Chapter 4: Amounts, Stoichiometry and Reactions in Aqueous Solutions

Sections 4.1 - 4.2: Introduction to Molarity

Molarity refers to the concentration of species in solution. The species is called **solute**. The solute dissolves in a liquid called the **solvent**. Hence, a **solution** is composed of solute and solvent. Qualitatively, a dilute solution has a low concentration of solute and a concentrated solution has a high concentration of solute. The terms dilute and concentrated are used in a comparative sense.

$$\text{Molarity} = \frac{\text{moles of solute (mol)}}{\text{volume of solution (L)}}$$

$$\text{moles (n)} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

$$\text{Molarity (M)} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)} \times \text{liters of solution (L)}}$$

Thus, mass (g) = Molarity (M) x molar mass (g/mol) x liters of solution (L)

The molarity (M) of a solution can be used to calculate:

(a) the amount of solute (moles or mass) in a given volume of solution, and

(b) the volume of solution containing a given amount (moles) of solute.

Example 1: Calculate the mass of $KMnO_4$ needed to prepare 700. mL of a 0.583 M solution of $KMnO_4$.

Molar Mass ($KMnO_4$) = 39.1 + 54.9 + 4 x 16.0 = 158.0 g/mol ($KMnO_4$).

$$\text{Molarity} = \frac{\text{mass of solute (g)}}{\text{molar mass (g/mol)} \times \text{volume of solution (L)}}$$

Hence, Mass (g) = Molarity (mol/L) x molar mass (g/mol) x volume (L)

Note that the volume must be expressed in liter (L) and 1 L = 1000 mL

$$\text{Mass (g)} = 0.583 \text{ mol/L} \times 158.04 \text{ g/mol} \times 0.700 \text{ L} = 64.5 \text{ g}$$

Therefore, 64.5 g of $KMnO_4$ dissolved in 700.mL of water produces a 0.583 M $KMnO_4$ solution.

Example 2: Calculate the volume in mL of a 0.215 M H_2SO_4 solution containing 0.948 g of H_2SO_4.

Molar Mass (H_2SO_4) = 2 x 1.0 + 32.1 + 4 x 16.0 = 98.1 g/mol (H_2SO_4).

$$\text{Molarity (mol/L)} = \frac{\text{mass of solute (g)}}{\text{molar mass (g/mol)} \times \text{volume of solution (L)}}$$

Hence,

$$\text{Volume of solution (L)} = \frac{\text{mass of solute (g)}}{\text{molar mass (g/mol)} \times \text{molarity (mol/L)}}$$

$$\text{Volume of solution (L)} = \frac{0.948 \text{ g}}{98.1 \text{ g/mol} \times 0.215 \text{ mol/L}} = 0.0449 \text{ L}$$

Hence, the volume of solution is 44.9 mL.

Example 3: A 75.0 mL solution of $AgNO_3$ contains 0.0385 moles of $AgNO_3$. Calculate the molarity of the solution.

$$\text{Molarity (M)} = \frac{\text{moles of solute (n)}}{\text{liters of solution (L)}} = \frac{0.0385 \text{ mol}}{0.075 \text{ L}}$$

Molarity = 0.513 M

In Section 4.2, practice the Interactive Problems.

Sections 4.3 - 4.4: Molarity of Ionic Solutions

Aqueous solutions of salts conduct electricity, showing that salts break into ions in solution. Salts break down into cations and anions. The ratio of cations and anions produced is given by the ion's chemical formula.

Polyatomic ions remain intact as ions in solution.

For example: Sodium sulfate (Na_2SO_4) dissolved in water breaks down into $Na^+_{(aq)}$ and $SO_4^{-2}{}_{(aq)}$ ions. SO_4^{-2} is a polyatomic ion. From the chemical formula Na_2SO_4, the ratio of ions in solution is: 2 $Na^+_{(aq)}$ to 1 $SO_4^{-2}{}_{(aq)}$

For example: Consider an aqueous solution of 1 M NaCl. In solution, NaCl breaks down into 1 $Na^+{}_{(aq)}$ and 1 $Cl^-_{(aq)}$ ions. The concentration of $Na^+_{(aq)}$ is 1 M and that of $Cl^-_{(aq)}$ is 1 M

For example: Now, consider a 1 M aqueous solution of Na_3PO_4. In solution, Na_3PO_4 breaks down into 3 $Na^+_{(aq)}$ and 1 $PO_4^{-3}{}_{(aq)}$. Hence, in 1 M Na_3PO_4 solution, the concentration of $Na^+_{(aq)}$ is 3 M and that of $PO_4^{-3}{}_{(aq)}$ is 1 M.

Thus, an ionic solution contains a mixture of ions, each with its own molarity.

Example: Calculate the molarity of ionic species present in 250 mL of an aqueous solution containing 1.85 g of ammonium sulfate.

Ammonium sulfate has a chemical formula: $(NH_4)_2SO_4$

The molar mass of $(NH_4)_2SO_4$ is: $2 \times 14.0 + 8 \times 1.0 + 32.1 + 4 \times 16.0 = 132.1$ g/mol

$$\text{Molarity}\ ((NH_4)_2SO_4) = \frac{\text{mass of solute (g)}}{\text{molar mass (g/mol)} \times \text{volume of solution (L)}}$$

$$\text{Molarity}\ ((NH_4)_2SO_4) = \frac{1.85\ \text{g}}{132.1\ \text{g/mol} \times 0.25\ \text{L}} = 0.056\ \text{mol/L}$$

Hence, the concentration of $(NH_4)_2SO_4$ is 0.056 M.

In solution, $(NH_4)_2SO_4$ breaks down as 2 $NH_4^+{}_{(aq)}$ and 1 $SO_4^{-2}{}_{(aq)}$

Thus, the concentration of $NH_4^+{}_{(aq)} = 2 \times 0.056 = 0.11$ M and
the concentration of $SO_4^{-2}{}_{(aq)} = 0.056$ M

In Section 4.4, practice Interactive Problems.

Sections 4.5 - 4.6: Diluting Solutions

When solutions of acids, bases and soluble ionic compounds are manufactured and transported, they are often prepared as **stock solutions**. These stock solutions are highly concentrated solutions. Different usages of these solutions often require different (typically lower) concentrations of solutes in these solutions. To prepare a solution of a substance from a stock solution we will need to dilute the stock solution.

Example: Commercially available concentrated HCl is usually a 12.0 M solution. How does one prepare 100. mL of 1.0 M HCl solution from 12.0 M HCl solution?

The obvious answer is to dilute the concentrated HCl (i.e. 12.0 M HCl) solution. But, what volume of 12.0 M HCl should be diluted? To answer this question, recall the definition of molarity.

$$\text{Molarity (M)} = \frac{\text{moles of solute (n)}}{\text{liters of solution (L)}}$$

Hence, moles of solute (n) = Molarity (M) x liters of solution (L)

This equation can be written as: moles of solute (n) = $M_i \times V_i$. The subscript "i" stands for "initial".

Thus, M_i = initial molarity
and V_i = initial volume

Now, add water to the stock solution. This changes the concentration from M_i to M_f and the volume from V_i to V_f. The subscript "f" stands for "final".

Thus, M_f = final molarity
and V_f = final volume

Because the number of moles of solute has not changed during dilution, moles of solute (n) = $M_f \times V_f$, and,

$$M_i \times V_i = M_f \times V_f$$

In the above example:

M_i = 12.0 M
V_i = To be calculated
M_f = 1.0 M
V_f = 100. mL

$$12.0\ \text{M} \times V_i = 1.0\ \text{M} \times 100.\ \text{mL}$$

$$V_i = \frac{1.0\ \text{M} \times 100.\ \text{mL}}{12.0\ \text{M}} = 8.3\ \text{mL}$$

(in the correct number of significant figures)

Note: You can use any volume units as long as V_i and V_f have the same units.

Hence, to prepare the diluted solution, take a clean 100 mL volumetric flask, add approximately 25 mL distilled water, then, <u>carefully</u> and <u>slowly</u>, add 8.3 mL 12.0 M HCl solution. Add more water to bring the total volume to 100. mL.

Example: A stock solution of 15.8 M HNO_3 is given. How many mL of the stock solution are required to make 0.250 liters of 0.12 M HNO_3?

$$M_i \times V_i = M_f \times V_f$$

$M_i = 15.8$ M
$V_i =$ To be calculated
$M_f = 0.12$ M
$V_f = 0.250$ L

$$15.8 \text{ M} \times V_i = 0.12 \text{ M} \times 0.250 \text{ L}$$

$$V_i = \frac{0.12 \text{ M} \times 0.250 \text{ L}}{15.8 \text{ M}}$$

$$V_i = 0.0019 \text{ L} = 1.9 \text{ mL}$$

In Section 4.6, practice Interactive Problems.

Sections 4.7 - 4.8: Theoretical Yield, Limiting Reactant and Percent Yield

Consider the chemical reaction:

$$2\ Sb_{(s)} + 3\ Cl_{2(g)} \rightarrow 2\ SbCl_{3(s)}$$

The reaction equation provides a recipe for the preparation of antimony chloride. It indicates that one needs 3 moles of chlorine gas for every two moles of antimony solid to form 2 moles of antimony chloride.

Whenever reactants are present in the relative amount of 3 moles of chlorine for 2 moles of antimony, we say that the reactants are present in the **stoichiometric amount**.

For example, if we are given 0.4 mol Sb with 0.6 mol Cl_2, we have reactants present in the stoichiometric amount (ratio of 0.4 to 0.6 = ratio of 2 to 3).

<u>Whenever we have reactants present in the stoichiometric amount, we can use any of the reactants to calculate how many moles of product are formed from a given amount of reactants.</u>

The maximum mass of products that can be formed from a given amount of reactants, assuming the reaction is complete, is called the **Theoretical Yield**.

For example: If we react 2 moles of Sb with 3 moles of Cl_2, we obtain 2 moles of $SbCl_3$. The mass of 2 moles of $SbCl_3$ is equal to 2 x (121.8 + 3 x 35.5) = 456.6 g. We say that when the reaction is carried out with 2 moles of Sb and 3 moles of Cl_2, the theoretical yield is equal to 456.6 g.

When the reaction is carried out with 4 moles Sb and 6 moles Cl_2, the theoretical yield is equal to 913.2 g. When the reaction is carried out with 0.2 moles Sb and 0.3 moles Cl_2, the theoretical yield is 45.7 g

When the reactants are present in the stoichiometric amount and the reaction is complete, there is no reactant left at the end of the reaction.

<u>What happens if reactants are not present in the stoichiometric amount</u>? When reactants are not present in the stoichiometric amount, then, some of the reactants are completely consumed in the reaction (**limiting reactant**) and some reactants remain at the end of the reaction (**excess reactant**).

For example: Consider the reaction between antimony and chlorine gas, again.

$$2\ Sb_{(s)} + 3\ Cl_{2(g)} \rightarrow 2\ SbCl_{3(s)}$$

When the reaction is started with 4 moles Sb and 6 moles Cl_2, we said above that the reactants are present in the stoichiometric amount.

When the reaction is started with 4 moles Sb and 7 moles Cl_2, then, at the end of the reaction, there will be one mole of Cl_2 left, since to completely react 4 moles Sb, only requires 6 moles Cl_2. We say that Cl_2 is in excess and Sb is the limiting reactant (it is the one limiting the amount of product that can be formed).

When the reaction is started with 5 moles Sb and 6 moles Cl_2, then, at the end of the reaction, there will be one mole of Sb left, since to completely react 6 moles Cl_2, only requires 4 moles Sb. We say that Sb is in excess and Cl_2 is the limiting reactant (it is the one limiting the amount of product that can be formed).

<u>Remember that the theoretical yield is always determined by the amount of limiting reactant</u>.

Example: Assume that the reaction between antimony and chlorine is carried out with 1.5 mol Sb and 1.5 mol Cl_2. Determine the limiting reactant, the excess reactant and the theoretical yield.

To determine which reactant is the limiting reactant, we calculate the amount of product formed from each of the given amounts of reactants. The limiting reactant is the reactant that produces the least amount of products.

1) Write the balanced chemical reaction.

$$2\ Sb_{(s)} + 3\ Cl_{2(g)} \rightarrow 2\ SbCl_{3(s)}$$

2) Determine the amounts of product formed from each reactant:

How much $SbCl_3$ is formed with 1.5 mol Sb?

Using the mole ratio derived from the balanced chemical reaction, we write:

$$\text{moles}\,(SbCl_3) = \frac{2\ \text{mol}\ SbCl_3}{2\ \text{mol}\ Sb}\ \text{x}\ 1.5\ \text{mol}\ Sb\ =\ 1.5\ \text{mol}\ SbCl_3$$

How much $SbCl_3$ is formed with 1.5 mol Cl_2?

Using the mole ratio derived from the balanced chemical reaction, we write:

$$\text{moles}\,(SbCl_3) = \frac{2\ \text{mol}\ SbCl_3}{3\ \text{mol}\ Cl_2}\ \text{x}\ 1.5\ \text{mol}\ Cl_2\ =\ 1.0\ \text{mol}\ SbCl_3$$

3) The limiting reactant is the reactant producing the smallest amount of product.

The limiting reactant is Cl_2. Hence, the excess reactant is Sb.

The theoretical yield is the amount of products formed from the limiting reactant. 1.5 mol Cl_2 produce 1.0 mol $SbCl_3$. The mass of 1.0 mol $SbCl_3$ is equal 228.3 g.

Hence, the theoretical yield is 228.3 g.

We found that the limiting reactant is chlorine gas and the excess reactant is antimony. So we know that all chlorine (1.5 mol) is consumed in the reaction. Now, let us calculate how much antimony is left after the reaction.

First, we calculate how much antimony is used to fully react with 1.5 mol Cl_2. We do this using the mole ratio.

$$\text{moles}\,(Sb) = \frac{2\ \text{mol}\ Sb}{3\ \text{mol}\ Cl_2}\ \text{x}\ 1.5\ \text{mol}\ Cl_2\ =\ 1.0\ \text{mol}\ Sb$$

Hence, 1.0 mol Sb is used out of an initial amount of 1.5 mol Sb. We conclude that 0.5 mol Sb remains after the reaction.

Experimental Yield, Theoretical Yield and Percent Yield

For a variety of reasons, the vast majority of chemical reactions carried out in laboratories or in chemical plants do not produce the maximum amount of product possible (theoretical yield).

Remember the **Theoretical Yield** is calculated from the amounts of reactants present and from the balanced chemical reaction.

The amount of products actually obtained from a given amount of reactants is called the **Experimental Yield**. The experimental yield is thus always smaller than the theoretical yield.

The **Percent Yield** is defined as the ratio of experimental yield to theoretical yield multiplied by 100.

$$\%\ \text{Yield} = 100\ \text{x}\ \frac{\text{Experimental Yield}}{\text{Theoretical Yield}}$$

Example: 68.5 g of CO $_{(g)}$ is reacted with 8.60 g of $H_{2\,(g)}$ to form 35.7 g of CH_3OH (methanol). Determine if any reactant is in excess. Calculate the theoretical yield and the percent yield.

1) Write the balanced chemical reaction.

$$CO_{(g)} + 2\,H_{2\,(g)} \rightarrow CH_3OH_{(l)}$$

2) Calculate the amount of methanol formed from 68.5 g CO.

Use the method of conversion factors.

$$g\,(CH_3OH) = 68.5\ g\ \ CO\ \text{x}\ \frac{1\ mol\ (CO)}{28.0\ g\ (CO)}\ \text{x}\ \frac{1\ mol\ (CH_3OH)}{1\ mol\ (CO)}\ \text{x}\ \frac{32.0\ g\ (CH_3OH)}{1\ mol\ (CH_3OH)}$$

$$g\ (CH_3OH) = 78.2\ g\ (CH_3OH)$$

3) Calculate the amount of methanol formed from 8.60 g H_2.

Use the method of conversion factors.

$$g(CH_3OH) = 8.60\ g\ H_2 \times \frac{1\ mol\ (H_2)}{2.0\ g\ (H_2)} \times \frac{1\ mol\ (CH_3OH)}{2\ mol\ (H_2)} \times \frac{32.0\ g\ (CH_3OH)}{1\ mol\ (CH_3OH)}$$

$$g\ (CH_3OH) = 68.8\ g\ (CH_3OH)$$

4) The mass of methanol formed with 8.60 g hydrogen is less than that formed with 68.5 g carbon monoxide. Hence, hydrogen is the limiting reactant and carbon monoxide is the excess reactant. The theoretical yield is the amount of methanol formed with 8.6 mol H_2 (limiting reactant). The theoretical yield is equal to 68.8 g.

5) The percent yield is then expressed as:

$$\%\ Yield = 100 \times \frac{35.7\ g}{68.8\ g} = 51.9\ \%$$

In Section 4.8, practice Interactive Problems.

Section 4.9: Volumetric Analysis

An acid – base reaction is often called a neutralization reaction. When just enough base is added to the acid, the acid is neutralized.

Acids and bases are classified as strong and weak. Hence, the nature of neutralization depends upon the strengths of acids and bases.

The acid – base reactions are classified into three categories:

Category 1: Strong acid – Strong base
Category 2: Weak acid – Strong base
Category 3: Strong acid – Weak base

In all three categories, there are a few common steps that should be applied to perform a volumetric analysis.

Category 1: Strong acid – Strong base

Step 1: List the ions present in a combined solution.
For example: In the reaction of HCl and NaOH,
The ions present in a combined solution will be H^+, Cl^-, Na^+, OH^-

Step 2: Figure out what reactions would occur with these ions and determine the net ionic equation. In the example, there are two possibilities.

H^+ combines with OH^- to form H_2O
Na^+ combines with Cl^- to form NaCl

The net ionic equation is: $H^+ + OH^- \rightarrow H_2O$

This is the net ionic equation, in general, for a reaction between a strong acid and a strong base.

Step 3: Perform volumetric calculation using the relationship: $M_i \times V_i = M_f \times V_f$
This equation can be written as

$$M_{H+} \times V_{H+} = M_{OH-} \times V_{OH-}$$

since the neutralization reaction is carried out under stoichiometric conditions only when the number of moles of OH^- is equal to the number of moles of H^+. (Remember moles = molarity x volume)

Example: Calculate the volume in mL of 0.100 M HCl needed to neutralize 50.0 mL of 0.250 M NaOH.

$M_{H+} = 0.100$ M
V_{H+} = To be calculated
$M_{OH-} = 0.250$ M
$V_{OH-} = 50.0$ mL

$$0.100 \text{ M} \times V_{OH-} = 0.250 \text{ M} \times 50.0 \text{ mL}$$

$$V_{OH-} = \frac{0.250 \text{ M} \times 50.0 \text{ mL}}{0.100 \text{ M}}$$

$$V_{OH-} = 125 \text{ mL}$$

Hence, 125 mL of 0.100 M HCl are needed to neutralize 50.0 mL of 0.250 M NaOH.

A titration is an analytical technique used to determine the volume or concentration of an acid or base. Volumetric analysis is a method involving calculations based on titration. A titration can be performed and the volumetric analysis can be done between

Category 1: Strong acid – Strong base
Category 2: Weak acid – Strong base
Category 3: Strong acid – Weak base

Considering the scope of the topic only category 1 is discussed in this chapter. For a more thorough discussion of titration reactions, see Chapter 16.

Note: The example below is not discussed in depth on the DVD.

Example: Let us determine the volume of a 0.250 M $Ba(OH)_2$ aqueous solution needed to neutralize 15.0 mL of a 0.175 M HCl solution.

We will solve this problem by two methods. The first method is similar to the one used in the example. The second method is identical to that used in stoichiometric problems in Chapter 3.

Method A:

The balanced chemical reaction is written as: $Ba(OH)_2 + 2\ HCl \rightarrow BaCl_2 + 2\ H_2O$

The net ionic reaction is written as: $H^+ + OH^- \rightarrow H_2O$

At the stoichiometric point, moles (H^+) = moles (OH^-)

or using molarities,

$$\text{Volume (HCl)} \times \text{Molarity } (H^+) = \text{Volume } (Ba(OH)_2) \times \text{Molarity } (OH^-)$$

$$\text{Hence, Volume } (Ba(OH)_2) = \text{Volume (HCl)} \times \frac{\text{Molarity } (H^+)}{\text{Molarity } (OH^-)}$$

Volume (HCl) = 0.015 L

Molarity (H^+) = Molarity (HCl) = 0.175 M (since 1 mol H^+ per 1 mol HCl)

$$\text{Molarity } (OH^-) = \frac{\text{moles } (OH^-)}{1\text{L}} = \frac{\text{moles } (Ba(OH)_2)}{1\text{L}} \times \frac{2 \text{ mol } (OH^-)}{1 \text{ mol } (Ba(OH)_2)}$$

$$\text{Molarity } (OH^-) = 2 \times \text{Molarity } (Ba(OH)_2) = 2 \times 0.250 \text{ M} = 0.500 \text{ M}$$

$$\text{Hence, Volume } (Ba(OH)_2) = 0.015 \text{ L} \times \frac{0.175 \text{ M}}{0.500 \text{ M}} = 0.00525 \text{ L} = 5.25 \text{ mL}$$

Method B:

Solve this problem as a normal stoichiometry problem. The balanced chemical reaction is written as:

$$Ba(OH)_2 + 2\ HCl \rightarrow BaCl_2 + 2\ H_2O$$

We want to calculate $V(Ba(OH)_2)$, the volume of $Ba(OH)_2$ solution necessary to completely react 15.0 mL of 0.175 M HCl. Use the method of conversion factors.

$$V(Ba(OH)_2) = 0.0150\ L\ (HCl) \times \frac{0.175\ mol\ (HCl)}{1L} \times \frac{1\ mol\ (Ba(OH)_2)}{2\ mol\ (HCl)} \times \frac{1\ L}{0.250\ mol\ (Ba(OH)_2)} \times \frac{1000\ mL}{1\ L}$$

Hence, $V(Ba(OH)_2) = 0.0150 \times 0.175 \times (1/2) \times (1/0.250) \times (1000)$ mL

$V(Ba(OH)_2) = 5.25$ mL

Note what conversion factors are used:

First, we write the given volume of HCl (0.0150 L)

Second, we transform L of HCl into mol (HCl) using HCl molarity (0.175 M)

Third, we use the mole ratio from the balanced reaction: $\frac{1\ mol\ (Ba(OH)_2)}{2\ mol\ (HCl)}$

to convert moles (HCl) to moles ($Ba(OH)_2$).

Fourth, we convert moles ($Ba(OH)_2$) to liters ($Ba(OH)_2$) using the molarity of the $Ba(OH)_2$ solution.

Use whatever method (A or B) works best for you!!

Interactive Problems are included in this section for your practice and understanding.

Sections 4.10 - 4.11: Gravimetric Analysis

Gravimetric analysis is a chemical analysis method based on the measurement of masses. It can be used, for instance, in combination with precipitation reactions to determine the amount of a species present in solution. We first isolate the precipitate by filtration and drying and subsequently weigh it. Using this mass and stoichiometry relationships we can determine the amount (mass or moles) of species present in solution.

Example 1: Calculate the mass of $Fe(OH)_3$ formed when 50.0 mL of a 0.300 M $NaOH_{(aq)}$ is added to 30.0 mL of 0.250 M $Fe(NO_3)_{3\ (aq)}$ solution.

Step 1: Write the balanced chemical reaction:

$$3\ NaOH_{(aq)} + Fe(NO_3)_{3\ (aq)} \rightarrow Fe(OH)_{3\ (s)} + 3\ NaNO_{3\ (aq)}$$

Step 2: Write the net ionic equation:

$$3\ OH^-_{(aq)} + Fe^{+3}_{(aq)} \rightarrow Fe(OH)_{3\,(s)}$$

Step 3: Determine the number of moles of each reactant:

$$\text{Moles } (OH^-) = \text{Moles } (NaOH) = 0.300\ M \times 0.050\ L = 0.0150\ mol$$

$$\text{Moles } (Fe^{+3}) = \text{Moles } (Fe(OH)_3) = 0.250\ M \times 0.030\ L = 0.00750\ mol$$

Step 4: Determine the limiting reactant. To do so calculate the number of moles of product formed from the given number of moles of each reactant.

From the hydroxide ion reactant.

$$\text{Moles } (Fe(OH)_3) = 0.0150\ mol\ (OH^-) \times \frac{1\ mol\ (Fe(OH)_3)}{3\ mol\ (OH^-)} = 0.00500\ mol$$

From the iron(III) ion reactant.

$$\text{Moles } (Fe(OH)_3) = 0.00750\ mol\ (Fe^{+3}) \times \frac{1\ mol\ (Fe(OH)_3)}{1\ mol\ (Fe^{+3})} = 0.00750\ mol$$

Hence, the limiting reactant is the hydroxide and the theoretical product is 0.00500 mol ($Fe(OH)_3$). The molar mass of iron(III) hydroxide is 106.9 g/mol. Hence, the mass of iron hydroxide collected should be 0.535 g.

Example 2: A liter sample of polluted water is analyzed for lead, believed to be in the lead(II) ionic form. This is done by adding excess sodium sulfate to the polluted water sample. The mass of $PbSO_4$ produced is 300.0 mg. Calculate the mass of Pb in one liter solution.

All the lead in water reacts with the sulfate ions to form a $PbSO_4$ precipitate.

The balanced chemical reaction is written as:

$$Na_2SO_4 + Pb(NO_3)_2 \rightarrow PbSO_4 + 2\ NaNO_3$$

The net ionic reaction is written as:

$$SO_4^{-2} + Pb^{+2} \rightarrow PbSO_4$$

$$\text{Molar mass } (PbSO_4) = 303.3\ g/mol$$

$$\text{Molar mass } (Pb) = 207.2\ g/mol$$

Using the method of conversion factors, we transform the mass of $PbSO_4$ into the mass of Pb.

$$\text{Mass (Pb)} = 300.0\ \text{mg PbSO}_4 \times \frac{1\ \text{g (PbSO}_4)}{1000\ \text{mg (PbSO}_4)} \times \frac{1\ \text{mol (PbSO}_4)}{303.3\ \text{g (PbSO}_4)} \times \frac{1\ \text{mol (Pb)}}{1\ \text{mol (PbSO}_4)} \times \frac{207.2\ \text{g (Pb)}}{1\ \text{mol (Pb)}}$$

Mass (Pb) = 0.2049 g Pb = 204.9 mg Pb

In Section 4.11, practice Interactive Problems.

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