

# Chapter 4

## Reactions in Aqueous Solution

# Topics

- **General properties of aqueous solutions**
- **Precipitation reactions**
- **Acid base reactions**
- **Oxidation reduction reactions**
- **Concentration of solutions**
- **Aqueous reactions and chemical analysis**

# 4.1 General Properties of Aqueous Solutions

A **solution** is a homogenous mixture of 2 or more substances

The solute is(are) the substance(s) present in the smaller amount(s). The substance dissolved

The **solvent** is the substance present in the larger amount. The substance does the dissolving

<u>Solution</u>	<u>Solvent</u>	<u>Solute(s)</u>
Soft drink (l)	H <sub>2</sub> O	Sugar, CO <sub>2</sub>
Air (g)	N <sub>2</sub>	O <sub>2</sub> , Ar, CH <sub>4</sub>
Soft Solder (s)	Pb	Sn

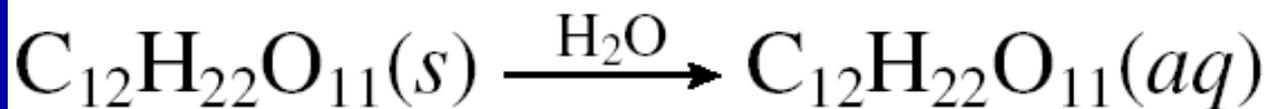
# Electrolytes and Nonelectrolytes

- **Electrolytes**- Substances that conduct an electric current in aqueous solution, or in the **molten state because the solutions contain ions**



- all ionic compounds are electrolytes because their aqueous solutions contain ions (they are also called “**salts**”)
- barium sulfate- will conduct when **molten**, but is **insoluble** in water!

- **Nonelectrolytes**- Substances that do not conduct an electric current when dissolved in water
  - Most of them are **molecular** materials, and their solutions ***do not contain ions***

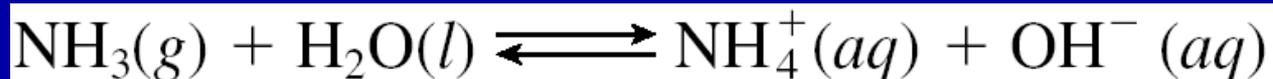


- Not all electrolytes conduct to the same degree
  - there are **strong** electrolytes, and **weak** electrolytes

- ***Dissociation*** - ionic compounds separate into constituent ions when dissolved in solution



- ***Ionization*** - formation of ions by molecular compounds when dissolved



# Strong and weak electrolytes

- Not all electrolytes conduct to the same degree
- Conductivity depends on: degree of dissociation or ionization
- *Strong Electrolyte*: 100% dissociation
  - All water soluble ionic compounds, **strong acids and strong bases**
- *Weak electrolytes*
  - Partially ionized in solution
  - Exist mostly as **the molecular form** in solution
  - **Weak acids and weak bases**

**TABLE 4.1****The Strong Acids**

<b>Acid</b>	<b>Ionization Equation</b>
Hydrochloric acid	$\text{HCl}(aq) \longrightarrow \text{H}^+(aq) + \text{Cl}^-(aq)$
Hydrobromic acid	$\text{HBr}(aq) \longrightarrow \text{H}^+(aq) + \text{Br}^-(aq)$
Hydroiodic acid	$\text{HI}(aq) \longrightarrow \text{H}^+(aq) + \text{I}^-(aq)$
Nitric acid	$\text{HNO}_3(aq) \longrightarrow \text{H}^+(aq) + \text{NO}_3^-(aq)$
Chloric acid	$\text{HClO}_3(aq) \longrightarrow \text{H}^+(aq) + \text{ClO}_3^-(aq)$
Perchloric acid	$\text{HClO}_4(aq) \longrightarrow \text{H}^+(aq) + \text{ClO}_4^-(aq)$
Sulfuric acid*	$\text{H}_2\text{SO}_4(aq) \longrightarrow \text{H}^+(aq) + \text{HSO}_4^-(aq)$
	$\text{HSO}_4^-(aq) \rightleftharpoons \text{H}^+(aq) + \text{SO}_4^{2-}(aq)$

\*Note that although each sulfuric acid molecule has two ionizable hydrogen atoms, it only undergoes the first ionization completely, effectively producing one  $\text{H}^+$  ion and one  $\text{HSO}_4^-$  ion per  $\text{H}_2\text{SO}_4$  molecule. The second ionization happens only to a very small extent.

- **Examples of weak electrolytes**

- **Weak acids**

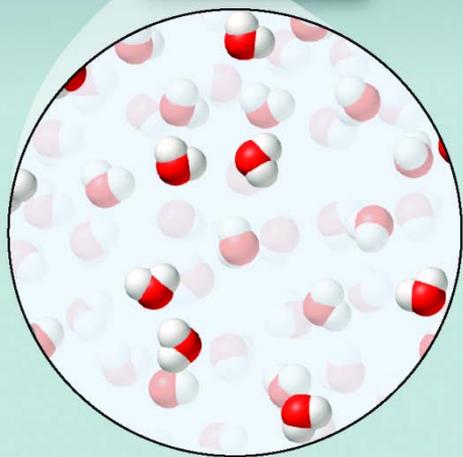


- **Weak bases**

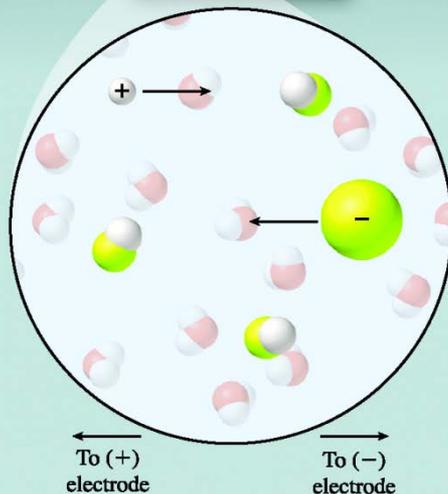


(Note: double arrows indicate a reaction that occurs in both directions - a state of *dynamic equilibrium* exists)

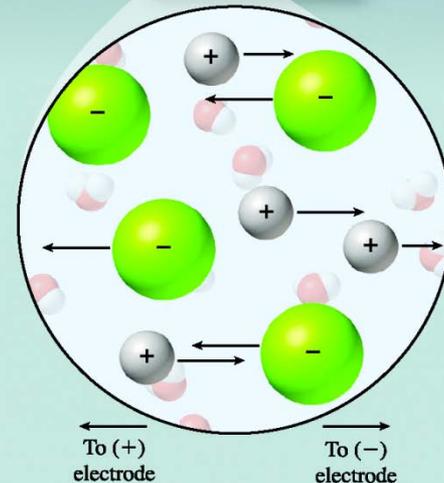
# Method to Distinguish Types of Electrolytes



**nonelectrolyte**



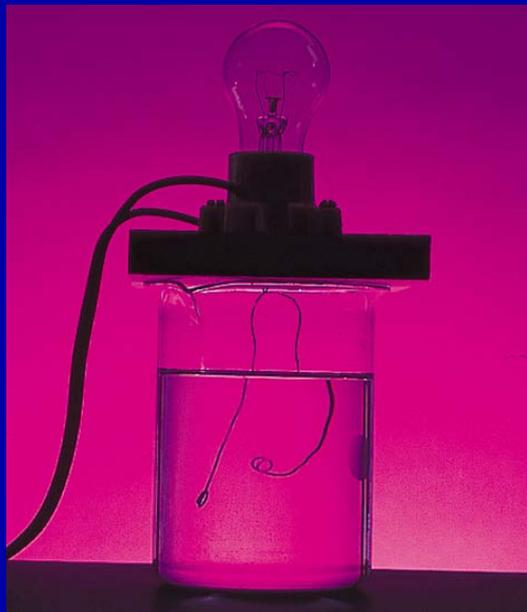
**weak electrolyte**



**strong electrolyte**

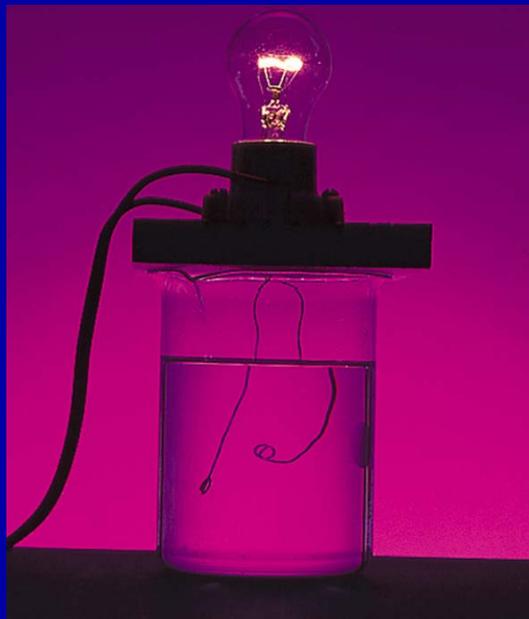
**Ethanol and  
table sugar**

**Nonelectrolyte**



**Acetic acid  
ammonia**

**Weak electrolyte**



**Sodium chloride  
Hydrochloric acid**

**Strong electrolyte**



**Classify the following as  
nonelectrolyte, weak electrolyte or  
strong electrolyte**

**– NaOH**

**strong electrolyte**

**– CH<sub>3</sub>OH**

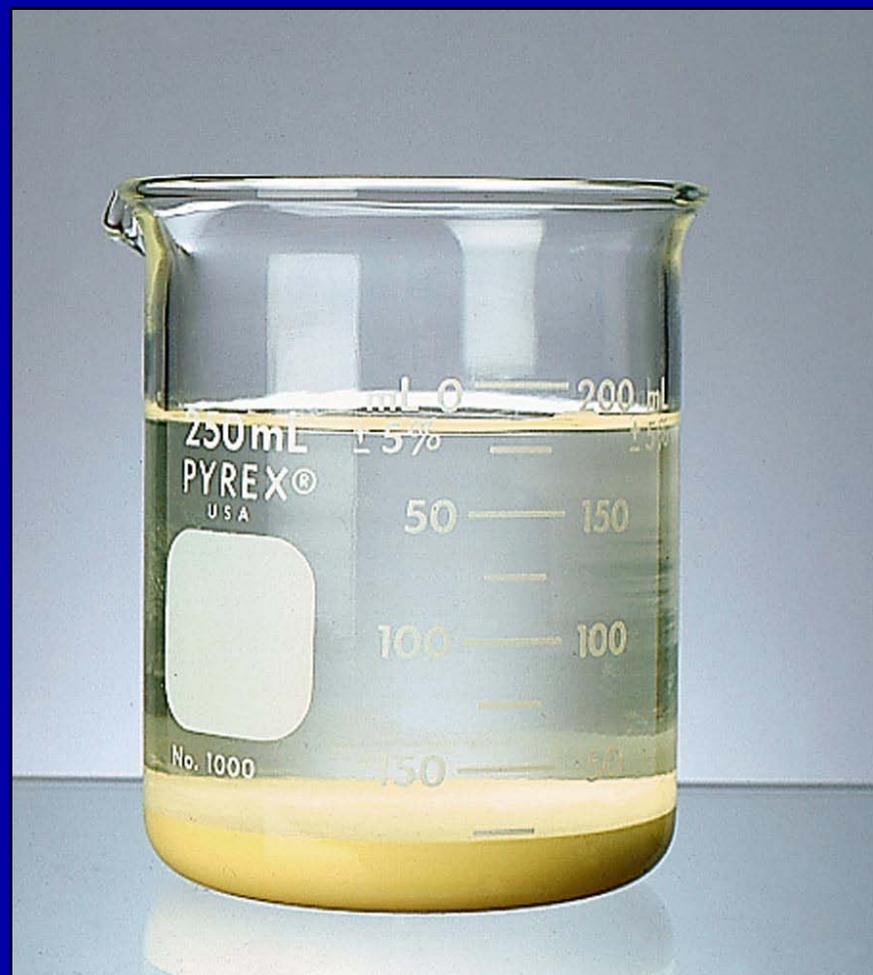
**nonelectrolyte**

**– H<sub>2</sub>CO<sub>3</sub>**

**weak electrolyte**

## 4.2 Precipitation Reactions

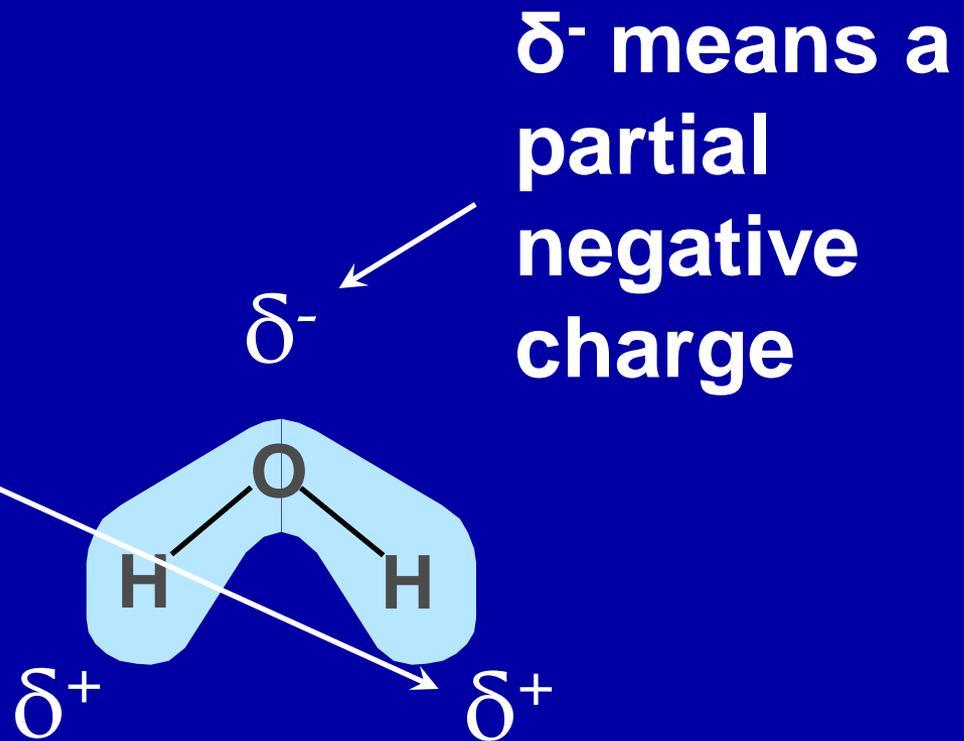
- When aqueous solutions of ionic compounds are mixed together **a solid forms (occurs when product is insoluble)**
- A solid that forms from mixed solutions is called precipitate
- **If the substance is not part of the solution, it is a precipitate**
- ***Solubility*** is the maximum amount of a solid that can dissolve in a given amount of solvent at a specified temperature
- Prediction based on solubility rules



# Water, the common solvent

## The Water Molecule, Polarity

Thus, water has a partial negative end (oxygen) and a partial positive end (Hydrogen) and it is called “polar” because of **the unequal charge distribution**



# Dissolving ionic salts in water and Hydration

- Ions have charges and attract the opposite charges on the water molecules.
- The process of breaking the ions of salts apart is called **hydration**

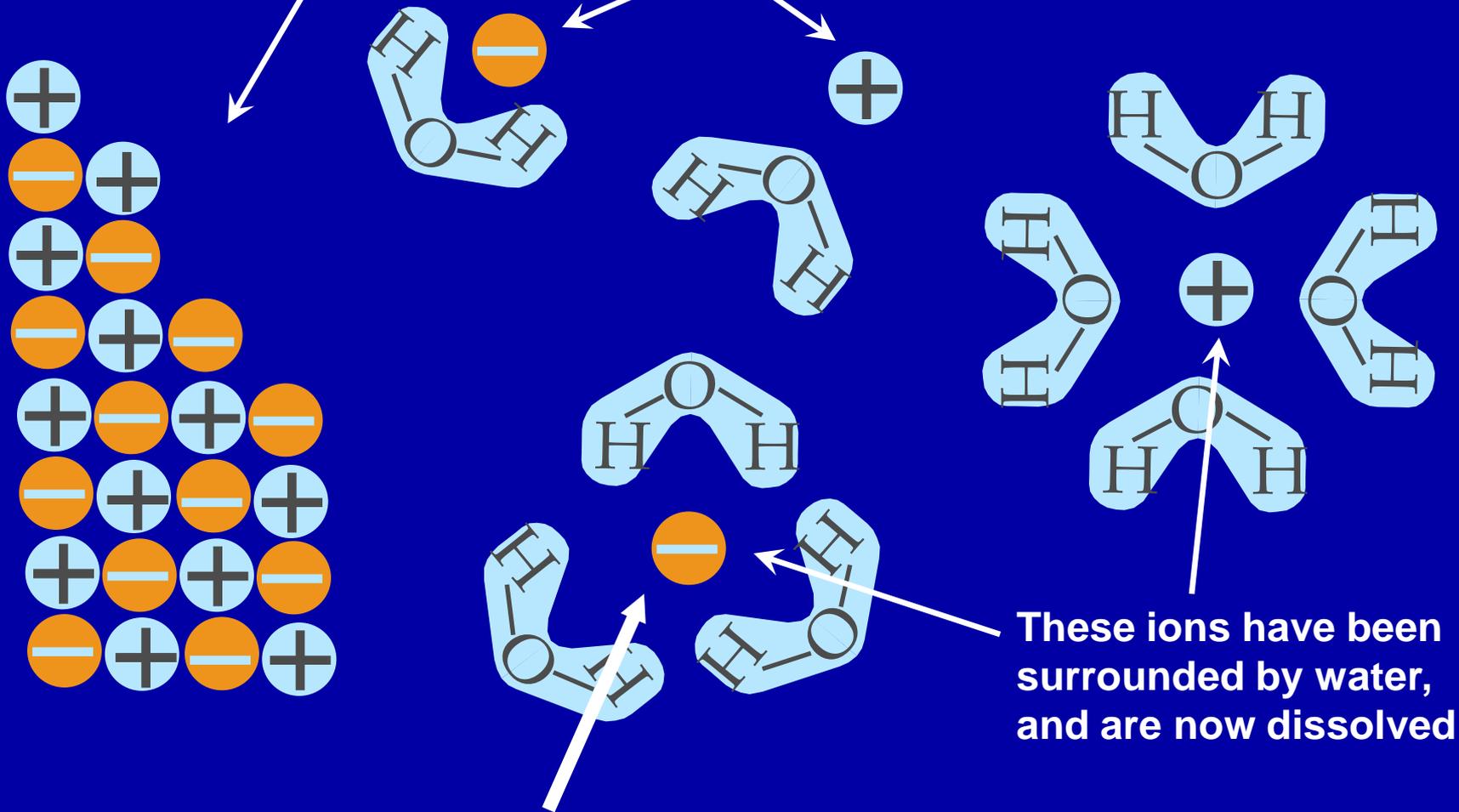


Designates hydration of ions

- **Hydration:** process by which water molecules remove and surround individual ions from the solid.

# How ionic solids dissolve in water

These ions have been pulled away from the main crystal structure by water's polarity.

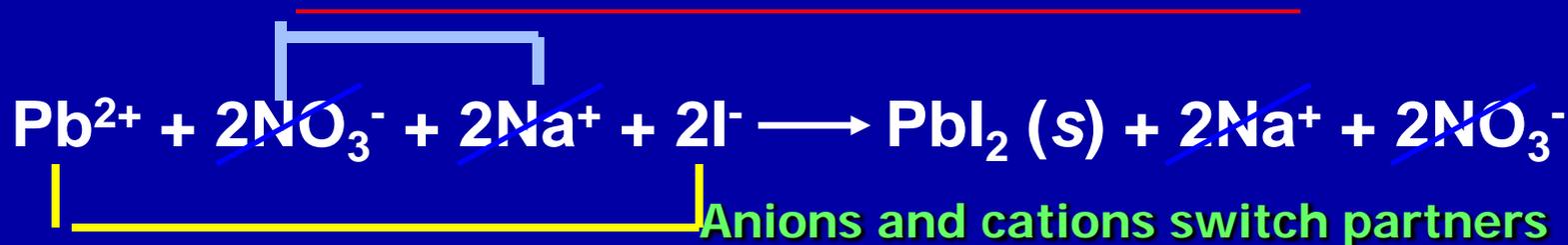
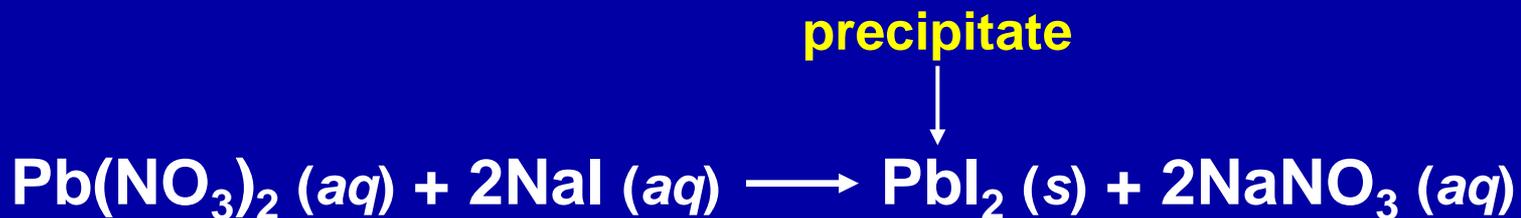


These ions have been surrounded by water, and are now dissolved

**Hydration**

# Precipitation Reactions

(Double replacement or metathesis reaction)



**PbI<sub>2</sub>**

Na<sup>+</sup> and NO<sub>3</sub><sup>-</sup> are Spectator ions

**TABLE 4.2****Solubility Guidelines: Soluble Compounds****Water-Soluble Compounds**

Compounds containing an alkali metal cation ( $\text{Li}^+$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$ ) or the ammonium ion ( $\text{NH}_4^+$ )

••• Compounds containing the nitrate ion ( $\text{NO}_3^-$ ), acetate ion ( $\text{C}_2\text{H}_3\text{O}_2^-$ ), or chlorate ion ( $\text{ClO}_3^-$ )

Compounds containing the chloride ion ( $\text{Cl}^-$ ), bromide ion ( $\text{Br}^-$ ), or iodide ion ( $\text{I}^-$ )

Compounds containing the sulfate ion ( $\text{SO}_4^{2-}$ )

**Insoluble Exceptions**

Compounds containing  $\text{Ag}^+$ ,  $\text{Hg}_2^{2+}$ , or  $\text{Pb}^{2+}$

Compounds containing  $\text{Ag}^+$ ,  $\text{Hg}_2^{2+}$ ,  $\text{Pb}^{2+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Sr}^{2+}$ , or  $\text{Ba}^{2+}$

**TABLE 4.3****Solubility Guidelines: Insoluble Compounds****Water-Insoluble Compounds**

Compounds containing the carbonate ion ( $\text{CO}_3^{2-}$ ), phosphate ion ( $\text{PO}_4^{3-}$ ), chromate ion ( $\text{CrO}_4^{2-}$ ), or sulfide ion ( $\text{S}^{2-}$ )

Compounds containing the hydroxide ion ( $\text{OH}^-$ )

**Soluble Exceptions**

Compounds containing  $\text{Li}^+$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$ , or  $\text{NH}_4^+$

Compounds containing  $\text{Li}^+$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$ , or  $\text{Ba}^{2+}$

# **Classify the following as soluble or insoluble in water**



**soluble**



**insoluble**

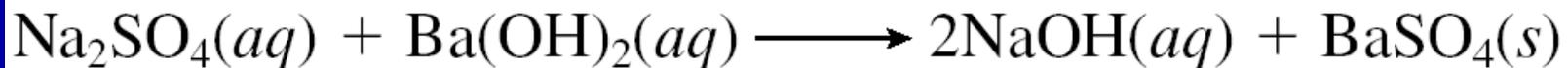


**insoluble**

## Solubility rules for common ionic compounds in water at 25° C

Soluble Compounds	Exceptions
Compounds containing alkali metal ions and $\text{NH}_4^+$ $\text{NO}_3^-$ , $\text{HCO}_3^-$ , $\text{ClO}_3^-$ $\text{Cl}^-$ , $\text{Br}^-$ , $\text{I}^-$ $\text{SO}_4^{2-}$	Halides of $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , $\text{Pb}^{2+}$ Sulfates of $\text{Ag}^+$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$ , $\text{Hg}^{2+}$ , $\text{Pb}^{2+}$
Slightly soluble Compounds	Exceptions
$\text{CO}_3^{2-}$ , $\text{PO}_4^{3-}$ , $\text{CrO}_4^{2-}$ , $\text{S}^{2-}$ $\text{OH}^-$	Compounds containing alkali metal ions and $\text{NH}_4^+$ Compounds containing alkali metal ions and $\text{Ba}^{2+}$ , $\text{Ca}^{2+}$ , $\text{Mg}^{2+}$ are marginally soluble

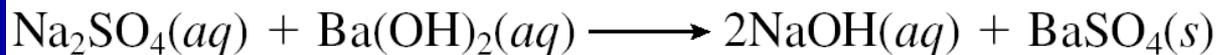
- **Molecular equation:** shows all compounds represented by their chemical formulas



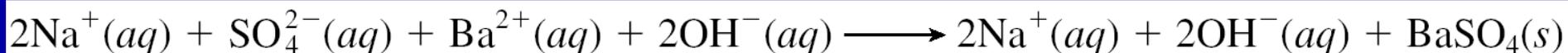
- **Ionic equation:** shows all strong electrolytes as ions and all other substances (non-electrolytes, weak electrolytes, gases) by their chemical formulas



## Molecular equation:



## Ionic equation:

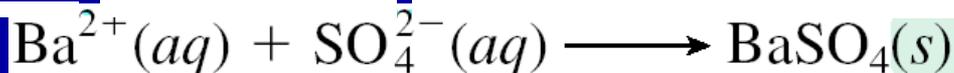


- **Net ionic equation: shows only the reacting species in the chemical equation**

**Eliminates spectator ions –**



## Net ionic equation:



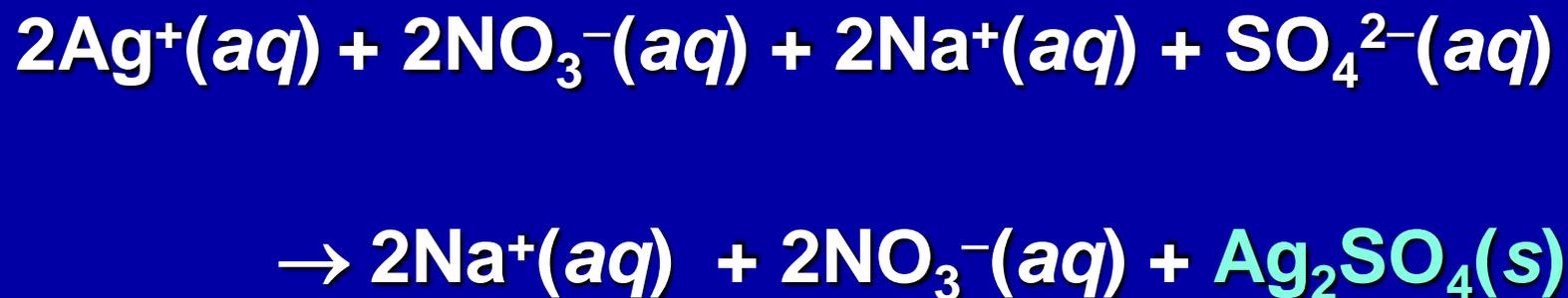
# Steps in writing a net ionic equation

**Aqueous solutions of silver nitrate and sodium sulfate are mixed. Write the net ionic reaction.**

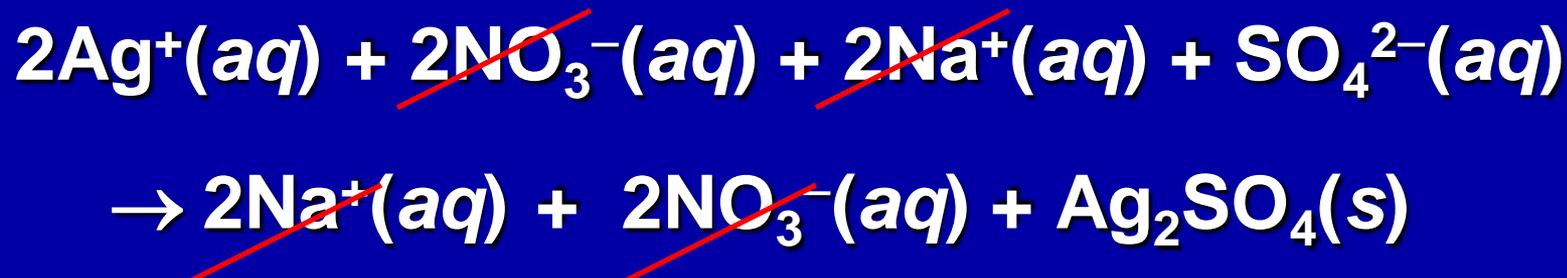
**Step 1:**



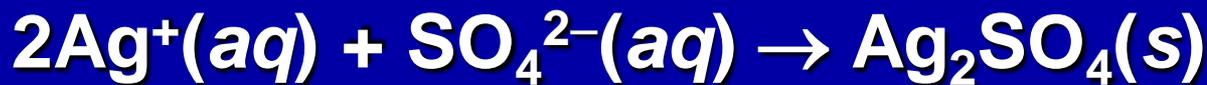
**Step 2: Use solubility table; all nitrates are soluble but silver sulfate is insoluble**



### Step 3: Cancel spectators



### Step 4: Write the net ionic reaction



## 4.3 Acid-Base Reactions

- Termed neutralization reactions.
- Involve an **acid** and a **base**.
- **Double replacement (or metathesis) reaction**
- A molecular compound (**water**) is a common product along with a **salt (ionic compound)**.

**TABLE 4.4****Strong Acids and Strong Bases**

<b>Strong Acids</b>	<b>Strong Bases</b>	<b>Strong Acids</b>	<b>Strong Bases</b>
HCl	LiOH	HClO <sub>3</sub>	CsOH
HBr	NaOH	HClO <sub>4</sub>	Ca(OH) <sub>2</sub>
HI	KOH	H <sub>2</sub> SO <sub>4</sub>	Sr(OH) <sub>2</sub>
HNO <sub>3</sub>	RbOH		Ba(OH) <sub>2</sub>

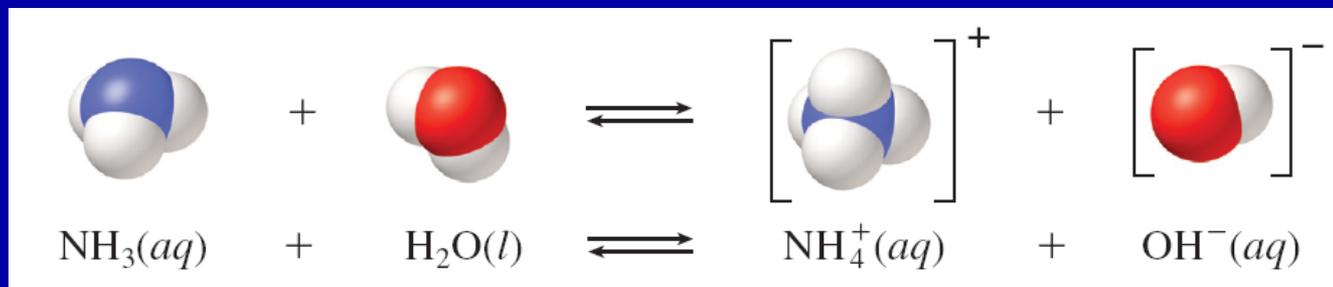
**All the other acids and bases are weak electrolytes (important for net ionic equations).**

## Definitions of acids and bases

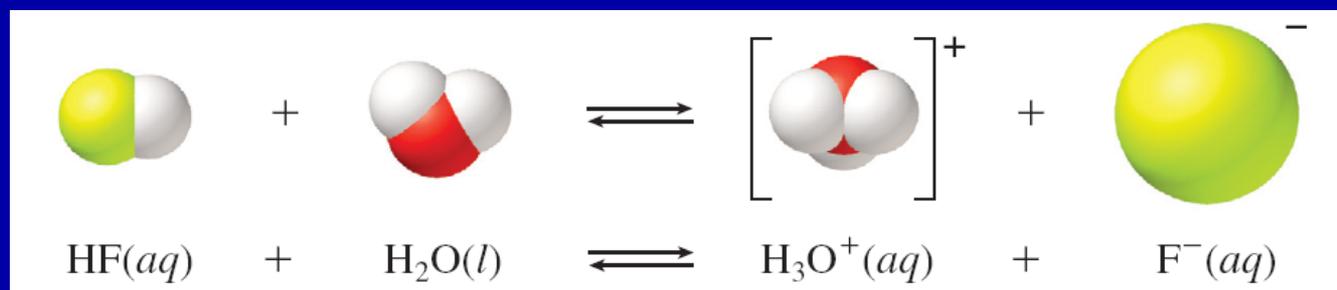
- ***Arrhenius acid*** - produces  $\text{H}^+$  in solution
- ***Arrhenius base*** - produces  $\text{OH}^-$  in solution
- More inclusive definitions:
  - ***Brønsted acid*** - proton donor
  - ***Brønsted base*** - proton acceptor

– Examples of a weak base and weak acid

▪ Ammonia with water:



▪ Hydrofluoric acid with water:



## Types of acids

- **Monoprotic**: one ionizable hydrogen



- **Diprotic**: two ionizable hydrogens



– **Triprotic**: three ionizable hydrogens



– **Polyprotic**: generic term meaning more than one ionizable hydrogen

# Types of bases

– **Monobasic:** One OH<sup>-</sup> group



– **Dibasic:** Two OH<sup>-</sup> groups



# Acid-Base Neutralization

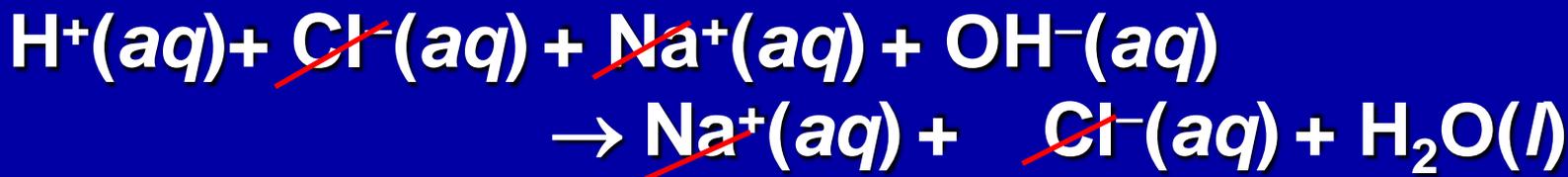
- **Neutralization:** Reaction between an acid and a base



**Molecular equation:** (Strong acid-strong base)



**Ionic equation:**

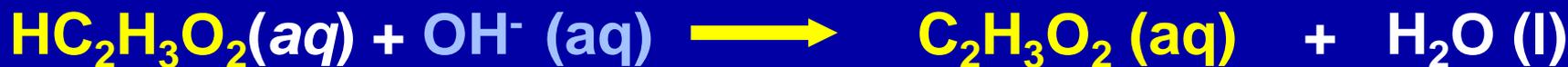
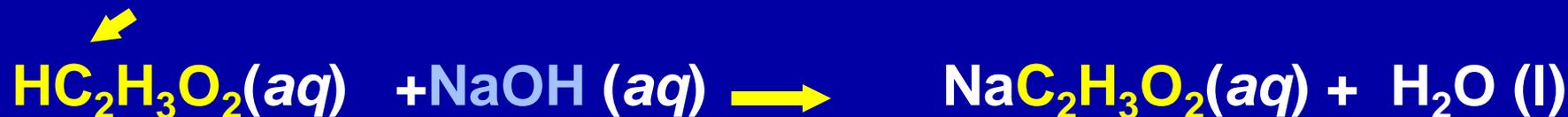


**Net ionic equation:**



# Describing weak acid- strong base reactions

Weak acid



## 4.4 Oxidation-Reduction Reactions

- Often called “**redox**” reactions
- Electrons are transferred between the reactants
  - One substance **is oxidized, loses electrons (electron donor)**
    - **Reducing agent**
  - Another substance is **reduced, gains electrons (electron acceptor)**
    - **Oxidizing agent**
- Oxidation numbers change during the reaction

– **Example**



– **Zinc is losing 2 electrons** and oxidized.

- **Reducing agent**

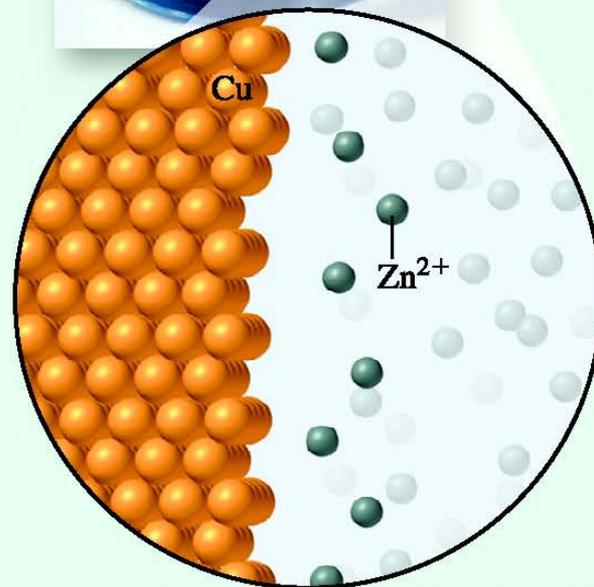
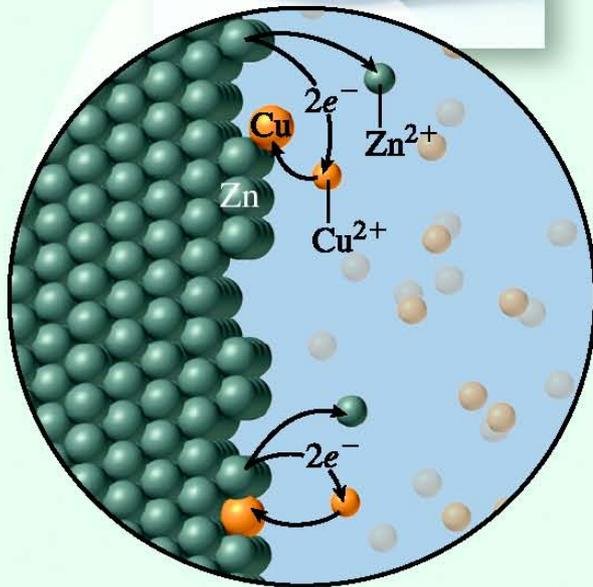
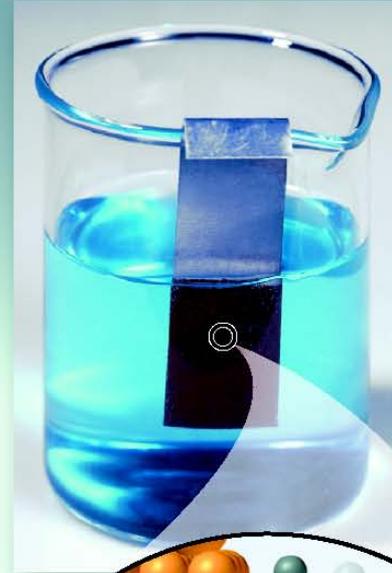
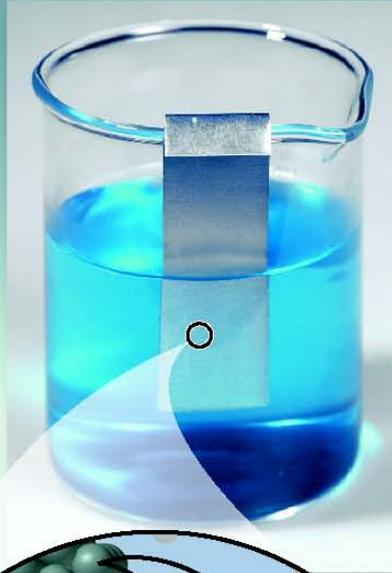


– **Copper ions are gaining the 2 electrons.**

- **Oxidizing agent**



# Reaction of Cu and $Zn^{2+}$ ions



# Example of redox reactions



Each sodium atom loses one electron:



Each chlorine atom gains one electron:



# Definitions

Loss of Electrons = Oxidation



Sodium is oxidized

Increase of the oxidation state

Gain of Electrons = Reduction



Chlorine is reduced

Decrease of the oxidation state

# Not All Reactions are Redox Reactions

- Reactions in which there has been **no change in oxidation number** are not redox reactions.

Examples:



# Assigning Oxidation States

- An “***oxidation state***” (*oxidation number*) is a positive or negative number assigned to an atom to indicate its degree of oxidation or reduction.

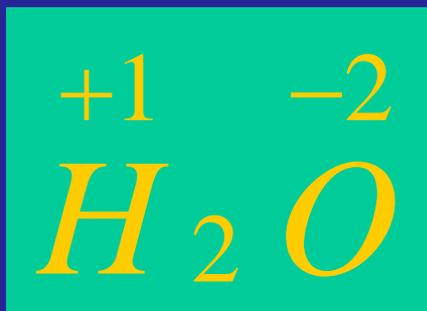
# Rules for Assigning Oxidation States

- 1) The oxidation State of any uncombined element is zero.
- 2) The oxidation State of a monatomic ion equals its charge.



# Rules for Assigning Oxidation States

- 3) The oxidation state of oxygen in compounds is -2, except in peroxides, such as  $\text{H}_2\text{O}_2$  where it is -1.
- 4) The oxidation state of hydrogen in compounds is +1, except in metal hydrides, like  $\text{NaH}$ , where it is -1.



# Rules for Assigning Oxidation states

5) The sum of the oxidation states of the atoms in the compound must equal 0.



$$2(+1) + (-2) = 0$$

H            O

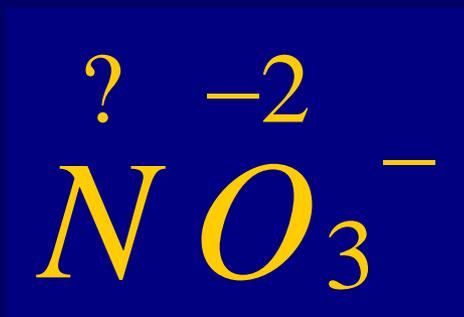


$$(+2) + 2(-2) + 2(+1) = 0$$

Ca            O            H

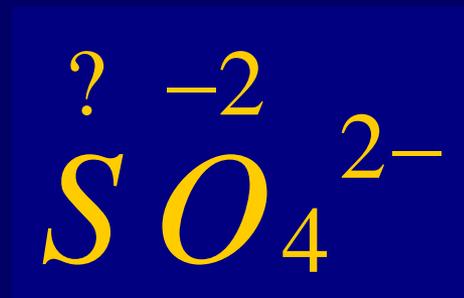
# Rules for Assigning Oxidation States

6) The sum of the oxidation States in the formula of a polyatomic ion is equal to its ionic charge.



$$\begin{array}{c} \text{X} + 3(-2) = -1 \\ \text{N} \quad \quad \text{O} \end{array}$$

$$\therefore \text{X} = +5$$



$$\begin{array}{c} \text{X} + 4(-2) = -2 \\ \text{S} \quad \quad \text{O} \end{array}$$

$$\therefore \text{X} = +6$$

What is the oxidation states of all atoms in the following ?



$$\text{F} = ?$$

$$7x - 1 + ? = 0$$

$$I =$$



$$\text{Na} = ? \quad \text{O} = ?$$

$$3x - 2 + 1 + ? = 0$$

$$I =$$



$$\text{O} = ? \quad \text{K} = ?$$

$$7x - 2 + 2x + 1 + 2x? = 0$$

$$\text{Cr} = ??$$

# Guidelines for Assigning Oxidation Numbers

**TABLE 4.5**

Elements with Reliable Oxidation Numbers in Compounds or Polyatomic Ions

Element	Oxidation Number	Exceptions
Fluorine	-1	
Group 1A or 2A metal	+1 or +2, respectively	
Hydrogen	+1	Any combination with a Group 1A or 2A metal to form a metal hydride. Examples: LiH and CaH <sub>2</sub> —the oxidation number of H is -1 in both examples.
Oxygen	-2	Anything higher on the list that necessitates its having a different oxidation number (see rule 2 for assigning oxidation numbers). Examples: H <sub>2</sub> O <sub>2</sub> and KO <sub>2</sub> —the oxidation number of O for H <sub>2</sub> O <sub>2</sub> is 1 and for KO <sub>2</sub> is - $\frac{1}{2}$ .
Group 7A (other than fluorine)	-1	Any combination with something higher on the list that necessitates its having a different oxidation number (see rule 2 for assigning oxidation numbers). Examples: ClF, BrO <sub>4</sub> <sup>-</sup> , and IO <sub>3</sub> <sup>-</sup> —the oxidation numbers of Cl, Br, and I are +1, +7, and +5, respectively.

# Oxidation of metals in aqueous solutions

## Displacement reactions

- A common reaction: active metal replaces (displaces) a metal ion from a solution



Does this reaction work?



**NOOOOOO**

- The activity series of metals is useful in order to predict the outcome of the reaction.

TABLE 4.6

## Activity Series

	Element	Oxidation Half-Reaction
 Increasing ease of oxidation	Lithium	$\text{Li} \longrightarrow \text{Li}^+ + e^-$
	Potassium	$\text{K} \longrightarrow \text{K}^+ + e^-$
	Barium	$\text{Ba} \longrightarrow \text{Ba}^{2+} + 2e^-$
	Calcium	$\text{Ca} \longrightarrow \text{Ca}^{2+} + 2e^-$
	Sodium	$\text{Na} \longrightarrow \text{Na}^+ + e^-$
	Magnesium	$\text{Mg} \longrightarrow \text{Mg}^{2+} + 2e^-$
	Aluminum	$\text{Al} \longrightarrow \text{Al}^{3+} + 3e^-$
	Manganese	$\text{Mn} \longrightarrow \text{Mn}^{2+} + 2e^-$
	Zinc	$\text{Zn} \longrightarrow \text{Zn}^{2+} + 2e^-$
	Chromium	$\text{Cr} \longrightarrow \text{Cr}^{3+} + 3e^-$
	Iron	$\text{Fe} \longrightarrow \text{Fe}^{2+} + 2e^-$
	Cadmium	$\text{Cd} \longrightarrow \text{Cd}^{2+} + 2e^-$
	Cobalt	$\text{Co} \longrightarrow \text{Co}^{2+} + 2e^-$
	Nickel	$\text{Ni} \longrightarrow \text{Ni}^{2+} + 2e^-$
	Tin	$\text{Sn} \longrightarrow \text{Sn}^{2+} + 2e^-$
	Lead	$\text{Pb} \longrightarrow \text{Pb}^{2+} + 2e^-$
	Hydrogen	$\text{H}_2 \longrightarrow 2\text{H}^+ + 2e^-$
	Copper	$\text{Cu} \longrightarrow \text{Cu}^{2+} + 2e^-$
	Silver	$\text{Ag} \longrightarrow \text{Ag}^+ + e^-$
	Mercury	$\text{Hg} \longrightarrow \text{Hg}^{2+} + 2e^-$
Platinum	$\text{Pt} \longrightarrow \text{Pt}^{2+} + 2e^-$	
Gold	$\text{Au} \longrightarrow \text{Au}^{3+} + 3e^-$	

## Balancing redox reactions

– **Electrons** (charge) must be balanced as well as **number and types of atoms**

– Consider this net ionic reaction:



– The reaction appears balanced as far as **number and type of atoms** are concerned, but look closely at the charge on each side.



- **Divide reaction into two half-reactions**



- **Multiply by a common factor to equalize electrons (the number of electrons lost must equal number of electrons gained)**



- Cancel electrons and write balanced net ionic reaction



**Predict whether each of the following will occur. For the reactions that do occur, write a balanced net ionic reaction for each.**

- **Copper metal is placed into a solution of silver nitrate**



- **A gold ring is accidentally dropped into a solution of hydrochloric acid**

**No reaction occurs, gold is below hydrogen on the activity series.**

# Other Types of Redox Reactions

## 1. Combination Reactions

- Many combination reactions may also be classified as redox reactions
- Consider:

Hydrogen gas reacts with oxygen gas



**Identify the substance oxidized and the substance reduced.**

## 2. Decomposition reactions

- Many decomposition reactions may also be classified as redox reactions
- Consider:

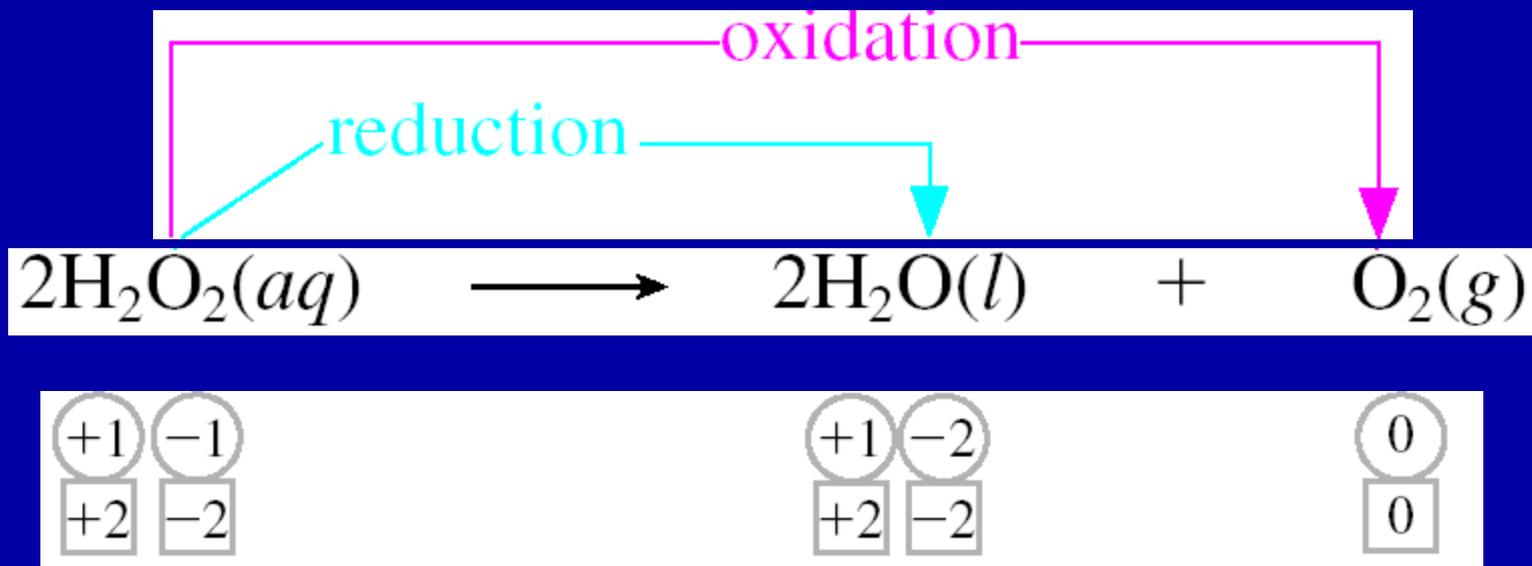
Potassium chlorate is strongly heated



Identify substances oxidized and reduced.

### 3. *Disproportionation* reactions

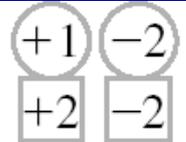
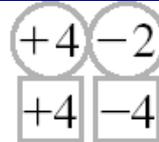
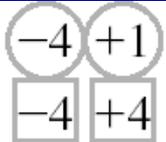
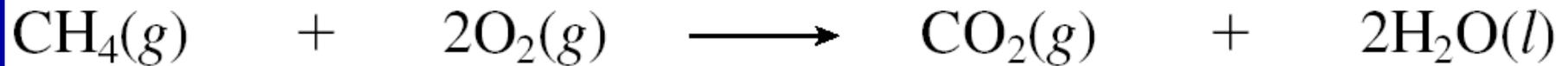
- One element undergoes both oxidation and reduction
- Consider:



## 4. Combustion reactions

- Common example, hydrocarbon fuel reacts with oxygen to produce carbon dioxide and water

– Consider:



# Oxidation Numbers on the Periodic Table

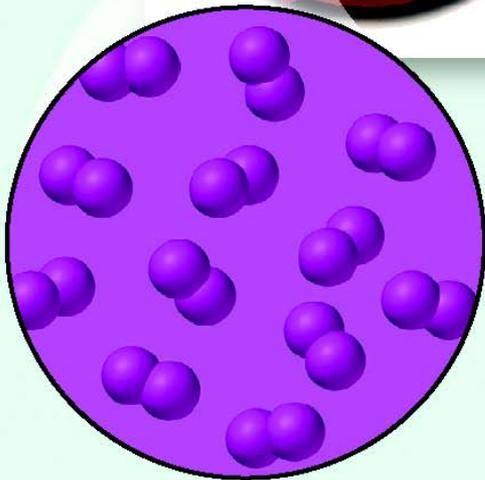
(most common in red)

1 1A												13 3A					14 4A	15 5A	16 6A	17 7A	18 8A
1 <b>H</b> +1 -1												5 <b>B</b> +3	6 <b>C</b> +4 +2 -4	7 <b>N</b> +5 +4 +3 +2 +1 -3	8 <b>O</b> +2 $-\frac{1}{2}$ -1 -2	9 <b>F</b> -1	10 <b>Ne</b>				
3 <b>Li</b> +1	4 <b>Be</b> +2											13 <b>Al</b> +3	14 <b>Si</b> +4 -4	15 <b>P</b> +5 +3 -3	16 <b>S</b> +6 +4 +2 -2	17 <b>Cl</b> +7 +6 +5 +4 +3 +1 -1	18 <b>Ar</b>				
11 <b>Na</b> +1	12 <b>Mg</b> +2	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B			11 1B	12 2B	31 <b>Ga</b> +3	32 <b>Ge</b> +4 -4	33 <b>As</b> +5 +3 -3	34 <b>Se</b> +6 +4 -2	35 <b>Br</b> +5 +3 +1 -1	36 <b>Kr</b> +4 +2				
19 <b>K</b> +1	20 <b>Ca</b> +2	21 <b>Sc</b> +3	22 <b>Ti</b> +4 +3 +2	23 <b>V</b> +5 +4 +3 +2	24 <b>Cr</b> +6 +5 +4 +3 +2	25 <b>Mn</b> +7 +6 +4 +3 +2	26 <b>Fe</b> +3 +2	27 <b>Co</b> +3 +2	28 <b>Ni</b> +2	29 <b>Cu</b> +2 +1	30 <b>Zn</b> +2	31 <b>Ga</b> +3	32 <b>Ge</b> +4 -4	33 <b>As</b> +5 +3 -3	34 <b>Se</b> +6 +4 -2	35 <b>Br</b> +5 +3 +1 -1	36 <b>Kr</b> +4 +2				
37 <b>Rb</b> +1	38 <b>Sr</b> +2	39 <b>Y</b> +	40 <b>Zr</b> +4	41 <b>Nb</b> +5 +4	42 <b>Mo</b> +6 +4 +3	43 <b>Tc</b> +7 +6 +4	44 <b>Ru</b> +8 +6 +4 +3	45 <b>Rh</b> +4 +3 +2	46 <b>Pd</b> +4 +2	47 <b>Ag</b> +1	48 <b>Cd</b> +2	49 <b>In</b> +3	50 <b>Sn</b> +4 +2	51 <b>Sb</b> +5 +3 -3	52 <b>Te</b> +6 +4 -2	53 <b>I</b> +7 +5 +1 -1	54 <b>Xe</b> +6 +4 +2				
55 <b>Cs</b> +1	56 <b>Ba</b> +2	57 <b>La</b> +3	72 <b>Hf</b> +4	73 <b>Ta</b> +5	74 <b>W</b> +6 +4	75 <b>Re</b> +7 +6 +4	76 <b>Os</b> +8 +4	77 <b>Ir</b> +4 +3	78 <b>Pt</b> +4 +2	79 <b>Au</b> +3 +1	80 <b>Hg</b> +2 +1	81 <b>Tl</b> +3 +1	82 <b>Pb</b> +4 +2	83 <b>Bi</b> +5 +3	84 <b>Po</b> +2	85 <b>At</b> -1	86 <b>Rn</b>				

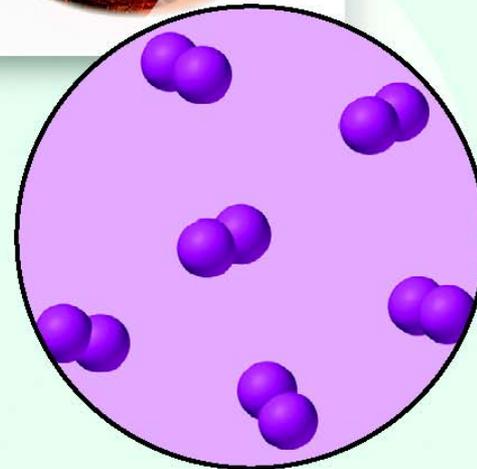
## 4.5 Concentration of Solutions

- **Concentration** is the amount of solute dissolved in a given amount of **solution**.
- Qualitative expressions of concentration
  - **Concentrated** – higher ratio of solute to solvent
  - **Dilute** - smaller ratio of solute to solvent

# Comparison of a Concentrated and Dilute Solution



**Concentrated solution:**  
More solute particles per unit volume



**Dilute solution:**  
Fewer solute particles per unit volume

# Molarity

Quantitative concentration term

- **Molarity** is the ratio of moles solute per liter of solution

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

- Symbols:  $M$  or  $[ ]$
- Different forms of molarity equation

$$M = \frac{\text{mol}}{L}$$

$$L = \frac{\text{mol}}{M}$$

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

# Molarity Calculation

If 4.0 g NaOH are used to make 500. mL of NaOH solution, what is the molarity (M) of the solution?

# Calculating Molarity

$$4.0 \text{ g NaOH} \times \frac{1 \text{ mole NaOH}}{40.0 \text{ g NaOH}} = 0.10 \text{ mole NaOH}$$

$$500. \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.500 \text{ L}$$

$$\frac{0.10 \text{ mole NaOH}}{0.500 \text{ L}} = \frac{0.20 \text{ mole NaOH}}{1 \text{ L}} = 0.20 \text{ M NaOH}$$

## Example

An acid solution is a 0.10 M HCl. How many moles of HCl are in 1500 mL of this acid solution?

$$1500 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 1.5 \text{ L}$$

$$1.5 \text{ L} \times \frac{0.10 \text{ mole HCl}}{1 \text{ L}} = 0.15 \text{ mole HCl}$$

## Example

How many grams of KCl are present in 2.5 L of 0.50 M KCl?

$$2.5 \text{ L} \times \frac{0.50 \text{ mole}}{1 \text{ L}} \times \frac{74.6 \text{ g KCl}}{1 \text{ mole KCl}} = 93 \text{ g KCl}$$

## Example

How many milliliters of a 0.10 M HCl solution that contains 0.15 mole HCl?

$$\begin{aligned} &0.15 \text{ mole HCl} \times \frac{1 \text{ L solution}}{0.10 \text{ mole HCl}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \\ &= 1500 \text{ mL HCl} \end{aligned}$$

## Example

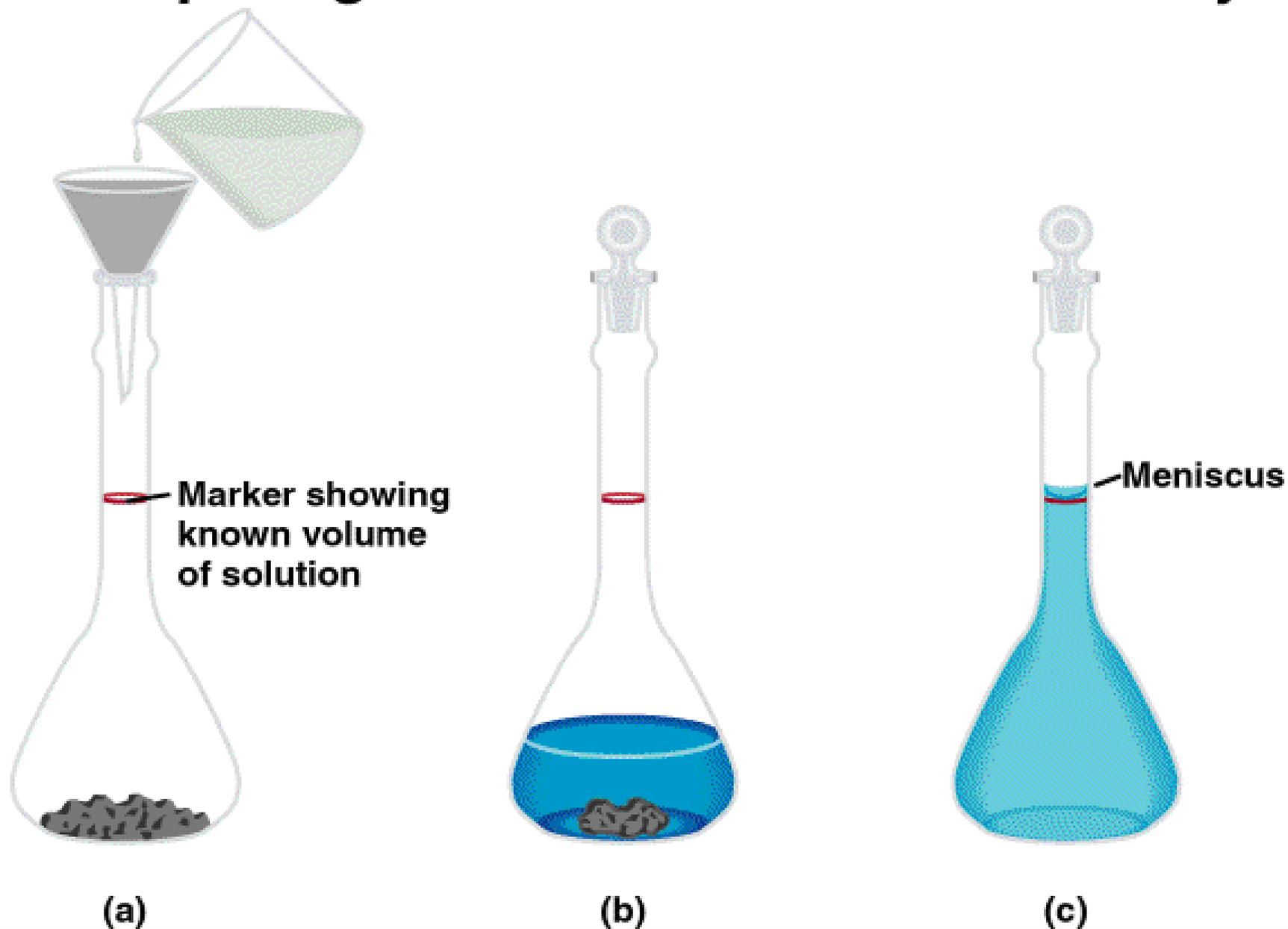
How many grams of NaOH are required to prepare 400. mL of 3.0 M NaOH solution?

$$400. \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.400 \text{ L}$$

$$0.400 \text{ L} \times \frac{3.0 \text{ mole NaOH}}{1 \text{ L}} \times \frac{40.0 \text{ g NaOH}}{1 \text{ mole NaOH}}$$

$$= 48 \text{ g NaOH}$$

# Preparing a Solution of Known Molarity



## Example

A sample of 0.14 M NaCl. What volume of sample contains 1.0 mg NaCl?

$$1.0 \text{ mg NaCl} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol NaCl}}{58.5 \text{ g NaCl}} \times \frac{1 \text{ L}}{0.14 \text{ mol NaCl}} = 1.2 \times 10^{-4} \text{ L}$$

$$\# \text{ mol} = M \times V \text{ (L)}$$



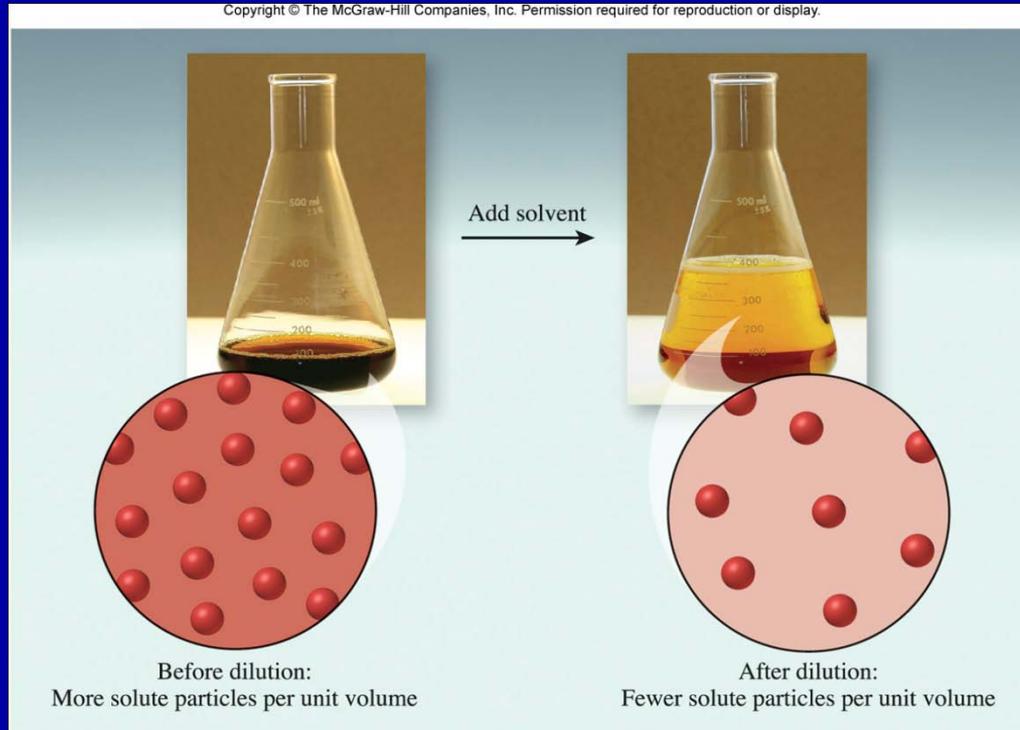
$$V(L) = \frac{\# \text{ mol}}{M}$$

$$V(L) = \frac{0.1 \times 10^{-3} \text{ g} / 58.5 \text{ g} / \text{mol}}{0.14 \text{ mol} / \text{L}} = 1.2 \times 10^{-4} \text{ L}$$

# Dilution

- Adding more solvent to a known solution.
- The moles of solute stay the same.
- #moles =  $M \times \text{volume (L)}$
  
- # moles before dilution (1) = # moles after dilution
  
- $M_1 V_1 = M_2 V_2$
  
- **Stock solution** is a solution of known concentration used to make more dilute solutions

# Preparing a less concentrated solution from a more concentrated solution by dilution



Moles of solute  
before dilution (i)

=

Moles of solute  
after dilution (f)

$$M_i V_i$$

=

$$M_f V_f$$

# Example

How would you prepare 60.0 mL of 0.200 M HNO<sub>3</sub> from a stock solution of 4.00 M HNO<sub>3</sub>?

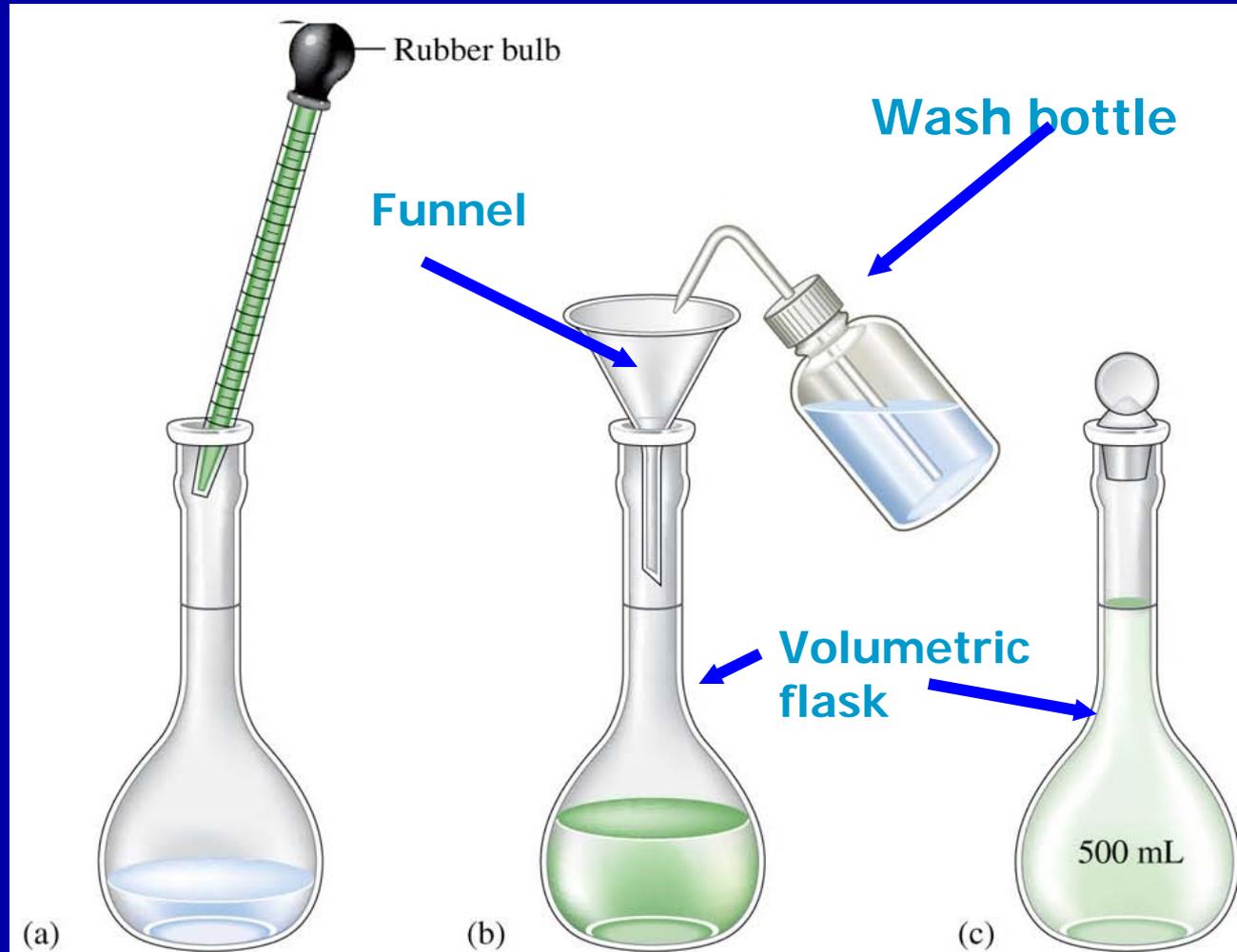
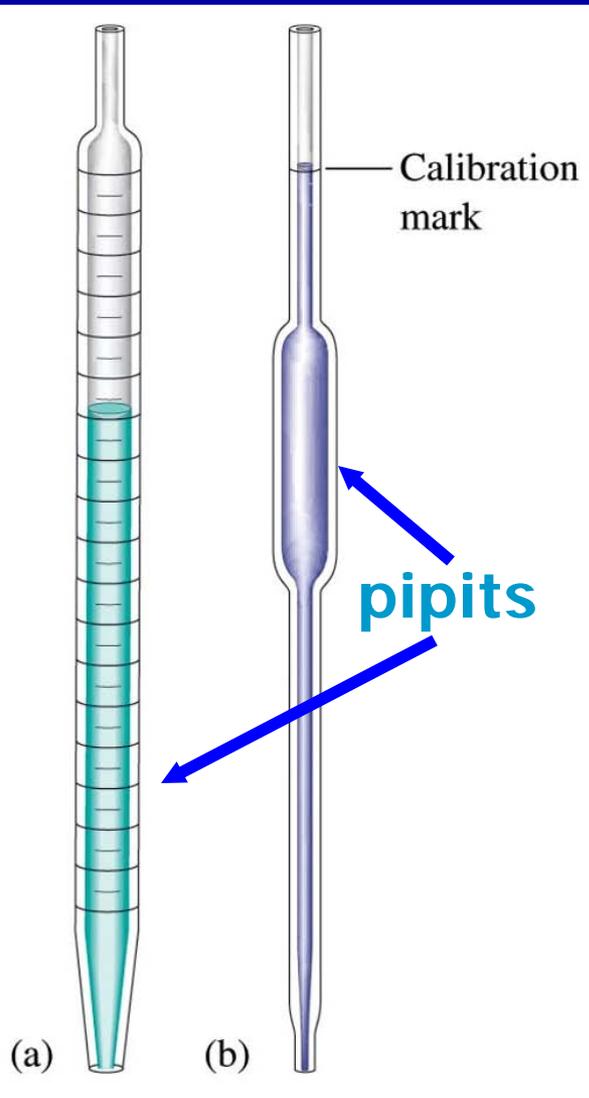
What volume of 4.00 M HNO<sub>3</sub> must be used to prepare 60.0 mL of 0.200 M HNO<sub>3</sub>?

$$M_1 = 4.00 \quad M_2 = 0.200 \quad V_2 = 60.0 \text{ mL} \quad V_1 = ? \text{ mL}$$

$$M_1 V_1 = M_2 V_2$$

$$V_1 = \frac{M_2 V_2}{M_1} = \frac{0.200 \text{ M} \times 60.0 \text{ mL}}{4.00 \text{ M}} = 3.00 \text{ mL}$$

# Dilution process



# Solution Stoichiometry

- Soluble ionic compounds dissociate completely in solution.
- Using mole ratios we can calculate the concentration of all species in solution.

1 mol NaCl dissociates into 1 mol  $\text{Na}^+$  and 1 mol  $\text{Cl}^-$

1 mol  $\text{Na}_2\text{SO}_4$  dissociates into 2 mol  $\text{Na}^+$  and 1 mol  $\text{SO}_4^{2-}$

1 mol  $\text{AlCl}_3$  dissociates into 1 mol  $\text{Al}^{3+}$  and 3 mol  $\text{Cl}^-$

**Find the concentration of all species in a 0.25 M solution of MgCl<sub>2</sub>**



**Given: MgCl<sub>2</sub> = 0.25 M**

$$[\text{Mg}^{2+}] = 0.25 \text{ M (1:1 ratio)}$$

$$[\text{Cl}^-] = 0.50 \text{ M (1:2 ratio)}$$

Using the square bracket notation,  
express the molar concentration for all  
species in the following solutions

**0.42 M Ba(OH)<sub>2</sub>**

**[Ba<sup>2+</sup> ] = 0.42 M (1:1 ratio)**

**[OH<sup>-</sup> ] = 0.84 M (2:1 ratio)**

**1.2 M NH<sub>4</sub>Cl**

**[NH<sub>4</sub><sup>+</sup> ] = 1.2 M (1:1 ratio)**

**[Cl<sup>-</sup> ] = 1.2 M (1:1 ratio)**

## **4.6 Aqueous Reactions and Chemical Analysis**

- **Types of quantitative analysis**
  - **Gravimetric analysis (mass analysis)**
    - **Example: precipitation reaction**
  - **Volumetric analysis (volume analysis)**
    - **Example: titration**

# Gravimetric Analysis

- One form: isolation of a precipitate
- Typical steps:
  - Determine mass of unknown solid
  - Dissolve unknown in water
  - Combine with excess amount of known substance to form a precipitate (excess drives reaction to completion)
  - Filter, dry and weigh the precipitate
  - Use formula and mass of ppt to find % of ion in unknown solid

**A 0.825 g sample of an ionic compound containing chloride ions and an unknown metal is dissolved in water and treated with excess silver nitrate. If 1.725 g of AgCl precipitate forms, what is the percent by mass of Cl in the original sample?**

## **Steps in solution:**

- **Find the % of Cl in AgCl**
- **Multiply the % of Cl by the mass of the precipitate to obtain the Cl in the sample**
- **Divide the mass of Cl in sample by total mass of sample (multiply by 100 for %)**

$$\% \text{ Cl} = \frac{35.45 \text{ g Cl}}{143.35 \text{ g AgCl}} \times 100 = 24.7\%$$

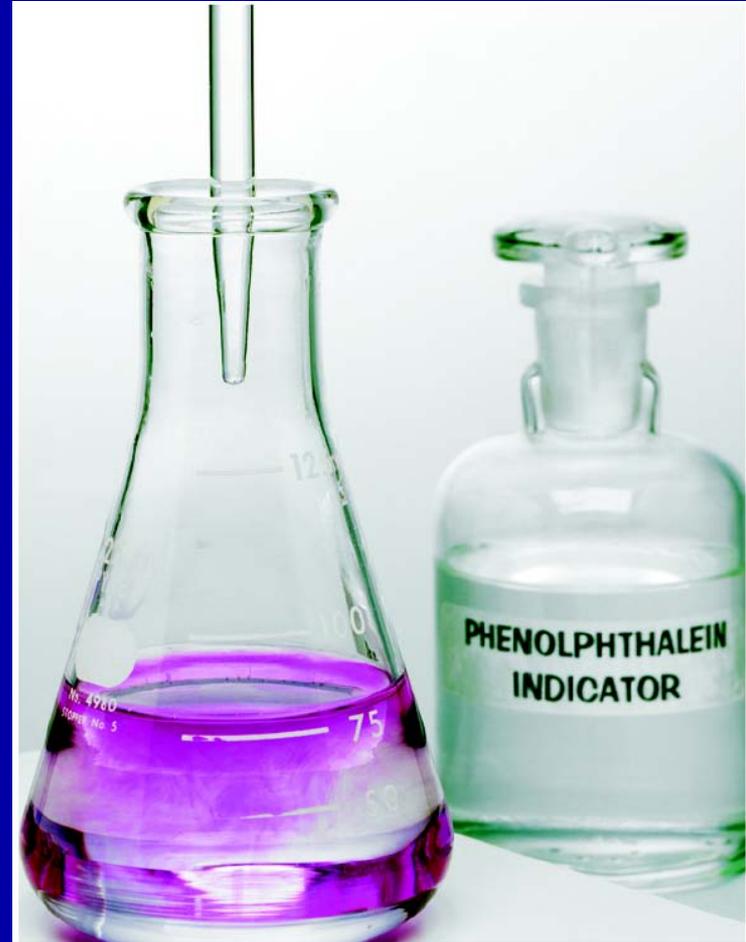
$$0.247 \times 1.725 \text{ g AgCl ppt} = 0.427 \text{ g Cl in sample}$$

$$\% \text{ Cl in unknown} = \frac{0.427 \text{ g Cl}}{0.825 \text{ g sample}} \times 100 = 51.7\% \text{ Cl}$$

# Volumetric analysis

- Commonly accomplished by *titration*
  - Addition of a solution of known concentration (standard solution) to another solution of unknown concentration.

# Apparatus for a Titration



# Volumetric analysis

- *Standardization* is the determination of the exact concentration of a solution.
- *Equivalence point* represents completion of the reaction.
- *Endpoint* is where the titration is stopped.
- *An indicator* is used to signal the endpoint.

**A student measured exactly 15.0 mL of an unknown monoprotic acidic solution and placed in an Erlenmeyer flask. An indicator was added to the flask. At the end of the titration the student had used 35.0 mL of 0.12 M NaOH to neutralize the acid. Calculate the molarity of the acid.**

$$0.035 \text{ L NaOH} \times \frac{0.12 \text{ mol NaOH}}{1 \text{ L}} \times \frac{1 \text{ mol acid}}{1 \text{ mol base}} = 0.0042 \text{ mol acid}$$

$$M = \frac{0.0042 \text{ mol}}{0.015 \text{ L}} = 0.28 \text{ M acid}$$

Calculate the molarity of 25.0 mL of a monoprotic acid if it took 45.50 mL of 0.25 *M* KOH to neutralize the acid.

$$\frac{0.25 \text{ mol KOH}}{\text{L}} \times 0.04550 \text{ L} \times \frac{1 \text{ mol acid}}{1 \text{ mol KOH}} = 0.01338 \text{ mol acid}$$

$$\frac{0.01338 \text{ mol acid}}{0.0250 \text{ L}} = 0.455 \text{ M}$$

# Titration

- A 50.00 mL sample of aqueous  $\text{Ca(OH)}_2$  requires 34.66 mL of 0.0980 M Nitric acid for neutralization. What is  $[\text{Ca(OH)}_2]$ ?



- # of moles of  $\text{HNO}_3$  = # of moles of  $\text{Ca(OH)}_2$  X  $\frac{2\text{HNO}_3}{\text{Ca(OH)}_2}$
- $0.098 \times 34.66 = M_{\text{Ca(OH)}_2} \times 50.00$
- $M_{\text{Ca(OH)}_2} = (0.098 \times 34.66 / 50.00)$

# Key Points

- **Electrolytes (strong, weak, and non)**
- **Precipitation reactions**
  - **Solubility rules**
- **Molecular, ionic, and net ionic reactions**
- **Acid-base neutralization reactions**
- **Oxidation-reduction reactions**