Stoichiometry 3.6

Empirical and Molecular Formulas

Molecular and Empirical Formulas

Molecular Formulas

- Chemical formulas that provide the actual number of each type of atom in molecule.

Empirical Formulas

- Chemical formulas that provide the relative number of each type of atom in molecule. (a ratio in simplest form)

Molecular and Empirical Formulas

Common Name	Molecular Formula	Empirical Formula
Hydrogen Peroxide	H_2O_2	HO
Ethylene	C_2H_4	CH_2
Dextrose	$C_6H_{12}O_6$	CH_2O

Molecular and Empirical Formulas

Common Name	Molecular Formula	Empirical Formula
Water	H_2O	
Acetylene	C_2H_2	
Hydrazine	N_2H_4	

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Water	H_2O	H_2O
Acetylene	C_2H_2	CH
Hydrazine	N_2H_4	NH_2

Ex1) Empirical Formula

Ex1) A sample of caffeine was found to contain 49.5% Carbon, 28.9% Nitrogen, 16.5% Oxygen, and 5.1% Hydrogen by mass.

Find the empirical formula for caffeine.

Ex1) Empirical Formula (cont.)

Step 1. Assume you have a 100g sample and convert each element into moles.

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$$\textit{Carbon}: 49.5\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 4.12\text{ mol C}$$

Ex1) Empirical Formula (cont.)

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$$\textit{Carbon}: 49.5\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 4.12\text{ mol C}$$

$$\textit{Hydrogen}: 5.1\text{ g H} \times \frac{1\text{ mol H}}{1.01\text{ g H}} = 5.0\text{ mol H}$$

Ex1) Empirical Formula (cont.)

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$$\textit{Hydrogen}: 5.1\ \text{g H} \times \frac{1\ \text{mol H}}{1.01\ \text{g H}} = 5.0\ \text{mol H}$$

$$\textit{Nitrogen}: 28.9\ \text{g N} \times \frac{1\ \text{mol N}}{14.01\ \text{g N}} = 2.06\ \text{mol N}$$

Ex1) Empirical Formula (cont.)

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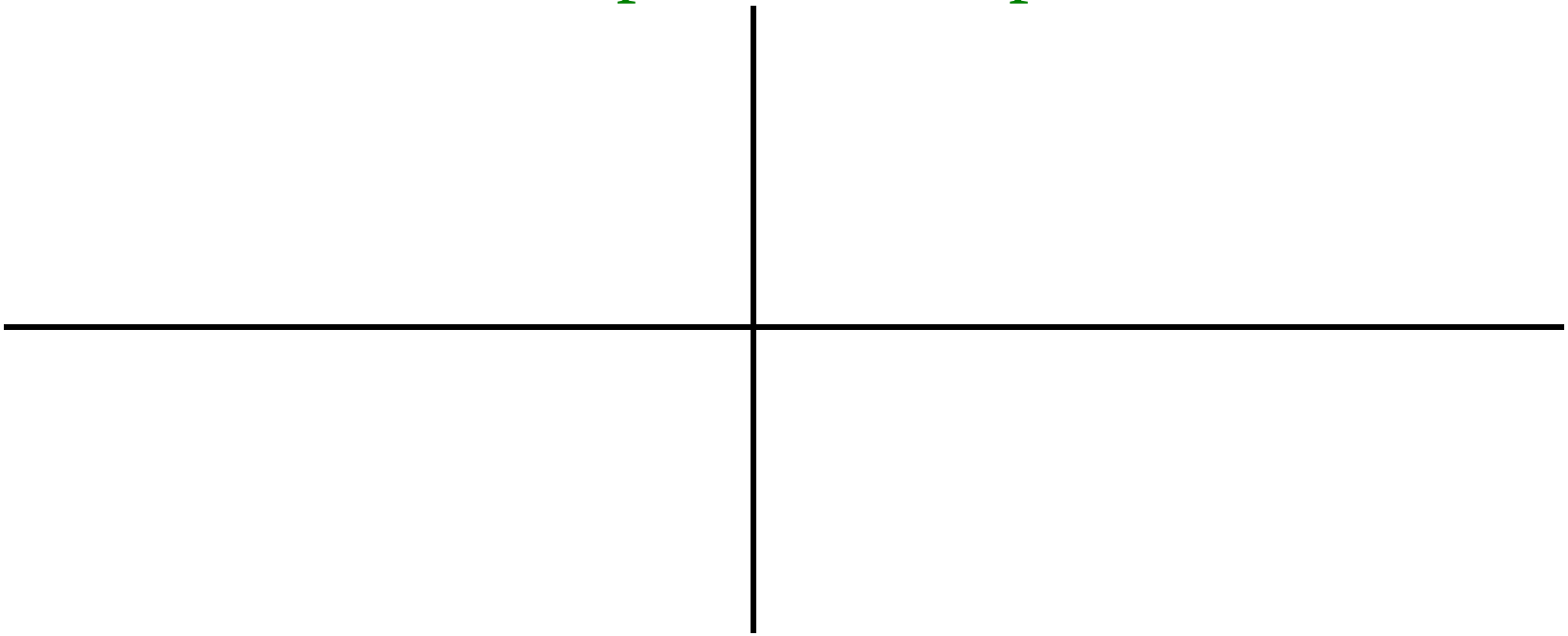
$$\textit{Nitrogen}: 28.9\text{ g N} \times \frac{1\text{ mol N}}{14.01\text{ g N}} = 2.06\text{ mol N}$$

$$\textit{Oxygen}: 16.5\text{g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 1.03\text{ mol O}$$

Ex1) Empirical Formula (cont.)

Step 2. Divide the number of moles of each element by the smallest value for moles calculated in the previous step.



Ex1) Empirical Formula (cont.)

Step 2. Divide the number of moles of each element by the smallest value for moles calculated in the previous step.

$$Carbon: \frac{4.12\ \text{mol}}{1.03\ \text{mol}} = 4$$

Ex1) Empirical Formula (cont.)

Step 2. Divide the number of moles of each element by the smallest value for moles calculated in the previous step.

$$Carbon: \frac{4.12\ \text{mol}}{1.03\ \text{mol}} = 4$$

$$Hydrogen: \frac{5.0\ \text{mol}}{1.03\ \text{mol}} \approx 5$$

Ex1) Empirical Formula (cont.)

Step 2. Divide the number of moles of each element by the smallest value for moles calculated in the previous step.

$$Carbon: \frac{4.12\ \text{mol}}{1.03\ \text{mol}} = 4$$

$$Hydrogen: \frac{5.0\ \text{mol}}{1.03\ \text{mol}} \approx 5$$

$$Nitrogen: \frac{2.06\ \text{mol}}{1.03\ \text{mol}} = 2$$

Ex1) Empirical Formula (cont.)

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$$Hydrogen: \frac{5.0\ \text{mol}}{1.03\ \text{mol}} \approx 5$$

$$Nitrogen: \frac{2.06\ \text{mol}}{1.03\ \text{mol}} = 2$$

$$Oxygen: \frac{1.03\ \text{mol}}{1.03\ \text{mol}} = 1$$

Empirical Formula = $C_4H_5N_2O$

Ex2) Molecular Formula

Ex2) The molar mass of caffeine is 194.2 g/mol. Find the molecular formula for caffeine.

$$\text{Empirical Formula}_{\text{caffeine}} = C_4H_5N_2O$$

$$\frac{\text{Molar Mass}}{\text{Empirical Formula Mass}} = \text{Multiplier for Empirical Formula}$$

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Ex2) The molar mass of caffeine is 194.2 g/mol. Find the molecular formula for caffeine.

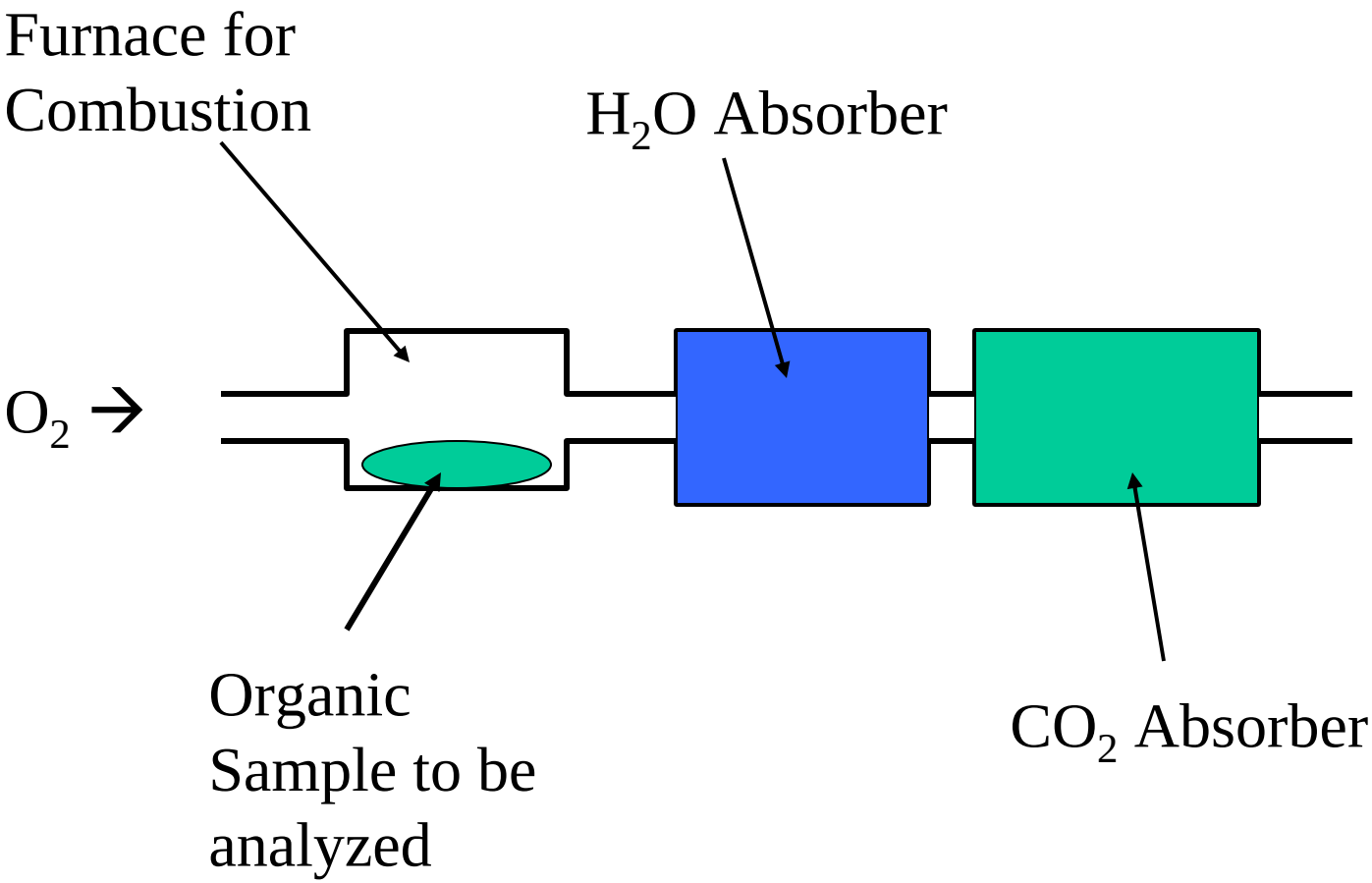
$$\text{Empirical Formula}_{\text{caffeine}} = C_4H_5N_2O$$

$$\frac{\text{Molar Mass}}{\text{Empirical Formula Mass}} = \text{Multiplier for Empirical Formula}$$

$$\frac{194.2\text{g}}{4(12.01\text{g})+5(1.01\text{g})+2(14.01\text{g})+1(16.00\text{g})} = 2$$

$\text{Molecular Formula}_{\text{caffeine}} = C_8H_{10}N_4O_2$

Combustion Analysis



Ex3) Combustion Analysis

Ex3) A 2.04 g sample containing C, H, and O underwent combustion analysis.

4.49 g of CO_2 and 2.45 g of H_2O were produced.

Find the Empirical Formula.

Your Thinking…

1) Use stoichiometry to find the grams of Carbon and Hydrogen produced.

   Subtract the total mass from the sum of the masses of carbon and hydrogen to find the mass of oxygen.

2) Find moles of carbon, hydrogen, and oxygen

3) Find the empirical formula

Step 1. Find the masses of C, H, and O.

Find mass of Carbon

Find mass of Hydrogen

Find mass of Oxygen

Step 1. Find the masses of C, H, and O.

Find mass of Carbon

$4.49\ g\ CO_2 \times \frac{\qquad}{\qquad} \times \frac{\qquad}{\qquad} \times \frac{\qquad}{\qquad} =$

Find mass of Hydrogen

Find mass of Oxygen

Step 1. Find the masses of C, H, and O.

Find mass of Carbon

$$4.49\ g\ CO_2 \times \frac{1\ mol\ CO_2}{44.0\ g\ CO_2} \times \text{——————} \times \text{——————} =$$

Find mass of Hydrogen

Find mass of Oxygen

Step 1. Find the masses of C, H, and O.

Find mass of Carbon

$$4.49\ \text{g}\ CO_2 \times \frac{1\ \text{mol}\ CO_2}{44.0\ \text{g}\ CO_2} \times \frac{1\ \text{mol C}}{1\ \text{mol}\ CO_2} \times \frac{\qquad}{\qquad} =$$

Find mass of Hydrogen

Find mass of Oxygen

Step 1. Find the masses of C, H, and O.

Find mass of Carbon

$$4.49\ g\ CO_2 \times \frac{1\ mol\ CO_2}{44.0\ g\ CO_2} \times \frac{1\ mol\ C}{1\ mol\ CO_2} \times \frac{12.0\ g\ C}{1\ mol\ C} = 1.22\ g\ C$$

Find mass of Hydrogen

$$2.45\ g\ H_2O \times \_\_\_\_\_ \times \_\_\_\_\_ \times \_\_\_\_\_ =$$

Find mass of Oxygen

Step 1. Find the masses of C, H, and O.

Find mass of Carbon

$$4.49\ g\ CO_2 \times \frac{1\ mol\ CO_2}{44.0\ g\ CO_2} \times \frac{1\ mol\ C}{1\ mol\ CO_2} \times \frac{12.0\ g\ C}{1\ mol\ C} = 1.22\ g\ C$$

Find mass of Hydrogen

$$2.45\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.0\ g\ H_2O} \times \frac{\quad}{\quad} \times \frac{\quad}{\quad} =$$

Find mass of Oxygen

Step 1. Find the masses of C, H, and O.

Find mass of Carbon

$$4.49\ g\ CO_2 \times \frac{1\ mol\ CO_2}{44.0\ g\ CO_2} \times \frac{1\ mol\ C}{1\ mol\ CO_2} \times \frac{12.0\ g\ C}{1\ mol\ C} = 1.22\ g\ C$$

Find mass of Hydrogen

$$2.45\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.0\ g\ H_2O} \times \frac{2\ mol\ H}{1\ mol\ H_2O} \times \frac{\qquad}{\qquad} =$$

Find mass of Oxygen

Step 1. Find the masses of C, H, and O.

Find mass of Carbon

$$4.49\ g\ CO_2 \times \frac{1\ mol\ CO_2}{44.0\ g\ CO_2} \times \frac{1\ mol\ C}{1\ mol\ CO_2} \times \frac{12.0\ g\ C}{1\ mol\ C} = 1.22\ g\ C$$

Find mass of Hydrogen

$$2.45\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.0\ g\ H_2O} \times \frac{2\ mol\ H}{1\ mol\ H_2O} \times \frac{1.0\ g\ H}{1\ mol\ H} = 0.27\ g\ H$$

Find mass of Oxygen

Step 1. Find the masses of C, H, and O.

Find mass of Carbon

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Find mass of Oxygen

O_{mass} = mass total – (C_{mass} + H_{mass}) = 2.04g – (1.22g C– 0.27g H)

O_{mass} = 0.55 g O

Ex3) Combustion Analysis (cont.)

Step 2. Find moles of carbon, hydrogen, and oxygen

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$$\textit{Carbon}: 1.22\text{g C} \times \frac{1\text{ mol C}}{12.01\text{g C}} = 0.10\text{ mol C}$$

Ex3) Combustion Analysis (cont.)

Step 2. Find moles of carbon, hydrogen, and oxygen

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Ex3) Combustion Analysis (cont.)

Step 2. Find moles of carbon, hydrogen, and oxygen

$$\textit{Carbon}: 1.22\text{g C} \times \frac{1\text{ mol C}}{12.01\text{g C}} = 0.10\text{ mol C}$$

$$\textit{Hydrogen}: 0.27\text{g H} \times \frac{1\text{ mole H}}{1.01\text{ g H}} = 0.27\text{ mol H}$$

$$\textit{Oxygen}: 0.55\text{g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 0.034\text{ mol O}$$

Ex3) Combustion Analysis (cont.)

Step 3. Divide the number of moles of each element by the smallest value of moles calculated.

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$$Carbon: \frac{0.10 \text{ mol}}{0.034 \text{ mol}} = 2.94 \approx 3$$

Ex3) Combustion Analysis (cont.)

Step 3. Divide the number of moles of each element by the smallest value of moles calculated.

$$Carbon: \frac{0.10\ \text{mol}}{0.034\ \text{mol}} = 2.94 \approx 3$$

$$Hydrogen: \frac{0.27\ \text{mol}}{0.034\ \text{mol}} = 7.94 \approx 8$$

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$$Oxygen: \frac{0.034\ \text{mol}}{0.034\ \text{mol}} = 1$$

Empirical Formula = C_3H_8O

Ex4) Molecular Formula

Ex4) Further experiments showed that molar mass of the hydrocarbon from example 3 is 60.11 g/mol. What is the molecular formula of this compound?

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$$\frac{60.11\ \text{g}}{3(12.01\text{g})+8(1.01\text{g})+1(16.00\text{g})} = 1$$

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Empirical Formula = Molecular Formula = C_3H_8O