Experiments Based On pH Change

pH SCALE

In order to express the hydronium ion (H_3O^+) concentration in a solution P.L. Sorensen (1909) devised a logarithmic scale. This scale is known as pH scale. The **pH of a solution** is defined as the negative logarithm of hydronium ion concentration in moles per litre.

$$pH = -\log[H_3O^+]$$
$$= log\frac{1}{H_3O^+}$$

Acidity, Alkalinity, Neutrality of Solutions

neutral solution : $H^+ = OH^- = 10^{-7}$M ; pH = – 7
acidic solution : $H^+ > OH^-$, $H^+ > 10^{-7}$M, pH < 7.
basic solution: $OH^- > H^+$, $H^+ < 10^{-7}$M, pH > 7. (also called an alkaline solution)

Strong and Weak Acids and Bases

strong acid—an acid that is a strong electrolyte and has a pH < 3.
For example, H_2SO_4, HCl, HBr, HI, HNO_3

weak acid—an acid that is a weak electrolyte or an ionic compound that partially reacts with water to form hydrogen ions in aqueous solution. It will have a pH greater than 3 but less than 7.
For example, H_2S, H_3PO_4, CH_3COOH, H_2CO_3.

strong base—a hydroxide that is a strong electrolyte and has a pH >11.
For example, $NaOH$, KOH, $Ba(OH)_2$.

weak base—a hydroxide that is a weak electrolyte or a compound that partially reacts with water to form hydroxide ions in aqueous solution. Its pH will be less than 11 but greater than 7. For example, carbonates, bicarbonates, ammonia (ammonium hydroxide), phosphates.

salt—an ionic compound produced by reacting an acid and a base. It will have a pH close to 7.

Ioconic Product Of Water

Pure water is very weakly ionised. So there is an equilibrium between ionised and unionised molecules.

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

The equilibrium constant, $K = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$

As concentration of water is very large so it practically remains constant.

Thus, $K[H_2O]^2 = [H_3O^+][OH^-]$

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

where, K_w is the ionic product of water. The value of K_w at 298 K is 1.0×10^{-14} mol^2 L^{-2}.

$\therefore \quad 1.0 \times 10^{-14} = [H_3O^+][OH^-]$

As one mole of water gives one mole of H_3O^+ and one mole of OH^- ions.

$\therefore \quad [H_3O^+] = [OH^-]$

Thus, $1.0 \times 10^{-14} = [H_3O^+]^2$

or $[H_3O^+] = 1.0 \times 10^{-7}$ mol L^{-1}

Now $pH = -\log[H_3O^+]$

$= -\log(1.0 \times 10^{-7}) = 7$

In general, it has been observed that at room temperature all neutral solutions have pH equal to 7, all acidic solutions have pH less than 7 and all basic solutions have pH more than 7.

Common Ion Effect

Common ion effect may be defined as the supression of degree of dissociation of a weak electrolyte by the addition of a small amount of some strong electrolyte having a common ion with that of the weak electrolyte.
Consider for example, NH_4OH which is a weak electrolyte and there is an equilibrium between unionised molecules and its ions.

$$NH_4OH \rightleftharpoons NH_4^+ + OH^-$$

When NH_4Cl, a strong electrolyte, is added to it, NH_4Cl ionises as

$$NH_4Cl \longrightarrow NH_4^+ + Cl^-$$

Due to the presence of common NH_4^+ ions the equilibrium (6.1) shifts in the backward direction and degree of dissociation of NH_4OH is supressed. So the concentration of OH^- ions decreases and hence concentration of H_3O^+ ions increases. Thus, pH of the solution is lowered.

Similarly consider acetic acid, a weak electrolyte

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$

When sodium acetate, a strong electrolyte is added to it, CHgCOONa ionises as :

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

Due to the presence of common CH_3COO^- ions the equilibrium (6.2) shifts in the backward direction and so concentration of H_3O^+ ions decreases and hence that of OH- ions in¬creases. Therefore, pH of solution increases.

Salts when dissolved in water may undergo hydrolysis producing acidic or basic solutions. **Hydrolysis** of salts may be defined as the interaction of ions of the salt with water producing acidic or basic solution.

Hydrolysis of salts of strong bases and weak acids produces alkaline solution on hydrolysis. For example, the aqueous solution of sodium acetate is alkaline due to the presence of excess hydroxyl ions in the solution.

$$\underset{\text{Sodium acetate}}{CH_3COONa + H_2O} \rightleftharpoons \underset{\substack{\text{Acetic acid}\\ \text{(weakly ionised)}}}{CH_3COOH} + Na^+ + OH^-$$

Hydrolysis of salts of strong acids and weak bases produces acidic solution due to the presence of excess hydronium ions in the solution. For example, an aqueous solution of ammonium chloride is acidic in nature.

$$NH_4Cl + 2H_2O \rightleftharpoons NH_4OH + H_3O^+ + Cl^-$$

Ammonium chloride (weakly ionised)

Hydrolysis of salts of weak acids and weak bases gives almost neutral solutions. For example,

$$CH_3COONH_4 + H_2O \rightleftharpoons CH_3COOH + NH_4OH$$

Ammonium acetate

Salts of strong acids and strong bases donot undergo hydrolysis and hence their aqueous solutions are neutral.

Table 6.1. Colour Changes and pH range of Certain Indicators

S.No.	Indicator	pH range	Colour in acidic medium	Colour in alkaline medium
1.	Thymol blue	1.2-2.8	Red	Yellow
2.	Methyl yellow	2.9-4.0	Red	Yellow
3.	Bromophenol blue	3.0-4.6	Yellow	Blue
4.	Congo red	3.0-5.0	Violet	Red
5.	Methyl orange	3.1-4.4	Red	Yellow
6.	Methyl red	4.2-6.3	Red	Yellow
7.	Phenol red	6.8-8.4	Yellow	Red
8.	Phenolphthalein	8.3-10.0	Colourless	Pink
9.	Thymolphthalein	9.4-10.5	Colourless	Blue

Universal Indicator

A universal indicator is prepared by mixing a number of common indicators together so that the mixture obtained can pass through a series of colour changes over a much wider pH range. For example, one such mixture which may show various colours at different pH is as given in Table 6.2.

Table 6.2. Colours of Universal Indicator at Different pH Values

pH	**Colour**

3.0	Red
5.0	Orange red
5.5	Orange
6.0	Orange-yellow
7.0-7.5	Greenish-yellow
8.0	Green
9.5	Blue
10.0	Violet

Such mixtures are commonly known as universal indicators. Universal indicators are available commercially as solutions and as test papers. A pH paper is a strip of paper which is prepared by dipping the strip in the solutions of different indicators and then drying them.
pH paper can be used to find the approximate pH of any solution. The pH paper is dipped in a given sample of the solution, the colour developed in the paper is compared with the colour chart and approximate pH of the solution can be predicted. A pH paper is shown in Fig. 6.1.

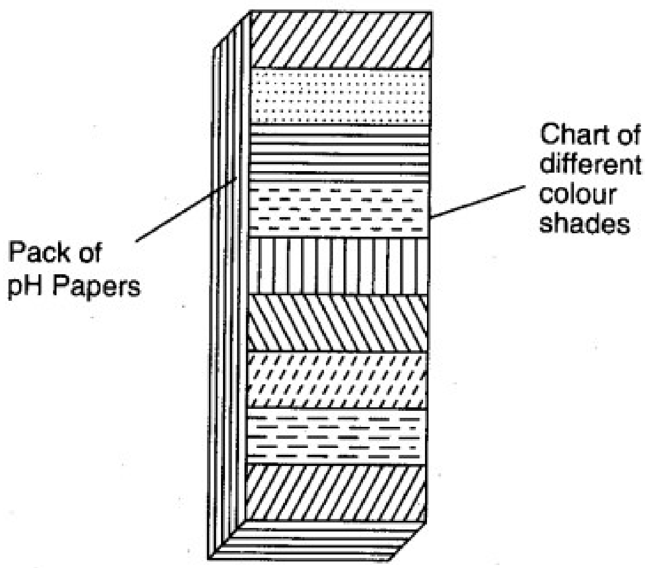

Fig. 6.1. Universal indicator as test papers.