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Acid base titration

Presented By;-
Mr. Samarpan Mishra (Assistant Professor)
Specialization:- Pharmaceutical Chemistry

Acid Base Titration

Definition:- An **acid–base titration** is a quantitative analytical method in which a solution of known concentration (titrant) is gradually added to a solution of unknown concentration (analyte) until the reaction between the acid and base is complete.

The point at which the reaction is complete is called the **equivalence point**, and it is usually detected using an **acid–base indicator** or a pH meter.

Reaction Examples

- **Strong Acid + Strong Base:- $HCl + NaOH \rightarrow NaCl + H_2O$**
- **Weak Acid + Strong Base:- $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$**
- **Strong Acid + Weak Base:- $HCl + NH_4OH \rightarrow NH_4Cl + H_2O$**

Acid–base indicators

Acid–base indicators are weak acids or weak bases whose **unionized** and **ionized** forms have **different colors**. They help detect the **endpoint** of a titration by a visible color change.

Table of common acid–base indicators with their **pH transition ranges** and **color changes**:

Indicator	pH Range	Color in Acidic Form	Color in Basic Form
Methyl Violet	0.0 – 1.6	Yellow	Violet
Thymol Blue (1st)	1.2 – 2.8	Red	Yellow
Methyl Orange	3.1 – 4.4	Red	Yellow
Methyl Red	4.4 – 6.2	Red	Yellow
Bromocresol Green	3.8 – 5.4	Yellow	Blue
Bromothymol Blue	6.0 – 7.6	Yellow	Blue
Phenol Red	6.4 – 8.0	Yellow	Red
Neutral Red	6.8 – 8.0	Red	Yellow
Phenolphthalein	8.2 – 10.0	Colorless	Pink
Thymol Blue (2nd)	8.0 – 9.6	Yellow	Blue
Alizarin Yellow R	10.1 – 12.0	Yellow	Red

Theories of acid base indicators

1. Ostwald's Theory (Ionization Theory)
2. Quinonoid Theory

1. Ostwald’s Theory (Ionization Theory)

Proposed by:- Wilhelm Ostwald

Concept:-

- Indicators are weak acids (HIn) or weak bases that partially ionize in solution.
- The **unionized form** and the **ionized form** show **different colors**.
- The color change depends on the degree of ionization, which changes with pH.

Example (Phenolphthalein):

$$HIn \rightleftharpoons H^+ + In^-$$

- HIn (unionized) → Colorless
- In^- (ionized) → Pink

Mechanism:

In acidic solution → H^+ from the medium suppresses ionization → color of HIn.

In basic solution → OH^- removes H^+ → more In^- formed → color of In^-.

Limitation:

Cannot explain **why** the color change occurs at a **particular pH range**.

2. Quinonoid Theory

The Quinonoid Theory explains the **color change of acid–base indicators** based on the structural changes in the indicator molecules. It was proposed by **Hantzsch and Ostwald**.

Principle

- Many organic indicators (like phenolphthalein, methyl orange) exist in **two tautomeric forms**:
    - I. **Benzenoid form** (aromatic, colorless or one color)
    - II. **Quinonoid form** (non-aromatic, conjugated, intense color)
- The change from benzenoid ↔ quinonoid form occurs depending on **H^+ ion concentration (pH)** of the solution.
- The two forms have **different absorption spectra**, so the color changes when one form predominates.

Explanation

- In acidic medium, the indicator exists predominantly in one form (e.g., benzenoid).
- In basic medium, ionization changes the structure to the quinonoid form, producing a different color.
- At the **transition pH range**, both forms are present in appreciable amounts, and the observed color is a mixture.

Example: Phenolphthalein

- **Acidic medium** → Benzenoid structure (colorless)
- **Basic medium** → Quinonoid structure (pink)

Limitations

- Not all indicators can be explained by quinonoid theory (only aromatic compounds showing tautomerism).
- Some indicators do not have a quinonoid structure but still show color change.

Classification Of Acid Base Titrations

S.No	Types
1.	Strong acid-strong base
2.	Weak acid-strong base
3.	Strong acid-weak base
4.	Weak acid-weak base

1. Strong Acid – Strong Base Titration

Theory:

- In a titration involving a strong acid and a strong base, both the acid and the base completely ionize in water, producing a high concentration of H^+ and OH^- ions, respectively.
- The reaction results in the formation of water and a neutral salt.
- The pH at the equivalence point is close to 7.0.

Example:

Reaction:- **HCl** (strong acid) + **NaOH** (strong base) → **NaCl** (salt) + **H_2O** (water)

Hydrochloric acid (HCl) reacts with sodium hydroxide (NaOH) to form sodium chloride (NaCl) and water.

Procedure:

1. A burette is used to carefully add the NaOH solution to the HCl solution until the equivalence point is reached, where the moles of H^+ ions equal the moles of OH^-
2. The pH at the equivalence point is near 7.0, indicating neutrality.
3. Indicator: Phenolphthalein, which is colorless in acidic solutions and turns pink in basic solutions.
4. As NaOH is added, the solution remains colorless until the endpoint is reached, where the pink color indicates neutralization.

2. Weak Acid – Strong Base Titration

Theory:

- In this case, the weak acid only partially ionizes, producing a relatively low concentration of H^+ ions, while the strong base completely ionizes, producing a high concentration of OH^-.
- The reaction forms a basic salt, and the pH at the equivalence point will be greater than 7.0 due to the basic nature of the salt.

Example:

Reaction:- CH_3COOH (weak acid) + $NaOH$ (strong base) → CH_3COONa (basic salt) + H_2O (water)

Acetic acid (CH_3COOH) reacts with sodium hydroxide ($NaOH$) to form sodium acetate (CH_3COONa) and water.

Procedure:

1. A burette is used to add NaOH to the acetic acid solution until the equivalence point is reached, where the moles of H^+ ions from the acid equal the moles of OH^- ions from the base.
2. The pH at the equivalence point is greater than 7.0, indicating a slightly basic solution due to sodium acetate.
3. **Indicator**: Phenolphthalein is used, turning from colorless to pink at the endpoint.

3. Strong Acid – Weak Base Titration

Theory:

- The strong acid completely ionizes in water, producing a high concentration of H^+ ions, while the weak base only partially ionizes, resulting in a lower concentration of OH^-
- The reaction forms an acidic salt, and the pH at the equivalence point will be less than 7.0.

Example:

Reaction:- **HCl** (strong acid) + **NH_3** (weak base) → **NH_4Cl** (acidic salt)

Hydrochloric acid (HCl) reacts with ammonia (NH_3) to form ammonium chloride (NH_4Cl).

Procedure:

1. A burette is used to add ammonia solution to the HCl solution until the equivalence point, where the moles of H^+ ions equal the moles of NH_3.
2. The pH at the equivalence point is less than 7.0, indicating a slightly acidic solution due to ammonium chloride.
3. **Indicator**: Methyl orange, which is red in acidic solutions and turns yellow in basic solutions, can be used to signal the endpoint.

4. Weak Acid – Weak Base Titration

Theory:

- In this case, both the weak acid and weak base only partially ionize, producing low concentrations of H^+ and OH^-
- The pH at the equivalence point can vary depending on the relative strengths of the acid and base and may be close to, less than, or greater than 7.0.
- The titration endpoint is less distinct, making precise pH measurement techniques important.

Example:

Reaction:- **CH_3COOH** (weak acid) + **NH_3** (weak base) → **CH_3COONH_4** (salt with variable pH)

Acetic acid (CH_3COOH) reacts with ammonia (NH_3) to produce ammonium acetate (CH_3COONH_4).

Procedure:

1. A burette is used to add the ammonia solution to the acetic acid solution until the equivalence point is reached.
2. The pH at the equivalence point can be close to, less than, or greater than 7.0, depending on the strengths of the acid and base.
3. **Indicator**: Due to the less distinct endpoints, a potentiometric titration using a pH meter is often used.

Neutralization Curve

Definition:- A **neutralization curve** (also called a titration curve) is a graph showing the variation of **pH** of the solution as one reagent (acid or base) is added to another during a titration.

x-axis → volume of titrant added

y-axis → pH of the solution

These curves help us:

1. Understand the progress of a titration
2. Identify the **equivalence point**
3. Select a suitable **indicator**

Example

Strong Acid vs. Strong Base

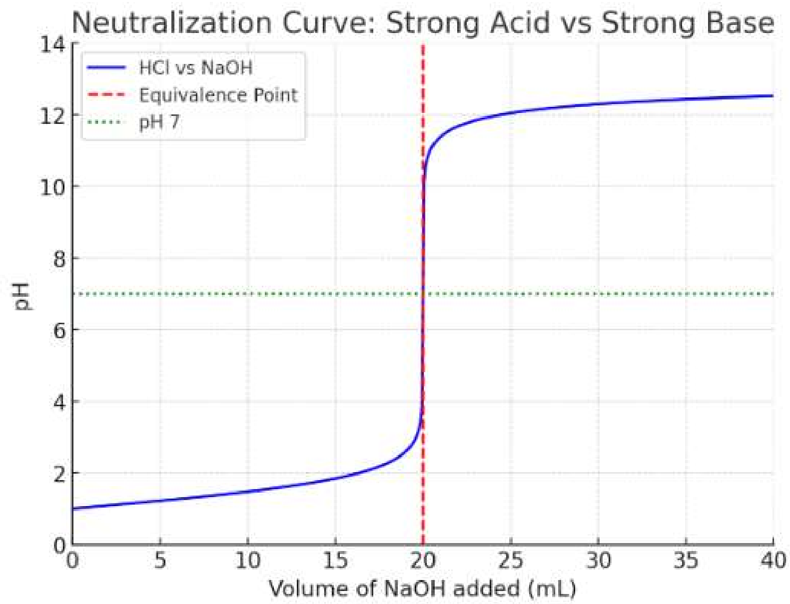
Example: HCl vs. NaOH

Features:

- ❑ Initial pH: very low (~1) for strong acid.
- ❑ pH rises slowly at first, then **steep vertical rise** near the equivalence point (pH ≈ 7).
- ❑ Final pH: high (~13).

Curve:

- ❑ Equivalence point at **pH 7** (neutral).
- ❑ Sharp change → many indicators work (phenolphthalein, methyl orange)



Neutralization Curve For **Strong Acid (Hcl) Vs. Strong Base (Naoh)**

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THANK YOU

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