REVIEW QUESTIONS
Chapter 8

1. Given the unbalanced equation shown below:

$$\mathbf{2}\ NaClO_3\ (s) \rightarrow \mathbf{2}\ NaCl\ (s) + \mathbf{3}\ O_2\ (g)$$

a) How many grams of O_2 can be produced from reaction of 12.0 moles of $NaClO_3$?

$$12.0 \text{ mol } NaClO_3 \times \frac{3 \text{ mol } O_2}{2 \text{ mol } NaClO_3} \times \frac{32.00 \text{ g}}{1 \text{ mol}} = 576 \text{ g } O_2$$

b) How many grams of NaCl are produced when 80.0 g of O_2 are produced?

$$80.0 \text{ g } O_2 \times \frac{1 \text{ mol}}{32.00 \text{ g}} \times \frac{2 \text{ mol NaCl}}{3 \text{ mol } O_2} \times \frac{58.45 \text{ g}}{1 \text{ mol}} = 97.4 \text{ g NaCl}$$

2. Given the balanced equations below, how many moles of each reactant must react to produce 13.70 mole of N_2?

$$4\ NH_3\ (g) + 6\ NO\ (g) \rightarrow 5\ N_2\ (g) + 6\ H_2O\ (g)$$

$$13.70 \text{ mol } N_2 \times \frac{4 \text{ mol } NH_3}{5 \text{ mol } N_2} = 10.96 \text{ mol } NH_3$$

$$13.70 \text{ mol } N_2 \times \frac{6 \text{ mol NO}}{5 \text{ mol } N_2} = 16.44 \text{ mol NO}$$

3. What mass of CO_2 can be produced from the reaction of 25.0 g of C_3H_8 with 75.0 g of O_2 according to the following equation:

$$C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O$$

$$\mathbf{25.0\ g\ C_3H_8 \times \frac{1\ mol}{44.11\ g} \times \frac{3\ mol\ CO_2}{1\ mol\ C_3H_8} = 1.70\ mol\ CO_2}$$

$$\mathbf{75.0\ g\ O_2 \times \frac{1\ mol}{32.00\ g} \times \frac{3\ mol\ CO_2}{5\ mol\ O_2} = 1.4\underline{0}6\ mol\ CO_2} \quad \longleftarrow \textbf{Limiting Reactant}$$

$$\mathbf{1.4\underline{0}6\ mol\ CO_2 \times \frac{44.01\ g}{1\ mol} = 61.9\ g\ CO_2}$$

4. How many grams of SO_2 are produced when 152 g of CS_2 react with 48.0 g of O_2 according to the following equation:

$$CS_2 + 3\,O_2 \rightarrow CO_2 + 2\,SO_2$$

$$\mathbf{152\ g\ CS_2 \times \frac{1\ mol}{76.13\ g} \times \frac{2\ mol\ SO_2}{1\ mol\ CS_2} = 3.99\ mol\ SO_2}$$

$$\mathbf{48.0\ g\ O_2 \times \frac{1\ mol}{32.00\ g} \times \frac{2\ mol\ SO_2}{3\ mol\ O_2} = 1.00\ mol\ SO_2} \quad \longleftarrow \textbf{Limiting Reactant}$$

$$\mathbf{1.00\ mol\ CO_2 \times \frac{64.06\ g}{1\ mol} = 64.1\ g\ SO_2}$$

5. When 50.0 g of $MgCO_3$ react completely with H_3PO_4, 15.8 g of CO_2 are produced. What is the percent yield for this reaction? The unbalanced equation is given below:

$$H_3PO_4 + MgCO_3 \rightarrow Mg_3(PO_4)_2 + CO_2 + H_2O$$

$$\mathbf{50.0\ g\ MgCO_3 \times \frac{1\ mol}{84.31\ g} \times \frac{3\ mol\ CO_2}{3\ mol\ MgCO_3} \times \frac{44.01\ g}{1\ mol} = 26.1\ g\ CO_2}$$

$$\mathbf{\%\ Yield = \frac{15.8\ g}{26.1\ g} \times 100 = 60.5\%}$$

6. The reaction of calcium and nitrogen gas (shown below) can be carried out with 90.0% yield. In an experiment, 56.6 g of calcium are reacted with 30.5 g of nitrogen gas. What mass of calcium nitride is formed in this experiment?

$$3\ Ca\ (s) + N_2\ (g) \rightarrow Ca_3N_2\ (s)$$

$$\mathbf{56.6\ g\ Ca \times \frac{1\ mol}{40.08\ g} \times \frac{1\ mol\ Ca_3N_2}{3\ mol\ Ca} \times \frac{148.26\ g}{1\ mol} = 69.8\ g\ Ca_3N_2}$$ ⟵ **Limiting Reactant**

$$\mathbf{30.5\ g\ N_2 \times \frac{1\ mol}{28.02\ g} \times \frac{1\ mol\ Ca_3N_2}{1\ mol\ N_2} \times \frac{148.26\ g}{1\ mol} = 161\ g\ Ca_3N_2}$$

$$\mathbf{69.8\ g\ Ca_3N_2 \times \frac{90.0\ g}{100\ g} = 62.8\ g\ Ca_3N_2}$$

7. Hydrogen gas has been suggested as a clean fuel because it produces only water when it burns. If the reaction has a 98% yield, what mass of hydrogen gas would form 75.0 kg of water?

$$\mathbf{2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)}$$

$$\mathbf{Theoretical\ yield = \frac{Actual\ yield}{\%\ yield} = \frac{75.0\ kg}{0.98} = 76.5\ kg\ H_2O}$$

$$\mathbf{76.5\ kg\ H_2O \times \frac{1\ mol}{18.02\ g} \times \frac{2\ mol\ H_2}{2\ mol\ H_2O} \times \frac{2.02\ g}{1\ mol} = 8.58\ kg\ H_2}$$

8. Diborane (B_2H_6) can be produced by the unbalanced reaction shown below. If the reaction has an 85.0% yield, how many grams of $NaBH_4$ are needed to produce 20.0 g of diborane?

$$\mathbf{3}\ NaBH_4\ (s) + \mathbf{4}\ BF_3\ (g) \rightarrow \mathbf{2}\ B_2H_6\ (g) + \mathbf{3}\ NaBF_4\ (s)$$

$$\mathbf{Theoretical\ yield = \frac{Actual\ yield}{\%\ yield} = \frac{20.0\ g}{0.850} = 23.5\ g\ B_2H_6}$$

$$\mathbf{23.5\ g\ B_2H_6 \times \frac{1\ mol}{27.68\ g} \times \frac{3\ mol\ NaBH_4}{2\ mol\ B_2H_6} \times \frac{37.84\ g}{1\ mol} = 48.2\ g\ NaBH_4}$$

9. Consider the reaction shown below:

$$B_2H_6\,(g) + 6\,Cl_2\,(g) \rightarrow 2\,BCl_3\,(g) + 6\,HCl\,(g) \qquad \Delta H = -755\text{ kJ}$$

a) Is this reaction endothermic or exothermic? **Exothermic**

b) How much heat is released when 85.0 g of B_2H_6 react?

$$\mathbf{85.0\text{ g }B_2H_6 \times \frac{1\text{ mol }O_2}{27.68\text{ g}} \times \frac{755\text{ kJ}}{1\text{ mol }B_2H_6} = 2320\text{ kJ}} \quad \textbf{(3 sig figs)}$$

10. Thermal decomposition of mercury (II) oxide occurs as shown below:

$$2\,HgO\,(s) \rightarrow 2\,Hg\,(l) + O_2\,(g) \qquad \Delta H = 181.6\text{ kJ}$$

a) How much heat is needed to decompose 555 g of HgO?

$$\mathbf{555\text{ g HgO} \times \frac{1\text{ mol}}{216.59\text{ g}} \times \frac{181.6\text{ kJ}}{2\text{ mol HgO}} = 233\text{ kJ}} \quad \textbf{(3 sig figs)}$$

b) If 275 kJ of heat is absorbed, how many grams of mercury form?

$$\mathbf{275\text{ kJ} \times \frac{2\text{ mol Hg}}{181.6\text{ kJ}} \times \frac{200.59\text{ g}}{1\text{ mol}} = 608\text{ g Hg}} \quad \textbf{(3 sig figs)}$$

11. Iron (II) sulfide reacts with HCl according to the reaction:

$$FeS\ (s) + 2\ HCl\ (aq) \rightarrow FeCl_2\ (aq) + H_2S\ (g)$$

A reaction mixture initially contains 20.0 g of FeS and 23.8 g of HCl. When the reaction has occurred as completely as possible, what mass of which reactant is left?

$$\textbf{20.0 g FeS} \times \frac{\textbf{1 mol}}{\textbf{87.91 g}} \times \frac{\textbf{2 mol HCl}}{\textbf{1 mol FeS}} \times \frac{\textbf{36.46 g}}{\textbf{1 mol}} = \textbf{16.6 g HCl}$$

$$\textbf{23.8 g HCl} \times \frac{\textbf{1 mol}}{\textbf{36.46 g}} \times \frac{\textbf{1 mol FeS}}{\textbf{2 mol HCl}} \times \frac{\textbf{87.91 g}}{\textbf{1 mol}} = \textbf{28.7 g FeS}$$

Calculations above show that FeS is limiting reactant since 28.7 g FeS is required to react with all of the HCl (23.8 g) and only 20.0 g is available. Therefore 7.2 g of HCl is in excess after the reaction is complete (23.8 g–16.6 g).

12. People often use sodium bicarbonate as an antacid to neutralize excess hydrochloric acid in an upset stomach. (a) Write a balanced equation for the reaction of aqueous sodium bicarbonate and aqueous hydrochloric acid; (b) What mass of hydrochloric acid (in grams) can 2.5 g of sodium bicarbonate neutralize?

$$\mathbf{NaHCO_3\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + CO_2\ (g) + H_2O\ (l)}$$

$$\textbf{2.5 g } \mathbf{NaHCO_3} \times \frac{\textbf{1 mol}}{\textbf{84.01 g}} \times \frac{\textbf{1 mol HCl}}{\textbf{1 mol } \mathbf{NaHCO_3}} \times \frac{\textbf{36.46 g}}{\textbf{1 mol}} = \textbf{1.1 g HCl}$$

13. A solution contains an unknown mass of dissolved barium ions. When sodium sulfate is added to the solution, a white precipitate forms. The precipitate is filtered, dried and found to have a mass of 258 mg. Based on this information, what mass of barium was in the original sample? (Assume that all of the barium was precipitated out of the solution by the reaction.)

$$\mathbf{Ba^{2+} + Na_2SO_4\ (aq) \rightarrow BaSO_4\ (s) + 2\ Na^{+}}$$

$$\textbf{258 mg } \mathbf{BaSO_4} \times \frac{\textbf{1 g}}{\mathbf{10^3}\textbf{ mg}} \times \frac{\textbf{1 mol}}{\textbf{233.39 g}} \times \frac{\textbf{1 mol } \mathbf{Ba^{2+}}}{\textbf{1 mol } \mathbf{BaSO_4}} \times \frac{\textbf{137.33 g}}{\textbf{1 mol}} = \textbf{0.152 g}$$

14. Consider the reaction shown below:

$$2\ N_2H_4\ (g) + N_2O_4\ (g) \rightarrow 3\ N_2\ (g) + 4\ H_2O\ (g)$$

A reaction flask initially contains 27.5 g of N_2H_4 and 74.9 g of N_2O_4. Determine the identity and mass of all the substances present in the reaction flask after all the reactants have reacted as much as possible. (Assume 100% yield.)

$$\mathbf{27.5\ g\ N_2H_4 \times \frac{1\ mol}{32.06\ g} \times \frac{1\ mol\ N_2O_4}{2\ mol\ N_2H_4} \times \frac{92.02\ g}{1\ mol} = 39.5\ g\ N_2O_4\ reacted}$$

Since amount of N_2O_4 reacted is less than amount available (74.9 g), it is in excess.

Therefore, at end of reaction:

$$\mathbf{excess\ N_2O_4 = 74.9\ g - 39.5\ g = 35.4\ g}$$

$$\mathbf{27.5\ g\ N_2H_4 \times \frac{1\ mol}{32.06\ g} \times \frac{3\ mol\ N_2}{2\ mol\ N_2H_4} \times \frac{28.02\ g}{1\ mol} = 36.1\ g\ N_2}$$

$$\mathbf{27.5\ g\ N_2H_4 \times \frac{1\ mol}{32.06\ g} \times \frac{4\ mol\ H_2O}{2\ mol\ N_2H_4} \times \frac{18.02\ g}{1\ mol} = 30.9\ g\ H_2O}$$

15. Magnesium ions can be precipitated from seawater by addition of sodium hydroxide. How many grams of NaOH must be added to a sample of seawater to completely precipitate 88.4 mg of magnesium present? (Write a balanced equation for the reaction first.)

$$\mathbf{Mg^{2} + 2\ NaOH\ (aq) \rightarrow Mg(OH)_2\ (s) + 2\ Na^{+}}$$

$$\mathbf{88.4\ mg\ Mg^{2+} \times \frac{1\ g}{10^3\ mg} \times \frac{1\ mol}{24.31\ g} \times \frac{2\ mol\ NaOH}{1\ mol\ Mg^{2+}} \times \frac{40.00\ g}{1\ mol} = 0.291\ g\ NaOH}$$

16. Pyrite (FeS_2), an impurity in some coals, reacts with oxygen to form the air pollutant sulfur dioxide, as shown below. What mass of SO_2 (in grams) is produced when 2.0×10^4 kg of coal containing 0.050 mass % pyrite is burned?

$$2\ FeS_2\ (s) + SO_2\ (g) \rightarrow 4\ SO_2\ (g) + 2\ FeO\ (s)$$

$$\mathbf{2.0 \times 10^4\ kg\ coal \times \frac{10^3\ g}{1\ kg} \times \frac{0.050\ g\ FeS_2}{100\ g\ coal} \frac{1\ mol}{119.97\ g} \times \frac{4\ mol\ SO_2}{2\ mol\ FeS_2} \times \frac{64.06\ g}{1\ mol} = 1.1 \times 10^4\ g\ SO_2}$$