Honors Chemistry | Notes

Name: Key Date: ______ Mods: ______

Unit 7: Empirical Formulas, Combustion & Hydrate Analysis

I. Empirical Formulas by Combustion Analysis

- Empirical formulas of compounds containing carbon, hydrogen, and oxygen can also be determined if given the results of hydrocarbon combustion [$C_XH_YO_Z$ (?) + O_2 (g) → CO_2 (g) + H_2O (g)]
  - Recall that the carbon in the hydrocarbon compound is converted to CO_2 and the hydrogen in the compound is converted to H_2O.
  - From the masses of CO_2 and H_2O produced, we can calculate the number of moles of C and H in the original compound.
  - If oxygen is also present in the compound, its mass can be determined by subtracting the masses of C and H from the compounds original mass

EXAMPLE PROBLEM

1. ***Isopropyl alcohol, a substance sold as rubbing alcohol, is composed of C, H, and O. Combustion of 0.255 g of isopropyl alcohol produces 0.561 g of CO_2 and 0.306 g of H_2O. Determine the empirical formula of isopropyl alcohol.***

Step #1: *Grams C* → Use the mass of CO_2, carbon dioxide's molar mass, and that fact that in 1 mole CO_2 there is 1 mole of C, to determine the grams of carbon in the original compound.

Step #2: *Grams H* → Use the mass of H_2O, water's molar mass, and that fact that in 1 mole H_2O there are 2 moles of H, to determine the grams of hydrogen in the original compound.

Step #3: *Grams O*→ Use the total mass of the original sample and subtract the determined masses of C and H to find the mass of just oxygen in the original compound.

Step #4: *Find moles of C, H, and O*→ Use the masses (in grams) and molar masses of C, H, and O respectively to determine the number of moles of each present. Next divide by small and multiply 'till whole to determine the empirical formula of the compound.

C: $\frac{0.561\ \cancel{g\ CO_2}}{} \times \frac{1\ \cancel{mol\ CO_2}}{44\ \cancel{g\ CO_2}} \times \frac{1\ \cancel{mol\ C}}{1\ \cancel{mol\ CO_2}} \times \frac{12\ g\ C}{1\ \cancel{mol\ C}} = 0.153\ g\ C$

H: $\frac{0.306\ \cancel{g\ H_2O}}{} \times \frac{1\ \cancel{mol\ H_2O}}{18\ \cancel{g\ H_2O}} \times \frac{2\ \cancel{mols\ H}}{1\ \cancel{mol\ H_2O}} \times \frac{1\ g\ H}{1\ \cancel{mol\ H}} = 0.034\ g\ H$

0.255 g sample
−0.153 g C
−0.034 g H
0.068 g O

$\frac{0.153\ \cancel{g\ C}}{} \times \frac{1\ mol\ C}{12\ \cancel{g\ C}} = \frac{0.0128\ mol\ C}{0.00425} = 3$

$\frac{0.034\ \cancel{g\ H}}{} \times \frac{1\ mol\ H}{1\ \cancel{g\ H}} = \frac{0.034\ mol\ H}{0.00425} = 8$

$\frac{0.068\ \cancel{g\ O}}{} \times \frac{1\ mol\ O}{16\ \cancel{g\ O}} = \frac{0.00425\ mol\ O}{0.00425} = 1$

empirical formula = C_3H_8O

II. 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 PROBLEM

2. ***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 $Na_2CO_3 \cdot xH_2O$ where the x indicates the number of moles of H_2O per mole of Na_2CO_3. When 2.558 g of washing soda is heated so that all the water of hydration is lost, only 0.948 g of Na_2CO_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

2.558 g hydrate
−0.948 g Na_2CO_3 (anhydrous salt)

1.61 g H_2O lost

$$\frac{0.948\ \cancel{g\ Na_2CO_3}}{} \left| \frac{1\ mol\ Na_2CO_3}{106\ \cancel{g\ Na_2CO_3}} \right. = \frac{0.00894\ mols\ Na_2CO_3}{0.00894} = 1$$

$$\frac{1.61\ \cancel{g\ H_2O}}{} \left| \frac{1\ mol\ H_2O}{18\ \cancel{g\ H_2O}} \right. = \frac{0.084\ mols\ H_2O}{0.00894} = 10$$

$Na_2CO_3 \cdot 10\ H_2O$ = Sodium Carbonate Decahydrate