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Complete stoichiometric calculations between reactants and products, given a balanced chemical equation.

1. Answer the following questions based on the balanced chemical equation written below:



a. If 0.38 moles of oxygen gas are consumed, then how many moles of carbon dioxide are produced?

$$(0.38 \text{ mol } \cancel{\text{O}_2}) \left(\frac{16 \text{ mol CO}_2}{25 \text{ mol } \cancel{\text{O}_2}} \right) = \boxed{0.24 \text{ mol CO}_2}$$

b. How many moles of fuel (C_8H_{18}) are required to produce 7.4 moles of water vapor?

$$(7.4 \text{ mol } \cancel{\text{H}_2\text{O}}) \left(\frac{2 \text{ mol } \text{C}_8\text{H}_{18}}{18 \text{ mol } \cancel{\text{H}_2\text{O}}} \right) = \boxed{0.82 \text{ mol } \text{C}_8\text{H}_{18}}$$

c. What is the mole ratio between carbon dioxide and fuel (C_8H_{18})?

$$\frac{16 \text{ mol CO}_2}{2 \text{ mol } \text{C}_8\text{H}_{18}}$$

d. If 27.9 g of fuel (C_8H_{18}) is burned, then how many liters of CO_2 are produced after cooling to STP?

$$(27.9 \text{ g } \text{C}_8\text{H}_{18}) \left(\frac{1 \text{ mol } \text{C}_8\text{H}_{18}}{114.2 \text{ g}} \right) \left(\frac{16 \text{ mol } \text{CO}_2}{2 \text{ mol } \text{C}_8\text{H}_{18}} \right) \left(\frac{22.4 \text{ L}}{1 \text{ mol } \text{CO}_2} \right) = \boxed{43.8 \text{ L CO}_2}$$

e. How many moles of water vapor are produced if 7.14×10^{25} molecules of oxygen gas are consumed?

$$(7.14 \times 10^{25} \text{ molecules } \cancel{\text{O}_2}) \left(\frac{1 \text{ mol } \cancel{\text{O}_2}}{6.02 \times 10^{23} \text{ molecules } \cancel{\text{O}_2}} \right) \left(\frac{18 \text{ mol } \text{H}_2\text{O}}{25 \text{ mol } \cancel{\text{O}_2}} \right) = \boxed{85.4 \text{ mol H}_2\text{O}}$$

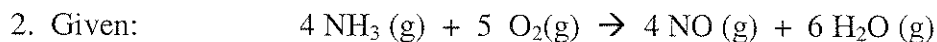
f. How many liters of oxygen (at STP) are required to react with 350 g of fuel (C_8H_{18})?

$$(350 \text{ g } \text{C}_8\text{H}_{18}) \left(\frac{1 \text{ mol } \text{C}_8\text{H}_{18}}{114.2 \text{ g}} \right) \left(\frac{25 \text{ mol } \text{O}_2}{2 \text{ mol } \text{C}_8\text{H}_{18}} \right) \left(\frac{22.4 \text{ L}}{1 \text{ mol}} \right) = \boxed{860 \text{ L O}_2}$$

g. How many molecules of carbon dioxide are produced when 2.50 L of water vapor are produced at STP?

$$(2.50 \text{ L } \text{H}_2\text{O}) \left(\frac{1 \text{ mol } \text{H}_2\text{O}}{22.4 \text{ L}} \right) \left(\frac{16 \text{ mol } \text{CO}_2}{18 \text{ mol } \text{H}_2\text{O}} \right) \left(\frac{6.02 \times 10^{23} \text{ molecules } \text{CO}_2}{1 \text{ mol } \text{CO}_2} \right) = \boxed{5.97 \times 10^{22} \text{ molecules CO}_2}$$

Determine the limiting and excess reagent from a balanced chemical reaction



- a. Determine the limiting and excess reagent when 3.5 moles of NH_3 are combined with 4.7 moles of oxygen gas.

$$\left. \begin{aligned} (3.5 \text{ mol } \cancel{\text{NH}_3}) \left(\frac{4 \text{ mol NO}}{4 \text{ mol } \cancel{\text{NH}_3}} \right) &= 3.5 \text{ mol NO} \\ (4.7 \text{ mol } \cancel{\text{O}_2}) \left(\frac{4 \text{ mol NO}}{5 \text{ mol } \cancel{\text{O}_2}} \right) &= 3.8 \text{ mol NO} \end{aligned} \right\} \begin{array}{l} 3.5 \text{ mol} < 3.8 \text{ mol} \\ \downarrow \\ \text{NH}_3 \text{ is LRA} \end{array}$$

- b. Based on your answer to (a), how many grams of NO can be produced when combining 3.5 moles of NH_3 with 4.7 moles of oxygen gas.

$$(3.5 \text{ mol } \cancel{\text{NO}}) \left(\frac{30.01 \text{ g NO}}{1 \text{ mol } \cancel{\text{NO}}} \right) = \boxed{110 \text{ g NO}}$$

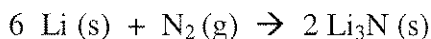
- c. Based on your answer to (a) how much excess reagent is leftover when the reaction runs to completion:

$$(3.5 \text{ mol } \cancel{\text{NH}_3}) \left(\frac{5 \text{ mol O}_2}{4 \text{ mol } \cancel{\text{NH}_3}} \right) = 4.4 \text{ mol O}_2 \text{ (consumed)}$$

$$4.7 \text{ mol O}_2 - 4.4 \text{ mol O}_2 = \boxed{.3 \text{ mol O}_2 \text{ leftover}}$$

[Skill 6.c] Determine the percent yield of a product given the experimental yield.

3. Lithium and nitrogen react to produce lithium nitride:



- a. Determine the theoretical yield of lithium nitride when 5.87 g of lithium is entirely consumed with excess nitrogen gas.

$$(5.87 \text{ g Li}) \left(\frac{1 \text{ mol Li}}{6.941 \text{ g}} \right) \left(\frac{2 \text{ mol Li}_3\text{N}}{6 \text{ mol Li}} \right) \left(\frac{34.83 \text{ g Li}_3\text{N}}{1 \text{ mol Li}_3\text{N}} \right) = \boxed{9.82 \text{ g Li}_3\text{N}}$$

- b. You run this reaction in the lab, reacting 5.87 g of lithium with excess nitrogen, and you recover an actual yield of 8.17 g lithium nitride. Determine the percent yield.

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

$$\% \text{ yield} = \frac{8.17 \text{ g}}{9.82 \text{ g}} \times 100\% = \boxed{83.2\%}$$