ACID BASE TITRATION

ACID-BASE THEORY:-

ARRHENIUS THEORY:-

✓**Acid** is a substance which gives hydrogen ion (H^+) in an aqueous medium.

✓**Base** is substance which gives hydroxy ions (OH^-) in an aqueous medium.

✓These theory are only applicable for aqueous titration.

LOWRY AND BRONSTED'S THEORY:-

✓**Acid** is a substance which donate a proton (hydrogen ion).

✓**base** is a substance which accept a proton.

LEWIS'S THEORY:-

✓**Acid** is a substance which accept electron pair.

✓**Base** is a substance which donate electron pair.

USANOVICH THEORY:-

- ✓**Acid** is a chemical species that reacts with a base thereby donating cations or accepting anions or electrons.

- ✓**Base** is a chemical species that reacts with an acid thereby donating anions or electrons or accepting cations.

LUX-FLOOD CONCEPT:-

- ✓**Acid** is a substance which accept oxide-ion.
- ✓**Base** is the substance which donate oxide ions.

DISSOCIATION CONSTANT

Dissociation constant of weak acid in aqueous solution

$HA \rightleftharpoons H^+ + A^-$

The dissociation constant of weak acid can be denoted as K_a

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

Now consider the dissociation constant for weak base

$B + H_2O \rightleftharpoons BH^+ + OH^-$

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

Water itself capable of dissociation, below enlist the reaction

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

The equilibirium constant:-

$$K = \frac{[H^+][OH^-]}{[H_2O]}$$

The $[H_2O]$ usually considered to be constant

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

Where k_w is ion product of water.

Dissociation concept of acid and base:-

- A strong acid dissociate (or ionizes) completely in aqueous solution to form hydronium ions (H_3O^+)

$$HCl + H_2O \longrightarrow H_3o + Cl^-$$

- A weak acid does not dissociate (or ionizes) completely in aqueous solution to form hydronium ions (H_3O^+)

$$CH_3COOH + H_2o \longleftrightarrow CH_3COOH + H_3O^+$$

- A strong base dissociate completely in aqueous solution to form hydronium ions (OH^-)

$$NaOH \longrightarrow Na^+ + OH^-$$

- A weak base does not dissociate completely in aqueous solution to form hydronium ions (OH^-)

$$NH_4OH \longrightarrow NH^{4+} + OH^-$$

EXAMPLES OF WEAK/ STRONG ACID:-

TYPE	EXAMPLES
Strong Acids	hydrochloric acid (HCl), sulfuric acid (H_2SO_4), nitric acid (HNO_3)
Weak Acids	acetic acid (CH_3COOH), hydrofluoric acid (HF), oxalic acid $(COOH)_2$
Strong Bases	sodium hydroxide (NaOH), potassium hydroxide (KOH), lithium hydroxide (LiOH)
Weak Bases	Liquid ammonia(NH_3)

ACID BASE TITRATION

An Acid-base titration is a quantitative analysis of acid and bases by which known concentration of an acid or base i.e. **titrant** neutralizes the unknown concentration of an acid or base i.e. **titrand**.

it is based on neutralization reaction between acid and bases that form water by combination of H_3O^+ ions, therefore known as water forming reaction.

$$\textbf{HA(Acid)} + H_2O \rightleftharpoons H_3O^+$$
$$H_3O^+ + OH \rightleftharpoons H_2O$$

The estimation of an alkaline solution by using a standard acidic solution is known as **Acidimetry.**
The estimation of an acidic solution by using a standard alkaline solution is known as **Alkalimetry.**

The titration progress can be monitored by visual indicators, P^H electrodes or both.

TYPE OF TITRATION & NEUTRALIZATION CURVE:-

The mechanism of neutralization can be understood by studying the change in the hydrogen ion concentration during the course of appropriate titration.

1) The equivalence point of an acid-base reaction (the point at which the amounts of acid and of base are just sufficient to cause complete neutralization).

2) The pH of the solution at equivalence point is dependent on the strength of the acid and strength of the base used in the titration.

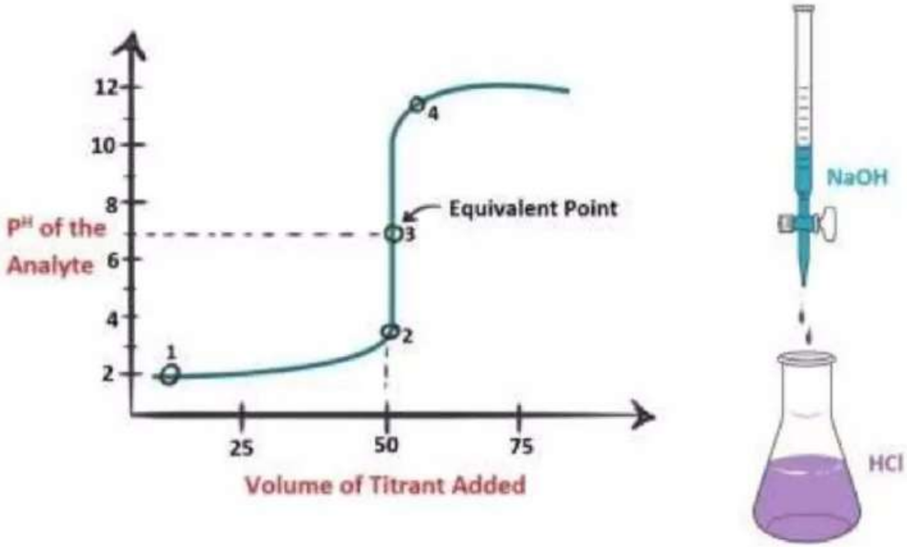
- For strong acid-strong base titration, $pH = 7$ at equivalence point
- For weak acid-strong base titration, $pH > 7$ at equivalence point
- For strong acid-weak base titration, $pH < 7$ at equivalence point

The neutralization curve can be categorized into four classes, which enlist below:-

1) Titration of a strong acid with strong base
2) Titration of a weak acid with a strong base
3) Titration of a strong acid with a weak base
4) Titration of a weak base with a weak acid

1) TITRATION OF A STRONG ACID WITH STRONG BASE:-

Suppose analyte is hydrochloric acid HCl (strong acid) and the titrant is sodium hydroxide NaOH (strong base).
If we start plotting the pH of the Titration curve of a strong acid with a strong base.



Point 1:-

At point 1, No NaOH added yet, so the pH of the analyte is low (because of it predominantly contains H_3O^+ from dissociation of HCl).

$$\mathbf{HCl + H_2O \longrightarrow H_3o + Cl^-}$$

As NaOH is added dropwise, H_3O^+ slowly starts getting consumed by OH^- produced by dissociation of NaOH.
Analyte is still acidic due to predominance of H_3O^+ ions.

Point 2:-

This is the pH recorded at a time point just before complete neutralization takes place.

Equivalent point

Point 3:-

This is the equivalence point where, moles of NaOH is equal to moles of HCl in the analyte.

At this point, H_3O^+ ions are completely neutralized by OH^- ions.

The solution only has salt (NaCl) and water and therefore the pH is neutral i.e. pH 7.

$$HCl + NaOH \xrightarrow{\text{Equivalent point}} NaCl + H_2O$$

Point 4:-

Addition of NaOH continues, pH starts becoming basic because HCl has been completely neutralized and now excess of OH^- ions are present in the solution (from dissociation of NaOH).

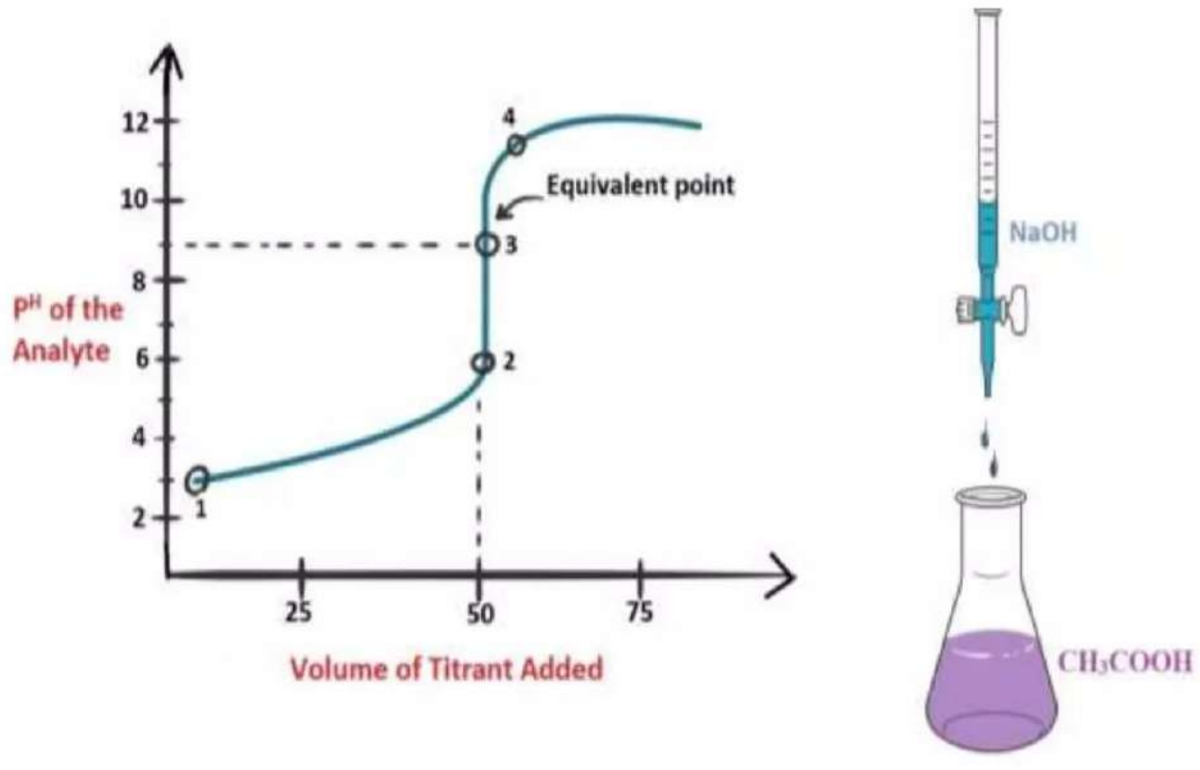
After this NaOH react with indicator and produce color that is end point.

$$NaOH \longrightarrow Na^+ + OH^-$$

2) TITRATION OF A WEAK ACID WITH A STRONG BASE:-

Let's assume our analyte is acetic acid CH_3COOH (weak acid) and the titrant is sodium hydroxide NaOH (strong base).

If we start plotting the pH of the analyte against the volume of NaOH that we are adding from the burette, we will get a titration curve as shown below.



Point 1:-

No NaOH added yet, so the pH of the analyte is low (it predominantly contains H_3O^+ from dissociation of CH_3COOH).

$$\mathbf{CH_3COOH + H_2O \rightleftharpoons CH_3COOH^- + H_3O^+}$$

As NaOH is added drop wise, H_3O^+ slowly starts getting consumed by OH^- produced by dissociation of NaOH).

But analyte is still acidic due to predominance of H_3O^+ ions.

Point 2:-

This is the pH recorded at a time point just before complete neutralization takes place.

Equivalent point

Point 3:-

This is the equivalence point. At this point, moles of NaOH added is equal to moles of CH_3COOH in the analyte. The H_3O^+ are completely neutralized by OH^- ions.

The solution contains only CH_3COONa salt and H_2O.

$$CH_3COOH + NaOH \xrightarrow{\text{Equivalent point}} CH_3COONa + H_2O^+$$

In the case of a weak acid versus a strong base, the pH is not neutral at the equivalence point.

The solution is basic (pH~9) at the equivalence point. Let's reason this out.

$$CH_3COONa \rightleftharpoons CH_3COO^- + Na^+$$

$$\downarrow H_20$$

$$CH_3COOH + OH^-$$

$\searrow$

Make solution basic at equivalent point

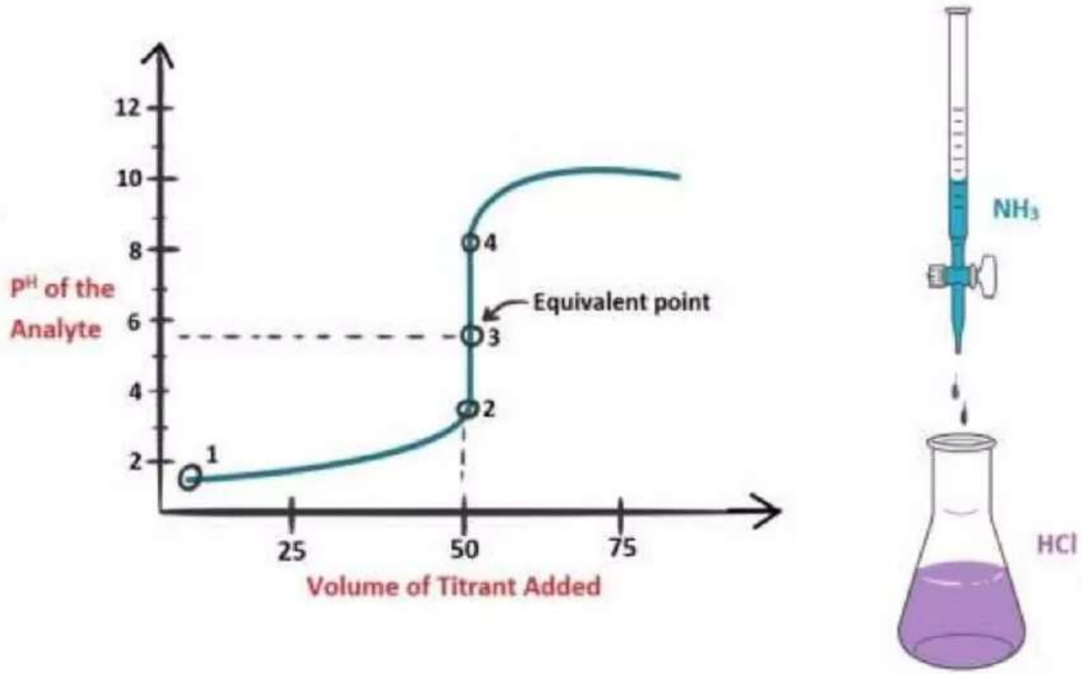
Point 4:-

Addition of NaOH continues, pH starts becoming basic because CH_3COOH has been completely neutralized and now excess of OH^- ions are present in the solution (from dissociation of NaOH).

After this NaOH react with indicator and produce color that is end point.

3) TITRATION OF A STRONG ACID WITH A WEAK BASE:-

Suppose our analyte is hydrochloric acid HCl (strong acid) and the titrant is ammonia NH_3 (weak base).
If we start plotting the pH of the analyte against the volume of NH_3 are adding from the burette, we will get a titration curve as shown below.



Point 1:-

No NH_3, added yet, so the pH of the analyte is low (it predominantly contains H_3O^+ from dissociation of HCl).

$$\mathbf{HCl + H_2O \longrightarrow H_3o + Cl^-}$$

As NH_3 is added drop wise, H_3O^+ slowly starts getting consumed by NH_3. Analyte is still acidic due to predominance of H_3O^+ ions.

$$\mathbf{NH_3 + H_3O \rightleftharpoons NH4^+ + H_2O}$$

Point 2:-

This is the pH recorded at a time point just before complete neutralization takes place.

Point 3:-

This is the equivalence point. At this point, moles of NH_3 added is equal to the moles of HCl in the analyte.

The H_3O^+ ions are completely neutralized by NH_3.

In the case of a weak base versus a strong acid, the pH is not neutral at the equivalence point.

The solution is in fact acidic (pH~5.5) at the equivalence point.
At the equivalence point, the solution only has ammonium ions NH_4 and chloride ions Cl^-. Where ammonium ion NH_4 is the conjugate acid of the weak base NH_3. So NH_4 is a relatively strong acid (weak base NH_3 has a strong conjugate acid), and thus NH_4 will react with H_2O to produce hydronium ions making the solution acidic.

$$NH_4^+ + H_2O \longrightarrow NH_3 + H_3O^+$$

Makes solution acidic at equivalent point

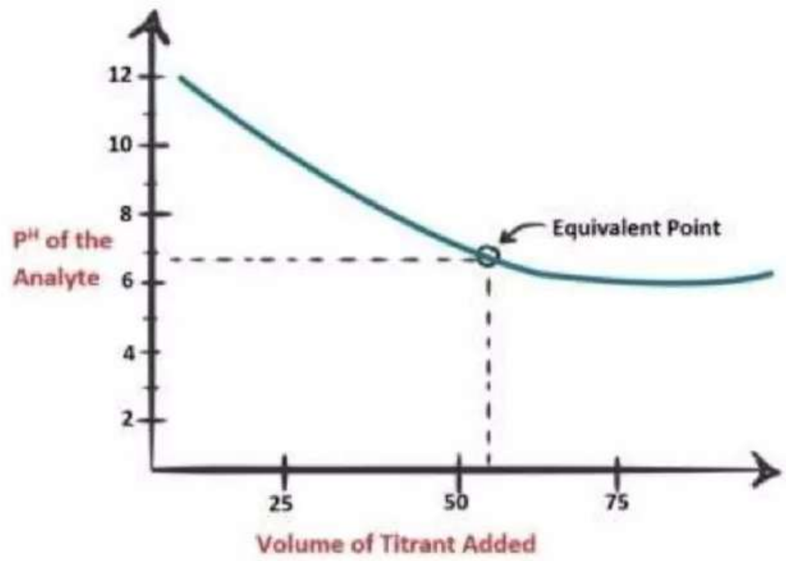
Point 4:-

After the equivalence point, NH_3 addition continues and is in excess, so the pH increases. NH_3 is a weak base so the pH is above 7.

After this NaOH react with indicator and produce color that is end point.

4) TITRATION OF A WEAK BASE WITH A WEAK ACID:-

Suppose our analyte is NH_3 (weak base) and the titrant is acetic acid CH_3COOH (weak acid).
If we start plotting the pH of the analyte against the volume of acetic acid that we are adding from the burette, we will get a titration curve as shown below.



CHOICE OF INDICATOR IN TITRATION

S.NO	TYPE OF TITRATION	PH	INDICATOR
1	Strong acid verses strong base	2-10	Any indicators can be used like Phenolphthaleine Methyl orange Methyl Red
2	Strong acid verses weak base	2-6	Methyl orange Methyl Red
3	Weak acid verses strong base	8-10	Phenolphthaleine Thymol Phthaleine
4	Weak acid verses Weak base	-	Mixed indicators are used like neutral red methylene blue for dilute ammonia or ethanoic acid.

S.No	INDIACATOR NAME	P^H	COLOR OBSERVATION	
			ACID	BASE
1	Thymol blue	1.2-2.8	Red	Yellow
2	Quinaldine Red	1.4-3.2	Colourless	Red
3	Methyl orange	2.9-4.6	Red	Orange
4	Methyl Red	4.2-6.3	Red	Yellow
5	Bromo Thymol blue	6-7.6	Yellow	Blue
6	Phenol Red	6.8-8.6	Yellow	Red
7	Phenolphthaleine	8.3-10	Colourless	Pink
8	Thymol Phthaleine	9.5-10.5	Colourless	Blue

MIXED INDICATOR:-

Mixture of two or more indicators are used to give a sharp end point over a narrow and selected range of P^H.

MIXTURE OF INDICATOR	COMPOSITION	P^H	COLOR CHANGE
Bromocresol Green : Methyl Green	1 Part of 0.1% aqueous solution of sodium salt : 1 Part of 0.2% Water.	4.3	Orange blue green
Bromocresol Green : Chlorophenol Red	1 Part of 0.1% aqueous solution of sodium salt : 1 Part of 0.1% aqueous solution of sodium salt.	6.1	Pale green blue violet
Bromomethyl blue : Neutral Red	1 Part of 0.1% solution in ethanol : 1 Part of 0.1% solution in ethanol.	7.2	Rose Pink green
Bromomethyl blue : Phenol Red	1 Part of 0.1% aqueous solution of sodium salt : 1 Part of 0.1% aqueous solution of sodium salt.	7.5	Yellow violet
Thymol blue : Cresol Red	3 Part of 0.1% aqueous solution of sodium salt : 1 Part of 0.1% aqueous solution of sodium salt.	8.3	Yellow violet

UNIVERSAL INDICATOR:-

A universal indicator is made of a solution of several compound like water, propane 1-ol, Phenolphthaleine, sodium salt, Thymol blue etc. that exihibit several color changes over a wide range of P^H values to indicate the acidity or basicity of solutions.

P^H RANGE	DESCRIPTION	COLOUR
< 3	Strong acid	Red
3–6	Weak acid	Orange or yellow
7	Neutral	Green
8–11	Weak alkali	Blue
> 11	Strong alkali	Violet or Indigo