Empirical & Molecular Formulas

I. Empirical Vs. Molecular Formulas

- *Molecular Formula* = actual/exact # of atoms in a compound (ex: Glucose = $C_6H_{12}O_6$)

- *Empirical Formula* = lowest whole # ratio of atoms in a compound (ex: Glucose = CH_2O)

II. Determining Empirical Formulas

- You can determine the empirical formula of a compound from % composition information or by knowing the mass of each element present in the whole compound

- Empirical Formula Rhyme (to help you remember the steps):

  Percent to Mass,
  Mass to Mole,
  Divide by small,
  Multiply 'till whole

- **Examples:**

1. *An experiment uses a catalyst that is 23.3 % cobalt, 25.3 % molybdenum, and 51.4 % chlorine. What is the empirical formula of this catalyst?*

Step #1: <u>Percent to mass</u> → Assume that you are working with a 100 gram sample, so the % of each element is equal to the mass of each element in grams (show this during step #2)

Step #2: <u>Mass to mole</u> → Convert the mass of each element to moles of each element using the molar mass.

Step #3: <u>Divide by small</u> → Divide each of the mole quantities by the smallest number of moles. Often, this will result in whole (or practically whole) numbers.

Step #4: <u>Multiply 'till whole</u> → If step #3 does not result in whole numbers, find the least common multiple that will achieve all whole numbers. These whole numbers represent the subscripts for each element in the empirical formula.

$\frac{23.3\text{ g Co}}{} \cdot \frac{1\text{ mol Co}}{58.93\text{ g Co}} = 0.3937 / 0.2637 = 1.5 \times 2 = 3$

$\frac{25.3\text{ g Mo}}{} \cdot \frac{1\text{ mol Mo}}{95.94\text{ g Mo}} = 0.2637 / 0.2637 = 1 \times 2 = 2$

$\frac{51.4\text{ g Cl}}{} \cdot \frac{1\text{ mol Cl}}{35.45\text{ g Cl}} = 1.450 / 0.2637 = 5.5 \times 2 = 11$

The empirical formula for this compound is: $Co_3Mo_2Cl_{11}$

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2. a) *Nicotine is a stimulant and an addictive chemical found in tobacco. An analysis of nicotine produces the following percent composition: 74.03% carbon, 17.27% nitrogen, and 8.70% hydrogen. What is the empirical formula of nicotine?*

$\frac{74.03\,g\,C}{} \bigg| \frac{1\,mol\,C}{12.01\,g\,C} = 6.164 / 1.2327 = 5$

$\frac{17.27\,g\,N}{} \bigg| \frac{1\,mol\,N}{14.01\,g\,N} = 1.2327 / 1.2327 = 1$

$\frac{8.70\,g\,H}{} \bigg| \frac{1\,mol\,H}{1.008\,g\,H} = 8.6310 / 1.2327 = 7$

C_5NH_7

b) *Further tests show that the molar mass of nicotine is 162.23 g/mol. Given this information, what is the molecular formula of nicotine?*

$\frac{162.23\,g/mol}{81.116\,g/mol} = 2$ $C_{10}N_2H_{14}$

3. *An ionic sample with a mass of 0.5000 g is determined to contain the elements indium and chlorine. If the sample has 0.2404 g of chlorine, what is the empirical formula of this ionic compound?*

0.5000
$- 0.2404$ Cl
$0.2598\,g$ In

$\frac{0.2402\,g\,Cl}{} \bigg| \frac{1\,mol\,Cl}{35.45\,g\,Cl} = 0.006776 / 0.002263 = 3$

$\frac{0.2598\,g\,In}{} \bigg| \frac{1\,mol\,In}{114.82\,g\,In} = 0.002263 / 0.002263 = 1$

$InCl_3$

Determining Empirical and Molecular Formulas

1. a) What is the empirical formula of a molecule containing 65.5% carbon, 5.5% hydrogen, and 29.0% oxygen?

$$\frac{65.5\,g\,C}{} \Big| \frac{1\,mol\,C}{12.01\,g\,C} = 5.4538/1.8125 = 3$$

$$\frac{5.5\,g\,H}{} \Big| \frac{1\,mol\,H}{1.008\,g\,H} = 5.4563/1.8125 = 3$$

$$\frac{29.0\,g\,O}{} \Big| \frac{1\,mol\,O}{16.00\,g\,O} = 1.8125/1.8125 = 1$$

C_3H_3O

b) If the molar mass of the compound in part a (above) is 110 grams/mole, what is the molecular formula?

$$\frac{110\,g/mol}{55.054\,g/mol} = 2$$

$C_6H_6O_2$

2. a) Find the empirical formula of a molecule containing 18.7% lithium, 16.3% carbon, and 65.0% oxygen.

$$\frac{18.7\,g\,Li}{} \Big| \frac{1\,mol\,Li}{6.94\,g\,Li} = 2.6945/1.3572 = 2$$

$$\frac{16.3\,g\,C}{} \Big| \frac{1\,mol\,C}{12.01\,g\,C} = 1.3572/1.3572 = 1$$

$$\frac{65.0\,g\,O}{} \Big| \frac{1\,mol\,O}{16.10\,g\,O} = 4.0625/1.3572 = 3$$

Li_2CO_3

b) If the molar mass of the compound in part a (above) is 73.8 grams/mole, what is the molecular formula?

$$\frac{73.8\,g/mol}{73.89} = 1$$

Li_2CO_3

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III. Empirical Formulas of Hydrates

- Empirical formulas of hydrates can be determined by essentially calculating the mol-to-mol ratio of the ionic salt-to-water molecules present in a hydrate's solid crystal structure
    - Recall copper (II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$)
        - this hydrate has 5 mols of H_2O for every 1 mol of $CuSO_4$
    - If you heat a hydrate enough, you can drive out all the water of hydration (water trapped in the hydrate solid leaving behind only the anhydrous ionic salt (in this case, $CuSO_4$)
    - Given the initial mass of the hydrate and either the mass of anhydrous salt remaining, or the mass of water lost, the empirical formula of the hydrate can be determined
- **Example:**

1. *Washing soda, a compound used to prepare hard water for washing laundry, is a hydrate which means that a certain number of water molecules are included in the solid structure. Its formula can be written $NaCO_3 \cdot xH_2O$ where the x indicates the number of moles of H_2O per mole of $NaCO_3$. When 2.558 g of washing soda is heated so that all the water of hydration is lost, only 0.948 g of $NaCO_3$ remains. What is the value of x? Using chemical nomenclature, how would you name this hydrate compound?*

Step #1: Mass of H_2O lost → Use the initial mass of hydrate sample and the mass of anhydrous salt left after heating to determine the mass of H_2O lost (water of hydration trapped in the hydrate)

Step #2: Grams of ionic salt & H_2O → Use the mass of H_2O and its molar mass as well as the mass of the ionic salt and its molar mass to determine the moles of each present in the hydrate. Then divide by small (smaller should always be the moles of ionic salt) to determine the # of moles H_2O per 1 mole of ionic salt present in the hydrate

hydrate 2.558g
AS 0.948g
water 1.610g

$0.948 g\ NaCO_3 \times \frac{1\ mol\ NaCO_3}{\cancel{82.99}\ 105.99\ g\ NaCO_3} = \cancel{0.01142}/\cancel{0.01142}\ 0.00894 = 1$

$1.610 g\ H_2O \times \frac{1\ mol\ H_2O}{18.016\ g\ H_2O} = 0.08937/0.01142 - \cancel{7.8}\ 10$

$NaCO_3 \cdot \cancel{8}\ 10\ H_2O$

Sodium carbonate ~~octa~~ deca hydrate

Empirical and Molecular Formulas: Hydrates

1. Epsom salts, a strong laxative used in veterinary medicine, is a hydrate. The formula for Epsom salts can be written as $MgSO_4 \bullet xH_2O$ where the x indicates the number of moles of H_2O per mole of $MgSO_4$. When 13.52 g of this hydrate is heated and all the water of hydration is lost, only 6.60 grams of the anhydrous salt remains.

   a. What is the value of x?

hydrate: 13.52g
AS: 6.60g $MgSO_4$
H_2O: 6.92g H_2O

$$\frac{6.60\text{g } MgSO_4}{} \left| \frac{1 \text{ mol } MgSO_4}{120.36 \text{ g } MgSO_4} \right. = 0.0548/0.0548 = 1$$

$$\frac{6.92\text{g } H_2O}{} \left| \frac{1 \text{ mol } H_2O}{18.016 \text{ g } H_2O} \right. = 0.3841/0.0548 = 7$$

X = 7 $MgSO_4 \cdot 7H_2O$

   b. Using chemical nomenclature, how would you name this hydrate compound?

magnesium sulfate heptahydrate

2. When 8.00 g of $Pb(C_2H_3O_2)_2 \bullet xH_2O$ are heated, 1.14 g of H_2O are driven off. Determine the chemical formula and name of this hydrate.

hydrate - 8.00g
AS - 6.86g $Pb(C_2H_3O_2)_2$
H_2O - 1.14g H_2O

$$\frac{6.86\text{g } Pb(C_2H_3O_2)_2}{} \left| \frac{1 \text{ mol}}{325.288 \text{ g}} \right. = 0.02109/0.02109 = 1$$

$$\frac{1.14\text{g } H_2O}{} \left| \frac{1 \text{ mol}}{18.016 \text{ g } H_2O} \right. = 0.06328/0.02109 = 3$$

$Pb(C_2H_3O_2)_2 \cdot 3H_2O$
lead(IV) acetate trihydrate

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