Packet #2 Unit 7

Unit 7: Part 2 → Using Chemical Formulas

NAME ____________________ PERIOD __________ DATE __________

CP Chemistry

A. PERCENT COMPOSITION

THE TERM PERCENTAGE COMPOSITION REFERS TO THE NUMBER OF PARTS PER HUNDRED OF EACH ELEMENT IN A COMPOUND. THIS CAN BE DETERMINED IN 2 WAYS – FROM...

A. A KNOWN CHEMICAL FORMULA B. LABORATORY DATA – MASS MEASUREMENTS

FORMULA →

$$\%\,composition = \frac{mass\ of\ element}{total\ mass} \times 100$$

PART A – KNOWN CHEMICAL FORMULA PRACTICE PROBLEMS

1. Cu_2S

2. FeO

3. Ammonium Sulfate

4. Copper(II) hydroxide

5. Lithium Phosphide

PART B – LABORATORY DATA

1. Analysis of a compound shows that it consists of 43.40 g of copper and 10.95 g of sulfur. What is the percentage composition of this compound?

2. A sample of the organic compound benzene is analyzed and found to consist of 13.74 g carbon and 1.15 g of hydrogen. What is the percentage composition of benzene?

Part C Practice: Show all work below

1. Sulfuric Acid
2. Barium hydroxide
3. Potassium phosphate
4. Analysis of an unknown compound shows that it consists of 21.8 g of oxygen, 4.09 g of aluminum, and 6.36 g of nitrogen. What is the percentage composition of this compound?
5. Analysis of an ore of calcium shows that it contains 13.61 g of calcium and 21.77 g of oxygen in a sample with a mass of 46.28. What is the percent composition of this compound?

Name ____________________ Period____ Date ____________

Unit 7: Part 3→
Determining
Chemical
Formulas

B. Empirical Formulas

- Smallest whole number ratio of atoms in a compound
  - C_2H_6 → reduce subscripts → CH_3

Here is a rhyme made up to help you remember the steps to finding the Empirical Formula
Percent to Mass
Mass to Mole
Divide by small
Multiply til whole

1. Find the mass or percent of each element
2. Convert to moles
3. Divide by the smallest #of moles to find each subscript
4. Is not whole #s, multiply subscripts by 2, 3, or 4 to get hole numbers

Example problem: Find the empirical formula for a sample of 25.9% N and 74.1% O.

C. Molecular Formulas

- "True Formula" – the actual number of atoms in a compound
  - If the Empirical Formula is CH_3 the Molecular Formula May Be C_2H_6

1. Find the Empirical Formula
2. Find the Empirical Formula Mass
3. Divide the true Molecular Mass by the Empirical Mass
4. Multiply each subscript by the answer from step 3

$$\frac{MF\ mass}{EF\ mass} = n \qquad (EF)_n$$

Example Problem: The empirical formula for ethylene is CH_2. Find the molecular formula if the molecular mass is 28.1 g/mol?

Empirical Formula Worksheet

1. Write the Empirical Formula for Each of the Following:

a. P_4O_6 ____________________

b. C_6H_9 ____________________

c. CH_2OHCH_2OH ____________________

d. $BrCl_2$ ____________________

e. $C_6H_8O_6$ ____________________

f. $C_{10}H_{22}$ ____________________

g. $Cu_2C_2O_4$ ____________________

h. Hg_2F_2 ____________________

2. Write the Empirical Formula for Each of the Following (show your work):

a. A compound composed of: 72% iron (Fe) and 27.6% oxygen (O) by mass.

b. A compound composed of: 9.93% carbon (C), 58.6% chlorine (Cl), and 31.4% fluorine (F). (This compound is commonly known as Freon)

c. A compound composed of: .556g carbon (C) and .0933g hydrogen (H).

3. Write the Molecular Formula for Each of the Following:

a. A compound with a molecular mass of 70.0 amu and an empirical formula of CH_2. ____________

b. A compound with a molecular mass of 46.0amu and an empirical formula of NO_2. ____________

4. Can the molecular formula of a compound ever be the same as the empirical formula for the compound? Explain you answer.

5. What is the empirical formula of a compound that has 3 times as many hydrogen atoms as carbon atoms, but only half as many oxygen atoms as carbon?

Name ______________________________ Period ________ Date __________

Formula Calculations - Ch. 7 (p.226-233 in M.C.)

PART A - PERCENTAGE COMPOSITION: Calculate the percent composition of the following compounds:

1. Find the percentage composition of iron(III) oxide. Formula: ______________

2. Find the mass percentage of water in copper(II) sulfate pentahydrate. Formula: __________

3. How many grams of iron can be obtained from a 268-g sample of iron(III) oxide?

PART B - EMPIRICAL & MOLECULAR FORMULA

4. A compound used to test for the presence of ozone in the stratosphere contains 96.2% thallium and 3.77% oxygen. What is its empirical formula?

5. The molecular mass of benzene, an important industrial solvent and know carcinogen, is 78.0 *g/mol* and its empirical formula is CH. What is the molecular formula of benzene?

6. Ascorbic acid, or vitamin C, has a percent composition of 40.9% C, 4.58% H, and 54.5% O. Its molecular mass is 176.1 *g/mol*. Find its <u>empirical</u> and <u>molecular</u> formulas. (HINT: Multiply by 2, 3, or 4 to get whole number subscripts.)

7. A sample of a compound containing carbon and hydrogen has a mass of 88.0 g. Experimental procedures show that 72.0 g of this sample is carbon and the remaining 16.0 g is hydrogen. What is the percentage composition of this compound?

8. What is the empirical formula of the above compound?

9. Determine the empirical formula of a substance that has 40.0% carbon and 53.3% oxygen and the rest hydrogen. If the molecular mass is 180 g/mol, what is the molecular formula?

10. A compound of copper and sulfur was produced in the lab by heating copper and sulfur together in a crucible. This data was collected:

Mass of crucible and cover	28.71g
Mass of crucible, cover and copper	30.25g
Mass of crucible, cover and copper-sulfur compound	30.64

a) calculate the percent composition of the compound
b) determine its empirical formula
c) is it an ionic or molecular compound?
d) Name the compound using both the Stock and traditional systems

EMPIRICAL FORMULA LAB

An empirical formula gives the simplest whole number ratio of the different atoms in a compound. The empirical formula does not necessarily indicate the exact number of atoms in a single molecule. This information is given by the molecular formula, which is always a simple multiple of the empirical formula.

In this experiment, you will determine the empirical formula of a magnesium-oxygen product, a compound that is formed when magnesium metal reacts with oxygen gas. According to the law of conservation of mass, the total mass of the products must equal the total mass of the reactants in a chemical reaction. Therefore,

$$\text{mass Mg} + \text{mass } O_2 = \text{mass } Mg_xO_y$$

Since you will measure the mass of magnesium and the magnesium-oxygen product, you will be able to calculate the mass of oxygen consumed during the reaction. Then, the ratio between the moles of tin and the moles of oxygen consumed can be calculated. Finally, the empirical formula can be written on the basis of this ratio.

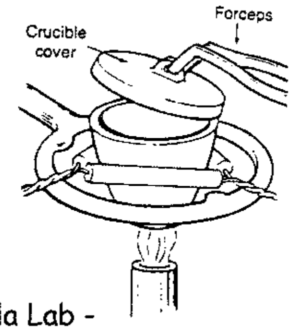
SAFETY

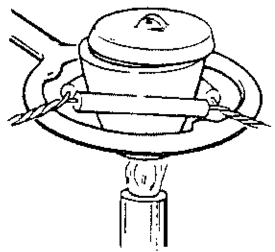
- Use tongs at all times when handling the hot crucible and lid.
- Do not look directly at the burning magnesium. Avoid inhaling the fumes produced while heating.

PROCEDURE

1. Clean and dry a crucible and lid and place them on a clay triangle as shown in Figure 1. To dry them, heat strongly for 2 to 3 minutes over the burner. Then let them cool to room temperature. **CAUTION**: *Hot porcelain. Use tongs to handle the crucible and lid throughout the experiment.*
2. Record the combined mass of the crucible and lid in your data table. **Measurement 1.**
3. Polish a strip of magnesium ribbon with steel wool until it is shiny. Cut the strip into small pieces and place them in the crucible. Record the combined mass of the crucible, lid, and magnesium. **Measurement 2.**
4. Heat the covered crucible gently over the burner. Lift the lid about every 20 seconds to allow air in. **CAUTION**: *Do not look directly at the burning magnesium metal. Avoid inhaling any fumes.*
5. When the magnesium appears to fully reacted (you see not luster), partially remove the crucible lid as shown in Figure 2. Continue heating for 1 minute.
6. Turn off the burner. After the crucible has **cooled to room temperature**, carefully add 2 drops of water to decompose any nitrides that may have formed. **CAUTION**: *Use care when adding water. Too much water can cause the crucible to crack.*
7. Cover the crucible completely. Resume heating with the burner for 2 minutes.

8. Turn off the burner. Cool the crucible, lid, and contents to room temperature. Record the combined mass of the crucible, cover, and product. **Measurement 3.**
9. Clean up your lab station. Discard the solid product in the designated waste beaker. Wash and towel-dry the crucible and lid. Wash you hands.





Empirical Formula Lab -

DATA

Construct a data table to display the measurements taken during the procedure. Make sure that each value is clearly labeled, including units.

ANALYSIS

Show all calculations and construct an Analysis Table to display your answers.

1. Use your data to calculate the mass of magnesium and the mass of oxygen in the product.
2. Determine the empirical formula of the magnesium-oxygen product. When calculating the mole ratio (**subscripts**), <u>round to the nearest whole number</u>. In practice problems, you should multiply by 2, 3, etc. to get a whole number ratio. In this case, you need to round in order to compensate for experimental error.
3. Use the masses from #1 to calculate the percent composition (%Mg and %O) of the product.

CONCLUSIONS

- Based on the charges of each element, write the formula for magnesium oxide. Does your experimental empirical formula agree with this formula?
- The literature value for the %Mg in this magnesium-oxygen compound is 60.3%. Use this value to calculate the percent error of your experimental %Mg. Comment on your degree of accuracy.
- Is your %Mg value too large or too small? What experimental errors might specifically account for <u>this type</u> of deviation? (This requires some in-depth analysis of how errors in your original data would affect your final answer, causing it to be too large or too small. It often helps to work backwards through your calculations. Then, you need to think about what might have happened during the experiment to create these errors. This requires higher-order thinking skills, but you can do it! Use your brain!)

Empirical Formula Lab - Results and Conclusions

Qualitative and Quantitative DATA:

Appearance of magnesium ribbon	
Mass of crucible and lid	
Mass of crucible, lid and Mg ribbon	
Mass of crucible lid, and product	
Appearance of the product	

ANALYSIS

Show all calculations and construct an Analysis Table to display your answers.

4. Use your data to calculate the mass of magnesium and the mass of oxygen in the product.

5. Determine the empirical formula of the magnesium-oxygen product. When calculating the mole ratio (**subscripts**), round to the nearest whole number. In practice problems, you should multiply by 2, 3, etc. to get a whole number ratio. In this case, you need to round in order to compensate for experimental error.

6. Use the masses from #1 to calculate the percent composition (%Mg and %O) of the product.

CONCLUSIONS

- Based on the charges of each element, write the formula for magnesium oxide. Does your experimental empirical formula agree with this formula?

- The literature value for the %Mg in this magnesium-oxygen compound is 60.3%. Use this value to calculate the percent error of your experimental %Mg. Comment on your degree of accuracy.

$$\% \text{ Error} = \frac{|\text{actual value} - \text{experimental value}|}{\text{actual value}} \times 100$$

- Is your %Mg value too large or too small? What experimental errors might specifically account for this type of deviation? (This requires some in-depth analysis of how errors in your original data would affect your final answer, causing it to be too large or too small. It often helps to work backwards through your calculations. Then, you need to think about what might have happened during the experiment to create these errors. This requires higher-order thinking skills, but you can do it! Use your brain!)

CP CHEMISTRY TEST STUDY GUIDE
Ch. 3 & 7 – The Mole

Test Date: ____________________

Read over your notes, and *rework* your homework assignments and do your review problems. Reviewing your labs may also be helpful. You are responsible for knowing the conversion factors and "formulas" for %composition, empirical formulas, etc. You will do AWESOME on this test if you can do the following things.

CHEMICAL FORMULAS

- Write the correct formula for ionic and molecular compounds and acids
- Name ionic and molecular compounds and acids
- Calculate the correct formula and molecular masses for compounds

MOLAR CONVERSIONS – p. 80-85, 221-226

- Explain the meaning of the mole.
- Explain how molar mass and Avogadro's number are used as conversion factors.
- Calculate molar mass of an element or compound.
- Perform molar conversions between grams, moles, & particles(molecules or atoms).

FORMULA CALCULATIONS – p. 226-233

- Explain the meaning of % composition, empirical formula, and molecular formula.
- Calculate % composition. Use it to determine the mass of an element in a sample.
- Calculate empirical and molecular formulas for unknown compounds from percent compositions and experimental data.

CP CHEMISTRY - THE MOLE REVIEW (from Chapter 7)

****ALL ANSWERS MUST INCLUDE THE PROPER UNITS & SIG FIGS.****

Show all work on a separate paper

SOLVE THE FOLLOWING MOLAR CONVERSION PROBLEMS:

1. How many grams would 8.1×10^{21} molecules of sucrose ($C_{12}H_{22}O_{11}$) weigh?
2. How many moles are in 53.8 g of magnesium chloride?
3. How many representative particles (molecules) are in 0.845 moles of $NaNO_3$?
4. How many molecules are in 50.0 g of calcium sulfide?
5. How many atoms are in a 2.0 kg bar of gold? (Note mass units.)

SOLVE THE FOLLOWING PERCENTAGE COMPOSITION PROBLEMS:

6. Find the percentage composition of sucrose ($C_{12}H_{22}O_{11}$).
7. Find the percentage composition of a sample containing 1.29 g of carbon and 1.71 g of oxygen.
8. Find the mass percentage of water in sodium carbonate decahydrate.
9. How many grams of zinc are in a 37.2-gram sample of zinc nitrate?

SOLVE THE FOLLOWING EMPIRICAL & MOLECULAR FORMULA PROBLEMS:

10. Find the empirical formula of a compound that contains 75% carbon and 25% hydrogen.
11. Find the empirical formula of a compound that contains 9.03 g magnesium and 3.48 g of nitrogen.
12. The empirical formula of a compound is NO_2. Its molecular mass is 92 g/mol. What is its molecular formula?
13. Glucose has an empirical formula of CH_2O. Find its molecular formula if its molecular mass is 180.0 g/mol.
14. A compound is composed of 34.2% sodium, 17.7% carbon, and 47.6% oxygen. Find its empirical formula. If its molecular mass is 134 g/mol, find its molecular formula.
15. A compound of copper and sulfur was produced in the lab by heating copper and sulfur together in a crucible. This data was collected:

Mass of crucible and cover	28.71g
Mass of crucible, cover and copper	30.25g
Mass of crucible, cover and copper-sulfur compound	30.64

a) calculate the percent composition of the compound
b) determine its empirical formula
c) is it an ionic or molecular compound?
d) Name the compound using both the Stock and traditional systems

Guide Sheet for Moles Problems

I. Calculating Molar Mass

1. multiply atomic mass of each element by number of atoms of that element in the formula (shown by the subscript)
2. find the sum of all the atomic masses --this is formula mass (unit is a.m.u.)
3. express formula mass in grams (unit is g/mol). This is the **Molar Mass.**

II. Calculating % Composition (from formula)

1. calculate formula mass
2. divide the total atomic mass of each element by the formula mass and multiply by 100

III. Calculating % Composition (from masses of each element)

1. divide the mass of each element by the total mass of the compound and multiply by 100

IV. Calculating Empirical Formula (from % Composition)

1. convert % of each element to grams based on 100 grams of the compound
2. multiply grams of each element by 1/molar mass that element
3. compare ratio of moles of each element and divide each by the smallest
4. if result in step 3 gives a ratio with decimal equivalent to 1/4, 1/3, 1/2, 2.3, 3/4 instead of whole numbers, convert to the fraction and multiply all ratios by the denominator or the fraction

V. Calculating Empirical Formula (from experimentally determined masses)

1. multiply the mass of each element (in grams) by 1/molar mass of that element
2. continue with steps 3 & 4 from IV above.

VI. Finding Molecular Formulas (when molar mass is known)

1. calculate the empirical formula
2. use the equation : (empirical formula mass)x = molar mass
3. find value for x: x = molar mass/empirical formula mass
4. multiply each subscript in empirical formula by value for x