

## TOPIC

# 4

## AQUEOUS REACTIONS AND SOLUTION STOICHIOMETRY

The content in this topic is the basis for mastering Learning Objectives 3.2, 3.3, 3.4, 3.8, and 3.9 as found in the Curriculum Framework.

You may have learned some of this material in first-year chemistry. When you finish reviewing this topic, be sure you are able to:

- Distinguish chemical formulas as strong, weak, and nonelectrolytes
- Recognize the formulas of strong and weak acids
- Write chemical equations for the ionization of strong and weak acids and bases
- Write balanced net ionic equations for precipitation reactions, neutralization reactions, and reactions between acids and bases
- Identify redox reactions and the electron transfer in redox reactions
- Use an activity series to write balanced net ionic equations for redox reactions between metals and metal ions
- Assign oxidation numbers to elements in a chemical formula
- Perform limiting reactant and solution stoichiometry calculations
- Interpret the results of a redox titration

### Section 4.1

## General Properties of Aqueous Solutions

A **solution** is a homogeneous mixture of two or more substances.

A **solvent** is the dissolving medium, usually the substance present in the greatest quantity in a solution.

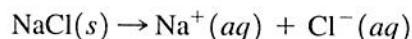
**Aqueous solutions** are solutions in which water is the solvent.

**Solutes** are other substances in the solution.

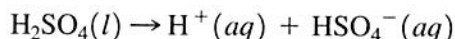
An **electrolyte** is a substance whose aqueous solution contains ions. The solution conducts electricity because the ions are free to migrate throughout the solution.

**Strong electrolytes** are substances that exist in solution, completely ionized. Ionic

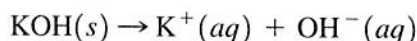
compounds and some molecular compounds called strong acids are strong electrolytes. For example, when solid sodium chloride dissolves in water, the ions dissociate completely:



Similarly, the **strong acid**, sulfuric acid,  $\text{H}_2\text{SO}_4$ , ionizes completely in aqueous solution:



**Strong bases**, such as potassium hydroxide, also dissociate completely:



### Your Turn 4.1

Write an equation for the ionization of gaseous hydrogen chloride in water. Write your answer in the space provided.

Table 4.1 lists the names and formulas of common strong acids and strong bases.

**Table 4.1** The names and formulas of common strong acids and strong bases.

Strong Acids		Strong Bases	
sulfuric acid*	$\text{H}_2\text{SO}_4$	lithium hydroxide	$\text{LiOH}$
nitric acid	$\text{HNO}_3$	sodium hydroxide	$\text{NaOH}$
perchloric acid	$\text{HClO}_4$	potassium hydroxide	$\text{KOH}$
chloric acid	$\text{HClO}_3$	rubidium hydroxide	$\text{RbOH}$
hydrochloric acid	$\text{HCl}$	cesium hydroxide	$\text{CsOH}$
hydrobromic acid	$\text{HBr}$	calcium hydroxide**	$\text{Ca}(\text{OH})_2$
hydroiodic acid	$\text{HI}$	strontium hydroxide**	$\text{Sr}(\text{OH})_2$
		barium hydroxide**	$\text{Ba}(\text{OH})_2$

\*Sulfuric acid is a diprotic acid (it has two  $\text{H}^+$  ions), and only the first  $\text{H}^+$  ion ionizes completely.

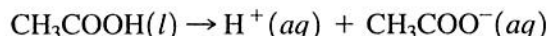
\*\* $\text{Ca}(\text{OH})_2$ ,  $\text{Sr}(\text{OH})_2$ , and  $\text{Ba}(\text{OH})_2$  are dibasic (they each have two  $\text{OH}^-$  ions).



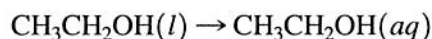


**Common misconception:** The terms “ionize” and “dissociate” are often used interchangeably by many texts, even though their meanings differ in a subtle way. Both imply that ions exist in solution when a solute dissolves. When an ionic compound dissolves, dissociation of ions already present occurs. An electrolyte that is a covalent compound ionizes in solution because the ions are not present in the pure compound.

**Weak electrolytes** exist mostly as molecules in solution, with a small fraction in the form of ions. Molecular compounds called weak acids and weak bases are weak electrolytes. For example, acetic acid,  $\text{CH}_3\text{COOH}$ , and ammonia,  $\text{NH}_3$ , are weak electrolytes. Only about 1% or less of each molecular compound's molecules ionize in aqueous solution. The percent dissociation is concentration dependent.



A **nonelectrolyte** is a substance that does not form ions in solution and its solution does not conduct electricity. A nonelectrolyte usually consists of a molecular compound, which when dissolved in water, usually consists of intact, un-ionized molecules. For example, when ethanol dissolves in water, its molecules remain intact:



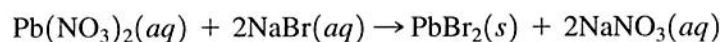
### Your Turn 4.2

←  
Classify each of these compounds as strong, weak, or nonelectrolytes: calcium chloride; ammonium sulfate; hydrocyanic acid,  $\text{HCN}$ ; glucose,  $\text{C}_6\text{H}_{12}\text{O}_6$ . Explain your reasoning. Write your answer in the space provided.

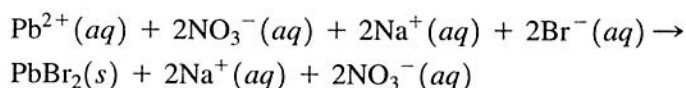
## Section 4.2

### Precipitation Reactions

A **precipitate** is an insoluble solid formed by a reaction in solution. For example, aqueous lead(II) nitrate reacts with aqueous sodium bromide to form solid lead(II) bromide and sodium nitrate. The **complete equation** for the reaction is:

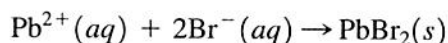


If the soluble strong electrolytes are shown as ions, a more accurate representation of the reaction is a **complete ionic equation**:



**Spectator ions** are ions that appear in identical form among both reactants and products of a complete ionic equation.

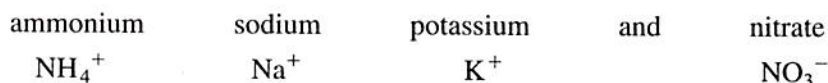
A **net ionic equation** omits spectator ions because they do not change from reactants to products:



### Predicting Products of Precipitation Reactions

To predict the products of precipitation reactions and to write their net ionic equations, use the following simplified guidelines:

The cations and anions that generally do not form precipitates are:



Therefore, the formulas of precipitates will not contain these four ions. There are no common exceptions. Ammonium, sodium, potassium, and nitrate ions are usually spectator ions in aqueous solution.

When predicting the products of precipitation reactions and writing their net ionic equations, apply the rules in Table 4.2. Make sure that the formulas for any precipitates never contain the above ions.

**Table 4.2** General rules for writing net ionic equations.

1.	Write all the reactants that are indicated to be solids (s), liquids (l), or gases (g).
2.	Rewrite the formulas of the aqueous reactants, omitting the spectator ions ( $\text{NH}_4^{+}$ , $\text{Na}^{+}$ , $\text{K}^{+}$ , and $\text{NO}_3^{-}$ ).
3.	Predict and write the product(s).
4.	If necessary, use ions to balance the mass and charge. Then balance using coefficients. Check to see if the charges are balanced.

#### Examples:

Write a net ionic equation for each of the following laboratory situations. Assume a reaction occurs in all cases. (Note: Although the Advanced Placement Exam does not require students to show the phases of reactants or products [g, l, s, aq], the phases are indicated in these examples for clarity.)

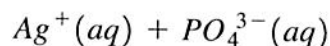


a. Aqueous solutions of silver nitrate and sodium phosphate are mixed.

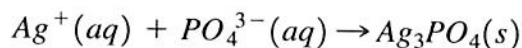
**Solution:**

1. No solids, liquids, or gases are indicated as reactants.

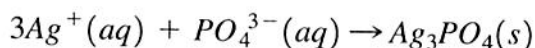
2. Nitrate and sodium ions are spectator ions, so omit them.



3. To predict the product, combine the ions to make a correct formula.



4. The  $\text{Ag}^+$  ion requires a coefficient of 3 to balance the mass and charge:



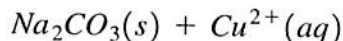
b. Aqueous copper(II) nitrate is added to solid sodium carbonate.

**Solution:**

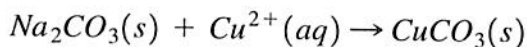
1. Sodium carbonate is the only solid indicated.



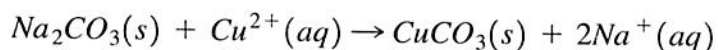
2. Omit nitrate, a spectator ion.



3. Because sodium ion is a spectator, copper(II) ion replaces sodium ion.



4. Two sodium ions added to the right side of the equation balances the mass and charge. Notice that, in this case,  $\text{Na}^+$  is not a spectator ion because it changes form from solid to aqueous during the reaction.



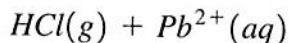
c. Hydrogen chloride gas is bubbled through an aqueous solution of lead(II) nitrate.

**Solution:**

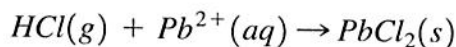
1. The statement identifies  $\text{HCl}$  as a gas.



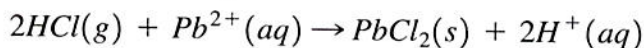
2. Nitrate is a spectator ion.



3. Combine lead(II) and chloride ions to form a precipitate.



4. Place two  $\text{H}^+$  ions on the right side of the equation to balance the mass and charge.



Write a net ionic equation for the reaction of aqueous sodium carbonate with solid nickel(II) nitrate. Write your answer in the space provided.

Your Turn 4.3

## Acids, Bases, and Neutralization Reactions

### Section 4.3

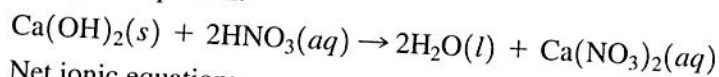
**Acids** are substances that ionize in aqueous solution to form  $\text{H}^+$  ions.

**Bases** are substances that react with  $\text{H}^+$  ions.

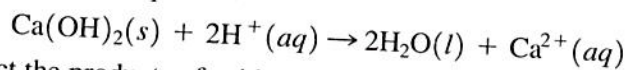
**Strong acids and strong bases are strong electrolytes.** They completely ionize in solution. Table 4.1 lists the names and formulas of common strong acids and bases.

**Weak acids and bases are weak electrolytes.** They ionize only slightly in water solution. For example, a neutralization reaction between an acid and a metal hydroxide produces water and a salt:

Complete equation:



Net ionic equation:



To predict the products of acid–base reactions and to write their net ionic equations, keep in mind that water is usually a product of neutralization reactions and apply the rules in Table 4.2 for writing net ionic equations. Because strong acids and bases are strong electrolytes, their ions are usually spectator ions in acid–base reactions.

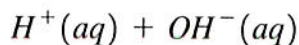
#### Examples:

Write a net ionic equation for each of the following laboratory situations. Assume a reaction occurs in all cases.

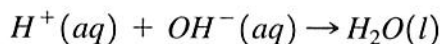
- Aqueous hydrochloric acid is mixed with a solution of sodium hydroxide.

#### Solution:

- Neither reactant is indicated as a solid, liquid, or gas.
- Both chloride (hydrochloric acid is a strong electrolyte) and sodium ions are spectator ions.



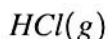
3. Water is a product of a neutralization reaction.



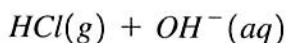
b. Gaseous hydrogen chloride is bubbled through a solution of sodium hydroxide.

**Solution:**

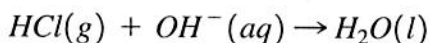
1. HCl is indicated as a gas.



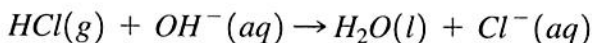
2. Sodium ion is a spectator ion.



3. Water is a product of a neutralization reaction.



4. One chloride ion on the right side of the equation balances the mass and charge.



Notice that the reactants in Questions a and b are the same but, because of the difference in phase of hydrochloric acid, the net ionic equations are different.

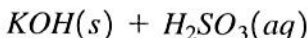
c. Solid potassium hydroxide is mixed with aqueous sulfurous acid.

**Solution:**

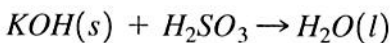
1. KOH is identified as a solid.



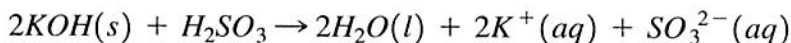
2. Sulfurous acid is a weak acid (it does not appear on the list of strong acids in Table 4.1), so it is shown in its molecular form.



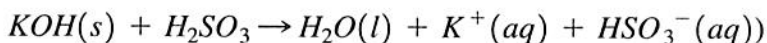
3. Water is a product of a neutralization reaction.



4. When the  $H^+$  ion of the acid reacts with the base, free sulfite ion is a product and  $2K^+$  are needed to balance the mass and charge.



(Note: Because this problem does not specify the relative amounts of acid and base added, it is also correct to assume a one-to-one mole ratio. The following equation is also acceptable:





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Your Turn 4.4

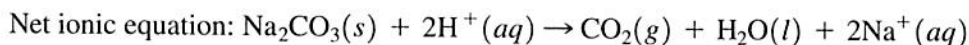
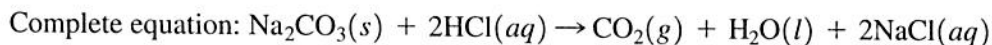
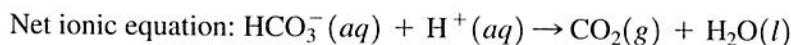
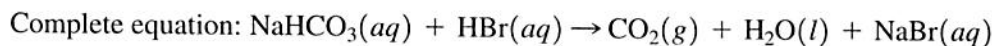
Write a net ionic equation for the reaction of aqueous potassium hydroxide with aqueous acetic acid. Write your answer in the space provided.

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**Acid–Base Reactions and Gas Formation**

Ionic compounds containing carbonate, sulfite, or sulfide ions produce gases when they react with acids.

Metal carbonates and metal hydrogen carbonates react with acids to produce carbon dioxide gas and water.

**Examples:**

Write a net ionic equation for each of the following laboratory situations. Assume a reaction occurs in all cases.

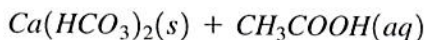
- a. Calcium hydrogen carbonate solid is added to an aqueous solution of acetic acid.

**Solution:**

1. Calcium hydrogen carbonate is identified as a solid.



2. Acetic acid is a weak acid because it is not listed among the strong acids in Table 4.1. So it is written in molecular form.

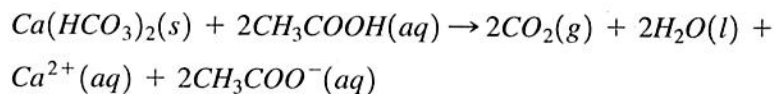


3. The products are carbon dioxide and water.





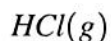
4. Calcium ions and acetate ions are needed to balance the mass and charge.



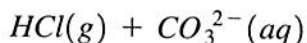
- b. Gaseous hydrogen chloride is bubbled through an aqueous solution of potassium carbonate.

**Solution:**

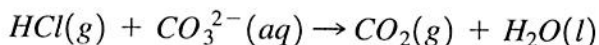
1. HCl is identified as a gas.



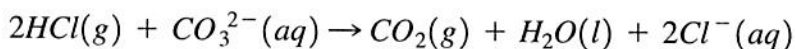
2. Potassium ion is a spectator ion.



3. The products are carbon dioxide gas and water.



4. Balance the mass and charge using two molecules of HCl and two chloride ions.



**Your Turn 4.5**

Write a net ionic equation for the reaction of aqueous nitric acid with solid lithium carbonate. Write your answer in the space provided.

**Section 4.4**

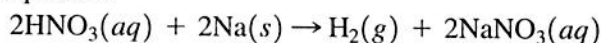
**Oxidation–Reduction Reactions**

**Oxidation–reduction reactions** (also called **redox reactions**) are reactions that transfer electrons between reactants.

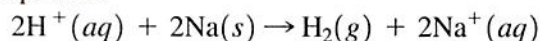
For example, acids react with active metals to produce hydrogen gas.

An acid reacts with a metal to produce hydrogen gas:

Complete equation:



Net ionic equation:



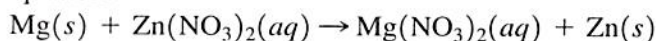
In the reaction, electrons are transferred from sodium atoms to hydrogen ions. The sodium atoms lose one electron each to become sodium ions. The hydrogen ions gain one electron each to become a diatomic hydrogen molecule.

**Oxidation** is the loss of electrons. In this case, sodium atoms lose electrons and are oxidized.

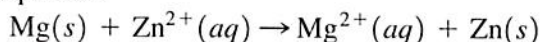
**Reduction** is the gain of electrons. In this case, hydrogen ions gain electrons and are reduced.

Metals also react with metal ions to exchange electrons:

Complete equation:



Net ionic equation:



The metal is oxidized and the metal ion is reduced.

An **activity series** is a list of metals arranged in order of decreasing ease of oxidation. Any metal on the list can be oxidized by ions below it. Table 4.3 shows an activity series for metals in aqueous solution. Notice that the metals copper, silver, and gold are below the  $\text{H}_2 \rightarrow 2\text{H}^+$  reaction. This means that Cu, Ag, and Au will not dissolve in acid solution, whereas other metals listed will react with acids.

**Table 4.3** An activity series. Metals at the top of the list are most likely to be oxidized (lose electrons).

Li(s)	→	Li <sup>+</sup>	(aq)	+	1 e <sup>-</sup>
K(s)	→	K <sup>+</sup>	(aq)	+	1 e <sup>-</sup>
Ca(s)	→	Ca <sup>2+</sup>	(aq)	+	2 e <sup>-</sup>
Na(s)	→	Na <sup>+</sup>	(aq)	+	1 e <sup>-</sup>
Mg(s)	→	Mg <sup>2+</sup>	(aq)	+	2 e <sup>-</sup>
Al(s)	→	Al <sup>3+</sup>	(aq)	+	3 e <sup>-</sup>
Zn(s)	→	Zn <sup>2+</sup>	(aq)	+	2 e <sup>-</sup>
Fe(s)	→	Fe <sup>2+</sup>	(aq)	+	2 e <sup>-</sup>
Ni(s)	→	Ni <sup>2+</sup>	(aq)	+	2 e <sup>-</sup>
Pb(s)	→	Pb <sup>2+</sup>	(aq)	+	2 e <sup>-</sup>
H <sub>2</sub> (g)	→	2H <sup>+</sup>	(aq)	+	2 e <sup>-</sup>
Cu(s)	→	Cu <sup>2+</sup>	(aq)	+	2 e <sup>-</sup>
Ag(s)	→	Ag <sup>+</sup>	(aq)	+	1 e <sup>-</sup>
Au(s)	→	Au <sup>3+</sup>	(aq)	+	3 e <sup>-</sup>



## Your Turn 4.6

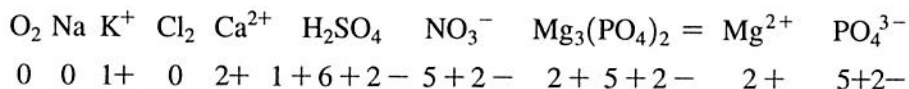
Aluminum metal and nickel metal are placed in a solution containing 1.0 M  $\text{Ni}(\text{NO}_3)_2$  and 1.0 M  $\text{Al}(\text{NO}_3)_3$ . Write a net ionic equation for the reaction that will ensue. Tell which reactant loses electrons and which gains electrons. Write your answer in the space provided.

**Oxidation state**, also called **oxidation number**, is a positive or negative whole number assigned to an element in a chemical formula based on a set of formal rules. The oxidation state is used to track electron transfer in redox reactions. Table 4.4 lists the rules for assigning oxidation numbers to elements in chemical formulas.

**Table 4.4** Simplified rules for determining oxidation numbers.

1. The oxidation number of combined oxygen is usually 2−, except in the peroxide ion,  $\text{O}_2^{2-}$ , where the oxidation number of oxygen is 1−.  
Examples: In  $\text{H}_2\text{O}$  and  $\text{H}_2\text{SO}_4$ , the oxidation state of oxygen is 2−. In  $\text{H}_2\text{O}_2$  and  $\text{BaO}_2$ , O is 1−.
2. The oxidation number of combined hydrogen is usually 1+, except in the hydride ion,  $\text{H}^-$ , where it is 1−.  
Examples: In  $\text{H}_2\text{O}$  and  $\text{H}_2\text{SO}_4$ , the oxidation state of hydrogen is 1+. In  $\text{NaH}$  and  $\text{CaH}_2$ , H is 1−.
3. The oxidation numbers of all individual atoms of a formula add to the charge on that formula. When in doubt, separate ionic compounds into common cation–anion pairs.

Examples:



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What is the oxidation number of each atom in potassium aluminum sulfate,  $KAl(SO_4)_2$ ? Write your answer in the space provided.

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→ Your Turn 4.7

## Concentrations of Solutions

### Section 4.5

A **solution** is a homogeneous mixture consisting of a solvent and one or more solutes.

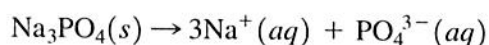
**Concentration** is the amount solute dissolved in a given amount of solvent or solution.

**Molar concentration**, also called **molarity** (abbreviated M), is the number of moles of solute dissolved in a liter of solution.

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

For example, 1.5 M  $Na_3PO_4(aq)$  means that every liter of solution contains 1.5 moles of sodium phosphate.

Because  $Na_3PO_4$  is a strong electrolyte, its ions dissociate completely in aqueous solution:



A 1.00 M solution of  $Na_3PO_4$  contains 3.00 M  $Na^+$  ions and 1.00 M  $PO_4^{3-}$  ions.

**Dilution**, adding water to a solution, decreases the molar concentration of each substance in the solution by a factor called the dilution factor. The dilution factor is the ratio of the original volume to the new volume.

#### Example:

*Enough water is added to 500 mL of 1.00 M  $Na_3PO_4$  to make the final volume 800 mL. What are the molar concentrations of each ion upon dilution?*

#### Solution:

1.00 M  $Na_3PO_4$  = 3.00 M  $Na^+$  ions and 1.00 M  $PO_4^{3-}$  ions.

$(3.00 \text{ M } Na^+ \text{ ions})(500 \text{ mL}/800 \text{ mL}) = 1.88 \text{ M } Na^+$ .

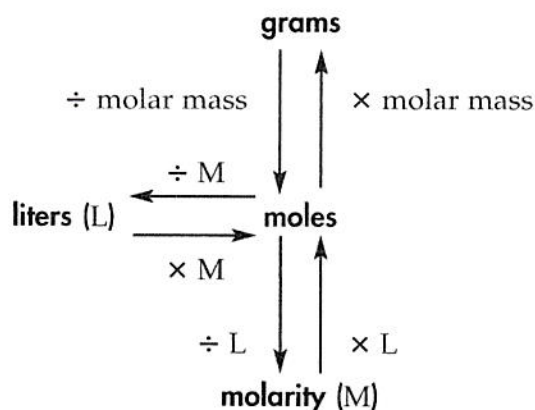
$(1.00 \text{ M } PO_4^{3-} \text{ ions})(500 \text{ mL}/800 \text{ mL}) = 0.625 \text{ M } PO_4^{3-}$ .



## Section 4.6

**Solution Stoichiometry and Chemical Analysis**

**Solution stoichiometry** involves calculations that relate moles of reactants and products to the volumes of solutions and their molar concentrations. Figure 4.1 illustrates the mole road for converting grams, moles, molarity, and liters.



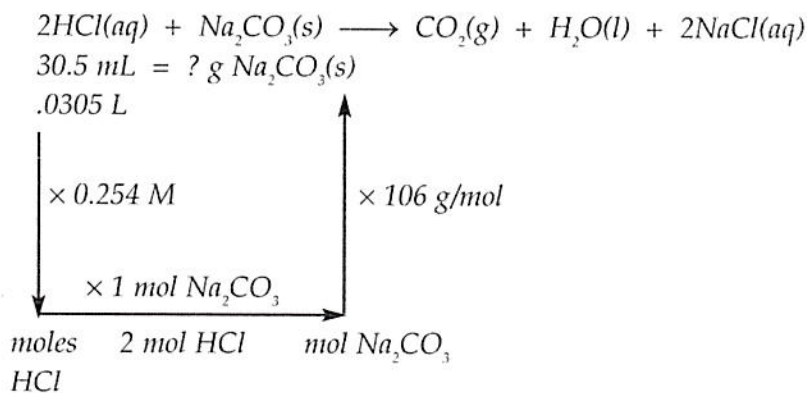
**Figure 4.1** Converting moles, grams, molarity, and liters.

**Example:**

How many grams of solid sodium carbonate are in a sample if it takes 30.5 mL of 0.254 M hydrochloric acid to completely react with the sample?

**Solution:**

Write and balance a chemical equation, convert mL to L by dividing by 1000, and then follow the road map to the solution.



$$\begin{aligned} \times \text{ g Na}_2\text{CO}_3 &= (30.5 \text{ mL HCl})(1000 \text{ mL/L})(0.254 \text{ mol l/L}) \\ &\quad (1 \text{ mol Na}_2\text{CO}_3/2 \text{ mol HCl})(106 \text{ g/mol}) = 0.411 \text{ g Na}_2\text{CO}_3 \end{aligned}$$

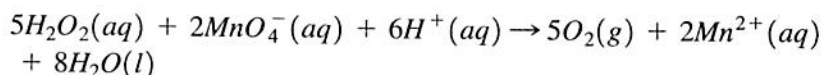
A **titration** is an analytical technique used to determine the unknown concentration of a solution by reacting it with a solution of known concentration called a standard solution.

The equivalence point of the titration is the point at which the moles of substance dissolved in the unknown solution completely react with the moles of substance in the standard solution.

The endpoint is the point at which an indicator changes color. The endpoint is designed to coincide closely to the equivalence point. From the results of the experiment the concentration of the unknown solution can be calculated.

**Example:**

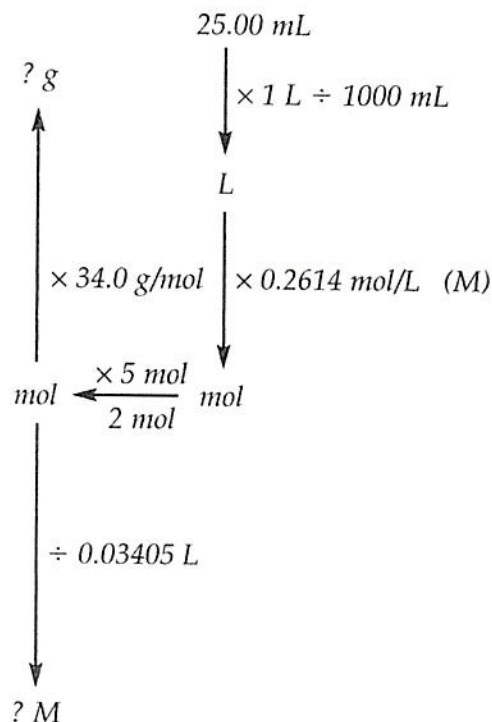
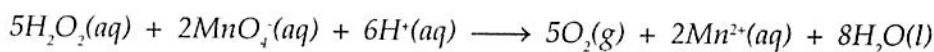
The percentage of hydrogen peroxide is to be determined by titration with a standard solution of potassium permanganate. The standard solution is prepared by dissolving 20.65 g of  $\text{KMnO}_4$  in enough water to make 500 mL of solution. Exactly 25.00 mL of this solution is titrated with an unknown solution of hydrogen peroxide. The endpoint is reached upon addition of 34.05 mL of  $\text{H}_2\text{O}_2$ . The density of the  $\text{H}_2\text{O}_2$  solution is 1.05 g per milliliter. The reaction is:



- What is the molar concentration of the standard  $\text{KMnO}_4$  solution?
- What is the molar concentration of the  $\text{H}_2\text{O}_2$  solution?
- How many grams of  $\text{H}_2\text{O}_2$  are used in the experiment?
- What is the percent of hydrogen peroxide in the solution?

**Solution:**

- Molarity of  $\text{KMnO}_4$  solution = moles of  $\text{KMnO}_4$ /liters solution  
 $= (20.65 \text{ g} / 158.0 \text{ g/mol}) / 0.5000 \text{ L} = 0.2614 \text{ M } \text{KMnO}_4$
- Follow the mole road from 25.00 mL of  $\text{KMnO}_4$  solution to molarity of  $\text{H}_2\text{O}_2$  solution:





Molarity of  $H_2O_2$  =

$$(25.00 \text{ mL} / 1000 \text{ mL/L})(0.2614 \text{ mol/L})(5 \text{ mol } H_2O_2 / 2 \text{ mol } MnO_4^-) / (0.03405 \text{ L}) = 0.4798 \text{ M}$$

c. On the road map, go from moles of  $H_2O_2$  to grams of  $H_2O_2$ .

$$(25.00 / 1000)(0.2614)(5/2) \text{ mol } H_2O_2 (34.0 \text{ g/mol}) = 0.555 \text{ g } H_2O_2$$

d.  $\% H_2O_2 = (g H_2O_2 / g \text{ solution})(100)$   $g \text{ solution} = 34.05 \text{ mL} \times 1.05 \text{ g/mL} = 35.8 \text{ g}$

$$\% H_2O_2 = (0.555 \text{ g } H_2O_2 / 35.8 \text{ g solution})(100) = 1.55\% H_2O_2$$

**Multiple Choice Questions**

1. Which of the following elements should react most readily with acid?
  - A) potassium
  - B) calcium
  - C) gold
  - D) copper
  - E) magnesium
2. The collection of ions, all of whose members do not commonly form precipitates, is
  - A)  $\text{Hg}_2^{2+}$ ,  $\text{Ag}^+$ ,  $\text{Pb}^{2+}$ ,  $\text{Ba}^{2+}$
  - B)  $\text{PO}_4^{3-}$ ,  $\text{OH}^-$ ,  $\text{S}^{2-}$ ,  $\text{CO}_3^{2-}$
  - C)  $\text{NO}_3^-$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{NH}_4^+$
  - D)  $\text{SO}_4^{2-}$ ,  $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$
  - E)  $\text{CrO}_4^{2-}$ ,  $\text{Cu}^{2+}$ ,  $\text{Fe}^{2+}$ ,  $\text{SO}_4^{2-}$
3. Which substance will NOT form a gas upon mixing with an aqueous acid?
  - A)  $\text{NaHCO}_3(\text{s})$
  - B)  $\text{Mg}(\text{s})$
  - C)  $\text{Ca}(\text{s})$
  - D)  $\text{K}_2\text{CO}_3(\text{s})$
  - E)  $\text{Al}_2\text{O}_3(\text{s})$
4. Which metal will NOT react with aqueous hydrochloric acid?
  - A) Fe
  - B) Al
  - C) Cu
  - D) K
  - E) Ni
5. How many milliliters of 0.40 M  $\text{FeBr}_3$  solution would be necessary to precipitate all of the  $\text{Ag}^+$  from 30 mL of a 0.40 M  $\text{AgNO}_3$  solution?  
$$\text{FeBr}_3(\text{aq}) + 3\text{AgNO}_3(\text{aq}) \rightarrow \text{Fe}(\text{NO}_3)_3(\text{aq}) + 3\text{AgBr}(\text{s})$$
  - A) 10
  - B) 20
  - C) 30
  - D) 60
  - E) 90

6. How many grams of baking soda, sodium hydrogen carbonate, are required to completely neutralize 1.00 L of 6.00 M sulfuric acid that has been spilled on the floor?
- A)  $(1/6.00)(2/1)(84.0)$   
B)  $(6.00)(84.0)$   
C)  $(6.00)(2/1)/(84.0)$   
D)  $(6.00)(2/1)(84.0)$   
E)  $(6.00)(1/2)(84.0)$
7. It takes 37.50 mL of 0.152 M sodium chromate to titrate 25.00 mL of silver nitrate. What is the molarity of the silver nitrate solution?
- A)  $(2)(37.50)(0.152)(25.00)$   
B)  $(25.00)/(37.50)(0.152)(2)$   
C)  $(0.152)(25.00)/(37.50)$   
D)  $(37.50)(0.152)/(25.00)$   
E)  $(2)(37.50)(0.152)/(25.00)$
8. Which substance will react with acid, at room temperature and pressure, to produce carbon dioxide gas?
- A) Mg  
B)  $\text{NaHCO}_3$   
C)  $\text{CH}_3\text{COOH}$   
D) NaOH  
E)  $\text{NH}_3$
9. What is the net ionic equation for the reaction of aqueous solutions of  $\text{CaCl}_2$  and  $\text{Na}_2\text{CO}_3$ ?
- A)  $\text{Ca}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq}) + 2\text{Na}^{+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + 2\text{Na}^{+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq})$   
B)  $\text{Cl}^{-}(\text{aq}) + \text{Na}^{+}(\text{aq}) \rightarrow \text{NaCl}(\text{s})$   
C)  $\text{CaCl}_2(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + 2\text{Cl}^{-}(\text{aq})$   
D)  $\text{Ca}^{2+}(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + 2\text{Na}^{+}(\text{aq})$   
E)  $\text{Ca}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CaCO}_3(\text{s})$
10. Magnesium burns in carbon dioxide to produce carbon and magnesium oxide. What is the ratio of carbon to magnesium oxide in the products?
- A) 1:1  
B) 2:1  
C) 1:2  
D) 2:3  
E) 3:2



**Free Response Questions**

1. A student accidentally spills a 1.00 L bottle of concentrated 18.0 M sulfuric acid on the floor of the laboratory. She attempts to neutralize the spill by pouring a 5.00 kg box baking soda, sodium hydrogen carbonate, onto the acid.
  - a. Write a balanced net ionic equation for the reaction. Assume the concentrated sulfuric acid is 100% pure and not in aqueous solution.
  - b. What is the limiting reactant?
  - c. How many grams of sodium hydrogen carbonate are required to neutralize all the acid?
  - d. How many moles of excess reactant remain after all the limiting reactant has been consumed?
  - e. Would a floor consisting of bare concrete require more, less, or the same amount of baking soda to neutralize the spill? Explain.
2. Three unknown metals are labeled A, B, and C. One is silver, one is aluminum acid, and the other is nickel. Using only aqueous solutions of 0.50 M iron(II) nitrate and 1.0 M hydrochloric acid, write a short, concise experimental procedure, the results of which will be sufficient to identify each of the unknown metals. Tell what you would expect to see and what it means. Write net ionic equations to illustrate your answers.