Chapter 4: The Major Classes of Chemical Reactions

Ionic Compounds are Strong Electrolytes

- **Electrolyte**
  - A substance that conducts a current when dissolved in water.
- **Strong Electrolytes**
  - Soluble ionic compound dissociate completely
  - may conduct a large current
  - Animation

Role of Water as a Solvent

- Why do *some* aqueous solutions conduct electricity and others do not?
- **Dissociation of Ionic Compounds**
  - Ionic compounds ***dissociate into ions*** when dissolved in water

    $NaCl_{(s)} + H_2O_{(l)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)}$

  - Resulting solution is called an electrolyte
    - Electrolytes conduct electricity.....Why?

The Dissolution of an Ionic Compound

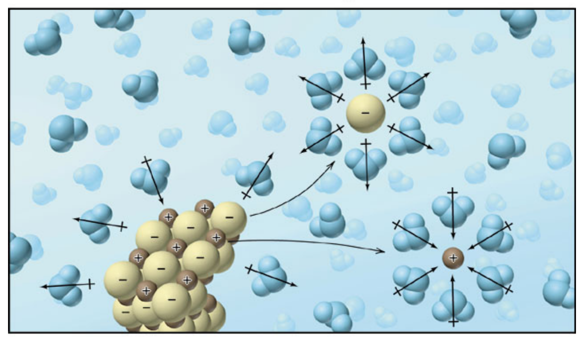


Fig. 4.3

Electron Distribution in Molecules of H_2 and H_2O

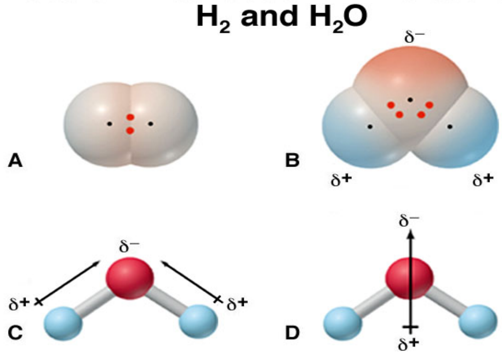


Fig. 4.2

Soluble vs. Insoluble Ionic Compounds

What determines the solubility of an Ionic Compound?

- Solubility of NaCl in water at 20°C = 365 g/L
- Solubility of $MgCl_2$ in water at 20°C = 542.5 g/L
- Solubility of $AlCl_3$ in water at 20°C = 699 g/L
- Solubility of $PbCl_2$ in water at 20°C = 9.9 g/L
- Solubility of AgCl in water at 20°C = 0.009 g/L
- Solubility of CuCl in water at 20°C = 0.0062 g/L

The Electrical Conductivity of Ionic Solutions

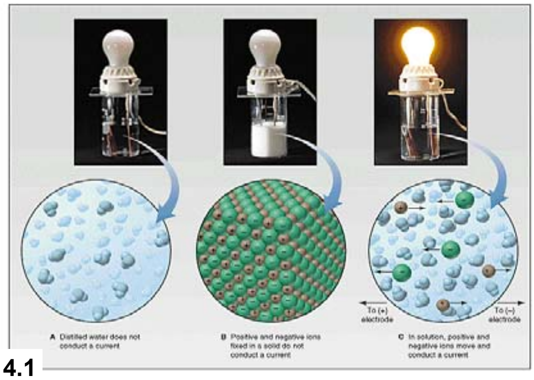


Fig. 4.1

Nonelectrolytes

- Their solutions do *not* conduct electricity.......Why?
- **Only neutral molecules present**
- ***Most*** **<u>molecular (covalent) substances</u>** produce ***neutral*** **molecules** in solution
  - e.g. Sucrose, glucose, methanol, ethanol....
- Many polar covalent molecules ***<u>ionize</u>*** *in* solution
  - E.g. HCl $_{(g)}$, Organic acids: e.g. CH_3COOH

Solubility of Covalent Compounds in Water

<u>Covalent compounds that are soluble in water</u>

- **Have polar hydroxyl group: -OH**
- **Forms strong electrostatic interactions with water**

<u>Examples</u>

- table sugar, sucrose: $C_{12}H_{22}O_{11}$
- Ethanol: CH_3CH_2-OH
- Ethylene glycol (antifreeze): $C_2H_6O_2$
- Methanol: CH_3-OH

Solubility of Covalent Compounds in Water

<u>Covalent compounds that are insoluble in water</u>

- Do not contain a polar center
- Have little or no interactions with water molecules

<u>Examples</u>

- Hydrocarbons in gasoline and oil

  Octane = C_8H_{18}
  Benzene = C_6H_6

- Oil spills: oil will not mix with the water and forms a layer on the surface!

Equations for the Dissociation of Ionic Compounds

- **Knowledge of the common polyatomic ions is a must**:
  - ✓e.g. nitrate, sulfate, phosphate, acetate, carbonate, hydroxide, Ammonium
- Rusty?
  - ✓Review ionic compounds (Chapter 2)

Write the equation for the dissociation of the following compounds in water

- Aluminum Chloride, $AlCl_3$
- Ammonium Sulfate, $(NH_4)_2SO_4$
- Ammonium Hydroxide, NH_4OH

Determining Moles of Ions in Aqueous Solutions of Ionic Compounds

Problem: How many moles of each ion are in each of the following:

a) **4.0 moles of sodium carbonate dissolved in water**

$$Na_2CO_{3\,(s)} \xrightarrow{H_2O} 2\,Na^+_{(aq)} + CO_3^{-2}{}_{(aq)}$$

b) 81.1 g of Iron (III) Chloride dissolved in water

$$FeCl_{3\,(s)} \xrightarrow{H_2O} Fe^{+3}_{(aq)} + 3\,Cl^-_{(aq)}$$

($FeCl_3$ = 162.2 g/mol)

Metathesis Reactions

(Double displacement or double replacement reactions)

AB (aq) + **CD** (aq) → **CB** + **AD**

Only occur if one of the following form

- **Precipitate (ppt)**
- **Gas**
- **Weak electrolyte** (e.g. acid-base reactions to form H_2O)

Double-Displacement Reactions

1) **Precipitation reactions - an insoluble product is formed**

$$Pb(NO_3)_{2\,(aq)} + 2\,NaI_{(aq)} \rightarrow PbI_{2\,(s)} + 2\,NaNO_{3\,(aq)}$$

$$Ba(NO_3)_{2\,(aq)} + Na_2SO_{4\,(aq)} \rightarrow BaSO_{4\,(s)} + 2\,NaNO_{3\,(aq)}$$

2) **Acid-base neutralization reactions - water is formed**

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + HOH_{(l)}$$

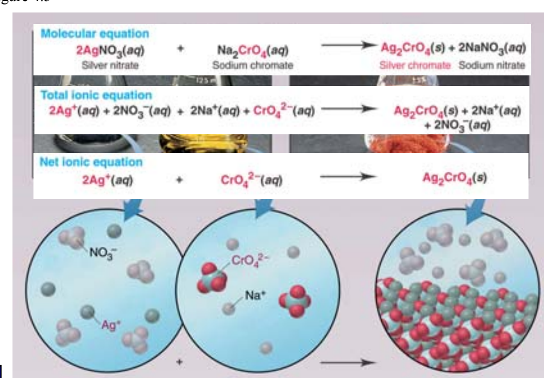
$$H_2SO_{4\,(aq)} + Ca(OH)_{2\,(s)} \rightarrow CaSO_{4\,(s)} + 2\,H_2O_{(l)}$$

3) **A carbonate or sulfite reacts with acid to form a gas**

$$Na_2CO_{3\,(aq)} + 2\,HBr_{(aq)} \rightarrow 2\,NaBr_{(aq)} + H_2O_{(l)} + CO_{2\,(g)}$$

$$K_2SO_{3\,(aq)} + 2\,HI_{(aq)} \rightarrow 2\,KI_{(aq)} + H_2O_{(l)} + SO_{2\,(g)}$$

Figure 4.5 **A precipitation reaction and its equation**



Reactions between Aqueous Ionic Compounds

- Predict what will happen if the following solutions are mixed:

$$Pb(NO_3)_{(aq)} + NaI_{(aq)} \rightarrow$$

The Reaction of $Pb(NO_3)_2$ and NaI

Fig. 4.5

Precipitation Reactions: A Solid Product is Formed

$Pb(NO_3)_{(aq)} + NaI_{(aq)} \rightarrow Pb^{+2}_{(aq)} + 2\ NO_3^-{}_{(aq)} + Na^+_{(aq)} + I^-_{(aq)}$

Vs.

$Pb(NO_3)_{2\,(aq)} + 2\ NaI_{(aq)} \rightarrow PbI_{2\,(s)} + 2\ NaNO_{3\,(aq)}$

- Why does a precipitate of **PbI_2** form?

Precipitation Reactions: Will a Precipitate Form?

$KCl_{(aq)} + NH_4NO_{3\,(aq)} = K^+_{(aq)} + Cl^-_{(aq)} + NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$

- **Will a ppt. Form??**
  - Solubility table/rules are needed
  - See Table 1, Chapter 4

Table 4.1 Solubility Rules For Ionic Compounds in Water

Soluble Ionic Compounds

1. All common compounds of Group 1A(1) ions (Li^+, Na^+, K^+, etc.) and ammonium ion (NH_4^+) are soluble.
2. All common nitrates (NO_3^-), acetates (CH_3COO^- or $C_2H_3O_2^-$) and most perchlorates (ClO_4^-) are soluble.
3. All common chlorides (Cl^-), bromides (Br^-) and iodides (I^-) are soluble, *except* those of Ag^+, Pb^{2+}, Cu^+, and Hg_2^{2+}.

Insoluble Ionic Compounds

1. All common metal hydroxides are insoluble, *except* those of Group 1A(1) and the larger members of Group 2A(2)(beginning with Ca^{2+}).
2. All common carbonates (CO_3^{2-}) and phosphates (PO_4^{3-}) are insoluble, *except* those of Group 1A(1) and NH_4^+.
3. All common sulfides are insoluble *except* those of Group 1A(1), Group 2A(2) and NH_4^+.

Predicting Whether a Precipitation Reaction Occurs & Writing Equations

Molecular Equation

$Ca(NO_3)_{2\,(aq)} + Na_2SO_{4\,(aq)} \rightarrow CaSO_{4\,(s)} + NaNO_{3\,(aq)}$

Total Ionic Equation

$Ca^{2+}_{(aq)} + 2\ NO_3^-{}_{(aq)} + 2\ Na^+_{(aq)} + SO_4^{-2}{}_{(aq)} \rightarrow CaSO_{4\,(s)} + 2\ Na^+_{(aq)} + 2\ NO_3^-{}_{(aq)}$

Net Ionic Equation

$Ca^{2+}_{(aq)} + SO^{-2}_{(aq)} \rightarrow CaSO_{4\,(s)}$

- Spectator Ions are **Na^+** and **NO_3^-**
- Balance by Charge and Mass!!

Precipitation Reactions: Will a Precipitate Form?

a) $Na_2SO_{4\,(aq)} + Ba(NO_3)_{2\,(aq)}$

b) $CaCl_{2\,(aq)} + Na_2CO_{3\,(aq)}$

Acids

- **Acids**
  - substances that produces H^+ (H_3O^+) ions when dissolved in water.

$HI_{(g)} + H_2O_{(L)} \rightarrow H_3O^+_{(aq)} + I^-_{(aq)}$

$HI_{(g)} \xrightarrow{H_2O} H^+_{(aq)} + I^-_{(aq)}$

The Hydrated Proton

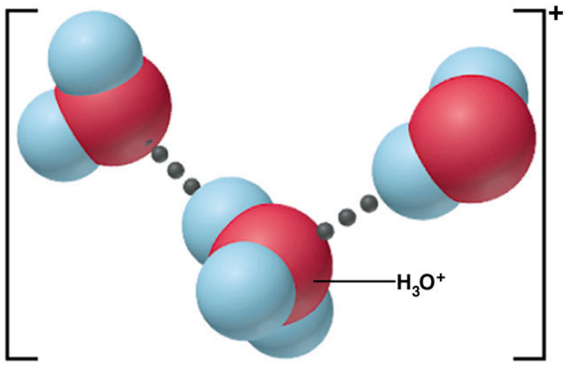


Fig. 4.4

Bases

- **Bases**
  - substances that produces OH^- ions when dissolved in water.

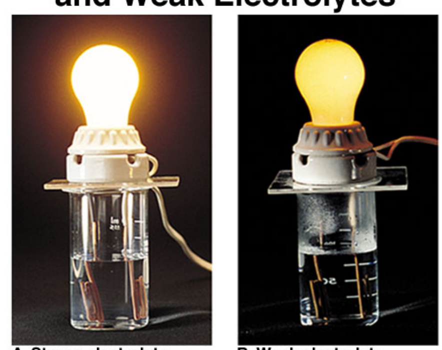
$$NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^{1-}_{(aq)}$$

$$NH_{3\,(g)} + H_2O_{(L)} \longrightarrow NH_4^+{}_{(aq)} + OH^{1-}_{(aq)}$$

Strong vs. Weak Acids and Bases

- **Acids and bases**
  - **May be strong or weak electrolytes**
  - **Strength determined by the degree of *ionization* in water**
  - **Strong acids and bases *ionize* completely, and are *strong* electrolytes.**
  - . Weak acids and bases *ionize* weakly and are *weak* electrolytes

The Behavior of Strong and Weak Electrolytes



A Strong electrolyte B Weak electrolyte

Fig. 4.7

Table 4.2 Selected Acids and Bases

Acids	Bases
Strong	**Strong**
hydrochloric acid, HCl	sodium hydroxide, NaOH
hydrobromic acid, HBr	potassium hydroxide, KOH
hydroiodic acid, HI	calcium hydroxide, $Ca(OH)_2$
nitric acid, HNO_3	strontium hydroxide, $Sr(OH)_2$
sulfuric acid, H_2SO_4	barium hydroxide, $Ba(OH)_2$
perchloric acid, $HClO_4$	
Weak	**Weak**
hydrofluoric acid, HF	ammonia, NH_3
phosphoric acid, H_3PO_4	
acetic acid, CH_3COOH (or $HC_2H_3O_2$)	

Strong Acids and the Molarity of H^+ Ions in Aqueous Solutions of Acids

Problem: What is the molarity of the sulfate and hydronium ions in a solution prepared by dissolving 155g of sulfuric acid into sufficient water to produce 2.30 Liters of acid solution?

$$H_2SO_{4\,(l)} + 2\ H_2O_{(l)} \longrightarrow 2\ H_3O^+_{(aq)} + SO_4{}^{-2}{}_{(aq)}$$

0.597 Molar in H^+ 0.687 Molar in $SO_4{}^{-2}$

Metathesis Reactions

(Double displacement or **double replacement reactions)**

AB (aq) + **CD** (aq) → **CB** + **AD**

Only occur if one of the following form

- **Precipitate (ppt)**
- **Gas**
- **Weak electrolyte** (e.g. acid-base reactions to form H_2O)

Acid - Base Reactions: Neutralization Rxns.

The generalized reaction between an Acid and a Base is:

$$HX_{(aq)} + MOH_{(aq)} \rightarrow MX_{(aq)} + H_2O_{(L)}$$

Acid + Base = Salt + Water

Writing Balanced Equations for Neutralization Reactions

Problem: Write balanced molecular and net ionic equations for the following chemical reactions:

a) Calcium hydroxide(aq) and hydrochloric acid(aq)

b) Lithium hydroxide(aq) and sulfuric acid(aq)

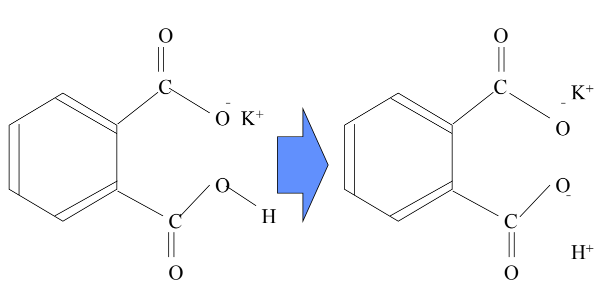
Finding the Concentration of Base from an Acid - Base Titration (I)

Problem: A titration is performed between **sodium hydroxide** and **potassium hydrogenphthalate (KHP)** to standardize the base solution, by placing **50.00 mg of solid potassium hydrogenphthalate** in a flask with a few drops of an indicator. A buret is filled with the base, and the initial buret reading is **0.55 ml**; at the end of the titration the buret reading is **33.87 ml**. What is the concentration of the base? **Molar mass of KHP is 204.2 g/mole**

$$HKC_8H_4O_{4\,(aq)} + OH^-_{(aq)} \rightarrow KC_8H_4O_4^-{}_{(aq)} + H_2O_{(aq)}$$

Answer: molarity of base = 0.07349 M

Potassium Hydrogenphthalate $KHC_8H_4O_4$



An Acid-Base Titration

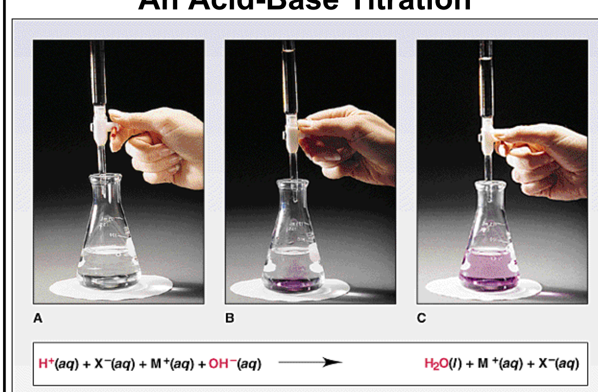


Fig. 4.8

Finding the Concentration of Base from an Acid - Base Titration (II)

$$\text{moles KHP} = \frac{50.00\ \text{mg KHP}}{\frac{204.2\ \text{g KHP}}{1\ \text{mol KHP}}} \times \frac{1.00\ \text{g}}{1000\ \text{mg}} = 0.00024486\ \text{mol KHP}$$

Volume of base = Final buret reading - Initial buret reading
= 33.87 ml - 0.55 ml = **33.32 ml of base**

one mole of acid = one mole of base; therefore 0.00024486 moles of acid will yield 0.00024486 moles of base in a volume of 33.32 ml.

$$\text{molarity of base} = \frac{0.00024486\ \text{moles}}{0.03332\ \text{L}} = 0.07348679\ \text{moles per liter}$$

molarity of base = 0.07349 M

Finding the Concentration of Acid from an Acid - Base Titration

Volume (L) of base (difference in buret readings)

↓ **M (mol/L) of base**

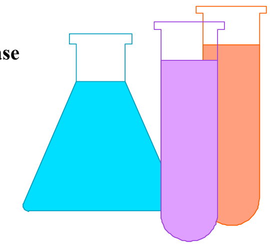
Moles of base

↓ **molar ratio**

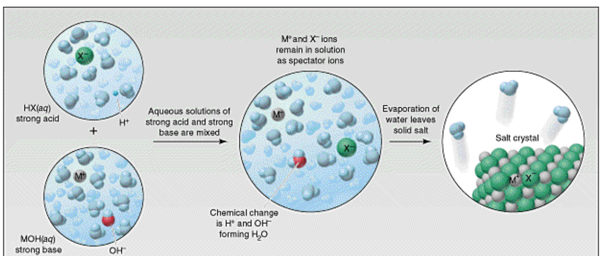
Moles of acid

↓ **volume (L) of acid**

M (mol/L) of acid



An Aqueous Strong Acid-Strong Base Reaction on the Atomic Scale



$$H^+(aq) + X^-(aq) + M^+(aq) + OH^-(aq) \xrightarrow{\text{mix}} H_2O(l) + M^+(aq) + X^-(aq) \xrightarrow{\Delta} H_2O(g) + MX(s)$$

Acid + Base → Salt + H_2O

Fig. 4.9

An Acid-Base Reaction That Forms a Gaseous Product

Molecular equation

$$NaHCO_3(aq) + CH_3COOH(aq) \longrightarrow CH_3COONa(aq) + CO_2(g) + H_2O(l)$$

Total ionic equation

$$Na^+(aq) + HCO_3^-(aq) + CH_3COOH(aq) \longrightarrow CH_3COO^-(aq) + Na^+(aq) + CO_2(g) + H_2O(l)$$

Net ionic equation

$$HCO_3^-(aq) + CH_3COOH(aq) \longrightarrow CH_3COO^-(aq) + CO_2(g) + H_2O(l)$$

The reaction of acid with carbonates or bicarbonates will produce carbon dioxide gas that is released from solution as a gas in the form of bubbles that leave the solution.

Fig. 4.10

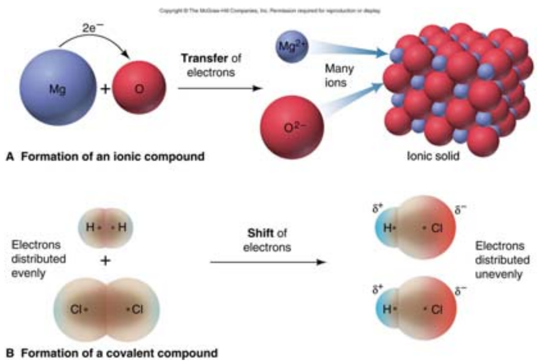
Oxidation-Reduction Reactions

- **"Redox Reactions"**
  - **Involve the transfer of one or more electrons from one substance to another**
  - **Examples**
    - **Formation of compounds from its elements and vice versa**
    - **Combustion reactions**
    - **Reactions that produce electricity in batteries**
    - **Cellular Respiration (energy production in cells)**
- **Objectives**
  - **Determine if a reaction is a redox reaction and identify the substances that are oxidized and reduced**
  - **To balance simple redox reactions**

Oxidation and Reduction

- **Oxidation**
  - Loss of Electrons
- **Reduction**
  - Gain of Electrons
- **L.E.O. the Lion said G.E.R.**

Figure 4.10 **The redox process in compound formation**



Oxidation Numbers

- Rules for Assigning Oxidation Numbers (Table 4.3, page 148, 3ed)

  Examples
  - **Ca, Ca^{2+}, $CaCl_2$, $CuSO_4$**
  - **H_2, H_2O, HNO_3, NO_3^{1-}, H_2SO_4, H_2SO_3, HCO_3^{1-}**
  - **Na_2O_2 , H_2O_2, ClO_2, FCl, MgH_2, BH_3**
- **Oxidation Number: Charge an atom would have *if* electrons in each of its bonds belonged *entirely* to the more *electronegative* element**

General Rules for Assigning an Oxidation Number

1. For an atom in its elemental form (Na, O_2 Cl_2, etc.) the **Ox. No. = 0**
2. For monotomic ions: **Ox. No. = ion charge**
3. The sum of Ox. No. values for the atoms in a compound equals zero.
4. **Polyatomic ions**: The sum of the Ox. No. values for the atoms in a equals the ion charge.

Specific Rules for Assigning an Oxidation Number

1. **Group 1A** = +1 in all compounds
2. **Group 2A** = +2 in all compounds
3. **Hydrogen** = +1 in combination with nonmetals
   = -1 in combination with metals or boron
4. **Fluorine** = -1 in all compounds
5. **Oxygen** = -1 in peroxides (O_2^{2-})
   = -2 in all other compounds (except with F)
6. **Group 7A** = -1 in combination with metals, nonmetals (except O), and other halogens lower in the group

Sample Problem 4.6 **Determining the Oxidation Number of an Element**

PROBLEM: **Determine the oxidation number (O.N.) of each element in…..**

(a) zinc chloride **(b) sulfur trioxide** **(c) nitric acid**

PLAN: **The O.N.s of the ions in a polyatomic ion add up to the charge of the ion and the O.N.s of the ions in the compound add up to zero.**

SOLUTION:

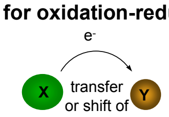
(a) **$ZnCl_2$**. The O.N. for zinc is +2 and that for chloride is -1.

(b) **SO_3**. Each oxygen is an oxide with an O.N. of -2. Therefore the O.N. of sulfur must be +6.

(c) **HNO_3**. H has an O.N. of +1 and each oxygen is -2. Therefore the N must have an O.N. of +5.

Terminology for oxidation-reduction (redox) reactions

Figure 4.12



X loses electron(s)	Y gains electron(s)
X is oxidized	Y is reduced
X is the reducing agent	Y is the oxidizing agent
Oxidation number increase	oxidation number decreases

Oxidizing Agents vs Reducing Agents

- **Oxidizing Agent**
  - A substance that causes oxidation
    - It is reduced in the process....Why?
- **Reducing Agent**
  - A substance that causes reduction
    - It is oxidized in the process....Why?
- **Redox Reactions**
  - Reaction in which oxidation numbers change

Use of Oxidation Numbers to Identify Oxidation and Reduction

- **Oxidation occurs if the oxidation number increases.........Why?**
- **Reduction occurs if the oxidation number decreases........Why?**
- **Practice..........**

Use of Oxidation Numbers to Identify Oxidation and Reduction

- **Identify the substances that are oxidized and reduced in the following examples**

$$Zn_{(s)} + 2\ HCl_{(aq)} \rightarrow ZnCl_{2\ (aq)} + H_{2\ (g)}$$

$$S_{8\ (s)} + 12\ O_{2\ (g)} \rightarrow 8\ SO_{3\ (g)}$$

$$NiO_{(s)} + CO_{(g)} \rightarrow Ni_{(s)} + CO_{2\ (g)}$$

- **Identify the oxidizing and reducing agents in each reaction**

Use of Oxidation Numbers to Identify Oxidation and Reduction

- **Identify the substances that are oxidized and reduced in the following examples**

$$2\ Ag\ NO_{3\ (aq)} + Cu_{(s)} \rightarrow Cu(NO_3)_{2(aq)} + 2\ Ag_{(s)}$$

- **Identify the oxidizing and reducing agents in each reaction**

Use of Oxidation Numbers to Identify Oxidation and Reduction

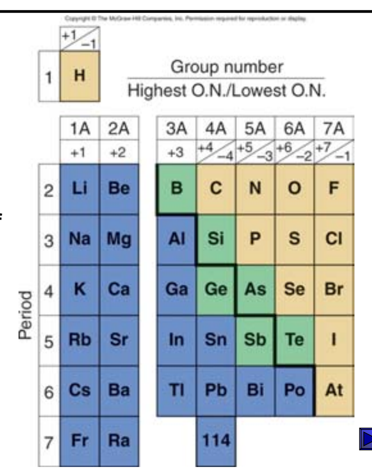
- **Identify the substances that are oxidized and reduced in the following examples**

$$N_{2(g)} + 2\ O_{2\ (g)} \rightarrow 2\ NO_{2\ (g)}$$

- **Identify the oxidizing and reducing agents in each reaction**

Figure 4.11

Highest and lowest oxidation numbers of reactive main-group elements



Main Group Elements

Period	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	H +1 -1							He
2	Li +1	Be +2	B +3	C +4,+2 -1,-4	N all from +5 → -3	O -1,-2	F -1	Ne
3	Na +1	Mg +2	Al +3	Si +4,-4	P +5,+3 -3	S +6,+4 +2,-2	-1 Cl +7,+5 +3,+1	Ar
4	K +1	Ca +2	Ga +3, +2	Ge +4,+2 -4	As +5,+3 -3	Se +6,+4 -2	-1 Br +7,+5 +3,+1	Kr +2
5	Rb +1	Sr +2	In +3,+2 +1	Sn +4,+2, -4	Sb +5,+3 -3	Te +6,+4 -2	-1 I +7,+5 +3,+1	Xe +6,+4 +2
6	Cs +1	Ba +2	Tl +3,+1	Pb +4,+2	Bi +3	Po +6,+4 +2,-2	-1 At +7,+5 +3,+1	Rn +2

Transition Metals

Possible Oxidation States

IIIB	IVB	VB	VIB	VIIB	VIIIB	VIIIB	VIIIB	IB	IIB
Sc +3	Ti +4,+3 +2	V +5,+4 +3+2	Cr +6,+3 +2	+2Mn +7,+6 +4,+3	Fe +3,+2	Co +3,+2	Ni +2	Cu +2,+1	Zn +2
Y +3	Zr +4,+3	Nb +5,+4 +2	Mo +6,+5 +4,+3	Tc +7,+5 +4	Ru +8,+5 +4,+3	Rh +4,+3	Pd +4,+2	Ag +1	Cd +2
La +3	Hf +4,+3	Ta +5,+4 +3	W +6,+5 +4	Re +7,+5 +4	+2Os +8,+6 +4,+3	Ir +4,+3 +1	Pt +4,+2	Au +3,+1	Hg +2,+1

Balancing REDOX Equations: The Oxidation Number Method

Step 1) Assign oxidation numbers to all elements in the equation.

Step 2) From the changes in oxidation numbers, identify the oxidized and reduced species.

Step 3) Compute the number of electrons lost in the oxidation and gained in the reduction from the oxidation number changes. Draw tie-lines between these atoms to show electron changes.

Step 4) Multiply one or both of these numbers by appropriate factors to make the electrons lost equal the electrons gained, and use the factors as balancing coefficients.

Step 5) Complete the balancing by inspection, adding states of matter.

REDOX Balancing Using Ox. No. Method

0 → -2: +2 e⁻ (O)

$\underline{2}\ H_{2(g)} + \underline{\ \ }\ O_{2(g)} \rightarrow \underline{2}\ H_2O_{(g)}$

0 → +1: - 1 e⁻ (H)

electrons lost must = electrons gained;

Therefore, multiply the hydrogen reaction by 2 to balance the equation

Sample Problem 4.8 Balancing Redox Equations by the Oxidation Number Method

PROBLEM: Use the oxidation number method to balance the following equation:

(a) $Cu(s) + HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + NO_2(g) + H_2O(l)$

SOLUTION:

Oxidation numbers: Cu 0; HNO_3 +1 +5 -2; $Cu(NO_3)_2$ +2 +5 -2; NO_2 +4 -2; H_2O +1 -2

(a) $Cu(s) + HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + NO_2(g) + H_2O(l)$

O.N. of Cu increases because it loses 2e⁻; it is oxidized and is the reducing agent.
O.N. of N decreases because it gains1e⁻; it is reduced and is the oxidizing agent.

loses 2e⁻ balance other ions

$Cu(s) + HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + NO_2(g) + H_2O(l)$

gains 1e⁻ x2 to balance e⁻

balance unchanged polyatomic ions

Sample Problem 4.8 Balancing Redox Equations by the Oxidation Number Method

continued

$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$ (4 corrected from 2)

Oxidation numbers: PbS +2 -2; O_2 0; PbO +2 -2; SO_2 +4 -2

(b) $PbS(s) + O_2(g) \rightarrow PbO(s) + SO_2(g)$

loses 6e⁻

$PbS(s) + 3/2\,O_2(g) \rightarrow PbO(s) + SO_2(g)$

gains 2e⁻ per O; need 3/2 O_2 to make $3O^{2-}$

Multiply by 2 to have whole number coefficients.

$2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$

REDOX Balancing Using Ox. No. Method

0   -1 e⁻   +1

$Ag_{(s)} + CN^-_{(aq)} + O_{2\,(g)} \rightarrow Ag(CN)_2^-{}_{(aq)} + OH^-_{(aq)}$

0   + 2 e⁻   - 2

To balance electrons we must put a **4 in front of the Ag**, since each oxygen looses two electrons, and they come two at a time! That requires us to put a 4 in front of the silver complex, yielding 8 cyanide ions.

$4\,Ag_{(s)} + 8\,CN^-_{(aq)} + O_{2\,(g)} \rightarrow 4\,Ag(CN)_2^-{}_{(aq)} + OH^-_{(aq)}$

Add **4 OH^-** to balance charge. Since there hydrogen is absent on the reactant side, add **2 H_2O** to balance the hydrogen and oxygen.

$4\,Ag_{(s)} + 8\,CN^-_{(aq)} + O_{2\,(g)} + 2\,H_2O_{(l)} \rightarrow 4\,Ag(CN)_2^-{}_{(aq)} + 4\,OH^-_{(aq)}$

Sample Problem 4.9   **Finding an Unknown Concentration by a Redox Titration**

PROBLEM: Calcium ion (Ca^{2+}) is required for blood to clot and for many other cell processes. An abnormal Ca^{2+} concentration is indicative of disease. To measure the Ca^{2+} concentration, 1.00mL of human blood was treated with $Na_2C_2O_4$ solution. The resulting CaC_2O_4 precipitate was filtered and dissolved in dilute H_2SO_4. This solution required 2.05mL of 4.88×10^{-4}M $KMnO_4$ to reach the end point.

$2\,KMnO_4(aq) + 5\,CaC_2O_4(s) + 8\,H_2SO_4(aq) \rightarrow 2\,MnSO_4(aq) + K_2SO_4(aq) + 5\,CaSO_4(s) + 10\,CO_2(g) + 8\,H_2O(l)$

(a) Calculate the amount (mol) of Ca^{2+}.

(b) Calculate the amount (mol) of Ca^{2+} ion concentration expressed in units of mg Ca^{2+}/100mL blood.

PLAN:

(a)

volume of $KMnO_4$ soln → multiply by M → mol of $KMnO_4$ → molar ratio → mol of CaC_2O_4 → ratio of elements in formula → mol of Ca^{2+}

Sample Problem 4.9   **Finding an Unknown Concentration by a Redox Titration**

continued   **SOLUTION:**

$2.05\text{mL soln} \times \frac{L}{10^3\text{ mL}} \times \frac{4.88\times10^{-4}\text{mol } KMnO_4}{L} = 1.00\times10^{-6}\text{mol } KMnO_4$

$1.00\times10^{-6}\text{mol } KMnO_4 \times \frac{5\text{mol } CaC_2O_4}{2\text{mol } KMnO_4} = 2.50\times10^{-6}\text{ mol } CaC_2O_4$

$2.50\times10^{-6}\text{ mol } CaC_2O_4 \times \frac{1\text{mol } Ca^{2+}}{1\text{mol } CaC_2O_4} = 2.50\times10^{-6}\text{ mol } Ca^{2+}$

PLAN:

(b)

mol Ca^{2+}/1mL blood → multiply by 100 → mol Ca^{2+}/100mL blood → multiply by M → g Ca^{2+}/100mL blood → 10^{-3}g = 1mg → mg Ca^{2+}/100mL blood

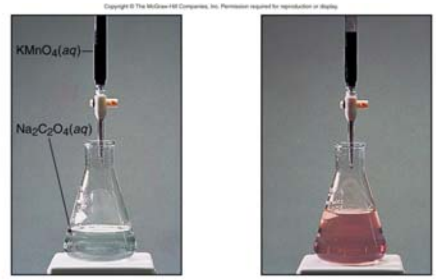
SOLUTION:

$\frac{2.50\times10^{-6}\text{ mol } Ca^{2+}}{1\text{mL blood}} \times 100 = \frac{2.50\times10^{-4}\text{ mol } Ca^{2+}}{100\text{mL blood}}$

$\frac{2.50\times10^{-4}\text{ mol } Ca^{2+}}{100\text{mL blood}} \times \frac{40.08\text{g } Ca^{2+}}{\text{mol } Ca^{2+}} \times \frac{\text{mg}}{10^{-3}\text{g}}$

$= 10.0\text{mg } Ca^{2+}/100\text{mL blood}$

Figure 4.13   **A redox titration**

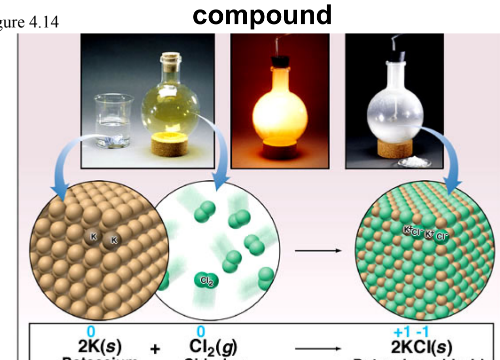


Net ionic equation:

$\overset{+7}{2MnO_4^-(aq)} + \overset{+3}{5C_2O_4^{2-}(aq)} + 16H^+(aq) \rightarrow \overset{+2}{2Mn^{2+}(aq)} + \overset{+4}{10CO_2(g)} + 8H_2O(l)$

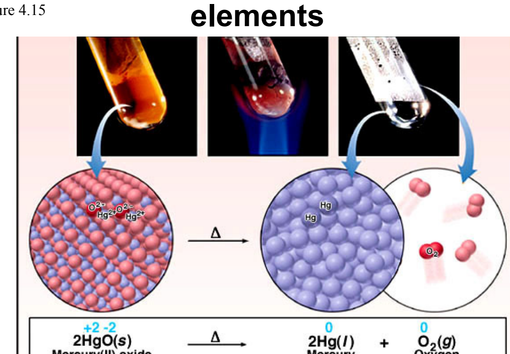
Combining elements to form an ionic compound

Figure 4.14



Decomposing a compound to its elements

Figure 4.15



Three Views of a Single-displacement Reaction

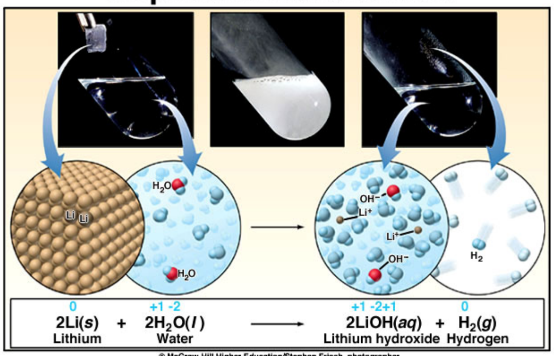
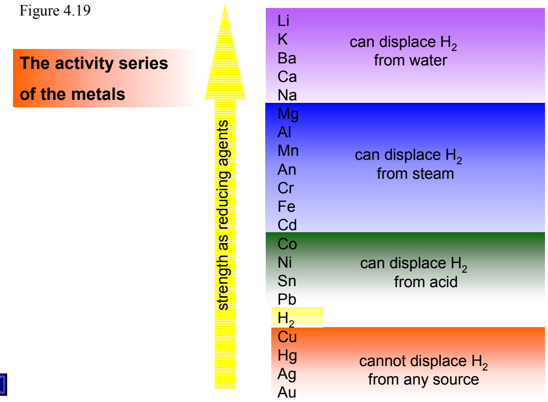


Fig. 4.17

Figure 4.19



Metals Displace Hydrogen from Water

Metals that will displace hydrogen from cold water:

$2\ Cs_{(s)} + 2\ H_2O \rightarrow H_{2\ (g)} + 2\ CsOH_{(aq)}$

$Ba_{(s)} + 2\ H_2O_{(l)} \rightarrow H_{2\ (g)} + Ba(OH)_{2\ (aq)}$

$2\ Na_{(s)} + 2\ H_2O_{(l)} \rightarrow H_{2\ (g)} + 2\ NaOH_{(aq)}$

Metals that will displace hydrogen from steam:

$Mg_{(s)} + 2\ H_2O_{(g)} \rightarrow H_{2\ (g)} + Mg(OH)_{2\ (s)}$

$2\ Cr_{(s)} + 6\ H_2O_{(g)} \rightarrow 3\ H_{2\ (g)} + 2\ Cr(OH)_{3\ (s)}$

$Zn_{(s)} + 2\ H_2O_{(g)} \rightarrow H_{2\ (g)} + Zn(OH)_{2\ (s)}$

Metals Displace Hydrogen from Acids

Reactions of metals above hydrogen in the activity series

$Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2\ (aq)} + H_{2\ (g)}$

$Zn_{(s)} + H_2SO_{4\ (aq)} \rightarrow ZnSO_{4\ (aq)} + H_{2\ (g)}$

$2\ Al_{(s)} + 6\ HCl_{(aq)} \rightarrow 2\ AlCl_{3\ (aq)} + 3\ H_{2\ (g)}$

$Cd_{(s)} + 2\ HBr_{(aq)} \rightarrow CdBr_{2\ (aq)} + H_{2\ (g)}$

$Sn_{(s)} + 2\ H_2SO_{4\ (aq)} \rightarrow Sn(SO_4)_{2\ (aq)} + 2\ H_{2\ (g)}$

Reactions of metals below hydrogen in the activity series

$Cu_{(s)} + HCl_{(aq)} \rightarrow$ No Reaction!

$Au_{(s)} + H_2SO_{4\ (aq)} \rightarrow$ No Reaction!

Three Views of Copper Displacing Silver Ions from Solution

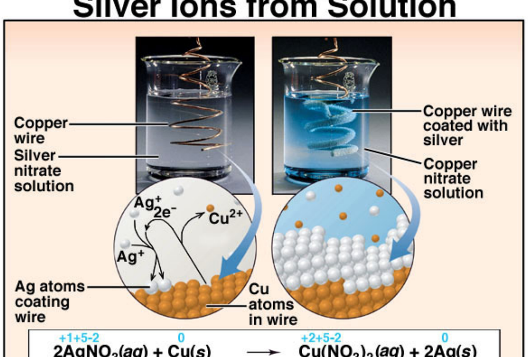


Fig. 4.19

Single-Displacement Reactions - Metals Replace Metal Ions from Solution

$Ba_{(s)} + Co(NO_3)_{2\ (aq)} \rightarrow Co_{(s)} + Ba(NO_3)_{2\ (aq)}$

$Cd_{(s)} + AgNO_{3\ (aq)} \rightarrow 2\ Ag_{(s)} + Cd(NO_3)_{2\ (aq)}$

$Mg_{(s)} + Pb(NO_3)_{2\ (aq)} \rightarrow Pb_{(s)} + Mg(NO_3)_{2\ (aq)}$

$Al_{(s)} + Ba(NO_3)_{2\ (aq)} \rightarrow$ No Reaction

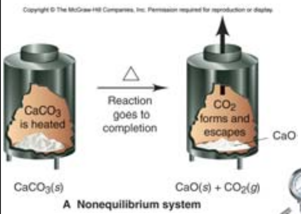
$4\ Cr_{(s)} + 3\ PtCl_{4\ (aq)} \rightarrow 3\ Pt_{(s)} + 4\ CrCl_{3\ (aq)}$

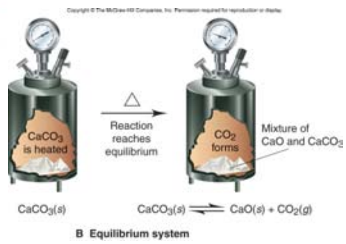
$Ca_{(s)} + Hg(NO_3)_{2\ (aq)} \rightarrow Hg_{(l)} + Ca(NO_3)_{2\ (aq)}$

$2\ Li_{(s)} + Na_2SO_{4\ (aq)} + H_2O_{(l)} \rightarrow 2\ NaOH_{(aq)} + H_{2\ (g)} + Li_2SO_{4\ (aq)}$

Figure 4.20

The equilibrium state





Many Chemical Reactions Are in a State of Dynamic Equilibrium

Solid - gas equilibrium processes

$$CaCO_{3\,(s)} \underset{\Delta}{\rightleftharpoons} CaO_{(s)} + CO_{2\,(g)}$$

Solution Equilibrium processes involving weak acids and bases

$$CH_3COOH_{(aq)} + H_2O_{(L)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

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