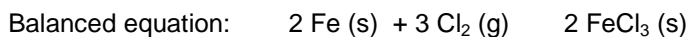
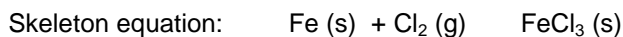


$$\text{ii) } n_{\text{C}_6\text{H}_{12}\text{O}_6} = 43.9 \text{ g C}_6\text{H}_{12}\text{O}_6 \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.16 \text{ g C}_6\text{H}_{12}\text{O}_6} = 0.244 \text{ mol C}_6\text{H}_{12}\text{O}_6$$

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.



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

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

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

$$n_{\text{Cl}_2} = 1.512 \text{ mol Fe} \frac{3 \text{ mol Cl}_2}{2 \text{ mol Fe}} = 2.268 \text{ mol Cl}_2$$

$$n_{\text{FeCl}_3} = 1.512 \text{ mol Fe} \frac{2 \text{ mol FeCl}_3}{2 \text{ mol Fe}} = 1.512 \text{ mol FeCl}_3$$

Step 4: Calculate the masses

i) Calculate the molar masses

$$\text{Cl}_2 = 2 (35.453) = 70.906 \text{ g mol}^{-1}$$

$$\text{FeCl}_3 = (55.85) + 3 (35.453) = 162.21 \text{ g mol}^{-1}$$

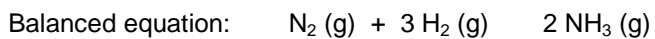
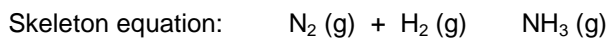
ii) Calculate their masses

$$m_{\text{Cl}_2} = 2.268 \text{ mol Cl}_2 \frac{70.906 \text{ g Cl}_2}{\text{mol Cl}_2} = 160.8 \text{ g Cl}_2$$

$$m_{\text{FeCl}_3} = 1.512 \text{ mol FeCl}_3 \frac{162.21 \text{ g}}{\text{mol FeCl}_3} = 245.3 \text{ g FeCl}_3$$

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.



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

Molar mass of $\text{N}_2 = 2 (14.0067) = 28.0134 \text{ g mol}^{-1}$

$$n_{\text{N}_2} = 75.4327 \text{ g N}_2 \frac{1 \text{ mol N}_2}{28.0134 \text{ g N}_2} = 2.69274 \text{ mol N}_2$$

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



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

$$\text{Molar Mass } \text{Fe}_2\text{O}_3 = 2 \times (55.85) + 3 \times (16.00) = 159.70$$

$$\text{Al}_2\text{O}_3 = 2 \times (26.98) + 3 \times (16.00) = 101.96$$

$$n_{\text{Fe}_2\text{O}_3} = 21.4 \text{ g Fe}_2\text{O}_3 \frac{\text{mol}}{159.70 \text{ g Fe}_2\text{O}_3} = 0.134 \text{ mol Fe}_2\text{O}_3$$

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

$$2 \text{ Al} \quad \frac{\text{mol}}{26.98 \text{ g Al}} = 0.268 \text{ mol Al}$$

$$2 \text{ Fe} \quad \frac{\text{mol}}{55.85 \text{ g Fe}} = 0.268 \text{ mol Fe}$$

$$1 \text{ Al}_2\text{O}_3 \quad \frac{\text{mol}}{101.96 \text{ g Al}_2\text{O}_3} = 0.134 \text{ mol Al}_2\text{O}_3$$

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.

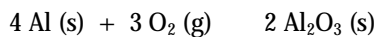
Example: Calculate the amount of each compound remaining after 5.00 g each of hydrogen and oxygen are reacted to form water?

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



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

$$\text{Molar Mass O}_2 = 2 \times (16.00) = 32.00$$

$$\text{Al}_2\text{O}_3 = 2 \times (26.98) + 3 \times (16.00) = 101.96$$

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

$$n_{\text{O}_2} = 15.0 \text{ g O}_2 \frac{\text{mol}}{32.00 \text{ g O}_2} = 0.469 \text{ mol O}_2$$

Step 3: Step 3: Determine the Limiting Reactant.

Assume Al is the Limiting Reactant. Then the number of moles of O₂ required to completely react with all the Al is:

$$n_{\text{O}_2} = 0.741 \text{ mol Al} \frac{3 \text{ mol O}_2}{4 \text{ mol Al}} = 0.556 \text{ mol O}_2 \text{ are required}$$

Compare the number of moles O₂ that are actually present with the number of moles O₂ that are required"

0.467 mol O₂ are given < 0.556 mol O₂ are required.

There are not enough moles O₂ initially present so O₂

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.

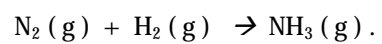
In an ideal world, real life and theory are one and the same, so that the actual yield and the theoretical yield should be the same. In real life however, many things can go improperly so that losses may be incurred, serving to reduce the actual yield (or the product may be contaminated with impurities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

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

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



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,



Step 6: Calculate the actual yield

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

Rearranging this equation:

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

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