Name ______________________ **Date**_________

Percent Composition, Hydrates, Empirical and Molecular Formulas!

Percent composition: the percentage by mass of each element in a compound.

- Percentage = (part / total)x100%

Hydrate: a chemical compound that contains chemically bound water molecules.

Empirical formula: simplest whole-number ratio of the atoms in the compound.

Molecular formula: whole-number multiple of the empirical formula.

- Shows the true composition.
- Molar mass of a compound = molar mass of the empirical formula x a whole number, *n*.

Part 1 (Review). Finding Percent composition. **Divide** the mass of the individual element within the compound by the entire mass of the compound. **Multiply by 100** to get the percent! (Part/whole *100%)

Ex. 1 mole of $Fe(NO_3)_3$ Contains 1 mole of iron, 3 moles of Nitrogen, and 9 moles of Oxygen.

Total molar mass = $\underset{Fe}{55.845} + 3*\underset{N}{14.007} + 9*\underset{O}{15.999} = 241.85$ g/mol

$$\frac{Mass\ of\ 1\ mole\ of\ Fe\ atoms}{Molar\ Mass\ of\ Fe(NO_3)_3} \rightarrow \frac{55.845 g/mol}{241.85 g/mol} * 100\% = 23.1\%\ Fe\ by\ mass!$$

$$\frac{Mass\ of\ 3\ mole\ of\ N\ atoms}{Molar\ Mass\ of\ Fe(NO_3)_3} \rightarrow \frac{3*14.007 g/mol}{241.85 g/mol} * 100\% = 17.4\%\ N\ by\ mass!$$

$$\frac{Mass\ of\ 9\ mole\ of\ O\ atoms}{Molar\ Mass\ of\ Fe(NO_3)_3} \rightarrow \frac{9*15.999 g/mol}{241.85 g/mol} * 100\% = 59.5\%\ O\ by\ mass!$$

<u>Now you try. Find the percent composition of Aluminum sulfate: $Al_2(SO_4)_3$</u>

Part 2 % composition → Empirical Formula

1. Convert % → g.
2. Convert g → mol : # g x $\frac{1 mol}{molar\ mass\ g} = \#mol$
3. Compare these amounts in mol to find the simplest whole-number ratio among the elements. (Divide each by the smallest number of moles. If you need to, double or even triple the answers to make them into whole number ratios.) These ratios will be the subscripts.
4. **Ex.** Copper bromide is 28.5% Copper and 71.5% Bromide by mass. What is its empirical formula?

   Cu: $28.5\ g\ Cu * \frac{1\ mol\ Cu}{63.546\ g\ Cu} = 0.448\ mol\ Cu \rightarrow \frac{0.448\ mol\ Cu}{0.448\ mol\ Cu} = 1$ → $\boxed{CuBr_2}$

   Br: $71.5\ g\ Br * \frac{1\ mol\ Br}{79.904\ g\ Br} = 0.895\ mol\ Br \rightarrow \frac{0.895\ mol\ Br}{0.448\ mol\ Cu} = 1.997 \approx 2$

Ex. Aluminum oxide is 52.9% Aluminum and 47.1% Oxygen by mass. What is its empirical formula?

Al: $52.9\ g\ Al * \frac{1\ mol\ Al}{26.982\ g\ Al} = 1.96\ mol\ Al \rightarrow \frac{1.96\ mol\ Al}{1.96\ mol\ Al} = 1 * 2 = 2$ → $\boxed{Al_2O_3}$

O: $47.1\ g\ O * \frac{1\ mol\ O}{15.999\ g\ O} = 2.94\ mol\ O \rightarrow \frac{2.94\ mol\ O}{1.96\ mol\ Al} = 1.5 * 2 = 3$

<u>Now you try. Find the empirical formula of a compound that is 46.6% Nitrogen and 53.4% Oxygen.</u>

Part 3 Hydrates!

Hydrates contain chemically bound water molecules. These waters can be removed by evaporation (heating). By subtracting the mass of the dehydrated compound from the mass of the hydrate, you can begin by finding the mass of the attached waters. If you know the percent composition of the hydrate, you can transform these into masses and start from this point as well.

1. Determine the mass of each the chemical compound and the bound/now freed waters.
2. Convert each mass into moles: # g x $\frac{1mol}{molar\ mass\ g} = \#mol$
3. Determine the mole ratio (Divide each by the smallest number of moles. If you need to, double or even triple the answers to make them into whole number ratios.) This ratios will tell you the number of water molecules bound to the compound.

Ex. 100g of $MgSO_4 \bullet xH_2O$ was heated and 48.8g of $MgSO_4$ remained: $MgSO_4 \cdot xH_2O\ (s) \rightarrow MgSO_4\ (s) + xH_2O\ (g)$

Mass of water: 100g M $MgSO_4 \bullet xH_2O$ – 48g $MgSO_4$ = 52g H_2O

$MgSO_4$: $48.8g\ MgSO_4 * \frac{1\ mol\ MgSO_4}{120.366\ g\ MgSO_4} = 0.405\ mol\ MgSO_4 \rightarrow \frac{0.405\ mol\ MgSO_4}{0.405\ mol\ MgSO_4} = 1$

H_2O: $51.2g\ H_2O * \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} = 2.84\ mol\ H_2O \rightarrow \frac{2.84\ mol\ H_2O}{0.405\ mol\ MgSO_4} = 7 \rightarrow x$

$\boxed{MgSO_4 \bullet 7H_2O}$

Now you try: 10.00g of $MgCO_3 \bullet xH_2O$ was heated and left 4.834 g of $MgCO_3$. What is the formula for the hydrate?

Part 4 Empirical → Molecular Formulas Empirical formulas have only the simplest whole number ratio of elements, the molecular formula contains whole number multiples of these ratios.

1. Find the molar mass of the empirical formula.
2. Divide the experimentally determined (measured in a lab) mass of the molecule by the molar mass of the empirical formula to determine what multiple of the empirical formula to use.
3. Multiply all subscripts by this multiple.

Ex. The empirical formula for a compound is P_2O_5. Its experimental molar mass is 284 g/mol. Determine the molecular formula of the compound.

Molar mass of the empirical formula P_2O_5:

2 × molar mass of P = 61.94 g/mol
+ 5 × molar mass of O = 80.00 g/mol
molar mass of P_2O_5 = 141.94 g/mol

$$n = \frac{\boldsymbol{experimental\ molar\ mass}}{\boldsymbol{molar\ mass\ of\ empirical\ formula}} = \frac{\mathbf{284\ g/mol}}{\mathbf{141.94\ g/mol}} = \mathbf{2.00}$$

n(empirical formula) = 2(P_2O_5) = $\boxed{P_4O_{10}}$

Now you try: A compound has an experimental molar mass of 78 g/mol. Its empirical formula is CH. What is its molecular formula?