Appendix D
Reaction Stoichiometry

D.1 INTRODUCTION

In Appendix A, the stoichiometry of elements and compounds was presented. There, the relationships among grams, moles and number of atoms and molecules were reviewed. A similar relationship exists for chemical reactions, and we will now extend this concept of stoichiometry to reactions. In reaction stoichiometry, we are interested in the quantitative relationships between the amounts of reactants and products in a reaction. We will find, as we did in Appendix A, that the mole is the central character in these calculations.

D.2 QUANTITATIVE RELATIONSHIPS IN REACTIONS

Chemical reactions surround us. Chemists use shorthand notation to describe them in a sentence called a chemical equation. The chemical equation that describes the combustion of benzene is

$$2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(g)$$

This equation implies that for every two benzene molecules that react, 15 dioxygen molecules must also react and 12 carbon dioxide molecules and six water molecules will be produced. In other words, it tells us about the “stoichiometry”, or the amounts of reactants and products involved. The coefficients, or numbers in front of each chemical formula, tell us the relative number of molecules involved in the reaction. They also tell us the relative number of *moles* involved in the reaction. Two moles of benzene will react with 15 moles of dioxygen to form twelve moles of carbon dioxide and six moles of water. It is important to note that the equation does not give us any direct information about the number of grams of each reactant or product, only the moles. If we want to know about a measurable quantity like grams, we will have to do some conversions.

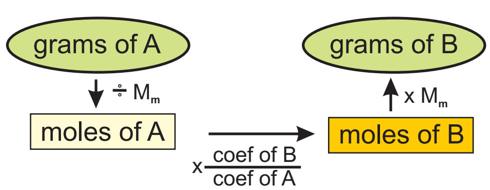
The first requirement for any stoichiometric calculation is a balanced equation. Once we have a balanced equation, the calculations will be performed by following three steps.

1. Convert the given quantitative information to moles. Experiments are always set up (thus chemistry problems are always written) such that number of moles of at least one reactant or product can be determined.
2. Use the balanced equation to convert from the moles of given substance to the moles of desired substance. Remember, we’re given quantitative information on one reactant or product, we desire quantitative information on another. This is the heart of all stoichiometry problems.
3. Convert from moles of the desired compound to the appropriate quantity.

These three steps are purposely vague. In the three previous appendices, we have discussed the conversion of moles to grams for a solid, moles to pressure, volume and temperature for a gas, and moles to volume and molarity for a solution. As you might guess, there are several variations on this three-step theme. The following examples will show some of the variety.

D.3 REACTION STOICHIOMETRY INVOLVING GRAMS

For calculations involving grams of reactants and products, our three-step scheme looks like this:



In the preceding diagram, A and B are products and/or reactants. The gram to mole conversion is achieved through application of the molar mass, the mole to mole conversion comes from the coefficients in the balanced equation.

Example 1

Consider the combustion of benzene:

$$2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(g)$$

What is the maximum mass of CO_2 (M_m = 44.0 g/mol) that can be produced from the combustion of 10.0 g of C_6H_6 (M_m = 78.1 g/mol)?

Solution:

The given information is grams of benzene. The desired information is grams of CO_2. The road map above tells us we must convert grams of benzene to moles, use the mole ratio from the equation to give moles of carbon dioxide, and then convert back to grams of carbon dioxide.

$$\mathbf{10.0\ g\ C_6H_6 \times \frac{1\ mol\ C_6H_6}{78.1\ g\ C_6H_6} = 0.128\ mol\ C_6H_6}$$

$$\mathbf{0.128\ mol\ C_6H_6 \times \frac{12\ mol\ CO_2}{2\ mol\ C_6H_6} = 0.768\ mol\ CO_2}$$

$$\mathbf{0.768\ mol\ CO_2 \times \frac{44.0\ g\ CO_2}{mol\ CO_2} = 33.8\ g\ CO_2}$$

Comment:

As we have seen several times in the previous appendices, each individual step is not difficult. Putting the steps together in a logical manner is the challenge. In this example, we have done each of the three steps separately. However, we could have strung together the conversion factors to save ourselves some writing. We will do that in the next example.

Example 2

How many grams of O_2 (M_m = 32.0 g/mol) are required to completely react with 10.0 g of C_6H_6?

Solution:

The given information is the mass of benzene, the desired information is the mass of dioxygen. Start with the given information and apply conversion factors following the road map.

$$\mathbf{10.0\ g\ C_6H_6 \times \frac{1\ mol\ C_6H_6}{78.1\ g\ C_6H_6} \times \frac{15\ mol\ O_2}{2\ mol\ C_6H_6} \times \frac{32.0\ g\ O_2}{1\ mol\ O_2} = 30.7\ g\ O_2}$$

Comment:

Stoichiometry problems do not always relate reactants to products. Here is a situation where both the given and desired information deals with reactants. In this example, we have done our three step calculation by stringing together the three conversion factors appropriate to the three steps. Note that the order of operation can be determined by using the units because the units of the denominator of each conversion factor must be the same as the units of the previous numerator. Using the units to help is called the factor-label method.

D.4 LIMITING REACTANTS

We know from experience that the amount of product that is formed depends on the amount of reactant that is consumed. You can drive a car only as long as it has gasoline. The gasoline is the limiting reactant because it dictates how much product (miles) can be achieved. The amount of gasoline determines not only how far you can go, but it also determines how much CO_2 and H_2O (the reaction products of the combustion of gasoline) can be made.

In any chemical reaction, the amount of products that are made is limited by the amount of reactants. When any one reactant runs out, the reaction stops. The reactant that runs out is called the *limiting reactant* or *limiting reagent*. Any reactants that do not run out are said to be in excess. In most chemical reactions, one or more of the reactants is in excess. In the gasoline combustion reaction, there is certainly more oxygen available then there is gasoline in the gas tank, and so the oxygen is in excess.

In calculating the amount of product formed in a reaction, we always have to identify the limiting reactant. In some cases it is obvious. In Example 1 above, we read that 10.0 g of C_6H_6 reacts with excess O_2. Clearly, C_6H_6 is the limiting reactant, and O_2 is the excess reactant. But consider the following example.

Calcium hydride reacts with water to form calcium hydroxide and hydrogen gas, *via* the following reaction

$CaH_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + 2H_2(g)$

If 10.0 g of CaH_2 reacts with 9.00 g of H_2O, what mass of $Ca(OH)_2$ can be formed? We start by determining the limiting reactant, but we cannot tell which reactant will limit the amount of product just by comparing the grams of each reactant, we must determine how much $Ca(OH)_2$ we can make from each reactant. Based on the amount of CaH_2 we start with, we can make:

$$10.0\text{ g }CaH_2 \times \frac{1\text{ mol }CaH_2}{42.10\text{ g }CaH_2} \times \frac{1\text{ mol }Ca(OH)_2}{1\text{ mol }CaH_2} \times \frac{74.10\text{ g }Ca(OH)_2}{1\text{ mol }Ca(OH)_2} = 17.6\text{ g }Ca(OH)_2$$

Based on the amount of H_2O available, we can make:

$$9.0\text{ g }H_2O \times \frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} \times \frac{1\text{ mol }Ca(OH)_2}{2\text{ mol }H_2O} \times \frac{74.10\text{ g }Ca(OH)_2}{1\text{ mol }Ca(OH)_2} = 18.5\text{ g }Ca(OH)_2$$

Even though we have enough water to make 18.5 g of $Ca(OH)_2$, there is only enough calcium hydride to make 17.6 g. In this case, CaH_2 is the limiting reactant, H_2O is in excess, and 17.6 g of $Ca(OH)_2$ would be produced.

Let's determine how much of the excess reactant remains. CaH_2 is the limiting reactant and all amounts are calculated from it, and so we must now determine how much water reacts with the CaH_2.

$$10.0\text{ g }CaH_2 \times \frac{1\text{ mol }CaH_2}{42.10\text{ g }CaH_2} \times \frac{2\text{ mol }H_2O}{1\text{ mol }CaH_2} \times \frac{18.0\text{ g }H_2O}{1\text{ mol }H_2O} = 8.56\text{ g }H_2O$$

The above is how much water reacts, we now determine how much remains by subtracting the amount that reacts from the initial amount.

9.00 - 8.56 = 0.44 g H_2O remains

Let's summarize limiting reactants with another example.

Example 3

25.0 g of Na_2SO_4 is added to 7.00 g of carbon and allowed to react according to the following equation:
$Na_2SO_4(s) + 4C(s) \rightarrow Na_2S(s) + 4CO(g)$

a) What is the limiting reactant?

b) How many grams of Na_2S can be formed?

c) How many grams of the excess reactant will be leftover?

Solution:

Perform the three-step calculation twice, starting from the information given for each reactant.

$$25.0\text{ g }Na_2SO_4 \times \frac{1\text{ mol }Na_2SO_4}{142.05\text{ g }Na_2SO_4} \times \frac{1\text{ mol }Na_2S}{1\text{ mol }Na_2SO_4} \times \frac{78.05\text{ g }Na_2S}{1\text{ mol }Na_2S} = 13.7\text{ g }Na_2S$$

$$7.00\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} \times \frac{1\text{ mol }Na_2S}{4\text{ mol C}} \times \frac{78.05\text{ g }Na_2S}{1\text{ mol }Na_2S} = 11.4\text{ g }Na_2S$$

Comparing the two calculations leads us to conclude that C is the limiting reactant and Na_2SO_4 is in excess. 11.4 g of Na_2S can be produced. Next, find the amount of excess reactant that reacts.

$$7.00\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} \times \frac{1\text{ mol }Na_2SO_4}{4\text{ mol C}} \times \frac{142.05\text{ g }Na_2SO_4}{1\text{ mol }Na_2SO_4} = 20.8\text{ g }Na_2SO_4$$

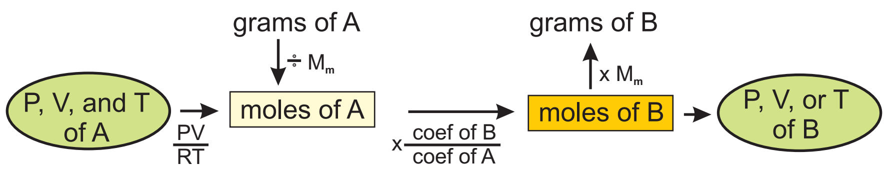
The amount remaining is given by the difference

25.0 g initially - 20.8 g consumed = 4.2 g of Na_2SO_4 remain

D.5 REACTIONS INVOLVING GASES

Examples 1 and 2 dealt with the combustion of benzene. In each, a known amount of benzene was burned, and we calculated the mass of CO_2 or O_2 produced. However, the measurable quantities of gases are pressure, volume and temperature, not mass. In this section, we use the treatment presented in Appendix B to introduce these quantities into our stoichiometric calculations.

If there is known quantitative information about the pressure, volume and temperature of a gas, we can calculate the moles of that gas, using PV = nRT. We can use this relationship in the first or last step of our calculation. We summarize this through the following road map or flowchart:



Notice that this flowchart is identical to the one that appeared earlier in this appendix, except that the use of P,V, and T information for a gas has been added as an entry into the scheme at the left and as a result from the scheme at the right. The following examples show how the ideal gas law can be used in reaction stoichiometry calculations.

Example 4

Octane combusts via the following chemical reaction:

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

How many liters of CO_2, collected at a pressure of 1.00 atm and a temperature of 25 °C, can be produced by the combustion of 35.0 kg of octane (the amount held by a typical car gasoline tank)?

Solution:

We first must recognize this as a reaction stoichiometry problem. We know the mass and therefore the number of moles of octane. We desire the volume of the carbon dioxide, so PV = nRT will be our third step. It is somewhat complicated to use the ideal gas law as a conversion factor, so we will string together the first two steps, and then do the third step separately, solving for volume.

$$35 \text{ kg } C_8H_{18} \times \frac{10^3 \text{ g}}{\text{kg}} \times \frac{1 \text{ mol } C_8H_{18}}{114 \text{ g } C_8H_{18}} \times \frac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}} = 2.45\text{x}10^3 \text{ mol } CO_2$$

Next apply the ideal gas law, PV = nRT.

$$V = \frac{nRT}{P} = \frac{(2.45 \times 10^3 \text{ mol})(0.0821 \text{ L}\cdot\text{atm}\cdot\text{K}^{-1}\cdot\text{mol}^{-1})(298 \text{ K})}{1.00 \text{ atm}} = 5.99\text{x}10^4 \text{ L}$$

Comment:

Burning a tank of gasoline generates enough carbon dioxide at 1.00 atm and 25 °C to fill a 16 ft $\times$ 16 ft room with an 8 ft ceiling.

Example 5

The following reaction is used to quickly inflate some car airbags:

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

How many grams of NaN_3 must be used if you wish to fill a 20.0 L airbag with dinitrogen to a pressure of 1.25 atm at 25 °C?

Solution:

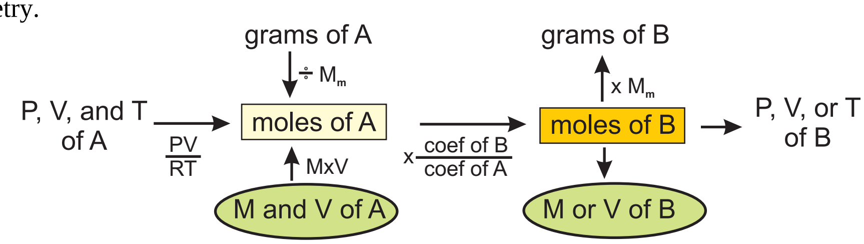
In this problem the known information involves a gaseous product (a pressure, volume and temperature are all given) and information on the reactant is desired. This time, the first of the three steps involves the ideal gas law, with the second and third steps being simple application of conversion factors.

$$n = \frac{PV}{RT} = \frac{(1.25\ \text{atm})(20.0\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm}\cdot\text{K}^{-1}\cdot\text{mol}^{-1})(298\ \text{K})} = 1.02\ \text{mol}\ N_2$$

$$1.02\ \text{mol}\ N_2 \times \frac{2\ \text{mol}\ NaN_3}{3\ \text{mol}\ N_2} \times \frac{65.0\ \text{g}\ NaN_3}{1\ \text{mol}\ NaN_3} = 44.2\ \text{g}\ NaN_3$$

D.6 REACTIONS INVOLVING SOLUTIONS

In the previous sections, we have discussed the quantitative relationships between reactants and products in a reaction. In all cases, the heart of the problem was the mole ratio, and the only difference in the problems lies in how we get to the mole ratio and what we do after we have applied it. We will now consider our last conversion to and from moles. If a reactant or product is dissolved in solution, the moles of that compound are related to the molarity and the solution volume, M = n/V. (See Appendix C if you do not remember this relationship.) We can add this route to our road map of reaction stoichiometry.



The following examples show how solution data can be manipulated along with masses and data on gases to give information on reaction stoichiometry.

Example 6

The Pb^{2+} ions from water soluble $Pb(NO_3)_2$ can be precipitated by the addition of KI, forming insoluble PbI_2:

$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$

If 25.0 mL of 0.375 M $Pb(NO_3)_2$ is reacted with excess KI, how many grams of PbI_2 will be produced?

Solution:

The moles of $Pb(NO_3)_2$ can be found since the volume and molarity of the solution are known. We convert the volume to liters then apply molarity to determine moles. The final two steps of the calculation are the mole to mole conversion from the balanced equation, and the conversion from moles of PbI_2 to grams, using the molar mass of 461.0 g/mol.

$$25.0\ \text{mL} \times \frac{0.375\ \text{mol}\ Pb(NO_3)_2}{1000\ \text{mL}} \times \frac{1\ \text{mol}\ PbI_2}{1\ \text{mol}\ Pb(NO_3)_2} \times \frac{461.0\ \text{g}\ PbI_2}{1\ \text{mol}\ PbI_2} = 4.32\ \text{g}\ PbI_2$$

Example 7

The silver ions in $AgNO_3$ solution can be precipitated by the addition of aqueous NaCl:

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

When 35.0 mL of a $AgNO_3$ solution of unknown concentration is reacted with excess NaCl solution, 8.53 g of AgCl is formed. What is the concentration of the $AgNO_3$ solution?

Solution:

In this problem, there is some information about the $AgNO_3$ solution and some information about the solid AgCl. After rereading the problem, it should become clear that the desired quantity is molarity of the $AgNO_3$ solution. Thus, we need to start at the other end, with the mass of AgCl. You should also recognize that whenever mass data is presented along with either a molar mass or a chemical formula (from which we can get a molar mass), we have an entry into our road map. Here, we will do the first two steps of the calculation in the usual manner, and then use the M = n/V relationship as our third step.

$$8.53\text{ g AgCl} \times \frac{1\text{ mol AgCl}}{143.35\text{ g AgCl}} \times \frac{1\text{ mol AgNO}_3}{1\text{ mol AgCl}} = 0.0595\text{ mol AgNO}_3$$

$$M = \frac{n}{V} = \frac{0.0595\text{ mol}}{0.0350\text{ L}} = 1.70\text{ mol/L} = 1.70\text{ M}$$

Comment:

The first two steps are as we have done many times now. The third step, using the molarity relationship, may at first seem a little unusual. Molarity is always the ratio of moles of a substance to the volume in liters. The first two steps tell us that there are 0.0595 moles of $AgNO_3$ contained in the 35.0 mL of solution. We simply take the ratio of these two numbers, first converting 35.0 mL to 0.0350 L, since molarity is moles per liter. Note that the experiment did not have to be done on a 1 L scale in order to calculate the molarity!

Example 8

Potassium permanganate solutions can react with acidic hydrogen peroxide solutions via the following balanced equation in water:

$$2KMnO_4 + 5H_2O_2 + 6HCl \rightarrow 2MnCl_2 + 5O_2(g) + 2KCl + 8H_2O(l)$$

When 15.0 mL of 0.0200 M $KMnO_4$ reacts with excess H_2O_2 and HCl, how many liters of O_2, collected at a total pressure of 1.00 atm and a temperature of 27 °C, will be formed?

Solution:

The volume and molarity data on $KMnO_4$ allow us to enter into the road map. The desired information is volume of O_2. We start by converting mL to L, and then proceed through the first two of the three steps in the calculation to find moles of O_2. We will do the third step, manipulation of PV = nRT, separately.

$$15.0\text{ mL} \times \frac{0.0200\text{ mol KMnO}_4}{1000\text{ m L}} \times \frac{5\text{ mol O}_2}{2\text{ mol KMnO}_4} = 7.50 \times 10^{-4}\text{ mol O}_2$$

$$V = \frac{nRT}{P} = \frac{(7.50 \times 10^{-4}\text{ mol})\ (0.0821\text{ L}\cdot\text{atm}\cdot\text{K}^{-1}\cdot\text{mol}^{-1})\ (300\text{ K})}{1.00\text{ atm}} = 0.0185\text{ L}$$

Comment:

As you can see, there are many possible routes along our road map for solving stoichiometry problems.

Example 9

How many mL of a 0.250 M NaOH solution are required to completely react with 40.0 mL of a 0.150 M H_2SO_4 solution. The overall reaction is:

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow 2H_2O(l) + Na_2SO_4(aq)$$

Solution:

We see that there is enough information to calculate moles of H_2SO_4. We desire information on NaOH. Start with the volume of 0.0400 L of H_2SO_4 solution and proceed as usual.

$$0.0400\text{ L} \times \frac{0.150\text{ mol H}_2\text{SO}_4}{1\text{ L}} \times \frac{2\text{ mol NaOH}}{1\text{ mol H}_2\text{SO}_4} \times \frac{1\text{ L}}{0.250\text{ mol NaOH}} \times \frac{10^3\text{ L}}{\text{mL}} = 48.0\text{ mL}$$

Comment:

The mL to L conversion was done at the beginning and the end of the problem since the volume information is in mL, but the concentration (molarity) is, of course, in moles per liter. Both molarities were used as conversion factors; 0.150 M converted volume of solution to moles of H_2SO_4, 0.250 M was “turned upside down” to convert moles of NaOH to liters of solution.

Example 10

Aqueous HCl and NaOH react in the following manner:

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$

When 30.0 mL of 0.100 M HCl are mixed with 20.0 mL of 0.125 M NaOH, what is concentration of the excess reagent?

Solution:

This is a limiting reactant problem. At first glance, it would appear quite different than the limiting reactant problems we saw earlier. However, upon careful reading of the experiment, we see that we have quantitative information on both reactants, enough to calculate moles of both. The desired quantity the concentration of the excess reactant. In essence, we are reacting an acid and a base, and need to determine which reactant is limiting, and find how much of the excess reactant is leftover, as was done in Example 3 above. In order to determine the limiting reactant, we calculate how much product can be made from each reactant. It doesn't matter which product we choose. Let's pick water.

$$30\text{ mL} \times \frac{1\text{ L}}{10^3\text{ mL}} \times \frac{0.100\text{ mol HCl}}{1\text{ L}} \times \frac{1\text{ mol H}_2\text{O}}{1\text{ mol HCl}} = 0.0030\text{ mol H}_2\text{O}$$

$$20\text{ mL} \times \frac{1\text{ L}}{10^3\text{ mL}} \times \frac{0.125\text{ mol NaOH}}{1\text{ L}} \times \frac{1\text{ mol H}_2\text{O}}{1\text{ mol NaOH}} = 0.0025\text{ mol H}_2\text{O}$$

Fewer moles of water can be made from the NaOH, so NaOH is the limiting reactant, HCl is the excess reactant. Notice that it was not necessary to go all the way through and calculate the grams of water. Clearly, if we multiply each result by 18.02 g/mol (the molar mass of H_2O), the conclusion is the same, NaOH is limiting. In order to calculate molarity of HCl, we need the number of <u>moles</u> of HCl which were leftover, and the <u>*total*</u> solution volume.

<u>Moles of HCl at the start</u>:

$$30\text{ mL} \times \frac{1\text{ L}}{10^3\text{ mL}} \times \frac{0.100\text{ mol HCl}}{1\text{ L}} = 0.0030\text{ mol HCl}$$

The number of moles of HCl consumed is based on the amount of limiting reactant consumed:

$$20\text{ mL} \times \frac{1\text{ L}}{10^3\text{ mL}} \times \frac{0.125\text{ mol NaOH}}{1\text{ L}} \times \frac{1\text{ mol HCl}}{1\text{ mol NaOH}} = 0.0025\text{ mol HCl react}$$

<u>Moles of HCl remaining</u>:

0.0030 - 0.0025 = 0.0005 moles of HCl remain

<u>Concentration of HCl at the end</u>

$$\frac{\text{moles of HCl}}{\text{total volume}} = \frac{0.0005\text{ mol HCl}}{0.050\text{ L}} = 0.010\text{ M}$$

Comment:

This problem is actually very similar to the limiting reactant problems we did before. The difference is that, instead of finding the grams of the leftover reactant, we had to find the concentration, which involved a calculation of the number of moles of the leftover reactant.

D.7 EXERCISES

Use the following molar masses to do the following problems:

C_4H_8 : 56.10 g/mol	C_4H_9OH : 74.12 g/mol
Fe_2O_3 : 159.70 g/mol	Al_2O_3 : 101.96 g/mol
V_2O_5 : 181.88 g/mol	NH_4VO_3 : 116.98 g/mol
NH_3 : 17.03 g/mol	V_2O_3 : 149.88 g/mol
Cu_2S : 159.17 g/mol	CuO : 79.55 g/mol
Cu_2O : 95.55 g/mol	$AgCl$: 143.4 g/mol

1. In the presence of acids, water can react with alkenes to form alcohols:

$$C_4H_8 + H_2O \rightarrow C_4H_9OH$$

If 250 g of C_4H_8 reacts with excess H_2O, how many grams of C_4H_9OH can be produced?

2. Aluminum reacts with iron(III) oxide in the "thermite reaction":

$$2Al(s) + Fe_2O_3(s) \rightarrow 2Fe(s) + Al_2O_3(s)$$

- **a)** If 10.0 g of Al reacts with excess Fe_2O_3, how many grams of Al_2O_3 can be produced?
- **b)** If 25.0 g of Al reacts with 10.0 g of Fe_2O_3, how many grams of Al_2O_3 can be produced?
- **c)** In the experiment in part b, what is the mass of the excess reactant remaining after complete reaction?

3. Vanadium(V) oxide reacts with ammonia and water as follows:

$$V_2O_5 + 2NH_3 + H_2O \rightarrow 2NH_4VO_3$$

- **a)** If 50.0 g of V_2O_5 is reacted with excess ammonia and water, how many grams of NH_4VO_3 can be produced?
- **b)** How many grams of NH_3 are required to completely react with 50.0 g of V_2O_5?

4. Vanadium(III) oxide can be made by reduction of vanadium(V) oxide with hydrogen:

$$V_2O_5(s) + 2H_2(g) \rightarrow V_2O_3(s) + 2H_2O(l)$$

- **a)** How many liters of H_2, measured at 1.00 atm and 30 ^oC, are required to completely react with 75.0 g of V_2O_5?
- **b)** If 10.0 g of V_2O_5 reacts with 1.65 L of H_2, measured at 1.00 atm and 30 ^oC, how many grams of V_2O_3 can be produced?

5. Copper(I) sulfide is prepared by heating copper and sulfur in the absence of air:

$$2Cu(s) + S(s) \rightarrow Cu_2S(s)$$

- **a)** How many grams of Cu_2S can be produced from the reaction of 25.0 g of Cu with excess S?
- **b)** How many grams of sulfur are required to form 75.0 g of Cu_2S?
- **c)** If a mixture of 135 g of Cu and 45 g of S is allowed to react, how many grams of Cu_2S could be produced?
- **d)** How many grams of the excess reactant remain in the experiment in part c?

6. Copper(I) oxide can be prepared by thermal decomposition of copper(II) oxide:

$$4CuO(s) \rightarrow 2Cu_2O(s) + O_2(g)$$

- **a)** How many grams of Cu_2O can be produced upon the decomposition of 450 g of CuO?
- **b)** How many liters of O_2, collected at 1.00 atm and 27 ^oC, can be produced by the decomposition of 450 g of CuO?

7. The silver ions in aqueous silver sulfate can be precipitated by addition of excess chloride:

$$Ag_2SO_4(aq) + 2NaCl(aq) \rightarrow 2AgCl(s) + Na_2SO_4(aq)$$

- **a)** How many grams of silver chloride can be formed when 35.0 mL of a 0.100 M Ag_2SO_4 solution is reacted with excess sodium chloride solution?
- **b)** If 22.7 mL of a silver sulfate solution of unknown concentration yields 0.985 g of AgCl upon reaction with excess sodium chloride solution, what is the concentration of the silver sulfate solution?

8. Zn metal reacts with hydrochloric acid to produce hydrogen gas and zinc(II) chloride:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

- **a)** If 15.0 g of Zn are added to excess HCl(aq), how many liters of $H_2(g)$, collected at 27 ^oC and 725 mm Hg, are produced?
- **b)** If excess Zn is added to 25.0 mL of 0.025 M HCl(aq), how many liters $H_2(g)$, collected at 27 ^oC and 725 mm Hg, can be produced?

9. Potassium permanganate and iron(II) chloride undergo an electron transfer reaction in acid solution:

$$KMnO_4(aq) + 5FeCl_2(aq) + 8HCl \rightarrow MnCl_2(aq) + 5FeCl_3(aq) + KCl(aq) + 4\ H_2O(l)$$

How many mL of 0.150 M $FeCl_2(aq)$ are needed to completely react with 13.7 mL of 0.110 M $KMnO_4$?

10. Citric acid reacts with sodium hydroxide in a proton transfer reaction:

$H_3C_6H_5O_7(aq) + 3NaOH(aq) \rightarrow 3H_2O(l) + Na_3C_6H_5O_7(aq)$

a) How many mL of 0.125 M NaOH(aq) are required to completely react with 25.0 mL of 0.0695 M citric acid?

b) If 37.5 mL of 1.25 M NaOH(aq) is needed to completely react with 22.5 mL of a citric acid solution, what is the concentration of the citric acid solution?

ANSWERS:

1. 330 g

2. **a)** 18.9 g **b)** 6.38 g **c)** 21.6 g

3. **a)** 64.3 g **b)** 9.34 g

4. **a)** 20.5 L **b)** 4.97 g

5. **a)** 31.3 g **b)** 15.1 g **c)** 169 g **d)** 11 g

6. **a)** 405 g **b)** 34.8 L

7. **a)** 1.00 g **b)** 0.151 M

8. **a)** 5.92 L **b)** 8.07 mL

9. 50.2 mL

10. **a)** 41.7 mL **b)** 0.694 M