Indicators

Indicators

- It is an organic dye that detects the end point by a visual change in colour.
- Neutralisation indicators
- Depends on pH
- pH range of an indicator

  Eg: methy orange

- Types of Indicator

  i) One colour indicator : eg. Phenolpthalein

  ii) Two colour indicator : eg. Methyl orange

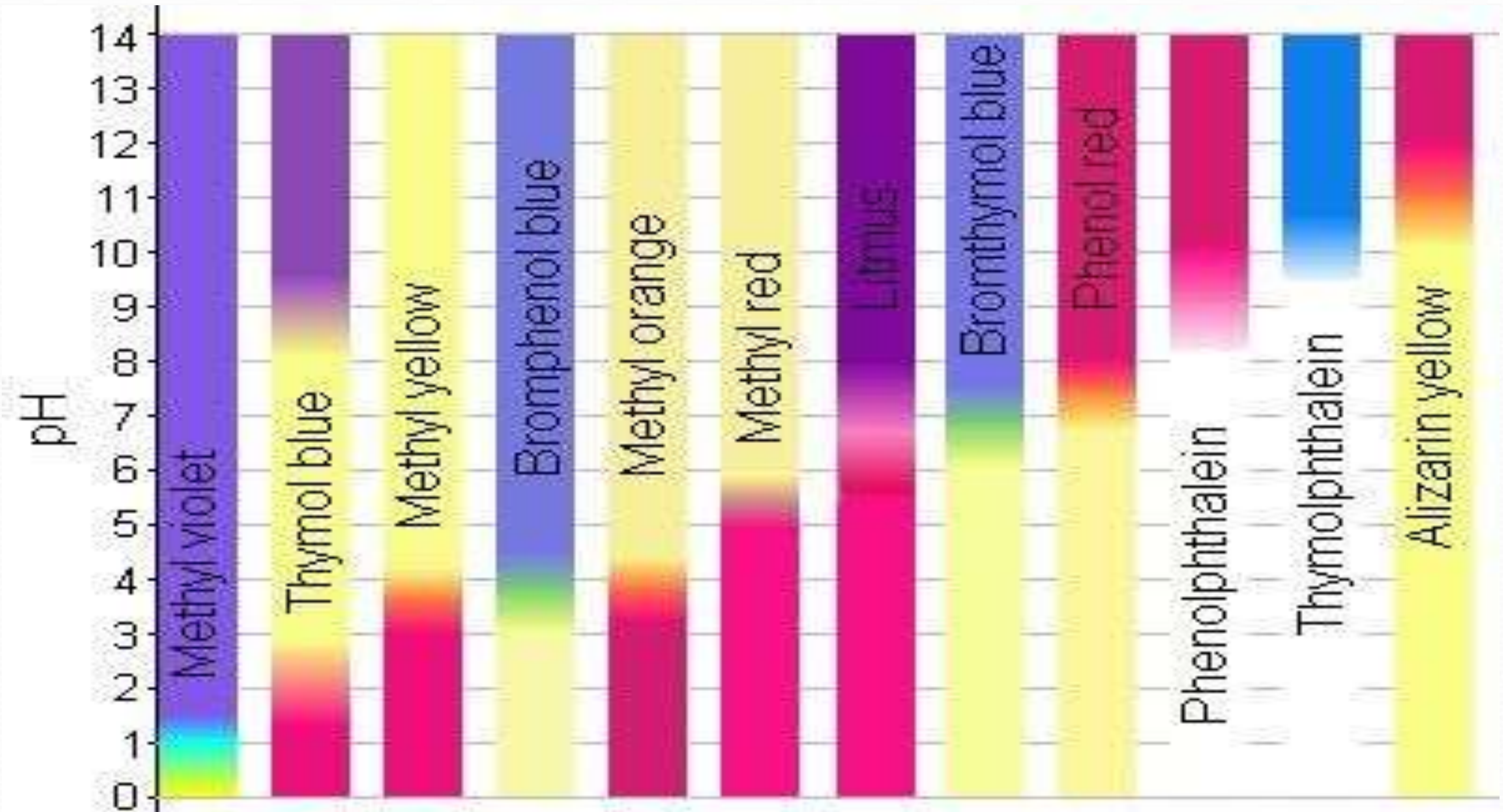
  iii) Mixed Indicator : eg. Neutral red + methylene Blue

  at pH=7.0 ,violet-blue to green from acid to base.

Commonly Used Indicators

Indicator	pH Range	Acid	Base
Thymol Blue	1.2-2.8	red	yellow
Thymol blue	8.0-9.6	yellow	blue
Methyl yellow	2.9-4.0	red	yellow
Methyl orange	3.1-4.4	red	orange
Bromcresol green	4.0-5.6	yellow	blue
Methyl red	4.4-6.2	red	yellow
Bromcresol purple	5.2-6.8	yellow	purple
Bromothymol Blue	6.2-7.8	yellow	blue
Phenol red	6.4-8.0	yellow	red
Cresol purple	7.6-9.2	yellow	purple
Phenolphthalein	8.0-10.0	colorless	red
Thymolphthalein	9.4-10.6	colorless	blue
Alizarin yellow GG	10.0-12.0	colorless	yellow

Colour changes of indicators



- Standard Solutions: strong acids or strong bases because they will react completely
  - Acids: hydrochloric (HCl), perchloric ($HClO_4$), and sulfuric (H_2SO_4)
  - Bases: sodium hydroxide (NaOH), potassium hydroxide (KOH)

Theory of Acid-Base indicators

- Two theories

1. Ostwald's theory
2. Quinonoid theory

Wilhelm Ostwald Theory of Indicators

- Many substances display colors that depend on the pH of the solutions in which they are dissolved.
- An acid/base indicator is a weak organic acid or a weak organic base whose undissociated form differs in color from its conjugate form.

e.g., the behavior of an acid-type indicator, HIn, is described by the equilibrium

$$\underset{\text{acid color}}{HIn + H_2O} \rightleftharpoons \underset{\text{base color}}{In^- + H_3O^+}$$

The equilibrium for a base-type indicator, In, is

$$\underset{\text{base color}}{In + H_2O} \rightleftharpoons \underset{\text{acid color}}{InH^+ + OH^-}$$

The equilibrium-constant expression for the dissociation of an acid-type indicator takes the form

$$K_a = \frac{[H_3O^+][In^-]}{[HIn]}$$

Rearranging leads to

$$[H_3O^+] = K_a \frac{[HIn]}{[In^-]}$$

The hydronium ion concentration determines the ratio of the acid to the conjugate base form of the indicator and thus determines the color developed by the solution.

acid color when

$$\frac{[HIn]}{[In^{-}]} \geq \frac{10}{1}$$

and its base color when

$$\frac{[HIn]}{[In^{-}]} \leq \frac{1}{10}$$

For the full acid color,

$$[H_3O^+] = 10K_a$$

and similarly for the full base color,

$$[H_3O^+] = 0.1K_a$$

To obtain the indicator pH range, we take the negative logarithms of the two expression:

$$\text{pH (acid color)} = -\log (10K_a) = pK_a + 1$$

$$\text{pH (basic color)} = -\log (0.1K_a) = pK_a - 1$$

$$\text{indicator pH range} = pK_a \pm 1$$

Wilhelm Ostwald Theory of Indicators and Range

- The colour changes shows that undissociate acid H-In or base In-OH have diff. colour from that of its ion.
- All indicators in general use are very weak organic acids or bases.
- Equation :– pH = pK In - Log [Acid form] / [Base form]

$$pH = pK_{In} - \log \frac{[\text{Acid form}]}{[\text{Base form}]}$$

Modern Quinoid Theory

- An Acid-Base indicator is a dynamic equilibrium mixture of two alternative tautomeri forms. Ordinarily one form is benzenoid while the other is quinoid.
- The two forms have different colours.
- Out of these one form exists in acidic solution while the other in alkaline solution.
- Change in pH causes the transition of benzenoid form to quinoid form and vice-versa and consequently a change in colour.