

REVIEW QUESTIONS

## Chapter 8

1. Given the unbalanced equation shown below:



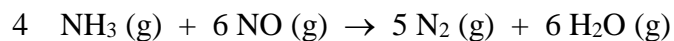
- a) How many grams of  $\text{O}_2$  can be produced from reaction of 12.0 moles of  $\text{NaClO}_3$ ?

$$12.0 \text{ mol NaClO}_3 \times \frac{3 \text{ mol O}_2}{2 \text{ mol NaClO}_3} \times \frac{32.00 \text{ g}}{1 \text{ mol}} = 576 \text{ g O}_2$$

- b) How many grams of  $\text{NaCl}$  are produced when 80.0 g of  $\text{O}_2$  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  $\text{N}_2$ ?



$$13.70 \text{ mol N}_2 \times \frac{4 \text{ mol NH}_3}{5 \text{ mol N}_2} = 10.96 \text{ mol NH}_3$$

$$13.70 \text{ mol N}_2 \times \frac{6 \text{ mol NO}}{5 \text{ mol N}_2} = 16.44 \text{ mol NO}$$

3. What mass of CO<sub>2</sub> can be produced from the reaction of 25.0 g of C<sub>3</sub>H<sub>8</sub> with 75.0 g of O<sub>2</sub> according to the following equation:

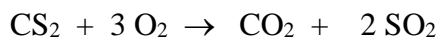


$$25.0 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol}}{44.11 \text{ g}} \times \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} = 1.70 \text{ mol CO}_2$$

$$75.0 \text{ g O}_2 \times \frac{1 \text{ mol}}{32.00 \text{ g}} \times \frac{3 \text{ mol CO}_2}{5 \text{ mol O}_2} = 1.406 \text{ mol CO}_2 \quad \leftarrow \text{Limiting Reactant}$$

$$1.406 \text{ mol CO}_2 \times \frac{44.01 \text{ g}}{1 \text{ mol}} = 61.9 \text{ g CO}_2$$

4. How many grams of SO<sub>2</sub> are produced when 152 g of CS<sub>2</sub> react with 48.0 g of O<sub>2</sub> according to the following equation:



$$152 \text{ g CS}_2 \times \frac{1 \text{ mol}}{76.13 \text{ g}} \times \frac{2 \text{ mol SO}_2}{1 \text{ mol CS}_2} = 3.99 \text{ mol SO}_2$$

$$48.0 \text{ g O}_2 \times \frac{1 \text{ mol}}{32.00 \text{ g}} \times \frac{2 \text{ mol SO}_2}{3 \text{ mol O}_2} = 1.00 \text{ mol SO}_2 \quad \leftarrow \text{Limiting Reactant}$$

$$1.00 \text{ mol CO}_2 \times \frac{64.06 \text{ g}}{1 \text{ mol}} = 64.1 \text{ g SO}_2$$

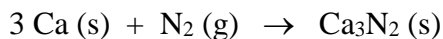
5. When 50.0 g of MgCO<sub>3</sub> react completely with H<sub>3</sub>PO<sub>4</sub>, 15.8 g of CO<sub>2</sub> are produced. What is the percent yield for this reaction? The unbalanced equation is given below:



$$50.0 \text{ g MgCO}_3 \times \frac{1 \text{ mol}}{84.31 \text{ g}} \times \frac{3 \text{ mol CO}_2}{3 \text{ mol MgCO}_3} \times \frac{44.01 \text{ g}}{1 \text{ mol}} = 26.1 \text{ g CO}_2$$

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

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?



$$56.6 \text{ g Ca} \times \frac{1 \text{ mol}}{40.08 \text{ g}} \times \frac{1 \text{ mol Ca}_3\text{N}_2}{3 \text{ mol Ca}} \times \frac{148.26 \text{ g}}{1 \text{ mol}} = 69.8 \text{ g Ca}_3\text{N}_2 \quad \leftarrow \text{Limiting Reactant}$$

$$30.5 \text{ g N}_2 \times \frac{1 \text{ mol}}{28.02 \text{ g}} \times \frac{1 \text{ mol Ca}_3\text{N}_2}{1 \text{ mol N}_2} \times \frac{148.26 \text{ g}}{1 \text{ mol}} = 161 \text{ g Ca}_3\text{N}_2$$

$$69.8 \text{ g Ca}_3\text{N}_2 \times \frac{90.0 \text{ g}}{100 \text{ g}} = 62.8 \text{ g Ca}_3\text{N}_2$$

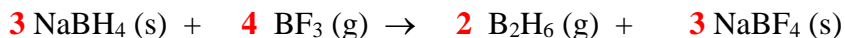
7. Hydrogen gas has been suggested as a clean fuel because it produces only water when it burns. If the reaction has a 98% yield, what mass of hydrogen gas would form 75.0 kg of water?



$$\text{Theoretical yield} = \frac{\text{Actual yield}}{\% \text{ yield}} = \frac{75.0 \text{ kg}}{0.98} = 76.5 \text{ kg H}_2\text{O}$$

$$76.5 \text{ kg H}_2\text{O} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{2 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \times \frac{2.02 \text{ g}}{1 \text{ mol}} = 8.58 \text{ kg H}_2$$

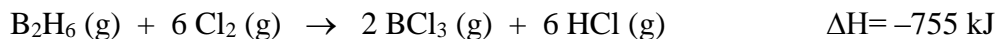
8. Diborane ( $\text{B}_2\text{H}_6$ ) can be produced by the unbalanced reaction shown below. If the reaction has an 85.0% yield, how many grams of  $\text{NaBH}_4$  are needed to produce 20.0 g of diborane?



$$\text{Theoretical yield} = \frac{\text{Actual yield}}{\% \text{ yield}} = \frac{20.0 \text{ g}}{0.850} = 23.5 \text{ g B}_2\text{H}_6$$

$$23.5 \text{ g B}_2\text{H}_6 \times \frac{1 \text{ mol}}{27.68 \text{ g}} \times \frac{3 \text{ mol NaBH}_4}{2 \text{ mol B}_2\text{H}_6} \times \frac{37.84 \text{ g}}{1 \text{ mol}} = 48.2 \text{ g NaBH}_4$$

9. Consider the reaction shown below:

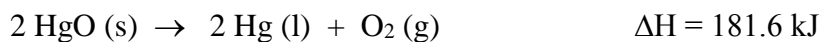


a) Is this reaction endothermic or exothermic? Exothermic

b) How much heat is released when 85.0 g of  $\text{B}_2\text{H}_6$  react?

$$85.0 \text{ g B}_2\text{H}_6 \times \frac{1 \text{ mol B}_2\text{H}_6}{27.68 \text{ g}} \times \frac{755 \text{ kJ}}{1 \text{ mol B}_2\text{H}_6} = 2320 \text{ kJ} \quad (3 \text{ sig figs})$$

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



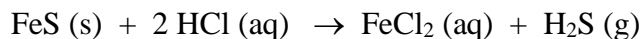
a) How much heat is needed to decompose 555 g of  $\text{HgO}$ ?

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

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

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

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



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?

$$20.0 \text{ g FeS} \times \frac{1 \text{ mol}}{87.91 \text{ g}} \times \frac{2 \text{ mol HCl}}{1 \text{ mol FeS}} \times \frac{36.46 \text{ g}}{1 \text{ mol}} = 16.6 \text{ g HCl}$$
$$23.8 \text{ g HCl} \times \frac{1 \text{ mol}}{36.46 \text{ g}} \times \frac{1 \text{ mol FeS}}{2 \text{ mol HCl}} \times \frac{87.91 \text{ g}}{1 \text{ mol}} = 28.7 \text{ g FeS}$$

**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?



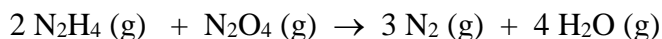
$$2.5 \text{ g NaHCO}_3 \times \frac{1 \text{ mol}}{84.01 \text{ g}} \times \frac{1 \text{ mol HCl}}{1 \text{ mol NaHCO}_3} \times \frac{36.46 \text{ g}}{1 \text{ mol}} = 1.1 \text{ g HCl}$$

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.)



$$258 \text{ mg BaSO}_4 \times \frac{1 \text{ g}}{10^3 \text{ mg}} \times \frac{1 \text{ mol}}{233.39 \text{ g}} \times \frac{1 \text{ mol Ba}^{2+}}{1 \text{ mol BaSO}_4} \times \frac{137.33 \text{ g}}{1 \text{ mol}} = 0.152 \text{ g}$$

14. Consider the reaction shown below:



A reaction flask initially contains 27.5 g of  $\text{N}_2\text{H}_4$  and 74.9 g of  $\text{N}_2\text{O}_4$ . 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.)

$$27.5 \text{ g N}_2\text{H}_4 \times \frac{1 \text{ mol}}{32.06 \text{ g}} \times \frac{1 \text{ mol N}_2\text{O}_4}{2 \text{ mol N}_2\text{H}_4} \times \frac{92.02 \text{ g}}{1 \text{ mol}} = 39.5 \text{ g N}_2\text{O}_4 \text{ reacted}$$

Since amount of  $\text{N}_2\text{O}_4$  reacted is less than amount available (74.9 g), it is in excess.

Therefore, at end of reaction:

$$\text{excess N}_2\text{O}_4 = 74.9 \text{ g} - 39.5 \text{ g} = 35.4 \text{ g}$$

$$27.5 \text{ g N}_2\text{H}_4 \times \frac{1 \text{ mol}}{32.06 \text{ g}} \times \frac{3 \text{ mol N}_2}{2 \text{ mol N}_2\text{H}_4} \times \frac{28.02 \text{ g}}{1 \text{ mol}} = 36.1 \text{ g N}_2$$

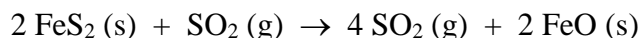
$$27.5 \text{ g N}_2\text{H}_4 \times \frac{1 \text{ mol}}{32.06 \text{ g}} \times \frac{4 \text{ mol H}_2\text{O}}{2 \text{ mol N}_2\text{H}_4} \times \frac{18.02 \text{ g}}{1 \text{ mol}} = 30.9 \text{ g H}_2\text{O}$$

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.)



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

16. Pyrite ( $\text{FeS}_2$ ), an impurity in some coals, reacts with oxygen to form the air pollutant sulfur dioxide, as shown below. What mass of  $\text{SO}_2$  (in grams) is produced when  $2.0 \times 10^4$  kg of coal containing 0.050 mass % pyrite is burned?



$$2.0 \times 10^4 \text{ kg coal} \times \frac{10^3 \text{ g}}{1 \text{ kg}} \times \frac{0.050 \text{ g FeS}_2}{100 \text{ g coal}} \times \frac{1 \text{ mol}}{119.97 \text{ g}} \times \frac{4 \text{ mol SO}_2}{2 \text{ mol FeS}_2} \times \frac{64.06 \text{ g}}{1 \text{ mol}} = 1.1 \times 10^4 \text{ g SO}_2$$