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

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

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

$$C_{12}H_{22}O_{11}(s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}(aq)$$

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

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

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

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

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

- 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

Acid	Ionization Equation
Hydrochloric acid	$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$
Hydrobromic acid	$HBr(aq) \longrightarrow H^+(aq) + Br^-(aq)$
Hydroiodic acid	$HI(aq) \longrightarrow H^+(aq) + I^-(aq)$
Nitric acid	$HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$
Chloric acid	$HClO_3(aq) \longrightarrow H^+(aq) + ClO_3^-(aq)$
Perchloric acid	$HClO_4(aq) \longrightarrow H^+(aq) + ClO_4^-(aq)$
Sulfuric acid*	$H_2SO_4(aq) \longrightarrow H^+(aq) + HSO_4^-(aq)$
	$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$

*Note that although each sulfuric acid molecule has two ionizable hydrogen atoms, it only undergoes the first ionization completely, effectively producing one H^+ ion and one HSO_4^- ion per H_2SO_4 molecule. The second ionization happens only to a very small extent.

- Examples of weak electrolytes

  - **Weak acids**

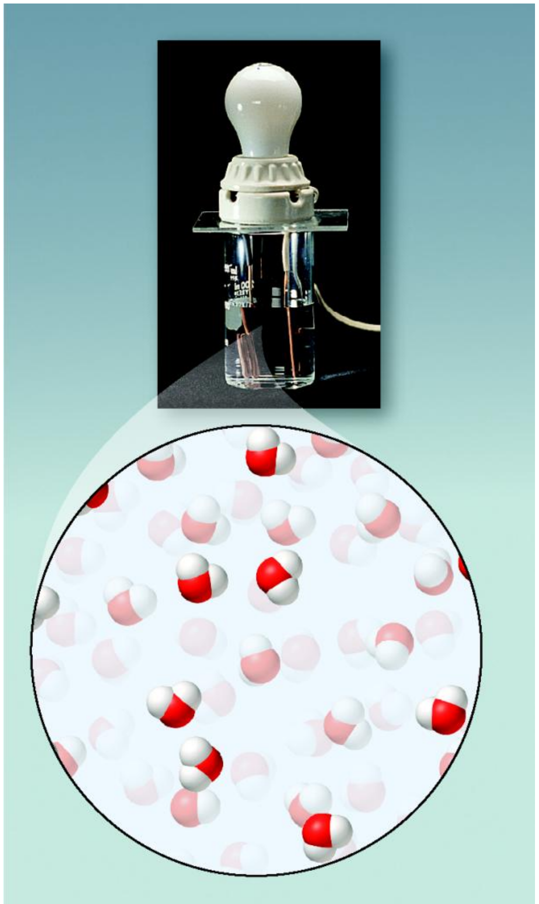
    $$HC_2H_3O_{2(aq)} \rightleftharpoons C_2H_3O_2^-{}_{(aq)} + H^+{}_{(aq)}$$

  - **Weak bases**

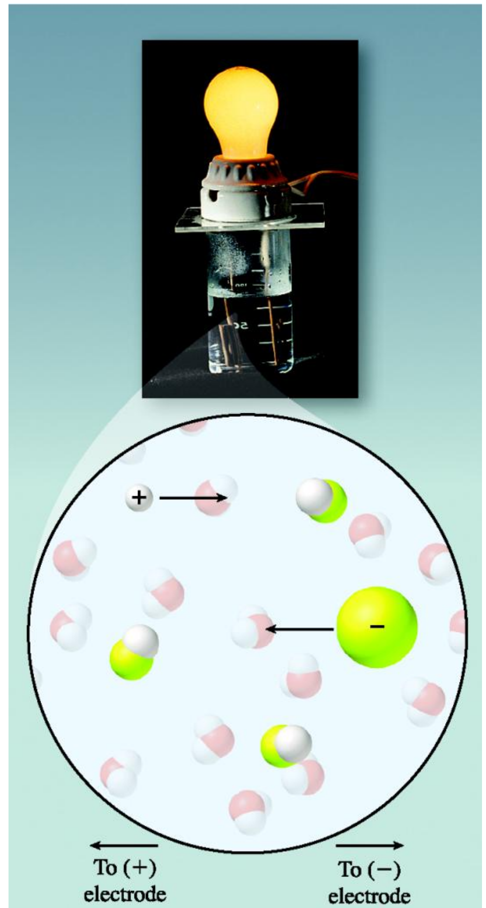
    $$NH_{3\ (aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

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

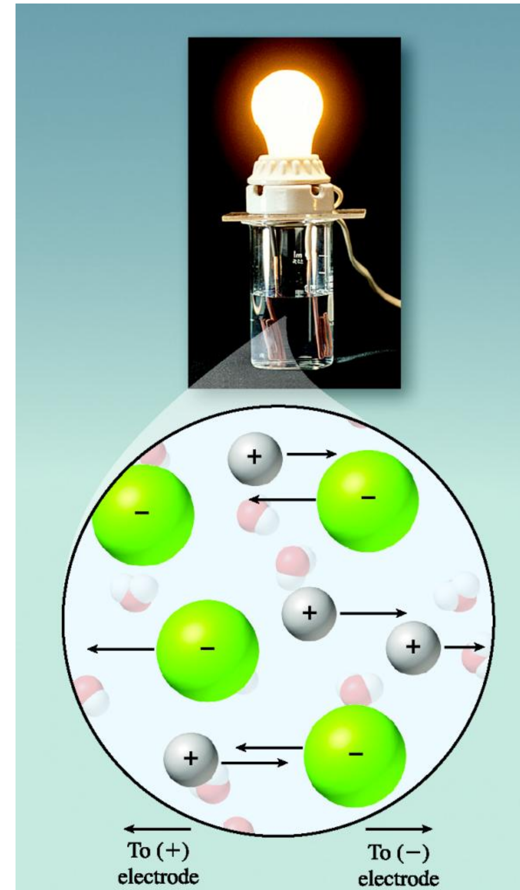
Method to Distinguish Types of Electrolytes



nonelectrolyte



weak electrolyte



strong electrolyte

Classify the following as nonelectrolyte, weak electrolyte or strong electrolyte

- NaOH

  strong electrolyte

- CH_3OH

  nonelectrolyte

- H_2CO_3

  weak electrolyte

4.2 Precipitation Reactions

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

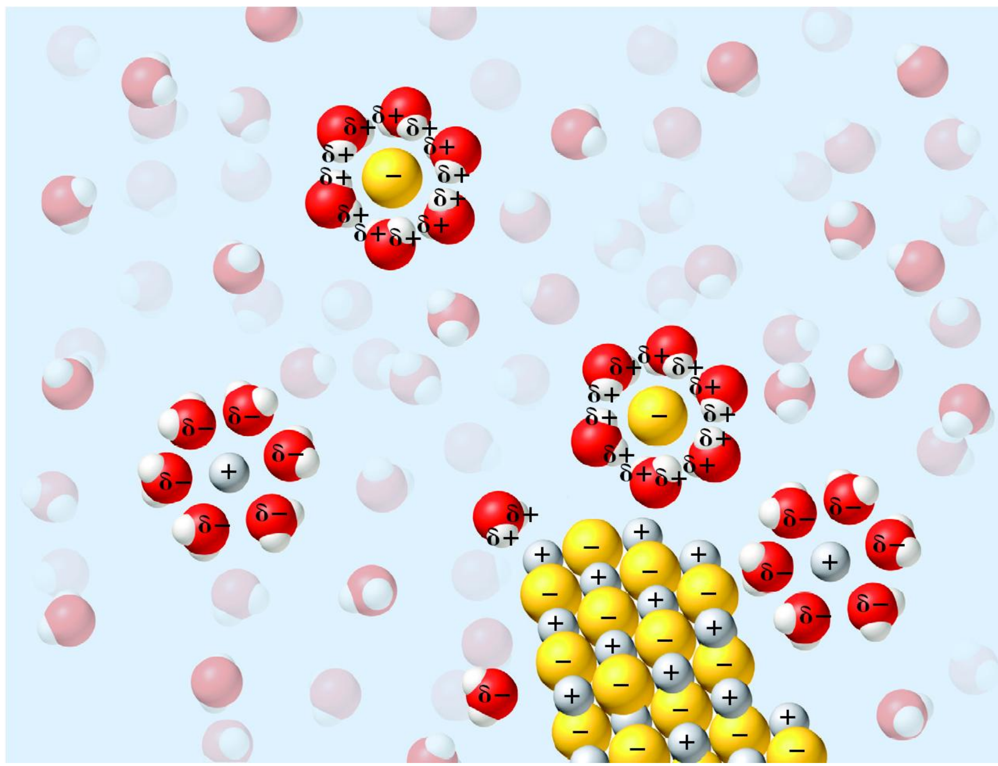
TABLE 4.2 Solubility Guidelines: Soluble Compounds

Water-Soluble Compounds	Insoluble Exceptions
Compounds containing an alkali metal cation (Li^+, Na^+, K^+, Rb^+, Cs^+) or the ammonium ion (NH_4^+)	
Compounds containing the nitrate ion (NO_3^-), acetate ion ($C_2H_3O_2^-$), or chlorate ion (ClO_3^-)	
Compounds containing the chloride ion (Cl^-), bromide ion (Br^-), or iodide ion (I^-)	Compounds containing Ag^+, Hg_2^{2+}, or Pb^{2+}
Compounds containing the sulfate ion (SO_4^{2-})	Compounds containing Ag^+, Hg_2^{2+}, Pb^{2+}, Ca^{2+}, Sr^{2+}, or Ba^{2+}

TABLE 4.3 Solubility Guidelines: Insoluble Compounds

Water-Insoluble Compounds	Soluble Exceptions
Compounds containing the carbonate ion (CO_3^{2-}), phosphate ion (PO_4^{3-}), chromate ion (CrO_4^{2-}), or sulfide ion (S^{2-})	Compounds containing Li^+, Na^+, K^+, Rb^+, Cs^+, or NH_4^+
Compounds containing the hydroxide ion (OH^-)	Compounds containing Li^+, Na^+, K^+, Rb^+, Cs^+, or Ba^{2+}

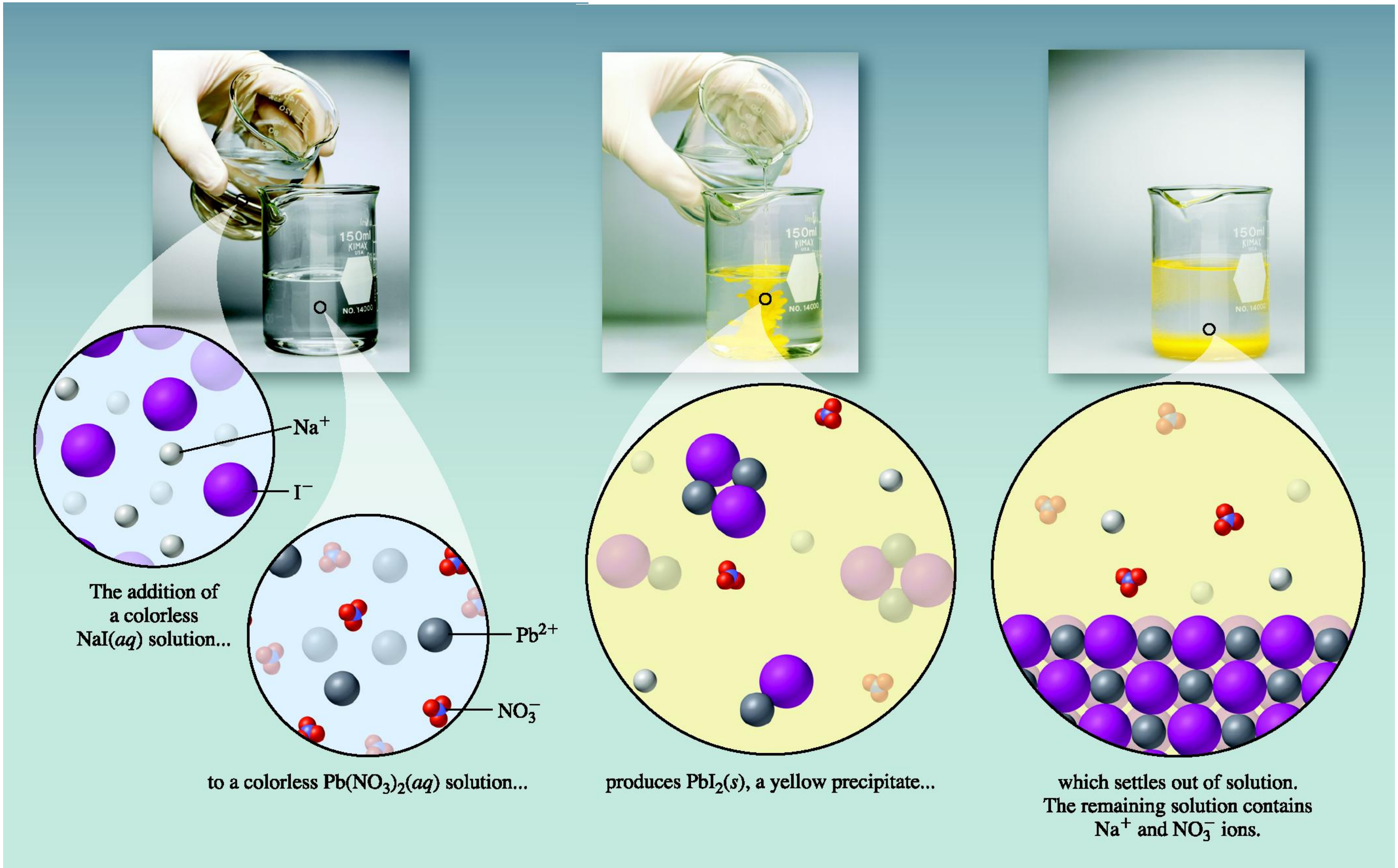
- ***<u>Hydration:</u>*** process by which water molecules remove and surround individual ions from the solid.



Identify the Precipitate

$Pb(NO_3)_2(aq) + 2NaI(aq) \rightarrow 2NaNO_3(?)(aq) + PbI_2(?)(s)$

Mixing Solutions of $Pb(NO_3)_2$ and NaI



Classify the following as soluble or insoluble in water

- $Ba(NO_3)_2$

  soluble

- AgI

  insoluble

- $Mg(OH)_2$

  insoluble

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

$$Na_2SO_4(aq) + Ba(OH)_2(aq) \longrightarrow 2NaOH(aq) + BaSO_4(s)$$

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

$$Na_2SO_4(aq) \longrightarrow 2Na^+(aq) + SO_4^{2-}(aq)$$

$$Ba(OH)_2(aq) \longrightarrow Ba^{2+}(aq) + 2OH^-(aq)$$

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

Molecular equation:

$$Na_2SO_4(aq) + Ba(OH)_2(aq) \longrightarrow 2NaOH(aq) + BaSO_4(s)$$

Ionic equation:

$$2Na^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2OH^-(aq) \longrightarrow 2Na^+(aq) + 2OH^-(aq) + BaSO_4(s)$$

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

$$\cancel{2Na^+(aq)} + SO_4^{2-}(aq) + Ba^{2+}(aq) + \cancel{2OH^-(aq)} \longrightarrow \cancel{2Na^+(aq)} + \cancel{2OH^-(aq)} + BaSO_4(s)$$

Net ionic equation:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$$

- **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:

$$2AgNO_3(aq)+Na_2SO_4(aq) \rightarrow 2NaNO_3(?)+Ag_2SO_4(?)$$

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

$$2Ag^{+}(aq) + 2NO_3^{-}(aq) + 2Na^{+}(aq) + SO_4^{2-}(aq)$$

$$\rightarrow 2Na^{+}(aq) + 2NO_3^{-}(aq) + Ag_2SO_4(s)$$

Step 3: Cancel spectators

$$2Ag^+(aq) + 2NO_3^-(aq) + 2Na^+(aq) + SO_4^{2-}(aq)$$

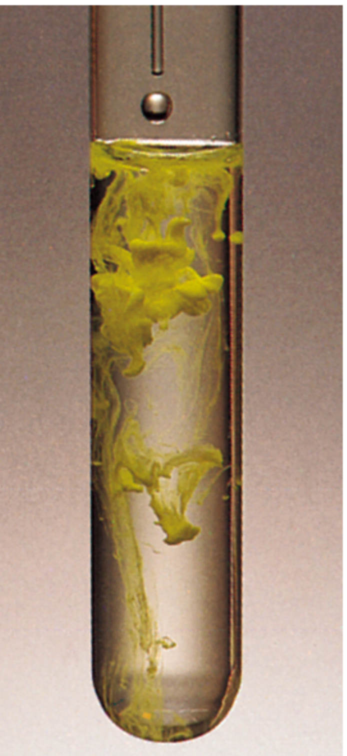
$$\rightarrow 2Na^+(aq) + 2NO_3^-(aq) + Ag_2SO_4(s)$$

Step 4: Write the net ionic reaction

$$2Ag^+(aq) + SO_4^{2-}(aq) \rightarrow Ag_2SO_4(s)$$

Precipitation Reactions

Precipitate: insoluble solid that separates from solution



PbI_2

precipitate ↓

$$Pb(NO_3)_2\,(aq) + 2NaI\,(aq) \longrightarrow PbI_2\,(s) + 2NaNO_3\,(aq)$$

molecular equation

$$Pb^{2+} + \cancel{2NO_3^-} + \cancel{2Na^+} + 2I^- \longrightarrow PbI_2\,(s) + \cancel{2Na^+} + \cancel{2NO_3^-}$$

ionic equation

$$Pb^{2+} + 2I^- \longrightarrow PbI_2\,(s)$$

net ionic equation

Na^+ and NO_3^- are ***spectator*** ions

4.3 Acid-Base Reactions

- Termed neutralization reactions.
- acid + base $\rightarrow$ salt + water
- 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

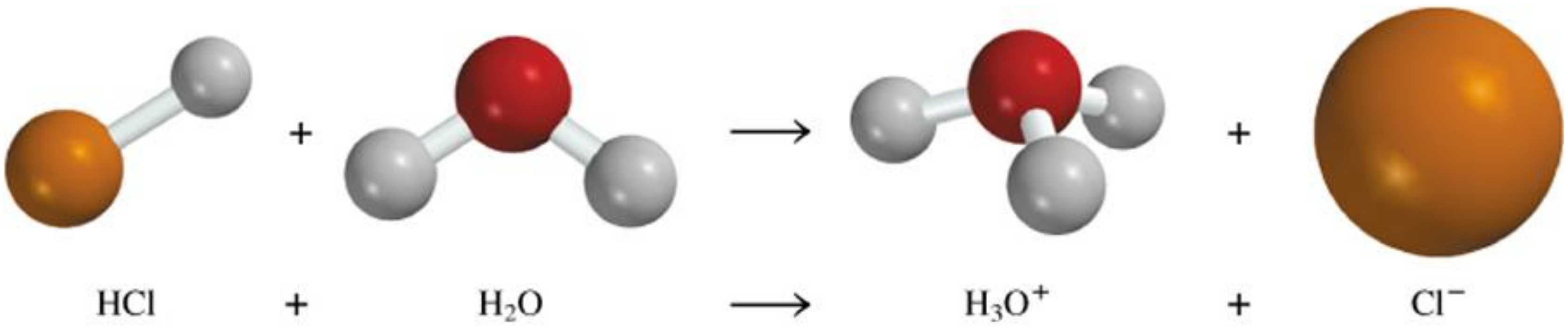
Strong Acids	Strong Bases	Strong Acids	Strong Bases
HCl	$LiOH$	$HClO_3$	$CsOH$
HBr	$NaOH$	$HClO_4$	$Ca(OH)_2$
HI	KOH	H_2SO_4	$Sr(OH)_2$
HNO_3	$RbOH$		$Ba(OH)_2$

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

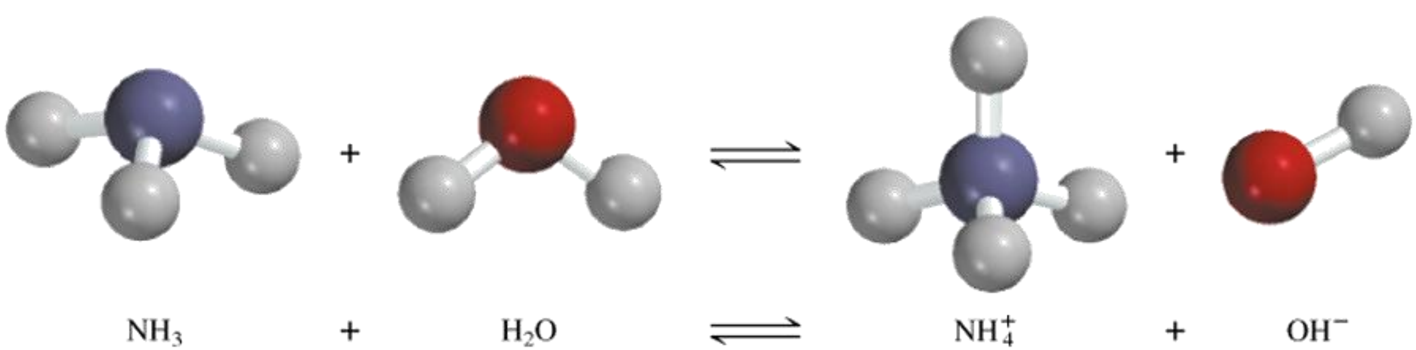
- Definitions of acids and bases
  - ***Arrhenius acid* -** produces H^+ in solution
  - **Arrhenius *base* -** produces OH^- in solution
  - More inclusive definitions:
    - ***Brønsted acid* -** proton donor
    - ***Brønsted base* -** proton acceptor

Acid and Base

Arrhenius Acid is a substance that produces H^+ (H_3O^+) in water



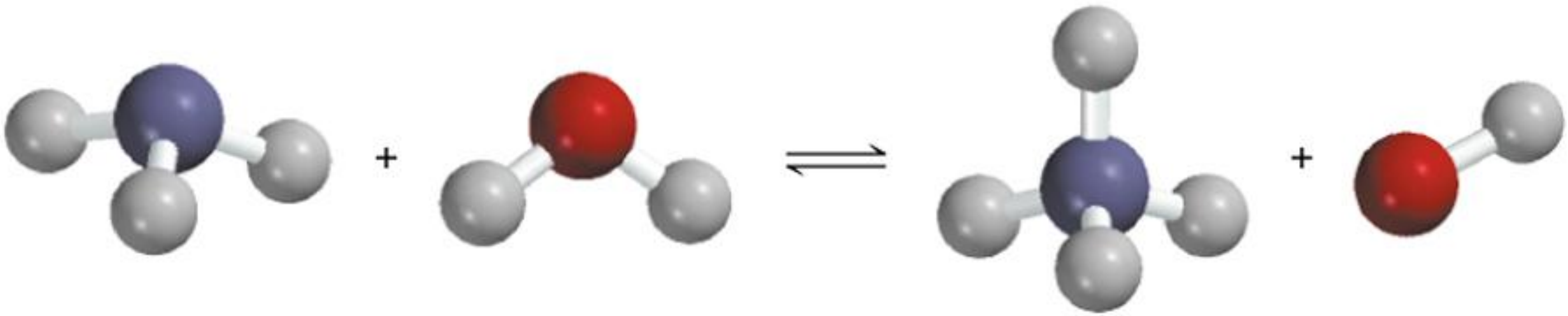
Arrhenius Base is a substance that produces OH^- in water



Continue Acid Base

Brønsted acid is a proton donor

Brønsted base is a proton acceptorA

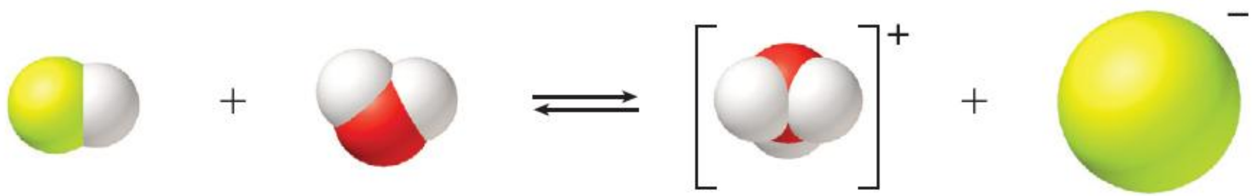


$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

base acid acid base

– Examples of a weak base and weak acid

- Hydrofluoric acid with water:



$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$$

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

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

  - ***Diprotic***: two ionizable hydrogens

$$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$$
$$HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$$

– ***Triprotic:*** three ionizable hydrogens

$$H_3PO_4 + H_2O \rightarrow H_3O^+ + H_2PO_4^-$$
$$H_2PO_4^- + H_2O \rightarrow H_3O^+ + HPO_4^{2-}$$
$$HPO_4^{2-} + H_2O \rightarrow H_3O^+ + PO_4^{3-}$$

– Polyprotic: generic term meaning more than one ionizable hydrogen

- **Types of bases**
  - Monobasic: One OH^- group

$$KOH \rightarrow K^+ + OH^-$$

  - Dibasic: Two OH^- groups

$$Ba(OH)_2 \rightarrow Ba^{2+} + 2OH^-$$

Acid-Base Neutralization

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

Acid + Base → Salt + Water

Molecular equation:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Ionic equation:

$$H^+(aq) + \cancel{Cl^-}(aq) + \cancel{Na^+}(aq) + OH^-(aq) \rightarrow \cancel{Na^+}(aq) + \cancel{Cl^-}(aq) + H_2O(l)$$

Net ionic equation:

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

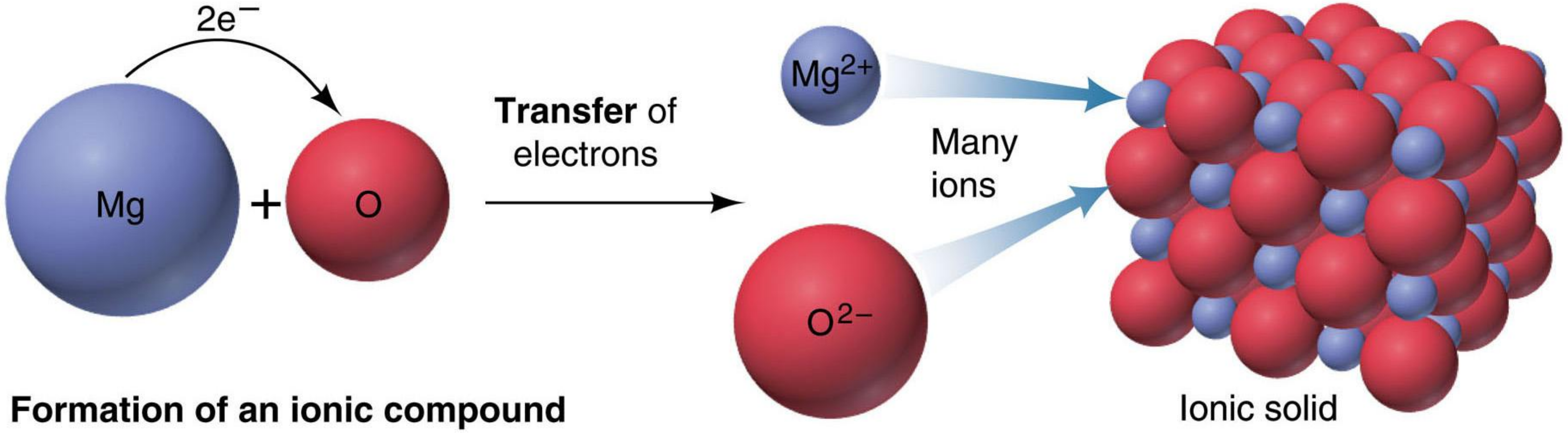
Solutions of acetic acid and lithium hydroxide are mixed. Write the net ionic reaction.

$$HC_2H_3O_2\,(aq) \;+\; OH^-(aq) \longrightarrow C_2H_3O_2^-(aq) \;+\; H_2O(l)$$

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

Oxidation-Reduction Reaction



A Formation of an ionic compound

electron transfer reactions)

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$$

Oxidation-Reduction Reaction

$$2Mg\ (s) + O_2\ (g) \longrightarrow 2MgO\ (s)$$

$$2Mg \longrightarrow 2Mg^{2+} + 4e^-$$

Oxidation half-reaction (lose e^-)

$$O_2 + 4e^- \longrightarrow 2O^{2-}$$

Reduction half-reaction (gain e^-)

– ***Example***

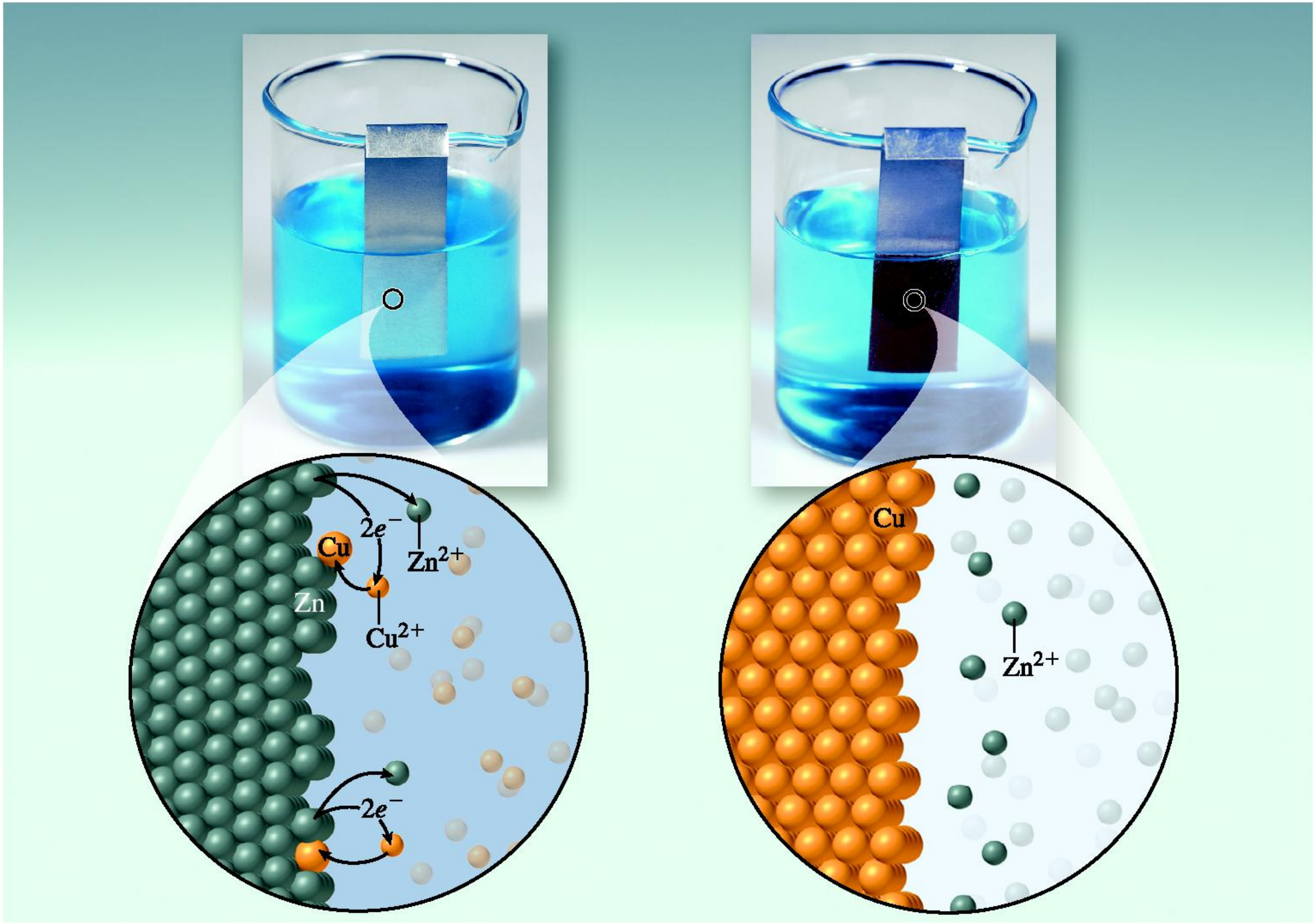
$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

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

– **Zinc is losing 2 electrons** and oxidized.
  - Reducing agent
  - $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

– **Copper ions are gaining the 2** electrons.
  - Oxidizing agent
  - $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Reaction of Cu and Zn^{2+} ions



Assigning The Oxidation State

1. Oxidation state of an atom in an element = **0**
   eg O_2, Cl_2, H_2, Na,…..etc
2. Oxidation state of monatomic element ions = **charge**
   e.g. Cl^-, Na^+, S^{-2},…..etc
3. Oxygen = **-2 in covalent compounds** (except in **<u>peroxides</u>** where it = **-1**)
4. H = **+1** in covalent compounds
5. Fluorine = **-1** in compounds

What is the oxidation numbers of the elements in the IF_7 ?

IF_7

F = -1

7x(-1) + ? = 0

I = +7

Oxidation Number

1. For Group 1A(1): O.N. = +1 in all compounds
2. For Group 2A(2): O.N. = +2 in all compounds
3. For Group 7A(17): O.N. = -1 in combination with metals, nonmetals (except O), and other halogens lower in the group

What is the oxidation number of the each element in the following Compounds?

$ZnCl_2$ for zinc is +2 and that for chloride is -1

SO_3 Each oxygen is an oxide with -2. sulfur is +6

HNO_3 H is +1 and each oxygen is -2. the N is +5

Assign oxidation numbers for all elements in each species

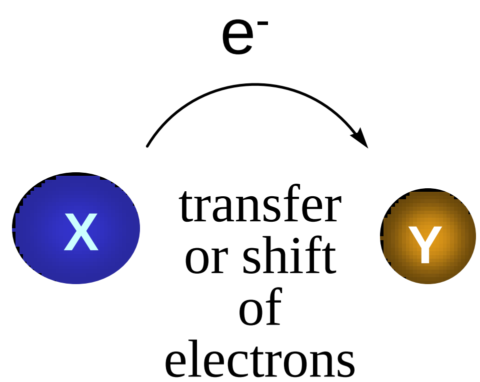
$MgBr_2$

Mg +2, Br –1

ClO_2^-

Cl +3 , O –2

Electron Transfer Terminology



X loses electron(s)

X is oxidized

X is the reducing agent

X increases its oxidation number

Y gains electron(s)

Y is reduced

Y is the oxidizing agent

Y decreases its oxidation number

Oxidation Reduction Reaction

Identify the ***oxidizing agent*** and ***reducing agent*** in each of the following

0 +1 +6 -2 +3 6 -2 0

a) $2Al(s) + 3H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

The oxidation number of Al increases; it is oxidized; it is the reducing agent.

The oxidation number of H decreases; it is reduced; H_2SO_4 is the oxidizing agent.

Continue

2 -2 2 -2 0 +4 -2

(b) $PbO(s) + CO(g) \longrightarrow Pb(s) + CO_2(g)$

The oxidation number of C increases; it is oxidized; CO is the reducing agent.

The oxidation number of Pb decreases; it is reduced; PbO is the oxidizing agent.

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

$$Mg(s) + CuCl_2(aq) \rightarrow Cu(s) + MgCl_2(aq)$$

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

TABLE 4.6 Activity Series

Increasing ease of oxidation (↑)	Element	Oxidation Half-Reaction
	Lithium	$Li \longrightarrow Li^{+} + e^{-}$
	Potassium	$K \longrightarrow K^{+} + e^{-}$
	Barium	$Ba \longrightarrow Ba^{2+} + 2e^{-}$
	Calcium	$Ca \longrightarrow Ca^{2+} + 2e^{-}$
	Sodium	$Na \longrightarrow Na^{+} + e^{-}$
	Magnesium	$Mg \longrightarrow Mg^{2+} + 2e^{-}$
	Aluminum	$Al \longrightarrow Al^{3+} + 3e^{-}$
	Manganese	$Mn \longrightarrow Mn^{2+} + 2e^{-}$
	Zinc	$Zn \longrightarrow Zn^{2+} + 2e^{-}$
	Chromium	$Cr \longrightarrow Cr^{3+} + 3e^{-}$
	Iron	$Fe \longrightarrow Fe^{2+} + 2e^{-}$
	Cadmium	$Cd \longrightarrow Cd^{2+} + 2e^{-}$
	Cobalt	$Co \longrightarrow Co^{2+} + 2e^{-}$
	Nickel	$Ni \longrightarrow Ni^{2+} + 2e^{-}$
	Tin	$Sn \longrightarrow Sn^{2+} + 2e^{-}$
	Lead	$Pb \longrightarrow Pb^{2+} + 2e^{-}$
	Hydrogen	$H_2 \longrightarrow 2H^{+} + 2e^{-}$
	Copper	$Cu \longrightarrow Cu^{2+} + 2e^{-}$
	Silver	$Ag \longrightarrow Ag^{+} + e^{-}$
	Mercury	$Hg \longrightarrow Hg^{2+} + 2e^{-}$
	Platinum	$Pt \longrightarrow Pt^{2+} + 2e^{-}$
	Gold	$Au \longrightarrow Au^{3+} + 3e^{-}$

- **Balancing redox reactions**
  - Electrons (charge) must be balanced as well as number and types of atoms
  - Consider this net ionic reaction:

$$Al(s) + Ni^{2+}(aq) \rightarrow Al^{3+}(aq) + Ni(s)$$

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

$$Al(s) + Ni^{2+}(aq) \rightarrow Al^{3+}(aq) + Ni(s)$$

- Divide reaction into two half-reactions

$$Al(s) \rightarrow Al^{3+}(aq) + 3e^{-}$$

$$Ni^{2+}(aq) + 2e^{-} \rightarrow Ni(s)$$

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

$$2\,[Al(s) \rightarrow Al^{3+}(aq) + 3e^{-}]$$

$$3\,[Ni^{2+}(aq) + 2e^{-} \rightarrow Ni(s)]$$

– Cancel electrons and write balanced net ionic reaction

$$2Al(s) \rightarrow 2Al^{3+}(aq) + \cancel{6e^{-}}$$

$$3Ni^{2+}(aq) + \cancel{6e^{-}} \rightarrow 3Ni(s)$$

$$2Al(s) + 3Ni^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Ni(s)$$

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

$$Cu(s) + 2\,Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2\,Ag(s)$$

- 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

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Identify the substance oxidized and the substance reduced.

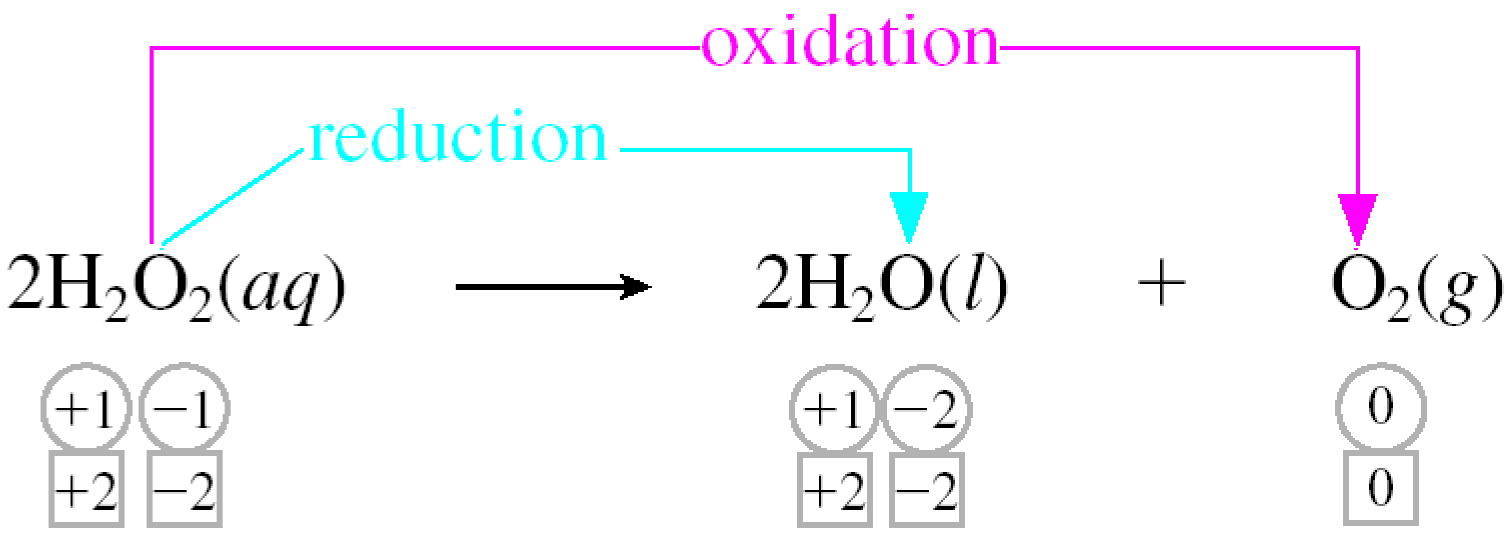
- **Decomposition reactions**
  - Many decomposition reactions may also be classified as redox reactions
  - Consider:

Potassium chlorate is strongly heated

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

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:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

$CH_4(g)$		$2O_2(g)$	$CO_2(g)$		$2H_2O(l)$	
−4	+1	0	+4	−2	+1	−2
−4	+4	0	+4	−4	+2	−2

Oxidation Numbers on the Periodic Table

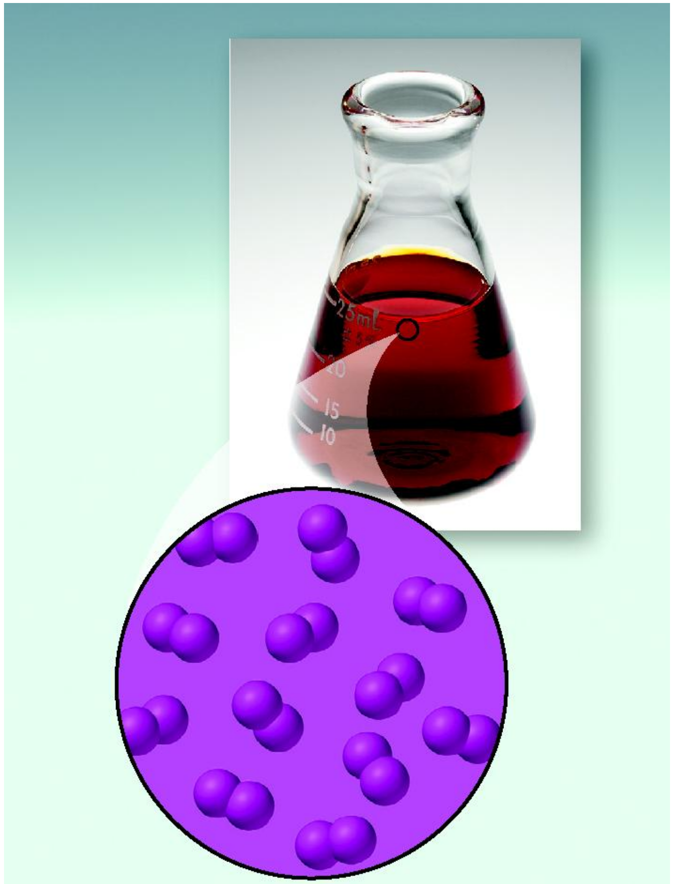
(most common in red)

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

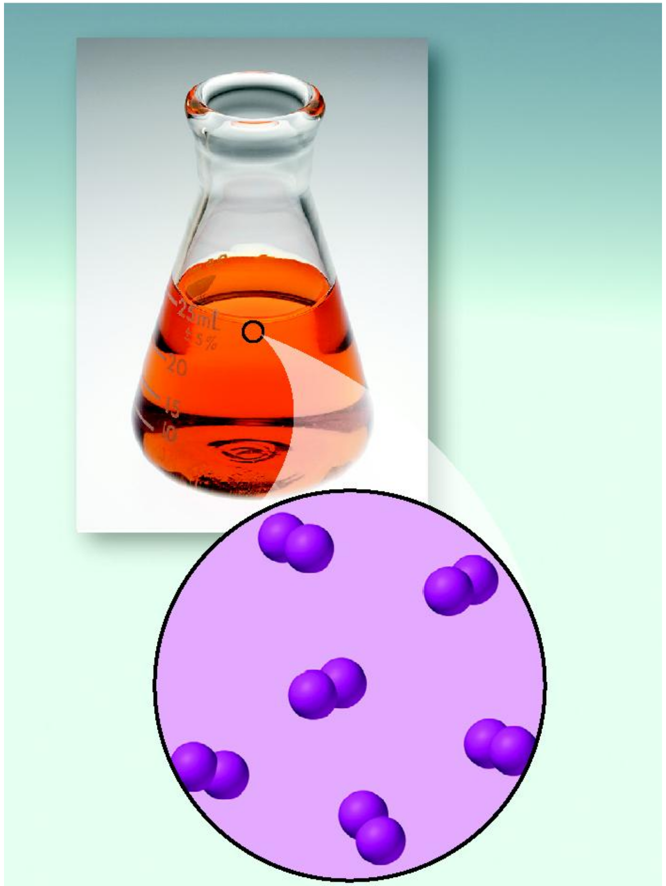
4.5 Concentration of Solutions

- ***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



Concentrated solution:
More solute particles per unit volume



Dilute solution:
Fewer solute particles per unit volume

- 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

$$M = \frac{mol}{\mathrm{L}} \qquad \mathrm{L} = \frac{mol}{M} \qquad mol = M \times \mathrm{L}$$

Example

Calculate the molarity of a solution prepared by dissolving 11.5 g of solid NaOH in 1.5 L of water

$$\text{Moles of NaOH} = \frac{\text{Mass (g)}}{\text{Molar mass (g/mol)}} = \frac{11.5\text{ g}}{40\text{ g/mol}} = 0.287\text{ mole}$$

$$\boldsymbol{M} = \boldsymbol{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{0.287\text{ mole}}{1.5\text{ L}}$$

0.19 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?

$$\text{Moles of NaOH} = \frac{75.00\ \text{g}}{40\ \text{g/mole}} = 1.88\ \text{mole}$$

$$\text{Volume of NaOH (L)} = \frac{\text{Moles}}{\text{Molarity}} = \frac{1.88\ \text{m}}{3.5\ \text{M}} = 0.537\ \text{L}$$

= 537 ml

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

moles of solute before dilution = moles of solute after dilution

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

For the next experiment the class will need 250. mL of 0.10 *M* $CuCl_2$. There is a bottle of 2.0 *M* $CuCl_2$. 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

$$M_c L_c = M_d L_d$$

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

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

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 Na^+ and $Cl^–$

Na_2SO_4 dissociates into $2Na^+$ and $SO_4^{2–}$

$AlCl_3$ dissociates into Al^{3+} and $3Cl^–$

Find the concentration of all species in a 0.25 *M* solution of $MgCl_2$

$$MgCl_2 \rightarrow Mg^{2+} + 2Cl^-$$

Given: $MgCl_2$ = 0.25 *M*

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

$[Cl^-]$ = 0.50 *M* (1:2 ratio)

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

0.42 *M* $Ba(OH)_2$

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

$[OH^-]$ = 0.84 *M* (2:1 ratio)

1.2 *M* NH_4Cl

$[NH_4^+]$ = 1.2 *M* (1:1 ratio)

$[Cl^-]$ = 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?

$$Ag^+ + Cl^- \rightarrow AgCl_{(s)}$$

$$\text{Moles of AgCl} = \frac{1.725\ \text{g}}{143.32\ \text{g/mol}} = 0.012\ \text{mole} = \text{mole of } Cl^-$$

$$\text{Mass of } Cl^- = 0.012\ \text{mol} \times 35.5\ \text{g/mol} = 0.43\ \text{g of } Cl^-$$

$$\%Cl- = \frac{0.43\ \text{g}}{0.825\ \text{g}} \times 100\% = 52\ \%$$

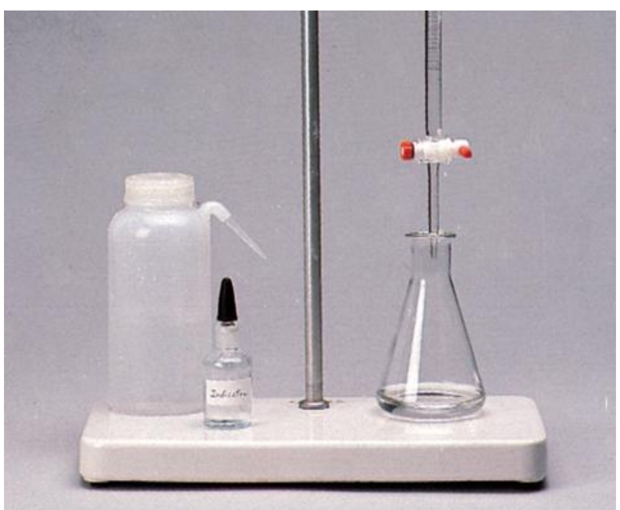
- **Volumetric analysis**
  - Commonly accomplished by ***titration***
    - Addition of a solution of known concentration (standard solution) to another solution of unknown concentration.
  - ***Standardization*** *is* the determination of the exact concentration of a solution.
  - ***Equivalence point*** represents completion of the reaction.
  - ***Endpoint*** is where the titration is stopped.
  - An ***indicator*** is used to signal the endpoint.

Acid Base Titration

Titration is analytical technique in which one can calculate the concentration of a solute in a solution.

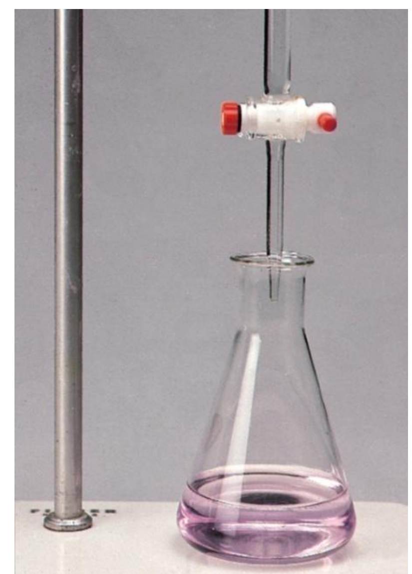
Equivalent point – the point at which the reaction is complete

Indicator – substance that changes color at (or near) the equivalent point **End Point**



Slowly add base to unknown acid UNTIL

the indicator changes color



Titration

Example

50 ml of HCl solution has been titrated with 0.1524 M NaOH solution. At the end point, 33.32 ml of NaOH was used in the titration. What is the concentration of the HCl solution?

WRITE THE BALANCED CHEMICAL EQUATION!

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

At the End Point

Moles of NaOH = Moles of HCl

$M1 \times V1 = M2 \times V2$

$0.1524 \times 0.03332 = M2 \times 0.050$

M2 = 0.1016 Molar

Example

What volume of a 1.420 *M* NaOH solution is Required to titrate 25.00 mL of a 4.50 *M* H_2SO_4 solution?

WRITE THE BALANCED CHEMICAL EQUATION!

$$H_2SO_4 + 2NaOH \longrightarrow 2H_2O + Na_2SO_4$$

At the equivalent Point

Moles of NaOH = 2 x Moles of H_2SO_4

M1 x V1 = 2 x M2 x V2

1.42 x V1 = 2 x 4.5 x 25.00

V1 = 158 ml

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