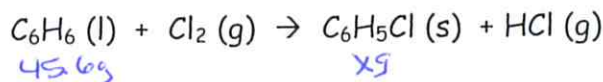


1. Chlorobenzene, C_6H_5Cl , is used in the production of chemicals such as aspirin and dyes. One way that chlorobenzene is prepared is by reacting benzene, C_6H_6 , with chlorine gas according to the following BALANCED equation.



$$C_6H_6: 6(12.01) + 6(1) = 78.10 \text{ g/mol}$$

$$C_6H_5Cl: 6(12.01) + 5(1.008) + 35.45 = 112.55 \text{ g/mol}$$

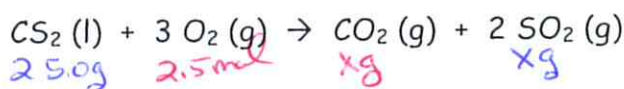
- a. What is the theoretical yield if 45.6 g of benzene react?

$$\frac{45.6 \text{ g } C_6H_6}{78.10 \text{ g } C_6H_6} \times \frac{1 \text{ mol } C_6H_6}{1 \text{ mol } C_6H_6} \times \frac{1 \text{ mol } C_6H_5Cl}{1 \text{ mol } C_6H_6} \times \frac{112.55 \text{ g } C_6H_5Cl}{1 \text{ mol } C_6H_5Cl} = 65.7 \text{ g } C_6H_5Cl$$

- b. If the actual yield is 63.7 g of chlorobenzene, calculate the percent yield.

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100 = \frac{63.7}{65.7} \times 100 = 97\%$$

2. When carbon disulfide burns in the presence of oxygen, sulfur dioxide and carbon dioxide are produced according to the following equation.



$$CS_2: 1(12.01) + 2(32.07) = 76.15 \text{ g/mol}$$

$$SO_2: 1(32.07) + 2(16) = 64.07 \text{ g/mol}$$

- a. What is the percent yield of sulfur dioxide if the burning of 25.0 g of carbon disulfide produces 40.5 g of sulfur dioxide?

$$\frac{25.0 \text{ g } CS_2}{76.15 \text{ g } CS_2} \times \frac{1 \text{ mol } CS_2}{1 \text{ mol } CS_2} \times \frac{2 \text{ mol } SO_2}{1 \text{ mol } CS_2} \times \frac{64.07 \text{ g } SO_2}{1 \text{ mol } SO_2} = 42.1 \text{ g } SO_2$$

$$\% \text{ yield} = \frac{40.5}{42.1} \times 100 = 96\%$$

- b. What is the percent yield of carbon dioxide if 2.5 mol of oxygen react and 32.4 g of carbon dioxide are produced?

$$\frac{2.5 \text{ mol } O_2}{3 \text{ mol } O_2} \times \frac{1 \text{ mol } CO_2}{1 \text{ mol } O_2} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 36.7 \text{ g } CO_2$$

$$CO_2: 1(12.01) + 2(16) = 44.01 \text{ g/mol}$$

$$\% \text{ yield} = \frac{32.4 \text{ g}}{36.7 \text{ g}} \times 100 = 88\%$$