COMPLEMENTARY CHEMISTRY COURSES
SEMESTER - I
15U1CPCHE1: GENERAL CHEMISTRY
(Common to Physical sciences and Life sciences)

Concept of Equilibrium

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Properties of Acids and Bases

What are the properties of acids and bases?

Acids

Acids taste sour, will change the color of an acid-base indicator, and can be strong or weak electrolytes in aqueous solution.

Bases

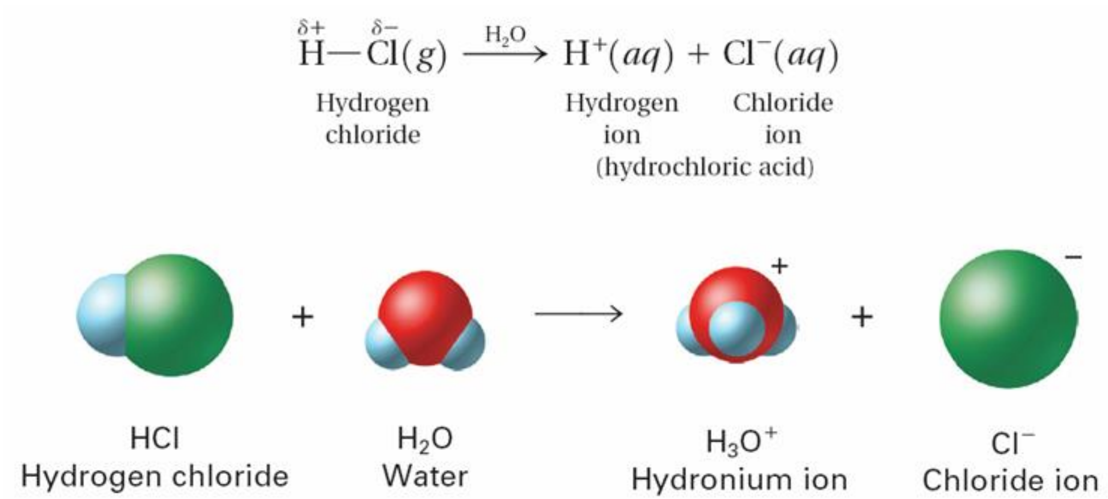
Bases taste bitter, feel slippery, will change the color of an acid-base indicator, and can be strong or weak electrolytes in aqueous solution.

Arrhenius Acids and Bases

How did Arrhenius define an acid and a base?

- Arrhenius said that acids are hydrogen-containing compounds that ionize to yield hydrogen ions (H^+) in aqueous solution. He also said that bases are compounds that ionize to yield hydroxide ions (OH^-) in aqueous solution.

Hydrochloric Acid



- Arrhenius Bases
  - Hydroxide ions are one of the products of the dissolution of an alkali metal in water.

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

Sodium metal	Water	Sodium hydroxide	Hydrogen

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

Sodium hydroxide	Sodium ion	Hydroxide ion

Brønsted-Lowry Acids and Bases

What distinguishes an acid from a base in the Brønsted-Lowry theory?

- The Brønsted-Lowry theory defines an acid as a hydrogen-ion donor, and a base as a hydrogen-ion acceptor.

Conjugate Acids and Bases

- A **conjugate acid** is the particle formed when a base gains a hydrogen ion.
- A **conjugate base** is the particle that remains when an acid has donated a hydrogen ion.

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

$NH_3(aq)$	$H_2O(l)$	$NH_4^+(aq)$	$OH^-(aq)$
Base	Acid	Conjugate acid	Conjugate base

$$HCl(g) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$$

$HCl(g)$	$H_2O(l)$	$H_3O^+(aq)$	$Cl^-(aq)$
Acid	Base	Conjugate acid	Conjugate base

- A **conjugate acid-base pair** consists of two substances related by the loss or gain of a single hydrogen ion.

$$\underset{\text{Base}}{NH_3(aq)} + \underset{\text{Acid}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate acid}}{NH_4^+(aq)} + \underset{\text{Conjugate base}}{OH^-(aq)}$$

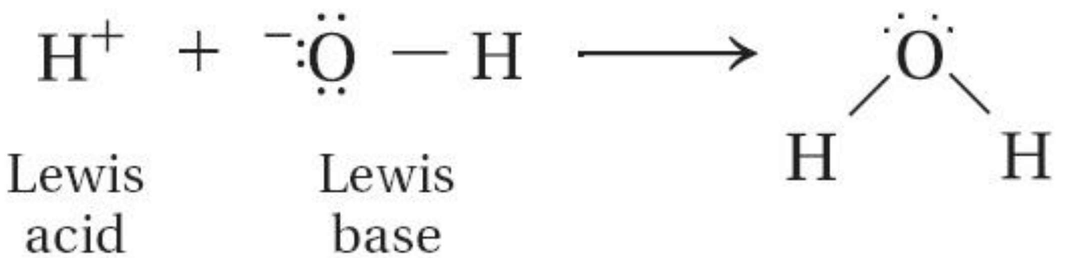
- A substance that can act as both an acid and a base is said to be **amphoteric.**

Lewis Acids and Bases

How did Lewis define an acid and a base?

- Lewis proposed that an acid accepts a pair of electrons during a reaction, while a base donates a pair of electrons

- A **Lewis acid** is a substance that can accept a pair of electrons to form a covalent bond.
- A **Lewis base** is a substance that can donate a pair of electrons to form a covalent bond.



Acid-Base definitions		
Type	Acid	Base
Arrhenius	H+ producer	OH- producer
Brownsted-lowry	H+ donor	H+ acceptor
Lewis	Electron pair acceptor	Electron pair donor

Strong and weak acids and bases

- Strong acid – fully dissociates in water, i.e. almost every molecule breaks up to form H^+ ions
- Some strong acids are…HCl, H_2SO_4, HNO_3

- Weak acid – partially dissociates in water
- Some weak acids are…carboxylic acids such as CH_3COOH, C_2H_5COOH

- Strong base – fully dissociates in water, i.e. almost every molecule breaks up to form OH^- ions
- Some strong bases are….NaOH, compounds which contain OH^- ions or O^{2-} ions

- Weak base – partially dissociates in water
- Some weak bases…nitrogen-containing compounds, such as NH_3

- Strengths can be determined by the acid or base dissociation

Acids

- Act as proton donors
- Electron pair acceptors
- Strong acids dissociate fully in water.
- Weak acids partially dissociate.

- K_a: acid dissociation constant

$$HA + H_2O \rightarrow H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

- Higher K_a values mean stronger acids

Bases

- Act as proton acceptors
- Electron pair donors
- Strong bases dissociate fully in water
- Weak bases partially dissociate
- K_b: base dissociation constant

pH and pOH

- $[H_3O^+]$ can vary greatly $\Rightarrow$ logarithmic scale used
- $pH = -\log [H_3O^+]$
- $pOH = -\log [OH^-]$
- pH > 7 basic
- pH = 7 neutral
- pH < 7 acidic
- Can also express dissociation constants in terms of logs:
  $$pK_a = -\log K_a$$
- $\therefore$ the higher the K_a the lower the pK_a
- Similarly for bases

Ionic product of water

$$2H_2O \rightarrow H_3O^+ + OH^-$$

K_w is the ion-product constant. K_w is the product of the molar concentrations of H_3O^+ and OH^- ions *at a particular temperature*

$$K_w = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

$[H_2O]$ is constant

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \textit{ at } 25°C$$

$[H_3O^+] = [OH^-]$ Solution is neutral

$[H_3O^+] > [OH^-]$ Solution is acidic

$[H_3O^+] < [OH^-]$ Solution is basic

$$pK_w = -\log K_w$$

- Can incorporate pH and pOH

$$K_w = [H_3O^+][OH^-]$$

$$-\log K_w = -\log [H_3O^+][OH^-]$$

$$pK_w = pH + pOH = 14 \quad (\text{at } 25°C)$$

Buffers

- Allow pH to be maintained over small additions of acid or base
- Made up of a weak acid and its conjugate base, e.g.

  $HA + H_2O \rightarrow A^- + H_3O^+$

  acid conjugate base
- The equilibrium will shift to the right on addition of a small amount to base and shifts to the left on addition of small amounts of acid
- Henderson-Hasselbalch equation allows determination of pH in buffer systems:

$$K_a = \frac{[A^-][H_3O^+]}{HA}$$

$$[H_3O^+] = K_a\left(\frac{[HA]}{[A^-]}\right)$$

$$-\log[H_3O^+] = -\log K_a - \log\left(\frac{[HA]}{[A^-]}\right)$$

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

Solubility Equilibria

- Solubility Product Expression and K_{sp}
  - K_{sp} is the ***solubility product constant***
  - Set up like other equilibrium expression
  - General example;

$$MX_n(s) \rightleftarrows M^{n+}(aq) + nX^{-}(aq)$$

$$K_{sp} = [M^{n+}][X^{-}]^n$$

  - Solids and liquids are not included in equilibrium expressions

The Common Ion Effect

- When a compound containing an ion **in common** with an already dissolved substance is added to a solution at equilibrium, the equilibrium shifts to the left.
- This phenomenon is known as the ***common ion effect.***
- Produced by the addition of a second solute.