Combustion Analysis

Combustion Analysis: Technique used to obtain the empirical formula of a hydrocarbon

➔ Remember a standard (unbalanced) combustion reaction? (This formula is unbalanced!)

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

How to Solve a Combustion Analysis Problem

1. Convert mass of CO_2 and mass of H_2O to moles of each compound.
2. Convert moles of CO_2 to moles of Carbon, and moles of H_2O to moles of hydrogen.
3. **If** the compound contains something which is not C or H, find its mass by subtraction, and convert the mass to moles.
4. Now you have mole numbers! Complete the empirical formula calculation (divide by small, multiply til whole).

Practice:

1. Upon combustion, a 0.8233 g sample of a compound containing only carbon, hydrogen, and oxygen produces 2.445 g CO_2 and 0.6003 g H_2O. What is the empirical formula of the compound?

$$CO_2 \rightarrow C: 2.445\,g\,CO_2 \times \frac{1\text{ mol }CO_2}{44.01\,g\,CO_2} \times \frac{1\text{ mol C}}{1\text{ mol }CO_2} \times \frac{12.01\,g\,C}{1\text{ mol}} = 0.6672\,g\,C$$

$$H_2O \rightarrow H: 0.6003\,g\,H_2O \times \frac{1\text{ mol }H_2O}{18.016\,g\,H_2O} \times \frac{2\text{ mol H}}{1\text{ mol }H_2O} \times \frac{1.008\,g\,H}{1\text{ mol H}} = 0.06717\,g\,H$$

$$0.8233\,g - 0.6672 - 0.06717 = 0.0889\,g\,O$$

$$\left.\begin{array}{l} C: 0.6672\,g / 12.01\,g/mol = 0.05555\text{ mol C} \\ H: 0.06717\,g / 1.008\,g/mol = 0.06664\text{ mol H} \\ O: 0.0889\,g / 16.00\,\frac{g}{mol} = 0.00556\text{ mol O} \end{array}\right] \div 0.00556 \quad \begin{array}{l} = 10 \\ = 12 \\ = 1 \end{array}$$

$$\boxed{C_{10}H_{12}O}$$

Practice Makes Perfect!

2. Combustion analysis determined that a compound containing only carbon and hydrogen produces 1.83 g CO_2 and 0.901 g H_2O. Find the empirical formula of the compound.

$$C: 1.83\text{ g } CO_2 \times \frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2} \times \frac{1\text{ mol C}}{1\text{ mol } CO_2} = 0.0416\text{ mol C}$$

$$H: 0.901\text{ g } H_2O \times \frac{1\text{ mol } H_2O}{18.016\text{ g } H_2O} \times \frac{2\text{ mol H}}{1\text{ mol } H_2O} = 0.100\text{ mol H}$$

$\div 0.0416$: $= 1$, $= 2.4$

$\times 2$: $= 2$, ≈ 5

$\boxed{C_2H_5}$

3. When the unbalanced reaction below occurs at STP, 1.5 L of CO_2 and 1.0 L of H_2O are created. What is the empirical formula of the hydrocarbon?

$$C_xH_y(g) + O_2 \rightarrow CO_2(g) + H_2O(g)$$

a. CH_2 b. C_2H_3 c. C_2H_5 d. C_3H_4

$$C: 1.5\text{ L} \times \frac{1\text{ mol } CO_2}{22.4\text{ L}} \times \frac{1\text{ C}}{1\text{ } CO_2} = \frac{1.5}{22.4}\text{ mol C}$$

$$H: 1.0\text{ L} \times \frac{1\text{ mol } H_2O}{22.4\text{ L}} \times \frac{2\text{ H}}{1\text{ } H_2O} = \frac{2}{22.4}\text{ mol H}$$

$\div \frac{1.5}{22.4}$: $= 1$, $= \frac{2}{22.4} \times \frac{22.4}{1.5} = \frac{2}{1.5} = \frac{4}{3}$

$\times 3$: $= 3$, $= 4$

4. Combustion analysis of 0.800 g of an unknown hydrocarbon yields 26 g CO_2 and 7.8 g H_2O. What is the formula of the hydrocarbon?

a. CH_2 b. C_2H_3 c. C_2H_5 d. C_3H_4

$$C: \overset{13}{\cancel{26}}\text{ g } CO_2 \times \frac{1\text{ mol}}{\underset{22}{\cancel{44}}\text{ g}} \times \frac{1\text{ C}}{1\text{ } CO_2} = \frac{13}{22} \approx 0.6$$

$$H: 7.8\text{ g } H_2O \times \frac{1\text{ mol}}{18\text{ g}} \times \frac{2\text{ H}}{1\text{ } H_2O} \approx \frac{16}{18} \approx 0.9$$

$\div 0.6$: $= 1$, $= \underbrace{1.5}_{9/6}$

$\times 2$: $= 2$, $= 3$