Textbook Chapter No.

03 Ionic Equilibria

Subtopics

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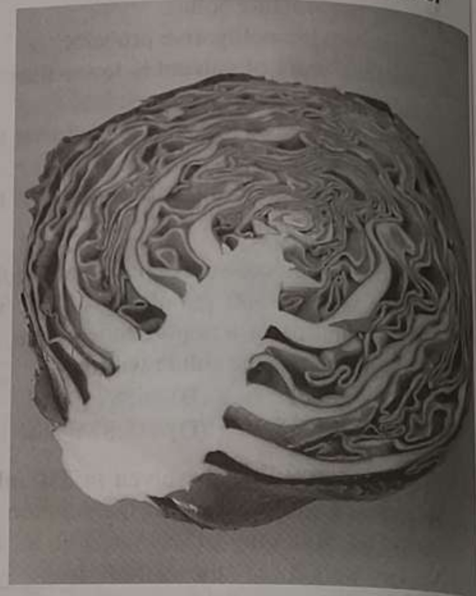
Red cabbage as a pH indicator



The red cabbage plant changes its colour according to the acidity or basicity of the soil in which it is cultivated. In acidic soils, the leaves are reddish; in neutral soils the leaves are purple, while in alkaline soils the leaves are greenish-yellow coloured. This is because they contain chemicals from the naturally coloured anthocyanin family of compounds.
The juice of red cabbage can be used as a home-made pH indicator. The juice is red, pink or magenta in acids, (pH < 7), purple in neutral solutions (pH ~ 7), and ranges from blue to greenish yellow in alkaline solutions (pH > 7).

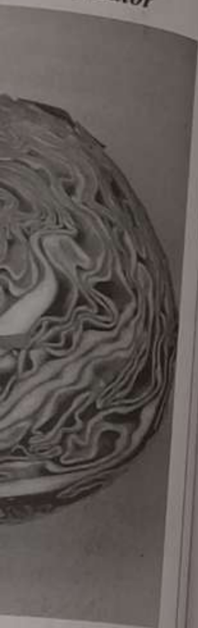
Quick Review

➢ **Types of electrolytes:**

Electrolytes

- **Strong electrolytes** — **e.g.** HCl, NaOH
- **Weak electrolytes** — **e.g.** CH_3COOH

pH indicator



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- **Various theories of acids and bases:**

Theory	Acid	Base
Arrhenius theory	A substance that contains hydrogen and produces H^+ ions in aqueous solution.	A substance that contains OH group and produces OH^- ions in aqueous solution.
Bronsted-Lowry theory	Any substance that can donate a proton (H^+) i.e., proton donor.	Any substance that can accept a proton i.e., proton acceptor.
Lewis theory	Any species that can accept a share in an electron pair.	Any species that can donate a share in an electron pair.

- **Classification of acids and bases:**

Acids and Bases

Based on extent of dissociation

- **Strong acids** — e.g. HNO_3
- **Weak acids** — e.g. HF
- **Strong bases** — e.g. KOH
- **Weak bases** — e.g. $Cu(OH)_2$

- **pH of solutions:**

Acidic solutions	$[H^+] > 1.0 \times 10^{-7}$ M	pH < 7.00
Basic solutions	$[H^+] < 1.0 \times 10^{-7}$ M	pH > 7.00
Neutral solutions	$[H^+] = 1.0 \times 10^{-7}$ M	pH = 7.00

- **Types of salts:**

Acidity and basicity of salts

Salts of strong acid and strong base
- Solution will be neutral.

e.g. NaCl, KNO_3, Na_2SO_4

Salts of strong acid and weak base
- Solution will be acidic.

e.g. NH_4Cl, NH_4NO_3, $(NH_4)_2SO_4$

Salts of weak acid and strong base
- Solution will be basic.

e.g. CH_3COONa, KCN, Na_2CO_3

Salts of weak acid and weak base
- Solution may be acidic, basic or neutral depending on K_a and K_b values of weak acid and weak base respectively.

e.g. CH_3COONH_4, NH_4CN, $(NH_4)_2CO_3$

- If $K_a > K_b$, solution is acidic.
  e.g. NH_4F
- If $K_a < K_b$, solution is basic.
  e.g. NH_4CN
- If $K_a = K_b$, solution is neutral.
  e.g. CH_3COONH_4

➢ **Buffer solutions:**

Buffer solutions

A buffer solution resists drastic changes in pH upon the addition of a small amount of either an acid or a base.

Types

Acidic buffer solution

- A solution containing a weak acid and its salt with a strong base.
  e.g.
  CH_3COOH and CH_3COONa in water.
- Henderson-Hasselbalch equation is:
  $$pH = pK_a + \log_{10}\frac{[salt]}{[acid]}$$

Basic buffer solution

- A solution containing a weak base and its salt with a strong acid.
  e.g.
  NH_4OH and NH_4Cl in water.
- Henderson-Hasselbalch equation is:
  $$pOH = pK_b + \log_{10}\frac{[salt]}{[base]}$$

Properties

- When a small amount of strong acid (or strong base) is added to a buffer solution, there is no significant change in the value of pH.
- The dilution of a buffer solution will not change its pH.
- The pH of a buffer solution does not change with time.

➢ **Different expressions for solubility product:**

Type of electrolyte	Example	Equation	K_{sp} expression	Molar solubi
AB (1 : 1 type salt)	AgCl	$AgCl \rightleftharpoons Ag^+ + Cl^-$	$K_{sp} = [Ag^+][Cl^-]$ $\mathbf{K_{sp} = S^2}$	$S = \sqrt{K_{sp}}$
AB_2 (1 : 2 type salt)	$PbCl_2$	$PbCl_2 \rightleftharpoons Pb^{2+} + 2Cl^-$	$K_{sp} = [Pb^{2+}][Cl^-]^2$ $K_{sp} = [S][2S]^2$ $\mathbf{K_{sp} = 4S^3}$	$S = \sqrt[3]{\frac{K_{sp}}{4}}$
A_2B (2 : 1 type salt)	Ag_2CrO_4	$Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^{2-}$	$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$ $K_{sp} = [2S]^2[S]$ $\mathbf{K_{sp} = 4S^3}$	$S = \sqrt[3]{\frac{K_{sp}}{4}}$
AB_3 (1 : 3 type salt)	$AlCl_3$	$AlCl_3 \rightleftharpoons Al^{3+} + 3Cl^-$	$K_{sp} = [Al^{3+}][3Cl^-]$ $K_{sp} = [S][3S]^3$ $\mathbf{K_{sp} = 27S^4}$	$S = \sqrt[4]{\frac{K_{sp}}{27}}$
A_2B_3 (2 : 3 type salt)	As_2S_3	$As_2S_3 \rightleftharpoons 2As^{3+} + 3S^{2-}$	$K_{sp} = [As^{3+}]^2[S^{2-}]^3$ $K_{sp} = [2S]^2[3S]^3$ $K_{sp} = 4S^2 \times 27S^3$ $\mathbf{K_{sp} = 108S^5}$	$S = \sqrt[5]{\frac{K_{sp}}{108}}$

➢ **Condition for the formation of a precipitate:**

Condition	Type of solution	Result
Ionic product = K_{sp}	Saturated solution	No precipitation
Ionic product > K_{sp}	Supersaturated solution	Precipitation
Ionic product < K_{sp}	Unsaturated solution	No precipitation

Formulae

1. **Degree of dissociation (α):**

$$\alpha = \frac{\text{Number of moles dissociated}}{\text{Total number of moles}}$$

2. **Ostwald's dilution law:**

$$\alpha \propto \frac{1}{\sqrt{C}} \quad \textbf{OR} \quad \alpha \propto \sqrt{V}$$

where c = concentration in mol dm^{-3}
V = volume in dm^3

3. **Acid dissociation constant (K_a):**

For weak acid HA, $K_a = \frac{[H^+][A^-]}{[HA]}$

$K_a = \alpha^2/V$ and $K_a = \alpha^2 c$

4. **Base dissociation constant (K_b):**

For weak base BOH, $K_b = \frac{[B^+][OH^-]}{[BOH]}$

$K_b = \alpha^2/V$ and $K_b = \alpha^2 c$

5. **Ionic product of water (K_w):**

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

6. **pH of solution:**

$pH = -\log_{10}[H_3O^+]$

7. **pOH of solution:**

$pOH = -\log_{10}[OH^-]$

8. **Relation between pH and pOH:**

$pH + pOH = 14$

9. **Henderson-Hasselbalch equation:**

Acidic buffer:

$$pH = pK_a + \log_{10}\frac{[\text{salt}]}{[\text{acid}]}$$

$pK_a = -\log_{10} K_a$

Basic buffer:

$$pOH = pK_b + \log_{10}\frac{[\text{salt}]}{[\text{base}]}$$

$pK_b = -\log_{10} K_b$

10. **Solubility product (K_{sp}):**

For salt B_xA_y:

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

11. **Molar solubility, S (mol/L):**

$$S = \frac{\text{Solubility in g/L}}{\text{Molar mass in g/mol}}$$

12. **Relation between K_{sp} and S:**

$K_{sp} = x^x y^y S^{x+y}$