## CHEM 106 Chapter 9 Stoichiometry

### MOLE-MOLE

1. Balance the equation. How many moles of $H_2S$ are expected when 2.55 moles methane react?

   $CH_4(g) + S(s) \rightarrow CS_2(l) + H_2S(g)$

   CH4 — 4 — 2

| | R | P |
|---|---|---|
| C | 1 | 1 |
| H | 4 | 2 4 |
| S | 1 | 1 3 4 |

2.55 mol — ? mol

$2.55 \text{ mol } CH_4 \times \frac{2 \text{ mol } H_2S}{1 \text{ mol } CH_4} = 5.10 \text{ mol } H_2S$

(3) exact

2. Given the balanced equation:

   $K_2Cr_2O_7(aq) + 6KI(aq) + 7H_2SO_4(aq) \rightarrow Cr_2(SO_4)_3(aq) + 4K_2SO_4(aq) + 3I_2(aq) + 7H_2O(l)$

   a) Calculate the number of moles of $H_2SO_4$ that will react with 2.0 mol KI.

   ?

   $2.0 \text{ mol KI} \times \frac{7 \text{ mol } H_2SO_4}{6 \text{ mol KI}} = 2.33333$

   (2) exact

   2.3 mol $H_2SO_4$

   b) Calculate the number of moles of $I_2$ that will be produced from 2.0 mol KI.

   ?

   $2.0 \text{ mol KI} \times \frac{3 \text{ mol } I_2}{6 \text{ mol KI}} = 1$

   (2)

   1.0 mol $I_2$

**MOLE-GRAM**

**3.** How many grams of aluminum metal should react to produce 0.0305 moles aluminum sulfate?

$2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

? g          0.0305 mol

$0.0305 \text{ mol } Al_2(SO_4)_3 \times \frac{2 \text{ mol Al}}{1 \text{ mol } Al_2(SO_4)_3} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 1.64578$

? mol Al

0.0610 mol Al

**1.65 g Al**

**4.** Oxygen gas can be produced by decomposing hydrogen peroxide:

$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

35.0g          ? mol

If 35.0 grams of pure hydrogen peroxide are decomposed, how many moles of oxygen are produced?

$35.0 \text{ g } H_2O_2 \times \frac{1 \text{ mol } H_2O_2}{34.016 \text{ g } H_2O_2} \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O_2} =$

$= 0.5144637 8$

**0.514 mol $O_2$**

**GRAM-GRAM** 80.908 g/mol 74.096g/mol

**5.** How many grams of HBr would react with 1.29 grams $Ca(OH)_2$ in the equation below?

$2HBr(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + CaBr_2(aq)$

?g 1.29g

$$1.29\,g\,Ca(OH)_2 \times \frac{1\,mol\,Ca(OH)_2}{74.096\,g\,Ca(OH)_2} \times \frac{2\,mol\,HBr}{1\,mol\,Ca(OH)_2} \times \frac{80.908\,g\,HBr}{1\,mol\,HBr}$$

0.0174 mol $Ca(OH)_2$

0.0348 mol HBr

2.81719

**2.82g HBr**

L.R.

**6.** What mass silver chloride is expected (theoretical yield) when 132.0 grams $AlCl_3$ react with excess silver nitrate?

133.33 g/mol 143.35

$AlCl_3(aq) + 3AgNO_3(aq) \rightarrow 3AgCl(s) + Al(NO_3)_3(aq)$

132.0g excess ?g

(4)

$$132.0\,g\,AlCl_3 \times \frac{1\,mol\,AlCl_3}{133.33\,g\,AlCl_3} \times \frac{3\,mol\,AgCl}{1\,mol\,AlCl_3} \times \frac{143.35\,g\,AgCl}{1\,mol\,AgCl}$$

$= 425.8\,g\,AgCl$ theoretical yield

e.g. A student carries out the reaction and obtains 407.3g AgCl. what is the percent yield?

$$\%yield = \frac{407.3\,g}{425.8\,g} \times 100\%$$

$\%y = 95.66\%$

**LIMITING REACTANT**

7. You combine 5.00 moles butane with 10.00 moles oxygen. The balanced equation is below.

$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$

5.00 mol 10.00 mol

a) How many moles of water ~~is~~ are produced?

b) What is the limiting reactant?

$$5.00\ mol\ C_4H_{10} \times \frac{10\ mol\ H_2O}{2\ mol\ C_4H_{10}} = 25.0\ mol\ H_2O$$

$$10.00\ mol\ O_2 \times \frac{10\ mol\ H_2O}{13\ mol\ O_2} = 7.692\ mol\ H_2O$$

L.R.

$O_2$ is the limiting reactant

**8.** How many grams silver bromide (AgBr) can be formed when solutions containing 50.0 g $MgBr_2$ and 100.0 g $AgNO_3$ are mixed? Which reactant is limiting? The molar masses are provided along with the balanced equation below.

184.11 g/mol    169.87 g/mol    187.77 g/mol

$MgBr_2(aq) + 2AgNO_3(aq) \rightarrow 2AgBr(s) + Mg(NO_3)_2(aq)$

50.0g    100.0g    ? g

Since two reactant amounts given, problem is L.R.

$MgBr_2 \rightarrow AgBr$

$$50.0\,g\,MgBr_2 \times \frac{1\,mol\,MgBr_2}{184.11\,g\,MgBr_2} \times \frac{2\,mol\,AgBr}{1\,mol\,MgBr_2} \times \frac{187.77\,g\,AgBr}{1\,mol\,AgBr} = 102\,g\,AgBr$$

theoretical yield

$AgNO_3 \rightarrow AgBr$

$$100.0\,g\,AgNO_3 \times \frac{1\,mol\,AgNO_3}{169.87\,g\,AgNO_3} \times \frac{2\,mol\,AgBr}{2\,mol\,AgNO_3} \times \frac{187.77\,g\,AgBr}{1\,mol\,AgBr} = 111\,g\,AgBr$$

$MgBr_2$ is the limiting reactant, since it will only yield 102g AgBr (the smaller of the two calculations).

$CCl_4$ $CS_2$ $Cl_2$

**9.** Carbon tetrachloride was prepared by reacting 100.g carbon disulfide and 100.g chlorine gas. The equation and molar masses are below.

Molar Masses (g/mol): 76.14 70.90 153.82 135.04

$$CS_2(l) + 3Cl_2(g) \rightarrow CCl_4(l) + S_2Cl_2(l)$$

100.g 100.g ?

Calculate the percent yield if 65.0 g $CCl_4$ was obtained. → Actual Yield

STEPS

$CS_2 \rightarrow CCl_4$

$$100.\text{g } CS_2 \times \frac{1\text{ mol } CS_2}{76.14\text{ g } CS_2} \times \frac{1\text{ mol } CCl_4}{1\text{ mol } CS_2} \times \frac{153.82\text{ g } CCl_4}{1\text{ mol } CCl_4} = 202\text{ g } CCl_4$$

$Cl_2 \rightarrow CCl_4$

$$100.\text{g } Cl_2 \times \frac{1\text{ mol } Cl_2}{70.90\text{ g } Cl_2} \times \frac{1\text{ mol } CCl_4}{3\text{ mol } Cl_2} \times \frac{153.82\text{ g } CCl_4}{1\text{ mol } CCl_4} = 72.3\text{ g } CCl_4$$

theoretical yield

$Cl_2$ = L.R.

$$\text{percent yield} = \frac{65.0\text{ g actual}}{72.3\text{ g theoretical}} \times 100\%$$

$$= 89.9\%\text{ yield}$$