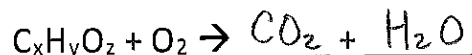


## Combustion Analysis

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

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



### 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:

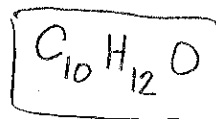
1. Upon combustion, a 0.8233 g sample of a compound containing only carbon, hydrogen, and oxygen produces 2.445 g  $CO_2$  and 0.6003 g  $H_2O$ . What is the empirical formula of the compound?

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

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

$$0.8233 \text{ g} - 0.6672 - 0.06717 = 0.0889 \text{ g } O$$

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



## Practice Makes Perfect!

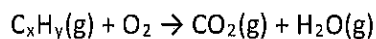
2. Combustion analysis determined that a compound containing only carbon and hydrogen produces 1.83 g CO<sub>2</sub> and 0.901 g H<sub>2</sub>O. Find the empirical formula of the compound.

$$\begin{array}{l}
 \text{C: } 1.83 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.0416 \text{ mol C} \\
 \text{H: } 0.901 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.100 \text{ mol H}
 \end{array}$$

$$\left. \begin{array}{l} = 1 \\ = 2.4 \end{array} \right\} \begin{array}{l} \div 0.0416 \\ \times 2 \end{array} \begin{array}{l} = 2 \\ \approx 5 \end{array}$$

$$\boxed{\text{C}_2\text{H}_5}$$

3. When the unbalanced reaction below occurs at STP, 1.5 L of CO<sub>2</sub> and 1.0 L of H<sub>2</sub>O are created. What is the empirical formula of the hydrocarbon?

a. CH<sub>2</sub>b. C<sub>2</sub>H<sub>3</sub>c. C<sub>2</sub>H<sub>5</sub>(d) C<sub>3</sub>H<sub>4</sub>

$$\begin{array}{l}
 \text{C: } 1.5 \text{ L} \times \frac{1 \text{ mol CO}_2}{22.4 \text{ L}} \times \frac{1 \text{ C}}{1 \text{ CO}_2} = \frac{1.5}{22.4} \text{ mol C} \\
 \text{H: } 1.0 \text{ L} \times \frac{1 \text{ mol H}_2\text{O}}{22.4 \text{ L}} \times \frac{2 \text{ H}}{1 \text{ H}_2\text{O}} = \frac{2}{22.4} \text{ mol H}
 \end{array}$$

$$\left. \begin{array}{l} = 1 \\ = 1.5 \end{array} \right\} \begin{array}{l} \div \frac{1.5}{22.4} \\ \times 3 \end{array} \begin{array}{l} = 3 \\ = 4 \end{array}$$

$$= \frac{2}{22.4} \times \frac{22.4}{1.5} = \frac{2}{1.5} = \frac{4}{3}$$

4. Combustion analysis of 0.800 g of an unknown hydrocarbon yields 26 g CO<sub>2</sub> and 7.8 g H<sub>2</sub>O. What is the formula of the hydrocarbon?

a. CH<sub>2</sub>(b) C<sub>2</sub>H<sub>3</sub>c. C<sub>2</sub>H<sub>5</sub>d. C<sub>3</sub>H<sub>4</sub>

$$\begin{array}{l}
 \text{C: } \overset{13}{26} \text{ g CO}_2 \times \frac{1 \text{ mol}}{44 \text{ g}} \times \frac{1 \text{ C}}{1 \text{ CO}_2} = \frac{13}{22} \approx 0.6 \\
 \text{H: } 7.8 \text{ g H}_2\text{O} \times \frac{1 \text{ mol}}{18 \text{ g}} \times \frac{2 \text{ H}}{1 \text{ H}_2\text{O}} \approx \frac{16}{18} \approx 0.9
 \end{array}$$

$$\left. \begin{array}{l} = 1 \\ = 1.5 \end{array} \right\} \begin{array}{l} \div 0.6 \\ \times 2 \end{array} \begin{array}{l} = 2 \\ = 3 \end{array}$$

$$= \frac{1.5}{9/6}$$