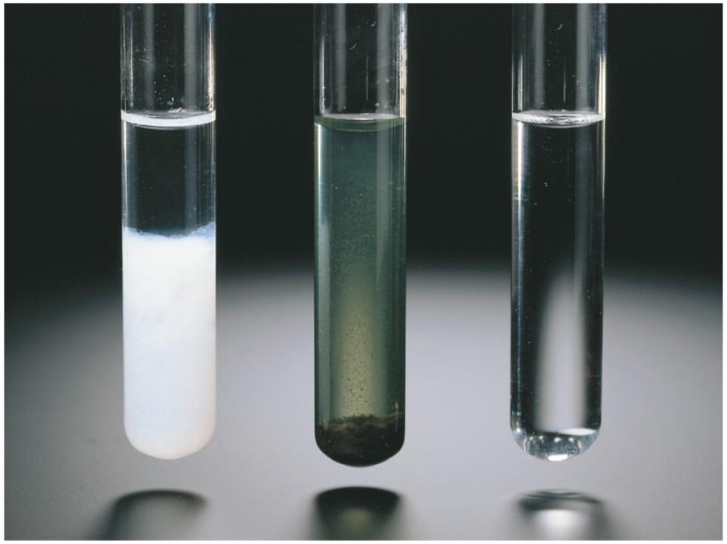
So you have a 30 minute quiz today

What are allowed?

- A pen
- A periodic table
- An A4-sized sheet of hand written notes
- A calculator

No other devices are allowed!

GENERAL CHEMISTRY I



Chapter 4
Reactions in Aqueous Solutions

Contents

4-1 The Nature of Aqueous Solutions

- Water (*Latin:* aqua)
  - Inexpensive
  - Can dissolve a vast number of substances
  - Many substances dissociate into ions
  - Aqueous solutions are found everywhere
    - Seawater
    - Living systems

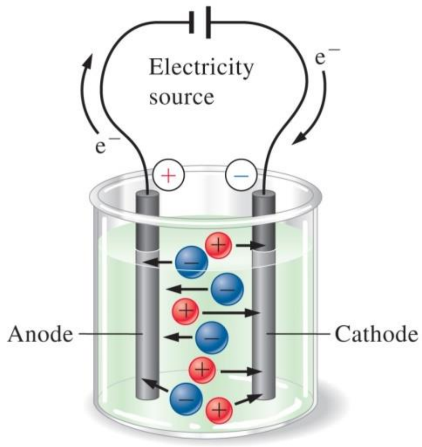
- A **solution** is a **homogeneous mixture** of two or more substances
  - A **solvent** is the component of a solution that is present in the **greatest amount**
  - A **solute** is the component of a solution that is **dissolved** in the solvent

Electrolytes and Nonelectrolytes

- Some solutes can disolve in water and **dissociate** into ions.

➔ Electric charge can be carried.

- **Electrolyte:** A substance that **disolves** in water and **separates** out into cations and anions



- Some solutes can disolve in water but **cannot dissociate** into ions.

➔ Electric charge cannot be carried.

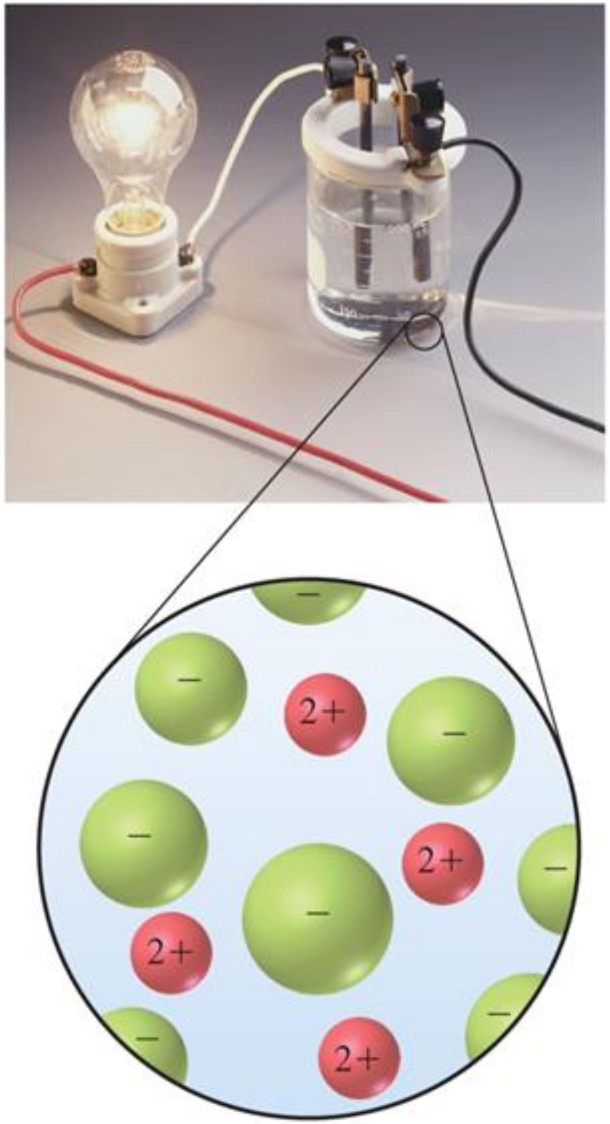
- **Nonelectrolyte**: A substance that **disolves** in water and **does not separates** out into cations and anions

Types of Electrolytes

- *Strong electrolyte* completely dissociates into cation and anion .
  - Good electrical conduction.

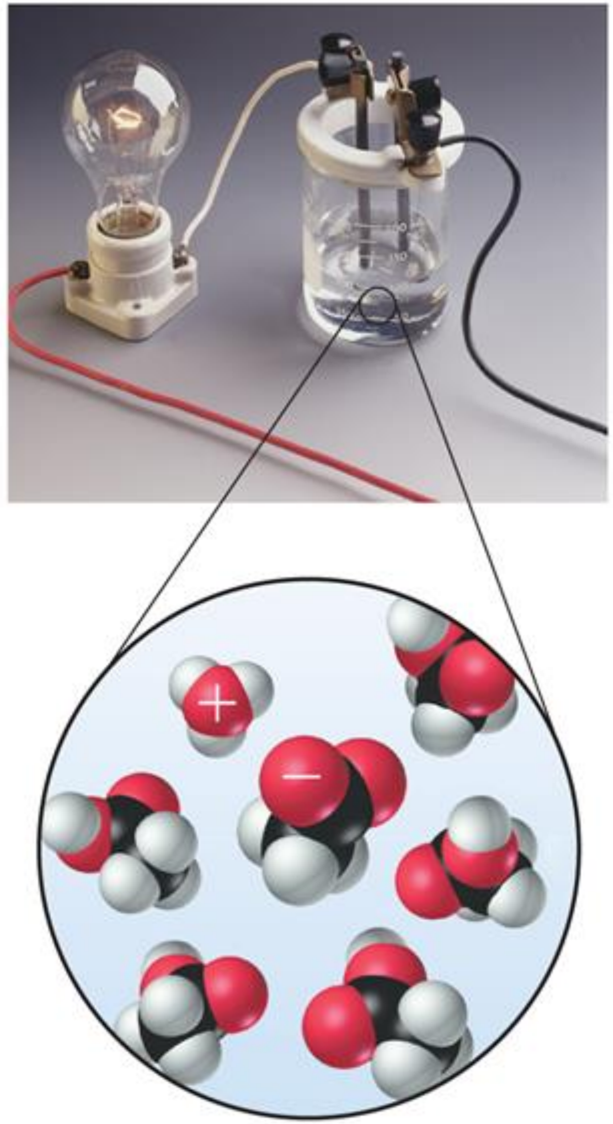
- *Weak electrolyte* partially dissociates into cation and anion .
  - Fair conductor of electricity.

- *Non-electrolyte* does not dissociate.
  - Poor conductor of electricity.



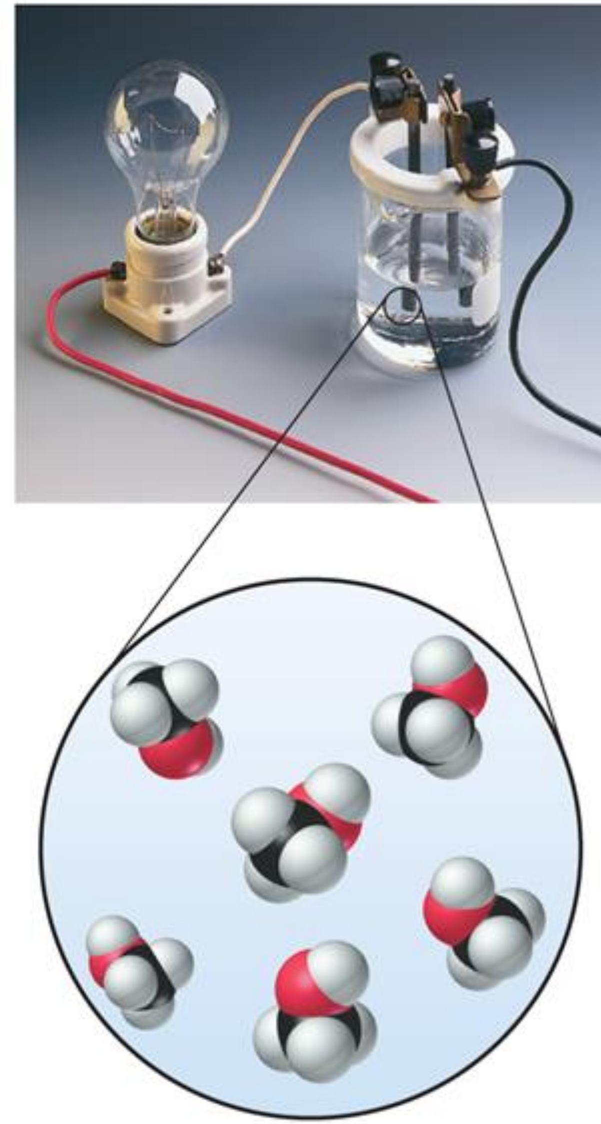
A strong electrolyte

$MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2\ Cl^-(aq)$



A weak electrolyte

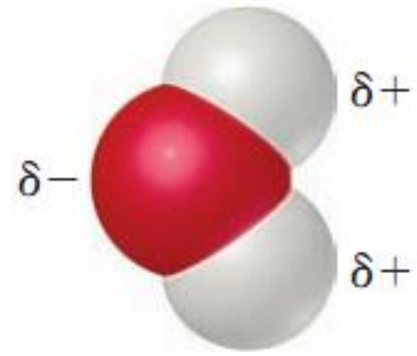
$CH_3COOH(aq) \leftrightharpoons CH_3CO_2^-(aq) + H^+(aq)$



A non-electrolyte

$CH_3OH(aq)$

How Compounds Dissolve in Water

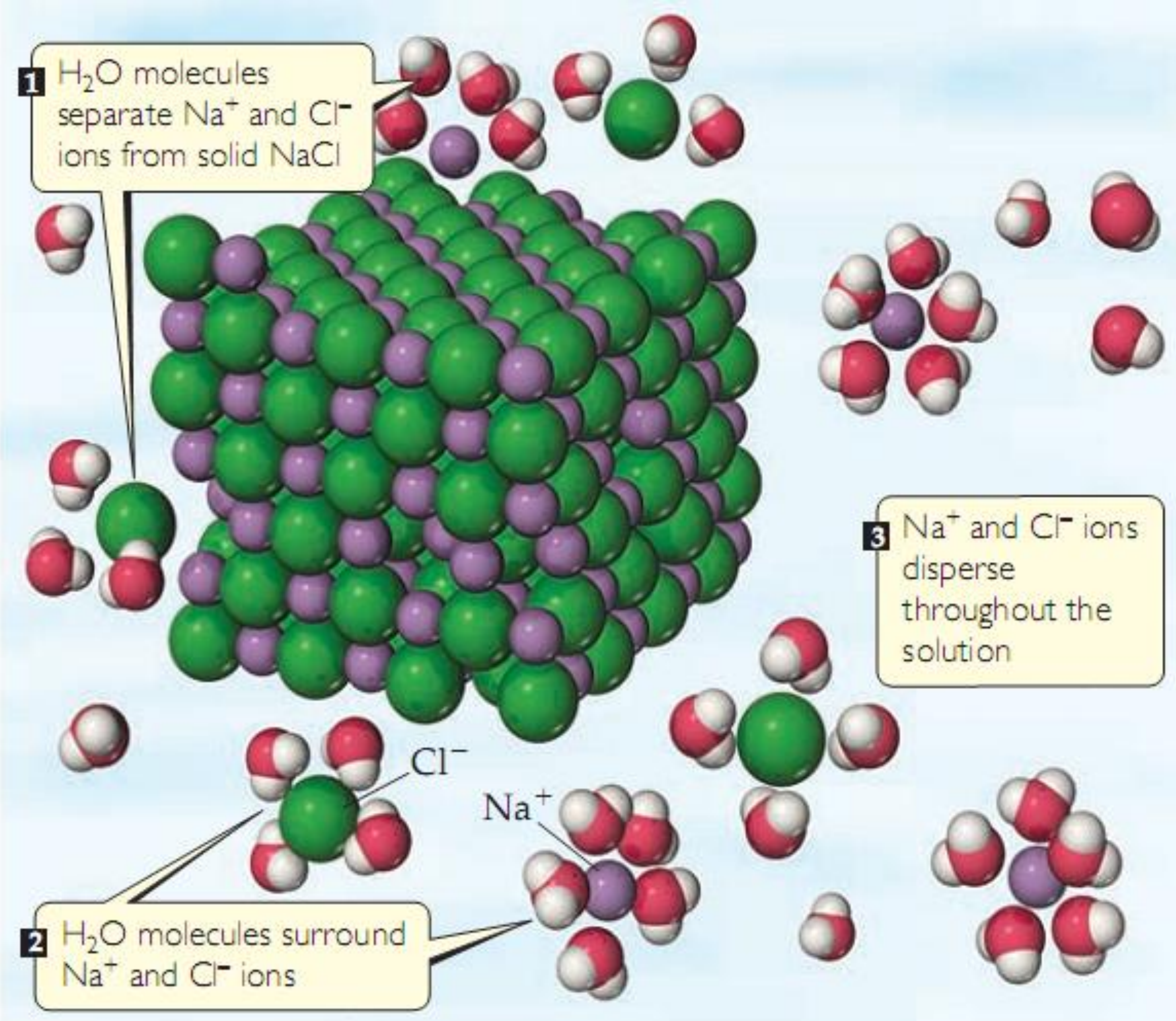


- the O atom is rich in electrons and has a partial negative charge

  ➔ O atoms of water tend to stick to cations and bring them to the solution

- each H atom has a partial positive charge

  ➔ H atoms of water tend to stick to anions and bring them to the solution


1 H₂O molecules separate Na⁺ and Cl⁻ ions from solid NaCl
2 H₂O molecules surround Na⁺ and Cl⁻ ions
3 Na⁺ and Cl⁻ ions disperse throughout the solution
Cl⁻
Na⁺

Dissociation and Ionization

An electrolyte produces ions in solution by one of two processes:

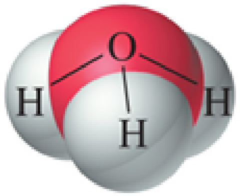
- **dissociation**, the separation of an entity into two or more entities
  - **Ionic** compounds
  - The ions separate from one another and become **surrounded** by water molecules

$$MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2\ Cl^-(aq)$$

- **ionization**, the generation of one or more ions.
  - **Molecular** compounds
  - The ions are generated by a **reaction** of the compound with water

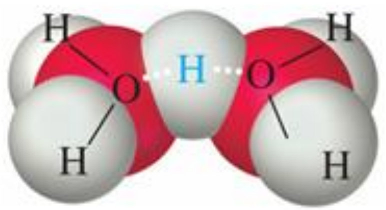
$$CH_3COOH(aq) + H_2O(l) \leftrightarrows CH_3COO^-(aq) + H_3O^+(aq)$$

Solvation: The Hydrated Proton



Hydronium ion
H_3O^+

The hydronium ion, interacts with other water molecules through electrostatic attractions.



A hydrated proton
$H_5O_2^+$

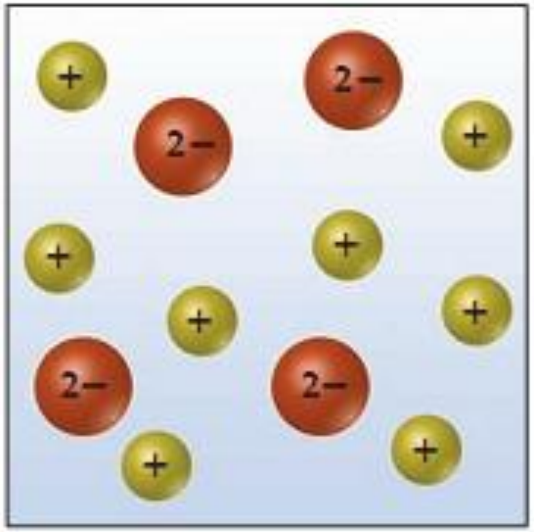
A hydrated proton
$H_9O_4^+$

EXAMPLE

The accompanying diagram represents an aqueous solution of either $MgCl_2$, KCl, or K_2SO_4.

Which solution does the drawing best represent?

$\boxed{K_2SO_4}$



PRACTICE

1. If you have an aqueous solution that contains 1.5 moles of HCl, how many moles of ions are in the solution?

(a) 1.0,
(b) 1.5,
(c) 2.0,
(d) 2.5,
(e) 3.0

2. If you were to draw diagrams representing aqueous solutions of

(a) $NiSO_4$, **(a) 6**
(b) $Ca(NO_3)_2$, **(b) 12**
(c) Na_3PO_4, **(c) 2**
(d) $Al_2(SO_4)_3$, **(d) 9**

how many anions would you show if each diagram contained six cations?

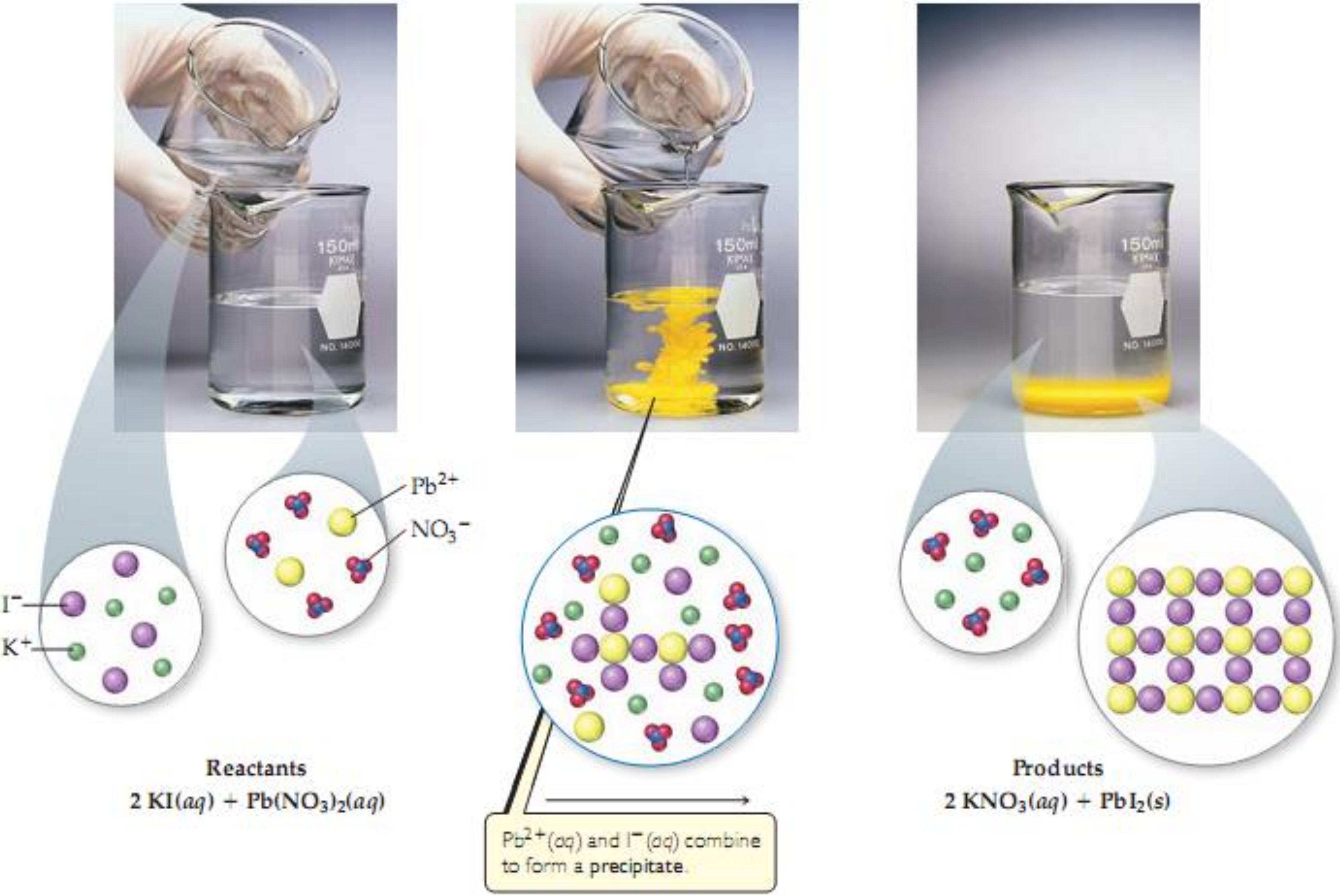
4-2 Precipitation Reactions

- Soluble ions can combine to form an *insoluble* compound.
- Precipitation occurs.
- A test for the presence of chloride ion in water.

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

Precipitation reactions transform ions into an insoluble salt in aqueous solution.

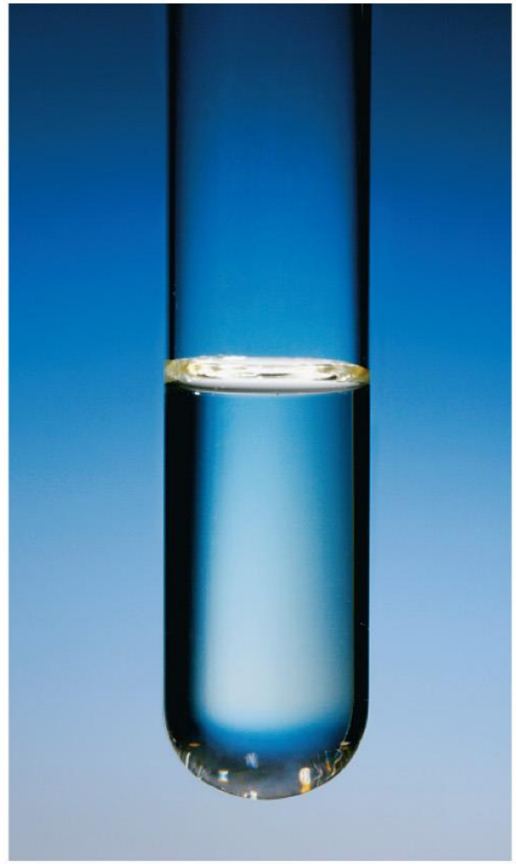
A **precipitate** is an insoluble solid formed by a reaction in solution.



Pb^{2+}
NO_3^-
I^-
K^+
Reactants
2 KI(aq) + Pb(NO$_3$)$_2$(aq)
$Pb^{2+}(aq)$ and $I^-(aq)$ combine to form a precipitate.
Products
2 KNO$_3$(aq) + PbI$_2$(s)

Silver Nitrate and Sodium Iodide

$AgNO_3$(aq)



NaI(aq)

AgI(s)

Na^+(aq) NO_3^-(aq)

Exchange (Metathesis) Reactions

General equation:

$$AX + BY \rightarrow AY + BX$$

Example:

$$AgNO_3\ (aq) + NaI \rightarrow AgI\ (s) + NaNO_3\ (aq)$$

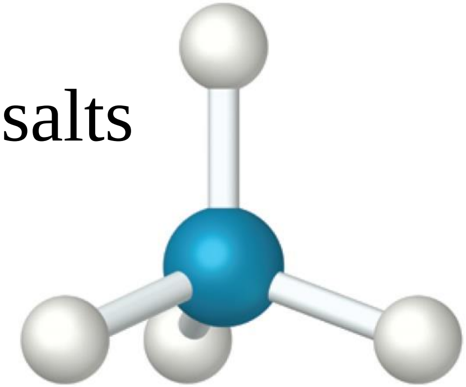
Complete and balance the equation for a metathesis reaction:

1. Use the chemical formulas of the reactants to determine which ions are present.
2. Write the chemical formulas of the products
3. Check the water solubilities of the products. For a precipitation reaction to occur, at least one product must be insoluble in water.
4. Balance the equation.

Solubility Rules

- Compounds that are *soluble*:
  - Alkali metal ion and ammonium ion salts

    Li^+, Na^+, K^+, Rb^+, Cs^+ NH_4^+



Ammonium ion

  - Nitrates, perchlorates and acetates

    NO_3^- ClO_4^- $CH_3CO_2^-$

Solubility Rules

- Compounds that are *mostly soluble*:
  - Chlorides, bromides and iodides Cl^- Br^- I^-
    - Except those of Pb^{2+}, Ag^+, and Hg_2^{2+}.
  - Sulfates SO_4^{2-}
    - Except those of Sr^{2+}, Ba^{2+}, Pb^{2+} and Hg_2^{2+}.
    - $Ca(SO_4)$ is slightly soluble.

Solubility Rules

- Compounds that are *insolube*:
  - Hydroxides and sulfides HO^- S^{2-}
    - Except alkali metal and ammonium salts
    - Sulfides of alkaline earths are soluble
    - Hydroxides of Sr^{2+} and Ca^2+ are slightly soluble.
  - Carbonates and phosphates CO_3^{2-} PO_4^{3-}
    - Except alkali metal and ammonium salts
  - Silver, Lead and Mercury Ag^+ Pb^{2+} Hg_2^{2+}
    - Except nitrates, acetates and perchlorates

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+}

EXAMPLE

Classify these ionic compounds as soluble or insoluble in water:

(a) sodium carbonate, Na_2CO_3,

(b) lead sulfate, $PbSO_4$.

(a) Soluble

(b) Insoluble

(a) Carbonates are insoluble except alkali metal salts

➔ Na_2CO_3 is soluble

(b) Sulfates are mostly soluble except those of Pb^{2+}

➔ $PbSO_4$ is insoluble

EXAMPLE

Given two separate aqueous solutions of $BaCl_2$ and K_2SO_4

(a) Predict the identity of the precipitate that forms when these are mixed.

(b) Write the balanced chemical equation for the reaction.

Solve

(a) The reactants contain Ba^{2+}, Cl^-, K^+, and SO_4^{2-} ions.

Exchanging the anions gives us $BaSO_4$ and KCl.

Most compounds of SO_4^{2-} are soluble but those of Ba^{2+} are not. Thus, $BaSO_4$ is insoluble and will precipitate from solution. KCl is soluble.

(b) $BaCl_2\,(aq) + K_2SO_4\,(aq) \rightarrow BaSO_4\,(s) + 2KCl\,(aq)$

PRACTICE

1. Which of the following compounds is insoluble in water?

(a) $(NH_4)_2S$,

(b) $CaCO_3$,

(c) NaOH,

(d) Ag_2SO_4,

(e) $Pb(CH_3COO)_2$.

(a) s

(b) insoluble

(c) s

(d) s

(e) s

2. Classify the following compounds as soluble or insoluble in water:

(a) cobalt(II) hydroxide,

(b) barium nitrate,

(c) ammonium phosphate.

(a) insoluble

(b) soluble

(c) soluble

PRACTICE

1. Will a precipitate form when solutions of $Ba(NO_3)_2$ and KOH are mixed?

(a) Yes

$\boxed{\text{(b) No}}$

$$Ba(NO_3)_2\ (aq) + KOH\ (aq) \rightarrow \mathbf{Ba(OH)_2\ (aq)} + KNO_3\ (aq)$$

2. When aqueous solutions of $Fe_2(SO_4)_3$ and LiOH are mixed

(a) What compound precipitates ?

(b) Write a balanced equation for the reaction.

$$Fe_2(SO_4)_3\ (aq) + LiOH\ (aq) \rightarrow \mathbf{Fe(OH)_3\ (s)} + Li_2SO_4\ (aq)$$

Net Ionic Equation

Overall Precipitation Reaction:

$$AgNO_3(aq) + NaI(aq) \rightarrow AgI(s) + NaNO_3(aq)$$

Complete ionic equation:

$$Ag^+(aq) + \cancel{NO_3^-(aq)} + \cancel{Na^+(aq)} + I^-(aq) \rightarrow AgI(s) + \overbrace{\cancel{Na^+(aq)} + \cancel{NO_3^-(aq)}}^{\text{Spectator ions}}$$

Net ionic equation:

$$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$$

EXAMPLE

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

Overall Precipitation Reaction:

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

Complete ionic equation:

$$\mathbf{Ca^{2+}(aq) + \cancel{2Cl^-(aq)} + \cancel{2Na^+(aq)} + CO_3^{2-}(aq) \rightarrow \overbrace{CaCO_3(s) + \cancel{2Na^+(aq)} + \cancel{2Cl^-(aq)}}^{\text{Spectator ions}}}$$

Net ionic equation:

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

PRACTICE

1. What happens when you mix an aqueous solution of sodium nitrate with an aqueous solution of barium chloride?

$\boxed{\textbf{(a)}}$ There is no reaction; all possible products are soluble.

(b) Only barium nitrate precipitates.

(c) Only sodium chloride precipitates.

(d) Both barium nitrate and sodium chloride precipitate.

(e) Nothing; barium chloride is not soluble and it stays as a precipitate.

2. Write the net ionic equation for the precipitation reaction that occurs when aqueous solutions of silver nitrate and potassium phosphate are mixed.

$$\boxed{Ag^{+}(aq) + PO_4^{\ 3-}(aq) \rightarrow Ag_3PO_4(s)}$$

4-3 Acid-Base Reactions

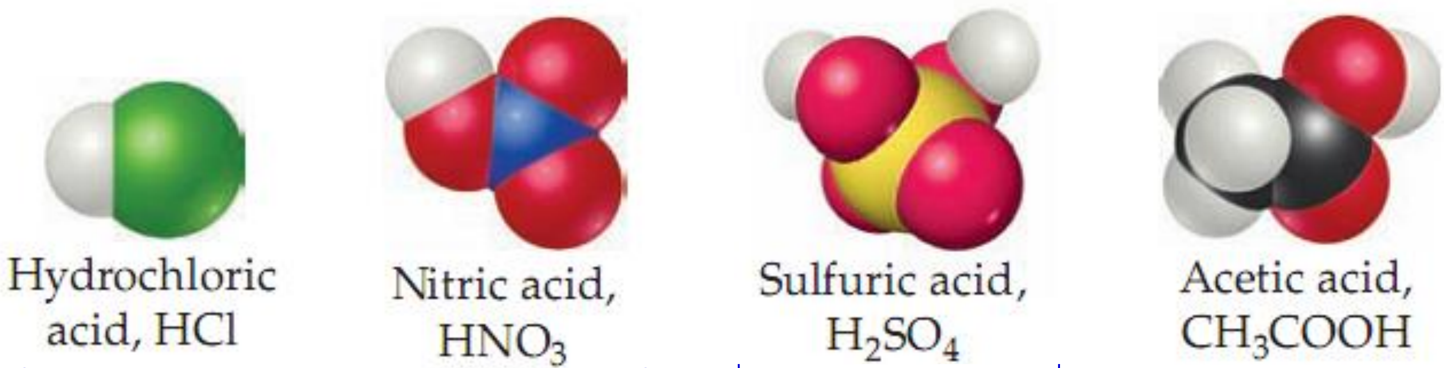
- Latin *acidus* (sour)
  - Sour taste
- Arabic *al-qali* (ashes of certain plants)
  - Bitter taste

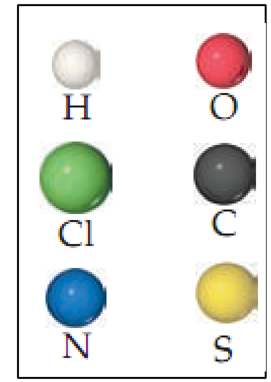
- Svante Arrhenius 1884 Acid-Base theory.

Acids are substances that dissociate in water to yield electrically charged atoms or molecules, called ions, one of which is a hydrogen ion (H^+).

Bases ionize in water to yield hydroxide ions (OH^-)

Acids





monoprotic acids *diprotic acids*

- Acids provide H^+ in aqueous solution.
  - Proton donor
- Strong acids completely ionize:

$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$ *yielding one H^+ per molecule*

$H_2SO_4\ (aq) \rightarrow H^+\ (aq) + HSO_4^-\ (aq)$ *only the first proton completely ionizes*

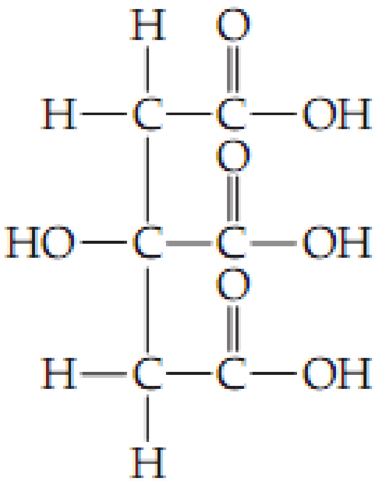
$HSO_4^-\ (aq) \leftrightarrows H^+\ (aq) + SO_4^{2-}\ (aq)$ *a mixture of $H^+(aq)$, $HSO_4^-\ (aq)$ and $SO_4^{2-}(aq)$*

- Weak acid ionization is not complete:

$CH_3COOH\ (aq) \rightarrow H^+\ (aq) + CH_3COO^-\ (aq)$ *yielding one of the four H*

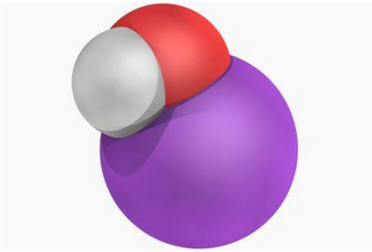
Give It Some Thought

The structural formula of citric acid, a main component of citrus fruits, is

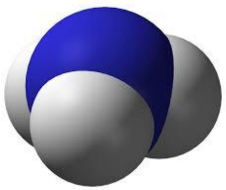


How many $H^+(aq)$ can be generated by each citric acid molecule dissolved in water?

Bases



Sodium hydroxide,
NaOH



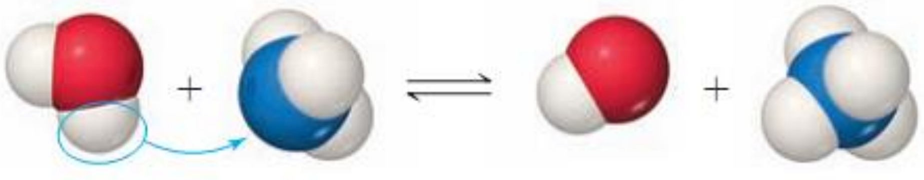
Amonia,
NH_3

- Bases provide OH^- in aqueous solution.
- Strong bases:

  $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

- Weak bases:

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



only about 1% of the NH_3 reacted form a mixture of $NH_3(aq)$, $OH^-(aq)$ and $NH_4^+(aq)$

Recognizing Acids and Bases.

- Acids have ionizable hydrogen ions.
  - CH_3COOH or $HC_2H_3O_2$
- Bases have OH^- combined with a metal ion.

  KOH

or can be identified by chemical equations

$$Na_2CO_3(s) + H_2O(l) \rightarrow HCO_3^-(aq) + 2\ Na^+(aq) + OH^-(aq)$$

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

Why isn't $Al(OH)_3$ classified as a strong base?

a. $Al(OH)_3$ is not basic in water.

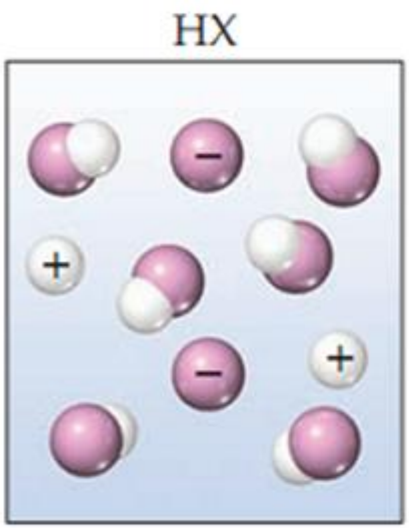
b. $Al(OH)_3$ is insoluble in water.

c. $Al(OH)_3$ is a strong acid in water, not basic.

d. $Al(OH)_3$ is a weak acid in water, not basic.

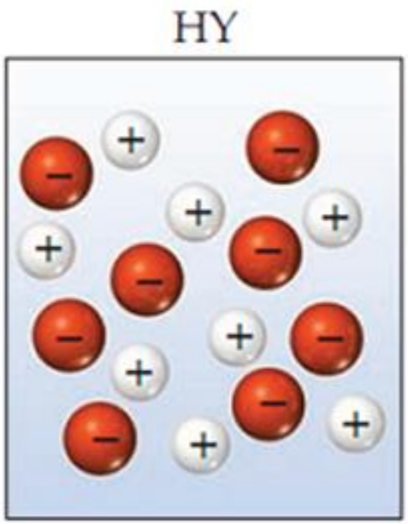
EXAMPLE

The following diagrams represent aqueous solutions of acids HX, HY, and HZ, with water molecules omitted for clarity. Rank the acids from strongest to weakest.

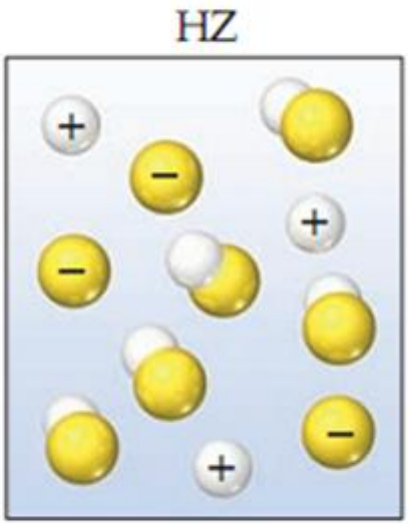


a mixture of molecules and ions

2 H+ ions 6 molecules



totally ionized



a mixture of molecules and ions

3 H+ ions 5 molecules

HY > HZ > HX

PRACTICE

A set of aqueous solutions are prepared containing different acids at the same concentration: acetic acid, chloric acid and hydrobromic acid. Which solution(s) are the most electrically conductive?

(a) chloric acid,

(b) hydrobromic acid,

(c) acetic acid,

(d) both chloric acid and hydrobromic acid,

(e) all three solutions have the same electrical conductivity

Identifying Strong and Weak Electrolytes

	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic	All	None	None
Molecular	Strong acids (see Table 4.2)	Weak acids, weak bases	All other compounds

Examples:

♦ $CaCl_2$	➢ ionic	➢ strong electrolyte
♦ HNO_3	➢ molecular, common strong acid	➢ strong electrolyte
♦ C_2H_5OH (ethanol)	➢ molecular, neither an acid nor a base	➢ nonelectrolyte
♦ HCOOH (formic acid)	➢ molecular, weak acid	➢ weak electrolyte
♦ KOH	➢ ionic	➢ strong electrolyte

PRACTICE

1. Which of these substances, when dissolved in water, is a strong electrolyte?

(a) ammonia,
(b) hydrofluoric acid,
(c) folic acid,
(d) sodium nitrate,
(e) sucrose.

2.Consider solutions in which 0.1 mol of each of the following compounds is dissolved in 1 L of water: $Ca(NO_3)_2$ (calcium nitrate), $C_6H_{12}O_6$ (glucose), $NaCH_3COO$ (sodium acetate), and CH_3COOH (acetic acid). Rank the solutions in order of increasing electrical conductivity, knowing that the greater the number of ions in solution, the greater the conductivity.

$$\boxed{\mathbf{C_6H_{12}O_6 < CH_3COOH < NaCH_3COO < Ca(NO_3)_2}}$$

To be continued..

Neutralization Reactions and Salts

Molecular equation:

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

Complete ionic equation

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

Net ionic equation

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

More Acid-Base Reactions

- Milk of magnesia $Mg(OH)_2$

$$Mg(OH)_2(s) + 2\ H^+(aq) \rightarrow Mg^{2+}(aq) + 2\ H_2O(l)$$

$$Mg(OH)_2(s) + 2\ CH_3COOH(aq) \rightarrow Mg^{2+}(aq) + 2\ CH_3COO^-(aq) + 2\ H_2O(l)$$

More Acid-Base Reactions

- Limestone and marble.

$$CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca^{2+}(aq) + H_2CO_3(aq)$$

But: $H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$

$$CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$$

TABLE 5.3 Some Common Gas-Forming Reactions

Ion	Reaction
HSO_3^-	$HSO_3^- + H^+ \longrightarrow SO_2(g) + H_2O(l)$
SO_3^{2-}	$SO_3^{2-} + 2\,H^+ \longrightarrow SO_2(g) + H_2O(l)$
HCO_3^-	$HCO_3^- + H^+ \longrightarrow CO_2(g) + H_2O(l)$
CO_3^{2-}	$CO_3^{2-} + 2\,H^+ \longrightarrow CO_2(g) + H_2O(l)$
S^{2-}	$S^{2-} + 2\,H^+ \longrightarrow H_2S(g)$
NH_4^+	$NH_4^+ + OH^- \longrightarrow NH_3(g) + H_2O(l)$

EXAMPLE

For the reaction between aqueous solutions of acetic acid (CH_3COOH) and barium hydroxide, $Ba(OH)_2$, write

(a) the balanced molecular equation,

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

(b) the complete ionic equation,

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

(c) the net ionic equation.

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

PRACTICE

1. Which is the correct net ionic equation for the reaction of aqueous ammonia with nitric acid?

(a) $NH_4^+ (aq) + H^+ (aq) \rightarrow NH_5^{2+} (aq)$

(b) $NH_4^+ (aq) + NO_3$- $(aq) \rightarrow NH_2^{2-} (aq) + HNO_3 (aq)$

(c) $NH_2^- (aq) + H^+ (aq) \rightarrow NH_3 (aq)$

(d) $NH_3 (aq) + H^+ (aq) \rightarrow NH_4^+ (aq)$

(e) $NH_4^+ (aq) + NO_3^- (aq) \rightarrow NH_4NO_3 (aq)$

2. For the reaction of phosphorous acid (H_3PO_3) and potassium hydroxide (KOH), write

(a) the balanced molecular equation

$$\mathbf{3\,KOH + H_3PO_3 \rightarrow 3\,H_2O + K_3PO_3}$$

(b) the net ionic equation

$$\mathbf{OH^- + H^+ \rightarrow H_2O}$$

Last week

4-1 The Nature of Aqueous Solutions

4-2 Precipitation Reactions

4-3 Acids, Bases, and Neutralization Reactions

This week

4-4 Oxidation-Reduction Reactions

4-5 Concentrations of Solutions

4-6 Solution Stoichiometry and Chemical Analysis

4-4 Oxidation-Reduction (Redox) Reactions

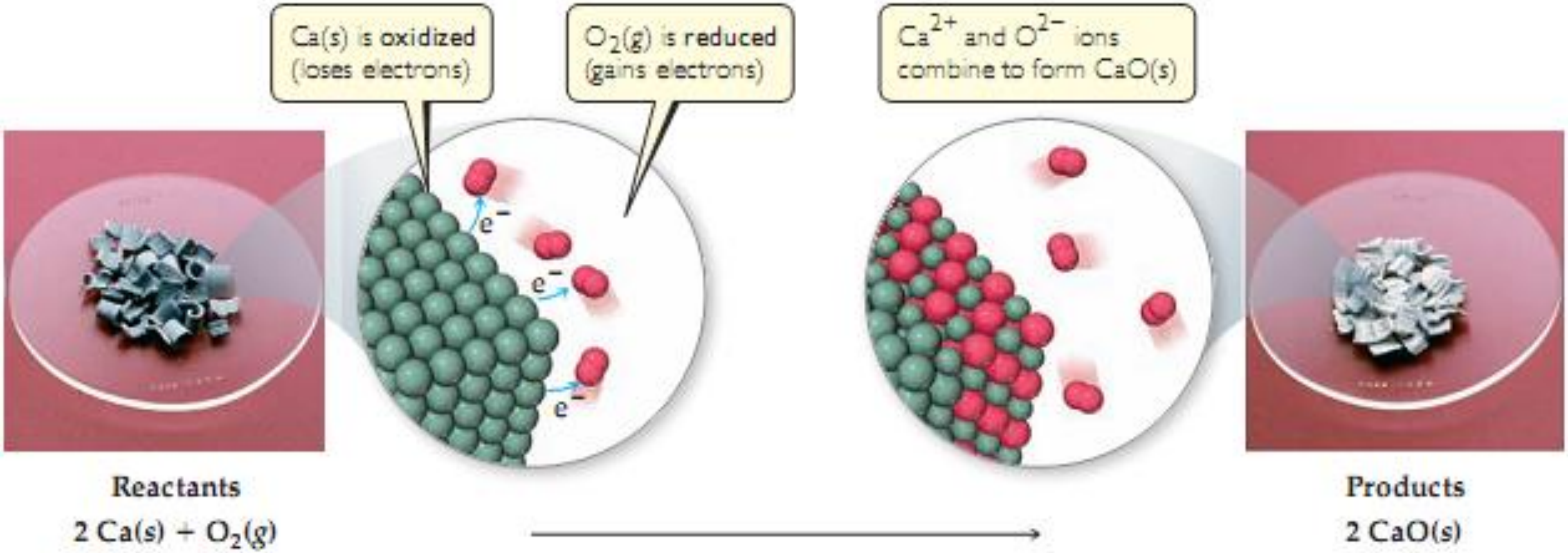
In **precipitation** reactions, **cations and anions** come together to form an **insoluble** ionic compound.

In **neutralization** reactions, **protons are transferred** from one reactant to another.

In **redox** reactions, **electrons are transferred** from one reactant to another

- Loss of electrons by a substance is called oxidation.
- The gain of electrons by a substance is called reduction
- Oxidation and reduction always occur together.

How many electrons does each oxygen atom gain during the course of this reaction?



Initially,

$2Cu + O_2 \rightarrow Cu_2O$ (pink or red)

Then

$2Cu_2O + O_2 \rightarrow 4CuO$ (black)

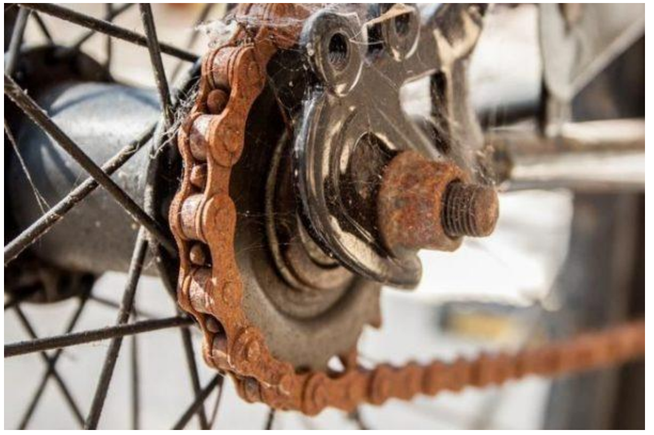
Due to pollution from coal firing:

$Cu + S \rightarrow 4CuS$ (black)

$2CuO + CO_2 + H_2O \rightarrow Cu_2CO_3(OH)_2$ (green)

$3CuO + 2CO_2 + H_2O \rightarrow Cu_3(CO_3)_2(OH)_2$ (blue)

$4CuO + SO_3 + 3H_2O \rightarrow Cu_4SO_4(OH)_6$ (green)



Rust consists of

- hydrated iron(III) oxides $Fe_2O_3.nH_2O$, and
- iron(III) oxide-hydroxide ($FeO(OH)$, $Fe(OH)_3$).

- Reduction of oxygen that is dissolved into water :

$$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$$

- Oxidation of iron in the presence of water:

$$2Fe \xrightarrow{water} 2Fe^{2+}(aq) + 4e^-$$

- Rust is formed in the presence of water via a redox reaction:

$$4\,Fe^{2+} + O_2 \rightarrow 4\,Fe^{3+} + 2\,O^{2-}$$

- Multistep acid–base reactions

$$Fe^{2+} + 2\,H_2O \rightleftharpoons Fe(OH)_2 + 2\,H^+$$
$$Fe^{3+} + 3\,H_2O \rightleftharpoons Fe(OH)_3 + 3\,H^+$$

- Dehydration equilibria

$$Fe(OH)_2 \rightleftharpoons FeO + H_2O$$
$$Fe(OH)_3 \rightleftharpoons FeO(OH) + H_2O$$
$$2\,FeO(OH) \rightleftharpoons Fe_2O_3 + H_2O$$

Oxidation Numbers

1. For an atom in its **elemental form**, the oxidation number is always **zero**.

2. For any **monatomic ion** the oxidation number equals the **ionic charge**.

3. **Nonmetals** usually have **negative** oxidation numbers, although they can sometimes be positive:

   (a) O usually has an oxidation number of -2 in both ionic and molecular compounds, except for peroxides, which contain the O_2^{2-} ion, giving each oxygen an oxidation number of -1.

   (b) H usually has an oxidation number of +1 when bonded to nonmetals and -1 when bonded to metals (or example, metal hydrides such as sodium hydride, NaH).

   (c) F has an oxidation number of -1 in all compounds. The other halogens have an oxidation number of -1 in most binary compounds. When combined with oxygen, as in oxyanions, they have positive oxidation states.

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

Oxidation and Reduction

- Oxidation
  - Oxidation number of some element *increases* in the reaction.
  - Electrons are on the right of the equation
- Reduction
  - Oxidation number of some element *decreases* in the reaction.
  - Electrons are on the left of the equation.

Some General Principles

- Example: Hematite is converted to iron in a blast furnace.

$$Fe_2O_3(s) + 3\ CO(g) \xrightarrow{\Delta} 2\ Fe(l) + 3\ CO_2(g)$$

- Oxidation and reduction always occur together.

Fe^{3+} is reduced to metallic iron.

CO(g) is oxidized to carbon dioxide.

- Assign oxidation number:

$$\overset{3+\ \ 2-}{Fe_2O_3}(s) + 3\ \overset{2+\ 2-}{CO}(g) \xrightarrow{\Delta} 2\ \overset{0}{Fe}(l) + 3\ \overset{4+\ 2-}{CO_2}(g)$$

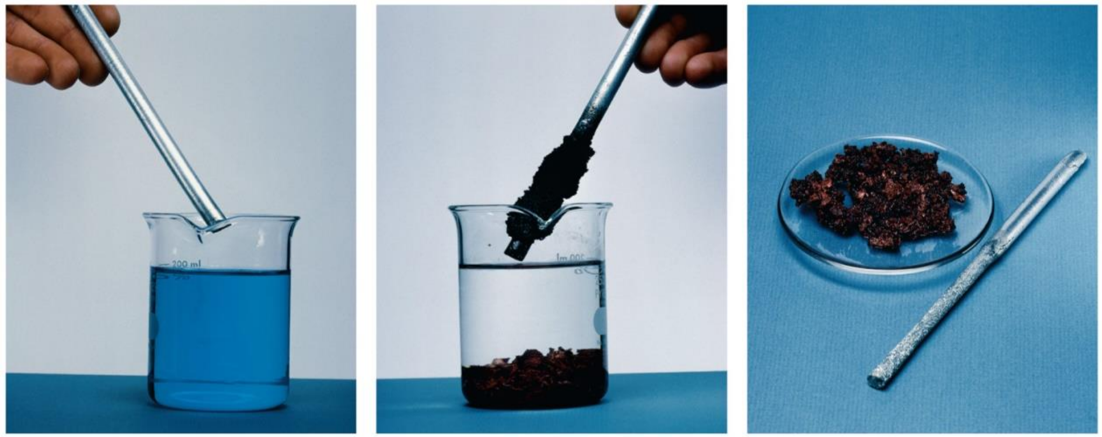
Fe^{3+} is reduced to metallic iron.

CO(g) is oxidized to carbon dioxide.

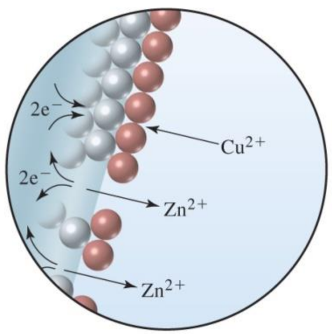
Thermite Reaction

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

An Oxidation Reduction Reaction



(a) (b) (c)



$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

Oxidation and Reduction Half-Reactions

- A reaction represented by two half-reactions.

Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2\ e^{-}$

Reduction: $Cu^{2+}(aq) + 2\ e^{-} \rightarrow Cu(s)$

Overall: $Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$

EXAMPLE

Determine the oxidation number of sulfur in

(a) H_2S,

(a) When bonded to a nonmetal, hydrogen has an oxidation number of +1. Because the H_2S molecule is neutral, the sum of the oxidation numbers must equal zero. Letting x equal the oxidation number of S, we have $2(+1) + x = 0$. Thus, **S has an oxidation number of -2**.

(b) S_8,

(b) Because S_8 is an elemental form of sulfur, **the oxidation number of S is 0**.

(c) SCl_2,

(c) Because SCl_2 is a binary compound, we expect chlorine to have an oxidation number of -1. The sum of the oxidation numbers must equal zero. Letting x equal the oxidation number of S, we have $x + 2(-1) = 0$. Thus, **the oxidation number of S must be +2**.

(d) Na_2SO_3,

(d) Sodium, an alkali metal, always has an oxidation number of +1 in its compounds. Oxygen commonly has an oxidation state of -2. Letting x equal the oxidation number of S, we have $2(+1) + x + 3(-2) = 0$. Therefore, the **oxidation number of S in Na_2SO_3 is +4**.

(e) SO_4^{2-}

(e) The oxidation state of O is -2. The sum of the oxidation numbers equals -2, the net charge of the $SO4^{2-}$ ion. Thus, we have $x + 4(-2) = -2$. Therefore the **oxidation number of S in this ion is +6**.

PRACTICE

1. In which compound is the oxidation state of oxygen -1?

(a) O_2,

(b) H_2O,

(c) H_2SO_4,

(d) H_2O_2,

(e) KCH_3COO

2. What is the oxidation state of the boldfaced element in

(a) $\mathbf{P}_2O_5$,

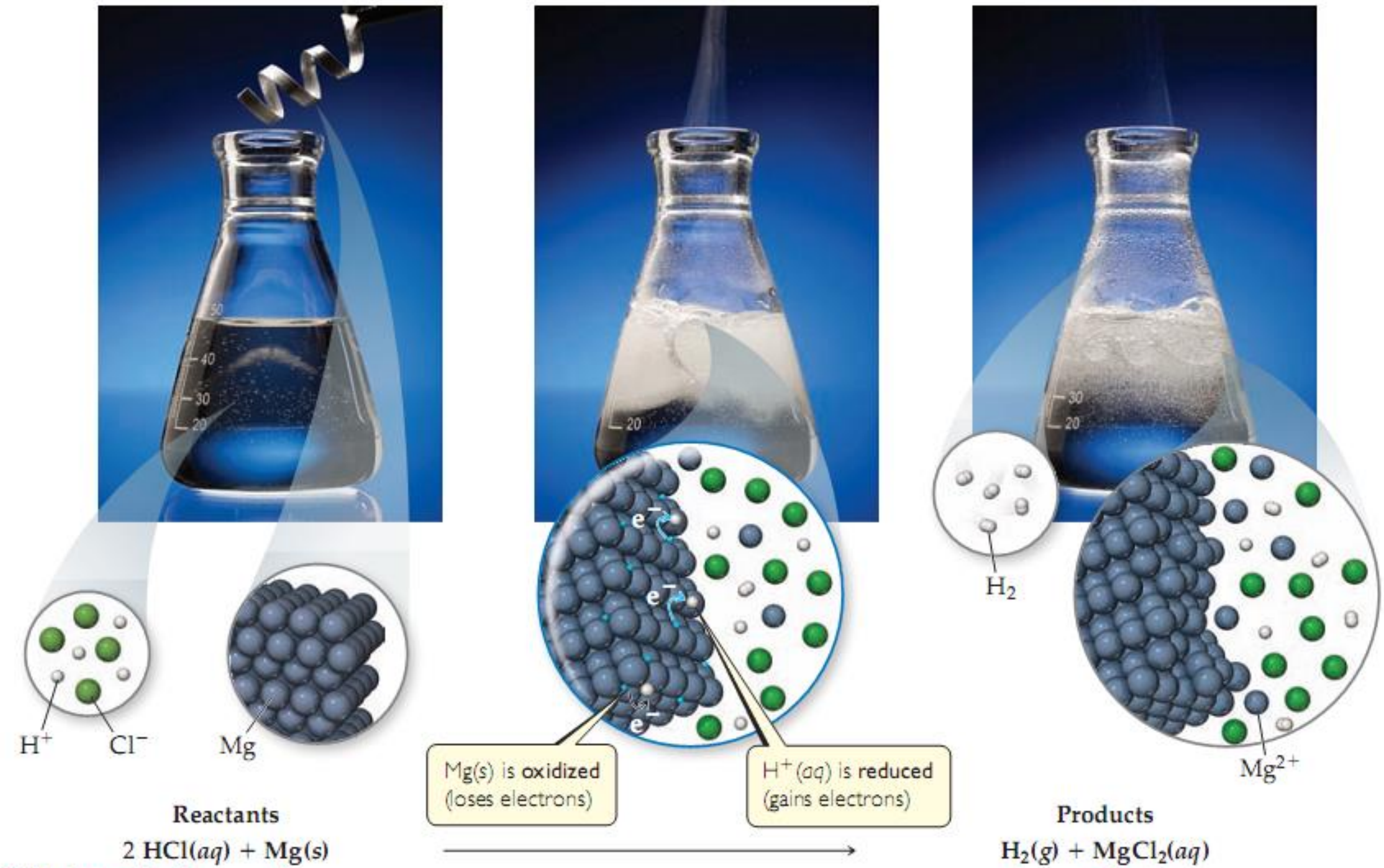
(b) $Na\mathbf{H}$,

(c) $\mathbf{Cr}_2O_7^{2-}$,

(d) $\mathbf{Sn}Br_4$,

(e) $Ba\mathbf{O}_2$?

How many moles of hydrogen gas would be produced for every mole of magnesium added into the HCl solution?



Oxidation of Metals by Acids and Salts

- Few can be balanced by inspection.
- Systematic approach required.
- The Half-Reaction (Ion-Electron) Method

$$\mathrm{A} + \mathrm{BX} \longrightarrow \mathrm{AX} + \mathrm{B}$$

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

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

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

Oxidation number $0 \quad +1 \ -1 \quad +2 \ -1 \quad 0$

$$\mathrm{Mg}(s) + 2\,\mathrm{H^+}(aq) \longrightarrow \mathrm{Mg^{2+}}(aq) + \mathrm{H_2}(g)$$

Balancing in Acid

- Write the equations for the half-reactions.
  - Balance all atoms except H and O.
  - Balance oxygen using H_2O.
  - Balance hydrogen using H^+.
  - Balance charge using e^-.
- Equalize the number of electrons.
- Add the half reactions.
- Check the balance.

EXAMPLE

Balancing the Equation for a Redox Reaction in Acidic Solution. The reaction described below is used to determine the sulfite ion concentration present in wastewater from a papermaking plant. Write the balanced equation for this reaction in acidic solution.

$$SO_3^{2-}(aq) + MnO_4^{-}(aq) \rightarrow SO_4^{2-}(aq) + Mn^{2+}(aq)$$

EXAMPLE

Determine the oxidation states:

$$\overset{4+}{SO_3^{2-}}(aq) + \overset{7+}{MnO_4^-}(aq) \rightarrow \overset{6+}{SO_4^{2-}}(aq) + \overset{2+}{Mn^{2+}}(aq)$$

Write the half-reactions:

$$SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + 2\, e^-(aq)$$

$$5\, e^-(aq) + MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$$

Balance atoms other than H and O:

Already balanced for elements.

EXAMPLE

Balance O by adding H_2O:

$$H_2O(l) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + 2\,e^-(aq)$$

$$5\,e^-(aq) + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4\,H_2O(l)$$

Balance hydrogen by adding H^+:

$$H_2O(l) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + 2\,e^-(aq) + 2\,H^+(aq)$$

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

Check that the charges are balanced: Add e^- if necessary.

EXAMPLE

Multiply the half-reactions to balance all e^-:

$$\cancel{5\,H_2O(l)} + 5\,SO_3^{2-}(aq) \rightarrow 5\,SO_4^{2-}(aq) + \cancel{10\,e^-(aq)} + \cancel{10\,H^+(aq)}$$

$$\underset{6}{\cancel{16}}\,H^+(aq) + \cancel{10\,e^-(aq)} + 2\,MnO_4^-(aq) \rightarrow 2\,Mn^{2+}(aq) + \underset{3}{\cancel{8}}\,H_2O(l)$$

Add both equations and simplify:

$$5\,SO_3^{2-}(aq) + 2\,MnO_4^-(aq) + 6H^+(aq) \rightarrow$$

$$5\,SO_4^{2-}(aq) + 2\,Mn^{2+}(aq) + 3\,H_2O(l)$$

Check the balance!

PRACTICE

1. Which of the following statements is true about the reaction between zinc and copper sulfate?

(a) Zinc is oxidized, and copper ion is reduced.

(b) Zinc is reduced, and copper ion is oxidized.

(c) All reactants and products are soluble strong electrolytes.

(d) The oxidation state of copper in copper sulfate is 0.

(e) More than one of the previous choices are true.

2. For the reaction between magnesium and cobalt(II) sulfate

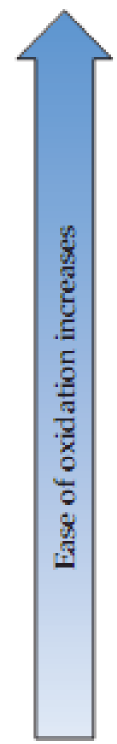
(a) Write the balanced molecular and net ionic equations .

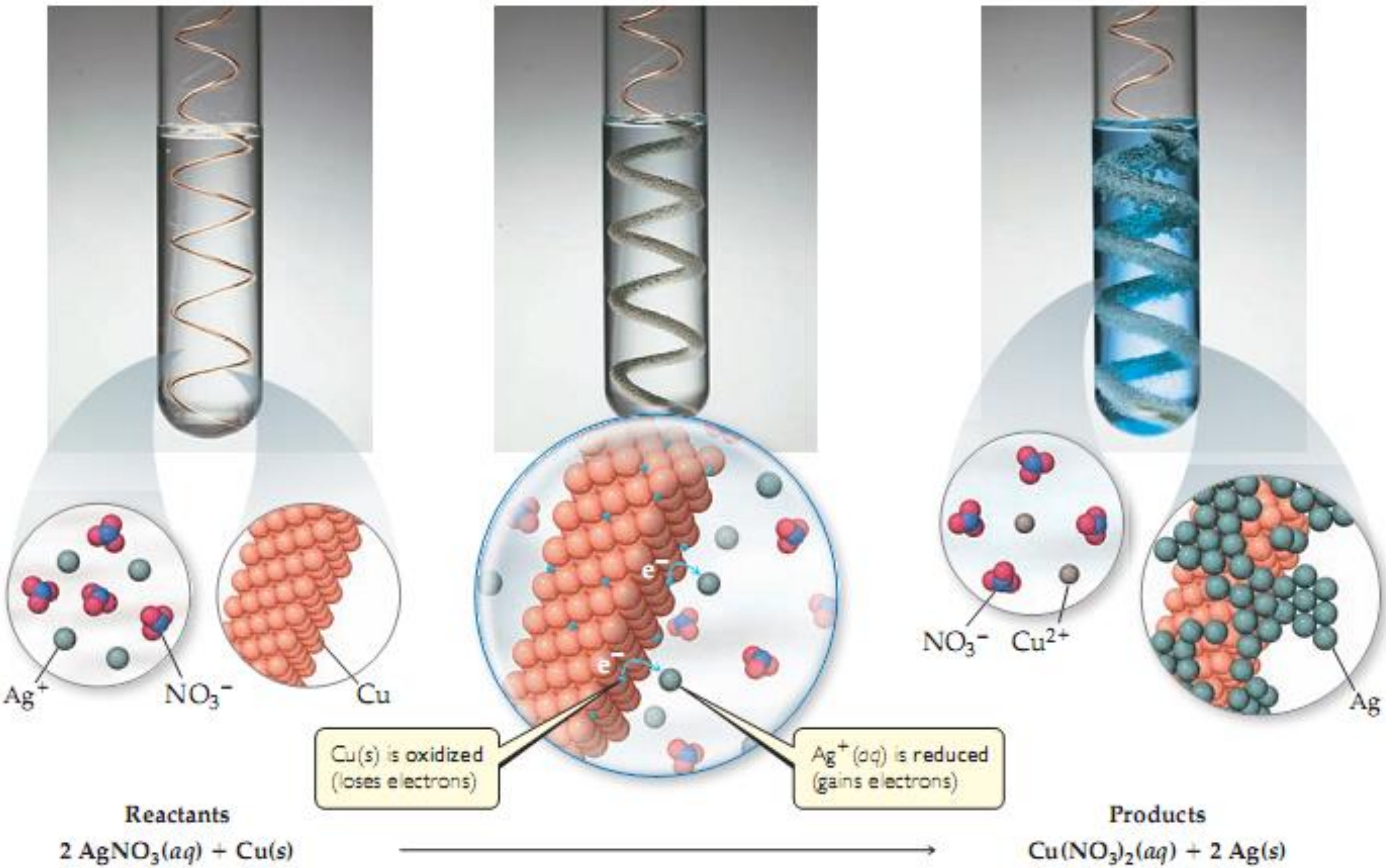
(b) What is oxidized and what is reduced in the reaction?

The Activity Series

Any metal on the list can be oxidized by the ions of elements below it

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	$H_2(g) \longrightarrow 2H^{+}(aq) + 2e^{-}$
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	$Au(s) \longrightarrow Au^{3+}(aq) + 3e^{-}$




Ag$^+$
NO$_3^-$
Cu
e$^-$
e$^-$
Cu(s) is oxidized
(loses electrons)
Ag$^+$(aq) is reduced
(gains electrons)
NO$_3^-$
Cu^{2+}
Ag
Reactants
2 AgNO$_3$(aq) + Cu(s)
Products
Cu(NO$_3$)$_2$(aq) + 2 Ag(s)

Give It Some Thought

Does a reaction occur when an aqueous solution of $NiCl_2$ (aq)

(a) is added to a test tube containing strips of metallic zinc,

(b) is added to a test tube containing $Zn(NO_3)_2$ (aq)?

EXAMPLE

Will an aqueous solution of iron(II) chloride oxidize magnesium metal? If so, write the balanced molecular and net ionic equations for the reaction.

Because Mg is above Fe in the table, the reaction occurs

$$Mg(s) + FeCl_2(aq) \longrightarrow MgCl_2(aq) + Fe(s)$$

Both $FeCl_2$ and $MgCl_2$ are soluble strong electrolytes and can be written in ionic form

$$Mg(s) + Fe^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Fe(s)$$

PRACTICE

1. Which of these metals is the easiest to oxidize?

(a) gold,

(b) lithium,

(c) iron,

(d) sodium,

(e) aluminum.

2. Which of the following metals will be oxidized by $Pb(NO_3)_2$?

(a) Zn,

(b) Cu,

(c) Fe

Balancing in Basic Solution

- OH^- appears instead of H^+.
- Treat the equation as if it were in acid.
  - Then add OH^- to each side to neutralize H^+.
  - Remove H_2O appearing on both sides of equation.
- Check the balance.

Disproportionation Reactions

- The same substance is both oxidized and reduced.
- Some have practical significance

  - Hydrogen peroxide

$$2\ H_2O_2(aq) \rightarrow H_2O(l) + O_2(g)$$

  - Sodium thiosulphate

$$2\ S_2O_3(aq) + H^+(aq) \rightarrow S(s) + SO_2(g) + H_2O(l)$$

Oxidizing and Reducing Agents.

- An oxidizing agent (oxidant):
  - Contains an element whose oxidation state *decreases* in a redox reaction

- A reducing agent (reductant):
  - Contains an element whose oxidation state *increases* in a redox reaction.

Oxidation States of Nitrogen

Compound or ion	Oxidation state
NO_3^-	+5
N_2O_4	+4
NO_2^-	+3
NO	+2
N_2O	+1
N_2	0
NH_2OH	−1
N_2H_4	−2
NH_3	−3

This species cannot be oxidized further

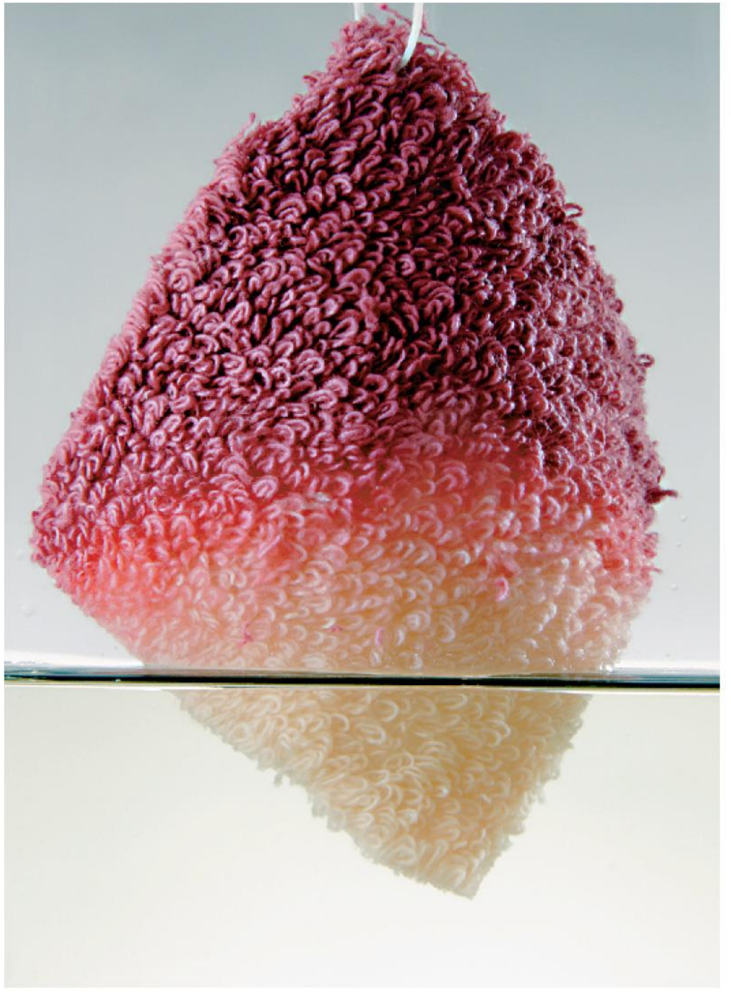
This species cannot be reduced further

Oxidation half-reaction (reducing agent)

Reduction half-reaction (oxidizing agent)

EXAMPLE

Identifying Oxidizing and Reducing Agents. Hydrogen peroxide, H_2O_2, is a versatile chemical. Its uses include bleaching wood pulp and fabrics and substituting for chlorine in water purification. One reason for its versatility is that it can be either an oxidizing or a reducing agent. For the following reactions, identify whether hydrogen peroxide is an oxidizing or reducing agent.



EXAMPLE

$$\overset{2+}{}\qquad\qquad\qquad\overset{3+}{}$$

$$H_2O_2(aq) + 2\ \overset{2+}{Fe^{2+}}(aq) + 2\ H^+ \rightarrow 2\ H_2O(l) + 2\ \overset{3+}{Fe^{3+}}(aq)$$

Iron is oxidized and
peroxide is reduced.

$$5\ H_2O_2(aq) + 2\ \overset{7+}{Mn}O_4^-(aq) + 6\ H^+ \rightarrow$$

$$8\ H_2O(l) + 2\ \overset{2+}{Mn^{2+}}(aq) + 5\ O_2(g)$$

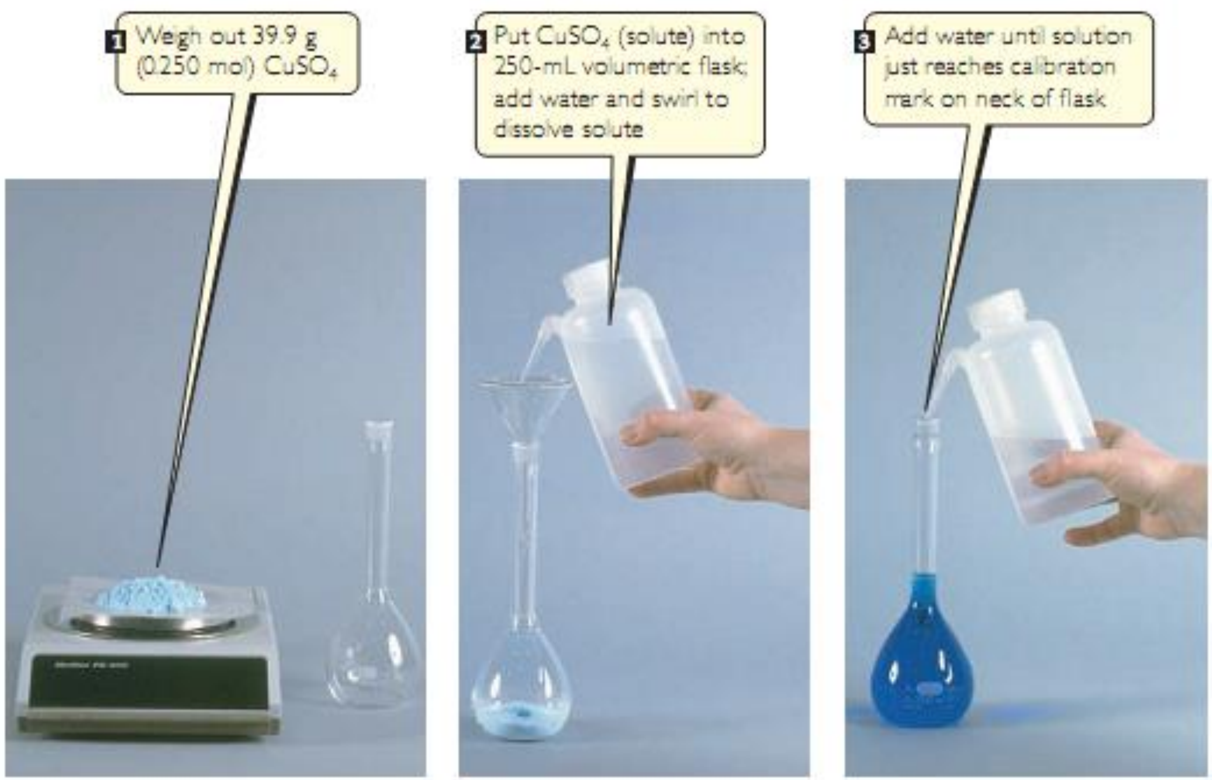
Manganese is reduced and
peroxide is oxidized.

4-5 Concentrations of Solutions.

- Molarity

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

Example: Preparing 0.250 L of a 1.00 M solution of $CuSO_4$.



EXAMPLE

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

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

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

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

PRACTICE

1. What is the molarity of a solution that is made by dissolving 3.68 g of sucrose ($C_{12}H_{22}O_{11}$) in sufficient water to form 275.0 mL of solution?

(a) 13.4 M,

(b) 7.43 x 10^{-2} M,

(c) 3.91 x 10^{-2} M

(d) 7.43 x 10^{-5} M

(e) 3.91 x 10^{-5} M.

2. Calculate the molarity of a solution made by dissolving 5.00 g of glucose ($C_6H_{12}O_6$) in sufficient water to form exactly 100 mL of solution.

Interconverting Molarity and Volume

If we know the molarity of an HNO_3 solution to be 0.200 M, we can calculate the number of moles of solute in a given volume, say 2.0 L.

$$\text{Moles } HNO_3 = (2.0 \text{ L} \cancel{\text{ soln}})\left(\frac{0.200 \text{ mol } HNO_3}{1 \text{ L} \cancel{\text{ soln}}}\right) = 0.40 \text{ mol } HNO_3$$

$$\text{Liters soln} = (2.0 \cancel{\text{ mol } HNO_3})\left[\frac{1 \text{ L soln}}{0.30 \cancel{\text{ mol } HNO_3}}\right] = 6.7 \text{ L soln}$$

$$\text{Liters} = \text{moles} \times 1/M = \cancel{\text{moles}} \times \text{liters}/\cancel{\text{mole}}.$$

Interconverting Percent Concentration and Molarity

A typical American beer contains 5.0% ethanol (CH_3CH_2OH) by volume in water (along with other components). The density of ethanol is 0.789 g/mL. What is the molarity of ethanol?

We would first consider 1.00 L of beer

5% = 5/100 = 0.05 L

$$\text{Moles ethanol} = (0.050\ \cancel{\text{L}})\left(\frac{1000\ \cancel{\text{mL}}}{\cancel{\text{L}}}\right)\left(\frac{0.789\ \cancel{\text{g}}}{\cancel{\text{mL}}}\right)\left(\frac{1\ \text{mol}}{46.0\ \cancel{\text{g}}}\right) = 0.858\ \text{mol}$$

There are 0.858 moles of ethanol in 1.00 L of beer, the concentration of ethanol in beer is **0.86 M**

Dilution

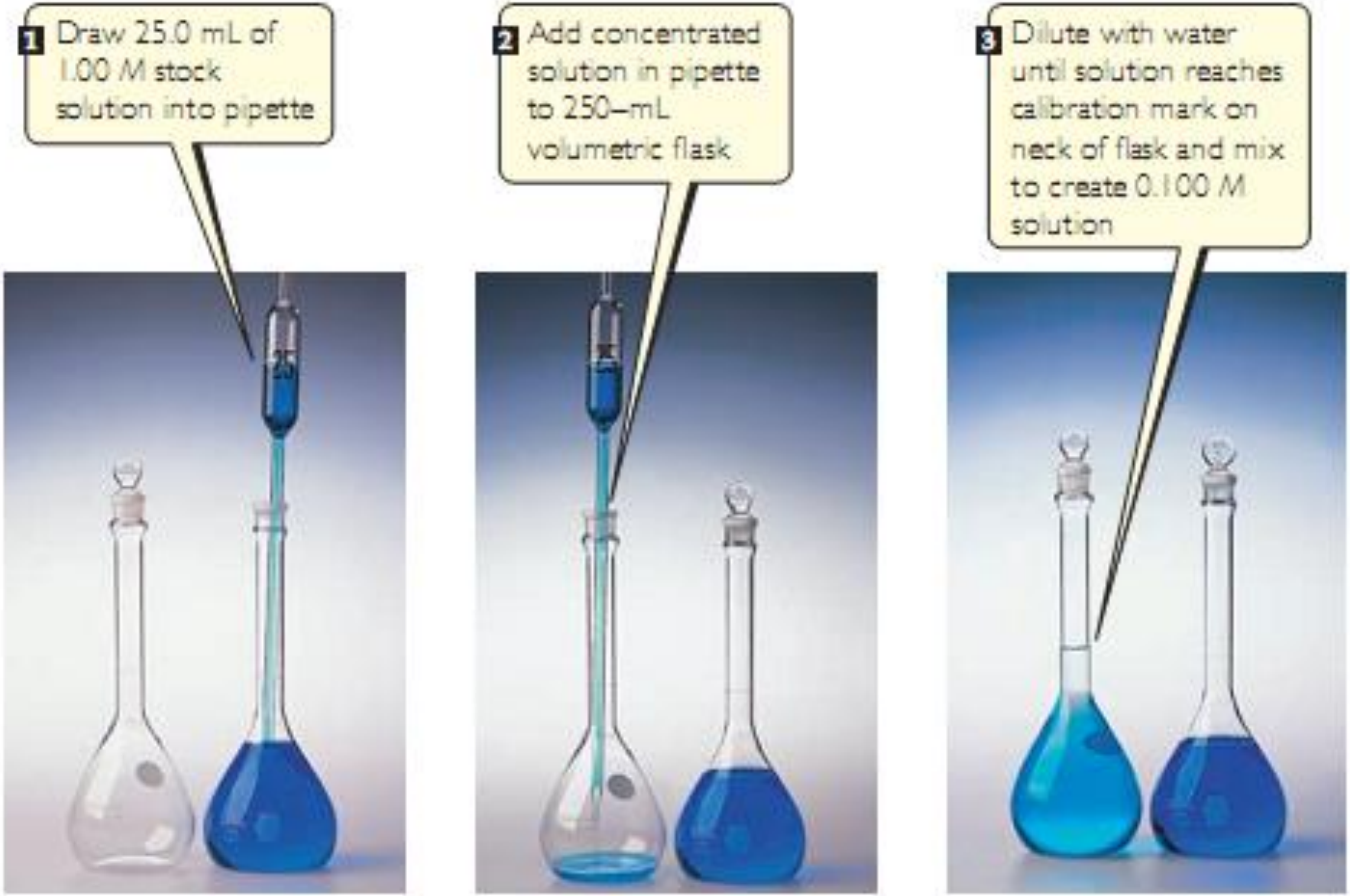
How we can prepare a 0.100 M $CuSO_4$ solution from 1.00 M $CuSO_4$

$$\text{Moles solute before dilution} = \text{moles solute after dilution}$$

$$\text{Moles } CuSO_4 \text{ in dilute soln} = (0.2500\ \cancel{\text{L soln}})\left(\frac{0.100\ \text{mol } CuSO_4}{\cancel{\text{L soln}}}\right)$$

$$= 0.0250\ \text{mol } CuSO_4$$

$$\text{Liters of conc soln} = (0.0250\ \cancel{\text{mol } CuSO_4})\left(\frac{1\ \text{L soln}}{1.00\ \cancel{\text{mol } CuSO_4}}\right) = 0.0250\ \text{L}$$


1 Draw 25.0 mL of 1.00 M stock solution into pipette
2 Add concentrated solution in pipette to 250–mL volumetric flask
3 Dilute with water until solution reaches calibration mark on neck of flask and mix to create 0.100 M solution

EXAMPLE

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

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

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

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

Converting liters to milliliters gives 15 mL

PRACTICE

1. What volume of a 1.00 M stock solution of glucose must be used to make 500.0 mL of a $1.75 * 10^{-2}$ M glucose solution in water?

(a) 1.75 mL, (b) 8.75 mL, (c) 48.6 mL, (d) 57.1 mL, (e) 28,570 mL.

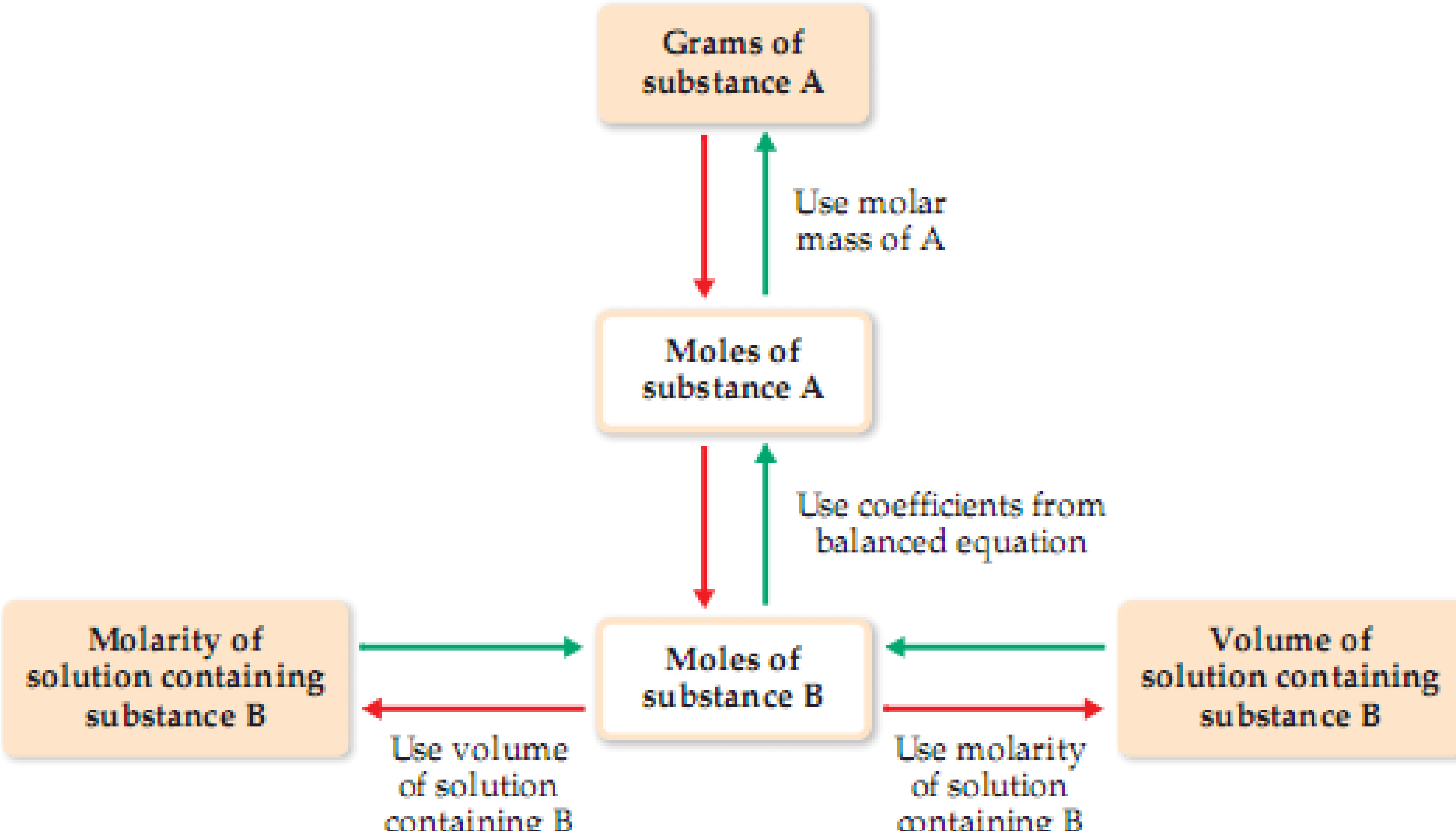
2.

(a) What volume of 2.50 M lead(II) nitrate solution contains 0.0500 mol of Pb^{2+}?

(b) How many milliliters of 5.0 M $K_2Cr_2O_7$ solution must be diluted to prepare 250 mL of 0.10 M solution?

(c) If 10.0 mL of a 10.0 M stock solution of NaOH is diluted to 250 mL, what is the concentration of the resulting stock solution?

4-6 Solution Stoichiometry and Chemical Analysis



EXAMPLE

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

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

$$V_{HNO_3} \times M_{HNO_3} \Rightarrow \text{mol } HNO_3 \Rightarrow \text{mol } Ca(OH)_2 \Rightarrow \text{g } Ca(OH)_2$$

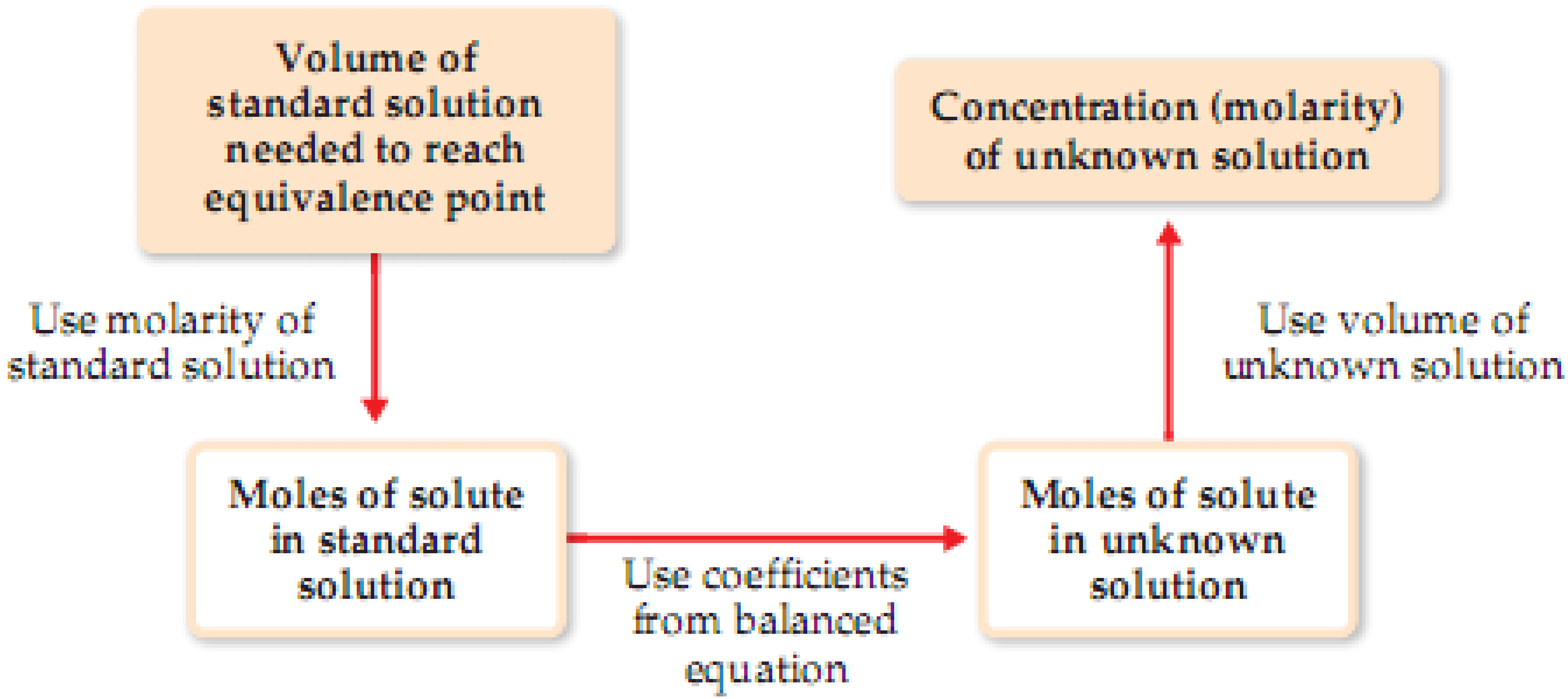
$$\begin{aligned}\text{Moles } HNO_3 &= V_{HNO_3} \times M_{HNO_3} = (0.0250\ \cancel{L})\left(\frac{0.100 \text{ mol } HNO_3}{\cancel{L}}\right)\\ &= 2.50 \times 10^{-3} \text{ mol } HNO_3\end{aligned}$$

$$2 \text{ mol } HNO_3 \simeq \text{ mol } Ca(OH)_2.$$

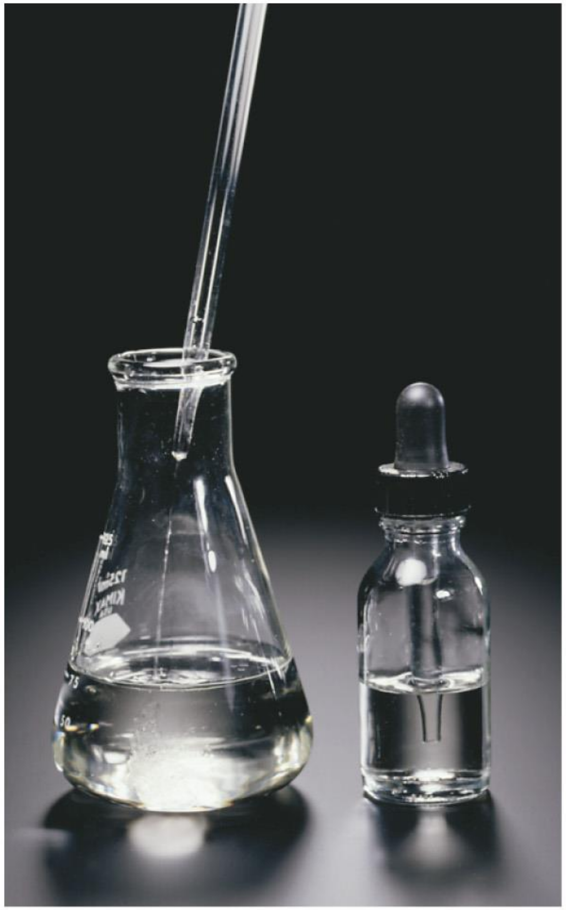
$$\begin{aligned}\text{Grams } Ca(OH)_2 &= (2.50 \times 10^{-3}\ \cancel{\text{mol } HNO_3})\\ &\times \left(\frac{1\ \cancel{\text{mol } Ca(OH)_2}}{2\ \cancel{\text{mol } HNO_3}}\right)\left(\frac{74.1 \text{ g } Ca(OH)_2}{1\ \cancel{\text{mol } Ca(OH)_2}}\right)\\ &= 0.0926 \text{ g } Ca(OH)_2\end{aligned}$$

Titrations.

- Titration
  - Carefully controlled addition of one solution to another.
- Equivalence Point
  - Both reactants have reacted completely.
- Indicators
  - Substances which change colour near an equivalence point.

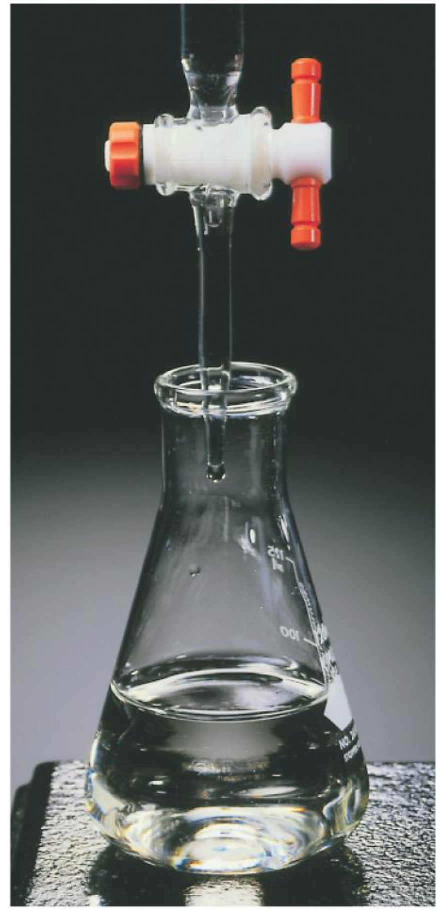

Volume of standard solution needed to reach equivalence point
Use molarity of standard solution
Moles of solute in standard solution
Use coefficients from balanced equation
Moles of solute in unknown solution
Use volume of unknown solution
Concentration (molarity) of unknown solution

Indicators

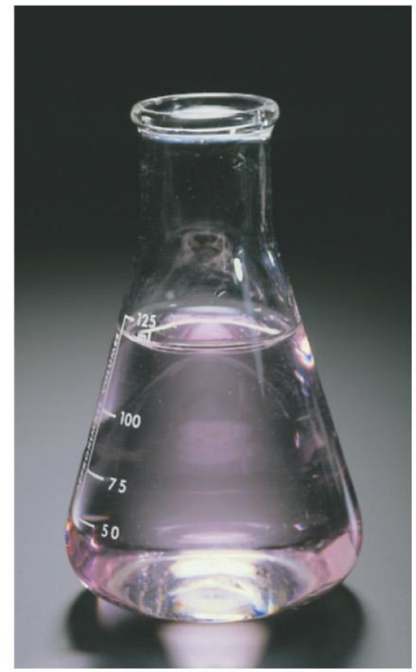


5.0 mL CH_3CO_2H

A few drops phenolpthalein



Add 0.1000 M NaOH



The “endpoint”
(close to the equivalence point)

EXAMPLE

Standardizing a Solution for Use in Redox Titrations. A piece of iron wire weighing 0.1568 g is converted to $Fe^{2+}(aq)$ and requires 26.42 mL of a $KMnO_4(aq)$ solution for its titration. What is the molarity of the $KMnO_4(aq)$?

$$5\ Fe^{2+}(aq) + MnO_4^-(aq) + 8\ H^+(aq) \rightarrow$$

$$5\ Fe^{3+}(aq) + Mn^{2+}(aq) + 4\ H_2O(l)$$

EXAMPLE

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

Determine $KMnO_4$ consumed in the reaction:

$$n_{MnO_4^-} = 0.1568\,\cancel{g}\,Fe \times \frac{1\,\cancel{mol}\,Fe}{55.847\,\cancel{g}\,Fe} \times \frac{1\,\cancel{mol}\,Fe^{2+}}{1\,\cancel{mol}\,Fe} \times$$

$$\frac{1\,\cancel{mol}\,MnO_4^-}{5\,\cancel{mol}\,Fe^{2+}} \times \frac{1\,mol\,KMnO_4}{1\,\cancel{mol}\,MnO_4^-} = 5.615 \times 10^{-4}\,mol\,KMnO_4$$

Determine the concentration:

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

-

Relative Concentrations in Solution

$$MgCl_2(s) \longrightarrow Mg^{2+}(aq) + 2\ Cl^-(aq)$$

In 0.0050 M $MgCl_2$:

Stoichiometry is important.

$[Mg^{2+}] = 0.0050$ M $[Cl^-] = 0.0100$ M $[MgCl_2] = 0$ M

EXAMPLE

Calculating Ion concentrations in a Solution of a Strong Electrolyte. What are the aluminum and sulfate ion concentrations in 0.0165 M $Al_2(SO_4)_3$?.

Write a Balanced Chemical Equation:

$$Al_2(SO_4)_3\,(s) \rightarrow 2\,Al^{3+}(aq) + 3\,SO_4^{2-}(aq)$$

Identify the Stoichiometric Factors :

$$\frac{2 \text{ mol } Al^{3+}}{1 \text{ mol } Al_2(SO_4)_3} \qquad \frac{3 \text{ mol } SO_4^{2-}}{1 \text{ mol } Al_2(SO_4)_3}$$

EXAMPLE

Aluminum Concentration:

$$[Al] = \frac{0.0165 \text{ mol } Al_2(SO_4)_3}{1 \text{ L}} \times \frac{2 \text{ mol } Al^{3+}}{1 \text{ mol } Al_2(SO_4)_3} = 0.0330 \frac{\text{mol } Al^{3+}}{1 \text{ L}}$$

Sulfate Concentration:

$$[SO_4^{2-}] = \frac{0.0165 \text{ mol } Al_2(SO_4)_3}{1 \text{ L}} \times \frac{3 \text{ mol } SO_4^{2-}}{1 \text{ mol } Al_2(SO_4)_3} = 0.0495 \text{ M } SO_4^{2-}$$

PRACTICE

1. A mysterious white powder is found at a crime scene. A simple chemical analysis concludes that the powder is a mixture of sugar and morphine ($C_{17}H_{19}NO_3$), a weak base similar to ammonia. The crime lab takes 10.00 mg of the mysterious white powder, dissolves it in 100.00 mL water, and titrates it to the equivalence point with 2.84 mL of a standard 0.0100 M HCl solution. What is the percentage of morphine in the white powder?

(a) 8.10%,

(b) 17.3%,

(c) 32.6%,

(d) 49.7%,

(e) 81.0%.

PRACTICE

2. A sample of an iron ore is dissolved in acid, and the iron is converted to Fe^{2+}. The sample is then titrated with 47.20 mL of 0.02240 M MnO_4^- solution. The oxidation-reduction reaction that occurs during titration is

$MnO4^-(aq) + 5\ Fe^{2+}(aq) + 8\ H^+(aq) \rightarrow Mn^{2+}(aq) + 5\ Fe^{3+}(aq) + 4\ H_2O(l)$

(a) How many moles of $MnO4^-$ were added to the solution?

(b) How many moles of Fe^{2+} were in the sample?

(c) How many grams of iron were in the sample?

(d) If the sample had a mass of 0.8890 g, what is the percentage of iron in the sample?

Homeworks

Excercises:

4.21

4.25

4.33

4.43

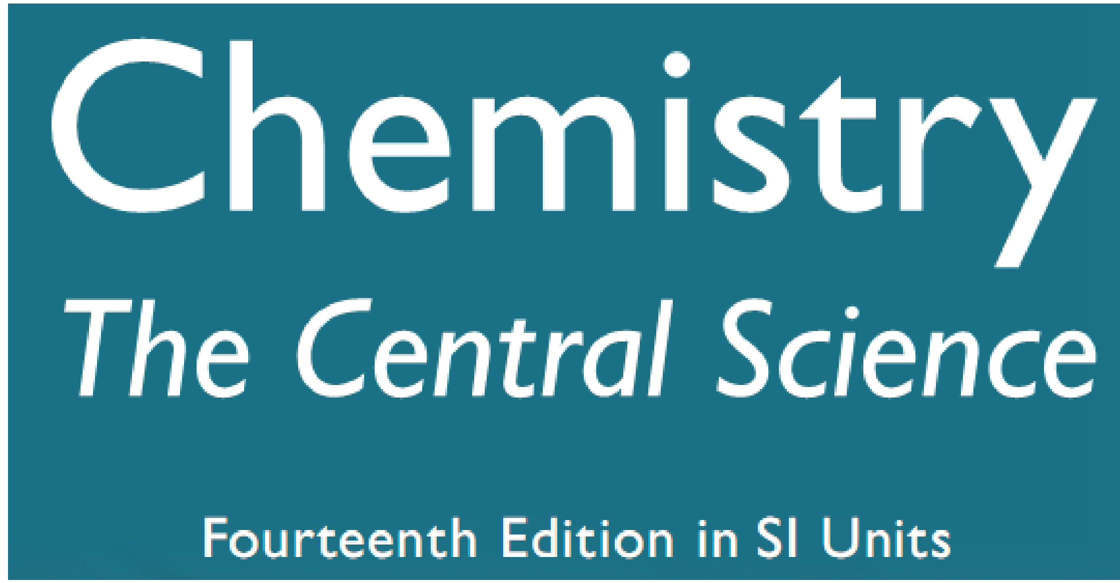
4.51

4.61

4.77

4.83

Due date: April 15th, 2020



You will have a 30 minute quiz next week

What are allowed?

- A pen
- A periodic table
- An A4-sized sheet of hand written notes
- A calculator

No other devices are allowed!