identify each element as a metal, nonmetal, or metalloid.

B. Metals and Nonmetals on the Periodic Table

On the incomplete periodic table in Figure 3, write the symbols of the missing elements. Use your text or a periodic table in the laboratory as a reference. Write the group number at the top of each column of the representative (main group) elements. Using different colors, indicate the alkali metals, alkaline earth metals, halogens and noble gases. Write the period numbers for the horizontal rows shown. Outline and label the areas occupied by the metals, nonmetals, and metalloids.

Activity 3 - Atoms and Elements Worksheet

Name___________________________

Section___________ **Date**______

Exercise A.

Complete the following table using the elements available in the lab:

Element Name (Atomic #)	Elemental Symbol	Physical Properties	Type of Element

After referring to a periodic table, review your answers regarding the "Type of Element" column. Are there any answers you would change based on their position in the periodic table?

Exercise B.

Complete the following periodic table as stated in the instructions to part B.

1																	2
3	4											5	6	7	8	9	10
11	12											13	14	15	16	17	18
19	20	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26	27	28	29	30	31 Ga	32 Ge	33 As	34 Se	35	36
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50	51 Sb	52 Te	53	54 Xe
55 Cs	56	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79	80	81 Tl	82	83 Bi	84 Po	85 At	86 Rn

Figure 3. Metals, Nonmetals, and Metalloids on the periodic table.

Questions and Problems

1. Complete the following table. All atoms are to be considered neutral (not ions).

Element	Atomic Number	Mass Number	Protons	Neutrons	Electrons
Fluorine		19			
Iron				30	
		27			13
			19	20	
Bromine		80			
Gold		197			
			53	74	

2. Complete the following table:

Isotope symbol	Protons	Neutrons	Electrons
${}^{40}_{20}Ca$			
	20	22	20
${}^{40}_{19}K$			
	20	24	20
${}^{46}_{20}Ca$			

3. Complete the list of names of elements and symbols:

Name of Element	Symbol	Name of Element	Symbol
Potassium			Na
Sulfur			P
Nitrogen			Fe
Magnesium			Cl
Copper			Ag

4. On the following list of elements, circle the symbols of the transition metals and underline the symbols of the halogens:

Mg	Cu	Br	Ag	Ni	Cl	Fe	F	Br	B	Na

5. Give the name of each of the following elements:

   a. Noble gas in Period 2 ______________________

   b. Halogen in Period 2 ______________________

   c. Alkali metal in Period 3 ______________________

   d. Halogen in Period 3 ______________________

   e. Alkali metal in Period 4 ______________________

6. From their positions on the periodic table, indicate by circling "M" or "NM" whether the following elements are metals or nonmetals:

Na	M NM	S	M NM	Cu	M NM
F	M NM	Fe	M NM	C	M NM
Ca	M NM	O	M NM	Zn	M NM

7. Consider the element chlorine:

   a. Explain how you would determine the number of neutrons in a chlorine atom that has a mass number of 37.

   b. Another atom of chlorine has a mass number of 35. How does it compare to the above atom of chlorine?

   c. Naturally occurring chlorine is a mixture of the above isotopes, with 75.77% ^{35}Cl and 24.23% ^{37}Cl. The atomic masses of these isotopes are 34.9689 and 36.9659, respectively. Calculate the average atomic mass of chlorine from the above data.

8. A neutral atom has a mass number of 80 and 45 neutrons. Write its complete isotope symbol.

9. An atom has eleven more protons and fifteen more neutrons than the atom in Question 8. Write the isotope symbol that describes this atom.

Activity 4 - Writing Formulas and Names

Goals

- ❑ Write the electron dot structure for an atom.
- ❑ Predict the charge of an ion from its electron dot structure.
- ❑ Use the periodic table to determine the ionic charge of a metal or nonmetal ion.
- ❑ Write the correct formula and name of an ionic or covalent compound.
- ❑ Write the correct formula and name of a compound containing polyatomic ions.

Pre-lab Questions *(answer these on a separate sheet using complete sentences)*

1. Where are the valence electrons in an atom located?
2. How is a positive ion formed from an atom? Why is the charge positive?
3. How is a negative ion formed from an atom? Why is the charge negative?
4. How are the group numbers on the periodic table related to the number of valence electrons? To ionic charge?
5. How do subscripts represent the charge balance in polyatomic ions?
6. According to what rubric are electrons shared in covalent compounds (i.e. what does electron sharing accomplish?
7. How do the names of covalent compounds differ from the names of ionic compounds?
8. What are polyatomic ions? How are they named?

Concepts to Review

Electronic structure (energy levels)
Formation of positive and negative ions
Balancing ionic charge
Ionic and covalent compounds
Writing formulas of ionic and covalent compounds
Naming ionic and covalent compound

Introduction

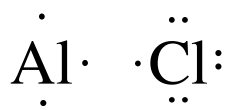
Most of the chemical reactivity of an element is determined by the **valence electrons**, which are the electrons in the highest energy level (or outermost electron shell). Usually in a compound, each atom has an **octet** of electrons (i.e. eight of these) in each of the valence shells. An octet of valence electrons provides atoms with the stable electron configuration found among the noble gases, a group of elements that are particularly stable and inert (unreactive). The first noble gas ($_2He$) does not have an octet since the second electron fills the first (n=1) valence shell, which can accommodate only two electrons.

Required Materials

A Periodic Table of the Elements.

A. Electron Dot Structures

When atoms of metals in groups 1A, 2A or 3A react with atoms of nonmetals in groups 5A, 6A and 7A, the metals lose valence shell electrons and the nonmetals gain valence shell electrons. We can predict the number of electrons lost or gained by analyzing the electron dot structures of the atoms. In an electron dot structure, the valence electrons are represented as dots around the symbol of the atom. For example, aluminum has 13 electrons, 2 in the first energy level, 8 in the second energy level and 3 in the third energy level. To describe this electronic structure we write the **electron arrangement** as 2-8-3. The last number represents the valence electrons so aluminum has three valence electrons and thus an electron dot structure with three dots. Chlorine (electron arrangement 2-8-7) has seven valence electrons and an electron dot structure with seven dots.



Main group metals (group A elements) with 1, 2 or 3 valence electrons *lose* their valence electrons to reach a stable electron configuration with a filled outer shell. For example, an aluminum atom loses its three valence electrons to reach stability and thus acquires an ionic charge of 3+. It is now an aluminum ion with a new electron arrangement of 2-8 (note the complete octet in the outer shell). Positive ions keep the same name as the element.

	Aluminum atom	Aluminum ion	
Symbol	Al	Al^{3+}	
Electron arrangement	2-8-3	2-8-0	(3 electons lost)
Number of protons	13 p^+	13 p^+	(same)
Number of electrons	13 e^-	10 e^-	(3 fewer electrons)
Net ionic charge	0	3+	

When nonmetals with 5, 6 or 7 valence electrons combine with metals, they *gain* electrons to complete their outer shells, and form stable (negatively charged) ions. For example, a chlorine atom gains one valence electron to become stable with an electron arrangement of 2-8-8. With the addition of one electron, chlorine becomes a chloride ion with an ionic charge of 1-. (When two elements combine to form a binary compound called a salt, the name of the negative ion ends in ***-ide***.)

	Chlorine atom	Chloride ion	
Symbol	Cl	Cl^-	
Electron arrangement	2-8-7	2-8-8	(electon added)
Number of protons	17 p^+	17 p^+	(same)
Number of electrons	17 e^-	18 e^-	(1 more electron)
Net ionic charge	0	1-	

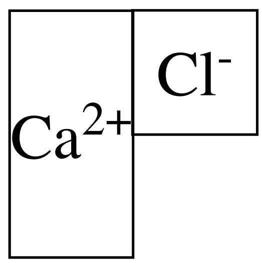
In the worksheet, *write the electron arrangements for atoms and their ions. Write the symbol, ionic charge, and name of each ion.*

B. Writing Ionic Formulas

The group number on the periodic table can be used to determine the ionic charges of elements in each family. **Nonmetals form ions only if they combine with a metal; if they combine with another non-metal, they form covalent (non-ionic) compounds.**

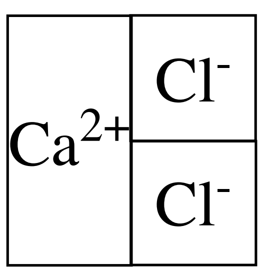
Group number	1A	2A	3A	4A	5A	6A	7A
Valence electrons	$1e^-$	$2e^-$	$3e^-$	$4e^-$	$5e^-$	$6e^-$	$7e^-$
Change	lose	lose	lose	none	gain	gain	gain
Ionic charge	1+	2+	3+	none	3-	2-	1-

In an ionic formula, the **total loss of electrons and total gain of electrons are equal.** The overall net charge is *zero*. This means that the total amount of positive charge must be made equal to the total amount of negative charge. To balance charge, we determine the smallest number of positive and negative ions that give an overall charge of *zero*. We can illustrate the process by showing the ions Ca^{2+} and Cl^- as geometric shapes that represent the amount of ionic charge:



The charge is balanced by using two Cl^- ions to match the shape of the Ca^{2+}. Charge balance occurs with one calcium ion and two chloride ions. This is shown by the subscripts in the formula $CaCl_2$. The subscript "1" for Ca is understood; it is never written. Note that *only the symbols are written in the formula, not the ionic charges.*

Balancing amount of ionic charge | Resulting formula



= $CaCl_2$

In the following worksheet, *write the symbols of the positive and negative ions using the periodic table to determine the charge. Determine the number of each ion that will give a charge balance. Write the correct formula using subscripts to indicate that two or more ions were needed. Write the names of the ionic compounds by placing the metal name first, then the nonmetal name ending in* **-ide**.

C. Ionic Charges for Transition Metals

Most of the transition metals form more than one type of positively charged ion (or cation). We will illustrate this variable valence (combining capacity) with *iron*. Iron forms two cations, one (Fe^{2+}) with a 2+ charge, and another (Fe^{3+}) with a 3+ charge. To distinguish between the two ions, a Roman numeral that gives the ionic charge of that particular ion follows the element name (see below). The Roman numeral is always included in the names of compounds with positive ions that can have variable charge (or oxidation) states. (In an older naming system, the ending *-ous* indicates the lower valence and the ending *-ic* indicates the higher one. Compound names using this system still appear on old reagent bottles and in old chemistry texts.)

Ions	Names	Formula of Compound	Name of Compound
Fe^{2+}	iron (II) ion or ferrous ion	$FeCl_2$	Iron (II) chloride or ferrous chloride
Fe^{3+}	iron (III) ion or ferric ion	$FeCl_3$	Iron (III) chloride or ferric chloride

Among the transition metals a few elements (zinc, silver and cadmium) form only a single type of ion; these have fixed ionic charges and are *not* variable, hence these *do not* use a Roman numeral in their names. Examples are: **Zn^{2+}, zinc ion; Ag^{+}, silver ion; Cd^{2+}, cadmium ion.**

D. Polyatomic Ions

When an ionic compound consists of three or more kinds of atoms, there is a generally a central atom (usually a metal), and a group of attached nonmetal atoms. Such ions are called *polyatomic ions*. A polyatomic ion is a group of covalently bonded atoms with an overall charge that is usually negative. The most common polyatomic ions consist of the nonmetals C, N, S, P, Cl or Br, combined with two to four oxygen atoms. Some examples are given below. The ions are named by replacing the ending of the nonmetal with -**ate** or -**ite**. The form for each central element's most common oxidation state takes the *-ate* ending; the *-ite* ending is for the ion with one less oxygen atom than the *-ate* ion has.
Ammonium ion, NH_4^+, is unusual because it has a positive charge and a metal-like name.

Common polyatomic ions		One oxygen less than common ion	
HO^-	hydroxide ion		
NO_3^-	nitrate ion	NO_2^-	nitrite ion
CO_3^{2-}	carbonate ion		
HCO_3^-	hydrogen carbonate ion or bicarbonate ion		
SO_4^{2-}	sulfate ion	SO_3^{2-}	sulfite ion
HSO_4^-	hydrogen sulfate ion or bisulfate ion	HSO_3^-	hydrogen sulfite ion or bisulfite ion
PO_4^{3-}	phosphate ion	PO_3^{3-}	phosphite ion

Note that there are polyatomic ions that consist of only two different types of atoms (e.g. hydroxide ion; see above) as well as those that have multiple copies of the same atom (e.g. azide ion: N_3^-).

In the worksheet, *write the formulas of compounds that contain ions of transition metals with variable valences. Write a correct name for each compound listed. Be sure to indicate the ionic charge if the transition metal has a variable valence by using a Roman numeral.*
To write the correct formula of a compound with a polyatomic ion, determine the ions required to achieve charge balance just as earlier. When two or more polyatomic ions are needed, enclose the formula of the ion in parentheses, and write the subscript **outside** the parentheses. *No change is made in the formula of the polyatomic ion itself.*

Consider the formula of the compound formed by Ca^{2+} and NO_3^- ions. The ions are $\mathbf{Ca^{2+}}$ and $\mathbf{NO_3^-}$. Since 2 nitrate ions will be needed to balance the charge on the calcium ion, we will need to indicate two NO_3^- ions using parentheses: $\mathbf{Ca(NO_3)_2}$. Note that the formula of the nitrate ion is not changed.

In the worksheet, *determine the positive ions and negative polyatomic ions needed for charge balance. Write the formula using parentheses if necessary. Name the compounds listed using the correct names of the polyatomic ions.*

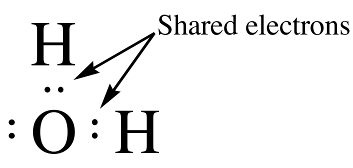
E. Covalent (Molecular) Compounds

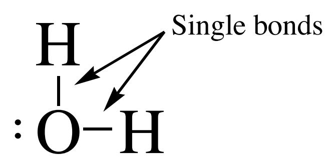
Covalent bonds form between nonmetal atoms located in Groups 4A, 5A, 6A or 7A. In a **covalent compound,** octets are achieved by sharing electrons between atoms. The sharing of one pair of electrons is referred to as a single bond. A double bond is the sharing of two pairs of electrons between atoms. In a triple bond, three pairs of electrons are shared. To write the formula of a covalent compound, determine the number of electrons needed to complete an octet. For example, nitrogen in Group 5 has five valence electrons so that it needs 3 more electrons for an octet; each nitrogen atom shares 3 electrons in covalent compounds.

Electron Dot Structures

The formulas of covalent compounds are determined by sharing valence electrons until each atom has an octet. For example, in water, oxygen shares two electrons with two hydrogen atoms. Oxygen has an octet and the hydrogen atoms are stable because they have two electrons in their outer (valence) shells. (Note that shared electron pairs are often represented as lines connecting the atoms that share them.)

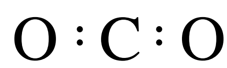
Dot Structure for H_2O



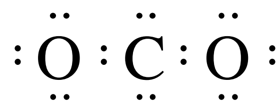


In another example, consider CO_2, a molecule that has two double bonds. In electron dot structures, carbon has 4 valence electrons and each oxygen atom has 6. Thus for CO_2, a total of 16 (4 + 6 + 6) electrons can be used in forming the octets by sharing electrons. We can use the following steps to determine the electron dot structure for CO_2:

1. Connect the atoms with pairs of electrons, which uses 4 electrons:



2. Add electrons to complete octets around all atoms:



3. Count the number of electrons used. We used 20. But only 16 are available. Therefore, we must economize by removing 2 pairs of electrons (4 electrons), and move 2 other pairs in between the C and the O's in order to maintain octets. This will form two double bonds:

Names of Covalent Compounds

Binary (two-element) covalent compounds are named by using *prefixes* that give the number of atoms of each element in the compound. The first nonmetal is named by the element name; the second ends in *-ide*. The prefixes are derived from Greek names: mono-, di-, tri-, tetra-, penta-, hexa-, hepta-, octa-, nona-and deca-. (Higher ones exist, but are rarely used.) Usually the prefix mono- is not shown for the first element.

Formula of Covalent Compound	Name
CO	carbon **mon**oxide
CO_2	carbon **di**oxide
PCl_3	phosphorus **tri**chloride
N_2O_4	**di**nitrogen **tetr**oxide ("a" dropped before the vowel "o" in "oxide")
SCl_6	sulfur **hexa**chloride

In the worksheet, *write the electron dot structure for each nonmetal. Then write electron dot structures for the covalent compounds. Name each of the covalent compounds, using the numerical prefixes when appropriate.*

Activity 4 - Writing Formulas and Names Worksheet

Name____________________________

Section____________ **Date**_______

Exercise A. Electron Dot Structures

1. Using the example given, complete this table.

Element	Atomic Number	Electron arrange-ment of atom	Electron dot structure of atom	Loss/gain of electrons by atom	Electron arrange-ment of ion	Ionic charge	Symbol of ion	Name of ion
Sodium	11	2-8-1	Na•	Lose 1 e^-	2-8	1+	Na^+	Sodium ion
Oxygen								
Aluminum								
Potassium								
Chlorine								
Calcium								
Nitrogen								
Sulfur								

2. Review the “Name of ion” column above. What distinguishes the naming of the metal cations from the naming of nonmetals anions?

Exercise B. Writing Ionic Formulas:

1. Use the periodic table to help complete the table below.

Name	Positive ion	Negative ion	Formula
Sodium oxide	Na^{+}	O^{2-}	Na_2O
Magnesium chloride			
Potassium chloride			
Calcium oxide			
Aluminum bromide			
Lithium phosphide			
Aluminum sulfide			
Aluminum nitride			
Calcium nitride			

2. Name the following ionic compounds:

Na_2S ____________________

MgF_2 ____________________

MgS ____________________

K_3N ____________________

Ca_3P_2 ____________________

$AlCl_3$ ____________________

3. Review the answers in problems 1 and 2 of exercise B above. What do the subscripts represent?

Exercise C. Ionic Charges for Transition Metals

1. Complete the table below.

Name	Positive ion	Negative ion	Formula
Iron (II) bromide	Fe^{2+}	Br^{-}	$FeBr_2$
Iron (II) chloride			
Iron (III) sulfide			
Copper (II) chloride			
Copper (II) sulfide			
Copper (II) nitride			
Zinc oxide			
Silver sulfide			

2. Name the following ionic compounds:

Cu_3P ______________________________

Fe_2O_3 ______________________________

FeI_3 ______________________________

$CuCl$ ______________________________

$ZnBr_2$ ______________________________

3. Consider your answers in problems 1 and 2 of exercise C above. What do the roman numerals in parentheses represent?

Exercise D. Polyatomic Ions

1. Complete the table below.

Name	Positive ion	Negative ion	Formula
Sodium nitrate	Na^+	NO_3^-	$NaNO_3$
Lithium carbonate			
Potassium sulfate			
Calcium bicarbonate			
Aluminum hydroxide			
Lithium sulfite			
Sodium phosphate			
Iron (II) phosphate			

2. Name the following ionic compounds:

$CaSO_4$ ____________________________

Cu_3PO_4 ____________________________

$Al(NO_3)_3$ ____________________________

Na_2CO_3 ____________________________

$MgSO_3$ ____________________________

$Fe(HCO_3)_2$ ____________________________

3. Consider *all* of the nomenclature exercises in exercises B, C and D. What are the rules for the correct placement of parentheses in the naming and writing chemical formulas of ionic compounds?

Exercise E. Lewis Dot Structures of Atoms and Molecules

1. ***Electron dot formulas of elements:*** Atoms are represented by symbol with valence e^-'s represented by dots. Complete the following table. Distribute dots on all four sides before pairing.

Hydrogen	Carbon	Nitrogen	Oxygen	Sulfur	Chlorine
$\mathrm{H}\cdot$					

2. ***Electron dot formulas of covalent compounds:*** Lewis dot structures must have the correct number of valence electrons displayed in bonded or nonbonded pairs along with the octet rule being obeyed (duet rule for H). Complete the following table for the given binary covalent compounds.

Compound	Electron dot structure	Name
HCl	$\mathrm{H}:\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}:$	Hydrogen chloride
SBr_2		Sulfur dibromide
PCl_3		
OF_2		
SO_3		

Questions and Problems

1. Write the correct formulas for the following ions:

sodium ion ___________	oxide ion ___________	calcium ion ___________
chloride ion __________	sulfate ion __________	iron (II) ion ___________

2. Write the correct name of the following compounds.

CuO	____________________
N_2O_4	____________________
$Al(NO_3)_3$	____________________
PCl_3	____________________
$FeCO_3$	____________________
$Cu(OH)_2$	____________________
Ag_2O	____________________

3. Identify the following compounds as ionic or covalent. (circle I or C) and write the corresponding molecular formula.

Sodium oxide	I C	__________________________
Iron (III) bromide	I C	__________________________
Sodium carbonate	I C	__________________________
Aluminum sulfite	I C	__________________________
Carbon tetrachloride	I C	__________________________
Nitrogen tribromide	I C	__________________________

4. Your friend wants to know what the formula $FeSO_4$ on her vitamin bottle means and what the name of this ingredient is. Help her understand the meaning of the symbols and the correct name associated with this formula (i.e write a brief answer to her question).

Activity 5 - Compounds and Their Formulas

Goals

- Identify the elements and number of atoms in the formula of a compound.
- Compare some physical properties of a compound with the properties of the elements from which it was formed.
- Determine the subscripts in the formula of a compound.
- Describe the types of elements in ionic and covalent compounds.
- Identify the bonding in a compound as ionic or covalent.

Pre-lab Questions *(answer these on a separate sheet using complete sentences)*

1. Why are color, texture, state, density, and melting point considered physical properties?
2. Why do the physical properties of the elements change when they combine to form a compound?
3. How is the number of atoms in a molecule indicated in the formula?
4. Why do compounds of metals and nonmetals consist of ions?
5. What is a covalent bond?
6. What compound in toothpaste is a preventative for cavities?

Concepts to Review

Formulas
Ions
Ionic and covalent bonds
Formation of ionic and covalent compounds
Naming ionic and covalent compounds

Introduction

Almost everything you see around you is made of compounds. A compound consists of two or more different elements that are chemically combined. Although there are currently (as of 2015) 118 elements known, there are millions of different compounds.

In a compound, there is a definite proportion of each element. This is represented in the formula, which gives the lowest whole number ratio of each kind of atom. For example, water has the formula H_2O.This means that two atoms of hydrogen and one atom of oxygen are combined in every molecule of water. Every water molecule is represented by this, and only this, formula.

A mixture consists of two or more substances (elements or compounds), which are not chemically combined. Thus, the components maintain their original physical properties, and they can be separated by physical methods such as use of a magnet, filtration, or evaporation.

Properties of Elements and Compounds

When we observe a compound or an element, we can describe physical properties such as color and luster. We can measure other physical properties such as density, melting point and boiling point. When elements undergo chemical combination, the physical properties change to the physical properties of the new compound, which is a novel substance different from its components. For example, when silver tarnishes, the physical property of the shiny silver metal changes to a dull gray color as silver combines with sulfur to form tarnish, Ag_2S. A chemical change has occurred when the reaction between the elements has caused a change in their physical properties.

Types of Bonds in Compounds

Atoms form compounds to become more stable, usually by forming octets in their outer shells. The attractions between the atoms in a compound are called *chemical bonds*. For example, when a metal combines with a nonmetal, the metal loses electrons to form a positive ion and the nonmetal gains electrons to form a negative ion. The attraction between the positive ions and the negative ions is called an *ionic bond*. When two nonmetals form a compound, they share electrons and form *covalent bonds*. The combinations of atoms in covalent compounds are called *molecules*.

Table 1. A selection of compounds and their corresponding bonding type.

Compound	Types of Elements	Characteristics	Type of Bonding
$NaCl$	Metal, nonmetal	Ions, (Na^+, Cl^-)	Ionic
CCl_4	Two nonmetals	Molecules	Covalent
$MgBr_2$	Metal, nonmetal	Ions, (Mg^{2+}, Br^-)	Ionic
NH_3	Two nonmetals	Molecules	Covalent

Safety

Wear safety glasses at all times!

Act in accordance with the laboratory safety rules of Cabrillo College.

Avoid contact with all chemical reagents and dispose of those used in experiments in appropriate waste containers.

Caution: Acids are corrosive; they will cause chemical burns to your skin. Know the location of solid sodium bicarbonate ($NaHCO_3$) in the lab as well as the aqueous solution of sodium bicarbonate. Use either the solid or the solution of sodium bicarbonate to neutralize any spills of the 6 *M* $HCl_{(aq)}$ solution. Should you happen to spill the acid solution on your skin, use the sodium bicarbonate solution to neutralize it right away and rinse off with water. Have a classmate notify the instructor immediately.

Caution: To sample the odor of a gas, first fill your lungs with fresh air and hold it while you use your hand to fan some of the vapors from the reaction tube toward you. Carefully note the odor.

Materials:

A selection of elements and compounds as identified on your experimental pages, samples of iron (Fe) filings, sulfur (S), iron filings and sulfur mixture (Fe + S), iron (II) sulfide (FeS), and 6 *M* HCl (hydrochloric acid).

Equipment: bar magnet, spatula, dropper, test tubes and test tube rack.

Experimental Procedure

A. Interpreting Formulas of Compounds

Observe the compounds in the laboratory display. Describe the physical properties of each compound. Write the formula for each compound. From the formula of each compound, state the number of atoms of each element present in that compound. From the display of elements, observe and record some of the physical properties of the individual elements.

B. Physical Properties of FeS and its Elements

1. Your instructor may do this part of the experiment as a demonstration. Obtain samples of Fe, S, a mixture of Fe and S, and FeS. These may be in prepared test tubes or sample containers. Describe the physical properties of each sample.
2. Using a chemistry handbook (Chemical Rubber Company—CRC), look up the density, melting and boiling points of Fe, S, and FeS. Record these values.
3. Test each of the samples for magnetic attraction by running a bar magnet under the sample in each container. (*Do not place the magnet directly into the samples!* The attracted particles cling to the magnet and make it difficult to clean.) If there is magnetic attraction, you will see particles follow the magnet. Record your observations.
4. (Optional) This part of the experiment involves a reaction that produces H_2S gas, which is toxic in more than trace amounts. Check with your instructor before proceeding. Place a small amount of each sample (enough to cover the tip of a spatula) in a test tube. **WORKING IN THE HOOD,** slowly add 15 drops of 6 *M* HCl (**corrosive**) into each test tube. Observe any reaction. **CAREFULLY** note any odor.
5. Describe each sample as an element, mixture, or compound.

Activity 5 - Compounds and Their Formulas

Name____________________________

Section____________ **Date**_______

Exercise A. Interpreting Formulas of Compounds

1. Complete the following table using the samples placed around the room.

Formula of compound	Physical properties of compound	Number of atoms of each element	Physical properties of the elements
Example: $CuSO_4$	Deep blue crystals	1 Cu, 1 S, 4 O	Cu - shiny, copper metal; S - yellow chunks; O - colorless gas

2. When elements combine to form compounds, are the physical properties of the compound the same as those of the elements? Explain.

3. Does the formula of a compound vary or is matter constant in composition? Explain your answer.

Exercise B. Physical Properties of FeS and Its Elements

1. Complete the following table from your observations of FeS and its elements.

Sample	**Physical properties** (color, state, luster?)	**Density**-d (g/mL)	**M.P.** (ºC)	**B.P.** (ºC)	**Magnetic attraction**	**Reaction with HCl? Odor?**	**Description of Sample** (element, mixture, compound)
Fe							
S							
Fe + S							
FeS							

Use the results in your chart to answer the following questions:

2. How does the attraction to the magnet differ for the elements, mixture, and compound? Explain.

3. Why do the physical properties of Fe and S differ from those of FeS?

4. Can the elements in the Fe + S mixture and the compound FeS be separated using the same methods? Explain.

Questions and Problems

1. Complete the table for the given compounds

Compound	Units in compound (ions or molecules)	Type of bonds (ionic or covalent)
LiBr	ions	ionic
$CaCl_2$		
CCl_4		
NH_3		
K_2S		
MgO		

2. List the number of atoms of each kind of element in the following formulas:

Formula	Number and Kind of Atoms in the Compound
H_2O	2 atoms H and 1 atom O
$CuCl_2$	______________________
Al_2S_3	______________________
$Ba(NO_3)_2$	______________________
$C_6H_{12}O_6$	______________________

3. Write formulas of the following compounds from the number of atoms given. The elements are listed in the order in which they appear in the formula.

1 atom of C and 2 atoms of O ______ CO_2 ______

1 atom of N and 3 atoms of H ______________________

1 atom of C and 4 atoms of Cl ______________________

2 atoms of Fe and 3 atoms of O ______________________

1 atom of Ba, 1 atom of S, 4 atoms of O ______________________

4. Identify the elements in each compound as a metal and nonmetal, or two nonmetals. Indicate the bonding in each as ionic or covalent.

	Elements	Type of Bonding
$BaCl_2$	metal and nonmetal	ionic
C_3H_8	______________________	____________________
Li_2O	______________________	____________________
PCl_3	______________________	____________________
NaBr	______________________	____________________
SO_3	______________________	____________________

Indicate whether each of the following is a chemical or physical change:

tearing a piece of paper in two ______________________________________

burning a match ______________________________________

grinding pepper ______________________________________

rusting iron nail ______________________________________

freezing water for ice cubes ______________________________________

Activity 6 - Introduction to Small Scale Chemistry

Goals

- Introduce small-scale techniques
- Record both qualitative and quantitative observations
- Draw conclusions from results

Introduction

Small-scale chemistry techniques involve a scaling down of reaction size and doing chemistry *in plastico* rather than using traditional glassware. Due to the decreased quantities of chemical reagents, the hazards, chemical waste, and costs of experiments are also decreased. The obvious physical differences in small-scale chemistry allow the opportunity to work with reactions that would be difficult or dangerous on a larger scale.

The surface tension of water allows us to replace our traditional test tube with drop size quantities, **the drop being its own container**. The volume of the drop can become reproducible with practice and consistent approaches.

Smaller amounts of reagents may require visual optimization using a hand lens (magnifying glass). This device is usually kept at hand during all small-scale experiments.

Consistent results that lead to meaningful conclusions require reproducibility. The most important technique in providing meaningful results in small-scale chemistry is avoiding contamination. The tip of any small-scale pipet must be kept clean by avoiding contact with anything else, e.g., another drop, the reaction surface, or another pipet. *Any contaminated pipet should be given to your instructor*.

Reproducible results require *precision*. Whenever more than one trial or more than one sample is being tested for comparison, all of the trials/samples must be treated as similarly as possible. Many times students cannot control the *accuracy* of a balance but they can make sure that they use the same one throughout an experiment in an appropriate and precise manner.

Principles of Observation

Observations in the laboratory can be either quantitative or qualitative. Quantitative observations are called **measurements** and usually require the recording of a number and the unit of measure. Qualitative observations describe what is seen in the laboratory. Words that are more than superficially descriptive must be used when qualitative observations are made, e.g., the word "blue" does not convey anything **specifically** meaningful to someone who was not present. Adjectives describing color, size, vigor of reaction, or any other observable are a critical tool in the recording of experimental observations.

When reacting two substances together it is imperative that you monitor the reaction immediately upon mixing and continue to observe until no further change occurs. Some chemical reactions are slow and some are fast. Sometimes more than one reaction can occur as laboratory conditions may influence results (air, light, fumes, etc.).

The interpretation of observations is what occurs in the formulation of **conclusions**. To simply state what you see is important. To extrapolate some meaningful conclusion from your data or sets of data requires more than observation.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact with all chemical reagents and dispose of those used in experiments using appropriate waste container.

Materials

10-mL graduated cylinder, top-loading centigram balances, deionized water, transfer pipets, calculator, green food dye solution

Experimental Procedure

A. Techniques:

***Microburet "Pipet" Techniques I* – The Basics**

The microburet, often referred to as the pipet, is a versatile tool.

1. Fill 2 wells of a 24-well tray with water that has been colored with green food dye. The green dye solution is in a plastic bottle obtained from Reagent Central. Be sure to promptly return the bottle containing the green dye solution to Reagent Central.

2. Fill 2 wells of the 24-well tray with distilled water. The distilled water is available in plastic squeeze bottles located in a cabinet at the side of the room.

3. To make a microburet select a long stem pipet and cut the stem with your scissors at a point about 2 cm from the bulb. NOTE: *make the cut at right angles!*

4. Begin by cleaning the microburet. Draw a little distilled water up into the bulb; shake it so that all internal surfaces have been wetted.

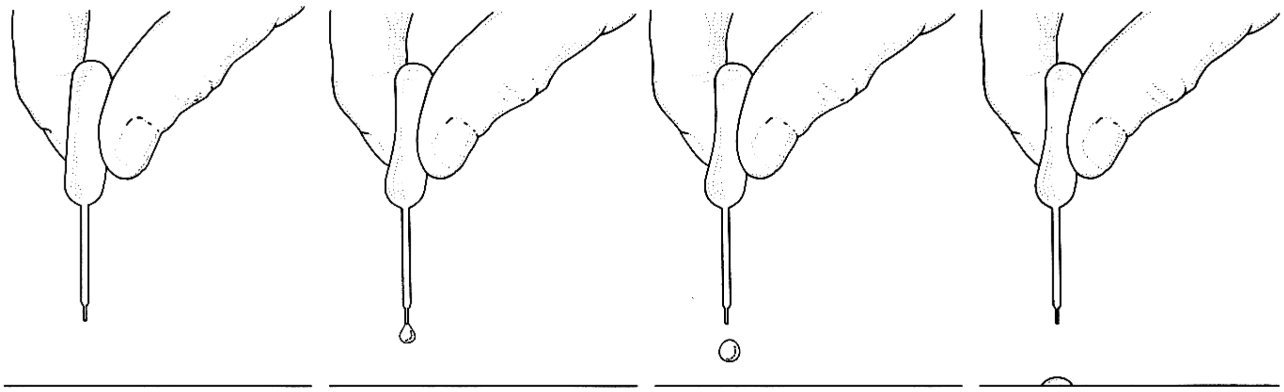
5. Holding the microburet vertically, expel the water into the waste cup. Press firmly to get the last drops out.

6. Now draw a little green food dye into the bulb, shake to rinse the bulb, and expel the dye into the waste cup.

7. Finally, squeeze the bulb and draw green food dye into the microburet.

Note: This sequence (steps 4, 5, 6, and 7) is known as a *good wash, rinse, and transfer technique*.

8. Practice producing pools of various sizes on the lab-top surface using the single drop *standard delivery technique*. To do so, hold the microburet *vertically* over the reaction surface as in the figure below, so its tip is about 1 to 2 centimeters above the reaction surface. Never touch the tip of the pipet to the reaction surface. "Touching" drops onto the reaction surface causes contamination of the contents of the microburet. Always expel the drops; never touch the drops onto the reaction surface. This technique (with the microburet held *vertically*) is called the *standard delivery technique*.



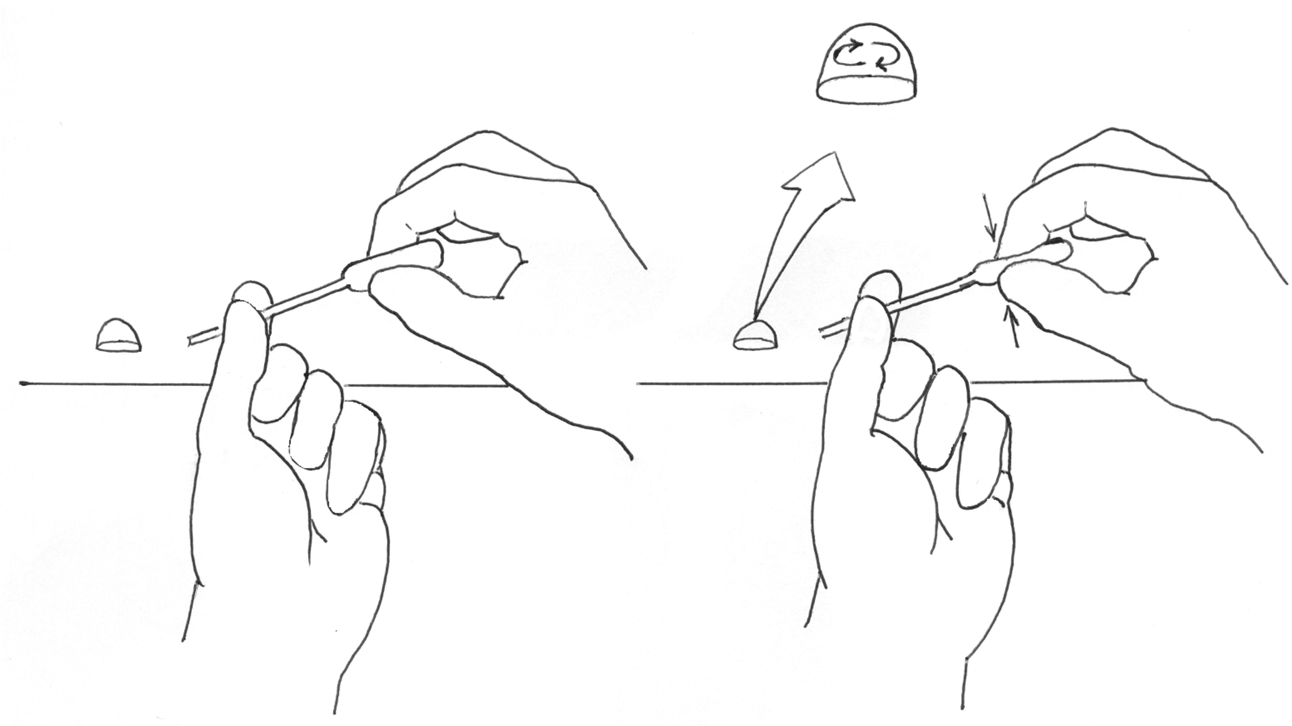
9. Make another microburet.

10. Follow the *good wash, rinse, and transfer technique* (steps 4-7) to fill the pipet with deionized water from one of the wells in the 24-well tray.

11. Expel a single drop of deionized water onto the reaction surface.

12. Now expel a drop of the green food dye solution on top of the drop of water, being careful not to touch the tip of the microburet to the deionized water drop.

13. Throughout this semester we will observe many chemical reactions between two substances conducted in this manner. While some visible change may accompany the initial mixing of two substances, often times thorough mixing is required to allow the reaction to go to completion. You can stir a hemispherical droplet on a reaction surface by gently blowing air from an empty long-stem pipet past it as shown below. Use one hand to steady the tip of the pipet and aim it at the side of the droplet. With the other hand gently squeeze the bulb repeatedly. The moving air will cause the droplet to swirl, mixing the contents without causing contamination.



14. Practice mixing drops of green food dye solution and deionized water until you are comfortable with the technique.

15. When finished, clean the reaction surface by first absorbing the liquid onto a small piece of microtowel. Dispose of the contaminated microtowel into the solid waste bins provided. Next wipe the surface with a paper towel and deionized water. Finally, dry the surface with a paper towel.

***Microburet "Pipet" Techniques II* – Effects of Drop Angle**

The microburet is a simple constant-drop-volume delivery device when properly used.

1. Refill the microburet with the green food dye solution.

2. Locate the 1 × 12 well strip in your chemistry kit. The well holds a volume of 0.40 mL when filled to the rim such that there is a slight convex bulge above the rim of the well.

3. Count how many drops it takes to fill a single 0.40-mL well using the *standard delivery technique*. To do so, hold the microburet *vertically* (90^o) over the well so its tip is about 1 to 2 centimeters above the well strip. Remember to never touch the tip of the pipet to the reaction surface. "Touching" drops onto the reaction surface causes contamination of the contents within the microburet. Always expel the drops, never touch the drops onto the reaction surface. Record your result.

4. Repeat the process to verify your result. Record your result.

5. Now tilt the microburet at 45^o angle. Count how many drops it takes to fill the well with the microburet at 45^o. Record your result.

6. Repeat the process to verify your result. Record your result.

7. Now tilt the microburet so that it is horizontal (0^o) over the well and see how many drops are required in this position. Record your result.

8. Repeat the process to verify your result. Record your result.

The point here is that in order to obtain consistent results, you must be consistent in your delivery technique, including the position of the microburet!

Using a balance

- ❑ Never place reagents directly on the pan of a balance.
- ❑ Use the same balance throughout an experiment.
- ❑ Record all of the numbers given in the digital output.
- ❑ Take into account the mass of the weighing container by:

"Weighing by difference": The empty container is weighed and its mass recorded. The mass of the object is obtained by first weighing the container with the object in it and then subtracting the mass of the empty container.

Or

"Taring the balance": The empty container is placed on the pan of the balance and then the "tare" button is pressed to zero the balance. The object may now be placed into the container and the balance will read the mass of the object only.

B. Discovery Experiment

Design an experiment to find the mass of one drop of water, delivered by the *standard delivery technique*, in units of milligrams. Your method must legitimately find the mass of a drop so that it has at least two significant figures. Repeat your experiment to verify your results. Be sure to record your procedure and results.

Activity 6 - Introduction to Small Scale Chemistry Worksheet

Name____________________________

Section____________ Date_______

Exercise A. Effect of Drop Angle

1. Complete the following table by recording your results:

Number of drops to fill 0.40-mL well

	Trial 1	Trial 2	Trial 3
Standard Delivery Technique Vertical (90°)			
Halfway (45°)			
Horizontal (0°)			

2. Based on the data in the table above, formulate a conclusion.

Exercise B. Determine mass of a drop of water delivered by the *Standard Delivery Technique*.

Description of the Procedure (using words):

1. Data and Calculations:

2. Mass of a single drop ____________________

Activity 7 - Chemical Changes[1]

Goals

- Observe and record chemical changes.
- Design and carry out experiments to identify chemicals in consumer products.
- Use proper small-scale techniques to produce reproducible results.

Pre-lab Questions *(answer these on a separate sheet using complete sentences)*

1. What do you expect to see if a chemical reaction occurs?
2. Why do chemists record what they see in the laboratory?
3. What are possible sources of errors in the laboratory?
4. How can visible changes help us describe invisible atoms and molecules?
5. How do you clean up your reaction surface after you have recorded your data and answered your questions?
6. Why are sodas called "carbonated beverages"?

Introduction

In today's lab you will practice combining prepared solutions in a reproducible manner to observe whether a change occurs. Chemical changes involve a change to the starting materials (the reactants) and are visible when a color changes, a solid comes out of solution or a gas is formed. By recording what you see when two solutions are mixed you testify to what you saw. This information can be used to identify unknowns or to describe changes on a submicroscopic level.

Description of liquids

Recording meaningful observations requires a descriptive vocabulary. Solutions may be colored and clear or colorless and clear. Proper description of a liquid should include both the color and clarity. Water for instance is described as clear and colorless. It is clear because there are no particulates floating in the liquid and light is transmitted through it. Water is lacking in color so it is colorless. Milk on the other hand is opaque and white. One should try to be as accurate as possible in the description of colors. The description of "blue" is incomplete, since periwinkle blue is a different observation than navy blue.

Descriptions of precipitates

A precipitate is a solid material that results from the chemical reaction of two liquids or solutions. Precipitates are described on the basis of color, consistency and distribution. Precipitates vary greatly in appearance. Any cloudy solution indicates a precipitate was formed. White solids may be opalescent white, ecru chunks or grayish white suspensions. A suspension consists of solid particles dispersed throughout the mixture, which looks different than precipitates that settle to the bottom. Other examples of appropriate adjectives include: milky, cloudy, sticky, clumpy, grainy, free-flowing… the list goes on.

In today's lab you will practice combining prepared solutions in a reproducible manner to observe whether a change occurs. You will study some of the chemicals in common consumer products. Based on your observations, you will try to make conclusions about the content of these consumer products.

---

[1] Adapted from: Waterman, E. L. *Chemistry: Small-Scale Chemistry Laboratory Manual;* Addison-Wesley/Prentice-Hall, Inc.: Upper Saddle River, New Jersey, 2002; pp 17-24.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

*Contact with silver nitrate ($AgNO_3$) will stain the skin.

Materials

Reagent Central chemicals include microburets filled with the following solutions:

Sodium hydrogen carbonate ($NaHCO_3$)

Sodium hydrogen sulfate ($NaHSO_4$)

Phenolphthalein (phen)

Starch

Ammonia (NH_3)

Sodium hypochlorite ($NaOCl$)
This will damage your clothing!

Lead (II) nitrate ($Pb(NO_3)_2$)

Copper (II) sulfate ($CuSO_4$)

FD&C No. 1 (blue dye)

Potassium iodide (KI)

Calcium chloride ($CaCl_2$)

Sodium carbonate (Na_2CO_3)

Silver nitrate ($AgNO_3$)

Sodium hydroxide (NaOH)

Hydrochloric acid (HCl) 

Equipment: Empty pipet for stirring Lab top reaction surface

Experimental Procedure

1. Use small-scale microburets to put 2 drops of each chemical on the X's in the indicated spaces below. For background contrast, view the drops on both black and white backgrounds provided by the X's. Stir each mixture by blowing air from an empty pipet (as previously described). Record what you see in your lab notebook or on your worksheet (according to your instructor's direction). Do not clean your surface yet.
2. Test several foods for the presence of starch. If you don't know how to do this, answer the questions on your worksheet. If you still don't know how to test for starch, ask your instructor.
3. Avoid contamination by cleaning up in a way that protects you and your environment. Carefully clean the small-scale reaction surface by absorbing the contents onto a small square of tissue paper or paper towel. Dispose of the paper in the appropriate waste container. Wipe the surface with a damp towel and then dry it. Wash your hands with soap and water before leaving the lab.

Reaction Template: Insert this page into the lab-top. Mix one drop of each solution, using a long stem pipet to blow air past the droplet to complete the mixing.

a. $NaHCO_3$ + HCl		h. $NaHSO_4$ + Na_2CO_3	
b. HCl + blue dye		i. Na_2CO_3 + phen	
c. $NaOCl$ + blue dye	Now add 1 drop of HCl	j. phen + $NaOH$	
d. $NaOCl$ + KI	Now add 1 drop of starch	k. $NaOH$ + $AgNO_3$	
e. KI + $Pb(NO_3)_2$		l. $AgNO_3$ + NH_3	Absorb this mixture onto a scrap of paper, expose it to sunlight, dispose of after making your observations.
f. $Pb(NO_3)_2$ + $CaCl_2$		m. $CuSO_4$ + NH_3	
g. $CaCl_2$ + $NaHSO_4$	Be patient! Some chemical reactions are slow!	n. $CuSO_4$ + $NaHCO_3$	

Activity 7 - Chemical Changes

Name____________________________

Section____________ **Date**______

Experimental Data: Record all your observations.

	Initial Observations Reactants (before mixing) X	+ Y	Final Observations Product (after mixing) Products
a.	$NaHCO_3$	HCl	
b.	HCl	blue dye	
c.	blue dye	NaOCl	 after HCl
d.	NaOCl	KI	 after starch
e.	KI	$Pb(NO_3)_2$	
f.	$Pb(NO_3)_2$	$CaCl_2$	
g.	$CaCl_2$	$NaHSO_4$	

Experimental Data Cont.: Record all your observations.

	Initial Observations Reactants (before mixing) X	+ Y	Final Observations Product (after mixing) Products
h.	$NaHSO_4$	Na_2CO_3	
i.	Na_2CO_3	phen	
j.	phen	NaOH	
k.	NaOH	$AgNO_3$	
l.	$AgNO_3$	NH_3	 after sunlight
m.	NH_3	$CuSO_4$	
n.	$CuSO_4$	$NaHCO_3$	

Testing Foods for Starch

1. Describe **how** you tested the available samples for starch. (Include reagents used.)

2. Complete the following table:

Sample	Observation	Conclusion (+/- for starch)
	Before: After:	
	Before: After:	
	Before: After:	
	Before: After:	

Questions and Problems

1. Sodium hydrogen carbonate, $NaHCO_3$, is also known as sodium bicarbonate or more commonly as baking soda. When an acid such as hydrochloric acid, HCl, is added to sodium bicarbonate, bubbles of carbon dioxide form. Write the formula for carbon dioxide. What common consumer products contain this gas?

2. Which of the other combinations of chemicals form bubbles?

3. Given that bicarbonate ion (HCO^{3-}) and carbonate ion (CO_3^{2-}) have similar chemical reactivities, what type of gas was likely formed from the other chemical combinations identified in question 2? Re-read the content of question 1. What is likely to be the nature of the reactant that does not contain carbonate ion in question 2 above, in other words, what type of chemical is it?

4. The body uses hydrochloric acid, HCl, to help digest food. In what organ is HCl is found?

5. What color does blue food dye turn when HCl is added?

6. Sodium hypochlorite, NaOCl, is a common ingredient in household bleaches and cleansers. What happened to the color of blue dye when both HCl and NaOCl are added?

7. Potassium iodide, KI, is the source of iodine (I_2) in iodized salt. What color is the KI + NaOCl mixture?

8. What color does starch change to in the presence of KI and NaOCl? Evaluate whether or not the mixture of KI and NaOCl can be used to distinguish solutions containing either starch or simple sugars (mono or disaccharides).

9. A precipitate is a solid that separates upon mixing solutions. These reactions typically result from the cations and anions of the reactants switching partners. Write the two reactants and possible products that could be formed for the combination that produced a bright yellow precipitate.

10. Which other mixtures produced precipitates? Describe their colors and textures with words like milky cloudy or grainy.

11. Write an equation showing the reactants and likely products for the reaction that was very slow to form a precipitate.

12. Which solutions produced a distinctive brown precipitate? Describe that color.

13. Look at the scrap of paper you used to absorb the silver nitrate and ammonia mixture. What evidence do you see that indicates that silver compounds are light sensitive? In what way is this chemical property utilized to form a long-term record of our daily experiences?

14. What were three observations that indicated the formation of a new substance (chemical change)?

15. Describe any other unique or interesting observations here.

16. Which foods contained starch? Is this consistent with what you would have predicted from your personal knowledge of food science? Explain.

Activity 8 – Chemical Names and Formulas[1]

Goals

- ❑ Write chemical names and formulas of common chemical compounds.
- ❑ Describe the colors and textures of common ionic compounds.
- ❑ Synthesize chemical compounds and write their names and formulas.

Pre-Lab Lecture Questions. *Answer these questions on a separate sheet using complete sentences.*

1. What is an ion? What is an ionic compound? How can we recognize ionic compounds? What is a salt?
2. Compare cations, anions and polyatomic ions. What do they all have in common? How are they different?
3. How can the periodic table help to remember the charges on the simple ions of the representative (main group) elements?
4. What is the chemical name of baking soda? Is there more than one name that can be used?
5. Why do some cation names include Roman numerals in parentheses?
6. Why do some chemical formulas include parentheses and others do not?
7. What is the precipitate formed when iron (III) chloride reacts with silver nitrate?

Concepts to Review

Names of Elements
Periodic Table
Atomic Structure
Transition Elements, Representative Elements

Introduction

Chemistry is the central science, a study of all that has mass and volume. An effort of this magnitude requires a clear language that communicates in a broad but consistent way. At first appearance, chemistry may appear difficult because there are common words that take on new meaning. For example, "salt" is a term widely used to describe table salt (also known as sodium chloride). In chemistry, a **salt** is simply **any compound composed of ions other than hydrogen ion, oxide ion, or hydroxide ion**. Sodium chloride is an example of a salt, as is potassium chloride, calcium carbonate and stannous fluoride. In chemistry, there is an effort to move away from using common names to identify the majority of compounds because this would require memorization of every single name. Considering the vast number of ionic compounds (over a million), a systematic method of nomenclature has been developed to designate these.

As a student of chemistry you will learn how to translate a chemical formula into the systematic name and vice versa. The observations and experiments in today's lab only involve compounds containing charged species—cations, anions and polyatomic ions. The various combinations of oppositely charged ions are called **ionic compounds.** Their chemical formulas represent the proportion of positive ion to negative ion that results in electrical neutrality, i.e., no net charge. The correct chemical formula for sodium chloride is NaCl. The 1:1 ratio of sodium to chloride ions tells us that sodium ions and chloride ions must have the same charge magnitude. (Note that when there is only one of an ion per formula, we do not use the number one as subscript to indicate this; i.e., we don't write Na_1Cl_1.) After looking at the table on the following page, we see that sodium is a cation with a 1+ charge and chloride is an anion with a 1- charge. Knowing both the magnitude and the sign of the charge is necessary for writing the correct formulas and the correct chemical names. Sodium oxide has a formula of Na_2O. Without looking at the table of ions, what must the charge of oxide be? If you recognized that there are two sodium 1+ ions for each oxide ion and deduced that oxide must have a 2- charge you are well on your way to describing ionic compounds!

---

[1] Adapted from: Waterman, E. L. *Chemistry: Small-Scale Chemistry Laboratory Manual;* Addison-Wesley/Prentice-Hall, Inc.: Upper Saddle River, New Jersey, 2002; pp 51-58.

It is common to see **precipitates** in the chemical reactions of ionic compounds in solution, i.e. insoluble solids coming out of solution. *A general rule is that precipitates usually do not contain sodium, potassium, acetate, nitrate or sulfate ions.*

Table 1. A Collection of Common Ions.

Name	**Formula**	**Name**	**Formula**	**Name**	**Formula**
Sodium	Na^+	Magnesium	Mg^{2+}		
Potassium	K^+	Calcium	Ca^{2+}		
Copper (I)	Cu^+	Copper (II)	Cu^{2+}		
Silver	Ag^+	Iron (II)	Fe^{2+}	Iron (III)	Fe^{3+}
Ammonium	NH_4^+	Lead (II)	Pb^{2+}	Lead (IV)	Pb^{4+}
		Tin (II)	Sn^{2+}	Tin (IV)	Sn^{4+}
Fluoride	F^-	Oxide	O^{2-}	Nitride	N^{3-}
Chloride	Cl^-	Sulfide	S^{2-}		
Bromide	Br^-	Sulfate	SO_4^{2-}		
Iodide	I^-				
Acetate	$C_2H_3O_2^-$				
Hydroxide	OH^-				
Nitrate	NO_3^-				
Nitrite	NO_2^-				
Hydrogen carbonate (**bicarbonate**)	HCO_3^-	Carbonate	CO_3^{2-}		
Dihydrogen phosphate	$H_2PO_4^-$	Hydrogen phosphate	HPO_4^{2-}	Phosphate	PO_4^{3-}

Safety

Wear safety glasses at all times!
Act in accordance with the laboratory safety rules of Cabrillo College.
Avoid contact with all chemical reagents and dispose of reactions using appropriate waste containers.

Contact with silver nitrate ($AgNO_3$) will stain the skin.

Materials

Reagent Central chemicals include a variety of pure ionic compounds and aqueous solutions of ionic compounds as identified on your experimental pages.

Equipment: Empty pipet for stirring Lab top reaction surface

Experimental Procedure

A. Compound Observations

1. View the samples of solid compounds available at Reagent Central. Write a description of the color and any other adjectives that might distinguish one compound from another. If the formula is given on the data sheet, provide the correct name. If the name is given, write the correct formula. Record observations and answers in your laboratory notebook and/or the data page provided.

B. Precipitation Reactions

1. Insert your experimental page inside of your reaction surface.
2. Place one drop of each solution in the indicated spaces below, taking care not to contaminate the microburets. Stir by blowing air from a dry pipet. Record any observable changes, describing what happened when the two solutions were mixed.
3. Any precipitates represent new compounds formed from swapping ion partners. Write the correct formulas for the two possible products. The precipitate will be the product that doesn't contain sodium, potassium, or nitrate ions. Write the name and formula of the precipitate on your worksheet.

Reaction Template: Insert this page into the labtop. Mix one drop of each solution, using a long stem pipet to blow air past the droplet to complete the mixing.

	$AgNO_3$	$Pb(NO_3)_2$	$CuSO_4$	$MgSO_4$	$FeCl_3$
$FeCl_3$	X	X			
KI	X	X			
NaOH	X	X	X	X	X
Na_2CO_3	X	X	X	X	X
Na_3PO_4	X	X	X	X	X

Activity 8 - Chemical Names and Formulas Worksheet

Name____________________________

Section____________ **Date**_______

Exercise A. Compound Observations

1. Write observations in boxes and provide missing formula or name.

Compound Name	Formula	Description of Solid	Compound Name	Formula	Description of Solid
Potassium iodide			Sodium carbonate		
Sodium chloride			Lead (II) nitrate		
Magnesium sulfate			Sodium acetate		
Copper (II) sulfate			Ammonium chloride		
$NaHCO_3$			$CaCl_2$		
$AgNO_3$			$FeCl_3$		
KF			NaH_2PO_4		

Exercise B. Precipitation Reactions

1. Write initial observations of the solutions as observed prior to mixing in the precipitation reactions. Review the material on page 49 regarding the description of liquids if you need assistance.

Solution	Formula	Description of Aqueous Solution
Silver nitrate	$AgNO_3$ (aq)	
Iron (III) chloride		
Sodium hydroxide		
Sodium carbonate		
Sodium phosphate		
Lead (II) nitrate		
Copper (II) sulfate		
Magnesium sulfate		
Potassium Iodide		

2. Write your observations on the products of these precipitation reactions. Review the material on page 49 regarding the description of precipitates if you need assistance.

Reaction	Description of Precipitate
$AgNO_3(aq)$ + $FeCl_3(aq)$	
$AgNO_3(aq)$ + $KI(aq)$	
$AgNO_3(aq)$ + $NaOH(aq)$	
$AgNO_3(aq)$ + $Na_2CO_3(aq)$	
$AgNO_3(aq)$ + $Na_3PO_4(aq)$	
$Pb(NO_3)_2(aq)$ + $FeCl_3(aq)$	
$Pb(NO_3)_2(aq)$ + $KI(aq)$	
$Pb(NO_3)_2(aq)$ + $NaOH(aq)$	
$Pb(NO_3)_2(aq)$ + $Na_2CO_3(aq)$	
$Pb(NO_3)_2(aq)$ + $Na_3PO_4(aq)$	
$CuSO_4(aq)$ + $NaOH(aq)$	
$CuSO_4(aq)$ + $Na_2CO_3(aq)$	
$CuSO_4(aq)$ + $Na_3PO_4(aq)$	
$MgSO_4(aq)$ + $NaOH(aq)$	
$MgSO_4(aq)$ + $Na_2CO_3(aq)$	
$MgSO_4(aq)$ + $Na_3PO_4(aq)$	
$FeCl_3(aq)$ + $NaOH(aq)$	
$FeCl_3(aq)$ + $Na_2CO_3(aq)$	
$FeCl_3(aq)$ + $Na_3PO_4(aq)$	

Questions and Problems

What did you learn? Give complete and legible answers for the following:

1. Write the formulas (including charges) and names of all the **cations** represented in this experiment.

2. Write the formulas (including charges) and names of all the **anions** represented in this experiment.

3. Write a simple rule for naming ionic compounds.

4. Write a simple rule for writing chemical formulas of ionic compounds.

5. When are Roman numerals used in naming compounds?

6. What does a numerical subscript following an element in a chemical formula mean?

7. When is it correct to use parentheses in chemical <u>formulas</u>?

8. Do this on a separate sheet of paper. Complete each of the 19 reactions by writing the correct formulas of the two possible products formed. Use subscripts on cations and anions to appropriately balance the charges in each product. One product will be a solid (designated by (s)) and the other will remain in solution (designated as aqueous by (aq)). Determine the phases of each product and indicate them using the proper abbreviation in parenthesis. The precipitate (solid) will not contain sodium or nitrate ions. The reaction equation does not need to be balanced with coefficients. For example:

   $AgNO_3(aq) + FeCl_3(aq) \rightarrow Fe(NO_3)_3(aq) + AgCl(s)$

Activity 9 - Nomenclature

Every compound has its own chemical formula and its own name. The nomenclature (naming system) for ionic and molecular compounds is different. Molecular compounds contain only nonmetals and ionic compounds contain **ions (charged particles)** comprised of metals and nonmetals.

<u>Ionic compounds:</u> These consist of any positive ion (**a cation**) except H^+ with any negative ion (**an anion**). If H^+ is the positive ion, it is an acid.

The **cation** may be a metal ion (e.g., Na^+) or a polyatomic ion (e.g., NH_4^+).

The **anion** may be a nonmetal ion (e.g., Cl^-) or a polyatomic ion (e.g., SO_4^{2-}).

A. Representative Metal + Nonmetal Compounds

Examples:	KBr	potassium bromide
	$AlCl_3$	aluminum chloride

- The metal cation is always first (the name of the element is unchanged).
- The nonmetal anion is second (the element name is given an *–ide* ending).
- The compound is electrically neutral without any charges in the formula.

B. Transition Metal + Nonmetal Compounds

In general (but not in every instance), the cations formed by the transition metals can have two different charges. **Memorize** those ions assigned by your instructor (flash cards can help you).

- If the transition metal forms only one ion, name the compound as in Case 1.

Examples:	$ZnCl_2$	zinc chloride
	Ag_2S	silver sulfide

- If the metal can form more than one type of ion, name the compounds according to one or both of the possible naming systems (each has two names!).

Examples: FeO ferrous oxide (old system) or iron (II) oxide (new system)
formed from Fe^{2+} and O^{2-}

Fe_2O_3 ferric oxide (old system) or iron (III) oxide (new system)
formed from Fe^{3+} and O^{2-}

Lead and **tin** form 2+ and 4+ ions. Even though they are **not** transition metals, they are named as such.

Archaic (old) system:

The **–ous** ending refers to the ion with the lower charge state (e.g., Fe^{2+} or Cu^+, cupr**ous**).

The **–ic** ending refers to the ion with the higher charge state (e.g., Fe^{3+} or Cu^{2+}, cupr**ic**).

Modern (new, IUPAC) system:

The modern names for Cu^+ and Cu^{2+} would be copper (I) ion and copper (II) ion.

Cases 1 and 2 involve ionic compounds that consist of only a metal cation and a nonmetal anion – two elements only. They are called **binary compounds** and consist of two monatomic ions. Ionic compounds can also be formed from more complex ions (polyatomic ions).

C. Ionic Compounds with Polyatomic Ions

The list of polyatomic ions (names and formulas) to be memorized is assigned by your instructor (again, index cards can be helpful). Don't worry – you will become more comfortable with these as you gain more experience. For all ionic compounds, the cation is named first, followed by the anion.

Examples:	$(NH_4)_2SO_4$	ammonium sulfate
	K_3PO_4	potassium phosphate
	$Fe_2(SO_4)_3$	iron(III) sulfate or ferric sulfate

Parentheses, ( ), are used only when **two or more** polyatomic ions comprise the positive portion or the negative portion (or both) of the compound. As examples, in ammonium sulfate two ammonium ions are required to balance the 2- charge on the sulfate ion to form $(NH_4)_2SO_4$, whereas in iron (III) sulfate, three 2- sulfate ions are required to balance the charge of two 3+ iron (III) ions to form $Fe_2(SO_4)_3$. An example of no need for parentheses is potassium phosphate (K_3PO_4).

D. Molecular compounds

These are compounds formed when two nonmetal atoms share electrons with other nonmetal atoms. Binary molecular compounds consist of two different atoms and should be named according to the rules below. Like ionic compounds, the more positive "ion" is first and the more negative "ion" is second, with the negative ion's name including an **–ide** ending. To determine which of the elements is the most positive (or negative), compare their relative electronegativities.

Unlike ionic compounds, **the number of each type of atom is specified with a prefix.**

1: mono	3: tri	5: penta	7: hepta	9: nona
2: di	4: tetra	6: hexa	8: octa	10: deca

If there is only one atom of the leading element, the mono prefix is not used.

Examples:	NO	nitrogen monoxide	N_2O	dinitrogen monoxide
	NO_2	nitrogen dioxide	IF_7	iodine heptafluoride
	O_2	oxygen	N_2	nitrogen

E. Acids:

Acids (from the Latin word *acidus*, meaning "sour") are an important class of compounds. One way to define these compounds is as a substance whose molecules each yield one or more hydrogen ions (H^+) when dissolved in water.

The formula for an acid is formed by adding a sufficient number of H to balance the charge on the anion. The name of the acid is related to the name of the anion and includes the label **acid**.

- Binary acids are an important class of acids. These follow the general formula HX. The anions whose names end in **–ide** have associated acids that have the **hydro–** prefix and an **–ic** ending (according to the old nomenclature system).

Example: anion = Cl^- corresponding acid = HCl (**hydro**chlor**ic acid,** or hydrogen chloride)

- Many of the most important acids are derived from oxyanions (polyatomic ions which contain oxygen). Oxyanions whose names end in **–ite** (sulfite, nitrite, chlorite, etc.) have associated acids whose names end in **–ous**.

Examples:	SO_3^{2-}	sulfite	H_2SO_3	sulfur**ous acid**
	ClO_2^-	chlorite	$HClO_2$	chlor**ous acid**

- Oxyanions whose names end in **–ate** (sulfate, phosphate, nitrate, chlorate, etc.) have corresponding acids whose names are given an **–ic** ending.

Examples:	SO_4^{2-}	sulfate	H_2SO_4	sulfuri**c acid**
	ClO_3^-	chlorate	$HClO_3$	chlor**ic acid**

- *Note that the sulfur containing acids use the root name of "sulfur-" rather than the shorter version "sulf-" used in the anions*. This is an exception and must be memorized. Phosphoric acid has three hydrogens attached to a phosphate ion and is like sulfur in that two syllables of the element name are used to name this acid.

- Table 1. Common Ions

Positive Ions (Cations)	Negative Ions (Anions)
+1 Charge	***-1 Charge***
Group 1A cations	Group 7A anions
Ammonium (NH_4^+)	Acetate ($C_2H_3O_2^-$)
Copper (I) or cuprous (Cu^+)	Cyanide CN^-
Hydrogen (H^+) "proton"	Dihydrogen phosphate ($H_2PO_4^-$)
Silver (Ag^+)	Hydrogen carbonate or bicarbonate (HCO_3^-)
Hydronium ion (H_3O^+)	Hydrogen sulfate of bisulfate (HSO_4^-)
	Hydroxide (OH^-)
	Nitrate (NO_3^-), nitrite (NO_2^-)
	Perchlorate (ClO_4^-), chlorate (ClO_3^-), Chlorite (ClO_2^-), hypochlorite (ClO^-)
	Permanganate (MnO_4^-)
	Thiocyanate (SCN^-)
+2 Charge	***-2 Charge***
Group 2A cations	Group 6A anions
Cadmium (Cd^{2+})	Carbonate (CO_3^{2-})
Chromium (II) or chromous (Cr^{2+})	Chromate (CrO_4^{2-}), dichromate($Cr_2O_7^{2-}$)
Cobalt(II) or cobaltous (Co^{2+})	Hydrogen phosphate (HPO_4^{2-})
Copper(II) or cupric (Cu^{2+})	Oxalate ($C_2O_4^{2-}$)
Iron(II) or ferrous (Fe^{2+})	Peroxide (O_2^{2-})
Lead(II) or plumbous (Pb^{2+})	Sulfate (SO_4^{2-}), sulfite (SO_3^{2-})
Manganese(II) or manganous (Mn^{2+})	
Mercury(I) or mercurous (Hg_2^{2+})	
Mercury(II) or mercuric (Hg^{2+})	
Nickel (Ni^{2+})	
Tin(II) or stannous (Sn^{2+})	
Zinc (Zn^{2+})	
+3 Charge	***-3 Charge***
Aluminum (Al^{3+})	Group 5A anions
Chromium(III) or chromic (Cr^{3+})	Phosphate (PO_4^{3-}), phosphite (PO_3^{3-})
Iron(III) or ferric (Fe^{3+})	Phosphide (P^{3-})
Titanium (III) (Ti^{3+})	
+4 Charge	
Lead(IV) or plumbic (Pb^{4+})	
Tin(IV) or stannic (Sn^{4+})	

Summary of metal cations with more than one possible charge state:
Cu^+, Cu^{2+}; Hg_2^{2+}, Hg^{2+}; Co^{2+}, Co^{3+}; Cr^{2+}, Cr^{3+}; Fe^{2+}, Fe^{3+}; Mn^{2+}, Mn^{3+}; Pb^{2+}, Pb^{4+}; Sn^{2+}, Sn^{4+}

Activity 9 - Nomenclature

Name______________________________

Section____________ **Date**______

Exercise A. Representative Metal + Nonmetal Compounds

1. Name the following:

NaF ____________________________________ CaS ____________________________________

SrI_2 ____________________________________ K_2O ____________________________________

Al_2O_3____________________________________ AlN ____________________________________

2. Give the formulas for the following (refer to the periodic table only):

Cesium phosphide ________________________ Calcium iodide___________________________

Barium fluoride __________________________ Magnesium nitride________________________

Lithium oxide ___________________________ Potassium sulfide________________________

Chloride ion______________________________ Aluminum ion ___________________________

Exercise B. Transition Metal + Nonmetal Compounds

1. Name the following using both naming systems:

Pb^{2+} _________________ Sn^{4+} ________________ Fe^{2+} _________________ Cu^{2+} ________________

_________________ ________________ _________________ ________________

2. Name the following:

$AgCl$ _____________________ $FeBr_3$_____________________ Cu_3N_____________________

3. Referring to question 2 above, what is the charge on the Ag? ________ Fe? __________Cu? _________

4. Give formulas for the following:

Chromium (III) oxide _______ Stannous fluoride __________ Ferrous iodide_____________

Ferric oxide _______________ Cuprous sulfide____________ Plumbic chloride __________

Exercise C. Ionic Compounds with Polyatomic Ions

1. Name the following:

$(NH_4)_2O$ _________________ $CuC_2H_3O_2$________________ Na_2SO_3__________________

$Fe(NO_3)_2$ _________________ $LiSCN$____________________ $NaHCO_3$_________________

1. Give the formulas for the following:

Cupric nitrate_______________ Zinc phosphate ____________ Silver carbonate ___________

Titanium (III) nitride ________ Mercury (II) cyanide________ Lead(IV) acetate___________

Potassium dichromate _______ Barium permanganate_______ Cadmium sulfate __________

Sodium chlorate____________ Cobalt (II) nitrite___________ Ammonium Phosphide______

Exercise D. Molecular compounds

2. Name the following:

SO_3 ________________________________ N_2O_5 ________________________________

N_2O_4 ________________________________ CO ________________________________

CO_2 ________________________________ Cl_2O ________________________________

P_2O_5 ________________________________ N_2 ________________________________

3. Give the formulas:

Bromine trichloride ______________________ Gallium nitride__________________________

Oxygen difluoride ______________________ Carbon tetrachloride _____________________

Sulfur hexafluoride ______________________ Silicon dioxide__________________________

Iodine pentabromide _____________________ Chlorine trifluoride ______________________

Hydrogen______________________________ Dibromine monoxide_____________________

4. Circle any of the common names that require memorization. The compounds marked in bold are those most commonly memorized, ask your instructor to specify the ones you will be tested on.

Methane, CH_4	**Water, H_2O**	**baking soda (Sodium bicarbonate), $NaHCO_3$**
Ethane, C_2H_6	**Ammonia, NH_3**	lye (Sodium hydroxide), NaOH
Propane, C_3H_8	Acetylene, C_2H_2	table salt (Sodium chloride), NaCl
Butane, C_4H_{10}	**Hydrogen peroxide, H_2O_2**	**Methanol (wood alcohol), CH_3OH**
Benzene, C_6H_6	**Ethanol (grain alcohol), C_2H_5OH**	

molecular elements: P_4, S_8, H_2, O_2, F_2, Br_2, I_2, N_2, Cl_2

Exercise E. Acids

5. Give the formula and name for the corresponding acids of the following anions.

Anion	Formula of anion	# of H^+ required to neutralize charge	Formula of Acid	Name of acid
Sulfide				
Carbonate				
Oxalate				
Phosphate				
Acetate				
Nitrite				

6. List of common acids (*ask your instructor to specify the ones you will be tested on*). Acids in boldface are STRONG acids/STRONG electrolytes.

HCl(aq)	**Hydrochloric acid**	HF(aq)	Hydrofluoric acid	H_3PO_4	Phosphoric acid
HBr(aq)	**Hydrobromic acid**	**HNO_3**	**Nitric acid**	H_2SO_3	Sulfurous acid
HI(aq)	**Hydroiodic acid**	**H_2SO_4**	**Sulfuric acid**	$HC_2H_3O_2$	Acetic acid

Activity 10 – Internet Exercise: Lewis Structures

Goals

- ❑ Draw Lewis Structures of molecules and polyatomic ions
- ❑ Determine the shape of molecules and polyatomic ions
- ❑ Discern between polar and non-polar bonds

Concepts to Review

Formulas
Ionic and covalent bonds
Naming ionic and covalent compounds

Instructions

Fill in your answers on the worksheet. To answer the questions that follow, use the mouse to zoom and move the models.

The first two pages of this handout are instructions on writing Lewis dot structures and using these dot structures to assign shapes. Your text is also a useful reference in completing this exercise.

An answer is provided for each problem identified by an asterisk next to the problem number. You should do these problems first, check your answers, and then do the rest.

These exercises were written by Harry Ungar and Linda Calciano and coded by the team of graduate and undergraduate students of the C4 Project at Cabrillo College. Supported by the National Science Foundation.

http://c4.cabrillo.edu/chem30a/exercises/Exer_1/

*Note that the above address contains an underscore between Exer and 1 (…/Exer_1/).

Writing Lewis Dot Structures

The Lewis structure of a covalent compound is important because it depicts all the valence electrons, those involved in chemical change. This will allow us to predict the shape of a molecule using VSEPR theory. A valid Lewis structure must have two qualities:

- The total number of valence electrons represented as bonds (shared) or nonbonded electrons (lone pairs). Note: The charge of a molecular ion will change the total number of electrons. For polyatomic cations, subtract one electron for every positive charge; for polyatomic anions, add one electron for every negative charge.
- All H atoms must share a duet of electrons and other atoms must share an octet of electrons. There are many exceptions to the octet "rule" but you will not be tested on them.

Steps for Writing a Correct Lewis Structure From a Formula

1. From the molecular formula and periodic table, determine the total number of valence electrons.
2. Choose the central atom by choosing the atom with the greatest bonding potential (single dots) or the least electronegative atom. (A chart of electronegativities is included in your supplemental materials at the back of your lab manual.) Hydrogen atoms are NEVER the central atom. Why? Halogens are similar to hydrogen usually found on the perimeter of the molecule.
3. Arrange the remaining atoms about the central atom and connect with lines representing a pair of shared e^-s (a single bond).
4. Assess how many electrons were used in bonding (connecting) and add any remaining electrons as nonbonded pairs.
5. Check the total number of electrons and the duets/octets.
6. If you don't have enough electrons to satisfy all the octets/duets, considering using multiple bonds—putting more than one pair of e-s between two atoms. Two shared pairs is called a double bond, three shared pairs are called a triple bond.

Assigning Shape to Molecules and Molecular Ions

Molecular species have certain geometries due to the arrangement of atoms about some central atom. According to the **Valence Shell Electron Pair Repulsion Theory (VSEPRT)**, the shape of a molecule or molecular ion will be whatever arrangement reduces repulsive forces between electron pairs. Bond angles are used to represent the distance between shared pairs (bonds)—the greater the bond angle, the greater the distance. Perfectly predictable or ideal bond angles only occur when all of the atoms attached to the central atom are the same and there aren't any nonbonded pairs of electrons attached to the central atom.

In general:

- Nonbonded pairs exert a greater repulsive force than bonded pairs.
- Bonded pairs are represented by lines. Double lines indicate double bonds consisting of 2 shared pairs. Triple lines indicate triple bonds, 3 shared pairs.
- Triple bonds exert a greater repulsive force than double or single and double bonds exert a greater repulsive force than single bonds.
- Nonbonded pairs and multiple bonds in a molecule lead to deviations in the ideal bond angle. E.g., the H-O-H angle in water is ~104 instead of 109; the H-C-O angle in H_2CO is greater than the H-C-H angle in that same molecule.

To assign shape using VSEPRT :

1. Write the correct Lewis dot structure. (Refer back to notes or text if necessary.)
2. Count the number of electron groups or regions by counting the number of atoms and the number of unshared/nonbonded pairs about the central atom of the molecule.
3. Arrange the nonbonded pairs and bonded atoms in such a way as to minimize electron pair epulsion.
4. Assign a molecular shape by looking at the shape formed by the bonded atoms only. You must know all the shapes by name – drawing them is not sufficient. (Learn the table below.)

Table 1. Molecular Shape as defined by the number of Electron Groups about the Central Atom

#Electron regions or groups	Bonded Atoms	Nonbonded e- Pairs	Molecular Shape	Bond angle (ideal or approx.)
2	2	0	Linear	180^o (ideal)
3	3	0	Trigonal Planar	120^o (ideal)
3	2	1	Angular or Bent	<120^o (approximate)
4	4	0	Tetrahedral	109.5^o (ideal)
4	3	1	Trigonal Pyramidal	<109.5^o (approximate)
4	2	2	Angular or Bent	<109.5^o (approximate)

Practice Steps 1-4 on the following series of polyatomic ions. Check the shape you assigned with the shape given in the table below.

Table 2. A Series of Chlorine Oxyanions: Same number of Electron Regions but Different Shapes.

Species	A= central atom X= bonded atom E= e- pair	Molecular Shape
Hypochlorite ion	AXE_3	Linear
Chlorite ion	AX_2E_2	Bent
Chlorate ion	AX_3E	Trigonal Pyramidal
Perchlorate ion	AX_4	Tetrahedral

Activity 10 - Internet Exercise: Lewis Structures

Name____________________________

Section____________ **Date**_______

Website address: http://c4.cabrillo.edu/chem30a/exercises/Exer_1/

Read and answer each question carefully and thoroughly. Please do not crowd your answers. Use scratch paper to work out each problem and write the final answer on this worksheet. Those parts of an exercise which have on-screen answers are identified by an asterisk; e.g., 1.2*. You should do these problems first, check your answers at the bottom of the web page, and then do the rest.

Exercise 1

a. Record the bonding arrangement of the atoms (the molecular framework) using stick bonds and element letters. Rotate the molecule using by holding down the shift key while moving the mouse to make sure you have seen all the atoms.

b. Very briefly describe the shape of each molecule; e.g., tetrahedral. For the more complex molecules, describe the shape around the most central atom, i.e. the one with the greatest number of bonds. Either C atom is central in Models 1.6 and 1.9.

c. Draw valid Lewis structures for the molecules. Note there are several exceptions to the octet rule: less than 8 electrons around boron and more than 8 around some sulfur and phosphorus compounds. Using the bonds depicted to help you add the appropriate number of lone (nonbonded) pairs.

1.1 Shape:_________________	1.2* Shape:____________________ ***(exception to octet rule)***	1.3* Shape:_________________
1.4 Shape:_________________	1.5 Shape:____________________	1.6 Shape:_________________
1.7 Shape:_________________	1.8 Shape:____________________	1.9 Shape:_________________
1.10 Shape:____________________ ***(exception to octet rule)***	1.11 Shape:____________________ ***(exception to octet rule)***	

Exercise 2

a. Record the bonding arrangement of the atoms in these molecular ions. Record the charge of the ion also as you will need this in part c. **Rotate the molecule using the mouse by click-dragging on the molecule to make sure you have seen all the atoms.**

b. Describe the shape of each ion; e.g., tetrahedral. For the more complex ions, describe the shape around the most central atom.

c. Draw correct Lewis structures.

2.1* Shape:__________________	2.2 Shape:__________________	2.3* Shape:__________________
2.4 Shape:__________________	2.5* Shape:_________________	2.6 Shape:__________________ ***(exception to octet rule)***
2.7 Shape:__________________ ***(exception to octet rule)***		

Exercise 3

a. Classify the bonds in these molecules as polar or nonpolar. Circle the correct answer and then write the chemical formula.

3.1* polar / nonpolar Formula:__________________

3.2 polar / nonpolar Formula:__________________

3.3 polar / nonpolar Formula:__________________

3.4 polar / nonpolar Formula:__________________

3.5 polar / nonpolar Formula:__________________

Exercise 4

a. Draw correct Lewis structures for the models.

b. Show the charge distribution in them.

c. Predict which molecules are polar. (Circle the correct answer.) **The polarity** of a molecule means that a molecule will visibly respond to the application of an electric field or any other polarizing force. Whether or not a species is polar can be readily assigned by viewing a 3-D model of a species. If all of the "things" about the central atom (could include bonded atoms and nonbonded pairs) are exactly the same, the molecule is nonpolar. If any side of the central atom experiences is different from another, the molecule is most always polar. For example, water is extremely polar and carbon tetrachloride is not. Ions are incredibly responsive to an electric field because they are charged, what you might call "beyond polar".

4.1* Polar? Y N	4.2 Polar? Y N	4.3 Polar? Y N
4.4 Polar? Y N	4.5 Polar? Y N *(check central atom to make sure!)*	

Exercise 5

a. Name the compounds.

5.1* ____

5.2 ____

5.3* ____

5.4 ____

5.5 ____

Exercise 6

a. Name the ions.

6.1* ____________________________ 6.4 ____________________________

6.2 ____________________________ 6.5 ____________________________

6.3* ____________________________

Exercise 7

a. Write the formula for each ionic compound that can be formed by combining the ion model on the left with each of the ions in the list to the right.

7.1* Write the formula of the negative ion displayed **here** ____________. Now write the formula of the compound made by combining the negative ion plus:

Potassium: ______________ Magnesium: ______________

Calcium: ______________ Aluminum: ______________

Sodium: ______________

7.2 Write the formula of the positive ion displayed **here**__________. Now write the formula of the compound made by combining the positive ion plus:

Chloride: ______________ Sulfate: _______________

Sulfide: ______________ Phosphate: _______________

Exercise 8

a. Match the model with its name. Not all names will be used.

b. Give the total charge on the ion.

	a. Model Name	b. Charge
8.1*	________________________	_______
8.2	________________________	_______
8.3	________________________	_______
8.4	________________________	_______
8.5*	________________________	_______
8.6	________________________	_______
8.7	________________________	_______
8.8	________________________	_______
8.9	________________________	_______

Names:
Acetate
Ammonium
Bicarbonate
Carbonate
Chlorate
Chloride
Cyanide
Dihydrogen phosphate
Hydrogen carbonate
Hydrogen phosphate
Hydrogen sulfate
Hydrogen sulfite
Hydronium
Hydroxide
Nitrate
Phosphate
Sulfate

Activity 11 - Energy Changes of Solution Formation

Goals

- Measure and record energy changes in the dissolution of a salt in water.
- Describe and explain energy changes in the dissolution process in terms of changes in entropy and enthalpy.
- Observe the change in freezing point of water upon addition of a solute.

Pre-Lab Lecture Questions. *Answer these questions on a separate sheet using complete sentences.*

1. What is the relationship between temperature and heat? Explain your answer.
2. Chemical and physical changes involve changes in energy. Give an example of a change that requires an input of energy and an example of a change that produces energy.
3. After reading the introduction section, define the properties represented by the symbols of T, G, H and S. What does "Δ" mean?
4. What does it mean if a reaction is spontaneous? How does spontaneity relate to Gibbs free energy?
5. Which of the following changes include increases in entropy?
    a. ATP + H_2O → ADP + phosphate ion + hydronium ion
    b. One starch molecule → 101,000 glucose molecules
    c. $N_2 + 3H_2 \rightarrow 2NH_3$
6. Is the making of ice cream increasing or decreasing entropy? If the ice cream ingredients mixed together is the chemical system, is it an exothermic or endothermic process to convert them to ice cream?

Introduction

Why are some changes spontaneous while others are not? Spontaneous reactions occur because this change will lead to an increase in the entropy (S) of the universe. A helpful picture is to think of entropy as a measure of disorder or chaos. In the context of any process under consideration, the universe can be separated into two components. One is the **system,** the portion of the universe that is the focus of attention (such as a beaker containing 100 mL of water and 2.0 g of NaCl). The other is the **thermal surroundings,** which is simply a source or sink of thermal energy (heat). **Note that the net entropy change** (ΔS) **in any process is the sum of the entropy changes in the system <u>and</u> the surroundings.**

There are two aspects to consider when assessing whether or not a process under consideration is spontaneous. One has to do with the organization of the materials under consideration. If the materials have become more "disorganized", **entropy** has increased and this aspect of the process favors spontaneity. Examples are a solid or liquid dissolving in a solvent, or a chemical reaction in which the number of product molecules is larger than the number of reactant molecules. In both these cases the chemical species under consideration have become more disorganized. (Another way to think of this is that the number of ways to arrange the components of the matter under consideration has increased without changing the appearance of the collective. For an apt example, the number of ways to arrange the components of an aqueous solution of NaCl without changing the appearance of the solution in any way far exceed the number of ways to arrange the components of solid NaCl and liquid water as separated phases.)

The other aspect of a process that has an effect on spontaneity is the transfer of thermal energy (heat) into or out of the system. When this transfer occurs under a condition of constant pressure, the quantity of heat transferred is called the **enthalpy change** (ΔH) associated with the process. Recall that the thermal surroundings are either a source or sink of heat. The entropy of the surroundings always increases when heat is transferred to it from the system (such processes are termed **exothermic;** ΔH is negative for these), and always decreases when the transfer of heat goes in the opposite direction (such processes are termed **endothermic;** ΔH is positive for these). This means that if a spontaneous reaction occurs in the system that is exothermic, the entropy increases in both the system and the surroundings. However when such a reaction is endothermic, whether or not it is spontaneous depends on the relative magnitudes of the entropy increase in the system and the entropy decrease in the surroundings, with the greater magnitude determining the outcome. Thus it is possible for an endothermic reaction to be spontaneous, but only if the increase in the entropy of the system is greater than the decrease in the entropy of the surroundings.

Any process for which ΔS is positive and ΔH is negative will lead to a spontaneous process, since the entropy of both the system and the surroundings has increased. However if only one of these driving forces is present, the spontaneity of the process depends on the relative magnitudes of each change. An American mathematician named J.W. Gibbs developed a relationship that considers both changes in the entropy of the system and that of the surroundings, but uses only quantities related to the system. It makes use of a term known as the *Gibbs free energy* (G). The relationship developed by Gibbs for a chemical process is as follows:

$$\Delta G = \Delta H - T\Delta S$$

In the above expression $\Delta G = G_{prod} - G_{react}$, $\Delta H = H_{prod} - H_{react}$, and $\Delta S = S_{prod} - S_{react}$. A negative Gibbs free energy change ($\Delta G < 0$) means that the products are lower in overall free energy than the reactants. A reaction with a negative ΔG is called e*xergonic* and is spontaneous. Both terms in the above relationship have units of energy and the overall change in G is considered the amount of energy available to do work (hence the term "free energy").

Many ionic compounds (salts) are soluble in water. For these the dissolution process is spontaneous and therefore ΔG is negative. To observe the contributions of enthalpy and entropy to ΔG, different salts will be dissolved in water and the final temperature of the solution will be measured. The typical ice/table salt mixture used in ice cream freezers will also be investigated.

Experimental Concerns

Use only small amount of quantities as described.

Use comparable amounts of salts.

Safety

Wear safety glasses at all times!
Act in accordance with the laboratory safety rules of Cabrillo College.
Avoid contact with all chemical reagents and dispose of reactions using appropriate waste containers.

Materials

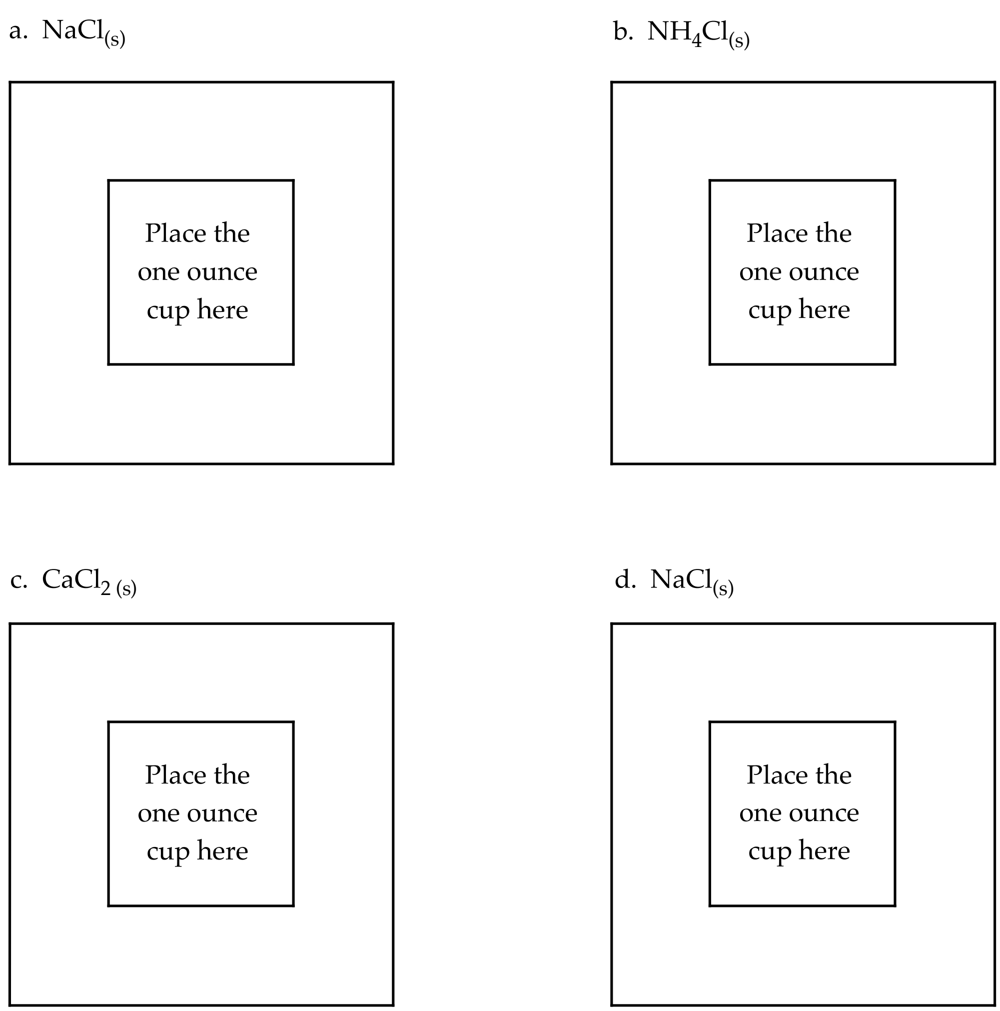
Rock salt, sodium chloride
Calcium chloride
ammonium chloride
crushed ice

Equipment: Lab top reaction surface, plastic spoon, 4 (1 oz) cups, alcohol thermometer

Reaction Template: Insert this page into the labtop.

Arrange four plastic cups according to the diagram. Do each experiment one at a time in the following way:

1. Transfer two level spoonfuls of deionized water into the cup. *(Use crushed ice in part (d).)*
2. Record the initial temperature of the water. Make sure you read the temperature to the limit of the device. (Hint: The last decimal place in the measurement is an estimated one.)
3. Add one level spoonful of the indicated solid chemical.
4. Stir gently and measure the highest or lowest temperature reached. Record this temperature as the final temperature.



Cleaning up

Avoid contamination by cleaning up in a way that protects you and your environment. Pour the contents of the cups down the drain, rinsing them with plenty of water. Rinse and dry the thermometer

before returning. Dispose of all paper towels in the waste bin. Wash your hands with soap and water before leaving.

Activity 11 - Energy Changes in Solution Formation

Name____________________________

Section____________ **Date**_______

Experimental Data and Calculations: Record your initial and final temperatures and then determine the overall change in temperature. Include the appropriate sign for the change in temperature.

Experimental Mixture	**Final** Temperature (^oC)	**Initial** Temperature (^oC)	$\Delta T = T_f - T_i$ (^oC)
a. $NaCl(s) + H_2O(l)$			
b. $NH_4Cl(s) + H_2O(l)$			
c. $CaCl_2(s) + H_2O(l)$			
d. $NaCl(s) + H_2O(l)$			

Questions and Problems

Use your results in the table above to answer the following questions.

1. Upon mixing, which mixture(s) cooled off?

2. Which mixture(s) warmed up?

3. Which mixture had the least temperature change?

When sodium chloride dissolves in water, the ions dissociate or come apart in solution:

$$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$$

4. Write ionic equations like the one above to show how ammonium chloride **and** calcium chloride dissociate when they dissolve in water.

5. Is the dissolving of ammonium chloride an exothermic or endothermic process? Explain. Rewrite the ionic equation from Question 4, but this time include "heat:" on the appropriate side of the equation.

6. Is the dissolving of ammonium chloride a spontaneous process? Does it occur with an increase or decrease in entropy? (Consider both the system and the surroundings.) What is the driving force for spontaneous change here?

7. Is the dissolving of calcium chloride an exothermic or endothermic process? Explain. Rewrite the ionic equation from Question 4, showing heat as a reactant or product.

8. Is the dissolving of calcium chloride a spontaneous process? Does it occur with an increase or decrease in entropy? (Consider both the system and the surroundings.) What is the driving force here?

9. Is the dissolving of sodium chloride an exothermic or endothermic process? Explain. Is it spontaneous? What is the driving force here?

10. Describe your observations when you mix sodium chloride with liquid water on the one hand, and solid water (ice) on the other. Any differences between the two results?

11. The principle behind an ice cream freezer is that when salt is added to the ice, the melting ice absorbs heat from the ice cream, and the ice cream freezes. What drives the spontaneous melting of ice when it is mixed with salt?

Activity 12 - Balancing Chemical Equations

Name____________________________

Section____________ **Date**_______

Tips for Writing Chemical Equations:

a. Make certain all formulas are correct.

b. Once all formulas are written correctly one may **<u>not change the subscripts</u>**, only the coefficients in order to balance the equation. Always choose the lowest whole number coefficients.

c. The symbols (s), (l), and (g) indicate the phase of each reactant or product: solid, liquid and gas, respectively.

d. Some elements exist in nature as diatomic molecules. The element names correspond to the diatomic formula because this is the elemental structure. These elements include hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine and iodine. The correct formula for the elemental name is diatomic not monatomic. A mnemonic device that may help you remember these elements is the name Hofbrincl H_2 O_2 F_2 Br_2 I_2 N_2 Cl_2.

Questions and Problems

1. Balance the following chemical equations:

a) $CH_4 + O_2 \longrightarrow CO_2 + H_2O$

b) $NaHCO_3 \longrightarrow Na_2CO_3 + CO_2 + H_2O$

c) $CH_4O + O_2 \longrightarrow CO_2 + H_2O$

d) $C_6H_6 + O_2 \longrightarrow CO_2 + H_2O$

e) $FeCr_2O_4 + C \longrightarrow Cr + Fe + CO$

f) $C_3H_8 + O_2 \longrightarrow CO_2 + H_2O$

g) $B_2H_6 + O_2 \longrightarrow B + H_2O$

h) $Ca(ClO_3)_2 \longrightarrow CaCl_2 + O_2$

i) $MnCl_2 + Al \longrightarrow Mn + AlCl_3$

j) $B_2O_3 + C \longrightarrow B_4C_3 + CO_2$

k) $Ca(ClO_3)_2 \longrightarrow CaCl_2 + O_2$

l) $H_2O_2 \longrightarrow H_2O + O_2$

2. Write balanced chemical equations for each of the following reactions:

a) hydrogen + oxygen $\longrightarrow$ water

b) potassium + water $\longrightarrow$ potassium hydroxide + hydrogen

c) fluorine + sodium chloride $\longrightarrow$ sodium fluoride + chlorine

d) aluminum + oxygen $\longrightarrow$ aluminum oxide

e) mercury (II) nitride $\xrightarrow{\Delta}$ mercury + nitrogen

f) The combustion of propane (C_3H_8)

Activity 13 - Measuring Mass: A Means of Counting[1]

Goals

- Properly use a top loading balance to determine the mass of a sample.
- Use molar masses to connect the measured mass of a sample to the number of particles in that sample.
- Use safe lab techniques to characterize matter.
- Apply principles to samples of both pure substances and mixtures.
- Apply dimensional analysis techniques to count small particles such as atoms and molecules.

Pre-Lab Lecture Questions. *Answer these questions on a separate sheet using complete sentences.*

1. What is the difference between weight and mass? How do you "properly" use a balance in the laboratory?
2. What determines the number of significant figures/digits in a measurement?
3. What determines the number of significant figures/digits in a calculation?
4. What is molar mass?
5. What is Avogadro's number?
6. Write as many different conversion factors as you can using the chemical formula of water, the molar mass of water, the definition of a mole, and Avogadro's number.
7. Read through the experimental procedure and classify substances as either pure or a mixture.

Concepts to Review

Classification of Matter: What is a pure substance (element, atom, molecule, compound?) and what is a mixture?
Significant Figures/Digits
Chemical Formulas
Unit Conversion Methods (Dimensional Analysis describing Atoms, Molecules and Ions)

Introduction

Our world contains groupings of objects everywhere: a dozen eggs, a pair of socks, a gross of pencils. These collections are convenient "packets" of individual pieces. The individual "pieces" of pure substances can be described by chemical formulas, e.g., H_2O is the chemical formula for water. This formula indicates that each molecule of water consists of two atoms of hydrogen combined with one atom of oxygen. The mass of this molecule is the sum of the masses of the atoms combined to form this compound. We cannot directly measure the mass of one molecule of water but we can recognize its relative mass and use a convenient "packet" of molecules to describe real world quantities. The **mole** is the chemist's standard collection of particles and is defined **as the amount of substance in a sample that contains as many units as there are atoms in exactly 12 grams of carbon-12**. That number of carbon-12 atoms is $\mathbf{6.022 \times 10^{23}}$ and is known as **Avogadro's number**.

1 mole carbon atoms = 12.0 g C = 6.022×10^{23} atoms C

1 mole H_2O = 2(1.008 g H) + 1(16.00 g O) = 18.02 g H_2O = 6.022×10^{23} molecules of water

Using these relationships, any mass of water can be converted into a number of molecules:

[1] Adapted from: Waterman, E. L. *Chemistry: Small-Scale Chemistry Laboratory Manual;* Addison-Wesley/Prentice-Hall, Inc.: Upper Saddle River, New Jersey, 2002; pp 59-62.

$$100.00 \text{ g } H_2O \left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right)\left(\frac{6.022 \times 10^{23}}{1 \text{ mol}}\right) = 3.34 \times 10^{24} \text{ molecules } H_2O$$

In this lab you will measure amounts of substances. You will then calculate the number of particles contained in the sample, numbers that cannot be counted—only calculated.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact with all chemical reagents and dispose of reactions using an appropriate waste container.

Materials:

Reagent Central solutions include:

Sucrose ($C_{12}H_{22}O_{11}$), sodium chloride (NaCl), chalk (calcium carbonate)

✔ **Check out** a sample containing:

Glass slides (assumed to be pure silicon dioxide), polystyrene peanuts, sulfur, fluorite, hematite, (or other minerals as provided by stockroom)

Equipment: Balance Plastic spoons

Experimental Procedure

1. Using a weighing paper or boat and balance, "weigh" one level teaspoon of sodium chloride and record its mass in your laboratory notebook and/or Table 1. This mass is the mass of your "sample." Using the same balance, measure the mass of one teaspoon of water and one of sucrose.
2. "Weigh" a glass slide, and record its mass in your laboratory notebook and/or Table 2. Repeat for the piece of chalk and a polystyrene peanut.
3. "Weigh" a piece of sulfur, and record its mass in your laboratory notebook and/or Table 3. Repeat for a piece of fluorite and a piece of hematite.
4. A nickel coin is a mixture of metals called an alloy. It consists of 75% copper and 25% nickel. Design and carry out an experiment to find out how many nickel atoms there are in one 5-cent piece. Record your experiment procedure in your laboratory notebook and/or in Table 4. Show all your calculations and give your final answer with the correct number of significant figures and in scientific notation.

Chemical Calculations

For each of the masses recorded:

1. Use the formula (see below) to determine the molar mass in units of g/mol.
2. Use the molar mass to determine the number of moles.
3. Use the number of moles of the substance and molar ratios to calculate the moles of each element.
4. Use the moles of each element in each sample along with Avogadro's number to calculate the number of atoms of each element.
5. Use the above calculations as a model to help you determine the number of nickel atoms in one 5-cent piece.

   Make sure your worksheet is complete, legible and turned in on time.

Activity 13 - Measuring Mass: A Means of Counting

Name____________________________

Section____________ **Date**_______

Experimental Data and Calculations

1. Complete the following tables:

Table 1. Counting Particles in Common Substances

Formula	**Name**	**Sample Mass (g)**	**Molar Mass (g/mol)**	**Moles in sample**	**Moles each element in sample**	**Atoms each element in sample**
NaCl						
	Water					
$C_{12}H_{22}O_{11}$						

Table 2. Counting Particles in Common Items.

Formula	**Name**	**Sample Mass (g)**	**Molar Mass (g/mol)**	**Moles in sample**	**Moles each element in sample**	**Atoms each element in sample**
SiO_2 (molecule)	Glass slides					
$CaCO_3$ (formula unit)	Chalk					
$C_{8000}H_{8000}$ (molecule)	Polystyrene					

Table 3. Counting Particles in Minerals.

Formula	**Name**	**Sample Mass (g)**	**Molar Mass (g/mol)**	**Moles in sample**	**Moles each element in sample**	**Atoms each element in sample**
S_8 (molecule)						
CaF_2 (formula unit)	Fluorite					
Fe_2O_3 (_______)	Hematite					

Table 4. Counting the Atoms of Nickel in a Nickel

Describe your experimental procedure:
Show all the steps of your calculations and your final answer including the correct number and units:

Activity 14 – Mole Worksheet

Name____________________________

Section____________ **Date**______

Questions and Problems

Solve the following problems, showing at least some of your work and making sure your answer is clearly boxed off. Your final answer should include the correct number of significant figures and the units. Use scientific notation if the answer is greater than 1000 or less than 1. Note: Make sure you have the correct chemical formula before doing any calculations. You will need a periodic table for this exercise.

1. A sample of mercury (II) bromide weighs 7.56 g.
    a. What is the molar mass of mercury (II) bromide?

    b. How many moles are in this sample?

2. What is the mass of 0.81 mol of Ammonium carbonate?

3. A sample of Chlorine gas contains 8.25 moles.
    a. How many ***molecules*** of chlorine are in the sample?

    b. How many chlorine ***atoms*** are in the sample?

4. Calculate the percent by mass of barium in barium sulfate.

5. What is the mass of 4.2×10^{23} molecules of carbon dioxide?

6. Use the equation below to solve the following problems:

$$2\ KMnO_4 \quad + \quad 16\ HCl \longrightarrow 5\ Cl_2 \quad + \quad 2\ KCl \quad + \quad 2\ MnCl_2 \quad + \quad 8\ H_2O$$

a. How many moles of HCl are required to react completely with 1.00 mole of $KMnO_4$?

b. How many moles of chlorine will be produced by 25.0 moles of $KMnO_4$ assuming that an excess of HCl is present?

c. How many moles of water will be produced if 40. g of HCl are completely reacted with excess potassium permanganate?

d. What is the maximum mass of manganese(II) chloride that will be produced if 40. g of HCl are completely reacted with excess Potassium permanganate?

7. A water solution of sulfuric acid (H_2SO_4) has a density of 1.67 g/mL and is 75 percent H_2SO_4 by mass. How many moles of H_2SO_4 are contained in 500. mL of this solution?

8. Cobalt chloride ($CoCl_2$) exists as a hydrate (has non-covalently bound waters of hydration) with a molecular mass of 237.93. Prolonged heating can drive off the waters of hydration. A 54.8 g sample of the hydrate was heated for 15 minutes, cooled and reweighed. The residual mass was found to be 33.2 g. Calculate the number of water molecules associated with each $CoCl_2$ in the hydrate.

Activity 15 - The Molar Mass of a Gas

Objective

The purpose of this experiment is to determine the number of grams per mole of a gas by measuring the pressure, volume, temperature, and mass of a sample.

Terms to Know

- *Molar Mass* - The number of grams per mole of a substance.
- *Ideal Gas Law* - The relation between pressure, volume, the number of moles, and temperature of a gas: PV = nRT
- *Vapor Pressure of Water* - The pressure exerted by water vapor as it saturates the environment above a sample of liquid water. Water vapor pressures depend upon the temperature, and tabulated values are found in reference tables.

Introduction

The molar mass of a gaseous compound can be determined by experiment even though the formula or composition of the compound is not known. In other words, it is possible to find the molar mass of a gas even if we do not know its identity. The molar mass is determined by using the fact that the number of moles of a gas sample is related to the pressure, volume, and temperature by the gas laws. The mass, pressure, volume and temperature of a gas sample are measured experimentally, and the data is used to calculate the molar mass.

To accomplish this, the number of moles in the sample is first calculated using the ideal gas law, PV = nRT. Recall that n represents the number of moles. The number of moles of a gas is found algebraically as:
n = PV/RT where R is the universal gas constant.

As an example, consider a 0.508 g sample of a gas that occupies 522 mL at 100 ºC and 0.960-atmosphere pressure. First summarize the data as:

Mass	P	V	T
0.508 g	0.960 atm	0.522 L	(100.0ºC + 273.2) = 373.2 K

The number of moles of gas is found using the ideal gas law. (Fill in the following blanks and calculate the number of moles.)

$$n = \frac{\underline{\quad\quad}\ \text{atm} \times \underline{\quad\quad}\ \text{L}}{0.0821\ \frac{\text{L atm}}{\text{mole K}} \times \underline{\quad\quad}\ \text{K}} = \underline{\quad\quad\quad}\ \text{mole}$$

The molar mass of the gas is found from the ratio of the mass of the sample to the moles in the sample. Divide the mass of the sample by the number of moles using the value calculated above.

$$\frac{0.508\ \text{g}}{\underline{\quad\quad}\ \text{mole}} = \frac{\underline{\quad\quad}\ \text{g}}{1\ \text{mole}}$$

In this experiment, a sample of gas is collected by water displacement; i.e. the gas bubbles into a container of water and as the gas accumulates, it displaces the water. The volume of the sample is the volume the gas occupies in the container. The temperature is found by measuring the temperature of the water in contact with the gas. The mass of the gas sample is measured by weighing a small tank or cylinder of gas, delivering the sample, and reweighing the cylinder.

The pressure of the sample is found by making sure that the pressure of the gas sample is the same as the atmospheric pressure. The atmospheric pressure is measured with a barometer. However, when a gas is collected by water displacement, it becomes saturated with water vapor. This means that once the gas sample is collected, it will be a mixture of the gas and water vapor. This does not affect the volume or temperature of the gas, but the measured pressure is the total pressure of the gas and the water vapor. To determine the pressure of the gas sample, the pressure of the water vapor must be subtracted from the total pressure. Appendix 2 lists the vapor pressure of water at various temperatures. To find the pressure of the gas sample in the experiment, look up the vapor pressure of water at the measured temperature and subtract this pressure from the measured atmospheric pressure:

$$P_{GAS} = P_{ATM} - P_{WATER}$$

Experimental Procedure

For this experiment you will need a 1-L beaker, a 250-mL Erlenmeyer flask, and a thermometer.
Place a piece of magic tape or a length of a gummed label along the neck of the flask near the top. Fill the beaker to the 750mL mark with tap water and transfer 250mL of this volume into the conical flask, filling it to the brim. Allow the water to stand for several minutes to be sure that it is at room temperature. Set up the apparatus as shown in Figure 1 below.

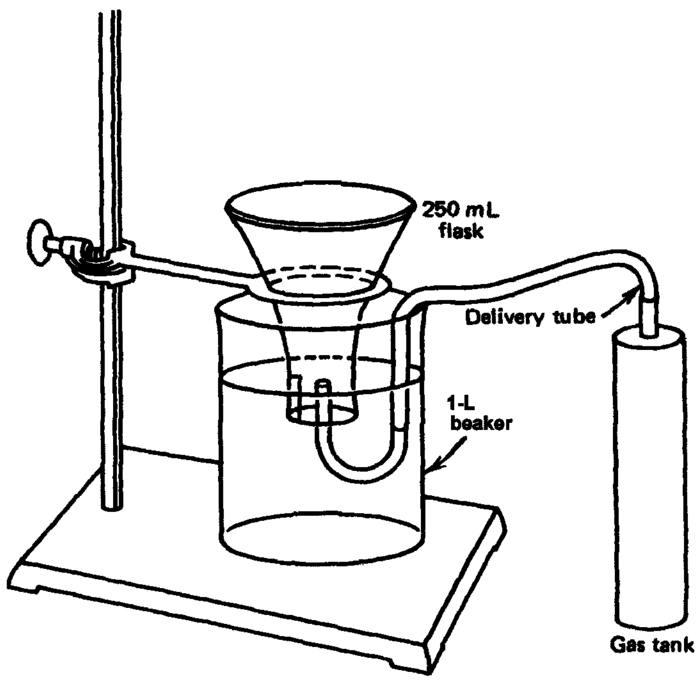


Figure 1. Apparatus for collecting gas samples

Gently stopper the flask with a rubber stopper and do not allow any air to become trapped in the flask under the stopper. Invert the flask through the iron ring that will hold it in position. Lower the ring with the flask so that the mouth of the flask is below the water level in the beaker. Secure the ring to support the flask. Use a spatula to remove the rubber stopper from the mouth of the flask. The stopper can remain in the beaker since it will not interfere with the experiment.

Obtain a small tank of gas. Use a paper towel to gently wipe the tank to make sure that it is clean. Weigh the tank to the nearest 0.01 g.

Attach a plastic delivery tube to the valve on the tank and insert the glass tube on the other end into the mouth of the flask in the beaker of water. Push the valve on the tank straight down to deliver gas to the flask. Continue to deliver gas until the flask is nearly full of gas, i.e. deliver gas until the gas level is near the neck of the flask. Do not allow any gas to escape from the flask. If any gas escapes, you will have to start the experiment over.

Carefully remove the delivery tube from the gas tank, gently clean and dry the tank with a paper towel and reweigh the tank to the nearest 0.01 g. Place a thermometer in the beaker.

Move the flask of gas up or down so that the level of water in the flask is the same as the water level in the beaker. This will adjust the pressure in the flask to atmospheric pressure. Use a pen or pencil to place a mark on the tape to indicate the water level.

Remove the flask from the apparatus and fill it with water to the level corresponding to the mark on the tape. This volume of water will correspond to the volume of the gas sample. Carefully pour the water from the flask into a graduated cylinder and measure the volume to the nearest 1 mL. Record this as the volume of the gas sample.

Record the temperature of the water in the beaker. Assume that the temperature of the gas sample is the same as the water. Record the atmospheric pressure and look up the vapor pressure of water at the temperature of the gas in the table at the end of this experiment. Repeat the experiment to obtain a second set of data. Do all your calculations before returning your equipment.

Calculate the molar mass of the gas using each set of data. First obtain the mass of the gas by subtracting the mass of the can after the sample was removed from the mass of the can before removing the sample. Calculate the pressure of the gas by subtracting the vapor pressure of water from the total atmospheric pressure. The vapor pressure of the water vapor in your gas sample is dependent on temperature and is found in the following table. If your temperatures are not exactly the same as those in the table, calculate the vapor pressure for your temperatures as measured to the tenths place on your lab thermometers. For example, at a temperature of 15.2 ^{o}C, the vapor pressure of water is between 12.8 and 13.6 torr. Since one degree Celsius (16 ^{o}C -15 ^{o}C) in this range equals 0.8 torr (13.6 torr-12.8 torr), 0.2 degree Celsius corresponds to an additional 0.16 torr, so that the vapor pressure of water at 15.2 ^{o}C is 13.0 torr (12.8 torr + 0.16 torr = 12.96 torr, or 13.0 torr).

Table 1. Vapor pressure of water as a function of temperature

Temperature (°C)	**Vapor Pressure (torr)**	**Temperature (°C)**	**Vapor Pressure (torr)**
0	4.6	25	23.8
5	6.5	26	25.2
10	9.2	27	26.7
15	12.8	28	28.3
16	13.6	29	30.0
17	14.5	30	31.8
18	15.5	40	55.3
19	16.5	50	92.5
20	17.5	60	149.4
21	18.6	70	233.7
22	19.8	80	355.1
23	21.2	90	525.8
24	22.4	100	760.0

Again, make sure your calculated answers (A-C) have significant figures that are consistent with the original data. Then convert your gas sample quantities into units that are consistent with the gas constant, R. Use algebra and the ideal gas law to calculate the molar mass. You can solve for the number of moles, n, and then divide the gram quantities by this number to get the molar mass. Another option is to derive the molar mass in units of g/mol directly from ideal gas law. If the results of the two trials are within 10 percent of one another, express your answer as the average of the two results.

If your two results differ by more than ±10%, run a third trial. Consult your instructor to see if you should include all three trials in your average.

Activity 15 – Molar Mass of a Gas

Name____________________________

Section____________ Date_______

Experimental Data: Record all your observations.

	Trial 1	Trial 2	Trial 3
Raw Data			
Mass of tank (g)			
Mass of tank – sample (g)			
Mass of sample (g)			
Volume of sample (mL)			
Temperature of sample (°C)			
Atmospheric pressure (torr)			
Vapor pressure of water (torr)			
Pressure of sample (torr)			
Conversion of Data to Appropriate Units			
Mass of sample (g)			
Volume of sample (L)			
Temperature of sample (K)			
Pressure of sample (atm)			
Calculate Values			
Moles of sample (mol)			
Molar Mass of sample (g/mol)			

Show a sample calculation for one trial and calculate the average molar mass:

Average Molar Mass (g/mol) ____________

Questions and Problems

1. Referring to the experimental determination of the molar mass, explain how and why each of the following factors would affect your calculated molar mass. Would the calculated value be greater than it should be, less than it should be, or not changed? Note that the relationships between the various factors involved in the calculation are:

$$\text{Molar mass} = \frac{g}{mol} = \frac{g}{n} = \frac{g}{\left(\frac{PV}{RT}\right)} = \frac{gRT}{PV}$$

   a. The measured temperature is a lower value than the actual temperature.

   b. The measured volume is a higher value than the actual volume.

   c. Some of the gas sample escapes from the tank before it reaches the flask.

2. The molar mass of a gas is determined by collecting a gas sample by water displacement.

   a. Using the following data, calculate the molar mass of the gas: Sample volume, 163 mL; temperature, 21.0 ^{o}C; mass, 0.281 g; total (i.e., barometric) pressure, 752 torr.

   b. The gas in part (a) contains 85.5% C and 14.5% H. First, determine the empirical formula, and then use the result from part (a) to determine the actual formula. First step: Assume 100. gram sample.

Activity 16 - Synthesis and Qualitative Analyses of Gases[1]

Goals

- Carry out chemical reactions that produce gases.
- Employ indicators to assist in the detection of gases.
- Deduce the identity of unknown gases by examining the chemical reactions of the reactants used to generate them.

Pre-Lab Lecture Questions. *Answer these questions on a separate sheet using complete sentences.*

1. What is the difference between a gas, a liquid and a solid at the molecular level?
2. What is observed when two solutions are mixed and a gas is produced?
3. Give the names and formulas for three gases that are part of our everyday lives.
4. What gas is the major component of air? Give its name and formula.
5. What gases are responsible for acid rain and where do they come from?
6. What experimental difficulties arise when the products of your chemical reactions include gases?
7. What is the chemical formula for bleach?

Concepts to Review

Indicators of Chemical Change
Nomenclature
Writing Chemical Equations

Introduction

We are swimming in a vast solution of gases as we move through our daily lives. Our bodies need oxygen from the atmosphere to live and we generate carbon dioxide as an end product of metabolism. Marsh gas or methane is also a natural product of the decomposition of biological matter. One of the most obvious visual evidences of chemical change is the production of bubbles-insoluble gases moving through a liquid sample.

The generation of gases is fundamental to all chemistry, both "natural" and industrial. The combustion of fossil fuels produces sulfur dioxide and the oxides of nitrogen. Both of these oxides are implicated as sources of pollution, with the nitrogen oxides being responsible for the brown haze we see and both of them reacting with atmospheric moisture to produce acid rain. Even something that can be a "natural product" (carbon dioxide) may cause environmental problems such as the greenhouse effect when its concentrations are unusually high.

In this lab you will mix six different combinations of reactants in the presence of four different indicators. The chemistry for each of the reactions includes both double displacement reactions (acid/base) and oxidation-reduction reactions. Similarly the chemical changes in the indicator drops include both double displacement and oxidation-reduction reactions. You are not expected to analyze or deeply understand the mechanism of the reactions but you are expected to make careful observations. Take a moment to identify the composition of the indicators. How are they similar? How are they different? A critical experimental concern in this experiment is that your products are gases and must be contained in order to be identified. Read through the experimental procedure carefully before beginning the lab.

[1] Adapted from: Waterman, E. L. *Chemistry: Small-Scale Chemistry Laboratory Manual;* Addison-Wesley/Prentice-Hall, Inc.: Upper Saddle River, New Jersey, 2002; pp 103-110.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact with all chemical reagents and dispose of reactions using appropriate waste containers.
Use microburets to dispense reagents in such a way that they do not make contact with other drops on the reaction surface.
Return any contaminated microburets to your instructor.

Materials:

Reagent Central chemicals include both indicator reagents and reactants:

Indicators	Chemical Reactants (center stage)	
Potassium iodide (KI)	Hydrochloric acid (HCl)	Sodium hydrogen sulfite ($NaHSO_3$)
KI + starch		Sodium nitrite ($NaNO_2$)
Bromthymol Blue indicator (BTB)		Sodium hydrogen carbonate ($NaHCO_3$)
KI + starch + sodium hypochlorite (NaOCl)		Sodium hypochlorite (NaOCl) "bleach"
	Acidic potassium permanganate ($KMnO_4$ + HCl)	Hydrogen peroxide (H_2O_2)
	Sodium hydroxide (NaOH)	Ammonium chloride (NH_4Cl)

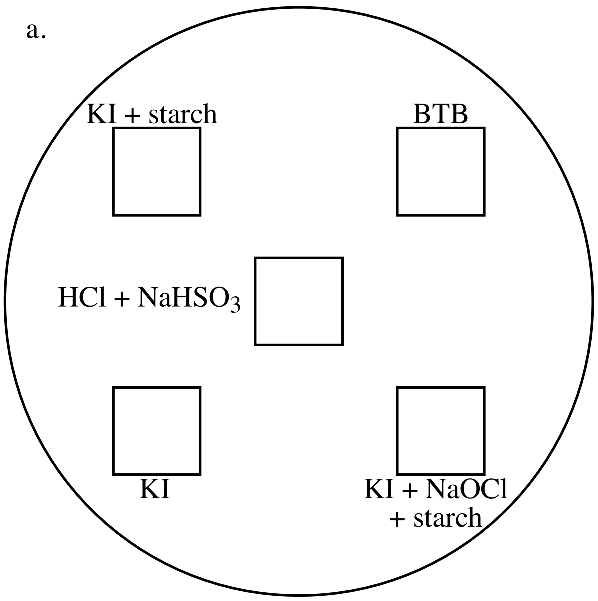
Equipment: Clear plastic cups Labtop reaction surface

Experimental Procedure

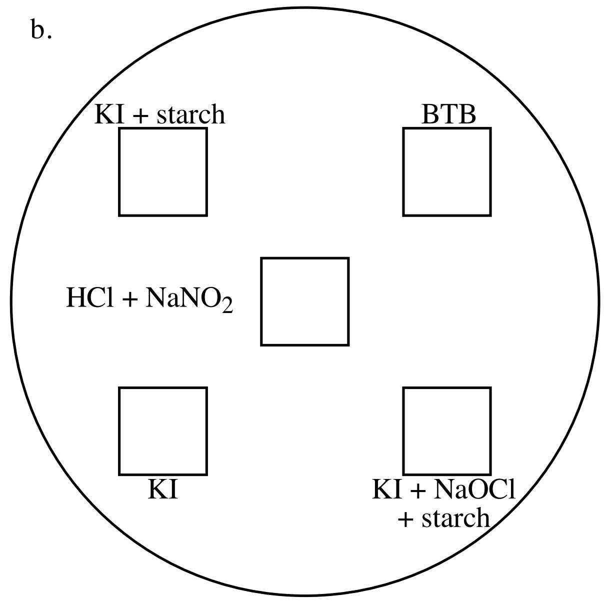
1. Insert your experimental page under your reaction surface such that the circles are visible. Place 1 drop of each indicator solution in the squares near the inside edge of each circle as depicted on your Experimental Template. When you are ready to make observations, mix 1 drop of each reactant (HCl + something else) in the center of each circle, and cover the entire circle with a cup. Be sure the cup does not touch any of the solutions.
2. Continue to observe the reactions in the center and in the indicator drops for few minutes, recording anything you see in your laboratory notebook and/or data page. Take care to notice any differences between the bubbles (size, how quickly they disappear, position), recording everything you see. Remember, "no visible change" or "no reaction" is an observation, a potentially important piece of evidence.
3. After all the reactions have gone to completion, neutralize the central reactions with a few drops of NaOH; clean the reaction surface thoroughly with a *damp* paper towel and then dry it. Clean the cups with a *dry* paper towel to absorb any stray moisture. Repeat this process to assure a clean surface for the next two reactions.
4. Continue onto the next experimental template and complete Steps 1-3 for the last two reactions.
5. Devise a method of detection for carbon dioxide gas. This is found in exhalations, as carbon dioxide is a waste byproduct of metabolism. Carbon dioxide combines with water to form carbonic acid. Test your method using any of the indicators available. Record your results.

Reaction Template 1: Insert this page into the labtop.

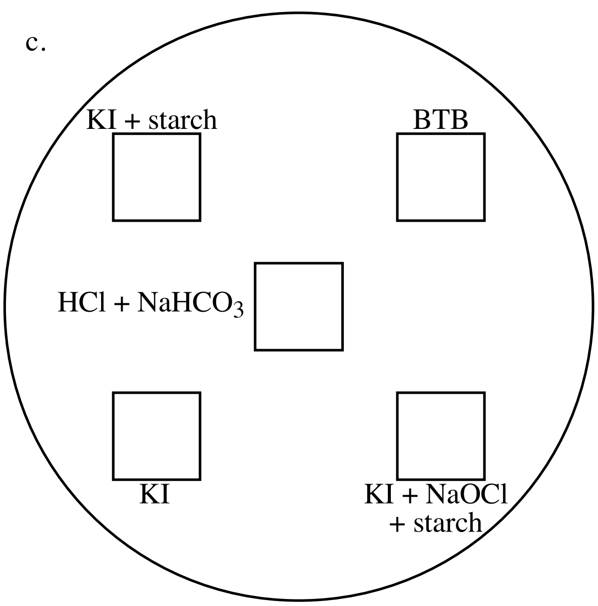
a.



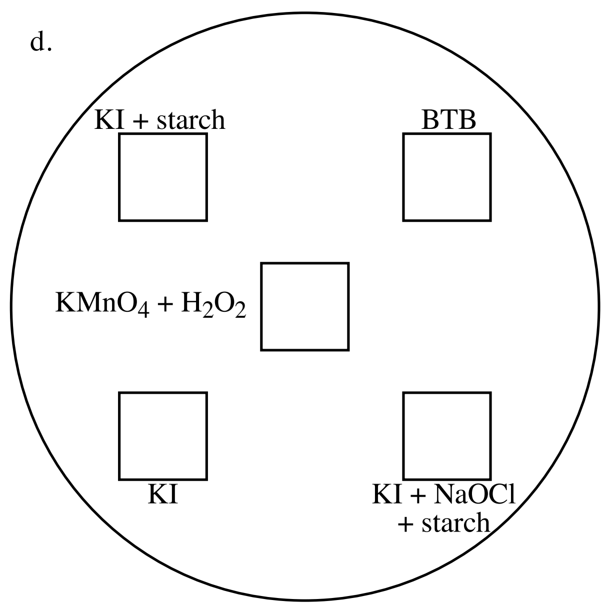
b.



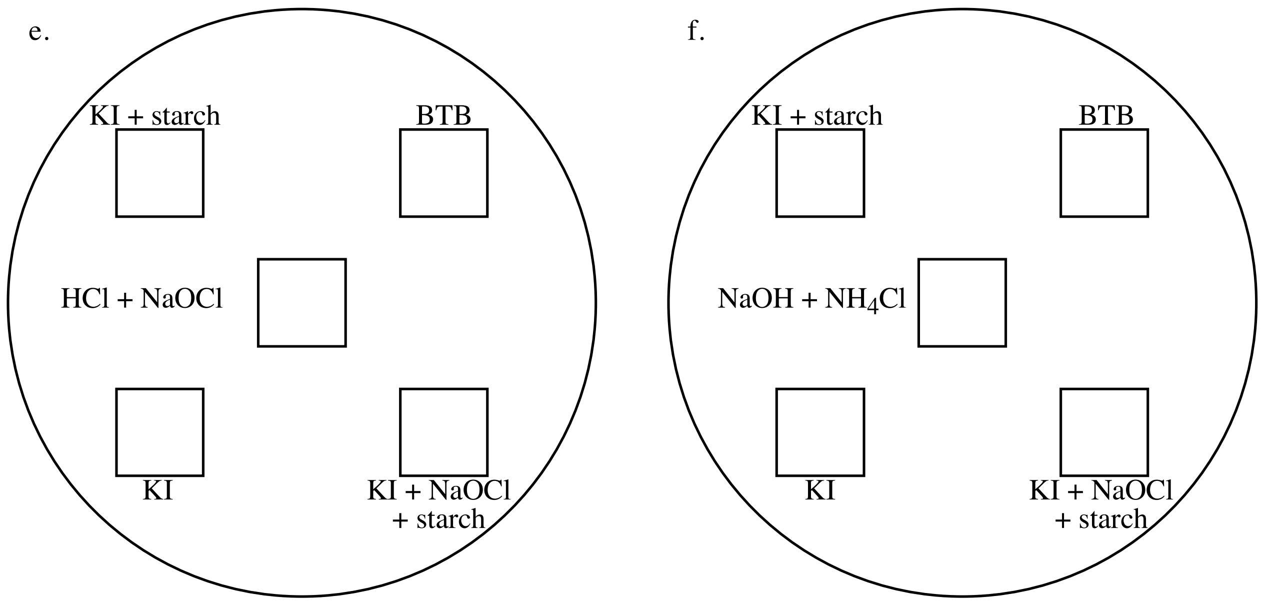
c.



d.



Reaction Template 2: Insert this page into the labtop.



Cleaning Up

Avoid contamination by cleaning up in a way that protects you and your environment. Carefully clean the plastic cups by wiping them with a dry paper towel. Clean the small-scale reaction surface by absorbing the contents onto a paper towel, rinsing the reaction surface with a damp paper towel, and drying it. Dispose of the paper towels in the wastebasket. Wash your hands thoroughly with soap and water.

Activity 16 - Synthesis and Qualitative Analysis of Gases

Name____________________________

Section____________ Date_______

Experimental Data: Observations of Gases

Complete the following table by recording any observed changes. Write "NVR" for no visible reaction.

MIXTURE	Bubbles? (describe)	BTB (green)	KI + starch (colorless)	KI + NaOCl + Starch (black)	KI (colorless)
a. HCl + $NaHSO_3$					
b. HCl + $NaNO_2$					
c. HCl + $NaHCO_3$					
d. $KMnO_4$ + H_2O_2					
e. HCl + NaOCl					
f. NaOH + NH_4Cl					

1. What is the physical evidence that mixtures (a-f) produced gases?

2. Are all the gases the same? Explain your answer using specific physical evidence from your experiment.

3. Given the word equations that describe all of the chemical reactions in this experiment, write and balance a chemical equation to describe the formation of each gas produced. **The letter of each word equation corresponds to the letter in the Experimental Data table**. Only ions will have charges in formula.

   a) Sodium hydrogen sulfite reacts with hydrochloric acid to produce sulfur dioxide gas, water, and sodium chloride.

   b) Sodium nitrite reacts with hydrochloric acid to produce nitrogen monoxide gas, water, sodium chloride, and sodium nitrate.

      – Nitrogen monoxide from the above reaction reacts with oxygen gas in the air to produce nitrogen dioxide.

   c) Sodium hydrogen carbonate reacts with hydrochloric acid to produce carbon dioxide gas, water, and sodium chloride.

   d) Five moles of hydrogen peroxide react with 2 moles of permanganate ion and 6 moles of hydrogen ion to produce oxygen gas, manganese (II) ion and water. Make sure you balance the product side of the equation without changing the number of moles of reactants.

   e) Chlorine gas, water and sodium chloride *are produced* when hydrochloric acid is added to sodium hypochlorite.

   f) Sodium hydroxide reacts with ammonium chloride to produce ammonia gas, water and sodium chloride.

4. Given the word equations below, write and balance chemical equations to describe each reaction of a gas with an indicator or *the water present in the indicator solution*. **The letter of each word equation corresponds to the letter in the Experimental Data table.** Only ions will have charges.

   a) Sulfur dioxide gas reacts with water to produce sulfurous acid ($H_2SO_3(aq)$). (Hint: This is a combination reaction.)

      – Sulfur dioxide also reacts with iodine and water to form iodide ion, sulfate ion and hydrogen ion.

b) Nitrogen dioxide gas reacts with water in the indicator to produce nitric acid ($HNO_3(aq)$) and nitrous acid ($HNO_2(aq)$).

c) Carbon dioxide combines with water to form carbonic acid ($H_2CO_3(aq)$).

e) Chlorine reacts with iodide ion to form iodine and chloride ion.

f) Ammonia gas reacts with water to produce ammonium hydroxide.

5. **Review** the indicator changes in your experimental data table and the reactions written in Problem 4. *When you answer the following questions, you are drawing conclusions from what you saw and the chemical equations in the previous questions.*

   a) What one of the three chemicals is responsible for the dark color associated with the KI/NaOCl/starch indicator mixture?

   b) What is the color of BTB in the presence of an "acid"? Explain. Look at the names of products given on the previous page for help here.

   c) What is the color of BTB in a solution produced by bubbling ammonia through water? Explain.

6. Describe the method used to detect carbon dioxide in the breath. Include your results.

7. Return to the two equations under 3b, and equation 4b. Use these three equations together to explain your observations for reaction b between hydrochloric acid and sodium nitrite.

Activity 17 – Solutions Worksheet

Name____________________________

Section____________ Date_______

Questions and Problems

For written answers, use complete sentences. For calculations, clearly ***show your work*** *and report your final answer with the correct number of significant figures.*

1. NaCl is more soluble in water than I_2. Explain.

2. How does an unsaturated solution differ from a saturated one?

3. The solubility of sucrose (common table sugar) at 70 °C is 320. g/100. g H_2O.

   a) How much sucrose can dissolve in 250.0 g of water at 70 °C?

   b) Will 620.0 g of sucrose dissolve in a teapot that contains 200.0 g of water at 70 °C? Explain.

4. If the solubility of sucrose at 0 °C is 180. g/100. g H_2O, will 300.0 g of sucrose dissolve in a pitcher of 150.0 g of iced tea at 0 °C? If not, how many grams will dissolve?

5. What is the difference between a mass/mass percent concentration and a mass/volume percent concentration? Show an example of both, using sucrose as the solute and water as the solvent, and 15.5 as the numerical value of the percentage.

6. A 15.0 mL sample of sodium chloride solution that has a mass of 15.78 g is placed in an evaporating dish and evaporated to dryness. The residue has a mass of 3.26 g. Calculate the following concentrations for the NaCl solution.

   a) mass/mass (m/m) percent

   b) mass/volume (m/V) percent

   c) molarity

7. A 3.0 % (m/V) KI solution has a volume of 25.0 mL. Calculate the concentration of this solution in units of M (moles/L).

8. How many grams of a 25% (m/m) NaCl solution contain 150.0 g of NaCl?

9. What is the molarity of a solution that contains 80.0 g of NaOH dissolved in 500.0 mL of solution?

10. How many milliliters of a 2.50 M $MgCl_2$ solution contain 17.5 g of $MgCl_2$?

11. Calculate the osmolarity (moles of particles per Liter of solution) of a 0.750 M solution of Calcium chloride ($CaCl_2$). Assume that $CaCl_2$ is a strong electrolyte (i.e. ionizes completely).

Activity 18 - Predicting the Products of Precipitation Reactions: Solubility Rules[1]

Goals

- Observe and record precipitation reactions.
- Derive general solubility rules from the experimental data.
- Describe precipitation reactions by writing net ionic equations.
- Understand the relationship between solubility and precipitation reactions.

Pre-Lab Lecture Questions. *Answer these questions on a separate sheet using complete sentences.*

1. What is a precipitate? To the best of your ability, describe what a precipitate looks like.
2. What is a precipitation reaction? What type of chemical reaction is this? What type of compound is usually involved in precipitation reactions?
3. How does the Law of Conservation of Mass relate to writing chemical equations?
4. What is the difference between chemical equations and net ionic equations? How are they similar? How are they different?
5. Name the chemical compound represented by $Pb(NO_3)_2$.

Concepts to Review

Types of Chemical Reactions
Writing Chemical Equations
Nomenclature
Definition of a Salt, Cation, Anion
Solutions
Small-scale laboratory techniques

Introduction

Water-soluble ionic compounds in solution consist of free ions surrounded by water. In the case of zinc chloride being dissolved in water to form an aqueous solution, the dissolution of the ionic compound can be depicted as follows:

$$ZnCl_{2(s)} \longrightarrow Zn^{2+}_{(aq)} + 2\ Cl^-_{(aq)}$$

These free aqueous ions can react with other ions. Mixing of dissolved ionic compounds can lead to a precipitation reaction, an example of a double displacement reaction. The **precipitate** is a **solid product**, a new ionic compound that is different from the reactants in both composition and solubility. Solubility is defined as the amount of substance (solute) that dissolves in a given amount of solvent. Solubility is a physical property that can be useful in predicting whether the mixing of aqueous ionic compounds will lead to a precipitation reaction. The mixing of a variety of combinations of ions and observing the resulting formation of precipitates leads to the formulation of general rules of solubility. Some examples of these rules include “All sodium salts are soluble in water” or “The mixing of two ionic compounds that contain a common ion will not lead to a precipitate”.

Let’s look at an example to see how these solubility rules can help us. As part of the lab, silver nitrate and sodium carbonate are mixed. A foggy white precipitate is formed. To write the chemical equation:

1. *Identify the reactants and write their correct formulas*:

[1] Adapted from: Waterman, E. L. *Chemistry: Small-Scale Chemistry Laboratory Manual;* Addison-Wesley/Prentice-Hall, Inc.: Upper Saddle River, New Jersey, 2002; pp 165-172.

$$AgNO_{3(aq)} + Na_2CO_{3(aq)}$$

silver nitrate sodium carbonate

2. *"Swap" and replace the cations of the reactant to form the products, two new ionic compounds:* the products will be **sodium nitrate** and **silver carbonate**.

$$AgNO_{3(aq)} + Na_2CO_{3(aq)} \longrightarrow \quad\quad + \quad\quad$$

silver nitrate sodium carbonate silver carbonate sodium nitrate

3. *Write the correct formulas for the products after the arrow.* (Use the names of the product to get the formula correct.)

$$AgNO_{3(aq)} + Na_2CO_{3(aq)} \longrightarrow Ag_2CO_3 + NaNO_3$$

silver nitrate sodium carbonate silver carbonate sodium nitrate

4. *Use solubility rules to determine which product is the precipitate*:

   "All sodium salts are soluble"; therefore silver carbonate must be the foggy, white precipitate.

$$AgNO_{3(aq)} + Na_2CO_{3(aq)} \longrightarrow Ag_2CO_{3(s)} + NaNO_{3(aq)}$$

5. *LAST but not least, balance the equation using whole number coefficients*:

$$2\ AgNO_{3(aq)} + Na_2CO_{3(aq)} \longrightarrow Ag_2CO_{3(s)} + 2\ NaNO_{3(aq)}$$

How do you convert a chemical equation into a correct net-ionic equation?

1. *Rewrite the correct chemical equations,* ***but*** *this time write any aqueous (not solid) ionic compounds as free cations and anions. This form is referred to as the total ionic equation.* Use the coefficients to correctly multiply the ions and keep the equation balanced. Subscripts may or may not multiply the number of ions formed. Look below, which subscripts changed coefficients when the aqueous compounds were written as free ions?

$$2\ Ag^+_{(aq)} + 2\ NO_3^-{}_{(aq)} + 2\ Na^+_{(aq)} + CO_3^{2-}{}_{(aq)} \longrightarrow Ag_2CO_{3(s)} + 2\ Na^+_{(aq)} + 2\ NO_3^-{}_{(aq)}$$

2. Once you are sure you have the correct total ionic equation, look at the equation again. It should be balanced and you will be able to see what ions actually changed. Any species that is exactly the same on both sides is considered to be a spectator ion. *To write the net ionic equation, eliminate the spectators and only write the species that changed*:

$$2\ Ag^+_{(aq)} + CO_3^{2-}{}_{(aq)} \longrightarrow Ag_2CO_{3(s)}$$

Why so many types of equations? Each equation type is useful for different reasons. The net ionic equation is preferred in double displacement reactions because it focuses solely on the product. Now that we know silver ion precipitates with carbonate ion, it will be the goal of this lab to determine whether this is normal behavior for these ions (the general rule) or unusual (an exception).

In this lab you will mix a variety of solutions. By using (soluble) sodium salts for at least one of the reactants you will be able to identify which product is the precipitate. To assist you in identifying the solid product, you will write chemical equations for all observable reactions. These chemical reactions will be translated into net ionic equations. After examining the chemical formulas of the precipitates, you will "conclude" by summarizing your results as a set of solubility rules.

6. Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

*Contact with silver nitrate ($AgNO_3$) will stain the skin.

Use microburets to dispense reagents in such a way that they do not make contact with other drops or the reaction surface.
Return any contaminated microburets to your instructor.

Materials:

Reagent Central chemicals include aqueous solutions of the compounds listed on your experimental page.

Equipment: Clean, dry transfer pipet for mixing. Lab-top reaction surface

Experimental Procedure

1. Insert your experimental page under your reaction surface. Place one drop of each solution in the squares on your experimental page. Record what happened after mixing (by blowing air past the drop using a clean, dry pipet). ("NVR" can indicate no visible reaction). Please include adjectives that describe both color and texture being as specific as possible (not all white precipitates look the same).
2. After all the reactions are completed and all observations recorded, take one last look at your surface. Have any of the squares changed over time? Record any noticeable changes (there won't be many). Clean your surface by absorbing the contents onto a paper towel. Rinse the reaction surface with a damp paper towel and dry it. Dispose of paper towels in the waste bin. Clean your area. Wash your hands thoroughly with soap and water.
3. Answer the questions, writing both chemical and net ionic equations correctly.
4. Draw general conclusions about the cations and anions in your experiment by formulating your own solubility rules.
5. Apply your rules to unknown combinations.

7. **Reaction Template:** Insert this page into the labtop. Mix one drop of each, using a long stem pipet to blow air past the droplet to complete the mixing.

Sol'ns	Na_2CO_3	NaCl	NaOH	$NaNO_3$	Na_3PO_4	Na_2SO_4
$AlCl_3$	✕	✕	✕	✕	✕	✕
NH_4Cl	✕	✕	✕	✕	✕	✕
$CaCl_2$	✕	✕	✕	✕	✕	✕
$CuSO_4$	✕	✕	✕	✕	✕	✕
$FeCl_3$	✕	✕	✕	✕	✕	✕
$Pb(NO_3)_2$	✕	✕	✕	✕	✕	✕
KI	✕	✕	✕	✕	✕	✕
$AgNO_3$	✕	✕	✕	✕	✕	✕

Activity 18 – Precipitation Reactions

Name____________________________

Section____________ **Date**______

Experimental Data:

1. Complete the following table by recording any observed precipitates (ppt.). Write “NR” for no reaction.

Sol'ns	Na_2CO_3	NaCl	NaOH	$NaNO_3$	Na_3PO_4	Na_2SO_4
$AlCl_3$						
NH_4Cl						
$CaCl_2$						
$CuSO_4$						
$FeCl_3$						
$Pb(NO_3)_2$						
KI						
$AgNO_3$						

2. Write out the chemical equation, the total ionic equation, and the net ionic equation for each reaction that formed a precipitate *on separate sheets of paper* and attach them.

3. Look at your data table and the net ionic equations in Question 2. What cations regularly formed precipitates (at least half of the combinations in a row)? Give both the correct name(s) and formula(s) for the cation(s).

4. What cations never formed precipitates (blank rows)? Give both the correct name(s) and formula(s) for the cation(s).

5. What anions regularly formed precipitates (at least half of the combinations in a column)? Give both the correct name(s) and formula(s) for the anion(s).

6. What anions never formed precipitates (blank columns)? Give both the correct name(s) and formula(s) for the anion(s).

7. Review your Questions 3 through 6 answers thus far. Complete the following set of solubility rules:

   Salts containing the following ions tend to be soluble:

   Salts containing the following ions tend to be insoluble:

8. According to your solubility rules from your data, predict whether the following combinations with produce a precipitate. Write “ppt.” in the squares where a reaction will occur.

Solutions	Potassium phosphate	Ammonium hydroxide
Magnesium chloride		
Sodium chloride		

Summary

9. Return to your solubility rules. Notice the phrase “tend to be soluble” or “tend to be insoluble”. What is the significance of the word “tend”?

10. Each type of equation, chemical, total ionic and net ionic is practical or useful in a way that the others are not. Describe what is useful about each type.

Activity 19 – Electrolytes

Goals

- Observe and distinguish the conductive nature of solutions.
- Classify substances as strong, weak or non- electrolytes using conductivity.
- Understand the role of electrolytes in electrical systems.
- Design a valid experiment and evaluate the results.

Pre-lab Questions *(answer these on a separate sheet using complete sentences)*

1. Define the term electrolyte.
2. Are all electrolytes the same in their ability to conduct electricity? Explain.
3. How can you detect electrical conductivity?
4. Does your body conduct electricity? If so, give your reasoning or evidence.
5. Go to your refrigerator, kitchen cupboard or bathroom cabinet and identify three to five electrolytes. This does not require any special equipment, only an understanding of bonding in compounds and what happens when they dissolve in water.
6. Turn to the materials section of this lab. Classify the compounds as either ionic or molecular/covalent. Provide the missing formulas.

Concepts to Review

Ionic and molecular/covalent compounds
Molarity
Nomenclature
Observation vs. Conclusion
Solutions

Introduction

The most common use of the word electrolyte is found in the description of modern sports drinks. These beverages are said to replace the vital components lost in our sweat during physical exertion. In general, **electrolytes** are anything that, when dissolved in water, produce an electrically conductive solution. It was the discovery that some solutes conduct electricity when dissolved in water that won Svante Arrhenius the Nobel Prize in chemistry in 1903.[1] The concept that matter could fall apart in aqueous solution and form ions seemed to contradict the atomic theory and was initially not accepted. Today an understanding of electrolytes is foundational to understanding biochemistry, electrochemistry and a whole host of other disciplines.

A simple experiment can be done with an open circuit and an energy source (electricity from an outlet or a battery). If the electrodes of the open circuit are placed in a solution (and they themselves are not touching), the solution itself may be conductive. A critical component to this circuit is some means to "see" the completion of the circuit. Light emitting devices (LEDs) are included to give visual evidence of electrical conductivity to the observer; when the circuit is completed a light turns on.

Most batteries contain some kind of electrolyte. The lead batteries in most automobiles contain a sulfuric acid solution. Hybrid cars use two types of batteries, both a high voltage nickel metal hydride (NiMH) battery to provide a variety of advanced capabilities, but also the traditional lead-acid battery for starting in cold weather and powering other auxiliary loads.[2]

In today's experiment, several stock solutions of the same molarity will be tested. Based on their ability to complete an open circuit and turn on an LED to some degree, you will classify a substance as an **electrolyte**

---

[1] Beaty, W. Ridiculed Discoverers, Vindicated Mavericks. http://www.amasci.com/weird/vindac.html#j1 (accessed 6/16/07)
[2] Johnson Controls Hybrid Technology. http://www.johnsoncontrols.com/publish/us/en/hybrid_technology.html (accessed 6/16/07)

or a **nonelectrolyte**. From this experiment you will also be able to distinguish within the class of electrolytes which ones are **strong** and which are **weak**.

Since you will be asked to distinguish between strong and weak electrolytes, pay very close attention to the intensity of the light emitted in the conductivity experiments. The distinction between strong electrolytes, weak electrolytes and nonelectrolytes may be made apparent in the optional exercise using Hot Wheels cars that depend on electrolytes as fuel. When using stock solutions as fuel for Hot Wheels, make sure the experiments are consistent and reproducible (*precise*) so that the comparisons will be valid (*accurate*).

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact with all chemical reagents and dispose of reactions using appropriate waste container.

Materials

<u>Compounds for Part A</u>
Calcium chloride
Sodium hydrogen sulfate
Sodium hydroxide
Sodium phosphate
Copper (II) sulfate
Zinc chloride
Sodium carbonate
Sodium hydrogen carbonate
Iron (III) chloride
Aluminum chloride

<u>Compounds for Part B</u>
Hydrogen peroxide
Hydrochloric acid ($HCl(aq)$)
Ethanol ($CH_3CH_2OH(l)$)
Methanol ($CH_3OH(l)$)
2-Propanol ($CH_3CH(OH)CH_3(l)$)
Ammonia

Equipment required: Glass slide, Hot plate, Hot Wheels racer, Conductivity apparatus

Experimental Procedure

A. Conductivity Tests

1. Place one drop of the first ten solutions (calcium chloride, zinc chloride, sodium hydrogen sulfate, sodium carbonate, sodium hydroxide, sodium hydrogen carbonate, sodium phosphate, iron (III) chloride, copper (II) sulfate, aluminum chloride) onto 1 or 2 glass slides.
2. Use the conductivity apparatus to test for conductivity. Record your observations
3. Carefully place the glass slide onto a hot plate. Begin heating at a low setting, only increasing to higher settings if the solutions do not change or evaporate in a timely fashion. Avoid inhaling any vapors that may be coming off of the slide.
4. After evaporation of the solutions, carefully take the slide off the hotplate. Handle with tweezers or other tools only, as the slide is hot and can burn the skin. Test the resulting residue for conductivity. Record you observations. Clean and dry the apparatus after each measurement.
5. When the slide is cool, clean thoroughly and proceed.

B. Conductivity Tests Continued

1. Place one drop of the remaining solutions onto 1 or 2 glass slides.
2. Use the conductivity apparatus to test for conductivity. Record your observations
3. Carefully place the glass slide onto a hot plate. Keep the heat on a low setting, being very cautious to not inhale any vapors coming off the slide. Only increase the heat if the solutions refuse to evaporate. Do not overheat!
4. After evaporation of the solutions, carefully take the slide off the hotplate. Handle with tweezers or another tool as it is hot and requires caution. Test the residue for conductivity. Record you observations. Clean and dry the apparatus after each measurement.
5. When the slide is cool, clean it thoroughly and return it to Reagent Central.

C. Hot Wheel Fuel Cell (Optional – May be done as Demonstration)

6. Get a Hot Wheel car from the cart or your instructor. Important features include a fuel cell on the top and back of the car, a start button on the top of the car and a switch on the bottom of the car. The switch has three settings, from left to right, (1) on, using fuel cell, (0) off and (2) on, using batteries (Figure 1). Familiarize yourself with the car. Remove the fuel cell. Can you make the car run without it? How? Now fill the fuel cell with DI water and replace the fuel cell. Observe the behavior of the car when you have the switch in all three positions. These cars can travel up to 50 ft. using the fuel cells so be careful where you point them. What did you observe? Add a pinch of "salt" to the DI water in the fuel cell and test the three switch positions again. What did you observe?

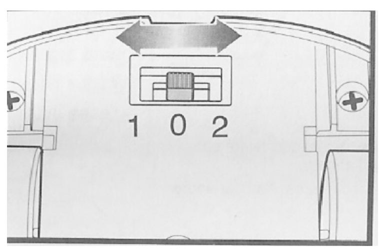


Figure 1. Switch on the underside of Hot Wheel car.

7. Once you are familiar with how the car works, test the hydrochloric acid, sodium hydroxide, ammonia and acetic acid solutions for "fuel efficiency". (They are all of the same molarity). This will require you to design an experiment that is as precise and accurate as possible. You are not confined to the geography of the lab space. Record your procedure and results.
8. Clean the "fuel cell" on your car (and the car itself if need be) and return it as soon as possible to the cart. Thoroughly clean your work area, making sure that all solutions have been removed using a damp towel and that the surfaces are also dried.

Activity 19 – Electrolytes

Name_____________________________

Section____________ **Date**_______

Exercises A and B. Conductivity Tests

1. Complete the following data table based on your experimental results.

Name and Formula	Appearance of Detector when Solution is Tested (**B**right, **D**im or **N**one)	Appearance of Detector when Dry Residue is tested (**B**right, **D**im or **N**one)

2. What are the differences between the first group of solutions tested and the second?

3. Classify all of the reagents in used as strong electrolyte, weak electrolyte or non-electrolyte by writing the **name** of the reagent in the appropriate column:

Strong Electrolyte	***Weak Electrolyte***	***Non-electrolyte***

Exercise C. Hot Wheel Fuel Cell

1. Describe what you observed regarding the Hot Wheels car. Why was this toy included in this experiment?

Activity 20 – Internet Exercise: Acids, Bases, and Salts

Goals

- Determine if a compound is acting as an acid or a base in a chemical equation.
- Identify conjugate acid/base pairs.
- Determine relative acidities of compounds.
- Write neutralization reactions.

Instructions

Access the website below and fill in your answers on the following worksheet.

A correct answer is provided for each problem identified by an asterisk next to the problem number. You should do these problems first, check your answers, and then do the rest in a similar fashion.

These exercises were written by Harry Ungar and Linda Calciano and coded by the team of graduate and undergraduate students of the C4 Project at Cabrillo College. Supported by the National Science Foundation.

http://c4.cabrillo.edu/chem30a/exercises/Exer_2/

Activity 20 - Internet Exercise: Acids, Bases and Salts

Name____________________________

Section____________ **Date**_______

Website address: http://c4.cabrillo.edu/chem30a/exercises/Exer_2/

Read and answer each question carefully and thoroughly. Please do not crowd your answers. Use scratch paper to work out each problem and write the final answer on this worksheet. Those parts of an exercise which have on-screen answers are identified by an asterisk; e.g., 1.1*. You should do these problems first, check your answers at the bottom of the web page, and then do the rest.

Exercise 1. Classify the following molecules or ions as:

A. Acid
B. Base
N. Neither

1.1* ____

1.2* ____

1.3 ____

1.4 ____

1.5* ____

1.6 ____

Exercise 2. Identify both of the conjugate acid-base pairs in each of the following sets of models. Use model numbers for identification.

Set	**Conjugate Acid-Base Pairs**	
A*	Models _________ and __________	Models __________ and __________
B	Models _________ and __________	Models __________ and __________
C	Models _________ and __________	Models __________ and __________
D	Models _________ and __________	Models __________ and __________

Exercise 3. Classify the following molecules and ions as:

A. Acid
B. Base
AB. Both
N. Neither

Set A*	3.1 _____	3.2 _____	3.3 _____	
Set B	3.4 _____	3.5 _____	3.6 _____	
Set C	3.7 _____	3.8 _____	3.9 _____	3.10 _____

Exercise 4. Match each of the following statements with one or more of the models below.

a.* Compound(s) or ion(s) whose conjugate acid is hydronium ion

b. Compound(s) or ion(s) whose conjugate acid is sulfuric acid

c. Compound(s) or ion(s) whose conjugate acid is carbonic acid

d. Compound(s) or ion(s) whose conjugate base is hydroxide ion

e.* Compound(s) or ion(s) whose conjugate base is carbonate ion

f. Compound(s) or ion(s) which are triprotic acids

g.* Compound(s) or ion(s) that would be produced by ionization of water

h. Compound(s) or ion(s) that would be produced by ionization of acetic acid

i. <u>Two</u> compounds or ions that would react to produce only water

Statement	**Model(s)**
a.*	
b.	
c.	
d.	
e.*	
f.	
g.*	
h.	
i.	

Exercise 5. Complete the following equations by drawing Lewis structures of the missing compounds which are indicated by the letters A through H.

	Missing Compounds	
5.1*	**Compound A**	**Compound B**
5.2	**Compound C**	**Compound D**
5.3*	**Compound E**	**Compound F**
5.4	**Compound G**	**Compound H**

Exercise 6. Complete the following equations by drawing Lewis structures of the missing compounds. The *product* of the first reaction, A, is the *reactant* (the starting material) of the second reaction. B is the *product* of the second reaction and the *reactant* of the third reaction.

Compound A	**Compound B**	**Compound C**

Exercise 7. Arrange the following acids in order of increasing acidity (from weakest to strongest). Use the model numbers. Use an acid-base table in your textbook for help.

Weakest Acid ___________ ___________ ___________ ___________ Strongest Acid

Exercise 8. Predict the approximate pH (neutral or higher or lower than 7) for aqueous solutions of the following salts. *Use your textbook for help*.

Circle one:

8.1* Neutral / Higher than pH 7 / Lower than pH 7

8.2 Neutral / Higher than pH 7 / Lower than pH 7

Exercise 9. Write balanced full, total ionic and net ionic equations for the complete neutralization reaction between NaOH and each of the following weak acids. (Remove all of the acidic protons.)

	***Model 9.1**
Full	
Total Ionic	
Net Ionic	

	Model 9.2
Full	
Total Ionic	
Net Ionic	

	Model 9.3
Full	
Total Ionic	
Net Ionic	

	Model 9.4
Full	
Total Ionic	
Net Ionic	

Activity 21– Proton Transfer: Acids and Bases Reacting with Water and One Another

Goals

- ❑ Understand what defines an acid and a base.
- ❑ Be able to distinguish the ionization of acids and bases from the neutralization of acids and bases.
- ❑ Characterize lab reagents and household products as acidic, basic or neutral using acid-base indicators.

Pre-Lab Lecture Questions. *Answer these questions on a separate sheet using complete sentences.*

1. Describe the bonds in a water molecule. Ionic or covalent? Nonpolar or polar?
2. What is an acid? What is a base?
3. What is a double displacement/replacement reaction?
4. What is an acid-base indicator? Give two examples.
5. What is the interpretation of an indicator color change?
6. Write the correct chemical equation describing the reaction between hydrogen chloride and water.
7. Write the correct chemical equation describing the reaction between ammonia and water.
8. Write the correct chemical equation describing the neutralization of ammonia by hydrogen chloride?
9. Understand the definitions of an Arrhenius acid and a Brønsted-Lowry acid.
10. Define monoprotic, diprotic and triprotic acids.

Concepts to Review

Nomenclature
Writing Chemical Equations
Small Scale Techniques
Arrhenius and Brønsted-Lowry definitions of acids and bases

Introduction

Early scientists characterized many substances by their taste. Substances that were sour were term **acids** (from the Latin word for sour, *acidus*). Bitterness was associated with alkaline or basic substances. Anything that was neither acidic nor basic was deemed **neutral**. These observations remain to this day but as information about acids and bases increased, the definitions changed. Svante Arrhenius discovered that acidic character was found in aqueous solutions that contained hydrogen ions attached to water molecules known as hydronium ions (H_3O^+). Acid concentration is still described in terms of hydronium ion concentration. Therefore, an **acid** can be defined as a substance that produces hydronium ions when dissolved in water.

All aqueous solutions contain hydroxide ions (OH^-). A **base** is defined as a substance that produces hydroxide ions when dissolved in water. Later, more general definitions of acids and bases were introduced because acids and bases can exist apart from aqueous media. These definitions are presented in the table below.

Definition	Alchemist	Arrhenius	Brønsted-Lowry	Lewis
Acid	Sour	A substance that produces hydronium ions when dissolved in water	Proton donor (H^+)	Electron pair acceptor
Base	Bitter Slippery	A substance that produces hydroxide ions when dissolved in water	Proton acceptor (H^+)	Electron pair donor

The inherent problem with using taste as a diagnostic tool in the lab has been recognized and is rarely used in analytical chemistry. One simple means to determine whether something is acidic or basic is to use a solution that has a different appearance in acid and base. The substances that exhibit different colors in the presence of acids on the one hand and bases on the other are called indicators. Naturally derived indicators like red cabbage juice abound in nature. Another example of a traditional natural indicator is found in litmus paper. Litmus is paper that is soaked in an extract of specific lichen. Red litmus turns blue in the presence of base and blue litmus turns red in the presence of acid. Chemists have developed synthetic indicators to suit a wide range of applications and conditions. Methyl red, a synthetic indicator, is red in acidic solutions and yellow in basic solutions.

In today's experiment you will test a variety of acids and bases using several indicators. After making observations of the colors associated with a variety of solutions in the presence of an indicator, you will represent the color changes using chemical equations. Chemical equations are symbolic representations of transformations of matter. These equations include the reactants (starting materials) written first, followed by an arrow, and finally the products (the outcome of combining the reactants):

$$\text{Reactants} \longrightarrow \text{Products}$$

To be valid, a chemical equation must have the correct formulas for all the reactants and products. In addition, the number and types of atoms on the reactant side of the equation must equal the number and types of atoms on the product side.

To help you in writing your equations, the Brønsted-Lowry definition of acids and bases may be the most useful. In this definition of acids and bases, only hydrogen ions (or protons (H^+)) are considered. An acid is defined as a proton donor; a base is a proton acceptor. The correct chemical equation describing the reaction between gaseous hydrogen chloride and water to produce hydrochloric acid is as follows:

$$HCl_{(g)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

(Note that not only do the number and type of atoms balance, but the sum of the charges on the reactant side equals the sum of charges on the product side as well.) The reaction between a molecular acid or base and water is known as an ionization reaction. In ionization reactions of acids, hydronium ion is always one of the products.

Ammonia (NH_3) is a common base that readily reacts with water to produce hydroxide ions along with another product. Can you write the equation for this reaction? Remember to write the correct formulas for the reactants first, insert an arrow and then write the correct formulas of the products. For a base reacting with water, one of the products must be hydroxide ion. Remembering that the base is a proton acceptor and the equation must be balanced, try to write the correct equation below:

ammonia + water $\longrightarrow$ hydroxide ion + ____________

Another important reaction occurs when acids and bases react with each other to form a salt and water. This is called a neutralization reaction and occurs when all of the protons on the acid are successfully neutralized by the hydroxide ions contributed by the base. An example of a neutralization reaction is the use of sodium hydroxide to neutralize battery acid (which contains sulfuric acid). Sulfuric acid is a diprotic acid (H_2SO_4) so it requires two moles of hydroxide ion to become neutralized:

$$H_2SO_{4(aq)} + 2\ NaOH_{(aq)} \longrightarrow Na_2SO_{4(aq)} + 2\ H_2O_{(l)}$$

Reaching the exact neutralization point using indicators is not possible. (Why not?) In this lab you will use indicators to determine the acid/base nature of a variety of substances. You will record your observations and then practice writing chemical equations describing ionization and neutralization. Make sure that you understand how to write these chemical equations before leaving lab today!

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses.
Avoid contact with all chemical reagents and dispose of reactions using appropriate waste container.

Materials:

Reagent Central solutions include:

Obvious Acids	Bases	Salt Solutions
Hydrochloric acid (HCl)	Sodium hydroxide (NaOH)	Sodium bicarbonate ($NaHCO_3$) Sodium carbonate (Na_2CO_3)
Nitric acid (HNO_3)	Potassium hydroxide (KOH)	Sodium acetate ($NaC_2H_3O_2$)
Sulfuric acid (H_2SO_4)		Sodium phosphate (Na_3PO_4)
Acetic acid ($HC_2H_3O_2$)	Ammonia (NH_3 also known as Ammonium hydroxide (NH_4OH)	Sodium hydrogen phosphate (Na_2HPO_4) Sodium dihydrogen phosphate (NaH_2PO_4)
Phosphoric acid (H_3PO_4)		Sodium hydrogen sulfate ($NaHSO_4$) Sodium hydrogen sulfite ($NaHSO_3$)

Indicators: Litmus paper, Bromthymol blue (BTB), cabbage extract (RCE), phenolphthalein (phen)

Equipment: Empty pipet for mixing Lab top reaction surface hot plates slides

Experimental Procedure

1. Insert your experimental page inside of your reaction surface. Place three drops of each solution on each box of your template. Use litmus to test the solution first. You will have both red and blue litmus available. The test requires a very small piece of paper touching the drop. If it does not change with the first piece, try the other color. Record any change as either "red to blue" or " blue to red". After testing using litmus, add the remaining three indicator solutions to your three drops (a unique indicator to each of the drops), making sure that the drops do not mix. Record any color changes using descriptive language.

2. Use BTB to characterize common household products as acids or bases. Make sure you describe the household product being tested.

3. Mix one drop of each solution on the glass slide, as depicted below. Carefully move the slide to a warm (**not hot**; setting should be on 1 or 2) hot plate. After the drops have evaporated, carefully remove the glass slide using a spatula or your forceps (tweezers). Record your observations.

NaOH	NaOH	KOH	NH_3
+	+	+	+
HCl	HNO_3	H_2SO_4	$HC_2H_3O_2$

4. Place one drop of ammonia on one end of a clean cotton swab. Waft the air above the swab towards your nose without placing the swab near your nose. Note any odor. Place two drops of HCl onto your reaction surface, and absorb this acid into the cotton swab containing the ammonia. Note and record its odor.

5. Answer all the questions on your worksheet.

Reaction Template: Insert this page into the labtop. Mix one drop of each, using a long stem pipet to blow air past the droplet to complete the mixing.

Indicator	Solution	Solution	Solution
Litmus BTB Phen RCE	a. $NaHCO_3$	f. $NaC_2H_3O_2$	k. H_3PO_4
Litmus BTB Phen RCE	b. $HC_2H_3O_2$	g. NaOH	l. NaH_2PO_4
Litmus BTB Phen RCE	c. HCl	h. HNO_3	m. Na_2HPO_4
Litmus BTB Phen RCE	d. Na_2CO_3	i. $NaHSO_4$	n. Na_3PO_4
Litmus BTB Phen RCE	e. H_2SO_4	j. NH_3	o. $NaHSO_3$

Activity 21– Proton Transfer Worksheet

Name_____________________________

Section____________ **Date**_______

Experimental Data:

1. Complete the following table with your observations.

Indicator	Observations	Observations	Observations
Litmus BTB Phen RCE	a. $NaHCO_3$	f. $NaC_2H_3O_2$	k. H_3PO_4
Litmus BTB Phen RCE	b. $HC_2H_3O_2$	g. NaOH	l. NaH_2PO_4
Litmus BTB Phen RCE	c. HCl	h. HNO_3	m. Na_2HPO_4
Litmus BTB Phen RCE	d. Na_2CO_3	i. $NaHSO_4$	n. Na_3PO_4
Litmus BTB Phen RCE	e. H_2SO_4	j. NH_3	o. $NaHSO_3$

2. Complete the following table:

Household Product	Color of product after contact with Litmus	BTB	Acid or Base?

3. Record your observations in below:

Reactants				
Appearance before evaporation				
Appearance after evaporation				

   a. Write neutralization equations for each of the mixtures in Part 3:

4. Describe the odor of ammonia before the addition of HCl.

   a. What did you observe after the addition of HCl? Give an explanation for your observation.

5. Answer the following questions:
    a. Describe an acid and base in terms of the color changes they produce in the indicators litmus paper and BTB.

    b. Name each acid you identified in this experiment. (Include the letter designations; e.g. **hydrochloric acid, c**)

    c. Name each base you identified in this experiment. (Include the letter, designations; e.g. **ammonia, j**)

    d. What is the common feature in the chemical formula of each acid?

    e. Define an acid and a base in terms of the color changes produced for phenolphthalein.

    f. Predict the colors of the household products in the presence of phenolphthalein.

g. You saw the acidic solutions turn yellow with BTB. What was not directly visible is the reaction between water and the acid (the proton transfer to water). Such interactions can be described by chemical equations:

c. $HCl + H_2O \longrightarrow H_3O^+ + Cl^-$ i. $Na^+ + HSO_4^- + H_2O \longrightarrow H_3O^+ + Na^+ + SO_4^{2-}$

Note that acids ionizing in water always have hydronium ion as a product. Write the ionic equations describing the reaction of water with the remaining acids in the same format given in the example above.

h. You saw basic solutions turn blue in the presence of BTB. What was not visible is the reaction between water and the base (proton transfer to the base). Such interactions are described by chemical equations:

a. $Na^+ + HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + Na^+ + OH^-$ j. $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Note that bases ionizing in water always have hydroxide ion as a product. Write the ionic equations of water and the remaining bases in the same format given in the example above.

i. Can you predict whether something is an acid or base by looking at its formula? Explain your answer.

Activity 22 - Acids and Bases Worksheet

Name____________________________

Section____________ Date______

Pre-Lab Lecture Questions. *Answer these questions on a separate sheet using complete sentences.*

1. Why do a lemon, grapefruit and vinegar taste sour?
2. What is the acid listed on the label of a bottle of vinegar?
3. What do antacids do? What are some bases listed on the labels of antacids?
4. Why are some aspirin products buffered?

Try It at Home

1. Some natural pigments act as indicators by forming different colors at different hydronium ion concentrations. Prepare an indicator by boiling some red cabbage leaves in water for 5 minutes. Cool the purple solution. Place small amounts of a household product such as vinegar, lemon juice, antacids, cleaners, shampoos and detergents in containers. Add a teaspoon of cabbage juice to each and observe the color. A pink-orange color indicates a pH range of 1-4; a pink-lavender, 5-6; purple, 7; green 8-11 and yellow, 12-13. Classify the products as acidic, neutral or basic. Try other highly colored vegetables or fruits to determine their use as indicators.
2. Place some cabbage indicator in a solution of baking soda made by adding 1 teaspoon of baking soda to a half a glass of water. Carefully add small amounts of vinegar. How does the color change? How do you know that the vinegar (an acid) neutralizes the baking soda (a base)?

Key Words

Use *complete sentences* to describe the following terms:

1. Electrolyte
2. Acid
3. Base
4. pH
5. Neutralization
6. Buffer

Electrolytes - Key Concepts

- Solutions of electrolytes are conductors of electrical current because electrolytes produce ions in aqueous solutions
- Strong electrolytes ionize completely, whereas weak electrolytes ionize only partially. Indicate incomplete ionization using double arrows " $\rightleftharpoons$ ".

Exercise A

Write an equation for the dissolving of the following salts as they combine with water to form an aqueous solution:

1. LiCl

2. $Mg(NO_3)_2$

3. Na_3PO_4

4. K_2SO_4

5. $MgCl_2$

Exercise B

Indicate whether aqueous solutions of the following solutes will contain ions, molecules, or both ions and molecules, and write an equation for their dissolution.

6. Glucose, $C_6H_{12}O_6$, a nonelectrolyte **Ions, molecules, or both? (circle one)**

Equation describing the dissolving of the solute:

7. NaOH, a strong electrolyte **Ions, molecules, or both? (circle one)**

Equation describing the dissolving of the solute:

8. K_2SO_4, a strong electrolyte **Ions, molecules, or both? (circle one)**

Equation describing the dissolving of the solute:

9. NH_3, a weak electrolyte that is also a base: **Ions, molecules, or both? (circle one)**

Equation describing the dissolving of the solute:

Acids and Bases – Key Concepts

- In water, an Arrhenius acid produces H^+, and an Arrhenius base produces OH^-.
- According to the Brønsted-Lowry theory, acids are proton (H^+) donors, and bases are proton acceptors.
- Protons form hydronium ions (H_3O^+) in water when they bond to water molecules.

Exercise C

Indicate whether the following characteristics describe and acid (A) or a base (B):

1. A B Turns blue litmus red
2. A B Contains more OH^- ions than H_3O^+ ions
3. A B Tastes bitter
4. A B Contains more H_3O^+ ions than OH^- ions
5. A B Tastes sour
6. A B Neutralizes bases
7. A B Turns red litmus blue
8. A B Neutralizes acids

Exercise D

Fill in the blank spaces with the formula or name of an acid or base

	Formula	**Name**
1.	HCl	
2.		Sodium hydroxide
3.		Sulfuric acid
4.		Nitric acid
5.	$Ca(OH)_2$	
6.	H_2CO_3	
7.	$Al(OH)_3$	
8.		Potassium hydroxide

Strengths of Acids and Bases – Key Concepts

- In strong acids, all the H^+ in the acid is donated to H_2O; in a weak acid, only a small percentage of acid molecules produce H_3O^+.
- Strong bases are hydroxides of Group 1A and 2A elements that ionize completely in water. An important weak base is ammonia, NH_3.

Exercise E

When do you use the double arrows in ionization equations? Write equations for the ionization of the following acids in water:

1. HCl, a strong acid

2. HF, a weak acid

3. HNO_3, a strong acid

Exercise F

When is water a reactant in the dissolving process? Write equations for the ionization of the following bases in water:

4. NaOH, a strong base

5. NH_3, a weak base

6. $Mg(OH)_2$, a strong base

Acid-Base Neutralization – Key Concepts

- Neutralization is a reaction of an acid and a base that produces water and a salt.
- The net ionic equation for the neutralization of any strong acid with any strong base is

$$H_3O^+ \quad + \quad OH^- \longrightarrow 2\,H_2O$$

- The net ionic equation for the neutralization of a weak acid by a strong base must include the acid written as a molecule:

$$HA \quad + \quad OH^- \longrightarrow A^- \quad + \quad H_2O$$

- In a balanced neutralization equation, the number of moles of OH^- utilized must equal the number of moles of protons available for reaction. In other words, one mole of a diprotic acid or a triprotic acid requires 2 or 3 moles of NaOH, respectively, to become neutralized. All of the protons must be converted to water.

Exercise G

Write neutralization equations for the reactions between the following acids and bases:

1. Hydrochloric acid and magnesium hydroxide

2. Sulfuric acid and sodium hydroxide

3. Nitric acid and potassium hydroxide

4. Phosphoric acid and sodium hydroxide

5. Sulfuric acid and ammonia

Ion Product of Water – Key Concepts

- In pure water, a small fraction of the water molecules transfer protons to each other, producing small, but equal amounts of H_3O^+ and OH^-. Both ions have a concentration of 1×10^{-7} M at room temperature.

- The ion product for water is denoted as K_w where $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25 °C. This equilibrium applies to *all* aqueous solutions, not only to pure water.

- In acidic solutions, $[H_3O^+]$ is greater than $[OH^-]$. In basic solutions, $[OH^-]$ is greater than $[H_3O^+]$. In neutral solutions, $[H_3O^+]$ is equal to $[OH^-]$. However K_w always holds as $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25 °C.

Example:

What is the $[H_3O^+]$ in a solution that has $[OH^-] = 1.0 \times 10^{-9}$?

Solution:

$$K_W = 1.0 \times 10^{-14} M^2$$

$$K_W = [H_3O^+][OH^-]$$

$$[H_3O^+] = \frac{K_W}{[OH^-]}$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14} M^2}{[OH^-]}$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14} M^2}{1.0 \times 10^{-9} M} = 1.0 \times 10^{-5} M$$

Exercise H

Calculate $[H_3O^+]$ when the $[OH^-]$ has the following values:

1. 1.0×10^{-10}

2. 1.0×10^{-5}

3. 1.0×10^{-7}

4. 1.2×10^{-4}

5. 3.5×10^{-8}

Exercise I

Calculate $[OH^-]$ when the $[H_3O^+]$ has the following values:

1. 1.0×10^{-3}

2. 1.0×10^{-10}

3. 1.0×10^{-6}

4. 2.8×10^{-13}

5. 8.6×10^{-7}

The pH Scale – Key Concepts

- For commonly encountered solutions, the pH scale ranges from 0 to 14. Its value is related to the hydronium ion concentration of the solution.

- A neutral solution has a pH of 7.00 at 25 °C. In acidic solutions, the pH is less than 7; in basic solutions, it is above 7.

- The mathematical definition is $pH = -\log [H_3O^+]$

- The number of decimal places in a pH value is the same as the number of sig. figs. in the concentration used to calculate that pH value.

Exercise J

Circle the most acidic pH in each group:

1. 5 2

2. 12 9 2

3. 3 7 10

4. 7.5 4.4 3.2

3. 0.2 1.5 2.3

6. 5.5 3.8 11.2 1.6

Exercise K

Calculate the pH of the following solutions at 25 °C. Indicate whether the solution is acidic, basic, or neutral.

	pH	Acidic, Basic, or Neutral
1. $[H_3O^+] = 1.0 \times 10^{-8}\ M$	______________	A B N
2. $[H_3O^+] = 0.0010\ M$	______________	A B N
3. $[OH^-] = 1.0 \times 10^{-12}\ M$	______________	A B N
4. $[OH^-] = 2.0 \times 10^{-5}\ M$	______________	A B N
5. $[OH^-] = 1.0 \times 10^{-7}\ M$	______________	A B N

Exercise L

Indicate whether the following pH values are acidic, basic, or neutral at 25 °C:

1. A B N plasma, pH = 7.40
2. A B N soft drink, pH = 2.80
3. A B N maple syrup, pH = 6.80
4. A B N beans, pH = 5.00
5. A B N tomatoes, pH = 4.20
6. A B N lemon juice, pH = 2.20
7. A B N saliva, pH = 7.00
8. A B N eggs, pH = 7.80
9. A B N lime (CaO, not citrus), pH = 12.40
10. A B N strawberries, pH = 3.00

Exercise M

Complete the following table for solutions at 25ºC:

	$[H_3O^+]$	$[OH^-]$	pH	acidic, basic, neutral
1.		1.0×10^{-12}		
2.			8.32	
3.	5.0×10^{-8}			
4.				neutral
5.			1.00	

Buffers – Key Concepts

- A buffer solution minimizes the change in pH when small amounts of acid or base are added.
- Virtually all buffers contain a weak acid and its conjugate base. The weak acid reacts with added OH^- and the conjugate base (which is also weak) reacts with added H_3O^+.
- Buffers are important in maintaining the pH of blood.

Exercise N

State whether or not mixtures 1 through 4 below represent a buffer system, and explain why or why not:

1. HCl and NaCl

2. K_2SO_4

3. H_2CO_3

4. H_2CO_3 and $NaHCO_3$

5. A buffer is prepared by adding 26.8 mL of 0.200 M HCl to 50.0mL of 0.200 M tris-(hydroxymethyl)aminomethane (Tris). The mixture is then diluted to a total volume of 200 mL with water. $TrisH^+$ has a pK_a value of 8.3 at $20^O C$. What is the pH of the above buffer solution? (You may use the Henderson-Hasselbalch equation (see below) in your calculation.) $pH = pKa + \log([Tris]/[TrisH^+])$

Activity 23 - Introduction to Radioactivity and Half-Life Experiment Kit

Introduction

Some substances contain radioactive nuclides (isotopes) which have a property called half-life. Half-life is the time it takes for half of the atoms of the radioactive nuclide to decay or change into another substance. The atoms do not decay in any set order. Some radioactive nuclides have a half-life of 5000 years. This means that after 5000 years, only half of the atoms in the sample will decay.

Geologists use half-life in radioactive dating of some rocks. They compare the amount of the radioactive nuclide in a rock to the amount that has decayed to form another substance. By knowing the half-life of the nuclide, the approximate age of the rock can be determined.

For example, uranium is radioactive. Through a series of steps it breaks down to lead (Pb) at a known rate. Its half-life is used to calculate the age of rock. The half-life of uranium (the amount of time it takes half of any amount of uranium to break down to lead) is very long, so if little lead is present relative to the amount of uranium, the rock is very young. In other words, the uranium has had no time to decay to lead.

By measuring the ratios between uranium and lead, we can estimate the ages of rocks that are millions of years old. The measurements are not perfect, but they have provided a time scale that is more accurate than any previous one. Similar calculations have been made using radioactive potassium and rubidium.

Carbon-14 is radioactive carbon formed from nitrogen in the atmosphere. It has been used to date (determine the age of) plant and animal remains. All living things used carbon in life processes, and a fraction of that carbon is carbon-14. After death, no additional carbon-14 is incorporated into the plant or animal tissue, so the existing carbon-14 gradually reverts to nitrogen. The rate of decay – its half-life – is approximately 5730 years. In 5730 years, half of the carbon-14 disappears and turns into nitrogen. In another 5730 years, half of that amount disappears, and so on until no carbon-14 remains.

Organic matter younger than 1000 years has lost too little radioactive carbon for the difference to be measured. Organic matter between 1000 to 50,000 years can be dated by the amount of carbon-14 it contains. This information is extremely useful to the geologist, anthropologist, and archeologist.

Using the simulats (simulated radioactive atoms), each having a white and black side to represent atoms of a radioactive nuclide, you will perform the following series of activities:

1. repeatedly shake and toss a given number of simulats on a flat surface;
2. remove the ones that land white side up and count the whites and blacks;
3. graph the results.

In addition, you will:

1. calculate the amount of carbon-14 found in an insect embedded in amber about 18,000 years ago;
2. calculate the amount of carbon-14 present in charcoal burned about 28,000 years ago; and
3. calculate the age of pollen from the amount of carbon-14 remaining in a sample of pollen.

Materials

1 Set plastic simulated radioactive atoms (simulates)
1 Shaker
1 Worksheet and Guide

Activity 23 - Radioactivity Simulation

Name___________________________

Section____________ **Date**______

Procedure - Exercise A

Work independently or in teams as directed by your instructor. It is necessary for each student to complete a worksheet while possibly sharing materials.

1. Assemble the shaker if necessary.
2. Place an entire set (about 200) simulats (simulated radioactive atoms) in the shaker. Place a hand over the open end and shake the contents. Count the simulats and record in with "0 toss" row.
3. Carefully toss or spill the entire contents on a flat surface in front of you.
4. Move all of the simulats that landed white side up to one side. Assume these have decayed. Count the number of white simulats and record this number in Table 1 in the "removed" column. Count the number of black simulats and record this number in the column labeled "No. of simulats left".
5. Replace only the black simulats in the shaker and repeat the procedure until you have done 8 tosses. Record the number removed and the number left each time in Table 1.
6. Plot the number of simulats left against the number of tosses (the time unit) in Figure 1 on the next page. Use a ruler to accurately plot your points.

Table 1. Data from Simulation of Radioactive Decay

No. of tosses	No. of simulats removed ("decayed")	No. of simulats left
0	0	Starting total =
1		
2		
3		
4		
5		
6		
7		
8		

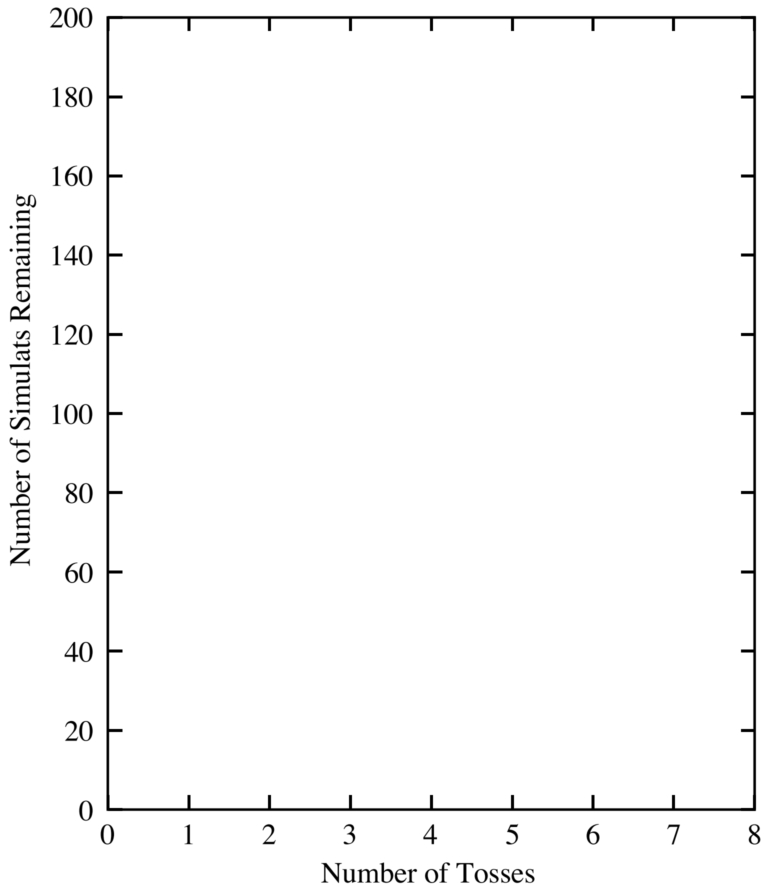


Figure 1. Graph of simulated decay results.

Procedure - Exercise B

Use the square to the right (Figure 2) to represent 100% of the carbon-14 (^{14}C) found in all living matter.

1. Divide the square in half with a vertical line to represent the amount of ^{14}C left after 5730 years (the half-life of ^{14}C).
2. Divide one of the halves again by drawing another line perpendicular to the first line. One of the resulting squares represents the amount of ^{14}C left after 11,460 years.
3. Continue to draw dividing lines to show the amount of ^{14}C left after 17,190 years, 22,920 years, and 28,650 years. Label each region accordingly.
4. On the graph in Figure 3 on the next page, plot time against the fraction of carbon-14 remaining in a once-living organic substance. This can be used to answer questions 5, 6 and 7.
5. What fraction of ^{14}C still remains in charcoal burned in a primitive man's campfire approximately 29,000 years ago? Use Fig. 2 or 3 to find the answer.

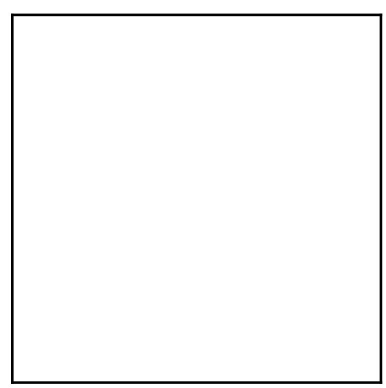


Figure 2. Depiction of sample size.

__________________remains

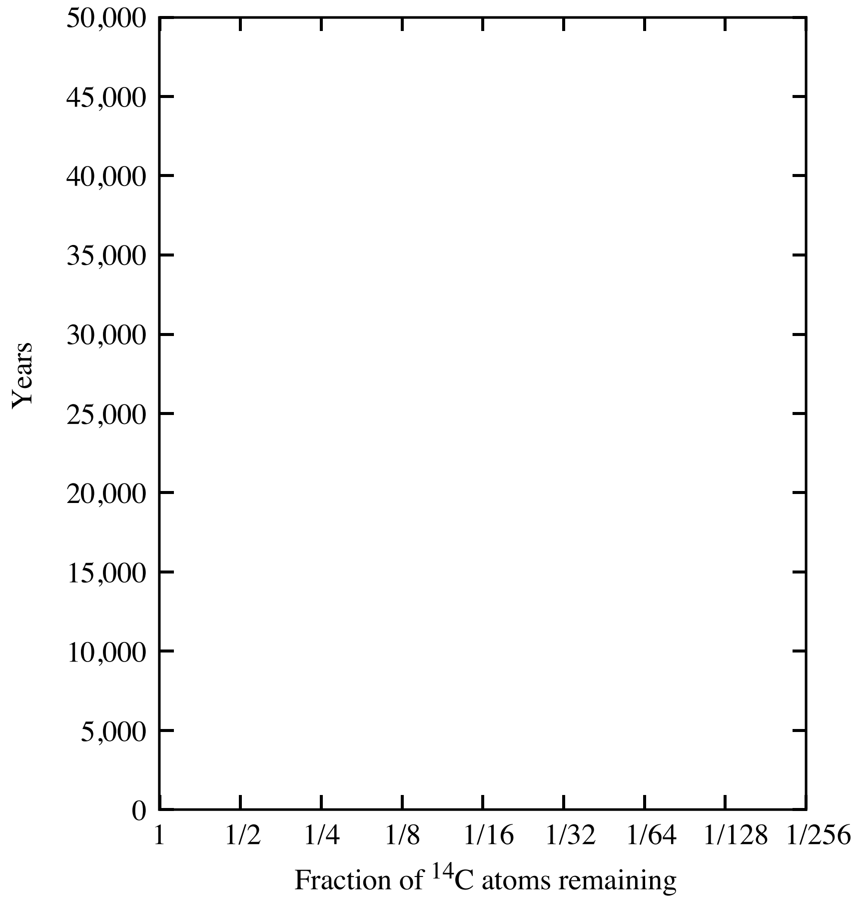


Figure 3. Relationship between half-life and sample remaining.

6. Using Figure 3, approximate the fraction of carbon-14 that still remains in an insect imbedded in amber 18,000 years ago. [Check your answer with the square in Figure 2]

_______________remains

7. Estimate the age of pollen found in peat swamps left by a glacier from your graph (Figure 3). Assume only 1/16 of the original carbon-14 was found in the pollen. [Check your answer with the square in Figure 2]

_______________years

Questions and Problems

1. How many simulats (atoms) changed or decayed by the end of Part 1?

2. Compared to the original number of simulats with which you started, approximately what fraction was left after each shake?

3. What does this indicate about the half-life? In other words, do atoms decay at about the same rate?

4. If each shake represented 500 years of time, what would be the half-life of the simulated radioactive nuclide?

5. What are some of the inaccuracies of this experiment in demonstrating half-life?

6. Will all of the carbon-14 in nature eventually disappear? Explain your answer

7. Precambrian time is older than 600,000,000 years. Can carbon-14 be used in dating organic substances from the era? Explain your answer.

8. Can carbon-14 be used for dating lava flows? Explain your answer.

Activity 24: Radioactivity - Worksheet

Name____________________________

Section____________ **Date**_______

1. Write the nuclear equations for the following radioactive processes:
    a. Silver-113 (beta emission)

    b. Iodine-122 (positron emission)

    c. Radium-226 undergoes alpha decay to produce radon-222. (Gamma rays are also a product.)

    d. Bismuth-209 (alpha decay)

    e. The bombardment of californium-249 with oxygen-16 atoms produces seaborgium-263 (Sg-263) and neutrons.

2. Rubidium-79 decays by positron emission and forms krypton-79, which is a gas. A weighed 100.00 mg sample of solid rubidium-79 was allowed to decay for 42 minutes, then weighed again. Its mass was 25.00 mg. What is the half-life of rubidium-79?

3. Iodine-131 is a radionuclide that is frequently used in nuclear medicine. The half-life is 8.0 days. How much of a 0.16-g sample of iodine-131 will remain undecayed in the body after a period of 32 days?

4. Iodine-135 is a nuclide found in radioactive fallout from nuclear weapon explosions. Its half-life is 6.70 hrs. How long will it take for 75% (3/4) of the iodine-135 atoms to undergo decay? Hint: Figure out the amount remaining first.

5. Using complete sentences, describe the difference between fusion and fission.

6. Identify two sources of naturally occurring radioactivity and two examples of synthetic radioactivity.

7. How do radionuclides used for diagnostic purposes differ form radionuclides used for therapeutic purposes?

8. List alpha, beta and gamma radiation in order of increasing penetrating power.

9. What is the difference between a PET and a CAT scan?

10. List and define three common units of radiation. If possible, include why that particular unit is useful.

Activity 25 - Endothermic and Exothermic Reactions

Introduction

Many chemical reactions give off energy. Chemical reactions that release energy are called *exothermic* reactions. Some chemical reactions absorb energy and are called *endothermic* reactions. You will study one exothermic and one endothermic reaction in this experiment.

In Part I, you will study the reaction between citric acid solution and baking soda. An equation for the reaction is:

$$H_3C_6H_5O_{7(aq)} + 3\ NaHCO_{3(s)} \longrightarrow 3\ CO_{2(g)} + 3\ H_2O_{(l)} + Na_3C_6H_5O_{7(aq)}$$

In Part II, you will study the reaction between magnesium metal and hydrochloric acid. An equation for this reaction is:

$$Mg_{(s)} + 2\ HCl_{(aq)} \longrightarrow H_{2(g)} + MgCl_{2(aq)}$$

Another objective of this experiment is for you to become familiar with Logger *Pro*. In this experiment, you will use Logger *Pro* to collect and display data as a graph or table, analyze your experimental data values, and print a graph or data table.

Safety

Wear safety glasses at all times!
Act in accordance with the laboratory safety rules of Cabrillo College.
Avoid contact with all chemical reagents and dispose of reactions using appropriate waste containers.

Materials

Power Macintosh or Windows PC	Styrofoam cup
Vernier computer interface	250-mL beaker
Logger *Pro*	citric acid, $H_3C_6H_5O_7$, solution
Temperature Probe	baking soda, $NaHCO_3$
50-mL graduated cylinder	hydrochloric acid, HCl, solution
balance	magnesium, Mg

Experimental Procedure

A. Citric Acid plus Baking Soda

1. Place a Styrofoam cup into a 250-mL beaker as shown in Figure 1. Measure out 30 mL of citric acid solution into the Styrofoam cup. Place a Temperature Probe into the citric acid solution.

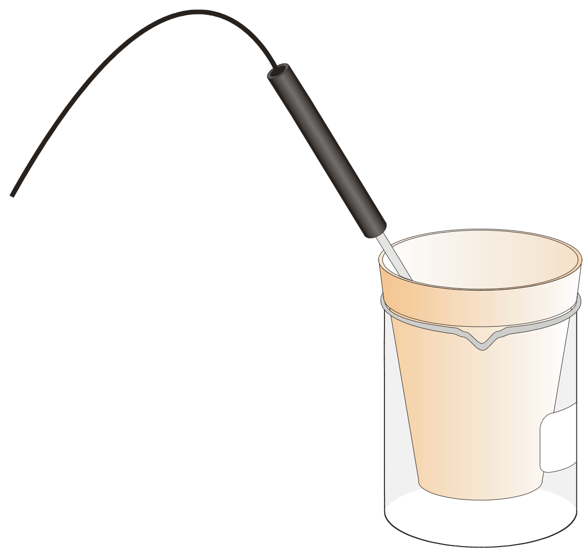


Figure 1. Experimental setup.

2. Prepare the computer for data collection by opening the Experiment 1 folder from the *Chemistry with Computers* folder of Logger *Pro*. Then open the experiment file that matches the probe you are using. The vertical axis has temperature scaled from -10 to 40°C. The horizontal axis has time scaled from 0 to 300 seconds.
3. Weigh out 10.0 g of solid baking soda on a piece of weighing paper.
4. The Temperature Probe must be in the citric acid solution for at least 30 seconds before this step. Begin data collection by clicking Collect. After about 20 seconds have elapsed, add the baking soda to the citric acid solution. Gently stir the solution with the Temperature Probe to ensure good mixing. Collect data until a minimum temperature has been reached and temperature readings begin to increase. You can click on Stop to end data collection or let the computer automatically end it after 300 seconds.
5. Dispose of the reaction products as directed by your teacher.
6. To analyze and print your data:
   a. Click the Statistics button, . In the statistics box that appears on the graph, several statistical values are displayed for Temp 1, including minimum and maximum. In your data table, record the maximum as the initial temperature and the minimum as the final temperature. Close the statistics box by clicking the upper-right corner of the box.
   b. To confirm the minimum and maximum temperatures, use the scroll bars in the Table window to scroll through the table to examine the data. Compare the minimum and maximum data points to those you recorded in the previous step.
   c. Print a copy of the Table window. Enter your name(s) and the number of copies.
   d. You will often want to change the scale of either axis of the graph. There are several ways to do this. To scale the temperature axis from 0 to 25°C instead of the present scaling, click the mouse on the "40" tickmark at the top of the axis. In place of the "40", type in "25" and press the Enter key. Notice that the entire axis readjusts to the change you made. Use the same

method to change the “-10” tickmark to “0”. Note: A second option is to click the Autoscale button, . The computer will automatically rescale the axes for you.

e. You can also expand any portion of the graph by zooming in on it. Select the area you want to zoom in on. Do this by moving the mouse pointer to the beginning of this section of data—press the mouse button and hold it down as you drag across the curve, leaving a rectangle. Then click the Zoom In button, . The computer will now create a new, full-size graph that includes just the region inside the rectangle. You can reverse this action by clicking the Undo Zoom button, .

f. When you again collect data in Part II of this experiment, the data will be collected as Latest run, the *most recent* set of data you have collected. The original Latest run will be lost if it is not saved or stored. Choose Store Latest Run from the Data menu to *store* Latest as Run 1, then save or print it later. Note that the line for Run 1 is thinner than it was for Latest. To hide the curve of your first data run, click the Temperature vertical-axis label of the graph, and uncheck Run 1. Click OK.

B. Hydrochloric Acid Plus Magnesium

7. Manually rescale the vertical axis to the original temperature scale of -10 to 40°C. To do so, click the mouse on the bottom tickmark and type in “-10”. Then click on the top tickmark and type in “40”.
8. Measure out 30 mL of HCl solution into the Styrofoam cup. Place the Temperature Probe into the HCl solution. Note: The Temperature Probe must be in the HCl solution for at least 45 seconds before doing Step 11. **CAUTION:** *Hydrochloric acid is caustic. Avoid spilling it on your skin or clothing. Wear chemical splash goggles at all times. Notify your teacher in the event of an accident.*
9. Obtain a piece of magnesium metal from the teacher.
10. Begin data collection by clicking Collect. After about 20 seconds have elapsed, add the Mg to the HCl solution. Gently stir the solution with the Temperature Probe to ensure good mixing. **CAUTION:** *Do not breathe the vapors!* Collect data until a maximum temperature has been reached and the temperature readings begin to decrease.
11. Dispose of the reaction products as directed by your teacher.
12. To analyze your Part II data:

a. Change the appearance of the graph by double-clicking anywhere on the graph to bring up the Graph Options dialog. Check the box in front of Point Protector Every 1 Point—a point protector will now outline each data point on the graph. Click OK.

b. Instead of scrolling through the Table window in this trial, click the Examine button, . The cursor will become a vertical line. As you move the mouse pointer across the screen, the temperature and time values corresponding to its position will be displayed in the box at the upper-left corner of the graph. Scroll across the initial 3-4 points to determine the initial temperature. Record the initial temperature in the data table. Move the mouse pointer across the peak of the temperature curve to determine the maximum temperature, and record it as the final temperature in your data table. To remove the examine box, click the upper-right corner of the box.

c. It is also possible to calculate statistics just for a portion of your collected data. To do so, you must first *select* the data you are interested in. For example, you might want to find the average (or mean) of the first few data points to use as an initial temperature, instead of using the minimum value. Select the flat portion of the curve—move the mouse pointer to time 0 and drag across the flat part of the curve. Now click the Statistics button, , and note the mean temperature value in the statistics box on the graph. This value is the mean of only the selected data points. When you are done, click on the upper-right corner of the statistics box to remove it.

13. To print a graph of temperature *vs*. time showing both data runs:

    a. Click the Temperature vertical-axis label of the graph. To display both temperature runs, check the Run 1 and Latest boxes. Click OK.

    b. Label both curves by choosing Make Annotation from the Analyze menu, and typing "Endothermic" (or "Exothermic") in the edit box. Then drag each box to a position near its respective curve.

    c. Print a copy of the Graph window. Enter your name(s) and the number of copies of the graph you want.

14. Save the temperature and time data from both data runs. Choose Save As from the File menu and give the file a distinct name. As directed by your teacher, choose a location for the file, and click OK.

Activity 25 - Endothermic and Exothermic Reactions Worksheet

Name____________________________

Section____________ **Date**______

Data Table

	Part I	Part II
Final temperature, t_2	°C	°C
Initial temperature, t_1	°C	°C
Temperature change, Δt	°C	°C

Observations

Processing the Data

1. Calculate the temperature change, Δt, for each reaction by subtracting the initial temperature, t_1, from the final temperature, t_2 ($\Delta t = t_2 - t_1$).

2. Tell which reaction is exothermic. Explain.

3. Which reaction had a negative Δt value? Is the reaction endothermic or exothermic? Explain.

4. For each reaction, describe three ways you could tell a chemical reaction was taking place.

5. Which reaction took place at a greater rate? Explain your answer.

Activity 26 - Electronic Configuration Worksheet

Name___________________________

Section___________ Date______

Scientists have learned that electrons in atoms occupy specific energy levels. These levels form an orderly pattern, and we can use these to figure out where the electrons are (i.e., which energy levels they occupy). This allows us to predict a great deal about how these electrons will behave in chemical reactions. This is a cornerstone of chemistry.

There are principal energy levels, called shells, numbered consecutively 1, 2, 3, … . As the shell number increases, the size and energy of the shell increases.

The first shell consists of a single sublevel; the second, two; the third, three; and so on. The sublevels are designated by letters: s, p, d, f, g, h, … , and the energy of each sublevel within an energy level increases in this order (for multi-electron atoms). Also note that as the energy increases, the levels generally become more and more closely spaced (see figure 1 on page 251).

Exercise:

1. For the principal shell 4, complete the list of sublevels: 4_s_, 4___, 4___, 4___
2. For the principal shell 6, complete the list of sublevels: ____, ____, ____, ____, ____, ____

Because the spacing between energy levels becomes smaller and smaller as energy increases, the sublevels of different energy levels overlap each other. For example, the energy of the 4s sublevel is lower than the 3d.

The next step is to figure out how many electrons can fit into each sublevel. Sublevels are divided into orbitals, which are mathematically defined regions in space in an atom, molecule, or ion, within which and electron can be found about 90% of the time. Each subshell contains a different number of orbitals: s has one; p, 3; d, 5; f, 7, … , but electrons in orbitals in a given sublevel have the same energy.

Exercise:

3. Fill in the answers (number of orbitals and the lines that represent orbitals), for the f, g, and h sublevels:

subshell	Number of orbital(s)	Diagram
s	_1_	___
p	_3_	___ ___ ___
d	___	___ ___ ___ ___ ___
f	___	
g	___	
h	___	

Experiments show that each orbital may hold a maximum of 2 electrons. When two electrons are in a single orbital, their spins are aligned in opposite directions _↑↓_. Thus, an orbital may have no electrons (empty) ____; one electron _↑_; or two electrons _↑↓_.

Electronic Energy Levels

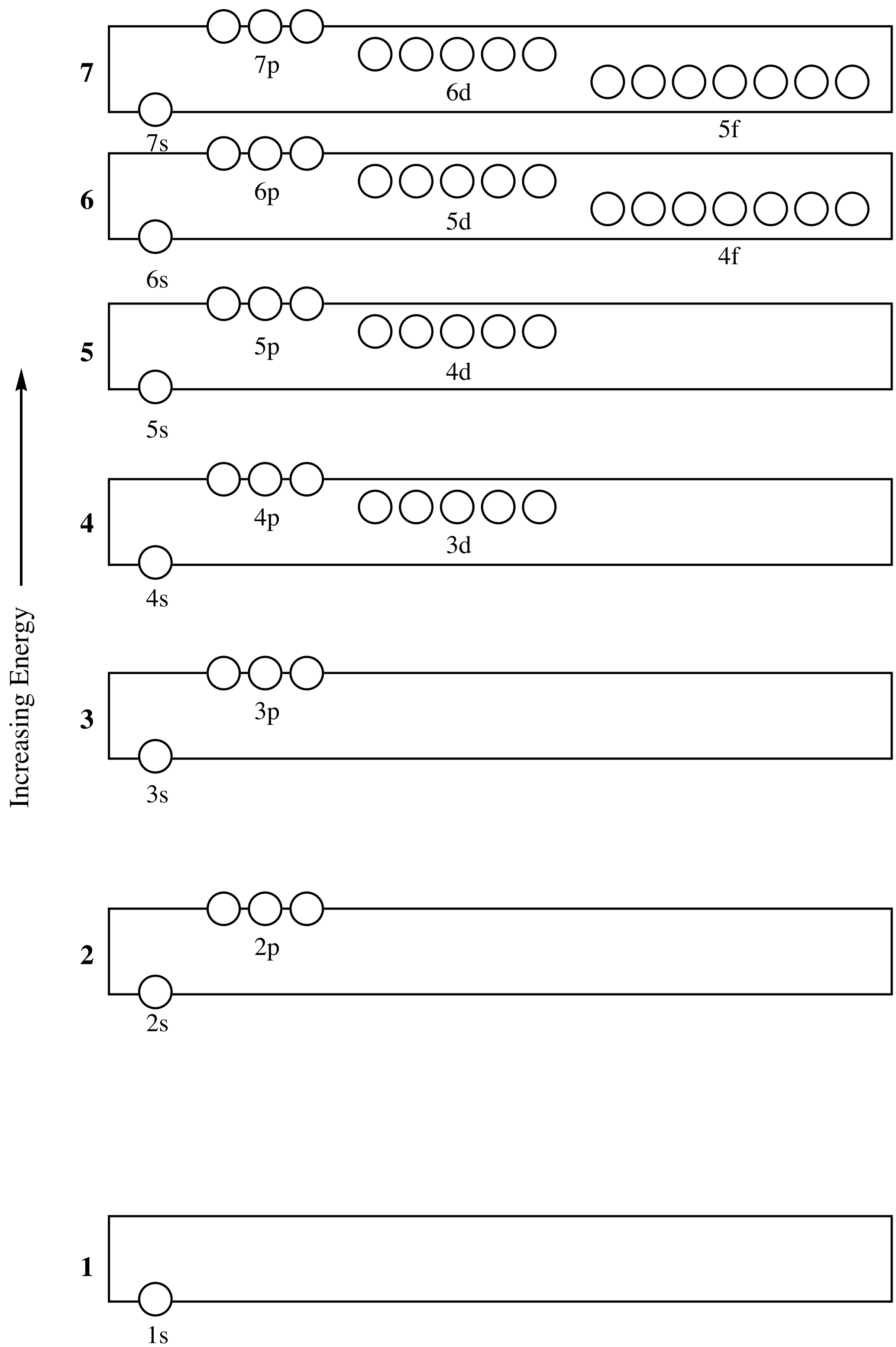


Figure 1. Schematic of the principal energy levels and sublevels through principal level 7.

Exercise:

4. Fill in the appropriate orbitals within these sublevels with the maximum number of electrons:

E.g.,	**2p** ↑↓ ↑↓ ↑↓	__6__ **electrons**
	2s	____ electrons
	4p	____ electrons
	3d	____ electrons
	5f	____ electrons
	6s	____ electrons
	5g	____ electrons
	6f	____ electrons

5. What is the maximum number of electrons in the principal energy level 1? ____________
6. What is the maximum number of electrons in the principal energy level 2? ____________
7. What is the maximum number of electrons in the principal energy level 3? ____________

We have basic rules. Now we will use this information to find the arrangement of electrons in different elements. These arrangements are called ***electron configurations*** when listed and ***orbital diagrams*** when drawn out.

Rules for drawing orbital diagrams:

a. The energy sequence of sublevels is the same for all elements.

b. Electrons go into the lowest energy orbital that has room available (Aufbau principle).

c. The maximum number of electrons in an orbital is 2.

d. Electrons fill orbitals within a sublevel with unpaired spins before spins are paired (Hund's rule).

Exercise:

8. Fill in the arrows representing electrons, as shown for H and He.

	H	**He**	Li	Be	B	C
2p	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___
2s	___	___	___	___	___	___
1s	↑	↑↓	___	___	___	___

Exercise:

Problem 8 continued.

	N	O	F	Ne
2p	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___
2s	___	___	___	___
1s	___	___	___	___

	Na	Mg	Al	Si	P	S
3p	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___
3s	___	___	___	___	___	___
2p	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___
2s	___	___	___	___	___	___
1s	___	___	___	___	___	___

	Cl	Ar
3p	___ ___ ___	___ ___ ___
3s	___	___
2p	___ ___ ___	___ ___ ___
2s	___	___
1s	___	___

The more compact list form (**electron configuration**) is:

H	He	Li	Be	B
$\mathbf{1s^1}$	$\mathbf{1s^2}$	$\mathbf{1s^2 2s^1}$	$\mathbf{1s^2 2s^2}$	$\mathbf{1s^2 2s^2 2p^1}$

Exercise:

9. Write the electron configuration for elements 6 – 18. Note: The last page of this handout includes step-by-step instructions for writing electron configurations. Note: Beginning with potassium, the fourth period elements fill in the 4s, 3d, and 4p sublevels **IN THAT ORDER** (see figure 1 on page 251).

10. Fill in the electron "arrows" for the following orbital diagrams. Note that the first 18 electrons through 3p have been omitted for clarity. ***Cr and Cu are exceptions, with configurations of $4s^1 3d^5$ and $4s^1 3d^{10}$, respectively.** This is because half-filled and/or fully-filled sublevels are more stable than subshells that are neither. Both silver and gold follow the same exceptional pattern as copper.

	K	Ca	Sc	Ti
3d	___ ___ ___ ___ ___	___ ___ ___ ___ ___	___ ___ ___ ___ ___	___ ___ ___ ___ ___
4s	___	___	___	___

	V	Cr*	Mn	Fe
3d	___ ___ ___ ___ ___	___ ___ ___ ___ ___	___ ___ ___ ___ ___	___ ___ ___ ___ ___
4s	___	___	___	___

	Co	Ni	Cu*	Zn
3d	___ ___ ___ ___ ___	___ ___ ___ ___ ___	___ ___ ___ ___ ___	___ ___ ___ ___ ___
4s	___	___	___	___

	Ga	Ge	As	Se
4p	___ ___ ___	___ ___ ___	___ ___ ___	___ ___ ___
3d	___ ___ ___ ___ ___	___ ___ ___ ___ ___	___ ___ ___ ___ ___	___ ___ ___ ___ ___
4s	___	___	___	___

	Br	Kr
4p	___ ___ ___	___ ___ ___
3d	___ ___ ___ ___ ___	___ ___ ___ ___ ___
4s	___	___

Exercise:

11. Which columns are representative elements? ______________________

12. Which subshells are being filled as you go across a row of representative elements? _____________

13. Which subshell is being filled as you go across a row of transition elements? _____________

14. Use the noble gas notation for the following electron configurations:

E.g.	**Si**	**[Ne]$3s^2 3p^2$**	Cd	[ ]
	Se	[ ]	Br	[ ]
	Sr	[ ]	I	[ ]
	Ba	[ ]	Sb	[ ]
	*Ag	[ ]		

Electronic Structure Representations

Write an electron configuration using the counting method or block from the periodic table:

1. Separate the periodic table into s, p, d and f blocks. Label where each block begins, s-block starts with 1s, p-block with 2p, d-block with 3d, etc. The s-block elements have their last electron in the s sublevel, the p-block elements have their last electron in the p sublevel and thus the label. (See the Figure 2.12 in the Timberlake text.)
2. Begin filling your electron sublevels starting in Period 1, at the element He you have filled a 1s sublevel with 2 electrons, so write $1s^2$ to begin your electron configuration (unless you are doing hydrogen).
3. Continue adding a filled sublevel to your electron configuration when you reach the end of a block ($2s^2$ $2p^6$ etc.).
4. Stop when you reach your desired element, recording how many places you moved in that block before stopping. This is the same number of electrons that fill that last sublevel (which is the same as the label of your block).

Writing quick, condensed electron configurations using the periodic table:

1. Find your element on the periodic table.
2. Find the noble gas that came just before your element.
3. Write the noble gas symbol in brackets to represent that number of electrons. [He] stands for $1s^2$ counting as 2 electrons , [Ne] stands for $1s^22s^22p^6$ accounting for 10 e-s, etc.
4. Write the remaining sublevel labels (2s, 3d, whatever) followed by the number of electrons in each orbital, written as a superscript, until you have the total number of electrons in your element. The sublevels are determined by their position in the periodic table. (Use the block version.)
5. Check your answer by adding the number of electrons represented by the noble gas to all of the superscripts.

How to draw an orbital diagram describing the electronic structure of an atom:

1. Determine the number of electrons by finding the atomic number on the periodic table.
2. Build an orbital diagram with the lowest energy orbital at the bottom of the diagram. Orbitals are represented by lines (boxes or circles). You must know the filling order and this can be obtained from using the blocks on the periodic table.
3. Fill the orbitals with the number of electrons.

   Remember:

   - Orbitals hold a maximum of two electrons.
   - These two electrons must be opposite in orientation.
   - Half-fill a sublevel (p,d,f) before pairing electrons.

Activity 27: Mass Titrations - Measuring Molar Concentrations[1]

Goals

- **Measure** the molar concentration of acids using mass titrations.
- **Compare** the accuracy of mass titrations with that of volumetric titrations.

Pre-Lab Lecture Questions. *Answer these questions on a separate sheet using complete sentences.*

1. What is the difference between mass and weight?
2. What is a titration?
3. What is a *dilute aqueous solution*?
4. Why do chemists do more than one trial when doing quantitative analysis?
5. What type of chemical reaction occurs when sodium hydroxide is added to hydrochloric acid?
6. What does *%(w/w)* mean?

Concepts to Review

Definitions of Mass
Chemical Reactions
Solutions

Introduction

A *mass titration* (also called a *weight titration)* is a method of finding molar concentration by weighing solutions rather than measuring their volumes. A mass titration is often more accurate and precise than a volumetric titration because a balance is usually a more accurate and precise instrument than a pipet or buret. The result of a mass titration depends only on weighings determined directly from the balance and not on volumes determined from pipets and burets.

You can use the balance to determine the volumes of solutions by weighing. You know that the density of water is 1.00 g/mL. The densities of dilute aqueous solutions can be safely assumed to be equal to the density of water. This means that the mass of a solution in grams is numerically equal to its volume in mL.

In this lab you will carry out several mass titrations. You will weigh a certain amount of acid solution and add an indicator. Then you will determine the mass of NaOH solution needed to reach the end point. To do this you will start by weighing a transfer pipet full of NaOH solution. There is no need to weigh the empty pipet, as you will only need to know the *difference* in mass before and after the titration. Next, you will carefully add NaOH solution to the acid solution until the indicator just changes color. Finally, you will determine the mass of the NaOH solution needed for the titration by weighing the NaOH pipet again, and taking the difference in the initial and final weighings.

Notice that you do not have to count drops, nor do you have to hold the pipet at any particular angle, nor do you have to be careful to expel any air bubbles. In short, you will work with none of the disadvantages of a drop-counting, "digital" titration.

[1] Adapted from: Waterman, E. L. *Chemistry: Small-Scale Chemistry Laboratory Manual;* Addison-Wesley/Prentice-Hall, Inc.: Upper Saddle River, New Jersey, 2002; pp 87-92.

Safety

Wear safety glasses. Both sodium hydroxide and the acid solutions are *very* dangerous if you get a drop into your eye.
Act in accordance with the laboratory safety rules of Cabrillo College.
Avoid contact with all chemical reagents and dispose of reactions using appropriate waste container.

Materials:

Small scale pipets of the following:

0.5 *M* NaOH	unknown HNO_3	unknown HCl
phenolphthalein	vinegar	
unknown H_2SO_4	lemon juice	

Equipment: Balance Plastic cups

Experimental Procedure

1. Weigh, to the nearest 0.01 g, approximately 0.5 g of acid solution as follows:
    a. Place an empty plastic cup on the balance pan. You can rinse it with deionized water from a wash bottle. Make sure the cup is dry on the outside, but it need not be dry on the inside. Then "tare" the balance so that it reads 0.00 g.
    b. Add acid solution from the transfer pipet until the balance gives a reading close to 0.5 g. Then record the mass of the acid solution to the nearest 0.01 g (i.e., record every digit you see on the readout).
2. Add 3 drops of phenolphthalein.
3. Place *two* full NaOH pipets in a second clean, dry plastic cup, and record their initial mass. (Use two pipets to make sure you have enough NaOH.)
4. Titrate the acid solution in the first plastic cup with NaOH from one or both of the pipets until *1 drop* turns the solution to a stable pink color. "Stable" means the pink color doesn't fade when you stir the titration solution, and it stays pink for at least 30 seconds.
5. Determine the mass of the NaOH solution used for the titration:
    a. Place *both* NaOH pipets back into the same clean, dry plastic cup used before.
    b. Weigh the cup and the pipets, and record their final mass.
    c. Subtract the final mass from the initial mass. This gives the net mass of NaOH solution delivered into the titration cup.
    d. Record the final mass as the initial mass for the next titration, if necessary.
6. Repeat steps 1-5 until you obtain consistent results. (The only way to tell if your result is reliable is to reproduce it.) Calculate the difference between your highest and lowest results for a given acid as a percentage of their average. This value should be less than 1%.
7. Carry out mass titrations to determine the molar concentrations of HCl, HNO_3, and H_2SO_4. Titrate each acid at least twice. Use the third column in the data table if necessary.
8. Next titrate the acetic acid in vinegar, and the citric acid in lemon juice. Again, titrate each solution at least twice.

Activity 27: Mass Titration

Name______________________________

Section_____________ **Date_______**

Experimental Data:

Record the precise molarity of the NaOH solution: __________*M*.

Molarity of Laboratory Acids

	HCl			HNO_3			H_2SO_4		
Mass of acid (g)									
Initial mass of NaOH (g)									
Final Mass of NaOH (g)									
Mass of NaOH used (g)									
Molarity of acid (*M*)									
Average acid molarity	*M*			*M*			*M*		

Molarity of Acids in Household Products

	$HC_2H_3O_2$ (acetic acid in vinegar)			$H_3C_6H_5O_7$ (citric acid in lemon juice)		
Mass of acid (g)						
Mass of NaOH (initial) (g)						
Mass of NaOH (final) (g)						
Mass of NaOH used (g)						
Molarity of acid (*M*)						
Average acid molarity (*M*)	*M*			*M*		

To calculate the molarity of acid in each solution:

$$[\text{acid}] = \frac{(\text{precise molarity of NaOH}) \times (\text{mass of NaOH solution in grams})}{(\text{mass of acid solution in grams}) \times n}$$

The quantity n in the denominator of the above fraction is the number of moles of *titratable* H^+ in each mole of the acid. For HCl, HNO_3 and $HC_2H_3O_2$, n = 1. For H_2SO_4, n = 2, and for $H_3C_6H_5O_7$, n = 3.

Cleaning Up

Clean the plastic cup used for the titration by emptying it into the liquid waste. Rinse with deionized water, and dry if necessary with a paper towel before handing it in. Paper towels go into the solid waste. Wash your hands with soap and water.

Questions for Analysis

1. Show that for a dilute aqueous solution, the number of milligrams (mg) is equal to the number of microliters (μL). What assumption about the solution must you make for this to be true?

2. Why did you *not* have to weigh the empty NaOH pipets before (or after) the titration?

3. Why did you not need to expel air bubbles from the NaOH pipet, or hold it at a vertical angle while titrating?

4. Calculate the %(w/w) of acetic acid in the vinegar.

5. Calculate the %(w/w) of citric acid in the lemon juice.

6. Assume that the true %(w/w) of acetic acid in vinegar is 5.00%. Calculate the percent error in the average value you obtained.

Lewis Structures and Shape Assignment

To draw a Lewis structure for a molecule or an ion:

1. Write the chemical symbols for the atoms in their bonding sequence (a reasonable structure or as given in the problem). You must choose the central atom first. The atom with the greatest bonding potential (most single dots in atomic Lewis structure) or lowest EN is the central atom. Here are more guidelines to help:
   - H and F are never in the center; O is rarely in the center but there are important exceptions.
   - Oxygen atoms are rarely bonded to other oxygen atoms (except in compounds like O_2, O_3, etc.).
   - The molecular formula sometimes reflects the molecular structure.
   - Carbon atoms are commonly bonded to other carbon atoms (where applicable).
   - C, N, O, & F will have 8 electrons around them; H has 2; B can have fewer than 8 but never more than 8.
2. Draw single bonds to connect the atoms.
3. Complete octets around all atoms except H, which gets a "duet"; B also tends to have 6 electrons around it.
4. Count available valence electrons (all outer-shell electrons); each "-" ionic charge = one more electron; each "+" ionic charge = one less electron
5. Commonly, you see:
   - H, F, Cl, Br, I usually form 1 bond/3 lone pairs
   - O, S, Se usually form 2 bonds/2 lone pairs
   - N, P, As usually form 3 bonds/1 lone pair (metalloids As&Te are similar to those of P&S)
   - C usually forms 4 bonds/0 lone pairs
   - *other configurations might be possible but are probably rare*
6. If octets require more electrons that are available but none of the atoms involved can form multiple bonds, leave the central atom with a sextet or less.
7. If octets aren't working with the most reasonable structure, adding extra electron pairs to the central atom (expanding the octet) is an option for ***large*** atoms only (beyond the second period – e.g., P, S, Cl, Se, Br, I).
8. Rearrange if your first instincts don't work out or if atoms are not in their most common bonding pattern.
9. If you have drawn the most reasonable structure and octets still require more electrons than are available, remove electron pairs as needed and move other pairs to form multiple bonds (original single bond plus 1 or 2 more bonds; each bond = 2 electrons).
10. Sometimes the same formula can correspond to different forms that are different compounds.
11. If the chemical species is an ion, put brackets around the structure and write the charge as a superscript outside the brackets.

VSEPR (Valence Shell Electron Pair Repulsion)

VSEPR is used to predict the shape of the molecule based upon the number of electron pairs around the central atom. For the purpose of predicting shape of the molecule or ion, treat multiple bonds from the central atom as a single electron pair about the central atom.

1. Draw the Lewis structure (see above).
2. Identify the central atom or atoms (you may have to do VSEPR analysis on more than one atom in a structure).
3. For each "central" atom, note the ***number of atoms bonded*** to it and the ***number of non-bonding electron pairs*** (or single, non-bonded electrons in the case of radicals) on it. The number of atoms bonded to the central atom and the number of bonds to it are ***not necessarily the same*** because of the possibility of multiple bonds.
4. Write a formula in the form $\mathbf{AX_nE_m}$, where the central atom is symbolized as **A**, each bonded atom as **X** (of number **n**), and each non-bonding electron pair as **E** (of number **m** – note: lone pairs take up space an contribute to shape!)
5. For example: $CO_2 = AX_2$; $CH_4 = AX_4$; $NH_3 = AX_3E$; $H_2O = AX_2E_2$; **$SF_4 = AX_4E$; $SF_6 = AX_6$ (not on test)**
6. Identify the shape around the central atom from the following table:

e^- groups	e^- group geometry	angles*	bond groups	lone pairs	geometry	VSEPR
2	linear	180°	2	0	linear	AX_2
3	trigonal planar	120°	3	0	trigonal planar	AX_3
			2	1	bent	AX_2E
4	tetrahedral	109.5°	4	0	tetrahedral	AX_4
			3	1	trigonal pyramidal	AX_3E
			2	2	bent	AX_2E_2

*there can be some distortion, depending on the nature of the groups attached to the central atom

1	Atomic Number
H	Element Symbol
2.2	Electronegativity
1.008	Atomic Weight

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H 2.2 1.008																	2 He 4.003
3 Li 0.98 6.941	4 Be 1.57 9.012											5 B 2.04 10.81	6 C 2.55 12.01	7 N 3.04 14.01	8 O 3.44 16.00	9 F 3.98 19.00	10 Ne 20.18
11 Na 0.93 22.99	12 Mg 1.31 24.30											13 Al 1.61 26.98	14 Si 1.9 28.09	15 P 2.19 30.97	16 S 2.58 32.07	17 Cl 3.16 35.45	18 Ar 39.95
19 K 0.82 39.10	20 Ca 1.00 40.08	21 Sc 1.36 44.96	22 Ti 1.54 47.87	23 V 1.63 50.94	24 Cr 1.66 52.00	25 Mn 1.55 54.94	26 Fe 1.83 55.84	27 Co 1.88 58.93	28 Ni 1.91 58.69	29 Cu 1.9 63.55	30 Zn 1.65 65.39	31 Ga 1.89 69.72	32 Ge 2.01 72.61	33 As 2.18 74.92	34 Se 2.55 78.96	35 Br 2.96 79.90	36 Kr 83.80
37 Rb 0.82 85.47	38 Sr 0.95 87.62	39 Y 1.22 88.91	40 Zr 1.33 91.22	41 Nb 1.6 92.91	42 Mo 2.16 95.95	43 Tc 1.9 97.91	44 Ru 2.2 101.1	45 Rh 2.28 102.9	46 Pd 2.2 106.4	47 Ag 1.93 107.9	48 Cd 1.69 112.4	49 In 1.78 114.8	50 Sn 1.96 118.7	51 Sb 2.05 121.8	52 Te 2.1 127.6	53 I 2.66 126.9	54 Xe 131.3
55 Cs 0.79 132.9	56 Ba 0.89 137.3		72 Hf 178.5	73 Ta 180.9	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po 209	85 At 210	86 Rn 222
87 Fr 223	88 Ra 226		104 Rf 261	105 Db 262	106 Sg 263	107 Bh 262	108 Hs 265	109 Mt 266	110 Ds 269	111 Uuu 272	112 Uub 277	113	114 Uuq 289	115	116 Uuh 289	117	118

Lanthanides	57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm 145	62 Sm 150.4	63 Eu 152.0	64 Gd 157.2	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
Actinides	89 Ac 227	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np 237	94 Pu 244	95 Am 243	96 Cm 247	97 Bk 247	98 Cf 251	99 Es 252	100 Fm 257	101 Md 258	102 No 259	103 Lr 262

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H Hydrogen 1.008																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.30											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.84	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.95	43 Tc Technetium 97.91	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3
55 Cs Cesium 132.9	56 Ba Barium 137.3		72 Hf Hafnium 178.5	73 Ta Tantalum 180.9	74 W Tungsten 183.8	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium 209	85 At Astatine 210	86 Rn Radon 222
87 Fr Francium 223	88 Ra Radium 226		104 Rf Rutherfordium 261	105 Db Dubnium 262	106 Sg Seaborgium 263	107 Bh Bohrium 262	108 Hs Hassium 265	109 Mt Meitnerium 266	110 Ds Darmstadtium 269	111 Uuu Unununium 272	112 Uub Ununbium 277	113	114 Uuq Ununquadium 289	115	116 Uuh Ununhexium 289	117	118
	Lanthanides		57 La Lanthanum 138.9	58 Ce Cerium 140.1	59 Pr Praseodymium 140.9	60 Nd Neodymium 144.2	61 Pm Promethium 145	62 Sm Samarium 150.4	63 Eu Europium 152.0	64 Gd Gadolinium 157.2	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0
	Actinides		89 Ac Actinium 227	90 Th Thorium 232.0	91 Pa Protactinium 231.0	92 U Uranium 238.0	93 Np Neptunium 237	94 Pu Plutonium 244	95 Am Americium 243	96 Cm Curium 247	97 Bk Berkelium 247	98 Cf Californium 251	99 Es Einsteinium 252	100 Fm Fermium 257	101 Md Mendelevium 258	102 No Nobelium 259	103 Lr Lawrencium 262