ii) $n_{C_6H_{12}O_6} = 43.9\ g\ C_6H_{12}O_6 \quad \frac{1\ mol\ C_6H_{12}O_6}{180.16\ g\ C_6H_{12}O_6} = 0.244\ mol\ C_6H_{12}O_6$

Step 3:

Example: Calculate the mass of chlorine required to react with 84.453 g iron. How much iron (III) chloride will be produced?

Step 1: Build the balanced chemical reaction.

Skeleton equation: $Fe\ (s) + Cl_2\ (g) \qquad FeCl_3\ (s)$

Balanced equation: $2\ Fe\ (s) + 3\ Cl_2\ (g) \qquad 2\ FeCl_3\ (s)$

Step 2: Calculate the number of moles of the given material.

$$n_{Fe} = 84.453\ g\ Fe\ \frac{1\ mol\ Fe}{55.85\ g\ Fe} = 1.512\ mol\ Fe$$

Step 3: Calculate the number of moles of the other chemicals in the equation.

$$n_{Cl_2} = 1.512\ mol\ Fe\ \frac{3\ mol\ Cl_2}{2\ mol\ Fe} = 2.268\ mol\ Cl_2$$

$$n_{FeCl_3} = 1.512\ mol\ Fe\ \frac{2\ mol\ FeCl_3}{2\ mol\ Fe} = 1.512\ mol\ FeCl_3$$

Step 4: Calculate the masses

i) Calculate the molar masses

$Cl_2 = 2\ (35.453) = 70.906\ g\ mol^{-1}$

$FeCl_3 = (55.85) + 3\ (35.453) = 162.21\ g\ mol^{-1}$

ii) Calculate their masses

$$m_{Cl_2} = 2.268\ mol\ Cl_2\ \frac{70.906\ mol\ Cl_2}{mol\ Cl_2} = 160.8\ g\ Cl_2$$

$$m_{FeCl_3} = 1.512\ mol\ FeCl_3\ \frac{162.21\ g}{mol\ FeCl_3} = 245.3\ g\ FeCl_3$$

Example: How many grams of hydrogen are needed to react completely with 75.4327 g nitrogen? How many gram of ammonia will be produced?

Step 1: Build the balanced chemical reaction.

Skeleton equation: N_2 (g) + H_2 (g) NH_3 (g)

Balanced equation: N_2 (g) + 3 H_2 (g) 2 NH_3 (g)

Step 2: Calculate the number of moles of the given material.

Molar mass of N_2 = 2 (14.0067) = 28.0134 g mol^{-1}

$$n_{N_2} = 75.4327\ g\ N_2\ \frac{1\ mol\ N_2}{28.0134\ g\ N_2} = 2.69274\ mol\ N_2$$

Step 3:

8.2 Limiting Reactants

Example: What mass of aluminum is required in the single displacement reaction of 21.4 g iron (III) oxide? What mass of products are formed?

Atomic Mass: Al = 26.98 Fe = 55.85 O = 16.00

Step 1: Write the balanced chemical equation

$$Fe_2O_3\,(s) + 2Al\,(s) \qquad 2Fe\,(s) + Al_2O_3\,(s)$$

Step 2: Calculate the number of moles of the given compound

Molar Mass Fe_2O_3 = 2 x (55.85) + 3 x (16.00) = 159.70

Al_2O_3 = 2 x (26.98) + 3 x (16.00) = 101.96

$$n_{Fe_2O_3} = 21.4\ g\ Fe_2O_3\ \frac{mol}{159.70\ g\ Fe_2O_3} = 0.134\ mol\ Fe_2O_3$$

Step 3: Calculate the number of moles of the other compounds in the chemical equation

Example: When 36.127 g benzene, C_6H_6 , is burnt in 115.723 g oxygen, how much product is formed?

Atomic Mass: H = 1.00797 C = 12.011 O = 15.9994

In this problem, we actually have more oxygen than we need. Thus, we have a basis for judging whether we have established the limiting reactant.

Example: Calculate the amount of each compound remaining after 5.00 g each of hydrogen and oxygen are reacted to form water?

Atomic Mass: H =

Example: If 20.0 g aluminum is reacted with 15.0 g oxygen, how much aluminum oxide will be obtained?

Atomic Mass: Al = 26.98 O = 16.00

Step 1: Write the balanced chemical equation

$$4\,Al\,(s) + 3\,O_2\,(g) \quad 2\,Al_2O_3\,(s)$$

Step 2: Calculate the number of moles of the given compounds

Molar Mass O_2 = 2 x (16.00) = 32.00

Al_2O_3 = 2 x (26.98) + 3 x (16.00) = 101.96

$$n_{Al} = 20.0\ g\ Al\ \frac{mol}{26.98\ g\ Al} = 0.741\ mol\ Al$$

$$n_{O_2} = 15.0\ g\ O_2\ \frac{mol}{32.00\ g\ O_2} = 0.469\ mol\ O_2$$

Step 3: Step 3: Determine the Limiting Reactant.

Assume Al is the Limiting Reactant. Then the number of moles of O_2 required to completely react with all the Al is:

$$n_{O_2} = 0.741\ mol\ Al\ \frac{3\ mol\ O_2}{4\ mol\ Al} = 0.556\ mol\ O_2\ are\ required$$

Compare the number of moles O_2 that are actually present with the number of moles O_2 that are required"

0.467 mol O_2 are given < 0.556 mol O_2 are required.

There are not enough moles O_2 initially present so O_2

8.3 Percent Yield

In the preceding stoichiometric calculations, given a certain mass of reactants, we are asked to calculate the amount of product we expect to get. This expected value for the amount of product is known as the **theoretical yield** as it is achieved by calculations based on a theoretical relationship between moles of reactants and products described by the balanced chemical equation (i.e., obtained through the stoichiometric calculations carried out above).

If the reaction is actually carried out in the lab, a certain experimental value for the mass of product will be obtained. This lab value for the mass is known as the **actual yield** or the experimental yield.

In an ideal world, real life and theory are one and the same, so that the actual yield and the theoretical yield should be the same. In real life however, many things can go improperly so that losses may be incurred, serving to reduce the actual yield (or the product may be contaminated with impurities raising the apparent mass of product, thereby increasing the apparent actual yield). Thus, chemists usually report a percent yield:

$$\boxed{Percent\ Yield = \frac{Actual\ Yield}{Theoretical\ Yield}\ 100\%}$$

This is the ratio of what you got to what you should have gotten.

Thus for example, in the preceding problem where 20.0 g Al are reacted with 15.0 g O

Example: Nitrogen gas and hydrogen gas react to produce ammonia gas according to the unbalanced reaction

$$N_2(g) + H_2(g) \rightarrow NH_3(g).$$

When 25.00 g each of nitrogen and hydrogen were reacted at a certain temperature and

Example: If 49.8 g Fe_3O_4 is mixed with 10.0 g C and they react according to the following unbalanced equation,

$$Fe_3O_4\,(s) \;+\; C\,(s) \qquad Fe\,(s) \;+\; CO\,(g),$$

Step 6: Calculate the actual yield

$$\%\,yield = \frac{Actual\ yield}{Theoretical\ yield}\ 100\%$$

Rearranging this equation:

$$Actual\ yield = (\%\,yield)\ \frac{Theoretical\ yield}{100\%}$$

$$Actual\ yield = (76.5\%)\ \frac{23.3\ g\ CO}{100\%} = 17.8\ g\ CO$$