# EMPIRICAL, MOLECULAR FORMULAS AND CHEMICAL ANALYSIS PROBLEMS

# PERCENT COMPOSITION

- Percent of each element a compound is composed of.
- Strategy:
  - Assume 1 mole of substance. Get molar mass of substance.
  - Find the mass of each element in a mole, divide by the total molar mass, multiply by a 100.
- Ex: Find the percent composition of $CH_4$
  - Molar mass of $CH_4$ = 16.043 g/mol
    - % C = 12.011 g/mol / 16.043 g/mol = 74.9%
    - % H = 4.032 g/mol / 16.043 g/mol = 25.1%

# EMPIRICAL VS MOLECULAR FORMULAS

- Empirical Formula the lowest ratio of atoms in a molecule.
- Molecular Formula is the actual # and types of atoms in a compound
  - Ex: Molecular formula for glucose is $C_6H_{12}O_6$
  - Glucose has an empirical formula of $CH_2O$.

# EMPIRICAL FORMULA FROM % COMPOSITION

- Can use percent composition information to obtain an empirical formula:
  - Ex: Mercury reacts with chlorine to produce a compound that is 73.9% Hg and 26.1% Cl by mass. What is the empirical formula for the compound?
  - Strategy:
  - 1) assume you have 100.0 g of the substance:
    - 73.9 g Hg
    - 26.1 g Cl
  - 2) convert mass to moles of each element
    - 73.9 g Hg x (1 mol/ 200.59 g) = 0.37 mol Hg
    - 26.1 g Cl x (1 mol/ 35.453 g) = 0.74 mol Cl
  - 3) Divide larger mole quantity by lowest mole quantity to get mole ratios:
    - 0.74 mol Cl/ 0.37 mol Hg = ~2 mol Cl : 1 mol Hg
    - **Therefore Empirical Formula is $HgCl_2$**

# NOW YOU TRY:

- A compound that is usually used as a fertilizer can also be used as a powerful explosive. The compound has the composition, 35.00% nitrogen, 59.96% oxygen and the remainder being hydrogen. What is its empirical formula? Given it is ionic, suggest a name for the compound.
- **Answer: $N_2O_3H_4$**
- **Ionic compound formula: $NH_4NO_3$**
- **Name: ammonium nitrate**

# EMPIRICAL TO MOLECULAR FORMULAS

- To get from E.F. to M.F. you will need the molar mass of the actual compound.
- Ex: A compound is made of only sulfur and nitrogen. It is 69.6% S by mass. Its molar mass is 184 g/mol. What is its formula?
- Strategy:
  1) Find the empirical formula first
     - **Answer: SN**
  2) Divide the molar mass of the substance by the molar mass of its empirical formula.
     - **MM of SN = 46 g/mol**
     - **MM of compound is 184 g/mol**
     - **184 g/mol / 46 g/mol = 4**
  3) Multiply each atom by this factor.
     - **Answer: $S_4N_4$**

# NOW YOU TRY:

- A substance has an empirical formula of $CH_2Br$ and a molar mass of 188 g mol-1. What is the molecular formula of the compound?
- **Answer: $C_2H_4Br_2$**

# CHEMICAL ANALYSIS

- Ex #1: Hydrates
  Washing soda, a hydrate, has a formula of $Na_2CO_3 \bullet xH_2O$, where *x* is the number of moles of $H_2O$ per mole of $Na_2CO_3$. When a 2.558 g sample of washing soda is heated, all the water of hydration is driven off, leaving 0.948 g $Na_2CO_3$. What is the formula for the hydrate (what is the value of *x*)?
- Strategy:
  - 1) determine how many moles of water the hydrate had by subtracting mass of $Na_2CO_3$ after it has been heated from the mass of the original sample:
    - 2.558 g – 0.948 g = 1.61 g $H_2O$
  - 2) Determine the mole ratio of $Na_2CO_3$ to $H_2O$ by dividing moles of $H_2O$ by moles of $Na_2CO_3$
    - 0.948 g $Na_2CO_3$ x (1mole/105.989 g) = .008945 mol $Na_2CO_3$
    - 1.61 g $H_2O$ x (1 mol/ 18.015 g) = .08937 mol $H_2O$
    - Therefore: .08936 mol $H_2O$ / .008945 mol $Na_2CO_3$ = 9.99 = ~10
    - That means that for every 1 mole of $Na_2CO_3$ there are 10 moles of $H_2O$. So the formula for the hydrate is: **$Na_2CO_3 \bullet 10\ H_2O$**

# CHEMICAL ANALYSIS

- Ex #2: Combustion

A 0.2000 gram sample of a compound (vitamin C) composed of only C, H, and O is burned completely with excess $O_2$ . 0.2998 g of $CO_2$ and 0.0819 g of $H_2O$ are produced. What is the empirical formula?

- strategy:
  - 1) Write as much of the reaction as you can:
    - $C\_H\_O\_ \quad + \quad O_2 \quad \rightarrow \quad CO_2 \quad + \quad H_2O$
    - 0.200 g 0.2998 g 0.0819 g

- 2) Think about what you know…
  - All the carbon in the $CO_2$ had to come from the original compound
  - All the Hydrogen in the $H_2O$ has to come from the original compound
  - The oxygen in the products came from two places: the original compound and the $O_{2=}$ gas it was combusted with. So we will deal with this later…
- 3) Determine the moles of C and H since you know where they came only from the original sample and you have a mass for both $CO_2$ and $H_2O$
  - 0.2998 g $CO_2$ x (1 mol / 44.009 g) = .0068122 mol $CO_2$ x (1 mole C/1 mole $CO_2$) = **.0068122 mole C**
  - 0.0819 g $H_2O$ x (1 mole/ 18.015 g) = .0045462 mol $H_2O$ x (2 mole H/1 mole $H_2O$)= .**0090924 mole H**

- 4) Determine the moles of Oxygen by subtracting the mass of C and mass of H from the mass of the original compound:
  - .0068122 mole C x (12.011g/mol) = .081821 g C
  - .0090924 mole H x (1.008 g/mol) = .0091651 g H
  - .200 g C_H_O – (.081821 g C) –( .0091651 g H) = .10901 g O x (1 mol/ 15.999 g) = **.0068134 mole O**
- 5) Determine the Empirical formula by dividing each mole quantity by the smallest one:
  - .0068134 mol O /.0068122 mol C = 1 mol O to 1 mol C
  - .0090924 mol H / .0068122 mol C = 1.33 mol H to mol C
  - Therefore $C\ H_{1.33}\ O$
  - **To get whole number, multiply by 3: $C_3H_4O_3$**