

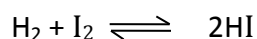
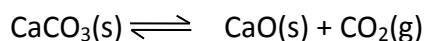
# Chemical Equilibrium

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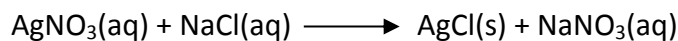
## Reversible and Irreversible Reactions

There are two types of chemical reactions, reversible and irreversible.

**Reversible Reactions:** Reactions which do not always proceed to completion and may be made to proceed in the opposite direction under suitable conditions are called reversible reaction. Reversible reactions occur in forward as well as in backward directions. For example,



**Irreversible Reactions:** Reactions which always proceed to completion in one direction only are called irreversible reactions. Irreversible reactions are unidirectional and do not take place in backward direction. For example,



## Equilibrium

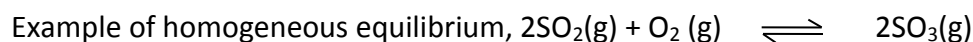
Equilibrium is defined as a state when rate of forward reaction is equal to the rate of backward reaction. This is a dynamic equilibrium. At equilibrium, there is no change in free energy ( $\Delta G = 0$ ).

## Types of Equilibria

- (i) Physical Equilibria e.g.,  $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{O}(\text{g})$
- (ii) Chemical Equilibria e.g.,  $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$
- (iii) Ionic Equilibria e.g.,  $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$

## Homogeneous and Heterogeneous Equilibrium

On the basis of phases the equilibrium is of two types, homogeneous equilibrium and heterogeneous equilibrium. In homogeneous equilibrium all reactants and products are present in same phase, while in heterogeneous equilibrium all reactants and products are present in two or more different phases.



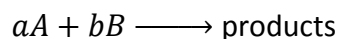
### General Characteristic of Equilibria

- (i) Equilibria can be attained only in a closed vessel.
- (ii) Equilibrium can be attained from either of the directions.
- (iii) An equilibrium is dynamic in nature.
- (iv) Equilibrium in any reaction can be established only if none of the products is allowed to escape.
- (v) Catalyst does not alter state of equilibrium. It only changes the time required to attain equilibrium.

### Law of Mass Action

According to the law of mass action, rate of chemical reaction is proportional to the product of molar masses (active masses) of reactants.

For the reaction,



$$\text{rate} \propto [A]^a [B]^b$$

$$\text{rate} = k[A]^a [B]^b$$

Active masses of solids are taken as unity whereas the active mass of liquids and gases are equal to their molar concentrations.

### Law of Chemical Equilibrium

For a reaction;  $aA + bB \rightleftharpoons xX + yY$

$$K = \frac{k_f}{k_b} = \frac{[X]^x [Y]^y}{[A]^a [B]^b}$$

Where,  $K$  = equilibrium constant,

$$\frac{[X]^x [Y]^y}{[A]^a [B]^b} = \text{concentration quotient (Q)}$$

### Equilibrium Constant

It is the ratio between the molar concentrations of products to that of reactants with each concentration term raised to power equal to its stoichiometric coefficient.

Equilibrium constant in terms of concentration ( $K_c$ )(mol/L)	Equilibrium constant in terms of pressure ( $K_p$ ) when all reactants and products are in gaseous state.
$K_c = \frac{[X]^x [Y]^y}{[A]^a [B]^b}$	$K_p = \frac{[pX]^x [pY]^y}{[pA]^a [pB]^b}$

Units of $K_c = (\text{mol L}^{-1})^{(x+y)-(a+b)}$	Units of $K_p = (\text{atm})^{(x+y)-(a+b)}$
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### Characteristics of Equilibrium Constant

(i) Its value depends on temperature. It has definite value for a given temperature.

It is dimensionless when numbers of moles of reactants are equal to number of moles of products. But when the number of moles of reactants are not equal to number of moles of products then it has specific units.

(ii) The value of equilibrium constant is independent of

- Presence of catalyst
- Initial concentration of reactants
- Direction from which equilibrium is attained

(iii) The magnitude of equilibrium constant is measure of competition of a reversible reaction.

(iv) Larger the value of equilibrium constant more will be the concentration of products

(v) When reaction is reversed value of  $K$  changes to  $K'$

$$K' = \frac{1}{K}$$

(vi) When reaction is divided by  $n$

$$K' = \sqrt[n]{K}$$

(vii) When reaction multiplied by  $n$

$$K' = (K)^n$$

(viii) When reaction is written in two steps each having equilibrium constant as  $K_1$  and  $K_2$  then

$$K = K_1 \times K_2$$

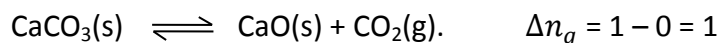
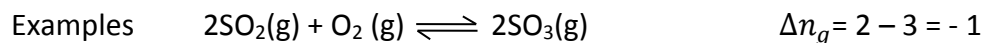
- The value of equilibrium constant for exothermic reaction decreases with increase in temperature whereas the value of equilibrium constant for endothermic reaction increases with increase in temperature.
- If  $\Delta H = 0$ , neither heat is evolved nor absorbed then equilibrium constant remains same at different temperatures.

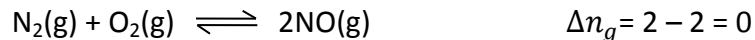
### Relation between $K_p$ and $K_c$

$$K_p = K_c \times RT^{(\Delta n_g)}$$

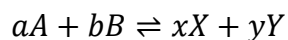
Where,  $\Delta n_g = (n_p - n_r)_{\text{gaseous}}$

$K_p = K_c$  when no. of gaseous reactants = no. of gaseous products.





### Concentration Quotient or Reaction Quotient (Q)



$$Q = \frac{[X]^x[Y]^y}{[A]^a[B]^b}$$

- (i) when  $Q = K$ , the reaction is in equilibrium.
- (ii) When  $Q > K$ , the net reaction proceed in backward direction.
- (iii) When  $Q < K$ , the net reaction proceed in forward direction.

### Standard free energy ( $\Delta G^\circ$ ) at equilibrium constant

$$\Delta G^\circ = -2.303 RT \log K_c$$

When,  $\Delta G^\circ = 0$ , then equilibrium constant  $K = 1$ , the reaction is in equilibrium.

$\Delta G^\circ$  is  $-ve$ , then equilibrium constant ( $K$ ) is  $+ve$ , reaction spontaneous.

$\Delta G^\circ$  is  $+ve$ , then equilibrium constant ( $K$ ) is  $<1$ , reverse reaction is feasible.

### Temperature dependence of Equilibrium Constant (K) (van't Hoff Equation)

$$\log \frac{K_2}{K_1} = \frac{\Delta H^\circ}{2.303R} \left[ \frac{T_2 - T_1}{T_1 T_2} \right]$$

Where,  $\Delta H$  = standard enthalpy change for the reaction.

$K_1$  and  $K_2$  are equilibrium constants at temperature  $T_1$  and  $T_2$ .

### Le-Chatelier's Principle

It states that if a system is in equilibrium and its concentration or temperature or pressure is subjected to change, the equilibrium shifts in a direction so as to undo the effect of change imposed.

### Factors influencing Equilibrium Reaction.

1. Concentration
2. Pressure
3. Inert Gas
4. Catalyst
5. Temperature

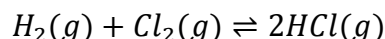
**Effect of change in concentration:** Rate of forward reaction increases when either the concentration of reactants is increased or the concentration of products is decreased.

**Effect of change in temperature:** For exothermic reaction decrease in temperature favours forward reaction and for endothermic reaction increase in temperature favours forward reaction. It also changes the equilibrium constant.

Endothermic Reaction $\Delta + R \rightleftharpoons P$	Exothermic Reaction $R \rightleftharpoons P + \Delta$
Increase in temperature, shifts reaction in forward direction	Increase in temperature, shifts reaction in backward direction.
Decrease in temperature, shifts reaction in backward direction.	Decrease in temperature, shifts reaction in forward direction.

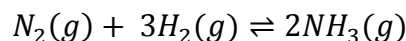
**Effect of change in pressure:** To find the effect of change in pressure the reactions are categorized as follows:

- (a) The reactions in which number of moles of reactants = number of moles of products. For example



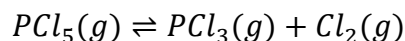
Such reactions are not affected by change in pressure.

- (b) The reactions in which number of moles of reactants > number of moles of products. For example



In such reactions forward reaction is favored at high pressure.

- (c) The reactions in which number of moles of reactants < number of moles of products. For example



In such reactions forward reaction is favoured at low pressure.

### Effect of adding inert gas

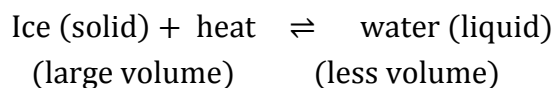
- (a) When inert gas is added at constant volume, the equilibrium is not affected.  
No. of moles of reactants = no. of moles of products.
- (b) When inert gas is added at constant pressure, there are two conditions.
- If the number of gaseous reactant molecules equal the number of gaseous product molecules, there is no affect on equilibrium.
  - If the number of gaseous product molecules > the number of gaseous reactant molecules, then rate of dissociation of reactants increases.

Inert gas	$\Delta n_g = 0$	$\Delta n_g = -ve$	$\Delta n_g = +ve$
At. constant V	No effect	No effect	No effect
At. constant P	No effect	← Backward	Forward →

**Catalyst:** Addition of catalyst decrease the activation and increases the rate of forward and backward rate equaly.

### Application of Le-Chatelier Principle in Melting of Solids

In general when a solid melts, the volume increases. Therefore, in such cases according to Le-Chatelier principle, increase in pressure causes more solidification.



Therefore, during melting of ice increase in pressure or temperature causes increase in rate of melting of ice.

### In Industrial Processes

It predicts the conditions for optimum yield in **Haber's process** (manufacture of  $\text{NH}_3$ ), **Contact process** (manufacture of  $\text{H}_2\text{SO}_4$ ) and **Berkland Eyde process** (manufacture of  $\text{HNO}_3$ ).

## *Ionic Equilibrium*

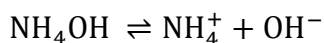
### Electrolytes

Electrolytes are the substances, which conduct electricity in molten state or in aqueous state. Depending upon the extent of ionisation, the electrolytes are divided into groups, i.e. strong electrolyte and weak electrolyte.

**Strong electrolytes** dissociate completely into their ions in solution and such solutions are very good conductors of electricity, e.g.  $\text{NaOH}$ ,  $\text{HCl}$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{KOH}$ ,  $\text{NaCl}$  etc.

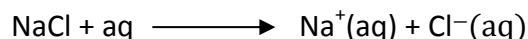
**Weak electrolytes** dissociate to small extent (partially) in the solution and the solution is poor conductor of electricity. e.g.  $\text{CH}_3\text{COOH}$ ,  $\text{H}_3\text{PO}_4$ ,  $\text{H}_3\text{BO}_3$ ,  $\text{NH}_4\text{OH}$ ,  $\text{HCN}$  etc.

- The solution of weak electrolytes contains ions, which are in equilibrium with unionised molecules, for example,



This equilibrium is known as ionic equilibrium and is dynamic in nature.

**Ionisation** : Separation of an electrolyte into its ions either on fusion or dissolution is called ionization.



Usually the term used, for weak electrolyte is dissociation and for strong electrolyte is ionization.

The solution of weak electrolytes contain ions, which are in equilibrium with unionized molecules.

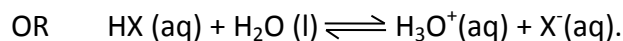
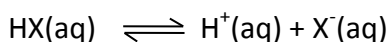
**Degree of Dissociation:** The fraction of total number of molecules undergoing dissociation is called degree of dissociation ( $\alpha$ ).

$$\alpha = \frac{\text{No. of molecules of electrolyte dissociated as ions}}{\text{total no. of molecules of electrolytes dissolved}}$$

Or  $\alpha = \frac{\text{amount of electrolyte dissociated}}{\text{initial amount}}$

**Arrhenius Concept of Acids and Bases:** According to Arrhenius theory, acids are substances that dissociates in water to give hydrogen ions  $\text{H}^+(\text{aq})$  and bases are substances that produce hydroxyl ions  $\text{OH}^-(\text{aq})$ .

The ionization of an acid ( $\text{HX}^-$ ) (aq) can be represented by the following equations:



Similarly, a base molecule like  $\text{MOH}$  ionizes in aqueous solution can be represented as follows:



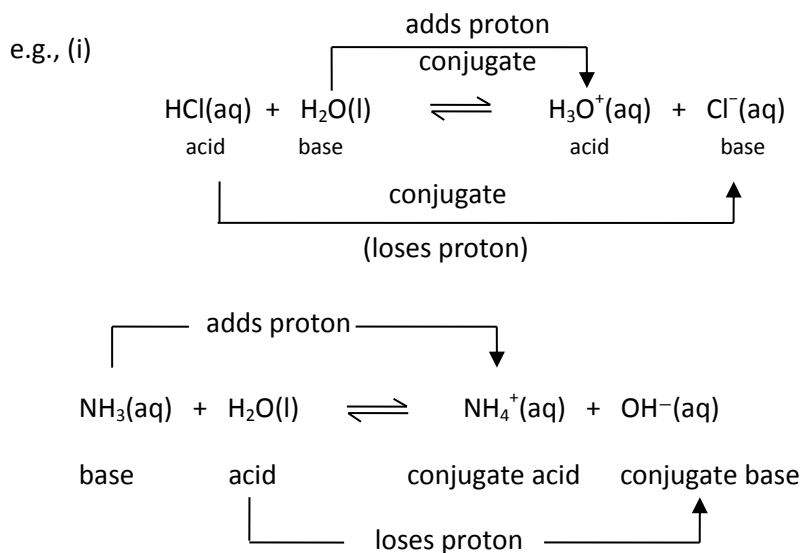
**Limitations:**

- (i) This concept applicable to only aqueous solution.
- (ii) Arrhenius concept does not account for the basicity of substances like, ammonia which do not possess a hydroxyl group.

**Bronsted Lowry Acids and Bases:**

According to Bronsted – Lowry theory, acid is a substance that is capable of donating a hydrogen ion  $\text{H}^+$  and bases are substances capable of accepting a hydrogen ion,  $\text{H}^+$ .

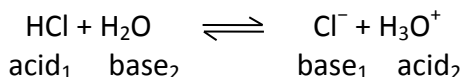
In short, acids are proton donors, bases are proton acceptors.



**Note:** The substances which behave like both as an acid and a base are called amphiprotic e.g.,  $\text{H}_2\text{O}$ ,  $\text{HSO}_3^-$ ,  $\text{HS}^-$  etc.

**Conjugate Acid of the Base:**

The acid formed by a base by accepting proton is called conjugate acid of the base, e.g.,  $\text{H}_3\text{O}^+$  is conjugate acid of  $\text{H}_2\text{O}$ , e.g.,



**Conjugate Base of the Acid:** The base formed by an acid by donating a proton is called conjugate base of that acid, e.g.,  $\text{Cl}^-$  is conjugate base of  $\text{HCl}$  as shown above.

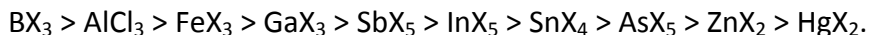
**Lewis Concept:** An acid is a substance which can accept a pair of electrons from any other substance e.g.,  $\text{BF}_3$ ,  $\text{AlCl}_3$  (incomplete octet),  $\text{SnCl}_4$ ,  $\text{SF}_4$  (central atom has vacant d-orbital) or cations  $\text{Fe}^{3+}$ ,  $\text{Cu}^{2+}$  etc.

A base is a substance capable to donating a pair of electrons to any other substance e.g.,  $\text{X}^-$ ,  $\text{OH}^-$ ,  $\text{CN}^-$  or neutral  $\text{R}\ddot{\text{O}}\text{H}$ ,  $\text{N}\ddot{\text{H}}_3$ ,  $\text{H}_2\ddot{\text{O}}$  etc.

**Lewis acid may be any of the following types of substances**

- (i) Molecules having an atom with incomplete octet.
- (ii) Simple cations.
- (iii) Molecules with central atom having empty d-orbitals
- (iv) Molecules with a multiple bond between atoms of different electronegativities.

**Strength of some Lewis acids-**



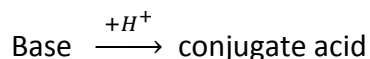
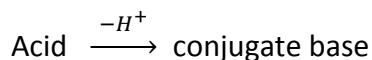
**Comparative Study of Different Concepts of Acids and Bases**

Property	Arrhenius Concept	Bronsted-Lowry Concept	Lewis Concept
Acids	Acid give $\text{H}^+$ in aq solution.	Acids are proton ( $\text{H}^+$ ) donors.	Acids are electron pair accepters.
Bases	Bases gives $\text{OH}^-$ in aq solution.	Bases are proton ( $\text{H}^+$ ) acceptors.	Bases are electron pair donors.
Fails to explain	(i) Behaviour of acids and bases in non aq. Solution. (ii) Acidic character of	Acidic and basic properties of compounds which do not have $\text{H}^+$ ions.	



	salts like $\text{AlCl}_3\text{BF}_3$ and basic character of $\text{NH}_3, \text{PH}_3$ .		
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**Conjugate acid-base pair:** A pair of acid and base, which differ by  $\text{H}^+$  ion only are called conjugate acid base pair.



- Weak acid has strong conjugate base and vice versa.
- The ions/molecules, which can act both as Bronsted acid and base are called amphoteric substances, e.g.  $\text{NH}_3, \text{H}_2\text{O}, \text{CH}_3\text{COOH}$  etc.

The strength of Bronsted acid depends on its tendency to donate a proton, whereas the strength of Bronsted base depends on its tendency to accept a proton.

- The decreasing order of strength of some acids is as follows:  
 $\text{HClO}_4 > \text{H}_2\text{SO}_4 > \text{HBr} > \text{HCl} > \text{HNO}_3 > \text{H}_3\text{O}^+$   
 The order of strength of conjugate bases is opposite of this.

### Ostwald Dilution Law

This law is applicable to weak electrolytes only. According to this law degree of dissociation of weak electrolyte is inversely proportional to the square root of concentration.

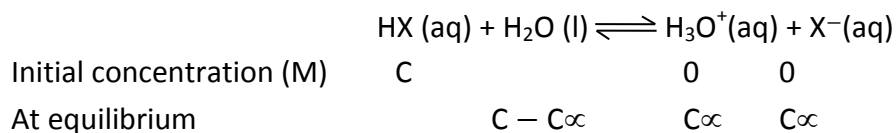
$$\alpha \propto \frac{1}{\sqrt{C}}$$

$$\alpha = \sqrt{\frac{K}{C}}$$

Where,  $\alpha$  = degree of dissociation,  $K$  = dissociation constant,  $c$  = concentration of solution.

### Ionization constants of Weak Acids:

Consider a weak acid  $\text{HX}$  that is partially ionized in the aqueous solution. The equilibrium can be expressed by:



Where  $C$  = initial concentration of the undissociated acid,  $\text{HX}$  at time  $t = 0$  &  $\alpha$  = degree of ionization.

Now,

$$\begin{aligned}K_a &= \frac{[\text{H}_3\text{O}^+][\text{X}^-]}{\text{HX}} \\&= \frac{C\alpha \cdot C\alpha}{C - C\alpha} \\&= \frac{C\alpha^2}{(1 - \alpha)} \\K_a &= \frac{C\alpha^2}{1 - \alpha}\end{aligned}$$

$K_a$  is called the dissociation or ionization constant.

$$\because \alpha \ll 1, \quad 1 - \alpha = 1$$

$$\therefore K_a = C\alpha^2$$

$$\alpha = \sqrt{\frac{K_a}{C}}$$

Higher the value of  $K_a$  stronger will be the acid.

#### Ionisation of Bases:



$$\begin{aligned}K_b &= \frac{[\text{M}^+(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{MOH}(\text{aq})]} \\&= \frac{C\alpha \cdot C\alpha}{C(1 - \alpha)} \\&= \frac{C\alpha^2}{1 - \alpha}\end{aligned}$$

$$\alpha = \sqrt{\frac{K_b}{C}} \quad \text{if } \alpha \ll 1$$

Where  $K_b$  is base dissociation constant. Higher the value of  $K_b$ , stronger will be the base.

**Relative strength of acids and bases:** It is the ratio of strengths of acid bases, for example,

- $\frac{\text{strength of acid } HA_1}{\text{strength of acid } HA_2} = \frac{\alpha_1}{\alpha_2} = \sqrt{\frac{K_{a1}}{K_{a2}}}$
- $\frac{\text{Strength of base } (BOH)_1}{\text{Strength of base } (BOH)_2} = \sqrt{\frac{K_{b1}}{K_{b2}}}$
- **For polyprotic acids**  $K_{a1} > K_{a2} > K_{a3}$  (Where,  $K_{a1}$ ,  $K_{a2}$  and  $K_{a3}$  are first, second and third dissociation constants).

#### The Ionisation Constant of Water and its Ionic Product:

Water being weak electrolyte is slightly ionized as follows:

It is the product of molar concentration of  $H^+$  and  $OH^-$

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

- Ionic product of water is  $10^{-14}$  at  $25^\circ C$ .
- Numerical value of ionic product of water ( $K_w$ ) increases with increase in temperature. ( $\because$  Dissociation of water increases with increase in temperature)
- Numerical value of ionic product of water ( $K_w$ ) is not affected by addition of a salt acid or base if temperature is constant.

### pH Scale

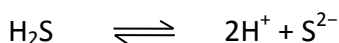
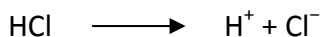
It is used to express and compare the acidic and basic strength of a solution. It is defined as the negative logarithm of hydrogen ion concentration (in moles per litre) present in it.

- $pH = -\log[H^+]$
- $pOH = -\log[OH^-]$
- For weak acids  $pH = \log\left[\frac{K_a}{\alpha}\right]$
- $pH \propto \frac{1}{\text{acidity}}$
- $pH + pOH = 14$  (at 298 K)
- pH of water decreases with increase in temperature ( due to increase in ionic product of water).
- If pH of a solution is zero, then solution is acidic.

### Common Ion Effect

The suppression of dissociation of weak electrolyte in presence of strong electrolyte having common ion is known as common ion effect, e.g. dissociation of  $CH_3COOH$  is suppressed in presence of  $CH_3COONa$ .

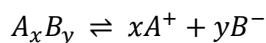
### Application of Common Ion Effect:



$H^+$  being common ion, concentration of  $S^{2-}$  will decrease which is used in precipitation of group 2 radicals for qualitative analysis.

### Solubility Product of a Sparingly Soluble Salt

It is defined as product of molar concentration of its ions in a concentrated solution, each concentration terms raised to the power equal to the number of ions produced on dissociation of one molecule of electrolyte.



$$\therefore K_{sp} = [A^+]^x[B^-]^y$$

- Solubility product of a substance has constant value at a given temperature.
- For precipitation of salt, ionic product  $>K_{sp}$ . No precipitation occurs when  $K_{sp} >$  ionic product.
- A compound with lower solubility product gets precipitated in preference to that having comparatively large solubility product.
- Uses of concept of solubility product are:
  - (i) It is used in purification of common salt. When HCl gas is passed in solution of impure NaCl, pure NaCl precipitates when ionic product of NaCl exceeds its solubility product.
  - (ii) It is used in precipitation of soap by adding NaCl. This process is called salting out of soap.
  - (iii) The concept of solubility product is used in qualitative salt analysis. The cations in inorganic salt arranged in increasing order of their solubility product.

### Buffer Solution

It is a type of solution which resists change in pH. They are said to have reserve acidity and alkalinity and their pH remains same when small amount of acid or base is added to them.

The capacity of a solution to resist change in its pH is called Buffer capacity or buffer index.

$$\text{Buffer capacity} = \frac{\text{no. of moles of acid or base added per litre of buffer solution}}{\text{change in pH}}$$

### Types of Buffers

- Buffers are of two types –acidic buffer and basic buffer.
- **Acidic buffer** are mixture of weak acid and its salt with a strong base in equimolar quantities. For example, mixture of acetic acid and sodium acetate. The pH of acidic buffer is  $<7$ .
- pH of acidic buffer can be calculated by using following equation called as **Henderson's equation**.

$$\text{pH} = \text{p}K_a + \log \left[ \frac{\text{salt}}{\text{acid}} \right]$$

$$\text{or} \quad \text{pH} = \text{p}K_a + \log \left[ \frac{\text{conjugate base}}{\text{acid}} \right]$$

- **Basic buffer** are mixture of weak base and its salt with strong acid in equimolar quantities. For example, mixture of ammonium hydroxide and ammonium chloride. The pH of basic buffer is  $>7$ .

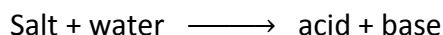
- pH of basic buffer can be calculated by using following equation.

$$pOH = pK_b + \log \left[ \frac{\text{salt}}{\text{base}} \right]$$

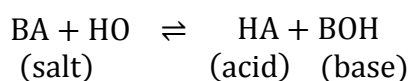
$$\text{or } pH = pK_b + \log \left[ \frac{\text{conjugate acid}}{\text{base}} \right]$$

### Salt Hydrolysis

It is simply the reaction of an ion or ions, with water to yield acidic or alkaline or sometimes neutral solution. In this process acid and base are given back.



Let's consider hydrolysis of salt BA



$$K_h = \frac{[\text{HA}][\text{OH}^-]}{[\text{BA}]}$$

Here  $K_h$  is known as hydrolysis constant. It is related to ionic product of water and dissociation constant of weak acid as follows

$$K_h = \frac{K_w}{K_a}$$

### Degree of Hydrolysis

It increases with dilution (decrease in concentration).

$$\text{Degree of hydrolysis } (h) = \frac{\text{no. of moles of salt hydrolysed}}{\text{total no. of moles of salt taken}}$$

### Comparative Study of Different Types of Salts and their Hydrolysis

Property	Salt of Strong Acid and Strong Base	Salt of Strong Acid and Weak Base	Salt of Weak Acid and Strong Base	Salt of Weak Acid and Weak Base
Hydrolysis	No hydrolysis	Hydrolysis	Hydrolysis	Hydrolysis
Nature of solution	Neutral	Acidic	Basic	Almost neutral
Hydrolysis constant ( $K_h$ )	—	$K_h = K_w/K_b$	$K_h = K_w/K_a$	$K_h = K_w/K_aK_b$
pH	pH=7	$pH = \frac{1}{2}(pK_w - \log C - pK_b)$	$pH = \frac{1}{2}(pK_w + pK_a + \log C)$	$pH = \frac{1}{2}(pK_w + pK_a - pK_b)$
Degree of hydrolysis (h)	—	$h = \sqrt{\frac{K_w}{CK_b}}$	$h = \sqrt{\frac{K_w}{K_aC}}$	$h = \sqrt{\frac{K_w}{K_a \times K_b}}$
Examples	NaCl, $K_2SO_4$	$NH_4Cl$	$CH_3COONa$ , $Na_2CO_3$	$CH_3COONH_4$ , $(NH_4)_2CO_3$ .

- Weaker the acid or base the greater will be the hydrolysis constant of salt and therefore, greater degree of hydrolysis.

### **Acid-Base Indicators**

Acid-base indicators are complex organic molecules that show a visible change in colour in solution when pH changes. Chemically they are either weak organic acids (e.g. phenolphthalein) or weak organic-base (e.g. methyl orange).

The selection of indicator for a given titration depends upon its range. Every indicator works only under specific pH range.

Name of indicator	Range of pH	Colour (in Acidic medium)	Colour (in Basic medium)
Phenolphthalein	8.3-10.0	Pink	Colour less
Methyl red	4.2-6.2	Red	Yellow
Methyl Orange	3.1-4.4	Pink	Yellow
Litmus	5.5-7.5	Red	Blue

Type of Titration	Indicator Used
Strong acid with strong base	Phenolphthalein
Strong acid with weak base	Methyl red, methyl orange
Weak acid with strong base	Phenolphthalein

### **Universal Indicator**

It is a mixture of certain indicators in required proportion which show colour changes over a considerable pH range.