Review

Avogadro's # = $N_A = 6.022 \times 10^{23}$ particles/mol = 1 mole

Ex. In 1 mol of CH_4, there are (is)...

... 6.022×10^{23} molecules CH_4

... 1 mole C = $1 \times (6.022 \times 10^{23})$ atoms C = 6.022×10^{23} atoms C

... 4 mole H = $4 \times (6.022 \times 10^{23})$ atoms H = 2.4088×10^{24} atoms H

Ex. In 3.5 mol SF_6, there are (is)...

... $3.5 \times (6.022 \times 10^{23})$ molecules SF_6 = 2.1077×10^{24} molecules SF_6

... 3.5 mole S = $3.5 \times (6.022 \times 10^{23})$ atoms S = 2.1077×10^{24} atoms S

... $3.5 \times 6 = 21$ mol F = $21 \times (6.022 \times 10^{23})$ atoms F = 1.26462×10^{25} atoms F

Ex. In 2 mol of the reaction

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

there are (is)...

$$\ldots\ 2 \text{ mol rxn} \times \frac{1 \text{ mol } C_6H_{12}O_6}{1 \text{ mol rxn}} = 2 \text{ mol } C_6H_{12}O_6$$

$$\ldots\ 2 \text{ mol rxn} \times \frac{6 \text{ mol } O_2}{1 \text{ mol rxn}} = 12 \text{ mol } O_2$$

$$\ldots\ 2 \text{ mol rxn} \times \frac{6 \text{ mol } H_2O}{1 \text{ mol rxn}} = 12 \text{ mol } H_2O$$

$$\ldots\ 2 \text{ mol rxn} * \left[\frac{1 \text{ mol } C_6H_{12}O_6}{1 \text{ mol rxn}} \times \frac{6 \text{ mol O}}{1 \text{ mol } C_6H_{12}O_6} + \frac{6 \text{ mol } O_2}{1 \text{ mol rxn}} * \frac{2 \text{ mol O}}{1 \text{ mol } O_2} + \frac{6 \text{ mol } CO_2}{1 \text{ mol rxn}} \times \frac{2 \text{ mol O}}{1 \text{ mol } CO_2} + \frac{6 \text{ mol } H_2O}{1 \text{ mol rxn}} \times \frac{1 \text{ mol}}{1 \text{ mol}}\right.$$

$$= 2 \text{ mol rxn} * \left[\frac{6 \text{ mol O} + 12 \text{ mol O} + 12 \text{ mol O} + 6 \text{ mol O}}{1 \text{ mol rxn}}\right]$$

$$= 72 \text{ mol O/mol rxn}$$

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Molar Mass = mass of one mole of a subtance; usually has units of g/mol

Ex. molar mass (PCl_5) = 1 × molar mass (P) + 5 × molar mass (Cl)

$= 1(30.974) + 5(35.453)$

$= 208.239$ g/mol

Percent Composition: % of total mass contributed by each element in a molecule

$$\%\ \text{composition}(x) = \frac{\#\ \text{atoms}(x)\ \text{in molecule} \times \text{molar mass}(x)}{\text{molar mass (molecule)}} = \frac{\text{mass}(x)\ \text{in molecule}}{\text{mass (molecule)}}$$

Ex. % composition of CO_2

$$\%(C) = \frac{1\ \text{atom C}}{1\ \text{molecule}\ CO_2} = \frac{1\ \text{mol C}}{1\ \text{mol}\ CO_2} \times \frac{12.011\ \text{g C}}{1\ \text{mol C}} \times \frac{1\ \text{mol}\ CO_2}{44.011\ \text{g}\ CO_2} = \frac{12.011\ \text{g C}}{44.011\ \text{g}\ CO_2} = 0.2729$$

$$\%(O) = \frac{2\ \text{atom O}}{1\ \text{molecule}\ CO_2} = \frac{2\ \text{mol O}}{1\ \text{mol}\ CO_2} \times \frac{16\ \text{g O}}{1\ \text{mol O}} \times \frac{1\ \text{mol}\ CO_2}{44.011\ \text{g}\ CO_2} = \frac{32\ \text{g O}}{44.011\ \text{g}\ CO_2} = 0.7271$$

$$= 1 - \%(C) = 1 - 0.2729 = 0.7271$$

Ex. % composition $CaCO_3$

$$\%(Ca) = \frac{40.08\ \text{g Ca}}{(40.08 + 12.011 + 3\times16)\ \text{g}\ CaCO_3} = \frac{40.08\ \text{g Ca}}{100.091\ \text{g}\ CaCO_3} = 0.4004$$

$$\%(O) = \frac{(3\times16)\ \text{g O}}{100.091\ \text{g}\ CaCO_3} = 0.4796$$

$$\%(C) = 1 - (0.4004 + 0.4796) = 0.12$$

Combustion Reactions: Burning a compound in O_2 to produce CO_2, H_2O, and inorganic products

Ex. $C_6H_{12}O_6 + 6O_2 \to 6CO_2 + 6H_2O$

Ex. $C_4H_{10}PO_2F + O_2 \to CO_2 + H_2O$ + inorganic products

If compound burned only contains C, H, O, we can write and balance a combustion reaction in a general way:

$$C_xH_yO_z + \left(\frac{2x + \frac{y}{2} - z}{2}\right)O_2 \to x\,CO_2 + \frac{y}{2}H_2O$$

Compounds that only contain C, H, O are called hydrocarbons.

Ex. Write and balance the chemical equation for the combustion of ethane, C_2H_6

$$C_2H_6 + O_2 \to CO_2 + H_2O$$

$$C_2H_6 + \frac{7}{2}O_2 \to 2CO_2 + 3H_2O$$

$$2C_2H_6 + 7O_2 \to 4CO_2 + 6H_2O$$

Ex. Write and balance the chemical equation for the combustion of butyric acid, $C_4H_8O_2$

$$C_4H_8O_2 + O_2 \to CO_2 + H_2O$$

$$C_4H_8O_2 + 5O_2 \to 4CO_2 + 4H_2O$$

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Combustion Analysis: determining the % composition of an unknown substance by combusting it (burning it in O_2)

Ex. 5.467 g of unknown compound Z are burned in excess O_2, producing 15.02 g of CO_2 and 2.458 g of H_2O. What is the percent composition of Z? Z contains only C, H, and O (it is a hydrocarbon).

$$Z + O_2 \rightarrow CO_2 + H_2O$$

$$5.467\,g\ Z + O_2 \rightarrow 15.02\,g\ CO_2 + 2.458\,g\ H_2O$$

$$15.02\,g\ CO_2 \times \frac{1\ mol\ CO_2}{44.011\,g\ CO_2} \times \frac{1\ mol\ C}{1\ mol\ CO_2} = \underset{\text{in } Z}{\boxed{0.34128\ mol\ C}} \times \frac{12.011\,g\ C}{1\ mol\ C} = \underset{\text{in } Z}{\boxed{4.0991\,g\ C}}$$

$$2.458\,g\ H_2O \times \frac{1\ mol\ H_2O}{18.0158\,g\ H_2O} \times \frac{2\ mol\ H}{1\ mol\ H_2O} = \underset{\text{in } Z}{\boxed{0.27287\ mol\ H}} \times \frac{1.0079\,g\ H}{1\ mol\ H} = \underset{\text{in } Z}{\boxed{0.2750\,g\ H}}$$

$$\text{mass O in Z} = \underset{\substack{\downarrow \\ \text{total mass} \\ Z}}{5.467\,g} - (\underset{\substack{\downarrow \\ \text{mass C} \\ \text{in Z}}}{4.0991\,g} + \underset{\substack{\downarrow \\ \text{mass H} \\ \text{in Z}}}{0.2750\,g}) = \boxed{1.0929\,g\ O\ \text{in}\ Z}$$

$$1.0929\,g\ O \times \frac{1\ mol\ O}{16\,g\ O} = \underset{\text{in } Z}{\boxed{0.0683\ mol\ O}}$$

% (C) = 4.0991 g / 5.467 g = 0.7498

% (H) = 0.2750 g / 5.467 g = 0.0503

% (O) = 1.0929 g / 5.467 g = 0.1999

-or-

$$C_{0.34128}H_{0.27287}O_{0.0683} \Rightarrow C_{4.997}H_{3.995}O_1 \Rightarrow C_5H_4O$$

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Do we know for sure that the unknown Z has the formula C_5H_4O?

What is the % composition of $C_{10}H_8O_2$?

$$\%(C) = (10 \times 12.011)/(10 \times 12.011 + 8 \times 1.0079 + 2 \times 16) = 0.7455 \approx 0.7498$$

$$\%(H) = (8 \times 1.0079)/161.122069 = 0.05004 \approx 0.0503$$

$$\%(O) = 1 - (0.7455 + 0.05004) = 0.20446 \approx 0.1999$$

remember, we rounded $C_{4.997}H_{3.995}O_1$ to get $C_5H_4O_1$

How do we know whether Z was C_5H_4O or $C_{10}H_8O_2$ or ... $C_{5n}H_{4n}O_n$?

Does it matter?

* CH_2O => formaldehyde => gas, colorless, smelly, highly toxic
* $C_2H_4O_2$ => acetic acid => vinegar, liquid, clear, tasty
* $C_3H_6O_2$ => lactic acid => solid, important in anaerobic exercise
* $C_6H_{12}O_2$ => glucose => Sugar, sweet, safe, solid, white

Empirical formula: the formula of a compound with the smallest whole-number ratio of elements in that compound

~~Exact Formula~~

Molecular Formula: the actual formula of a compound, giving the number of atoms of each element present in one molecule of that compound

Ex.

Molecular Formula		Empirical Formula
CH_2O →	$(CH_2O)_1$ →	CH_2O
$C_2H_4O_2$ →	$(CH_2O)_2$ →	CH_2O
$C_3H_6O_2$ →	$(CH_2O)_3$ →	CH_2O
$C_6H_{12}O_2$ →	$(CH_2O)_6$ →	CH_2O

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Back to our combustion analysis...

$$5.467\,g\ Z + O_2 \rightarrow 15.02\,g\ CO_2 + 2.458\,g\ H_2O$$

$$Z = (C_5H_4O)_n \rightarrow \text{empirical formula or molecular formula?}$$

What is n? To ~~know~~ find n, we must know either

1) the molar mass of Z, or
2) the number of moles of Z combusted (which gives us the molar mass of Z)

1) Assume the molar mass of Z = 241.683 g/mol. What is the molecular formula of Z?

$$\frac{\text{molar mass } Z}{\text{molar mass } C_5H_4O} = \frac{241.683\ g/mol}{80.561\ g/mol} \approx 3 = n$$

The molecular formula of Z is $(C_5H_4O)_3 \Rightarrow C_{15}H_{12}O_3$

2) Assume ~~[illegible]~~ 0.034 mol of Z were combusted in the reaction above. What is the molecular formula of Z?

$$\frac{5.467\,g\ Z}{0.034\,mol\ Z} = 160.794\ g/mol \text{ of } Z = \text{molar mass of } Z$$

$$\frac{\text{molar mass } Z}{\text{molar mass } C_5H_4O} = \frac{160.794\ g/mol}{80.561\ g/mol} \approx 2 = n$$

The molecular formula of Z is $(C_5H_4O)_2 \Rightarrow C_{10}H_8O_2$

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