Chapter 4 Reactions in Aqueous Solution

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Solutions

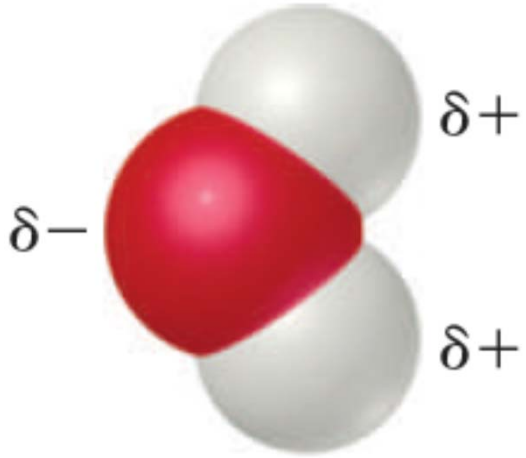
- Solutions are defined as homogeneous mixtures of two or more pure substances.
- The solvent is present in greatest abundance.
- All other substances are solutes.
- When water is the solvent, the solution is called an aqueous solution.

Electrolytes and Nonelectrolytes

- A substance (such as NaCl) whose aqueous solutions contain ions is called an electrolyte.
  - When NaCl dissolves in water, the solution contains Na^+ and Cl^- ions, each surrounded by water molecules
- A substance (such as $C_{12}H_{22}O_{11}$) that does not form ions in solution is called a nonelectrolyte.
  - The sucrose dissolves in water, the solution contains only neutral sucrose molecules surrounded by water molecules.

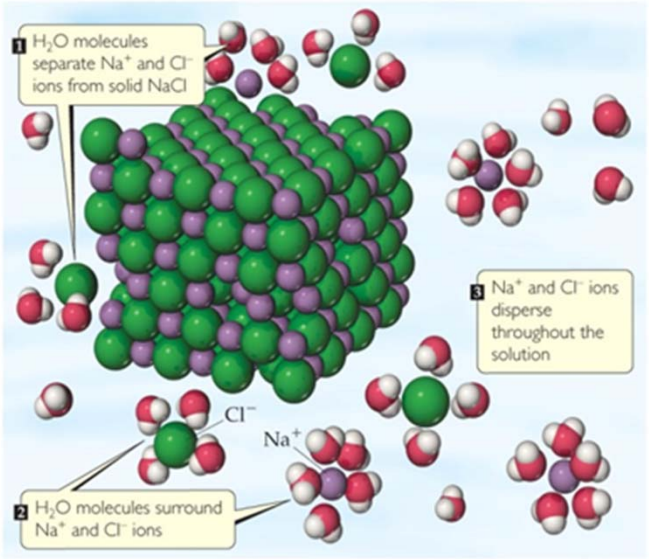
Charge Distribution in a Water Molecule

- There is an uneven distribution of electrons within the water molecule.
  - This causes the oxygen side of the molecule to have a partial negative charge (δ^-) and the hydrogen side to have a partial positive charge (δ^+).

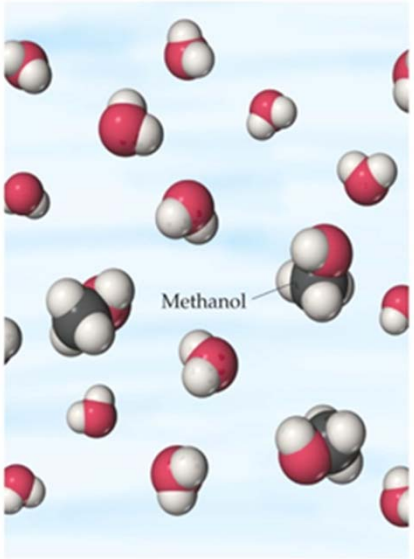


Aqueous Solutions

- Substances can dissolve in water by different ways:
  - Ionic Compounds dissolve by dissociation, where water surrounds the separated ions.
  - Molecular compounds interact with water, but most do NOT dissociate.
  - Some molecular substances react with water when they dissolve.



(a) Ionic compounds like sodium chloride, NaCl, form ions when they dissolve.



(b) Molecular substances like methanol, CH_3OH, dissolve without forming ions.

Strong and Weak Electrolytes

- Electrolytes differ in the extent to which they conduct electricity.
- Strong electrolytes are those solutes that exist in solution completely or nearly completely as ions.
  - all water-soluble ionic compounds (NaCl)
  - a few molecular compounds (such as HCl)

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

- Weak electrolytes are those solutes that exist in solution mostly in the form of neutral molecules with only a small fraction in the form of ions.
  - in a solution of acetic acid (CH_3COOH), most of the solute is present as $CH_3COOH(aq)$ molecules.
  - Only a small fraction (about 1%) of the CH_3COOH has dissociated into $H^+(aq)$ and $CH_3COO^-(aq)$ ions.

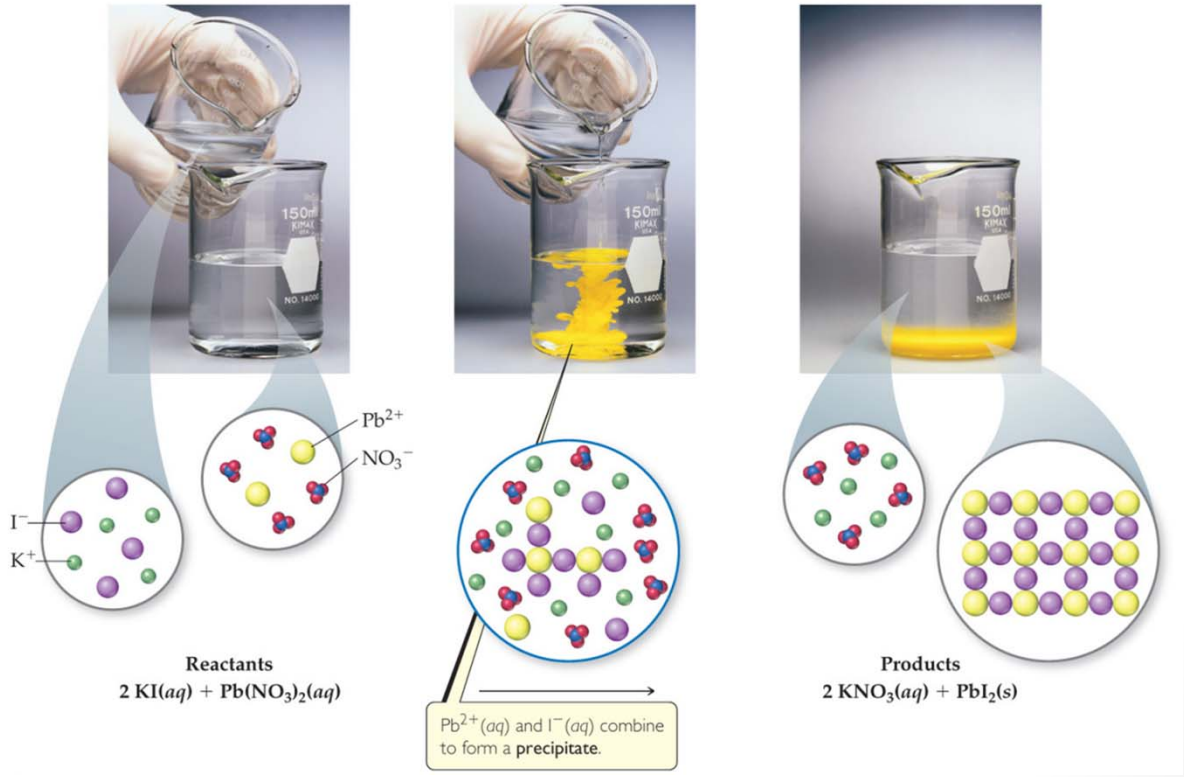
$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

Strong and Weak Electrolytes

- CH_3COOH is extremely soluble in water but is a weak electrolyte.
- $Ca(OH)_2$ is not very soluble in water, but the amount that does dissolve dissociates almost completely. Thus, $Ca(OH)_2$ is a strong electrolyte.
- Water-soluble ionic compounds are strong electrolytes.
  - Ionic compounds can usually be identified by the presence of both metals and nonmetals. (EX: NaCl, $FeSO_4$, and $Al(NO_3)_3$)
  - But, ionic compounds containing the ammonium ion, NH_4^+ [for example, NH_4Br and $(NH_4)_2CO_3$) are weak electrolytes.

Precipitation Reactions

- Reactions that result in the formation of an insoluble product are called precipitation reactions.
- A precipitate is an insoluble solid formed by a reaction in solution.



Solubility Guidelines for Ionic Compounds

- The **solubility** of a substance at a given temperature is the amount of the substance that can be dissolved in a given quantity of solvent at the given temperature.
- Any substance with a solubility less than 0.01 mol/L will be considered **insoluble**.

Table 4.1 Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		Important Exceptions
Compounds containing	NO_3^-	None
	CH_3COO^-	None
	Cl^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	Br^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	I^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}
Insoluble Ionic Compounds		**Important Exceptions**
Compounds containing	S^{2-}	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^+ and the alkali metal cations
	OH^-	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}

Exercise 4.2

- Classify these ionic compounds as soluble or insoluble in water: (a) sodium carbonate, Na_2CO_3, (b) lead sulfate, $PbSO_4$.

Sol:

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Metathesis (Exchange) Reactions

- Such reactions are called either exchange reactions or metathesis reactions

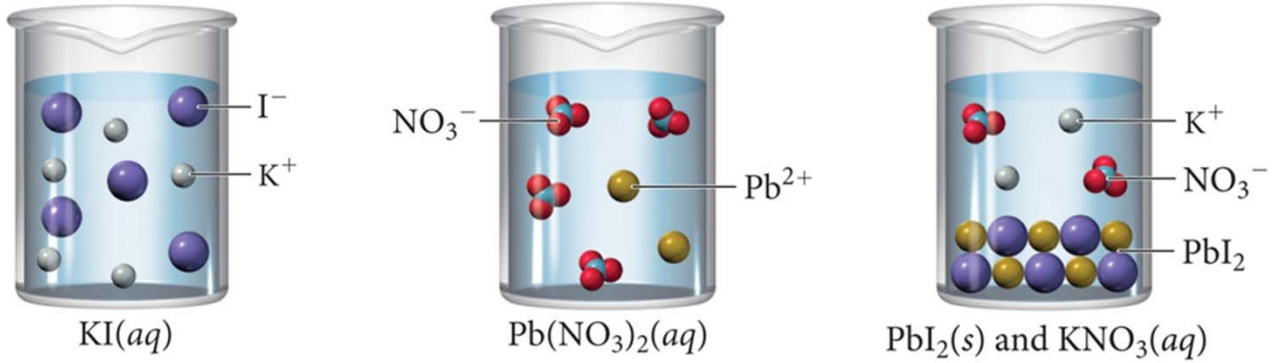
Example: $AgNO_3(aq) + KCl(aq) \longrightarrow AgCl(s) + KNO_3(aq)$

- Metathesis comes from a Greek word that means “to transpose.”

Predicting Precipitation Reactions

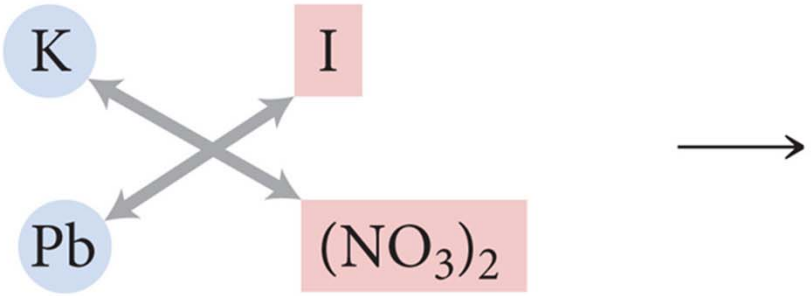
1. Determine what ions each aqueous reactant has.
2. Determine formulas of possible products.
   - Exchange ions.
     (+) ion from one reactant with (–) ion from other
   - Balance charges of combined ions to get the formula of each product.
3. Determine solubility of each product in water.
   - Use the solubility rules.
   - If product is insoluble or slightly soluble, it will precipitate.
4. If neither product will precipitate, write **no reaction** after the arrow.
5. If any of the possible products are insoluble, write their formulas as the products of the reaction using (s) after the formula to indicate solid. Write any soluble products with (aq) after the formula to indicate aqueous.
6. Balance the equation.
   - Remember to only change coefficients, not subscripts.

Predicting Precipitation Reactions



Original compounds

Possible products



KNO_3

PbI_2

$$\underset{\text{(soluble)}}{2\,KI(aq)} + \underset{\text{(soluble)}}{Pb(NO_3)_2(aq)} \longrightarrow \underset{\text{(soluble)}}{2\,KNO_3(aq)} + \underset{\text{(insoluble)}}{PbI_2(s)}$$

When a potassium iodide solution is mixed with a lead(II) nitrate solution, a yellow lead(II) iodide precipitate forms.

Question

(a) $NaOH(aq) + MgCl_2(aq) \longrightarrow ?$

(b) $BaS(aq) + CuSO_4(aq) \longrightarrow ?$

(c) $(NH_4)_2SO_4(aq) + ZnCl_2(aq) \longrightarrow ?$

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Representing Aqueous Reactions

- Ways to Write Metathesis Reactions
  - Molecular equation
  - Complete ionic equation
  - Net ionic equation

Molecular Equation

- An equation written in this fashion, showing the complete chemical formulas of reactants and products, is called a molecular equation because it shows chemical formulas without indicating ionic character.

EX:

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

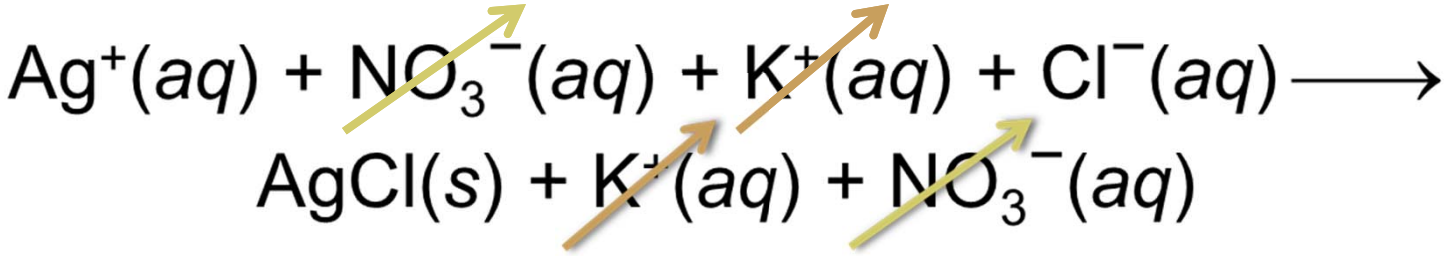
Complete Ionic Equation

- An equation written in this form, with all soluble strong electrolytes shown as ions, is called a complete ionic equation.

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

Net Ionic Equation

- To form the net ionic equation, cross out anything that does not change from the left side of the equation to the right.
- The ions crossed out are called spectator ions



K^+ and NO_3^- are called spectator ions

Exercise 4.4
Writing a Net Ionic Equation

- Write the net ionic equation for the precipitation reaction that occurs when aqueous solutions of calcium chloride and sodium carbonate are mixed.

Sol:

$$CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2\ NaCl(aq)$$

$$Ca^{2+}(aq) + 2Cl^-(aq) + 2Na^+(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s) + 2Na^+(aq) + 2Cl^-(aq)$$

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

Insoluble Ionic Compounds		Important Exceptions
Compounds containing	S^{2-}	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations

Acids, Bases, and Neutralization Reactions

- Acids are substances that ionize in aqueous solution to form hydrogen ions H^+ (aq).
- A hydrogen atom consists of a proton and an electron, H^+ is simply a proton.
- Acids are often called proton donors.
- Molecules of different acids ionize to form different numbers of H^+ ions
  - Monoprotic acids, yielding one H^+ per molecule of acid. (EX: HCl and HNO_3)
  - Diprotic acid, one that yields two H+ per molecule of acid. (EX: H_2SO_4)

$$H_2SO_4(aq) \longrightarrow H^+(aq) + HSO_4^-(aq) \qquad [4.9]$$

$$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq) \qquad [4.10]$$

Aqueous solutions of sulfuric acid contain a mixture of H^+(aq), HSO_4^-(aq), and SO_4^{2-}(aq).

Bases

- Bases are substances that accept (react with) H^+ ions.
- Bases produce hydroxide ions (OH^-) when they dissolve in water.
- Ionic hydroxide compounds, such as NaOH, KOH, and $Ca(OH)_2$, are among the most common bases.
- Compounds that do not contain OH^- ions can also be bases. For example, ammonia (NH_3) is a common base.

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

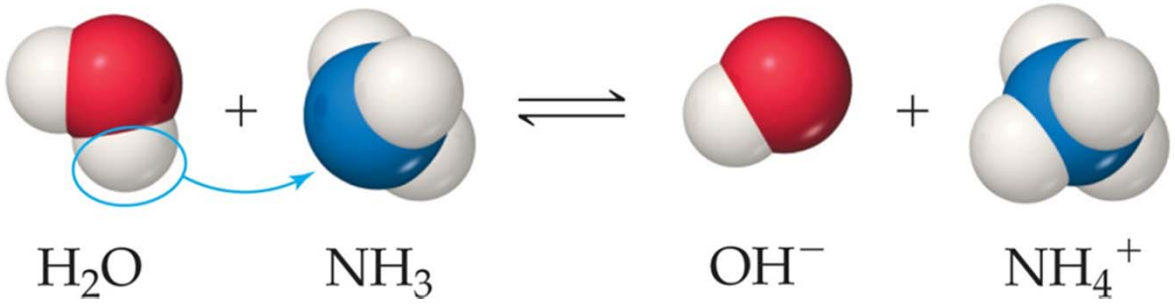
Strong or Weak?

- Strong acids completely dissociate in water; weak acids only partially dissociate.
- Strong bases dissociate to metal cations and hydroxide anions in water; weak bases only partially react to produce hydroxide anions.

Table 4.2 Common Strong Acids and Bases

Strong Acids	Strong Bases
Hydrochloric acid, HCl	Group 1A metal hydroxides [LiOH, NaOH, KOH, RbOH, CsOH]
Hydrobromic acid, HBr	
Hydroiodic acid, HI	Heavy group 2A metal hydroxides [$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$]
Chloric acid, $HClO_3$	
Perchloric acid, $HClO_4$	
Nitric acid, HNO_3	
Sulfuric acid (first proton), H_2SO_4	

Acid-Base Reactions



- In an acid–base reaction, the acid (H_2O above) donates a proton (H^+) to the base (NH_3 above).
- When a solution of an acid and a solution of a base are mixed, a **neutralization reaction** occurs.
- When the base is a metal hydroxide, water and a **salt** (an ionic compound) are produced.

$$\underset{\text{(acid)}}{HCl(aq)} + \underset{\text{(base)}}{NaOH(aq)} \longrightarrow \underset{\text{(water)}}{H_2O(l)} + \underset{\text{(salt)}}{NaCl(aq)}$$

Writing Chemical Equations for a Neutralization Reaction

- For the reaction between aqueous solutions of acetic acid (CH_3COOH) and barium hydroxide, $Ba(OH)_2$, write (a) the balanced molecular equation, (b) the complete ionic equation, (c) the net ionic equation.

Molecular equation

$$2\,CH_3COOH(aq) + Ba(OH)_2(aq) \longrightarrow 2\,H_2O(l) + Ba(CH_3COO)_2(aq)$$

Complete ionic equation

$$2\,CH_3COOH(aq) + Ba^{2+}(aq) + 2\,OH^-(aq) \longrightarrow 2\,H_2O(l) + Ba^{2+}(aq) + 2\,CH_3COO^-(aq)$$

Net ionic equation

$$CH_3COOH(aq) + OH^-(aq) \longrightarrow H_2O(l) + CH_3COO^-(aq)$$

Gas-Forming Reactions

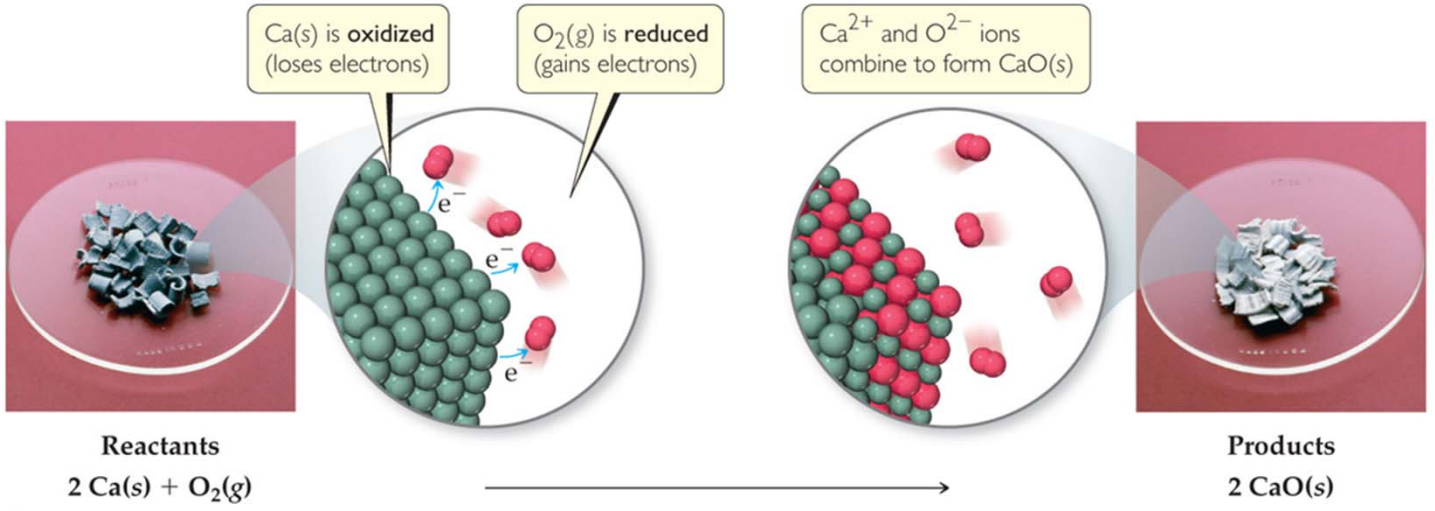
- When a carbonate or bicarbonate reacts with an acid, the products are a salt, carbon dioxide, and water.

$$CaCO_3(s) + 2\ HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

$$NaHCO_3(aq) + HBr(aq) \longrightarrow NaBr(aq) + CO_2(g) + H_2O(l)$$

Oxidation-Reduction Reactions

- Loss of electrons is oxidation.
- Gain of electrons is reduction.
- One cannot occur without the other.
- The reactions are often called redox reactions.



Oxidation Numbers

- To determine if an oxidation–reduction reaction has occurred, we assign an oxidation number to each element in a neutral compound or charged entity.
- Rules to Assign Oxidation Numbers:
  - For an atom in its elemental form, the oxidation number is always zero.

  Ex: each H atom in the H_2 molecule has an oxidation number of 0.
  each P atom in the P_4 molecule has an oxidation number of 0.
  - For any monatomic ion the oxidation number equals the ionic charge.

  Ex: K^+ has an oxidation number of +1, S^{2-} has an oxidation number of -2
    - In ionic compounds the alkali metal ions (group 1A) always have a 1+ charge and therefore an oxidation number of +1.
    - The alkaline earth metals (group 2A) are always +2,
    - Aluminum (group 3A) is always +3 in ionic compounds.

Oxidation Numbers

- Nonmetals usually have negative oxidation numbers, although they can sometimes be positive.
  - The oxidation number of oxygen is usually -2 in both ionic and molecular compounds. The major exception is in compounds called peroxides, which contain the O_2^{2-} ion, giving each oxygen an oxidation number of -1.
  - The oxidation number of hydrogen is usually +1 when bonded to nonmetals and -1 when bonded to metals (Ex: NaH).
  - The oxidation number of fluorine is -1 in all compounds. The other halogens have an oxidation number of -1 in most binary compounds. When combined with oxygen, as in oxyanions, however, they have positive oxidation states. (Ex: $HClO_4$, H: +1, O: -2, Cl: +7)

Oxidation Numbers

- The sum of the oxidation numbers of all atoms in a neutral compound is zero.
- The sum of the oxidation numbers in a polyatomic ion equals the charge of the ion.

Ex: hydronium ion H_3O^+, which is a more accurate description of $H^+(aq)$, the oxidation number of each hydrogen is +1 and that of oxygen is -2.

Thus, the sum of the oxidation numbers is 3(+1) + (-2) = +1, which equals the net charge of the ion.

Determining Oxidation Numbers

- Exercise 4.8:

Determine the oxidation number of sulfur in (a) H_2S, (b) S_8, (c) SCl_2, (d) Na_2SO_3, (e) SO_4^{2-}.

Oxidation of Metals by Acids and Salts

- The reaction between a metal and either an acid or a metal salt conforms to the general pattern

$$A + BX \longrightarrow AX + B$$

Examples:

$$Zn(s) + 2\,HBr(aq) \longrightarrow ZnBr_2(aq) + H_2(g)$$

$$Mn(s) + Pb(NO_3)_2(aq) \longrightarrow Mn(NO_3)_2(aq) + Pb(s)$$

These reactions are called **displacement reactions**

$$Mg(s) + 2\,HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

Oxidation number: 0 +1 −1 +2 −1 0

- The oxidation number of Mg changes from 0 to +2, an increase that indicates the atom has lost electrons and has therefore been oxidized.
- The oxidation number of H^+ in the acid decreases from +1 to 0, indicating that this ion has gained electrons and has therefore been reduced.

The Activity Series

- A list of metals arranged in order of decreasing ease of oxidation is called an activity series.

Table 4.5 Activity Series of Metals in Aqueous Solution

Metal	Oxidation Reaction
Lithium	$Li(s) \longrightarrow Li^+(aq) + e^-$
Potassium	$K(s) \longrightarrow K^+(aq) + e^-$
Barium	$Ba(s) \longrightarrow Ba^{2+}(aq) + 2e^-$
Calcium	$Ca(s) \longrightarrow Ca^{2+}(aq) + 2e^-$
Sodium	$Na(s) \longrightarrow Na^+(aq) + e^-$
Magnesium	$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$
Aluminum	$Al(s) \longrightarrow Al^{3+}(aq) + 3e^-$
Manganese	$Mn(s) \longrightarrow Mn^{2+}(aq) + 2e^-$
Zinc	$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$
Chromium	$Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^-$
Iron	$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$
Cobalt	$Co(s) \longrightarrow Co^{2+}(aq) + 2e^-$
Nickel	$Ni(s) \longrightarrow Ni^{2+}(aq) + 2e^-$
Tin	$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^-$
Lead	$Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^-$
Hydrogen	$H_2(g) \longrightarrow 2\,H^+(aq) + 2e^-$
Copper	$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$
Silver	$Ag(s) \longrightarrow Ag^+(aq) + e^-$
Mercury	$Hg(l) \longrightarrow Hg^{2+}(aq) + 2e^-$
Platinum	$Pt(s) \longrightarrow Pt^{2+}(aq) + 2e^-$
Gold	$Au(s) \longrightarrow Au^{3+}(aq) + 3e^-$

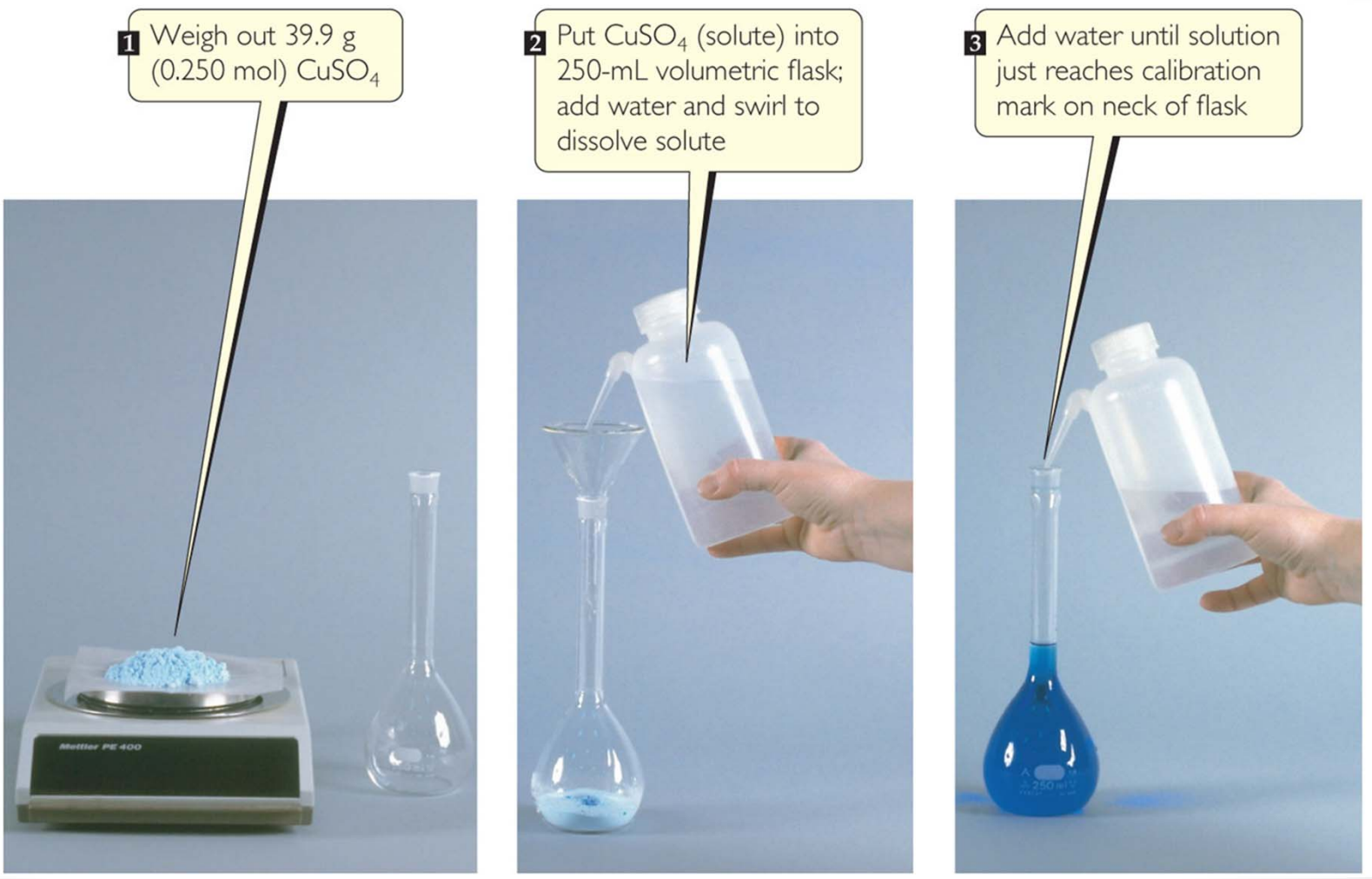
- The metals at the top of the table, such as the alkali metals and the alkaline earth metals, are most easily oxidized; that is, they react most readily to form compounds. They are called the active metals.
- The metals at the bottom of the activity series, such as the transition elements from groups 8B and 1B, are very stable and form compounds less readily. These metals are called noble metals.

Concentrations of Solutions

- Molarity (symbol M) expresses the concentration of a solution as the number of moles of solute in a liter of solution

$$\text{Molarity} = \frac{\text{moles solute}}{\text{volume of solution in liters}}$$

Preparing 0.250 L of a 1.00 M solution of $CuSO_4$



Calculating Molarity

- Calculate the molarity of a solution made by dissolving 23.4 g of sodium sulfate (Na_2SO_4) in enough water to form 125 mL of solution.

Sol:

$$\text{Moles } Na_2SO_4 = (23.4 \text{ g } Na_2SO_4)\left(\frac{1 \text{ mol } Na_2SO_4}{142.1 \text{ g } Na_2SO_4}\right) = 0.165 \text{ mol } Na_2SO_4$$

$$\text{Liters soln} = (125 \text{ mL})\left(\frac{1 \text{ L}}{1000 \text{ mL}}\right) = 0.125 \text{ L}$$

$$\text{Molarity} = \frac{0.165 \text{ mol } Na_2SO_4}{0.125 \text{ L soln}} = 1.32 \frac{\text{mol } Na_2SO_4}{\text{L soln}} = 1.32\ M$$

Calculating Molar Concentrations of Ions

- What is the molar concentration of each ion present in a 0.025 M aqueous solution of calcium nitrate $Ca(NO_3)_2$?

Sol:

$$Ca(NO_3)_2 \rightarrow Ca^{2+} + 2NO_3^-$$

$[Ca^{2+}]$: 0.025 M

$$\frac{\text{mol } NO_3^-}{\text{L}} = \left(\frac{0.025 \text{ mol } Ca(NO_3)_2}{\text{L}}\right)\left(\frac{2 \text{ mol } NO_3^-}{1 \text{ mol } Ca(NO_3)_2}\right) = 0.050\ M$$

$[NO_3^-]$: 0.05M

Using Molarity to Calculate Grams of Solute

- How many grams of Na_2SO_4 are required to make 0.350 L of 0.500 M Na_2SO_4?

Sol:

$$M_{Na_2SO_4} = \frac{\text{moles } Na_2SO_4}{\text{liters soln}}$$

$$\text{Moles } Na_2SO_4 = \text{liters soln} \times M_{Na_2SO_4}$$

$$= (0.350 \cancel{\text{ L soln}})\left(\frac{0.500 \text{ mol } Na_2SO_4}{1 \cancel{\text{ L soln}}}\right)$$

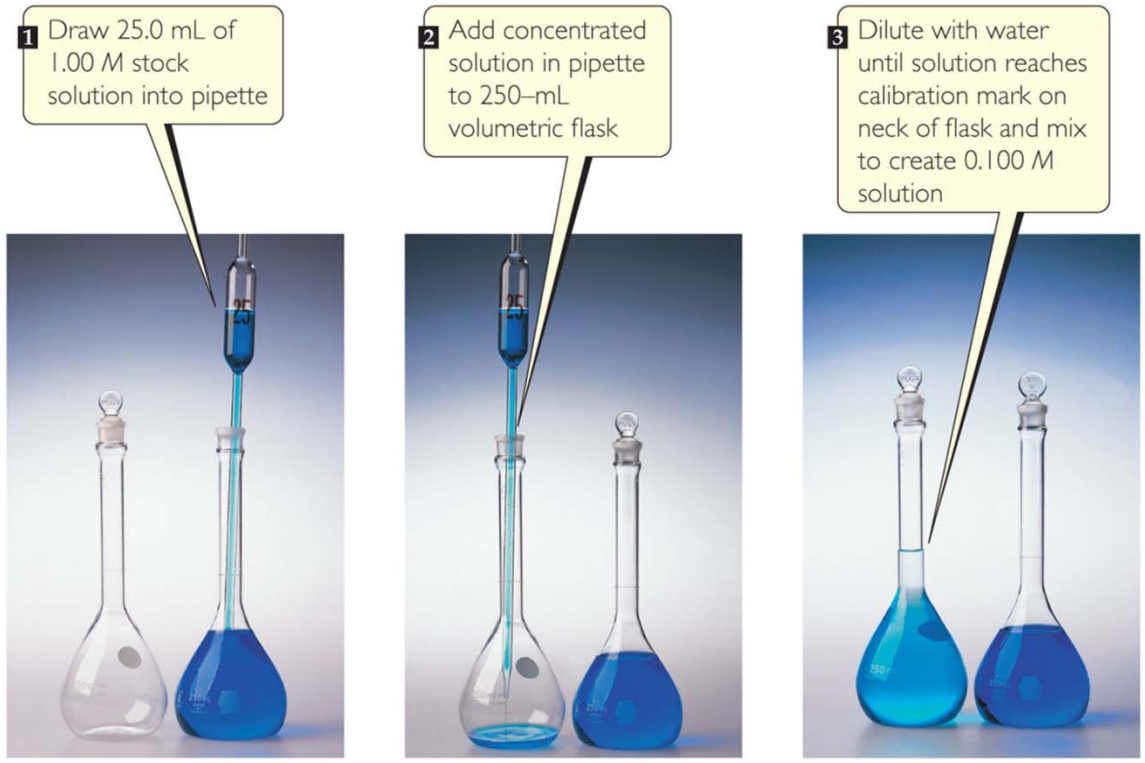
$$= 0.175 \text{ mol } Na_2SO_4$$

$$\text{Grams } Na_2SO_4 = (0.175 \cancel{\text{ mol } Na_2SO_4})\left(\frac{142.1 \text{ g } Na_2SO_4}{1 \cancel{\text{ mol } Na_2SO_4}}\right)$$

$$= 24.9 \text{ g } Na_2SO_4$$

Dilution

- One can also dilute a more concentrated solution by using a pipet to deliver a volume of the solution to a new volumetric flask, and adding solvent to the line on the neck of the new flask.



Dilution

- Solutions are stored as concentrated stock solutions.
- To make solutions of lower concentrations from these stock solutions, more solvent is added.
- The amount of solute doesn't change, just the volume of solution:

  Moles solute before dilution = moles solute after dilution

- The concentrations and volumes of the stock and new solutions are inversely proportional:

$$M_c \cdot V_c = M_d \cdot V_d$$

where M_c and M_d are the molarity of the concentrated and dilute solutions, respectively, and V_c and V_d are the volumes of the two solutions.

Preparing a Solution by Dilution

- How many milliliters of 3.0 M H_2SO_4 are needed to make 450 mL of 0.10 M H_2SO_4?

Sol:

$$\text{Moles } H_2SO_4 \text{ in dilute solution} = (0.450 \text{ L soln})\left(\frac{0.10 \text{ mol } H_2SO_4}{1 \text{ L soln}}\right)$$

$$= 0.045 \text{ mol } H_2SO_4$$

Calculating the volume of the concentrated solution that contains 0.045 mol H_2SO_4:

$$\text{L conc soln} = (0.045 \text{ mol } H_2SO_4)\left(\frac{1 \text{ L soln}}{3.0 \text{ mol } H_2SO_4}\right) = 0.015 \text{ L soln}$$

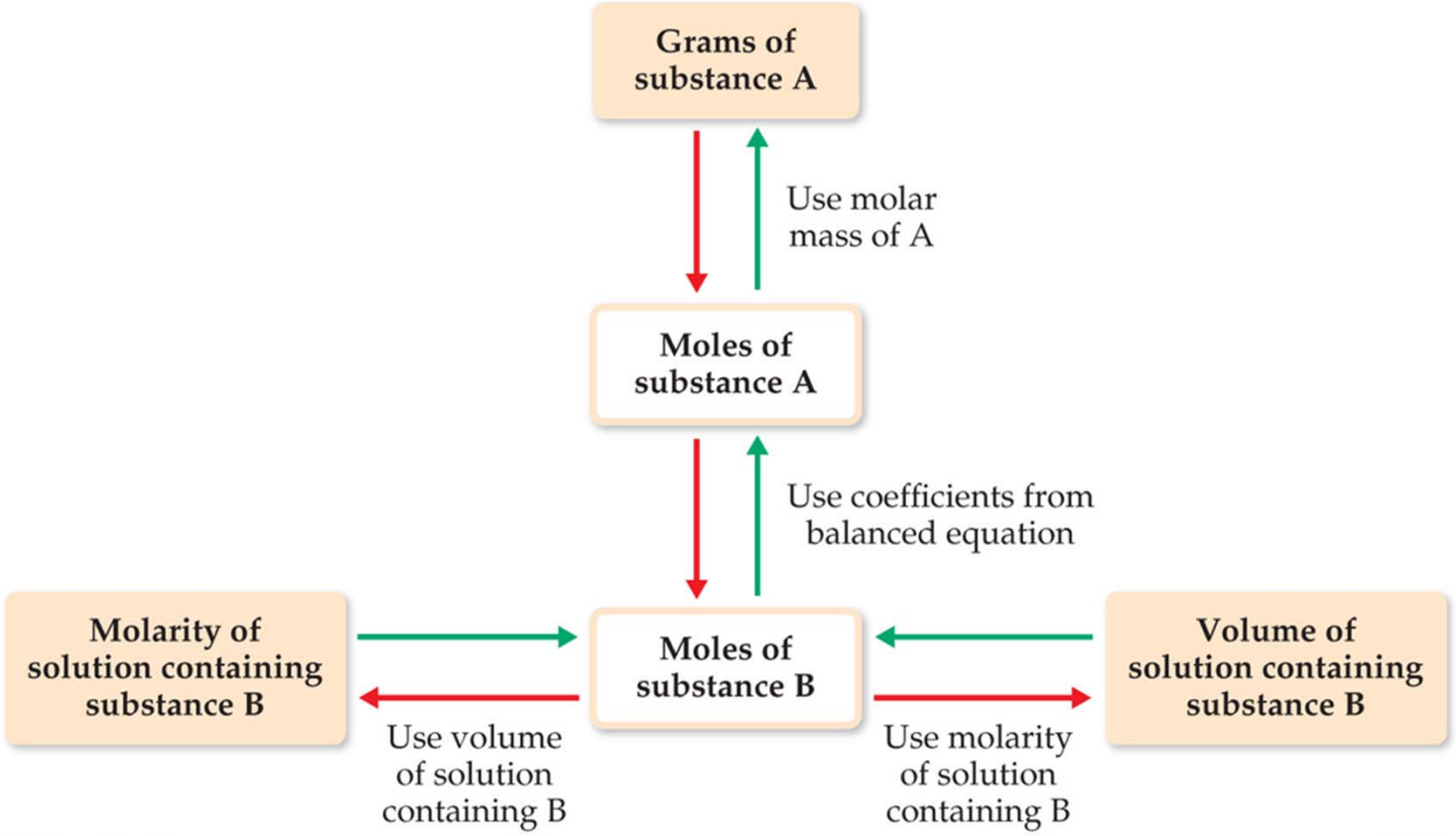
Converting liters to milliliters gives 15 mL.

Or

$$(3.0\ M)(V_{conc}) = (0.10\ M)(450 \text{ mL})$$

$$(V_{conc}) = \frac{(0.10\ M)(450 \text{ mL})}{3.0\ M} = 15 \text{ mL}$$

Using Molarities in Stoichiometric Calculations



Using Mass Relations in a Neutralization Reaction

- How many grams of $Ca(OH)_2$ are needed to neutralize 25.0 mL of 0.100 M HNO_3?

Sol:

$$2\,HNO_3(aq) + Ca(OH)_2(s) \longrightarrow 2\,H_2O(l) + Ca(NO_3)_2(aq)$$

$$\text{Moles } HNO_3 = V_{HNO_3} \times M_{HNO_3} = (0.0250\,\cancel{L})\left(\frac{0.100\text{ mol } HNO_3}{\cancel{L}}\right)$$

$$= 2.50 \times 10^{-3}\text{ mol } HNO_3$$

Thus, 2 mol HNO_3 $\simeq$ mol $Ca(OH)_2$. Therefore,

$$\text{Grams } Ca(OH)_2 = (2.50 \times 10^{-3}\,\cancel{\text{mol } HNO_3}) \times \left(\frac{1\,\cancel{\text{mol } Ca(OH)_2}}{2\,\cancel{\text{mol } HNO_3}}\right)\left(\frac{74.1\text{ g } Ca(OH)_2}{1\,\cancel{\text{mol } Ca(OH)_2}}\right)$$

$$= 0.0926\text{ g } Ca(OH)_2$$

Titrations

- A titration is an analytical technique in which one can calculate the concentration of a solute in a solution.
- In a titration, a substance in a solution of known concentration is reacted with another substance in a solution of unknown concentration.
- Equivalence point —the point in the titration when the number of moles of OH^- added equals the number of moles of H+ initially in solution
- The equivalence point is typically signaled by an indicator , a dye whose color depends on the acidity or basicity of the solution

Determining Solution Concentration by an Acid–Base Titration

- In one analysis, 45.7 mL of 0.500 M H_2SO_4 is required to neutralize 20.0 mL of NaOH solution. What is the concentration of the NaOH solution?

Sol:

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

$$\text{Moles } H_2SO_4 = (45.7\ \cancel{\text{mL soln}})\left(\frac{1\ \cancel{\text{L soln}}}{1000\ \cancel{\text{mL soln}}}\right)\left(\frac{0.500\ \text{mol } H_2SO_4}{\cancel{\text{L soln}}}\right)$$

$$= 2.29 \times 10^{-2}\ \text{mol } H_2SO_4$$

$$\text{Moles NaOH} = (2.29 \times 10^{-2}\ \cancel{\text{mol } H_2SO_4})\left(\frac{2\ \text{mol NaOH}}{1\ \cancel{\text{mol } H_2SO_4}}\right)$$

$$= 4.57 \times 10^{-2}\ \text{mol NaOH}$$

$$\text{Molarity NaOH} = \frac{\text{mol NaOH}}{\text{L soln}}$$

$$= \left(\frac{4.57 \times 10^{-2}\ \text{mol NaOH}}{20.0\ \cancel{\text{mL soln}}}\right)\left(\frac{1000\ \cancel{\text{mL soln}}}{1\ \text{L soln}}\right) = 2.29\,\frac{\text{mol NaOH}}{\text{L soln}} = 2.29\ M$$

Question

A piece of iron wire weighing 0.1568 g is converted to $Fe^{2+}(aq)$ and requires 26.24 mL of a $KMnO_4(aq)$ solution for its titration. What is the molarity of the $KMnO_4(aq)$?

$$5\,Fe^{2+}(aq) + MnO_4^-(aq) + 8\,H^+(aq) \longrightarrow 5\,Fe^{3+}(aq) + Mn^{2+}(aq) + 4\,H_2O(l) \qquad (5.27)$$

Sol:

$$5\,Fe^{2+}(aq) + MnO_4^-(aq) + 8\,H^+(aq) \longrightarrow 5\,Fe^{3+}(aq) + Mn^{2+}(aq) + 4\,H_2O(l)$$

First, determine the amount (in moles) of $KMnO_4$ consumed in the titration.

$$\begin{aligned} ?\ \text{mol } KMnO_4 &= 0.1568\ \text{g Fe} \times \frac{1\ \text{mol Fe}}{55.847\ \text{g Fe}} \times \frac{1\ \text{mol } Fe^{2+}}{1\ \text{mol Fe}} \\ &\quad \times \frac{1\ \text{mol } MnO_4^-}{5\ \text{mol } Fe^{2+}} \times \frac{1\ \text{mol } KMnO_4}{1\ \text{mol } MnO_4^-} \\ &= 5.615 \times 10^{-4}\ \text{mol } KMnO_4 \end{aligned}$$

The volume of solution containing the 5.615×10^{-4} mol $KMnO_4$ is 26.24 mL = 0.02624 L, which means that

$$\text{concn } KMnO_4 = \frac{5.615 \times 10^{-4}\ \text{mol } KMnO_4}{0.02624\ \text{L}} = 0.02140\ \text{M } KMnO_4$$