Unit 8

Stoichiometry

Chemistry Assignments and Objectives

EQ: Why do you have to refill your car's gas tank?

Lesson 1 – Learning Targets

1. Determine the mole ratios from a balanced chemical equation.
2. Explain and use the steps used in solving stoichiometric problems.

Lesson 1 – Homework Problems

1. A reaction between methane and sulfur produces carbon disulfide (CS_2), a liquid often used in the production of cellophane.

$$\underline{2}\ CH_4\ (g) + \underline{\quad}\ S_8\ (s) \rightarrow \underline{2}\ CS_2\ (l) + \underline{4}\ H_2S\ (g)$$

a) Balance the equation

b) Calculate the mol CS_2 produced when 1.50 mol S_8 is used. $1.50\,mol\,S_8 \times \frac{2\,mol\,CS_2}{1\,mol\,S_8} = 3.00\,mol\,CS_2$

c) How many mol H_2S is produced? $1.50\,mol\,S_8 \times \frac{4\,mol\,H_2S}{1\,mol\,S_8} = 6.00\,mol\,H_2S$

2. Titanium is a transition metal used in many alloys because it is extremely strong and lightweight. Titanium tetrachloride ($TiCl_4$) is extracted from titanium oxide using chlorine and coke (carbon).

$$TiO_2\ (s) + C(s) + 2Cl_2\ (g) \rightarrow TiCl_4\ (s) + CO_2\ (g)$$

What mass of Cl_2 gas is needed to react with 1.25 mol of TiO_2?

$$1.25\,mol\,TiO_2 \times \frac{2\,mol\,Cl_2}{1\,mol\,TiO_2} \times \frac{70.90\,g\,Cl_2}{1\,mol\,Cl_2} = 177\,g\,Cl_2$$

3.Sodium chloride is decomposed into the elements sodium and chlorine by means of electrical energy. How many grams of chlorine gas can be obtained from 2.50 mol NaCl? (Note: Write/balance the equation.)

$2\,NaCl(s) \rightarrow 2\,Na(s) + Cl_2(g)$ $\quad 2.50\,mol\,NaCl \times \frac{1\,mol\,Cl_2}{2\,mol\,NaCl} \times \frac{70.90\,g\,Cl_2}{1\,mol\,Cl_2} = 88.6\,g\,Cl_2$

4. One in a series of reactions that inflate air bags in automobiles is the decomposition of sodium azide (NaN_3). Determine the mass of N_2 produced if 100.0 g NaN3 is decomposed.

$$2\,NaN_3\ (s) \rightarrow 2\,Na\ (s) + 3\,N_2\ (g)$$

$$100.0\,g\,NaN_3 \times \frac{1\,mol\,NaN_3}{65.02\,g\,NaN_3} \times \frac{3\,mol\,N_2(g)}{2\,mol\,NaN_3} \times \frac{28.02\,g\,N_2}{1\,mol\,N_2(g)} = 64.64\,g\,N_2$$

Lesson 2 – Learning Targets

1. Identify the limiting reactant in a chemical equation.
2. Identify the excess reactant in a chemical equation.
3. Determine how much of the excess reactant remains after a reaction is done.
4. Realize that the amount of the limiting reactant is what determines the amount of product formed.
5. Calculate the mass of product formed when the amounts of more than one reactant are given in a problem.

Lesson 2 – Homework Problems

1. The reaction between solid sodium and iron(III)oxide is one in a series of reactions that inflates an automobile airbag.

$$6\,Na\,(s) + Fe_2O_3\,(s) \rightarrow 3\,Na_2O\,(s) + 2\,Fe\,(s)$$

If 100.0 g Na and 100.0 g Fe_2O_3 are used in this reaction, determine:

a) The limiting reactant Fe_2O_3 $100.0\,g\,Na \times \frac{1\,mol\,Na}{22.99\,g\,Na} = \frac{4.350\,mol\,Na}{6} = .725$

b) The excess reactant Na $100.0\,g\,Fe_2O_3 \times \frac{1\,mol}{159.70\,g} = \frac{.6262\,mol}{1\,mol} = .6262$

c) The mass of solid iron produced $.6262\,mol\,Fe_2O_3 \times \frac{2\,mol\,Fe}{1\,mol\,Fe_2O_3} \times \frac{55.85\,g\,Fe}{1\,mol\,Fe} = 69.94\,g\,Fe$

2. Photosynthesis reactions in green plants use carbon dioxide and water to produce glucose ($C_6H_{12}O_6$) and oxygen. **Write the balanced chemical equation for the reaction.** If a plant has 88.0 g carbon dioxide and 64.0 g water available for photosynthesis, determine: $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

a) The limiting reactant CO_2 $88.0\,g\,CO_2 \times \frac{1\,mol\,CO_2}{44.01\,g\,CO_2} = \frac{2.00\,mol\,CO_2}{6} = .333$ ← limiting

c) The excess reactant H_2O $64.0\,g\,H_2O \times \frac{1\,mol\,H_2O}{18.016\,g\,H_2O} = \frac{3.55\,mol\,H_2O}{6} = .592$

d) The mass of glucose produced $88.0\,g\,CO_2 \times \frac{1\,mol\,CO_2}{44.01\,g\,CO_2} \times \frac{1\,mol\,C_6H_{12}O_6}{6\,mol\,CO_2} \times \frac{180.156\,g\,C_6H_{12}O_6}{1\,mol\,C_6H_{12}O_6}$

60.0g

Lesson 3 – Learning Targets

1. Realize that the amount of product calculated is called a theoretical yield.
2. Use laboratory data to determine a percent yield.

Lesson 3 – Homework Problems

1. Aluminum hydroxide is often present in antacids to neutralize stomach acid (HCl). If 14.0 g aluminum hydroxide is present in an antacid tablet, determine the theoretical yield of aluminum chloride produced when the tablet reacts with stomach acid. If the actual yield of aluminum chloride from this tablet is 22.0 g, what is the percent yield?

$$Al(OH)_3\,(s) + 3HCl\,(aq) \rightarrow AlCl_3\,(aq) + 3\,H_2O\,(l)$$

$14.0\,g\,Al(OH)_3 \times \frac{1\,mol\,Al(OH)_3}{78.004\,g\,Al(OH)_3} \times \frac{1\,mol\,AlCl_3}{1\,mol\,Al(OH)_3} \times \frac{133.33\,g\,AlCl_3}{1\,mol\,AlCl_3} = 23.9\,g\,AlCl_3$ theoretical

$\frac{22.0}{23.9} \times 100 = 92.1\%$

2. When copper wire is placed into a silver nitrate solution, silver crystals and copper(II)nitrate form. Write the balanced equation for the reaction. If a 20.0-g sample of copper is used, determine the theoretical yield of silver. If 60.0 g silver is actually recovered from the reaction, determine the percent yield of the reaction.

$Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag + Cu(NO_3)_2(aq)$

$20.0\,g\,Cu \times \frac{1\,mol\,Cu}{63.55\,g\,Cu} \times \frac{2\,mol\,Ag}{1\,mol\,Cu} \times \frac{107.87\,g\,Ag}{1\,mol\,Ag} = 67.9\,g\,Ag$ theoretical

$\frac{60.0\,g}{67.9\,g} \times 100 = 88.4\%$

Unit 8 Review

final stoic: $3.91\ mol\ Na \times \frac{1\ mol\ H_2}{2\ mol\ Na} \times \frac{22.4\ L\ H_2}{1\ mol\ H_2} = \boxed{44.0\ L\ H_2}$

Limiting Reactant Practice

1. Sodium reacts with water to form sodium hydroxide and hydrogen gas according to the equation:

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

If 90.0 g of sodium is dropped into 80.0 g of water, how many liters of hydrogen at STP would be produced?

$90.0\ g\ Na \times \frac{1\ mol\ Na}{22.99\ g\ Na} = \frac{3.91\ mol}{2} = 1.96$ limiting

$80.0\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} = \frac{4.44\ mol\ H_2O}{2} = 2.22$ excess

2. Phosphorus burns in oxygen gas to produce tetraphosphorus decaoxide according to the equation:

$$P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(g)$$

If 2.50 g of phosphorous is ignited in a flask containing 750 mL of oxygen at STP, how many grams of P_4O_{10} is formed?

$2.50\ g\ P_4 \times \frac{1\ mol\ P_4}{124\ g\ P_4} = \frac{.020\ mol}{1} = .020$ excess

$.750\ L\ O_2 \times \frac{1\ mol\ O_2}{22.4\ L} = \frac{.033\ mol}{5} = .0067$ limiting

Final Stoic: $.033\ mol\ O_2 \times \frac{1\ mol\ P_4O_{10}}{5\ mol\ O_2} \times \frac{284\ g\ P_4O_{10}}{1\ mol\ P_4O_{10}} = \boxed{1.90\ g\ P_4O_{10}}$

Stoichiometry Practice

1. At STP, how many moles of hydrogen gas are needed to react with 15.1 g chlorine gas to produce hydrogen chloride gas?

$H_2 + Cl_2 \rightarrow 2HCl$

$15.1\ g\ Cl_2 \times \frac{1\ mol\ Cl_2}{70.90\ g\ Cl_2} \times \frac{1\ mol\ H_2}{1\ mol\ Cl_2} = .213\ mol\ H_2$

2. At STP, how many grams of oxygen gas are needed to react completely with 16.2 moles of hydrogen gas to produce water?

$2H_2 + O_2 \rightarrow 2H_2O$

$16.2\ mol\ H_2 \times \frac{1\ mol\ O_2}{2\ mol\ H_2} \times \frac{32.00\ g\ O_2}{1\ mol\ O_2} = 259\ g\ O_2$

3. How many grams of solid iodine are produced by the decomposition of 25.4 g of hydrogen iodide gas at STP?

$$2\ HI(g) \rightarrow I_2(s) + H_2(g)$$

$25.4\ g\ HI \times \frac{1\ mol\ HI}{127.908\ g\ HI} \times \frac{1\ mol\ I_2}{2\ mol\ HI} \times \frac{253.8\ g\ I_2}{1\ mol\ I_2} = 25.2\ g\ I_2$

4. What volume of chlorine gas, measured at STP, can be produced by the decomposition of 73.0 g of hydrogen chloride gas given the following equation?

$$2HCl(g) \rightarrow H_2(g) + Cl_2(g)$$

$73.0\ g\ HCl \times \frac{1\ mol\ HCl}{36.458\ g\ HCl} \times \frac{1\ mol\ Cl_2}{2\ mol\ HCl} \times \frac{22.4\ L\ Cl_2}{1\ mol\ Cl_2} = 22.4\ L\ Cl_2$

5. Find the number of grams of sulfur that is produced when 1.87 L of oxygen gas at STP is also produced by the decomposition of sulfur trioxide gas according to this chemical equation:

$$2SO_3(g) \rightarrow 2S(s) + 3O_2(g)$$

$1.87\ L\ O_2 \times \frac{1\ mol\ O_2}{22.4\ L\ O_2} \times \frac{2\ mol\ S}{3\ mol\ O} \times \frac{32.07\ g\ S}{1\ mol\ S} = 1.78\ g\ S$

Percent Yield Practice

1. Methane combusts with oxygen and produces carbon dioxide and water. How many liters of carbon dioxide will be made from 1.00×10^4 grams of methane at STP? What is the percent yield if 13, 500 liters are actually produced?

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

$1.00\times10^4\ g\,CH_4 \times \frac{1\ mol\ CH_4}{16.042\ g\ CH_4} \times \frac{1\ mol\ CO_2}{1\ mol\ CH_4} \times \frac{22.4\ L\ CO_2}{1\ mol\ CO_2} = 1.40 \times 10^4\ L\ CO_2$

$\frac{13,500}{14000} \times 100 = 96.$

3. Pure zirconium is obtained using the two-step Van Arkel process. In the first step, impure zirconium and iodine are heated to produce zirconium iodide (ZrI_4). In the second step, ZrI_4 is decomposed to produce pure zirconium.

$$ZrI_4\,(s) \rightarrow Zr\,(s) + 2\,I_2\,(g)$$

Determine the percent yield of zirconium if 45.0 g ZrI_4 is decomposed and 5.00 g pure Zr is obtained.

$45.0\ g\ ZrI_4 \times \frac{1\ mol\ ZrI_4}{598.52\ g\ ZrI_4} \times \frac{1\ mol\ Zr}{1\ mol\ ZrI_4} \times \frac{91.22\ g\ Zr}{1\ mol\ Zr} = 6.85\ g\ Zr$

$\frac{5.00}{6.85} \times 100 = 73.0\%$

3. a) If 10.0 L of nitrogen reacts with 16.0 L of hydrogen, how many liters of ammonia gas would be produced at STP. (1: excess; 3: limiting)

$$N_2\,(g) + 3\,H_2\,(g) \rightarrow 2\,NH_3\,(g)$$

$16.0\ L\ H_2 \times \frac{2\ mol\ NH_3}{3\ mol\ H_2} = 10.7\ L\ NH_3$

For gases, volume ratio is also mole ratio.

b) If 9.8 L of ammonia at STP is actually recovered, calculate the percent yield of the reaction.

$\frac{9.8}{10.7} \times 100 = 92\%$ yield