



# Chapter 4

## Reactions in Aqueous Solutions

# 4.1 General Properties of Aqueous Solutions

- **Solution** - a homogeneous mixture
  - **Solute:** the component that is dissolved
  - **Solvent:** the component that does the dissolving

Generally, the component present in the greatest quantity is considered to be the solvent. *Aqueous* solutions are those in which *water* is the solvent.

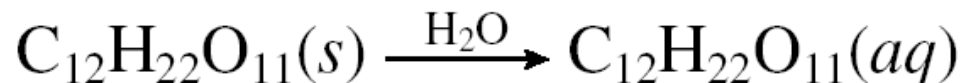
- Electrolytes and Nonelectrolytes
  - **Electrolyte**: substance that dissolved in water produces a solution that conducts electricity

- Contains ions



- **Nonelectrolyte**: substance that dissolved in water produces a solution that does not conduct electricity

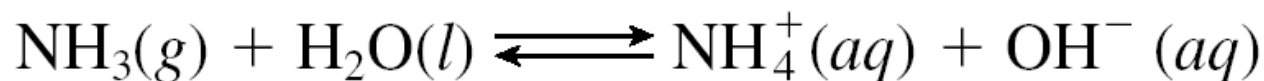
- Does not contain ions



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



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



- Strong and weak electrolytes
  - ***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

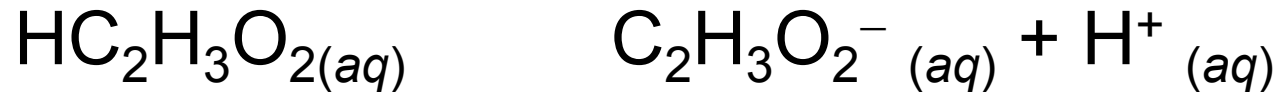
**TABLE 4.1****The Strong Acids**

| <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**  $\rightleftharpoons$

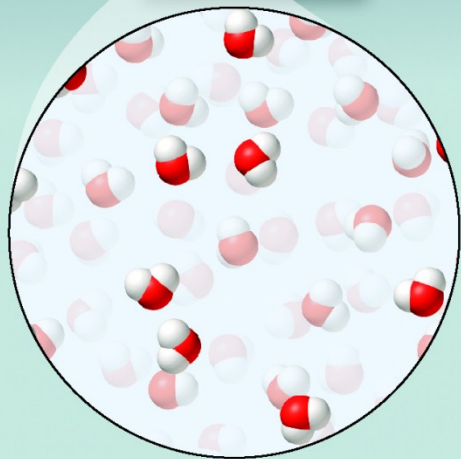


– **Weak bases**  $\rightleftharpoons$

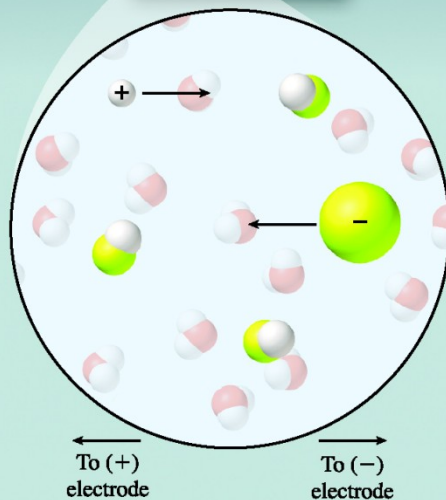
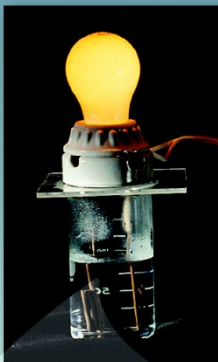


(Note: double arrows indicate a reaction that occurs in both directions - a state of *dynamic equilibrium* exists)

# Method to Distinguish Types of Electrolytes

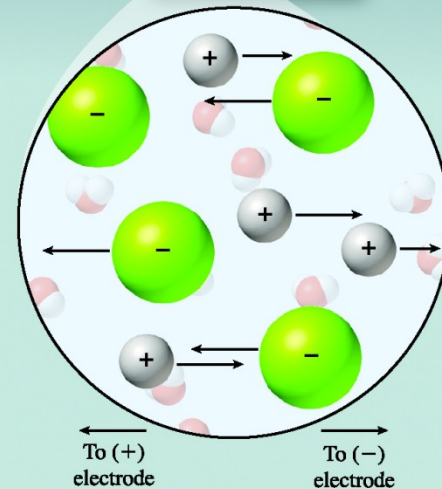


nonelectrolyte



weak electrolyte

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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

- **Precipitation** (formation of a solid from two aqueous solutions) occurs when product is insoluble
- Produce insoluble ionic compounds
- **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

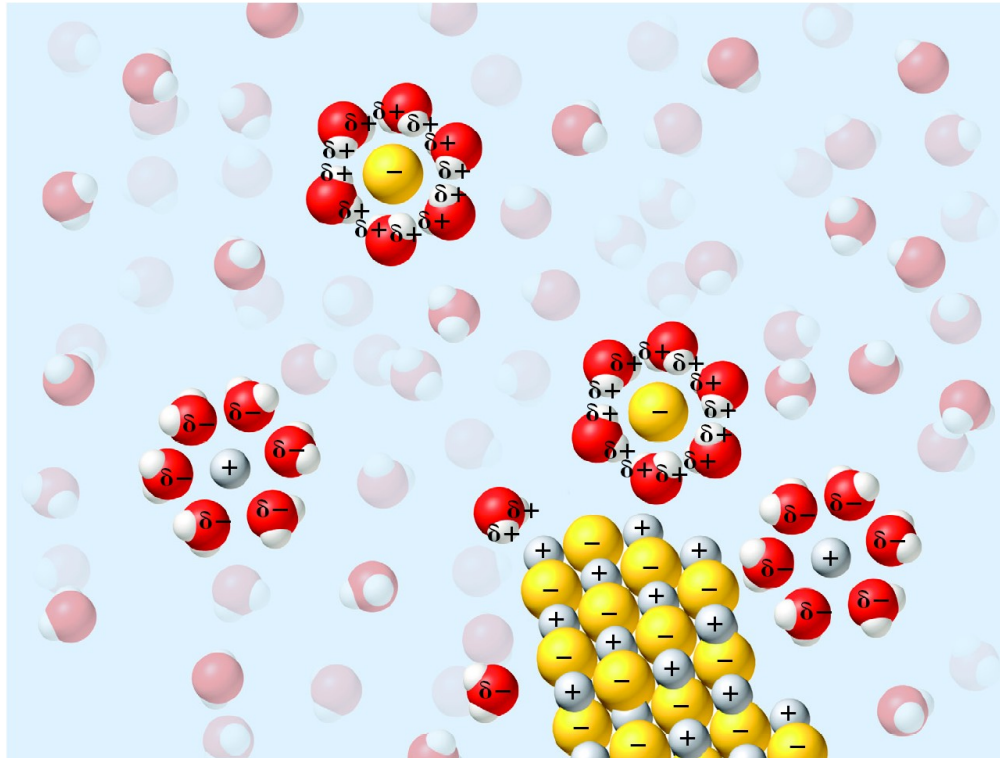
**TABLE 4.2****Solubility Guidelines: Soluble Compounds**

| <b>Water-Soluble Compounds</b>   | <b>Insoluble Exceptions</b>  |
|--|--|
| 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 $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , or $\text{Pb}^{2+}$  |
| Compounds containing the sulfate ion ( $\text{SO}_4^{2-}$ )  | Compounds containing $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , $\text{Pb}^{2+}$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , or $\text{Ba}^{2+}$ |

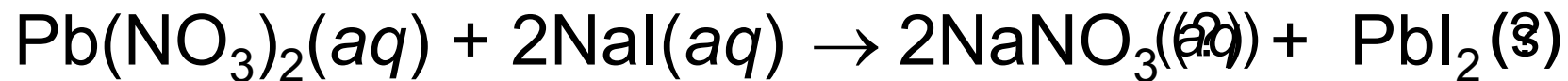
**TABLE 4.3****Solubility Guidelines: Insoluble Compounds**

| <b>Water-Insoluble Compounds</b>  | <b>Soluble Exceptions</b>   |
|---|---|
| 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 $\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$ , $\text{Rb}^+$ , $\text{Cs}^+$ , or $\text{NH}_4^+$  |
| Compounds containing the hydroxide ion ( $\text{OH}^-$ )  | Compounds containing $\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$ , $\text{Rb}^+$ , $\text{Cs}^+$ , or $\text{Ba}^{2+}$ |

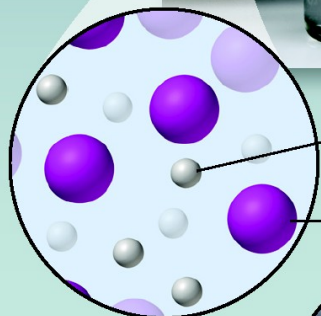
- **Hydration:** process by which water molecules remove and surround individual ions from the solid.



## Identify the Precipitate



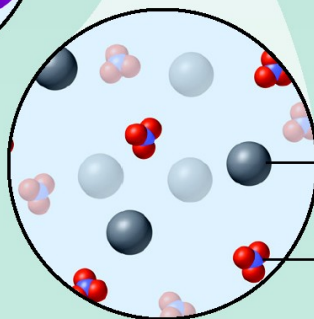
# Mixing Solutions of $\text{Pb}(\text{NO}_3)_2$ and $\text{NaCl}$



$\text{Na}^+$

$\text{I}^-$

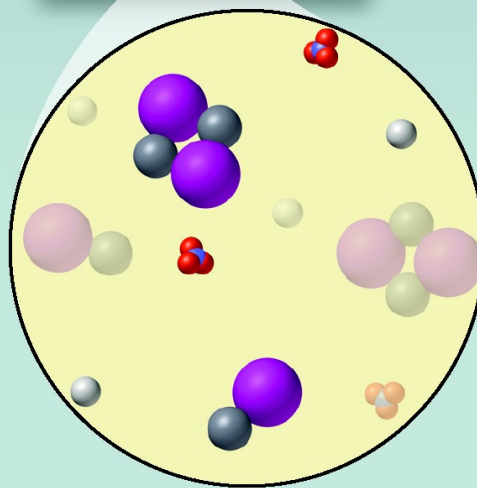
The addition of a colorless  $\text{NaI}(\text{aq})$  solution...



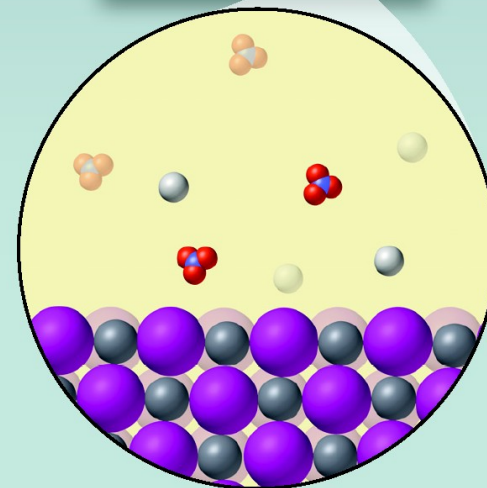
$\text{Pb}^{2+}$

$\text{NO}_3^-$

to a colorless  $\text{Pb}(\text{NO}_3)_2(\text{aq})$  solution...

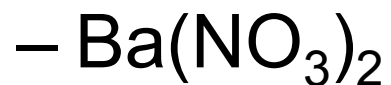


produces  $\text{PbI}_2(\text{s})$ , a yellow precipitate...



which settles out of solution. The remaining solution contains  $\text{Na}^+$  and  $\text{NO}_3^-$  ions.

Classify the following as soluble or insoluble in water



soluble

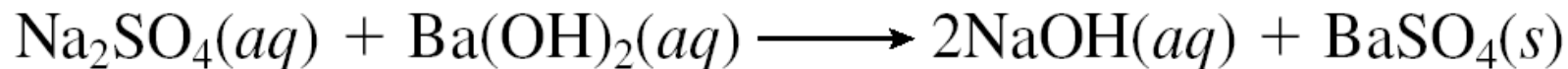


insoluble



insoluble

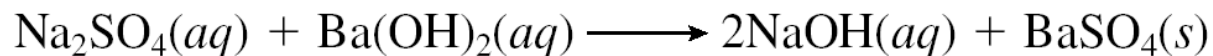
- ***Molecular equation:*** shows all compounds represented by their chemical formulas



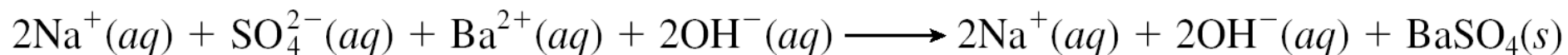
- ***Ionic equation:*** shows all strong electrolytes as ions and all other substances (non-electrolytes, weak electrolytes, gases) by their chemical formulas



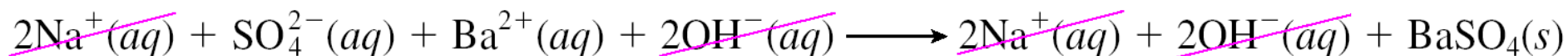
Molecular equation:



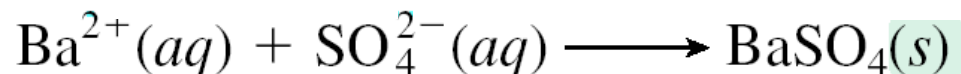
Ionic equation:



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



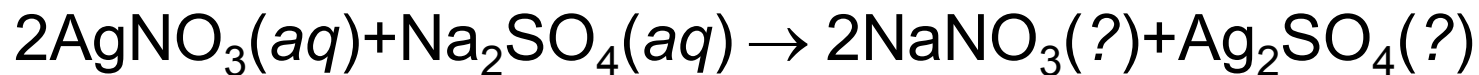
Net ionic equation:



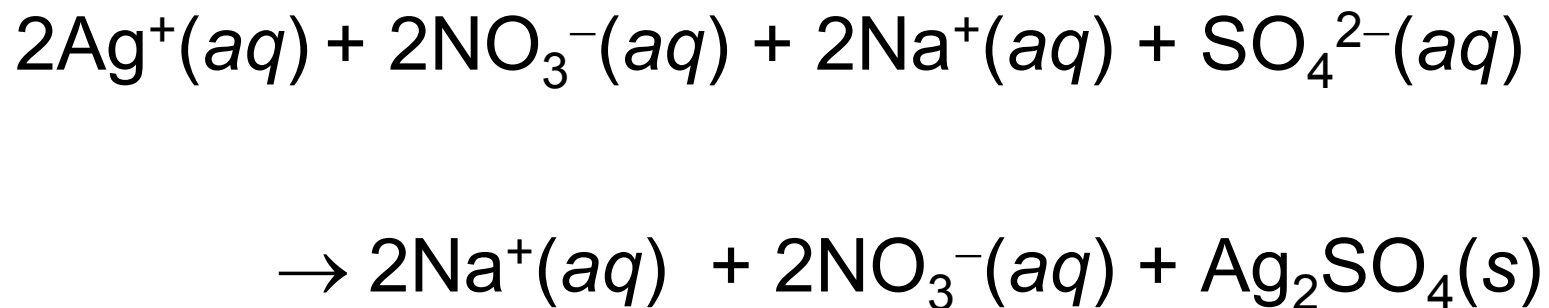
- Steps in writing a net ionic equation
  - Write the balanced molecular equation.
    - Predict products by exchanging cations and anions in reactants.
  - Separate strong electrolytes into ions.
  - Cancel spectator ions.
  - Use the remaining species to write the net ionic equation.

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

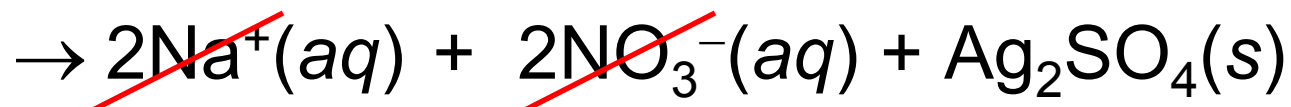
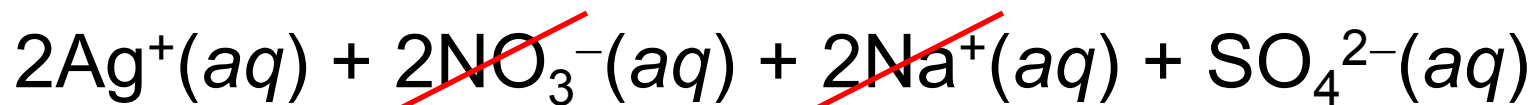
Step 1:



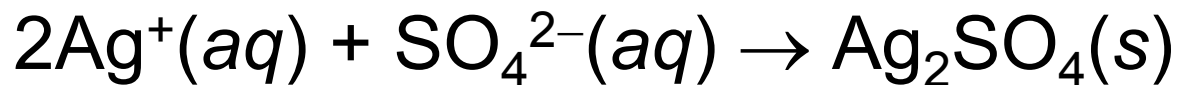
Step 2: Use solubility table; all nitrates are soluble but silver sulfate is insoluble



### Step 3: Cancel spectators



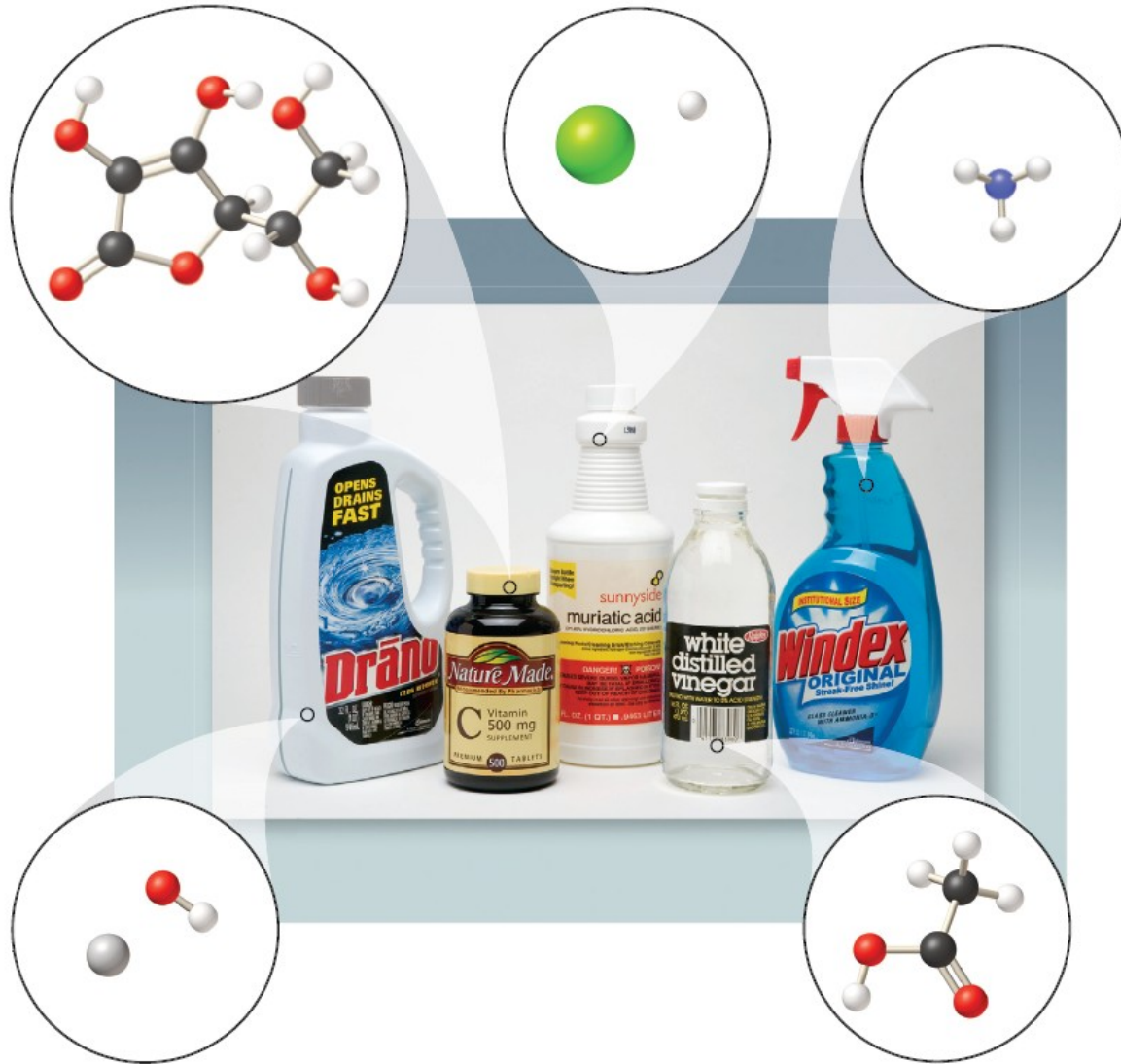
### Step 4: Write the net ionic reaction



# 4.3 Acid-Base Reactions

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

# Common Acids and Bases



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

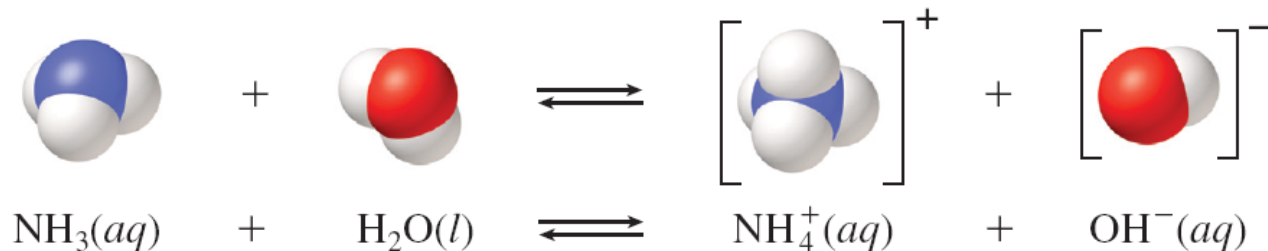
| <b>Strong Acids</b> | <b>Strong Bases</b> | <b>Strong Acids</b>            | <b>Strong Bases</b> |
|---------------------|---------------------|--------------------------------|---------------------|
| HCl                 | LiOH                | HClO <sub>3</sub>              | CsOH                |
| 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).

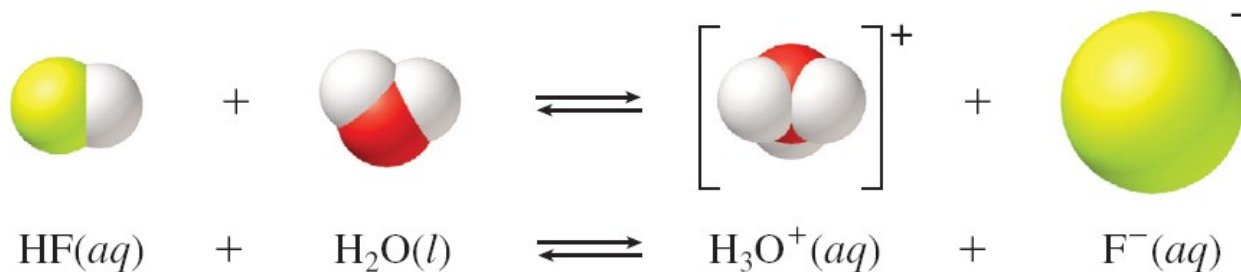
- Definitions of acids and bases
  - **Arrhenius acid** - produces  $H^+$  in solution
  - **Arrhenius base** - produces  $OH^-$  in solution
  - More inclusive definitions:
    - **Brønsted acid** - proton donor
    - **Brønsted base** - proton acceptor

– Examples of a weak base and weak acid

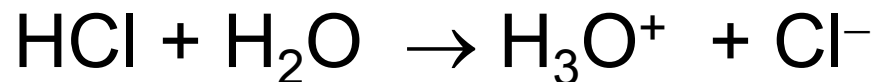
• Ammonia with water:



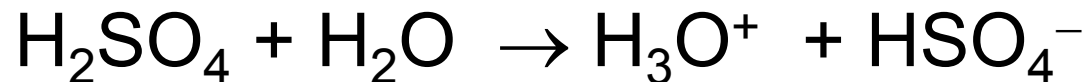
• Hydrofluoric acid with water:



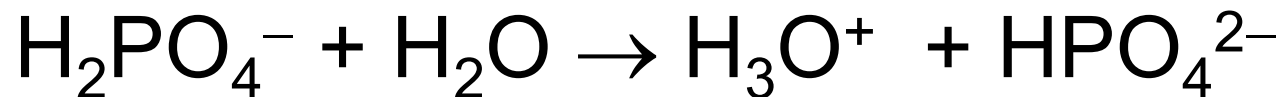
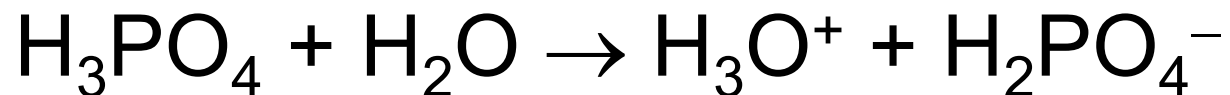
- Types of acids
  - ***Monoprotic***: one ionizable hydrogen



- ***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

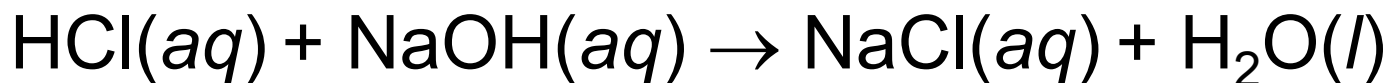


# Acid-Base Neutralization

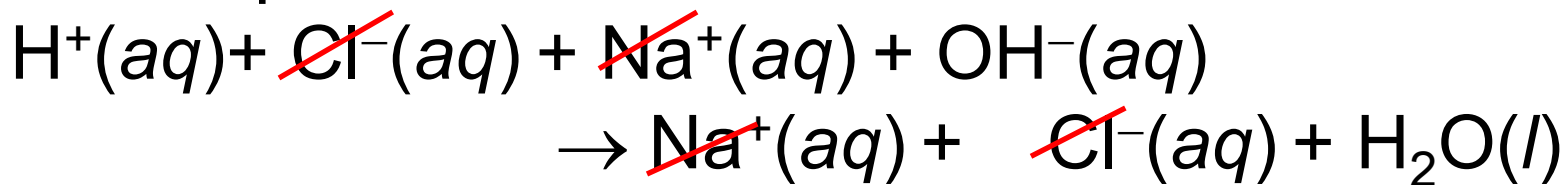
- Neutralization: Reaction between an acid and a base



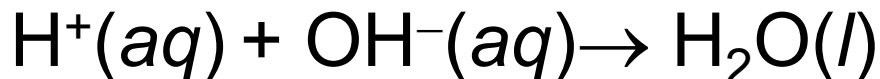
Molecular equation:



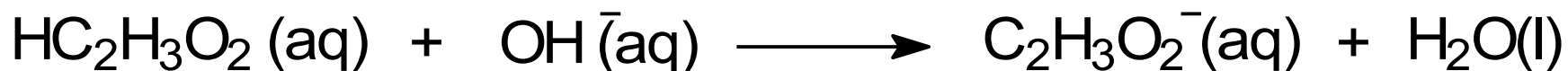
Ionic equation:



Net ionic equation:



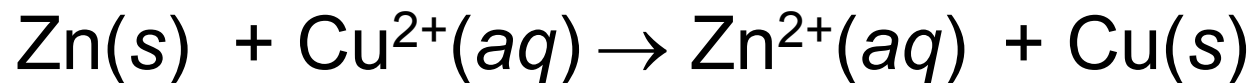
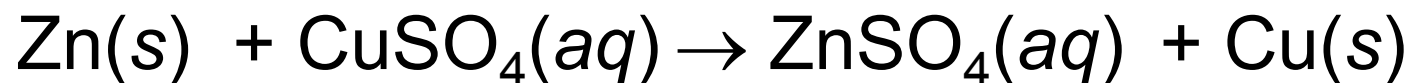
Solutions of acetic acid and lithium hydroxide are mixed. Write the net ionic reaction.



# 4.4 Oxidation-Reduction Reactions

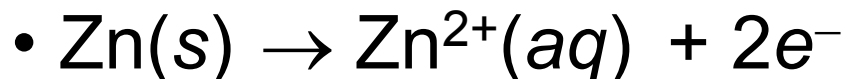
- Often called “redox” reactions
- Electrons are transferred between the reactants
  - One substance is oxidized, loses electrons
    - Reducing agent
  - Another substance is reduced, gains electrons
    - Oxidizing agent
- Oxidation numbers change during the reaction

– Example



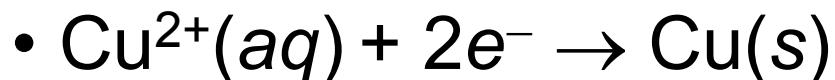
– Zinc is losing 2 electrons and oxidized.

- Reducing agent

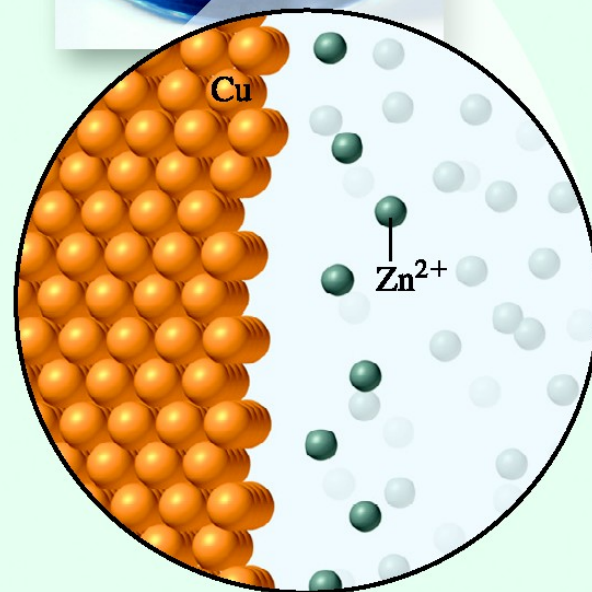
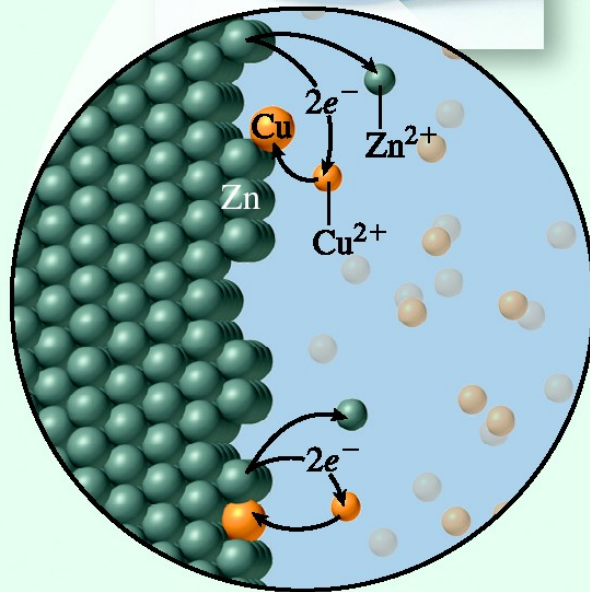
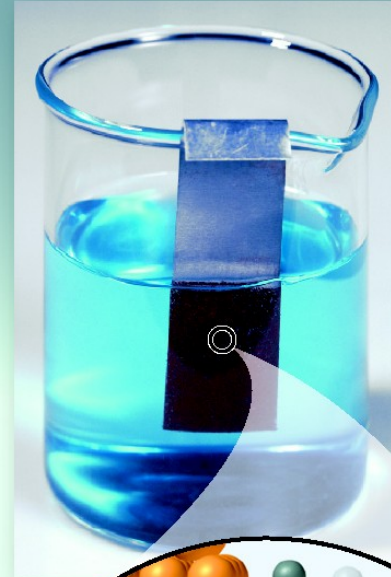
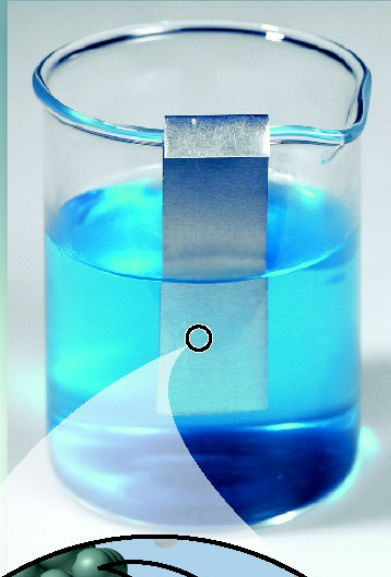


– Copper ions are gaining the 2 electrons.

- Oxidizing agent



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



- Rules for assigning oxidation numbers

1. Elements (uncombined) are 0.



2. Oxidation numbers must sum to the overall charge of the species.



$$? + 4(-2) = -2$$

Solve: ? - 8 = -2            ? = + 6 (S)

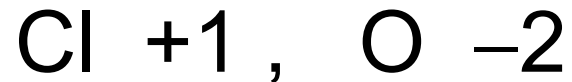
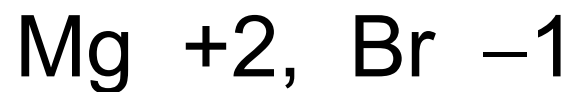
# Guidelines for Assigning Oxidation Numbers

**TABLE 4.5**

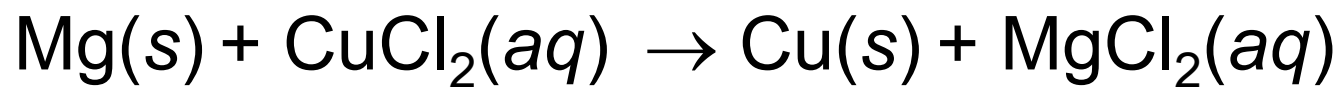
Elements with Reliable Oxidation Numbers in Compounds or Polyatomic Ions

| 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                     | Any combination with something 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 -½. |
| 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.     |

Assign oxidation numbers for all elements  
in each species



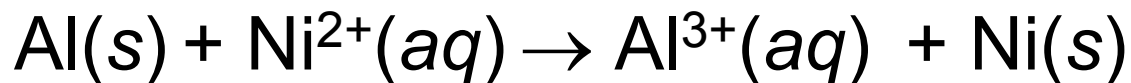
- Displacement reactions
  - A common reaction: active metal replaces (displaces) a metal ion from a solution



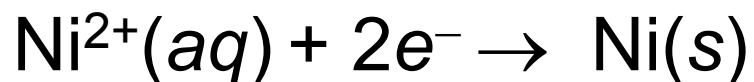
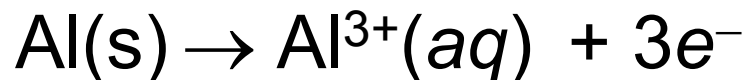
- The activity series of metals is useful in order to predict the outcome of the reaction.

| ↑<br>Increasing ease of oxidation | Element   | Oxidation Half-Reaction                       |
|-----------------------------------|---|---|
|                                   | 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                              | $\text{Au} \longrightarrow \text{Au}^{3+} + 3e^-$ |   |

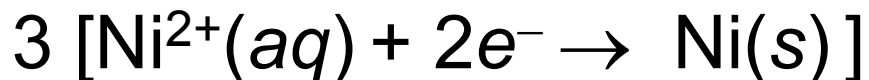
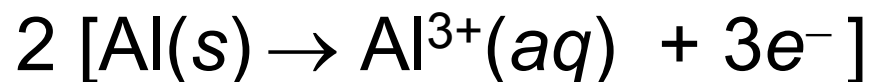
- Balancing redox reactions
  - Electrons (charge) must be balanced as well as number and types of atoms
  - Consider this net ionic reaction:
$$\text{Al}(s) + \text{Ni}^{2+}(aq) \rightarrow \text{Al}^{3+}(aq) + \text{Ni}(s)$$
  - The reaction appears balanced as far as number and type of atoms are concerned, but look closely at the charge on each side.



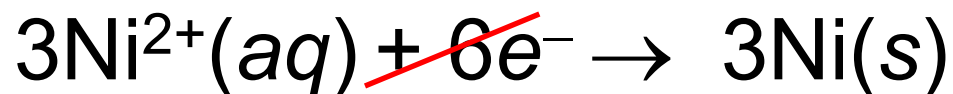
- Divide reaction into two half-reactions



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

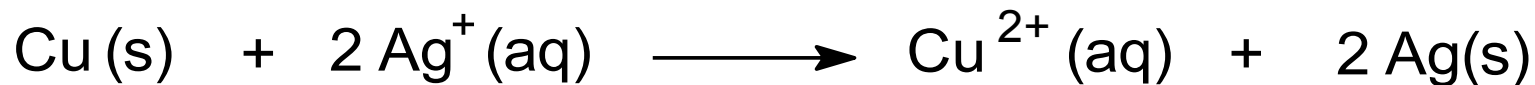


- Cancel electrons and write balanced net ionic reaction



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

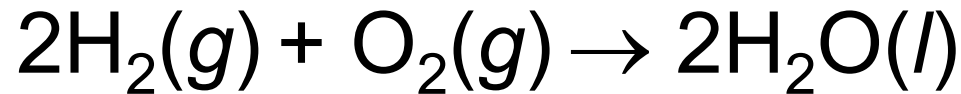
- Copper metal is placed into a solution of silver nitrate



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

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

- Combination Reactions
  - Many combination reactions may also be classified as redox reactions
  - Consider:  
Hydrogen gas reacts with oxygen gas



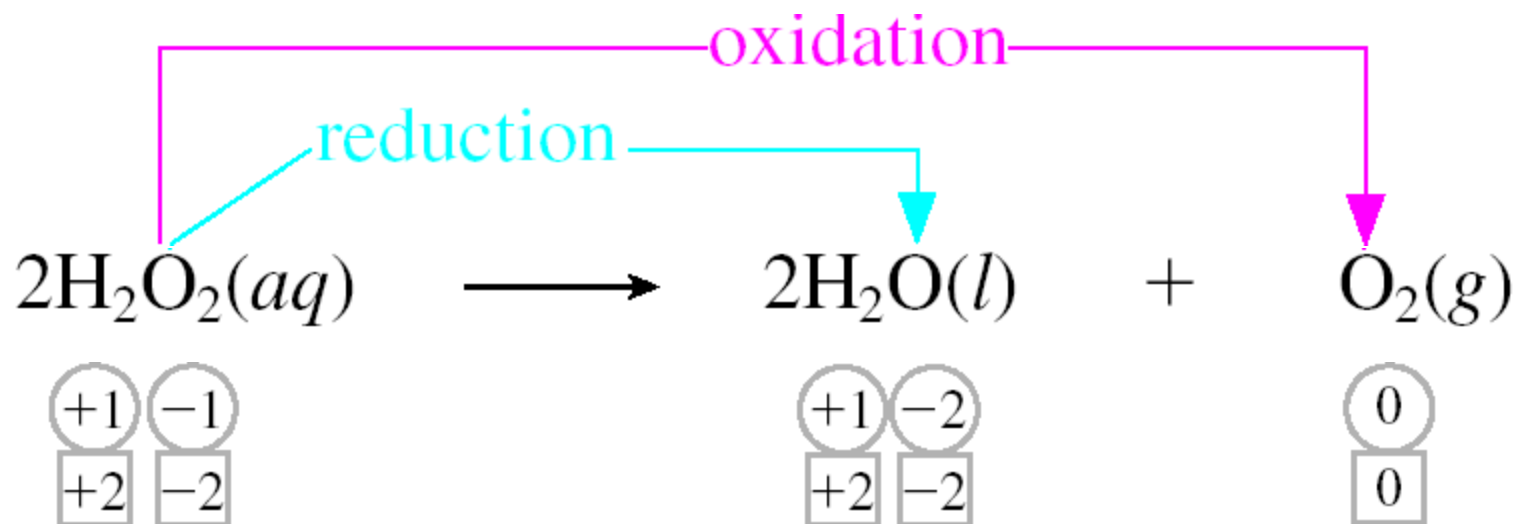
Identify the substance oxidized and the substance reduced.

- Decomposition reactions
  - Many decomposition reactions may also be classified as redox reactions
  - Consider:  
Potassium chlorate is strongly heated



Identify substances oxidized and reduced.

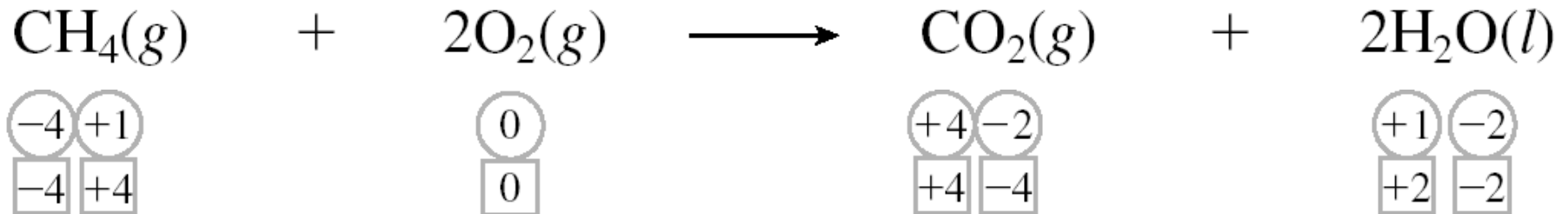
- ***Disproportionation*** reactions
  - One element undergoes both oxidation and reduction
  - Consider:



- **Combustion** reactions

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

- Consider:



# Oxidation Numbers on the Periodic Table

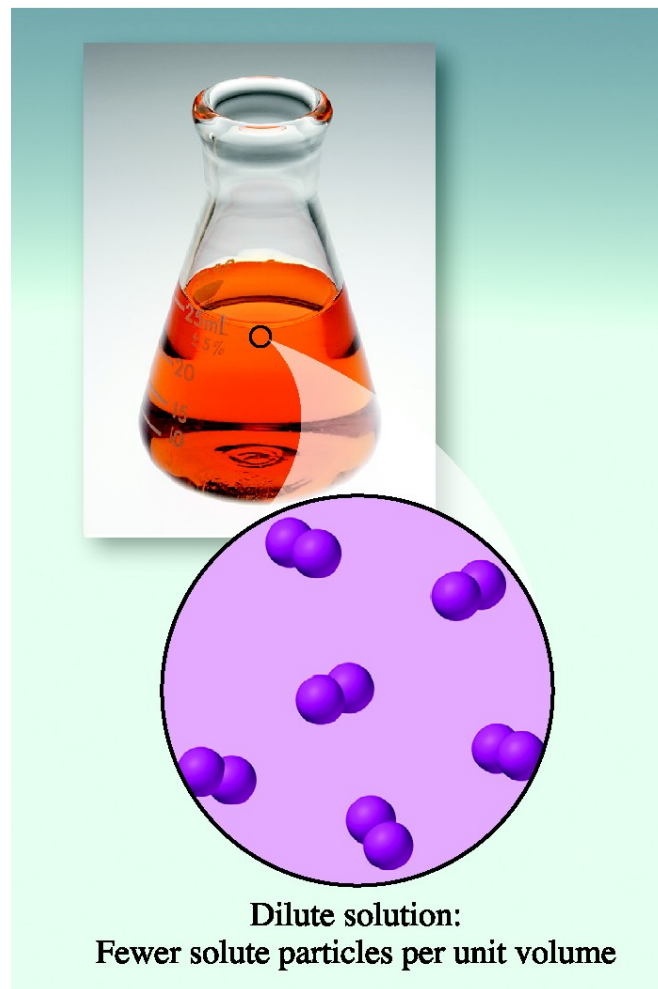
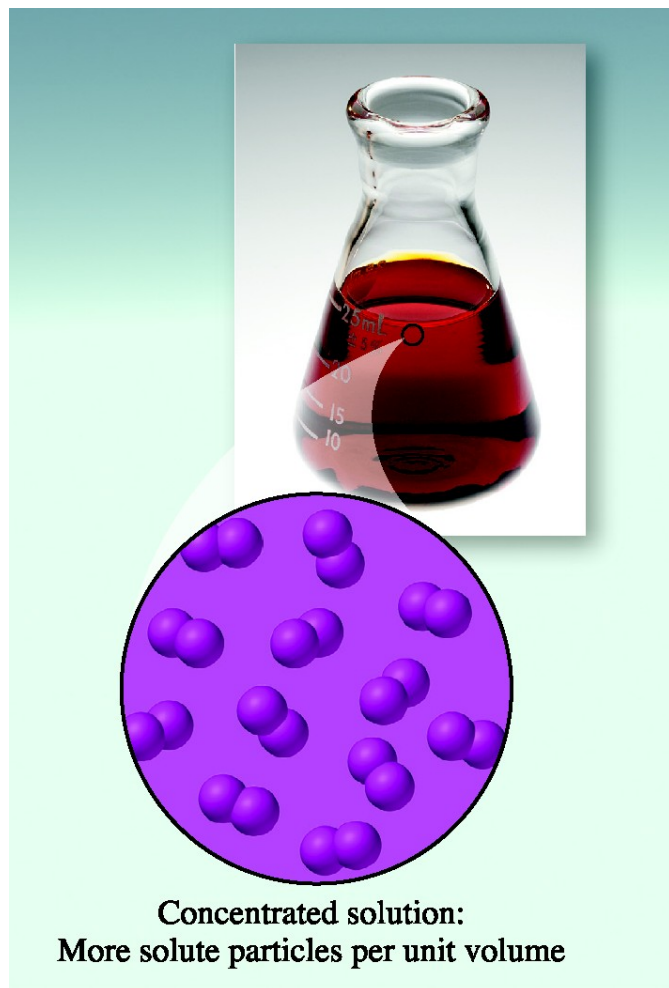
| 1<br>1A                   |                       |                       |                                   |  |   |   |   |                                   |                             |                             | 13<br>3A                    | 14<br>4A                        | 15<br>5A  |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

(most common in red)

# 4.5 Concentration of Solutions

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



- 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

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

Calculate the molarity of a solution prepared by dissolving 45.00 grams of KI into a total volume of 500.0 mL.

Calculate the molarity of a solution prepared by dissolving 45.00 grams of KI into a total volume of 500.0 mL.

$$\frac{45.00 \text{ g KI}}{500.0 \text{ mL}} \times \frac{1 \text{ mol KI}}{166.0 \text{ g KI}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.5422 \text{ M}$$

How many milliliters of 3.50 *M* NaOH can be prepared from 75.00 grams of the solid?

How many milliliters of 3.50 *M* NaOH can be prepared from 75.00 grams of the solid?

$$75.00 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} \times \frac{1 \text{ L}}{3.50 \text{ mol NaOH}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 536 \text{ mL}$$

- Dilution
  - Process of preparing a less concentrated solution from a more concentrated one.

moles of solute before dilution = moles of solute after dilution

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

$$M_c \times L_c = M_d \times L_d$$

For the next experiment the class will need 250. mL of 0.10 *M* CuCl<sub>2</sub>. There is a bottle of 2.0 *M* CuCl<sub>2</sub>. Describe how to prepare this solution. How much of the 2.0 *M* solution do we need?

Concentrated: 2.0 *M* use ? mL ( $L_d$ )

Dilute: 250. mL of 0.10 *M*

$$M_c L_c = M_d L_d$$

$$(2.0 \text{ M}) (L_c) = (0.10 \text{ M}) (250. \text{ mL})$$

$$L_c = 12.5 \text{ mL}$$

12.5 mL of the concentrated solution are needed; add enough distilled water to prepare 250. mL of the solution.

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

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

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

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

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



$$[\text{Ba}^{2+}] = 0.42 \text{ M (1:1 ratio)}$$

$$[\text{OH}^-] = 0.84 \text{ M (2:1 ratio)}$$



$$[\text{NH}_4^+] = 1.2 \text{ M (1:1 ratio)}$$

$$[\text{Cl}^-] = 1.2 \text{ M (1:1 ratio)}$$

# 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 %)

THE  
MIDDLE

THE  
MIDDLE

THE  
MIDDLE

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

# Apparatus for a Titration



A student measured exactly 15.0 mL of an unknown monoprotic acidic solution and placed it 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.

$$\begin{array}{r}
 \text{M}_1 V_1 = \text{M}_2 V_2 \\
 M \times 15.0 \text{ mL} = 0.12 \text{ M} \times 35.0 \text{ mL} \\
 M = \frac{0.12 \text{ M} \times 35.0 \text{ mL}}{15.0 \text{ mL}} \\
 M = 0.28 \text{ M}
 \end{array}$$

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.

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

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

# Key Points

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

# Key Points

- Balancing redox reactions by the half reaction method
- Various types: decomposition, combination
- Molarity
- Solution stoichiometry
  - Gravimetric analysis
  - Volumetric analysis