

9-2-05

CHM 123 - Lecture (Friday 10:30 a)

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Limiting Reactants

The reactant that is completely used in a chemical reaction is called the limiting reactant.

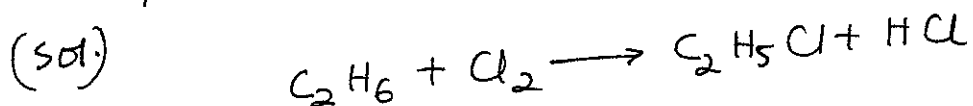
Percent Yield:

- Amount of product calculated to form when all the limiting reactant is consumed is the theoretical yield.

- Actual yield: amount actually obtained from the reaction.

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

(Q) Calculate % yield of $\text{C}_2\text{H}_5\text{Cl}$ if the reaction of 125g of C_2H_6 with 255g of Cl_2 produced 206g of $\text{C}_2\text{H}_5\text{Cl}$?



Determining limiting reactant:

$$\text{Mass of } \text{C}_2\text{H}_5\text{Cl} \left\{ \begin{array}{l} \text{from } \text{C}_2\text{H}_6 \end{array} \right\} = 125\text{g} \left(\frac{1 \text{ mol } \text{C}_2\text{H}_6}{30\text{g}} \right) \left(\frac{1 \text{ mol } \text{C}_2\text{H}_5\text{Cl}}{1 \text{ mol } \text{C}_2\text{H}_6} \right) \left(\frac{64.5\text{g}}{1 \text{ mol } \text{C}_2\text{H}_5\text{Cl}} \right) = 258\text{g}.$$

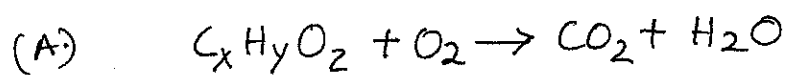
Similarly do the calculation for $\text{Cl}_2 \Rightarrow$ results in Cl_2 being the limiting reactant:

$$\text{Mass of } \text{C}_2\text{H}_5\text{Cl} \left\{ \begin{array}{l} \text{from } \text{Cl}_2 \end{array} \right\} \Rightarrow \underline{232\text{g}} \Rightarrow \text{theoretical yield}$$

$$\% \text{ yield of } \text{C}_2\text{H}_5\text{Cl} = \frac{\text{Actual yield}}{\text{theoretical yield}} \times 100 = \frac{206\text{g}}{232\text{g}} \times 100 = \text{---}\%$$

Combustion Analysis

(Q) Propionic acid contains C, H and O. If 0.206g of the acid burns completely in O_2 and gives 0.421g of CO_2 and 0.172g of H_2O , what is the empirical formula of the acid?



Determining masses of C and H in original sample:

$$\text{Mass of C} = \underset{CO_2}{0.421g} \left(\frac{1 \text{ mol } CO_2}{44.01g} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol } CO_2} \right) \left(\frac{12.01g}{1 \text{ mol C}} \right) = 0.115g \text{ of C}$$

$$\text{Mass of H} = \underset{H_2O}{0.172g} \left(\frac{1 \text{ mol } H_2O}{18.02g} \right) \left(\frac{2 \text{ mol H}}{1 \text{ mol } H_2O} \right) \left(\frac{1.00794g}{1 \text{ mol H}} \right) = 0.02g \text{ of H}$$

The carbon and hydrogen have to come only from the hydrocarbon (acid). Using the masses of C and H and the total amount of acid (0.206g), the mass of oxygen can be calculated. Once all 3 masses are obtained, use method given in previous notes to obtain empirical formula.

If absent from lecture today,
Handout given in class (Stoichiometry worksheet) can be picked up in BRWN 1144. Problem 3 $\rightarrow$ relevant to Quiz next week.
