Chapter 4: Reaction in Aqueous Solutions

*4.1: Ionic Theory of Solutions:

- Solution: Homogeneous mixture forms from two or more substances.
- Solute: The substance present in a smaller amount in solution.
- Solvent: The substance present in larger amount in solution.

mixture
- Homogeneous (solution)
  - solvent (مذيب) large amount
  - solute (مذاب) small amount
- Heterogeneous

Arrhenius 1884:- Ionic theory of solution (substance produce freely moving ions when dissolved ~~completely~~ in water) ⇒ coduct electricity.

1 electrolyte (cations^{+}, anions^{-}): substance dissolve completely in water produce ions ⇒ conduct electricity
Ex: salts, acids, Bases

2 Non electrolyte: substance when dissolve in water does not produce ions ⇒ not conduct electricity
Ex: Sucrose:- $C_{12}H_{22}O_{11}$, Methane:- CH_3OH.

* Electrolytes:

1. strong electrolyte: 100% dissolve in water (dissociation 100%)

   Ex: salts $NaCl \xrightarrow{H_2O} Na^+ + Cl^-$

   strong acids and Bases $HCl \xrightarrow{H_2O} H^+ + Cl^-$

   $NaOH \xrightarrow{H_2O} Na^+ + OH^-$

2. weak electrolyte: weak Acids and Bases (partial dissociation in water)

   Ex: $CH_3COOH + H_2O \rightleftarrows CH_3COO^- + H_3O^+$

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

- Solubility :- is the ability to dissolve in water at specific temp.

soluble		Insoluble exception
G1A [Li^+, Na^+, K^+]		—
NH_4^+		—
$C_2H_3O^-$, NO_3^-		—
- I, Cl, Br	دائماً ~~solub~~ ذائب إلا اذا أجى مع $\rightarrow$	Ag^+, Hg^{+2}, Pb^{+2}, Hg_2^{+2}
SO_4^{-2}	دائماً ذائب الا إذا أجى مع $\rightarrow$	Ca^{+2}, Sr^{+2}, Ba^{+2}, Ag^+, Hg_2^{+2}, Pb^{+2}

Insoluble		Soluble exception
CO_3^{-2}, PO_4^{-3}, S^{-2}	$\rightarrow$	G1A, NH_4^+
OH^-	$\rightarrow$	G1A, Ca^{+2}, Sr^{+2}, Ba^{+2}

Example:

Hg_2Cl_2: insoluble $\quad$ Li_3PO_4: soluble

KI: soluble $\quad$ $BaCO_3$: insoluble

$Pb(NO_3)_2$: Soluble

$Al(OH)_3$: insoluble

*4.2: Molecular and Ionic Equations:

Molecular equation → Complete ionic equation → Net ionic equation (Balanced)

$$Ca(OH)_{2(aq)} + Na_2CO_{3(aq)} \longrightarrow CaCO_{3(s)} + 2NaOH_{(aq)}$$: Molecular eq.

$$Ca^{+2} + \cancel{2OH^-} + \cancel{2Na^+} + CO_3^{-2} \longrightarrow CaCO_{3(s)} + \cancel{2Na^+} + \cancel{2OH^-}$$: Complete ionic eq.

$$Ca^{+2} + CO_3^{-2} \longrightarrow CaCO_{3(s)}$$: Net ionic eq.

OH^-, Na^+ : spectator ions

Example: $\underline{\underline{2}}HClO_{4\,(aq)} + Ca(OH)_{2\,(aq)} \longrightarrow Ca(ClO_4)_{2\,(aq)} + 2H_2O_{(l)}$

$2H^+_{(aq)} + \cancel{2ClO_4^-}_{(aq)} + \cancel{Ca^{+2}}_{(aq)} + 2OH^-_{(aq)} \longrightarrow \cancel{Ca^{+2}}_{(aq)} + \cancel{2ClO_4^-}_{(aq)} + 2H_2O_{(l)}$

$2H^+_{(aq)} + 2OH^-_{(aq)} \longrightarrow 2H_2O_{(l)}$ net ionic equation

spectator ions: Ca^{+2}, ClO_4^-

Example: $\underline{HC_2H_3O_2}_{\,\underline{\underline{(s)}}} + NaOH_{(aq)} \longrightarrow \underline{NaC_2H_3O_2}_{\,(aq)} + H_2O_{(l)}$

$HC_2H_3O_2 + \cancel{Na^+} + OH^- \longrightarrow \cancel{Na^+} + C_2H_3O_2^- + H_2O$

$HC_2H_3O_2\,_{(s)} + OH^-_{(aq)} \longrightarrow C_2H_3O_2^-\,_{(aq)} + H_2O_{(l)}$

spectator ions: Na^+

Examples:

a) Of the species below, only .... is not an electrolyte:

A- HCl B- Rb_2SO_4 C- Ar D- KOH E- NaCl

b) a strong electrolyte is one that:.......... completely in solution:

A- reacts B- associates C- disappears D- ionizes

c) The balance net ionic equation for precipitation of $CaCO_3$ when aqueous solution of Na_2CO_3 and $CaCl_2$ are mixed is:

$$Na_2CO_{3(aq)} + CaCl_{2(aq)} \longrightarrow CaCO_{3(s)} + 2NaCl_{(aq)}$$

$$\cancel{2Na^+} + CO_3^{-2} + Ca^{+2} + \cancel{2Cl^-} \longrightarrow CaCO_{3(s)} + \cancel{2Na^+} + \cancel{2Cl^-}$$

$$CO_3^{-2} + Ca^{+2} \longrightarrow CaCO_{3(s)}$$ Net ionic equation

⇒ Types of chemical reactions:

1. 4.3: Precipitation reaction.
2. 4.4: Acid-Base reaction.
3. 4.5: Oxidation reduction reaction.

* 4.3: Precipitation Reaction:

is an insoluble solid compounds formed during a chemical reaction in solution

Ex: $MgCl_{2\,(aq)} + 2AgNO_{3\,(aq)} \longrightarrow 2AgCl_{(s)} + Mg(NO_3)_{2\,(aq)}$

$2Ag^+ + 2Cl^- \longrightarrow 2AgCl_{(s)}$ spectators → Mg^{+2}, NO_3^-

Ex: $2NaCl_{(aq)} + Fe(NO_3)_{2\,(aq)} \longrightarrow 2NaNO_{3\,(aq)} + FeCl_{2\,(aq)}$

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Ex: $Al_2(SO_4)_{3\,(aq)} + 6NaOH_{(aq)} \longrightarrow 2Al(OH)_{3\,(s)} + 3Na_2SO_{4\,(aq)}$

$2Al^{+3} + \cancel{3SO_4^{-2}} + \cancel{6Na^+} + 6OH^- \longrightarrow 2Al(OH)_{3\,(s)} + \cancel{6Na^+} + \cancel{3SO_4^{-2}}$

$2Al^{+3}_{(aq)} + 6OH^-_{(aq)} \longrightarrow 2Al(OH)_{3\,(s)}$

* 4.4: Acid-Base reaction:

- Acids : HCl → mono protic acid $\quad H_3PO_4$ → poly protic acid (tri protic)

  H_2SO_4 → Poly protic acid (diprotic)

- Bases: NH_3, $NaOH$, KOH, $Ca(OH)_2$

- Acid-Base indicators: dye used to distinguish between acidic and Basic solutions by means of the color changes it undergoes in these solutions.

  Ex: Red cabbage, litmus, phenolphthalein

- Acid-Base definition:

  Arrheniuse:-

  acid: Produce H^+ in water. $\quad HNO_3 \xrightarrow{H_2O} H^+ + NO_3^-$

  Base: Produce $O\bar{H}$ in water. $\quad NaOH \xrightarrow{H_2O} O\bar{H} + Na^+$

  Bronsted and Lowrye-

  acid: proton (H^+) donor. $\quad HNO_3 + H_2O \longrightarrow H_3O^+ + NO_3^-$

  Base: Proton (H^+) acceptor. $\quad NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$

Acids

- strong acids (strong electrolyte)
  Complete ionization in water Ex: $HCl, HNO_3, HClO_4, HI, HBr, H_2SO_4, HClO_3$
- weak acids (weak electrolyte)
  Partial ionization in water Ex: $CH_3COOH, HCN, HNO_2, HF, \ldots$

Bases

- strong Bases (strong electrolyte)
  complete ionization in water Ex: $LiOH, NaOH, Ba(OH)_2, KOH, Ca(OH)_2, Mg(OH)_2$
- weak Bases (weak electrolyte)
  Partial ionization in water Ex: $NH_3, Al(OH)_3, \ldots$

Exercise: classify each of the following species to bronsted acid or base:

A SO_4^{-2} Base

B HI Acid

C $H_2PO_4^-$ Acid + Base

D HCO_3^- Acid + Base

Conclusion: there are some compounds that can be Acid or base, and we call them

Amphoteric

Neotralization reaction:

$$\underset{strong}{Acid} + \underset{strong}{Base} \longrightarrow \underset{ionic}{salt} + H_2O$$

Ex: $\overset{strong}{2HCl} + \overset{strong}{Ca(OH)_2} \longrightarrow CaCl_2 + 2H_2O$

$2H^+ + \cancel{2Cl^-} + \cancel{Ca^{+2}} + 2OH^- \longrightarrow \cancel{Ca^{+2}} + \cancel{2Cl^-} + 2H_2O$

$2H^+ + 2OH^- \longrightarrow 2H_2O$

Ex: $\overset{weak}{HCN} + \overset{strong}{KOH} \longrightarrow KCN + H_2O$

$HCN + \cancel{K^+} + OH^- \longrightarrow \cancel{K^+} + CN^- + H_2O$

$HCN + OH^- \longrightarrow CN^- + H_2O$

Ex: $H_2SO_4 + 2NH_3 \longrightarrow (NH_4)_2SO_4$

$2H^+ + \cancel{SO_4^{-2}} + 2NH_3 \longrightarrow 2NH_4^+ + \cancel{SO_4^{-2}}$

$NH_3 + H^+ \longrightarrow NH_4^+$

Ex: $HNO_2 + NaOH \longrightarrow NaNO_2 + H_2O$

$HNO_2 + \cancel{Na^+} + OH^- \longrightarrow \cancel{Na^+} + NO_2^- + H_2O$

$HNO_2 + OH^- \longrightarrow NO_2^- + H_2O$

Acid-Base Reactions with Gas Formation:

- Certain salts, notably carbonates, sulfites, sulfides, react with acids to form a gaseous Product.

⇒ Consider: $Na_2CO_{3(aq)} + 2HCl_{(aq)} \longrightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$ How!!

⇒ normaly in this reaction an exchange happens beween cations and anions like the following reaction:

$$Na_2CO_{3(aq)} + 2HCl_{(aq)} \longrightarrow 2NaCl_{(aq)} + H_2CO_{3(aq)}$$

→ carbonic acid

⇒ But carbonic acid is unstable and decomposes to water and carbon dioxide gas, as the following reaction:

$$Na_2CO_{3(aq)} + 2HCl_{(aq)} \longrightarrow 2NaCl_{(aq)} + \underbrace{H_2O_{(l)} + CO_{2(g)}}_{H_2CO_{3(aq)}}$$

⇒ The net ionic equation for this reaction is:

$$CO_{3(aq)}^{-} + 2H^{+}_{(aq)} \longrightarrow H_2O_{(l)} + CO_{2(g)}$$

* some ionic compounds that evolve gases when treated with acids:

ionic compound	gas
carbonate (CO_3^{-2})	CO_2
sulfite (SO_3^{-2})	SO_2
sulfide (S^{-2})	H_2S

Example: write the molecular equation and the net ionic equation for the reaction of Zinc sulfide with hydrochloric acid:

$$ZnS_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_2S_{(g)}$$

$$ZnS + 2H^+ + \cancel{2Cl^-} \longrightarrow Zn^{+2} + \cancel{2Cl^-} + H_2S$$

$$ZnS_{(s)} + 2H^+_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + H_2S_{(g)}$$

Example: H_2SO_4 is neutralized by NaOH in aqueous solution, the net ionic equation is:

A) $SO_{4(aq)}^{-2} + 2Na^+_{(aq)} \longrightarrow Na_2SO_{4(aq)}$

B) $SO_{4(aq)}^{-2} + 2Na^+_{(aq)} \longrightarrow Na_2SO_{4(s)}$

C) $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$

D) $H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow 2H_2O_{(l)} + SO_{4(aq)}^{-2}$

$$H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$$

$$2H^+ + \cancel{SO_4^{-2}} + \cancel{2Na^+} + 2OH^- \longrightarrow \cancel{2Na^+} + \cancel{SO_4^{-2}} + 2H_2O$$

$$2H^+ + 2OH^- \longrightarrow 2H_2O$$

$$H^+ + OH^- \longrightarrow H_2O$$

Note: Double Replacement Reactions:

1] Precipitation Reaction.

2] Acid and Base Neutralization.

3] Reaction Leading Gas.

$$AB + CD \longrightarrow CB + AD$$

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*4.5: Oxidation - Reduction Reaction:

- Electron transfer rxn
- Oxidation reaction: loss of electrons.
- Reduction reaction: gain of electrons.

عامل مختزل - Reducing agent: هو الذي تأكسد وأجبر غيره على الاختزال

عامل مؤكسد - Oxidizing agent: هو الذي اختزل وأجبر غيره على التأكسد

$$Fe^{\circ}_{(s)} + \overset{+2}{Cu}SO_{4(aq)} \longrightarrow \overset{+2}{Fe}\underline{\underline{SO_4}}_{(aq)} + Cu^{\circ}_{(s)}$$

$$Fe^{\circ}_{(s)} + Cu^{+2}_{(aq)} \longrightarrow Fe^{+2}_{(aq)} + Cu_{(s)}$$

$Fe \longrightarrow Fe^{+2} + 2\bar{e}$: Half Oxidation, Reducing agent

$Cu^{+2} + 2\bar{e} \longrightarrow Cu$: Half Reduction, Oxidizing agent

- Oxidation number: actual charge of the atoms if $\bar{e}$s completely transferd.

$$2Ca^{\circ}_{(s)} + O^{\circ}_{2(g)} \longrightarrow 2\overset{+2}{Ca}\overset{-2}{O}_{(s)}$$

*Oxidation Number rules:

1] oxidation number of atoms in element = Zero
Ex: Ca, K, Na, Li, Cl_2, O_2

2] In mono atomic ions, The oxidation number equal charge of the ion
Ex: $Cl^- = -1$, $Na^{\oplus} = +1$, $Ca^{+2} = +2$

3] Oxygen = -2 except in peroxid $O_2 = -1$ $\overset{+1}{H_2}\overset{-2}{O}$ $\overset{+1}{H_2}\overset{-1}{O_2}$

4] Hydrogen = +1 except in products with G1A, G2A = -1
Ex: $CaH_2 \rightarrow H = -1$
NaH, $LiH \rightarrow H = -1$

5] F always -1, other halogens (Cl, Br, I) = -1 except with halogens above it or higher (EN), then oxidation number = positive
Ex: ClF, BrI

F → Cl → Br → I

Higher EN ⟶ Lower EN

6] The sum of oxidation number of the atoms in compound = 0
The sum of oxidation number of the atoms in polyatomic ions = the charge of the ion

$\overset{+1}{Na}\overset{-1}{Cl}$ SO_4^{-2} $-2 = -2\times4 + 6$
$0 = +1 - 1$

Some common Oxidation reduction reactions:

1] combination reaction: $A + B \longrightarrow C$

$2\overset{0}{Na}_{(s)} + \overset{0}{Cl_2}_{(g)} \longrightarrow 2\overset{+1}{Na}\overset{-1}{Cl}_{(s)}$

$\overset{+2}{Ca}\overset{-2}{O}_{(s)} + S\overset{-2}{O_2}_{(g)} \longrightarrow \overset{+2}{Ca}S\overset{-2}{O_3}_{(s)}$ ⟶ The charge is the same, not oxidation reduction reaction combination.

$2\overset{0}{Al}_{(s)} + 3\overset{0}{Br_2} \longrightarrow 2\overset{+3}{Al}\overset{-1}{Br_3}$

2] Decomposition Reaction: $C \xrightarrow[\text{تسخين}]{\text{طاقة}} A + B$

$2\overset{+2}{Hg}\overset{-2}{O}_{(s)} \xrightarrow{\Delta} 2\overset{0}{Hg} + \overset{0}{O_2}$

$\overset{+2}{Ca}C\overset{-2}{O_3} \xrightarrow[\text{burning}]{\Delta} \overset{+2}{Ca}\overset{-2}{O} + C\overset{-2}{O_2}$ ⟶ not oxidation reduction reaction.

3] Combustion Reaction: $A + O_2 \longrightarrow B$

$4\overset{0}{Fe} + 3\overset{0}{O_2} \longrightarrow 2\overset{+3}{Fe_2}\overset{-2}{O_3}$

4] Displacement Reaction: $A + BC \longrightarrow B + AC$

$\overset{0}{Cu} + 2\overset{+}{Ag}\overset{-}{NO_3} \longrightarrow \overset{+2}{Cu}(\overset{-}{NO_3})_2 + 2\overset{0}{Ag}$ ⟶ metal displacement.

$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$ ⟶ hydrogen displacement.

$Cl_2 + 2KBr \longrightarrow 2KCl + Br_2$ ⟶ halogen displacement.

*4.6. Balancing of simple oxidation reduction reaction:

$Zn_{(s)} + Ag^{+}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + Ag$

$2 \cdot (Ag^{+} + \bar{e} \longrightarrow Ag) \cdot 2Ag^{+} + 2\bar{e} \longrightarrow 2Ag$ Half Reduction

$Zn \longrightarrow Zn^{+2} + 2\bar{e}$ Half Oxidation

$Zn + 2Ag^{+} \longrightarrow Zn^{+2} + 2Ag$

$0 + 2 \longrightarrow +2 + 0$

$2 \longrightarrow 2$

$Mg^{0}_{(s)} + N^{0}_{2\,(g)} \longrightarrow Mg^{+2}_{3} N^{-3}_{2\,(s)}$

$3 \times (Mg \longrightarrow Mg^{+2} + 2\bar{e}) = 3Mg \longrightarrow 3Mg^{+2} + 6\bar{e}$

$N_2 + 6\bar{e} \longrightarrow N_2^{-3}$ $\cancel{6\bar{e}}$ $\cancel{-6 \longrightarrow 0}$

$3Mg + N_2 \longrightarrow Mg_3N_2$

$3Mg^{+2} + N_2^{-3}$

*4.7: Molar Concentration: Molarity (M)

its amole of solute dissolved in one litter solution.

$$\boxed{\text{Molarity } (M) = \frac{\text{mol Solute}}{\text{Litter Solution}} = \frac{mol}{V} = M = \frac{mol}{L}}$$

Example: $NaNO_3 = 0.38\,g$ placed in $50.0\,mL$, what is the molarity? (m.w $NaNO_3 = 85\,g/mol$)

$$M = \frac{mol}{V} = \frac{4.47\times10^{-3}}{0.05\,L} = \boxed{0.0894\,M}$$

$$mol\ NaNO_3 = \frac{0.38}{85} = 4.47\times10^{-3}\ mol$$

Example: An experiments calls for the addition to a reaction vessel of 0.184 g of sodium hydroxide, NaOH, in aqueous solution, How many (mL) of 0.150 M NaOH should be added?

$$\frac{0.184}{40} = 4.6\times10^{-3}\ mol$$

$$M = \frac{mol}{V}$$

$$0.15 = \frac{4.6\times10^{-3}}{V} \rightarrow \boxed{V = 0.03067\,L}$$

$$\boxed{V = 30.7\ mL}$$

*4.8: Diluting Solutions:

Prepare less concentration from high concentration.

$\Longrightarrow$ add solvent

$$M = \frac{mol}{V} \rightarrow mol = \underline{M \times V}$$

constant

mole before = mole after

$$\boxed{M_i \times V_i = M_f \times V_f}$$

Example: 14.8 M NH_3 ~~是~~ you are given, How many mL of this solution do you require to give 100.0 mL of 1 M NH_3 when diluted?

14.8 M_i ?? mL_i

1 M_f 100.0 mL_f

$$14.8 \times V_i = 1 \times 100 \text{ mL}$$

$$V_i = \frac{100}{14.8} = \boxed{6.76 \text{ mL}}$$

*4.9: Gravimetric Analysis:

Gravimetric → weight

Analytical chemistry

Quantitave Analysis

Qualtative Analysis

- Gravimetric: quantitave analysis in which amount of a species in material is determined by converting the species to a product can be isolated completely and weighted.

Example: A 1L sample of polluted water was analyzed for lead (II) ion, Pb^{2+}, by adding an excess of sodium sulfate to it. The mass of lead (II) sulfate that precipitated was 229.8 mg. what is the mass of lead in litter of the water? Give the answer as mg of lead per L of Soln

$PbSO_4$

$$\%Pb = \frac{207.2\,\cancel{g/mol}}{303.3\,\cancel{g/mol}} * 100\% = 68.32\%$$

$$\text{mass } Pb = 229.8 \times 0.6832 = \boxed{157 \text{ mg}}$$

*4.10: Volumetric Analysis: (Determine Volume)

- Volumetrics: method of analysis based on titration.
- Titration: is a procedure for determining the amount of substance "A" by adding a carefully measured volume of solution with known concentration of "B" until the reaction of "A" and "B" is just complete. ⇒ when indicator change it color.

※ Acid-Base titration:

$HCl + NaOH \longrightarrow NaCl + H_2O$

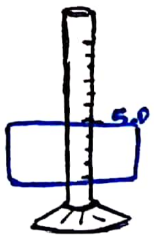
5mL   10L

M=??   M=0.5

⇒ Complete rxn mol A = mol B from balance chemical equation.

Equivelence point: Theoritical point where amount of A = amount of B.

End point: The point where the color of indicator changes.



5mL HCl

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

H_2SO_4: 35 mL, 0.175 M $NaOH$: ? mL, 0.250 M

$$\text{mol } H_2SO_4 = 0.175 \times 35\times10^{-3} = 6.125\times10^{-3} \text{ mol} \times \frac{2\ NaOH}{1\ H_2SO_4}$$

$$\text{mol } NaOH = 12.25\times10^{-3} \text{ mol}$$

$$V = \frac{mol}{M} = \frac{12.25\times10^{-3}}{0.25} = 49\times\cancel{10^{-3}}\ L \times \cancel{10^{3}}$$

$$\boxed{= 49\ mL}$$

Example: A flask contains a solution with an unknown amount of HCl. This solution is titrated with 0.207M NaOH. It takes 4.47mL of the NaOH solution to complete the reaction. what is the mass of the HCl?

$NaOH + HCl \longrightarrow NaCl + H_2O$

$NaOH \rightarrow 0.207M,\ 4.47mL$

$$\text{mol } NaOH = 0.207 \times 4.47\times10^{-3} = 0.925 \text{ mol}\times10^{-3}$$

$$\text{mol } HCl = 0.925 \text{ mol}\times10^{-3}$$

$$\text{mass } HCl = 0.925\times10^{-3}\times 36.45 = 33.7\times10^{-3}\ g$$

$$\boxed{= 0.0337\ g}$$