CHAPTER TWO

ACID-BASE EQUILIBRIA

ACID-BASE CONCEPTS

The Arrhenius Concept of Acids and Bases

Acid-is a substance that releases hydrogen ion(H^+)when dissolved in water

✓Increases the concentration of hydrogen ion when added to water

Example- HCl,HI,HNO_3,$HClO_4$, H_2SO_4

$HCl(g) + H_2O \rightarrow H_3O^+ + Cl^-$

Base-is a substance that release hydroxide ion(OH^-) when added to water

Example-KOH, NaOH, LiOH, Ca(OH^-)

$KOH(s) + H_2O \rightarrow OH^- + K^+$

The general acid base reaction according to Arrhenius is

Acid + base =salt + water

Limitation of Arrhenius

- The theory can be applied only to reaction that occur in water.it doesn’t explain reaction that occur in gaseous phase
- It doesn’t give any reason why some compounds that contain hydrogen such us HCl dissolves in water to form acidic solution where as NH_3 and CH_4 do not
- The theory state that only compounds that contain OH^- ions can be classified as base. It can not explain why other compounds such as Na_2CO_3.KF,etc have the characteristics properties of base.

Example: Based on their dissociations in water solution, classify each of the following compounds as Arrhenius acid, Arrhenius base, or as a compound that cannot be classified as an Arrhenius acid or Arrhenius base

a) H_3PO_4 (l) + H_2O (l)$H_2PO_4^-$ (aq) + H_3O^+ (aq)

b) NaCl (s) + H_2O (l) Na^+ (aq) + Cl^- (aq)

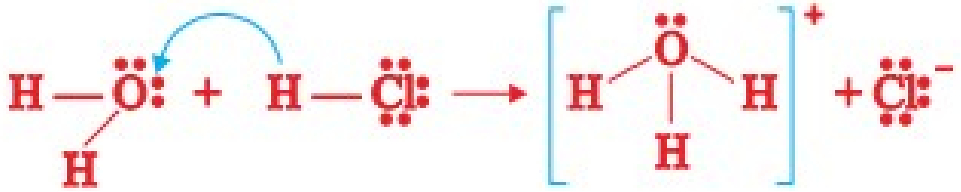
c)$Ca(OH)_2$ + H_2O (l) Ca_2^+ (aq) + $2OH^-$ (aq)

Brønsted-Lowry Concept of Acids and Bases

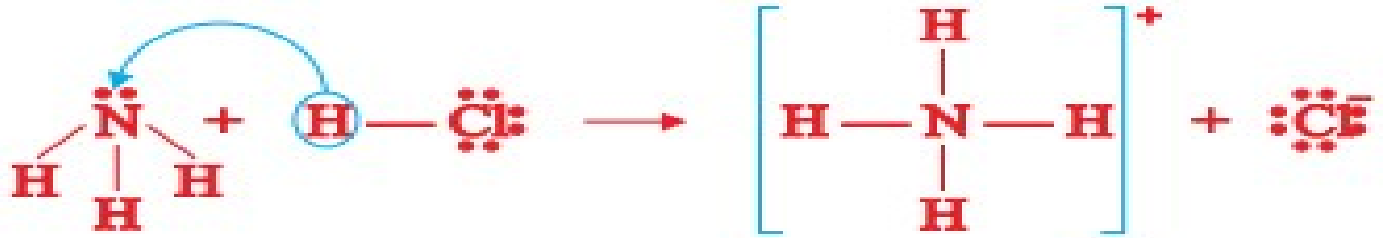
- ✓ Acid is proton donor
- ✓ Base is proton acceptor
- All Arrhenius acids are bronsted-lowery acids but not the reverse
- Arrhenius bases contain hydroxide ion OH- which is bronsted-lowery base. bronsted-lowery acids should contain hydrogen to be donated and base should contain lone pair of electrons to accept a proton (hydrogen ion)

Brønsted-Lowry Concept of Acids and Bases

- Consider the ionization of hydrochloric acid in water:



The Brønsted-Lowry concept is not limited to reactions in aqueous solution. For example, HCl is an acid in liquid ammonia, NH3(l), just as it is in water.



Brønsted-Lowry Concept of Acids and Bases

Every acid has a conjugate base, and every base has a conjugate acid. Thus, for any conjugate acid-base pair:

- **Conjugated acid** is formed when a base accepts a proton (one H^+)

The conjugate acid has one more H and one less minus charge than the base.

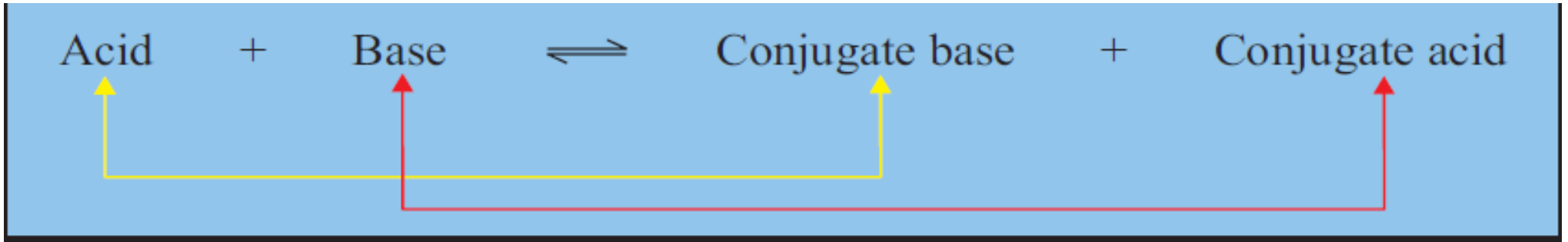
$$Base + H^+ = Conjugated\ acid$$

- **Conjugated base** is formed when an acid loses a proton(one H^+)

$$Acid\text{-}H^+ = conjugate\ base$$

- The conjugate base has one less H and one more minus charge than the acid

Brønsted-Lowry Concept of Acids and Bases



HF	+	H_2O	⇌	F^-	+	H_3O^+
$HCOOH$	+	CN^-	⇌	$HCOO^-$	+	HCN
NH_4^+	+	CO_3^{2-}	⇌	NH_3	+	HCO_3^-
$H_2PO_4^-$	+	OH^-	⇌	HPO_4^{2-}	+	H_2O
H_2SO_4	+	$N_2H_5^+$	⇌	HSO_4^-	+	$N_2H_6^{2+}$
HPO_4^{2-}	+	SO_3^{2-}	⇌	PO_4^{3-}	+	HSO_3^-

Identify the Brønsted-Lowry acids and bases and their respective conjugates in each of the following reactions.

a) $H_2S + NH_3$ $NH_4^+ + HS^-$

b) $OH^- + H_2PO_4^-$ $H_2O + HPO_4^{2-}$

c) $NH_3 + HCO_3^-$ $NH_4^+ + CO_3^{2-}$

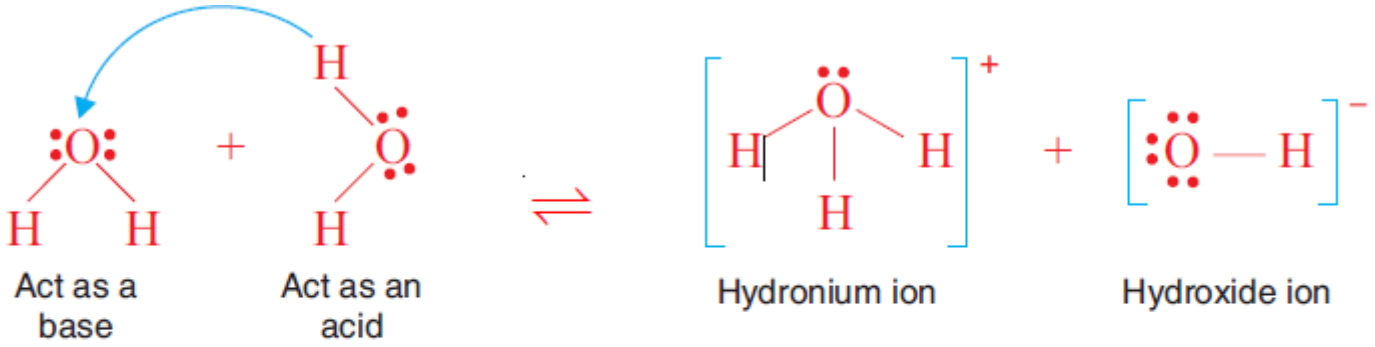
d) $H_3PO_4 + H_2O$ $H_2PO_4^- + H_3O^+$

Strengths of Conjugate Acid-Base Pairs

- ✓The stronger the acid, the weaker is its conjugate base
- ✓Similarly, the stronger the base, the weaker is its conjugate acid
- ✓For example, HCl is a strong acid, and its conjugate base Cl^-,is a weak base
- ✓Acetic acid, CH3COOH,is a weak acid, and its conjugate base, $CH3COO^-$,is a strong base

Auto-ionization of Substances(self ionization)

- Is the interaction of two molecules of the same substances to form positive and negative ions by a proton transfer is known as **Auto-ionization of Substances(self ionization)**



This can be represented by the following equation:

$$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

Amphiprotic Species

- Are substances that behave as either a bronsted acid or bronsted base
- Many molecules and ions gain or lose a proton, under appropriate conditions. Such species are said to be **amphiprotic**

Example of Amphiprotic Species :

$$HCO_3^-(aq) \rightleftharpoons H^+(aq) + CO_3^{2-}(aq)$$

$$HCO_3^-(aq) + H^+(aq) \rightleftharpoons H_2CO_3(aq)$$

Identify the amphiprotic species in each of the following reactions.

- A) $HPO_4^{2-}(aq) + H_3O^+(aq) \rightleftharpoons H_2O(l) + H_2PO_4^-(aq)$

  $HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + PO_4^{3-}(aq)$
- B) $NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

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

Weakness of bronsted-lowery theory

- ✓ It is restricted an acid base concept only to proton transfer
- ✓ It doesn’t explain why species like CO_2,SO_2,Cu^{2+} function as acid even if they don’t contain proton

Lewis Concept of Acids and Bases

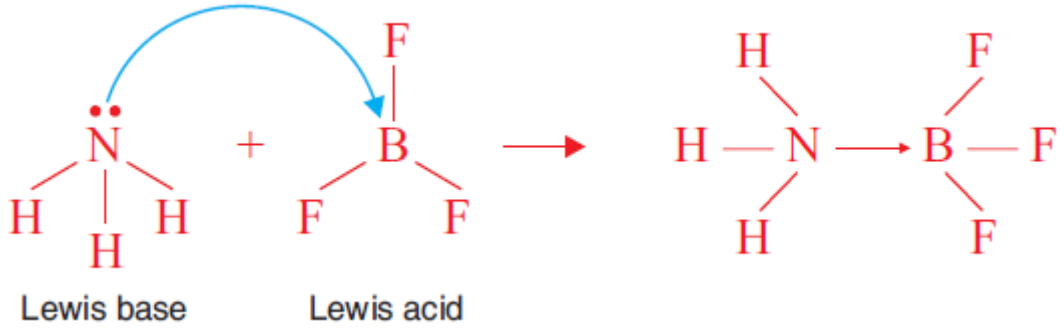
- ✓The Lewis acid-base concept includes many reactions that do not involve proton- transfer reactions
- ✓**Acid** is electron pair acceptor

  -is electron deficient species
- ✓**Base-** is an electron pair donor.

  - it is an electron rich species

Lewis Concept of Acids and Bases

Consider, for example, the reaction between boron trifluoride (BF3) and ammonia to form a compound, BF3 ¨ :NH3.



Identify Lewis acids and Lewis bases in each of the following reactions.

a) $SiCl_4 + 2Cl^- \rightarrow [SiCl_6]^{2-}$

b) $PF_5 + F^- \rightarrow [PF_6]^-$

c) $CO_2 + H_2O \rightarrow H_2CO_3$

IONIC EQUILIBRIA OF WEAK ACIDS AND BASES

Ionization of Water

Water self-ionize by transferring a proton from one water molecules to another and producing a hydronium ion and hydroxide ion

$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

$K_c =$

where, K_c is the equilibrium constant, in terms of concentration

$K_c[H_2O]^2 = K_w = [H_3O^+][OH^-]$

Note that K_w is the product of the molar concentrations of H_3O^+ and OH^-

ions at a particular temperature.

IONIC EQUILIBRIA OF WEAK ACIDS AND BASES

- ✓Since we use $H^+(aq)$ and $H3O^+(aq)$ interchangeably to represent the hydrated proton, the equilibrium constant can also be expressed as

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

- ✓In pure water at 25°C, the concentrations of H^+ and OH^- ions are equal and found to be$[H+] = 1.0 \times 10^{-7}$ M and $[OH^-] = 1.0 \times 10^{-7}$ M at 25°C,

$Kw= [H^+][OH^-] = (1.0 \times 10^{-7})\ (1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$

- ✓the value of K_c in this case Kw, changes as temperature changes

IONIC EQUILIBRIA OF WEAK ACIDS AND BASES

- ✓In any aqueous solution at 25°C, no matter what it contains, the product of $[H^+]$ and $[OH^-]$ must always equal 1.0×10^{-14}

There are three possible situations.

- A neutral solution, where $[H^+] = [OH^-]$.
- An acidic solution, where $[H^+] > [OH^-]$.
- A basic solution, where $[OH^-] > [H^+]$

Calculate $[H^+]$ or $[OH^-]$, as required, for each of the following solutions at 25°C,and state whether the solution is neutral, acidic, or basic.

a)$[OH^-] = 1.0 \times 10^{-4}$ M b) $[H^+] = 1.0 \times 10^{-7}$ M

Continuous.....

pH SCALE

- ✓It is a scale from 0 to 14
- ✓from this pH scale from 0-6 is acid,as pH decrease acidity increase.
- ✓Also pH scale from 8-14 is base , as pH increase base increase.
- ✓The pH of a solution is defined as the negative logarithm of the hydrogen ion concentration (in mol/L):
- ✓$pH = -\log [H3O^+]$ or $pH = -\log [H^+]$
- ✓Note that the negative logarithm gives us positive numbers for pH
- ✓pH decreases as the concentration of H^+ ions increases; in other words, the more acidic the solution, the lower its pH; the more basic the solution, the higher its pH

pH and pOH

- ✓ $pH==-, =10^{-pH}$
- ✓ $pOH==-=10^{-pOH}$
- ✓ pH = pOH=7, neutral solution
- ✓ pH >7 or pOH<7, basic solution
- ✓ pH <7 or pOH>7, acidic solution
- ✓ pH + pOH=14

Calculate:

a) the pH and pOH of a juice solution in which $[H_3O^+]$ is 5.0×10^{-3} M

b) the $[H_3O^+]$ and $[OH^-]$ of human blood at pH = 7.40

Continuous.....

At 40°C, the value of Kw is 2.92 × 10–14. Calculate the [H+] and [OH–] of pure water at 40°C.

Continuous.....

Measures of the Strengths of Acids and Bases in Aqueous Solution

- ✓ **Strength** refers to the degree of ionization or dissociation of acid or bases in water
- ✓ **Strong acid**-ionizes almost completely when dissolved in water

Example-HI,HBr,H_2SO_4,HNO_3

- ✓ **Weak acid** ionize slightly when dissolved in water

Example-HF,H_2CO_3,HCN,HNO_2,CH_3COOH

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Measures of the Strengths of Acids and Bases in Aqueous Solution

Strength of acids are affected by

1.Number of oxygen-as number of oxygen increase strength of acids increases.

Example-$HClO_4$>$HClO_3$>$HClO_2$

2.Electron negativity of the central atom

-as electron negativity increases strength of acid increase

3.Size of the central atom for binary acid.

as the size increases strength of the acid also increases

Example-HI>HBr>HCl>HF

Measures of the Strengths of Acids and Bases in Aqueous Solution

Strong base-ionize completely

-Example KOH,NaOH,LiOH

Weak base-ionize slightly in water

Example NH_3,$Mg(OH)_2$,$Fe(OH)_2$

The Strengths of Acids and Bases in Aqueous Solution is can be measured by:

1. Concentration of hydrogen and hydroxide ions

A solution of a stronger base and acid will contain a larger concentration of hydroxide ions and hydrogen ion than a solution of a weaker base and acid if both solutions are of equal concentration

Measures of the Strengths of Acids and Bases in Aqueous Solution

2) pH and pOH

- ✓The smaller pH value and the higher pOH value, the stronger is the acid.
- ✓The higher pH value and the lower pOH value, the stronger is the base

3) Percent ionization

Percent ionization of acid= **%**

Percent ionization of acid= **%**

- ✓If a higher percentage of the original acid and molecules dissociate, then the acid and base is a strong acid and strong base

Measures of the Strengths of Acids and Bases in Aqueous Solution

4) Dissociation (Ionization)Constants

Acid Dissociation Constant, Ka

The acid-dissociation constant is a quantitative measure of the strength of the acid in a given solvent

$HA(aq) + H_2O(l)\ H_3O^+(aq) + A^-(aq)$

K=

K =

K and are constant and replaced by K_a

$K_a =$,

The larger Ka value ,the strong acid

Measures of the Strengths of Acids and Bases in Aqueous Solution

Base Dissociation Constant, K_b

$$B(aq) + H_2O(l) \; BH^+(aq) + OH^-(aq)$$

$$K=$$

$$K =$$

K and are constant and replaced by k_b

$$K_b=,$$

The larger Kb value ,the strong base

Calculate the pH of a 0.50 M HF solution at 25°C. The ionization of HF is given by $HF(aq) + H_2O(l)\ H_3O^+(aq) + F^-(aq)$

Continuous.....

Continuous.....

Measures of the Strengths of Acids and Bases in Aqueous Solution

Pka=- ,the larger value of Pka, the weaker acid

Pkb=- the larger value of Pkb, the weaker base

End of Part one

Thank you