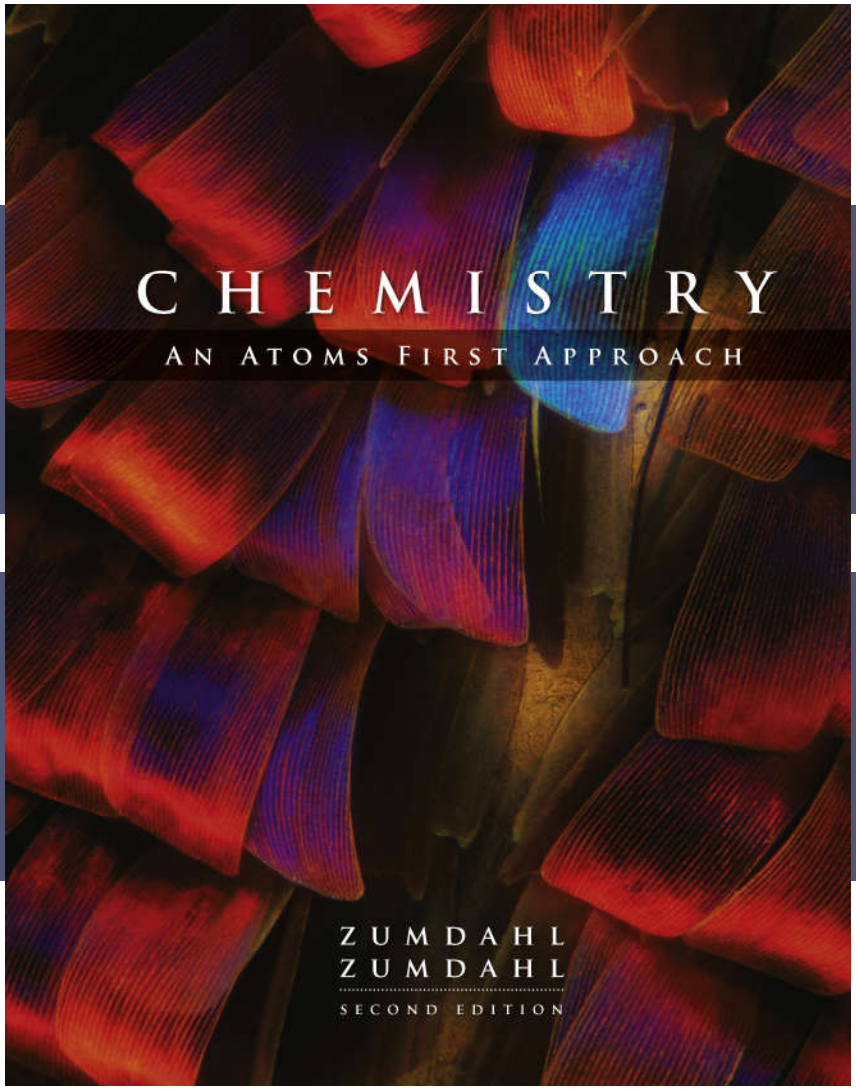

Chapter 6

Types of Chemical Reactions and Solution Stoichiometry

Chapter 6
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- (6.1) Water, the common solvent
- (6.2) The nature of aqueous solutions: Strong and weak electrolytes
- (6.3) The composition of solutions
- (6.4) Types of chemical reactions
- (6.5) Precipitation reactions
- (6.6) Describing reactions in solution
- (6.7) Stoichiometry of precipitation reactions
- (6.8) Acid–base reactions

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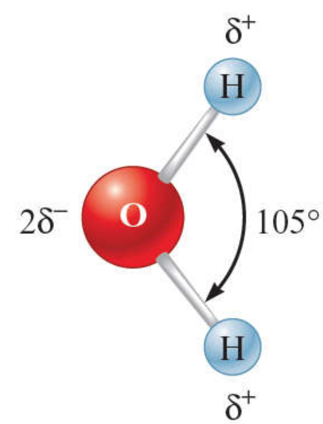
- (6.9) Oxidation–reduction reactions
- (6.10) Balancing oxidation–reduction equations

Questions to Consider

- Why is sugar solution not used as an electrolyte?
- What kind of reaction would one classify photosynthesis as?

Importance of Water

- An integral part of many life-sustaining reactions
- Cooling effect of water is used to reduce the temperature of:
  - Automobile engines
  - Nuclear power plants
  - Many industrial processes
- Used as a means of transportation
- Vital to the growth of crops
- It is a **polar molecule**



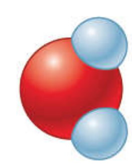
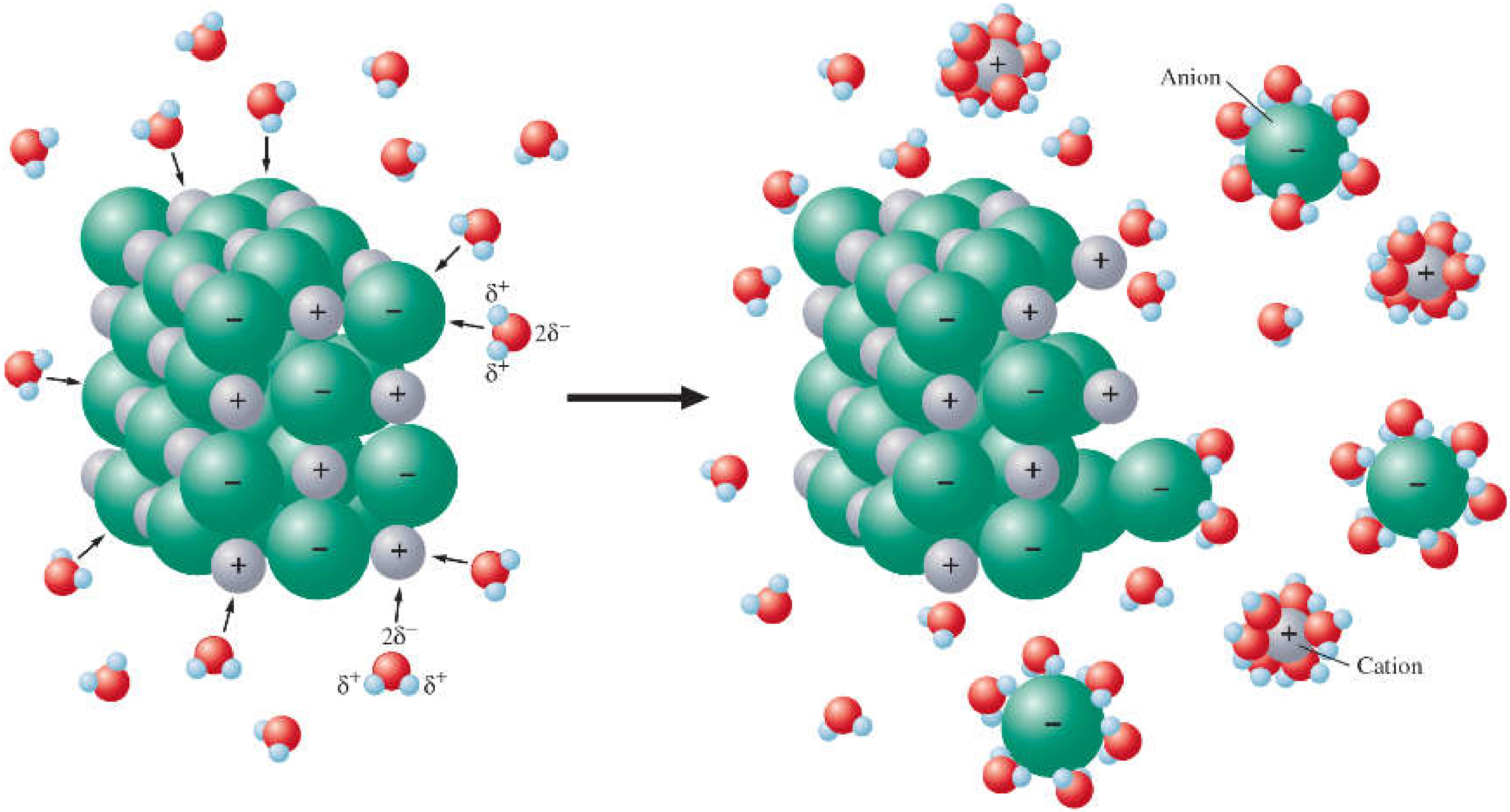


Figure 6.2 - Water Dissolving a Salt



Hydration

- The interaction between water molecules and the ions of a salt, causing dissolution
  - Salts are broken up into individual cations and anions
- Consider the hydration of ammonium nitrate

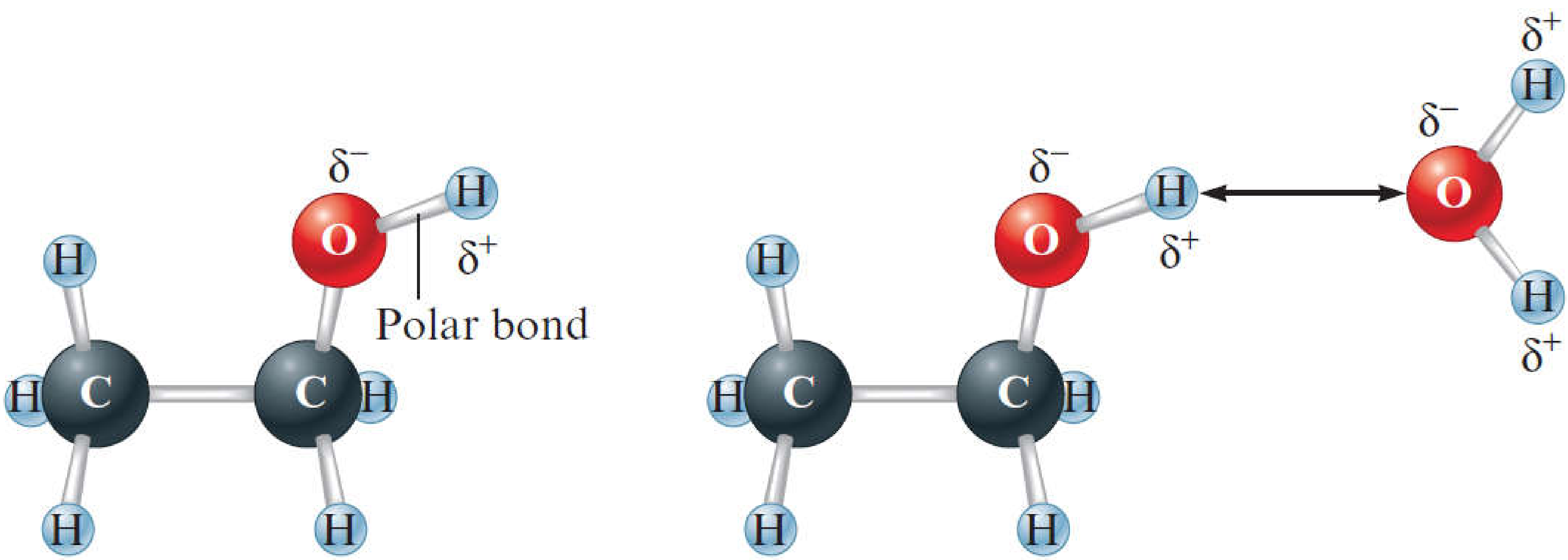
$$NH_4NO_3(s) \xrightarrow{H_2O(l)} NH_4^+(aq) + NO_3^-(aq)$$

  - (*aq*) indicates the hydration of ions

Solubility

- Varies among different substances
- Differences in solubility of ionic compounds depend on:
  - Attractions of ions to one another
  - Attraction of ions to water molecules
- When ionic solids dissolve in water, ions undergo hydration and are dispersed
- Many non-ionic solids are soluble
- Animal fat is not soluble in pure water

Figure 6.3 - Molecular Structures of Ethanol and Water



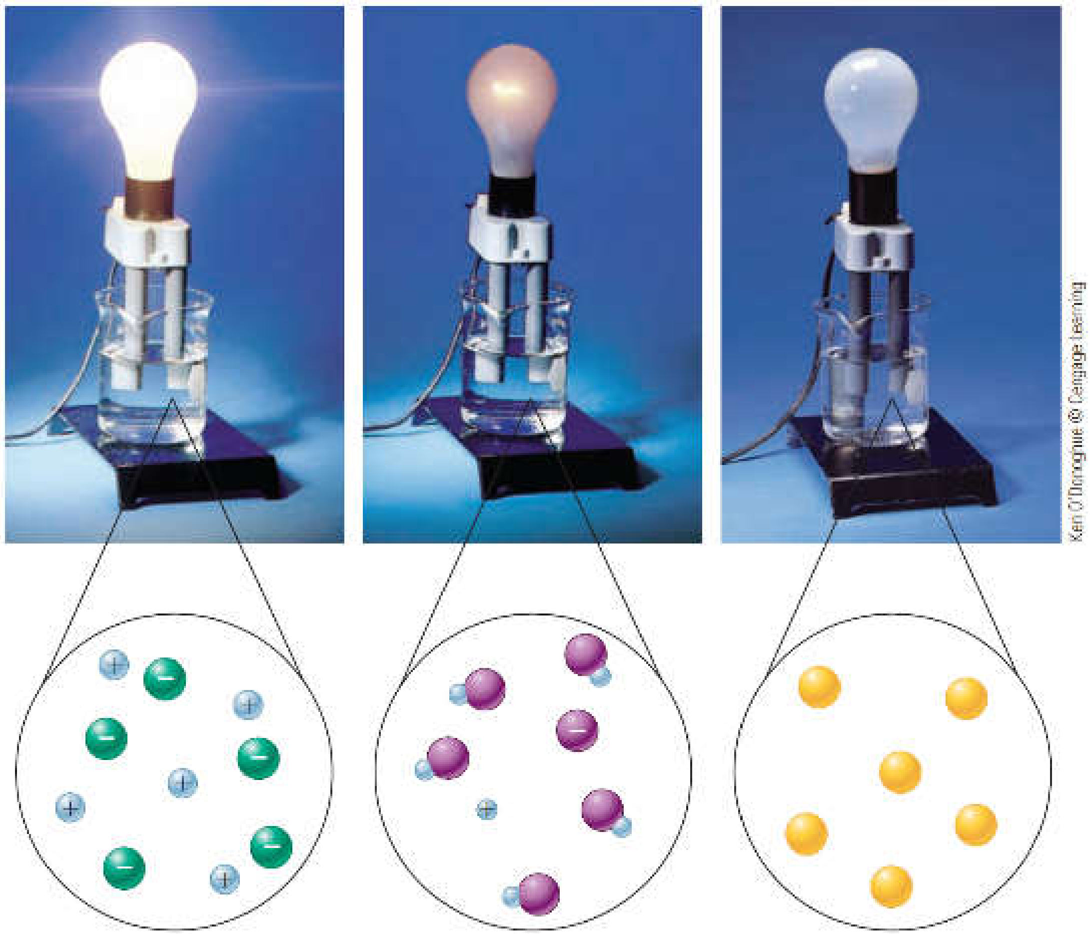
Nature of Aqueous Solutions

- **Solute**: Substance dissolved
- **Solvent**: Liquid water
- **Electrical conductivity**: Ability of a solution to conduct electricity
  - Solutions with high electrical conductivity are **strong electrolytes**
  - Solutions with low electrical conductivity are **weak electrolytes**
  - **Nonelectrolytes** do not conduct electricity

Conductivity of Solutions

- According to Svante Arrhenius, the conductivity of a solution depends on the number of ions present
  - Strong electrolytes readily produce ions in aqueous solution
  - Weak electrolytes produce a relatively lesser number of ions in aqueous solution
  - Nonelectrolytes are those that do not produce ions in aqueous solution

Figure 6.4 - Electrical Conductivity of Aqueous Solutions



Strong Electrolytes

- Substances that are completely ionized upon dissolution
- Classified as:
  - Soluble salts
    - Sodium chloride produces Na^+ and Cl^- ions when dissolved
  - **Strong acids**
    - Undergo ionization reactions to produce H^+ ions
    - Represented in the aqueous form in equations
    - Completely dissociate into ions
      - H_2SO_4 produces two H^+ ions under certain conditions

$$H_2SO_4 \xrightarrow{H_2O} H^+(aq) + HSO_4^-(aq)$$

Strong Electrolytes

- **Strong bases**
  - Soluble ionic compounds that possess OH^- ions
  - Bitter to taste
  - Slippery to touch

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

Weak Electrolytes

- Substances that produce lesser ions when dissolved
- Classified as:
  - **Weak acids**
    - Acids that produce a low amount of ions in aqueous solution

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

  - **Weak bases**
    - Bases that produce a low amount of ions in aqueous solution

$$NH_3(aq) + H_2O(l) \longrightarrow NH_4^+(aq) + OH^-(aq)$$

Nonelectrolytes

- Substances that do not produce any ions when dissolved
- Consider ethanol (C_2H_5OH) dissolved in water
  - Molecules are dispersed, but do not break up into ions
  - Resulting solution is not capable of conducting electricity

Stoichiometry of Chemical Reactions

- Performing stoichiometric calculations requires the following information:
  - The nature of the reaction
    - Depends on the exact nature of chemicals when dissolved
  - The amounts of chemicals present in the solutions
    - Expressed as concentrations

Molarity

- Defined as the moles of solute per volume of solution in liters

$$M = \text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

- A solution that is 1.0 molar (1.0 *M*) contains 1.0 mole of solute per liter of solution

Interactive Example 6.1 - Calculation of Molarity I

- Calculate the molarity of a solution prepared by dissolving 11.5 g of solid NaOH in enough water to make 1.50 L of solution

Solution

- Objective
  - To find the molarity of NaOH solution
- Information available
  - 11.5 g NaOH
  - 1.50 L solution
- Information needed
  - Moles solute
  - $\text{Molarity} = \dfrac{\text{Mol solute}}{\text{L solution}}$

Solution

- Step 1 - Determine the moles of NaOH (40.00 g/mol)

$$11.5\ \cancel{\text{g NaOH}} \times \frac{1\ \text{mol NaOH}}{40.00\ \cancel{\text{g NaOH}}} = 0.288\ \text{mol NaOH}$$

- Step 2 - Determine the molarity of the solution

$$\text{Molarity} = \frac{\text{mol solute}}{\text{L solution}} = \frac{0.288\ \text{mol NaOH}}{1.50\ \text{L solution}} = 0.192\ M\ \text{NaOH}$$

Determining the Number of Moles in a Solute

- A sample is taken
- The number of moles present in the sample are determined
  - The molarity of the solution is multiplied by the volume of the sample

$$\text{Liters of solution} \times \text{molarity} = \cancel{\text{liters of solution}} \times \frac{\text{moles of solute}}{\cancel{\text{liters of solution}}}$$

$$= \text{moles of solute}$$

Interactive Example 6.4 - Concentration of Ions II

- Calculate the number of moles of Cl^- ions in 1.75 L of 1.0×10^{-3} *M* $ZnCl_2$

Solution

- Objective
  - To find the moles of Cl^- ion in the solution
- Information available
  - 1.0×10^{-3} M $ZnCl_2$
  - 1.75 L
- Information required
  - Balanced equation for dissolving $ZnCl_2$

Solution

- Step 1 - State the balanced equation for dissolving the ions

$$ZnCl_2(s) \xrightarrow{H_2O} Zn^{2+}(aq) + 2Cl^-(aq)$$

- Step 2 - Calculate the molarity of the solution

$$2 \times (1.0 \times 10^{-3} M) = 2.0 \times 10^{-3} M \ Cl^-$$

- Step 3 - Determine the moles of Cl^-

$$1.75 \text{ L solution} \times 2.0 \times 10^{-3} M \ Cl^- = 1.75 \ \cancel{\text{L solution}} \times \frac{2.0 \times 10^{-3} \text{ mol } Cl^-}{\cancel{\text{L solution}}}$$

$$= 3.5 \times 10^{-3} \text{ mol } Cl^-$$

Dilution

- The addition of water to a concentrated solution in order to achieve the desired molarity
- Pieces of apparatus used in the process of dilution
  - Volumetric (transfer) pipettes
    - Specific sizes - 5 mL, 10 mL, 25 mL, etc
  - Measuring pipettes
    - Used to measure volumes for which a volumetric pipette is not available

Interactive Example 6.7 - Concentration and Volume

- What volume of 16 *M* sulphuric acid must be used to prepare 1.5 L of a 0.10 *M* H_2SO_4 solution?

Solution

- Objective
  - To find the volume of H_2SO_4 required to prepare the solution
- Information available
  - 1.5 L of 0.10 *M* H_2SO_4 is required
  - 16 *M* H_2SO_4 is available
- Information needed
  - Moles of H_2SO_4 in the required solution

Solution

- Step 1 - Determine the moles of H_2SO_4 required

$$M \times V = \text{mol}$$

$$1.5 \text{ L solution} \times \frac{0.10 \text{ mol } H_2SO_4}{\text{L solution}} = 0.15 \text{ mol } H_2SO_4$$

- Step 2 - Determine the volume of 16 *M* H_2SO_4

$$V \times \frac{16 \text{ mol } H_2SO_4}{\text{L solution}} = 0.15 \text{ mol } H_2SO_4$$

Solution

$$V = \frac{0.15\ \cancel{\text{mol H}_2\text{SO}_4}}{\dfrac{16\ \cancel{\text{mol H}_2\text{SO}_4}}{1\ \text{L solution}}} = 9.4 \times 10^{-3}\ \text{L or } 9.4\ \text{mL solution}$$

- To make 1.5 L of 0.10 *M* H_2SO_4 using 16 *M* H_2SO_4, 9.4 mL of concentrated acid must be diluted with 1.5 L of water

Moles of the Solute are not Altered by Dilution

- An alternate way to expresses this concept is by the following equation:

$$M_1V_1 = M_2V_2$$

  - M_1 and V_1 represent the molarity and volume of the solution before dilution
  - M_2 and V_2 represent the molarity and volume of the diluted solution
- Validating this concept:

$$\begin{aligned} M_1 \times \mathrm{V}_1 &= \text{mol solute before dilution} \\ &= \text{mol solute after dilution} = M_2 \times V_2 \end{aligned}$$

Concept Check

- Which of the following solutions contains the greatest number of ions?
  - 400.0 mL of 0.10 M NaCl
  - 300.0 mL of 0.10 M $CaCl_2$
  - 200.0 mL of 0.10 M $FeCl_3$
  - 800.0 mL of 0.10 M sucrose

Concept Check

- A 0.50 *M* solution of sodium chloride in an open beaker sits on a lab bench. Which of the following would decrease the concentration of the salt solution?
  - (a) Adding water to the solution
  - (b) Pouring some of the solution down the sink drain
  - (c) Adding more sodium chloride to the solution
  - (d) Letting the solution sit out in the open air for a couple of days

Types of Solution Reactions

- Precipitation reactions
- Acid–base reactions
- Oxidation–reduction reactions

Precipitation Reaction

- A reaction in which the mixing of two solutions results in the formation of an insoluble substance which separates from the solution
  - The solid formed is called a **precipitate**
- Identification of the precipitate requires knowledge of the chemical composition of the solutions

Section 6.5
Precipitation Reactions

Figure 6.13 - Formation of a Precipitate

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Identification of the Precipitate

- Consider a mixture containing aqueous solutions of potassium chromate and barium nitrate
  - Identify the reactants
    - K_2CrO_4 (*aq*)
    - $Ba(NO_3)_2$ (*aq*)
  - Recognize that when a solid dissolves in water, the ions are separated
    - K_2CrO_4 (*aq*) contains K^+ and CrO_4^{2-} ions
    - $Ba(NO_3)_2$ (*aq*) contains separated Ba^{2+} and NO^{3-} ions

Identification of the Precipitate

- Therefore, the mixed solution contains the following ions:
  - K^+, CrO_4^{2-}, Ba^{2+}, and NO_3^-
- The compound formed must have a zero net charge
  - Both anions and cations are present in the product
- Most ionic materials contain only two types of ions
  - One type of cation and one type of anion
  - Possible combinations
    - K_2CrO_4, KNO_3, $BaCrO_4$, or $Ba(NO_3)_2$

Identification of the Precipitate

- K_2CrO_4 and $Ba(NO_3)_2$ are the reactants
- The substance formed is yellow in color
  - KNO_3 is white
  - $BaCrO_4$ is yellow
- The solid formed is almost certainly $BaCrO_4$
- If KNO_3 is mixed in the same quantity of water as in the mixed solution, it will not form a solid
- Thus,

$$K_2CrO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow BaCrO_4(s) + 2KNO_3(aq)$$

Simple Rules for the Solubility of Salts

- Most nitrate (NO_3^-) salts are soluble
- Most salts containing the alkali metal ions and the ammonium ion are soluble
- Most Cl^-, Br^-, and I^- salts are soluble (except Ag^+, Pb^{2+}, and Hg_2^{2+})
- Most sulfate salts are soluble (except $BaSO_4$, $PbSO_4$, Hg_2SO_4, and $CaSO_4$)

Simple Rules for the Solubility of Salts

- Most hydroxides are only slightly soluble (NaOH, KOH are soluble; $Ba(OH)_2$, $Ca(OH)_2$ are marginally soluble)
- Most sulfide (S^{2-}), carbonate (CO_3^{2-}), chromate (CrO_4^{2-}), and phosphate (PO_4^{3-}) salts are only slightly soluble, except for those containing the cations in Rule 2

Simple Rules for the Solubility of Salts

Solubility Rules for Ionic Compounds

Rule	Applies to	Rule	Exceptions
1	Li^+, K^+, Na^+, NH_4^+	Group IA and ammonium compounds are soluble.	
2	$C_2H_3O_2^-$, NO_3^-	Acetates and nitrates are soluble.	
3	Cl^-, Br^-, I^-	Most chlorides, bromides, and iodides are soluble	$AgCl$, Hg_2Cl_2, $PbCl_2$, $AgBr$, $HgBr_2$, Hg_2Br_2, $PbBr_2$, AgI, HgI_2, Hg_2I_2, PbI_2
4	SO_4^{2-}	Most sulfates are soluble.	$CaSO_4$, $SrSO_4$, $BaSO_4$, Ag_2SO_4, , Hg_2SO_4, $PbSO_4$
5	CO_3^{2-}	Most carbonates are insoluble.	Group IA carbonates, $(NH_4)_2CO_3$
6	PO_4^{3-}	Most phosphates are insoluble.	Group IA phosphates, $(NH_4)_3PO_4$
7	S^{2-}	Most sulfides are insoluble.	Group IA sulfides, $(NH_4)_2S$
8	OH^-	Most hydroxides are insoluble.	Group IA hydroxides, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

Rule 1 “wins” over rule 3, etc. Follow the rules in the above order.

Interactive Example 6.8 - Predicting Reaction Products

- Predict what will happen when aqueous solutions of Na_2SO_4 and $Pb(NO_3)_2$ are mixed

Solution

- The ions present in the solution before any reaction occurs are Na^{+}, SO_4^{2-}, Pb^{2+}, and NO^{3-}
- The possible salts that could form precipitates are

$$Na^{+} + SO_4^{2-} + Pb^{2+} + NO^{3-} \longrightarrow$$

- $NaNO_3$ is solluble but $PbSO_4$ is insoluble
  - $PbSO_4$ forms the precipitate

$$Na_2SO_4(aq) + Pb(NO_3)_2(aq) \longrightarrow PbSO_4(s) + 2NaNO_3(aq)$$

Concept Check

- Which of the following ions form compounds with Pb^{2+} that are generally soluble in water?
  - S^{2-}
  - Cl^{-}
  - NO_3^{-}
  - SO_4^{2-}
  - Na^{+}

Formula Equation (Molecular Equation)

- Gives the overall reaction stoichiometry but not necessarily the actual forms of the reactants and products in solution
- It is not a detailed representation

$$K_2CrO_4(aq) + Ba(NO_3)_2(aq) \longrightarrow BaCrO_4(s) + 2KNO_3(aq)$$

Complete Ionic Equation and Net Ionic Equation

- A **complete ionic equation** represents as ions all reactants and products that are strong electrolytes

$$2K^{+}(aq) + CrO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^{-}(aq) \longrightarrow BaCrO_4(s) + 2K^{+}(aq) + 2NO^{3-}(aq)$$

  - Ions that do not participate directly are called **spectator ions**

- A **net ionic equation** includes only those solution components undergoing a change

$$Ba^{2+}(aq) + CrO_4^{2-}(aq) \longrightarrow BaCrO_4(s)$$

Interactive Example 6.9 - Writing Equations for Reactions

- Write the formula equation, the complete ionic equation, and the net ionic equation for the following reaction:
  - Aqueous potassium chloride is added to aqueous silver nitrate to form a silver chloride precipitate plus aqueous potassium nitrate

Solution

- Formula equation

$$KCl(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + KNO_3(aq)$$

- Complete ionic equation

$$K^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \longrightarrow AgCl(s) + K^+(aq) + NO_3^-(aq)$$

- K^+ ↑ Spectator Ion
- NO_3^- ↑ Spectator ion
- $AgCl(s)$ ↑ Solid, not written as separate ions
- K^+ ↑ Spectator ion
- NO_3^- ↑ Spectator ion

Solution

- Cancelling the spectator ions gives the net ionic equation

$$\cancel{K^+}(aq) + Cl^-(aq) + Ag^+(aq) + \cancel{NO_3^-}(aq) \longrightarrow AgCl(s) + \cancel{K^+}(aq) + \cancel{NO_3^-}(aq)$$

$$Cl^-(aq) + Ag^+(aq) \longrightarrow AgCl(s)$$

Concept Check

- Write the correct formula equation, complete ionic equation, and net ionic equation for the reaction between cobalt(II) chloride and sodium hydroxide

Solving Stoichiometry Problems for Reactions in a Solution

- Identify the species present in the combined solution, and determine what reaction, if any, occurs
- Write the balanced net ionic equation for the reaction
- Calculate the moles of reactants
- Determine which reactant is limiting
- Calculate the moles of product(s), as required
- Convert to grams or other units, as required

Concept Check (Part I)

- 10.0 mL of a 0.30 *M* sodium phosphate solution reacts with 20.0 mL of a 0.20 *M* lead(II) nitrate solution (assume no volume change).
  - What precipitate will form?
  - What mass of precipitate will form?

Concept Check (Part II)

- 10.0 mL of a 0.30 *M* sodium phosphate solution reacts with 20.0 mL of a 0.20 *M* lead(II) nitrate solution (assume no volume change).
  - What is the concentration of nitrate ions left in solution after the reaction is complete?

Concept Check (Part III)

- 10.0 mL of a 0.30 *M* sodium phosphate solution reacts with 20.0 mL of a 0.20 *M* lead(II) nitrate solution (assume no volume change).
  - What is the concentration of phosphate ions left in solution after the reaction is complete?

Interactive Example 6.10 - Determining the Mass of Product Formed I

- Calculate the mass of solid NaCl that must be added to 1.50 L of a 0.100 *M* $AgNO_3$ solution to precipitate all the Ag^+ ions in the form of AgCl

Solution

- Objective
  - To find the mass of solid NaCl required to precipitate the Ag^+
- Information available
  - 1.50 L of 0.100 *M* $AgNO_3$
- Information required
  - Moles of Ag^+ in the solution
- Step 1 - Determine the ions present in the solution
  - Ag^+, NO_3^-, Na^+, and Cl^-

Solution

- Step 2 - State the balanced equation for the reaction
  - $NaNO_3$ is soluble, and AgCl is insoluble
  - Therefore,

$$Ag^{+}(aq) + Cl^{-}(aq) \longrightarrow AgCl(s)$$

- Step 3 - Determine the moles of Ag^+ ions present in the solution

$$1.50\ \cancel{L} \times \frac{0.100\ \text{mol}\ Ag^{+}}{\cancel{L}} = 0.150\ \text{mol}\ Ag^{+}$$

Solution

- Step 4 - Determine the moles of Cl^- required to react with all the Ag^+
  - Because Ag^+ and Cl^- react in a 1:1 ratio, 0.150 mole of Cl^- and thus 0.150 mole of NaCl are required
- Step 5 - Determine the mass of NaCl required

$$0.150 \ \cancel{\text{mol NaCl}} \times \frac{58.44 \text{ g NaCl}}{\cancel{\text{mol NaCl}}} = 8.77 \text{ g NaCl}$$

Acid–Base Reactions (Brønsted–Lowry)

- An acid is a **proton** donor
- A base is a **proton** acceptor
- Consider a solution containing $HC_2H_3O_2(aq)$ and $KOH(aq)$
  - The principal species are $HC_2H_3O_2$, K^+, and OH^-
  - Precipitation does not occur as KOH is soluble
  - OH^- ions are capable of attracting protons from $HC_2H_3O_2$ molecules

Acid–Base Reactions (Brønsted–Lowry)

- The net ionic reaction is

$$OH^-(aq) + HC_2H_3O_2(aq) \longrightarrow H_2O(l) + C_2H_3O_2^-(aq)$$

  - For purposes of stoichiometric calculations, it can be assumed that the OH^- ion reacts completely with any weak acid
- Acid–base reactions are also called **neutralization reactions**

Performing Calculations for Acid–Base Reactions

- List the species present in the combined solution before any reaction occurs, and decide what reaction will occur
- Write the balanced net ionic equation for this reaction
- Calculate moles of reactants
- Determine the limiting reactant where appropriate
- Calculate the moles of the required reactant or product
- Convert to grams or volume (of solution), as required

Interactive example 6.12 - Neutralization Reactions I

- What volume of a 0.100 *M* HCl solution is needed to neutralize 25.0 mL of 0.350 *M* NaOH?

Solution

- Objective
  - To find the volume of 0.100 *M* HCl required for neutralization
- Information available
  - 25 mL of 0.350 *M* NaOH
  - 0.100 *M* HCl
  - The chemical reaction

$$H^{+}(aq) + OH^{-}(aq) \longrightarrow H_2O(l)$$

Solution

- Step 1 - Identify the ions present in the combined solution
  - H^+, Cl^-, Na^+, and OH^-
  - The two possible reactions are

$$Na^+(aq) + Cl^-(aq) \longrightarrow NaCl(s)$$

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

- Step 2 - Determine the balanced equation for the reaction

$$H^+(aq) + OH^-\ (aq) \longrightarrow H_2O\ (l)$$

Solution

- Step 3 - Determine the moles of reactant present in the solution

$$25.0\ \text{mL NaOH} \times \frac{1\ \text{L}}{1000\ \text{mL}} \times \frac{0.350\ \text{mol OH}^-}{\text{L NaOH}} = 8.75 \times 10^{-3}\ \text{mol OH}^-$$

- Step 4 - Identify the limiting reactant
  - This problem requires the addition of just enough H^+ to react exactly with the OH^- ions present
    - The limiting reactant is not significant here

Solution

- Step 5 - Determine the moles of H^+ required
  - Since H^+ and OH^- ions react in a 1:1 ratio, 8.75×10^{-23} moles of H^+ are required to neutralize the OH^- ions present
- Step 6 - Determine the volume of HCl required

$$V \times \frac{0.100 \text{ mol H}^+}{\text{L}} = 8.75\times10^{-3} \text{ mol H}^+$$

$$V = \frac{8.75 \times 10^{-3} \ \cancel{\text{mol H}^+}}{\dfrac{0.100 \ \cancel{\text{mol H}^+}}{\text{L}}} = 8.75\times10^{-2} \text{ L}$$

Acid–Base Titrations

- **Volumetric analysis** is used to determine the amount of a certain substance by doing a titration
- **Titration** is the delivery of a measured volume of solution of known concentration (titrant) into a solution containing the substance being analyzed (analyte)

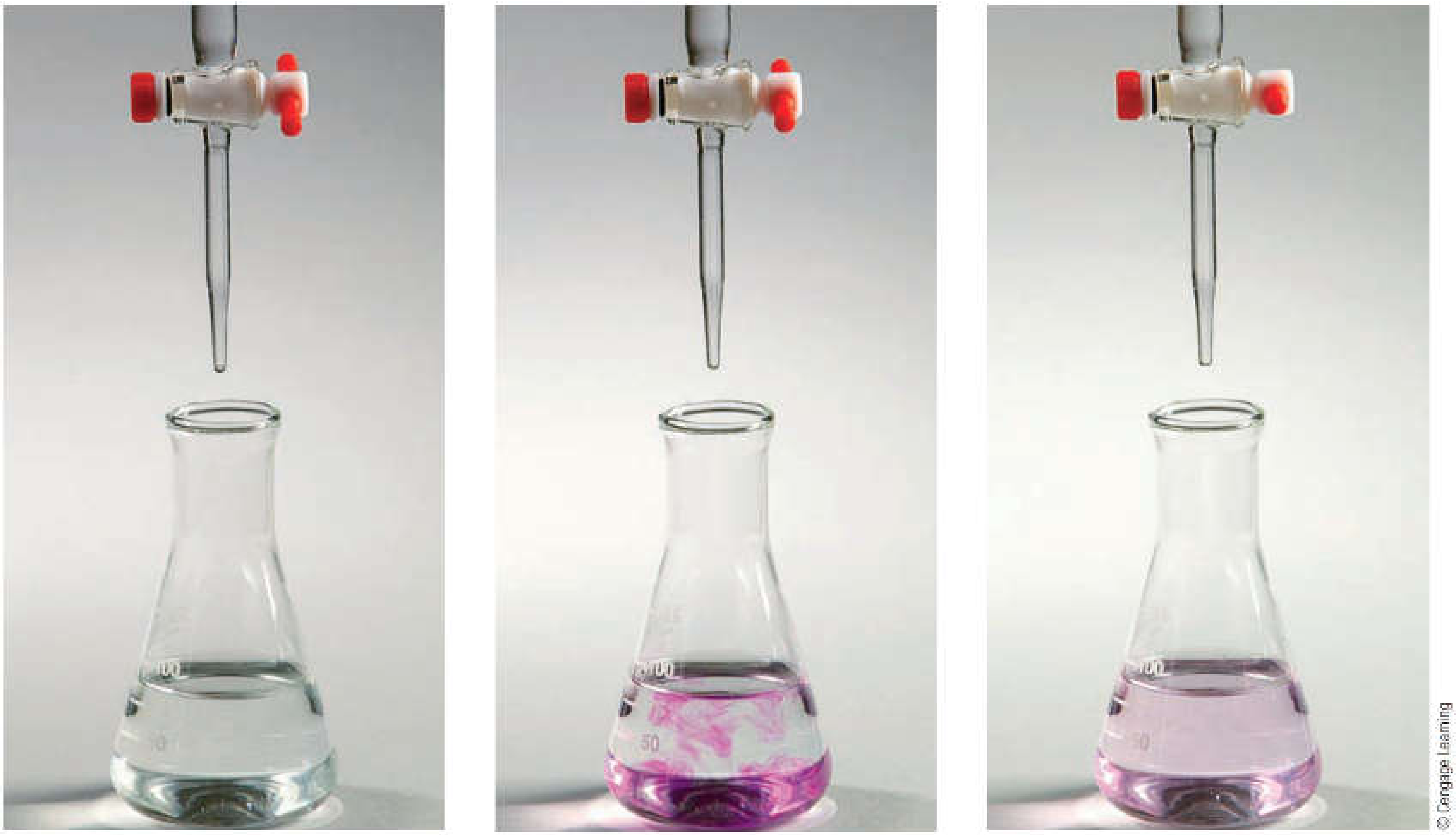
Acid–Base Titrations

- **Equivalence point** or **stoichiometric point** is the stage in titration at which enough titrant has been added to react exactly with the analyte
  - An **indicator** is a substance added at the beginning of the titration that changes color at the equivalence point
    - **Endpoint** is the stage where the indicator actually changes color

Requirements for a Successful Titration

- The exact reaction between the titrant and analyte must be known (and rapid)
- The stoichiometric (equivalence) point must be marked accurately
- The volume of titrant required to reach the stoichiometric point must be known accurately

Figure 6.18 - Titration of an Acid with a Base



Interactive Example 6.15 - Neutralization Analysis

- An environmental chemist analyzed the effluent (the released waste material) from an industrial process known to produce the compounds carbon tetrachloride (CCl_4) and benzoic acid ($HC_7H_5O_2$), a weak acid that has one acidic hydrogen atom per molecule. A sample of this effluent weighing 0.3518 g was shaken with water, and the resulting aqueous solution required 10.59 mL of 0.1546 *M* NaOH for neutralization. Calculate the mass percent of $HC_7H_5O_2$ in the original sample

Solution

- Objective
  - To find the mass percent of $HC_7H_5O_2$ in the original sample
- Information available
  - 0.3518 g effluent (original sample)
  - 10.59 mL 0.1546 *M* NaOH for neutralization of $HC_7H_5O_2$
  - The chemical reaction

$$HC_7H_5O_2(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_7H_5O_2^-(aq)$$

Solution

- Step 1 - Identify the species present in the combined solution
  - $HC_7H_5O_2$, Na^+, and OH^-
- Step 2 - Determine the balanced equation for the reaction

$$HC_7H_5O_2(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_7H_5O_2^-(aq)$$

- Step 3 - Calculate the moles of OH^- required

$$10.59\ \cancel{\text{mL NaOH}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.1546\ \text{mol OH}^-}{\cancel{\text{L NaOH}}} = 1.637 \times 10^{-3}\ \text{mol OH}^-$$

Solution

- Step 4 - Identify the limiting reactant
  - This problem requires the addition of just enough OH^{+} ions to react exactly with the $HC_7H_5O_2$ present
- Step 5 - Determine the mass of $HC_7H_5O_2$ present

$$1.637 \times 10^{-3} \text{ mol } \cancel{HC_7H_5O_2} \times \frac{122.12 \text{ g } HC_7H_5O_2}{1 \text{ mol } \cancel{HC_7H_5O_2}} = 0.1999 \text{ g } HC_7H_5O_2$$

- Step 6 - Determine the mass percent of $HC_7H_5O_2$ in the effluent

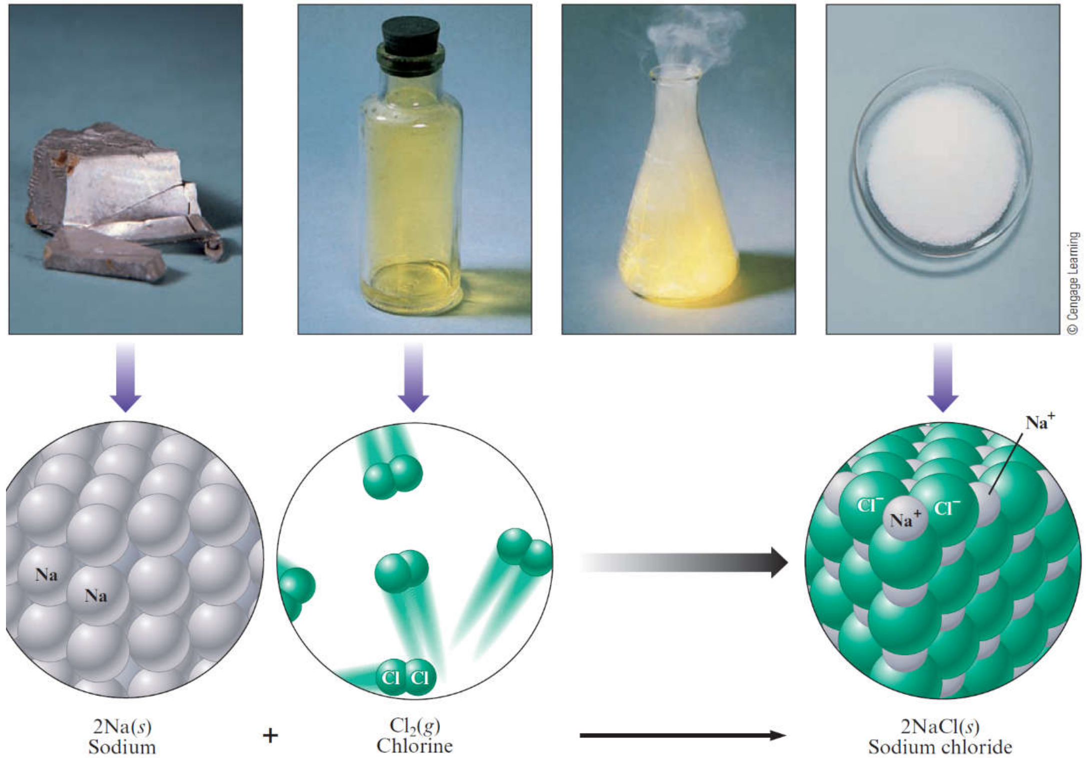
$$\frac{0.1999 \text{ g}}{0.3158 \text{ g}} \times 100\% = 56.82\%$$

Concept Check

- For the titration of sulfuric acid (H_2SO_4) with sodium hydroxide (NaOH), how many moles of sodium hydroxide would be required to react with 1.00 L of 0.500 *M* sulfuric acid to reach the endpoint?

Oxidation–Reduction Reactions (Redox Reactions)

- Reactions in which one or more electrons are transferred



Oxidation States

- Also called oxidation numbers
- Help keep track of electrons in oxidation–reduction reactions
- Defined as imaginary charges that atoms would have if the shared electrons were divided equally between identical atoms bonded to one another
- Oxidation states on ions are represented as $+n$ or $-n$

Table 6.2 - Rules for Assigning Oxidation States

The Oxidation State of . . .	Summary	Examples
• An atom in an element is zero	Element: 0	Na(*s*), O_2(*g*), O_3(*g*), Hg(*l*)
• A monatomic ion is the same as its charge	Monatomic ion: charge of ion	Na^+, Cl^-
• Fluorine is −1 in its compounds	Fluorine: −1	HF, PF_3
• Oxygen is usually −2 in its compounds Exception: peroxides (containing O_2^{2-}), in which oxygen is −1	Oxygen: −2	H_2O, CO_2
• Hydrogen is +1 in its covalent compounds	Hydrogen: +1	H_2O, HCl, NH_3

The Characteristics of Oxidation–Reduction Reactions

- At times, the transfer of electrons may not be evident
- Consider the combustion of methane

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

Oxidation State:

- CH_4: C = –4, H = +1 (each H)
- O_2: 0
- CO_2: C = +4, O = –2 (each O)
- H_2O: H = +1 (each H), O = –2

  - Carbon undergoes a change in oxidation state from –4 in CH_4 to +4 in CO_2

$$\underset{-4}{CH_4} \longrightarrow \underset{+4}{CO_2} + 8e^-$$

The Characteristics of Oxidation–Reduction Reactions

- Oxygen undergoes a change from an oxidation state of 0 in O_2 to –2 in H_2O and CO_2

$$\underset{\uparrow \atop 0}{2O_2} + 8e^- \longrightarrow CO_2 + 2H_2O$$

$$4(-2) = -8 \text{ (O in } CO_2 \text{ and } H_2O)$$

- **Oxidation** is an increase in oxidation state
  - Loss of electrons
- **Reduction** is a decrease in oxidation state
  - Gain of electrons

The Characteristics of Oxidation–Reduction Reactions

- In the combustion of methane:
  - The carbon in methane is oxidized
  - Oxygen is reduced
  - CH_4 is the **reducing agent**
  - O_2 is the **oxidizing agent**
- When the oxidizing or reducing agent is named, the whole compound is specified

Concept Check

- Which of the following are oxidation–reduction reactions? Identify the oxidizing agent and the reducing agent

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

b) $Cr_2O_7^{2-}(aq) + 2OH^-(aq) \longrightarrow 2CrO_4^{2-}(aq) + H_2O(l)$

c) $2CuCl(aq) \longrightarrow CuCl_2(aq) + Cu(s)$

Interactive Example 6.17 - Oxidation–Reduction Problems

- Metallurgy, the process of producing a metal from its ore, always involves oxidation–reduction reactions. In the metallurgy of galena (PbS), the principal lead-containing ore, the first step is the conversion of lead sulfide to its oxide (a process called roasting):

$$2\text{PbS}(s) + 3\text{O}_2(g) \longrightarrow 2\text{PbO}(s) + 2\text{SO}_2(g)$$

Interactive Example 6.17 - Oxidation–Reduction Problems

- The oxide is then treated with carbon monoxide to produce the free metal:

$$\mathrm{PbO}(s) + \mathrm{CO}(g) \longrightarrow \mathrm{Pb}(s) + \mathrm{CO_2}(g)$$

- For each reaction, identify the atoms that are oxidized and reduced, and specify the oxidizing and reducing agents

Solution

- Assigning oxidation states to the first reaction

$$2PbS(s) + 3O_2(g) \longrightarrow 2PbO(s) + 2SO_2(g)$$

Oxidation states: Pb in PbS: +2, S in PbS: –2; O in O_2: 0; Pb in PbO: +2, O in PbO: –2; S in SO_2: +4, O in SO_2: –2 (each O)

  - Sulfur is oxidized
  - Oxygen is reduced
  - The reducing agent is PbS

Solution

- Assigning oxidation states to the second reaction

$$PbO(s) + CO(g) \longrightarrow Pb(s) + CO_2(g)$$

$$\underset{+2}{Pb}\underset{-2}{O}(s) + \underset{+2}{C}\underset{-2}{O}(g) \longrightarrow \underset{0}{Pb}(s) + \underset{+4}{C}\underset{-2\ (\text{each O})}{O_2}(g)$$

  - Lead is reduced
  - Carbon is oxidized
  - PbO is the oxidizing agent
  - CO is the reducing agent

Balancing Oxidation–Reduction Reactions by Oxidation States

- Write the unbalanced equation
- Determine the oxidation states of all atoms in the reactants and products
- Show electrons gained and lost using "tie lines"
- Use coefficients to equalize the electrons gained and lost
- Balance the rest of the equation by inspection
- Add appropriate states

Example 6.18 - Balancing Oxidation–Reduction States

- Balance the reaction between solid lead(II) oxide and ammonia gas to produce nitrogen gas, liquid water, and solid lead

Solution

- Step 1 - State the unbalanced equation

$$PbO(s) + NH_3(g) \longrightarrow N_2(g) + H_2O(l) + Pb(s)$$

- Step 2 - Assign oxidation states to each atom

$$\underset{+2}{Pb}\overset{-2}{O} + \overset{-3}{N}\underset{+1}{H_3} \longrightarrow \overset{0}{N_2} + \underset{+1}{H_2}\overset{-2}{O} + \underset{0}{Pb}$$

Solution

- Step 3 - Use “tie lines” to show electrons gained and lost

3 e^- lost (each atom)

$$\overset{+2}{PbO} + \overset{-3}{NH_3} \longrightarrow \overset{0}{N_2} + H_2O + \overset{0}{Pb}$$

2 e^- gained

  - The oxidation states of all other atoms are unchanged

Solution

- Step 4 - Use coefficients to equalize electrons gained and lost

3 e⁻ lost (each atom) multiply by 2

$$\overset{+2}{PbO} + \overset{-3}{N}H_3 \longrightarrow \overset{0}{N_2} + H_2O + \overset{0}{Pb}$$

2 e⁻ gained multiply by 3

$$3PbO + 2NH_3 \longrightarrow N_2 + H_2O + 3Pb$$

Solution

- Step 5 - Balance the rest of the elements

$$3PbO + 2NH_3 \longrightarrow N_2 + 3H_2O + 3Pb$$

- Step 6 - Add appropriate states

$$3PbO(s) + 2NH_3(g) \longrightarrow N_2(g) + 3H_2O(l) + 3Pb(s)$$