Precise

CHEMISTRY

Vol. **I**

BASED ON LATEST BOARD PAPER PATTERN

#itna hi kaafi hain

Std. XII
Science

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Precise
CHEMISTRY (Vol. I)
Std. XII

Salient Features

☞ Written as per **Latest Board Paper Pattern**

☞ Subtopic-wise segregation for powerful concept building

☞ Complete coverage of Textual Exercise Questions, Intext Questions and Numericals

☞ Marks provided to the Questions as per relevant weightage wherever deemed necessary

☞ Relevant Previous Years' Board Questions:
   March 2013 to July 2023

☞ Each chapter contains:
   - 'Quick Review' of the chapter for quick revision
   - 'Important Formulae' and 'Solved Examples' to cover numerical aspect in detail
   - 'Exercise' to provide Theory questions, Numericals and MCQs for practice
   - Includes selective questions from NCERT textbook for practice

☞ Includes Important Feature to elucidate concept: *Reading Between the Lines*

☞ Q.R. codes provide:
   - The Video/pdf links boosting conceptual retention

☞ Includes Board Question Paper of February 2024 (Solution in pdf format through QR code)

Precise Chemistry Vol. I, Std. XII Sci. is intended for every Maharashtra State Board aspirant of Std. XII, Science. The scope, sequence, and level of the book are designed to match the latest textbook issued by the Maharashtra State board.

With the examination in focus, the *Precise Series* has been specifically designed to make preparation easier, by providing a methodical and organized perspective of the curriculum, thus greatly improving the chances of scoring well.

Chemistry is a science that has the potential to unlock the understanding of the natural world by allowing us to appreciate the changes that characterize matter interactions.

In order to make sure that students fully grasp the nub of the subject, it is important to present such concepts meaningfully and in an easy to read format.

This book has been crafted to provide an **exam-centric approach** to the curriculum, while **retaining the essence** of the subject. Each chapter is thus structured to provide a **conceptual foundation**, in addition to offering **ample practice** for acing the board examination. **Chemical formulae, bonding structures** and **chemical equations** form the basic building blocks of Chemistry and students are advised to memorise them perfectly.

To quote the Nobel Prize winner, chemist Irving Langmuir, "*A chemist who does not know Mathematics is seriously handicapped!*", a solid theoretical foundation must always be put to the test by solving numericals.

Students should take advantage of the **extensive array of numericals** provided in the book to ascertain their command on problem solving.

Questions of Board Examination from **March 2013 to July 2023** are provided so that students would get an idea about the types of questions that are asked in Board Examinations.

A holistic preparation is the key to mastering any subject and conquering the board examination.

Our **Precise Chemistry Vol. I, Std. XII Sci.** adheres to our vision and achieves several goals: **building concepts, developing competence to solve numericals, recapitulation and self-study** – all while facilitating effective preparation of the chapter.

Publisher

Edition: Sixth

The journey to create a complete book is strewn with triumphs, failures and near misses. If you think we've nearly missed something or want to applaud us for our triumphs, we'd love to hear from you.
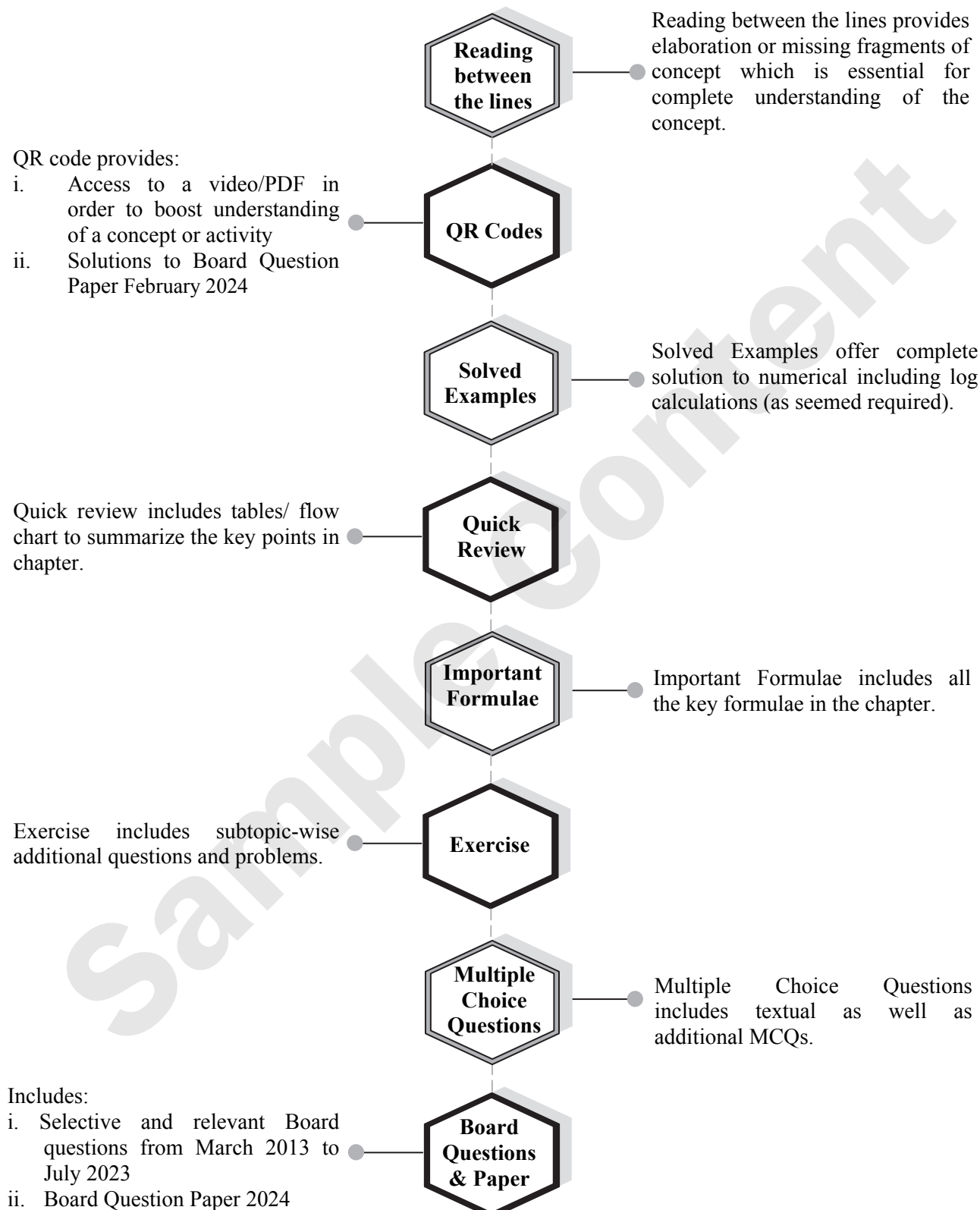Please write to us on: mail@targetpublications.org

KEY FEATURES

Reading between the lines

Reading between the lines provides elaboration or missing fragments of concept which is essential for complete understanding of the concept.

QR Codes

QR code provides:
i. Access to a video/PDF in order to boost understanding of a concept or activity
ii. Solutions to Board Question Paper February 2024

Solved Examples

Solved Examples offer complete solution to numerical including log calculations (as seemed required).

Quick Review

Quick review includes tables/ flow chart to summarize the key points in chapter.

Important Formulae

Important Formulae includes all the key formulae in the chapter.

Exercise

Exercise includes subtopic-wise additional questions and problems.

Multiple Choice Questions

Multiple Choice Questions includes textual as well as additional MCQs.

Board Questions & Paper

Includes:
i. Selective and relevant Board questions from March 2013 to July 2023
ii. Board Question Paper 2024

PAPER PATTERN

- There will be single question paper of 70 Marks and practical examination of 30 Marks in Chemistry.
- Duration of the question paper will be 3 hours.

Section A: **(18 Marks)**

This section will contain Multiple Choice Questions and Very Short Answer(VSA) type of questions.

There will be 10 MCQs and 8 VSA type of questions, each carrying **One** mark.

Students will have to attempt all the questions.

Section B: **(16 Marks)**

This section will contain 12 Short Answer (SA-I) type of questions, each carrying **Two** marks.

Students will have to attempt any 8 questions.

Section C: **(24 Marks)**

This section will contain 12 Short Answer (SA-II) type of questions, each carrying **Three** marks.

Students will have to attempt any 8 questions.

Section D: **(12 Marks)**

This section will contain 5 Long Answer (LA) type of questions, each carrying **Four** marks.

Students will have to attempt any 3 questions.

Distribution of Marks According to the Type of Questions

Type of Questions		
MCQ	1 Mark each	10 Marks
VSA	1 Mark each	8 Marks
SA - I	2 Marks each	16 Marks
SA - II	3 Marks each	24 Marks
LA	4 Marks each	12 Marks

CONTENTS

[Reference: Maharashtra State Board of Secondary and Higher Secondary Education, Pune - 04**]**

Note: 1. * mark represents Textual question.

2. # mark represents Intext question.

3. + mark represents Textual examples.

4. ☼ symbol represents textual questions that need external reference for an answer.

Scan the adjacent QR Code to know more about our *"HSC 25 Question Papers & Activity Sheets with Solutions"* book for Std. XII (Sci.) and Gear up yourself to score more in the XII Board Examination.

Scan the adjacent QR Code to know more about our *"Board Questions with Solutions"* book for Std. XII (Sci.) and Learn about the types of questions that are asked in the XII Board Examination.

3 Ionic Equilibria

Contents and Concepts

3.1 Introduction

Q.1. Define ionic equilibrium. Give four examples of ionic equilibrium. [3 Marks]

Ans: *The equilibrium between ions and unionized molecules in solution is called **ionic equilibrium**.*

e.g.

i. Equilibrium between H^+ and OH^- ions and unionized water molecules.
ii. Ionization of weak acids and weak bases.
iii. Reactions between ions of salt and ions of water.
iv. Solid salt and its ions in water.

3.2 Types of electrolyte

Q.2. Define electrolytes. [1 Mark]

Ans: *The substances which give rise to ions when dissolved in water are called **electrolytes**.*

Q.3. Define nonelectrolytes. [1 Mark]

Ans: *The substances which do not ionize and exist as molecules in aqueous solutions are called **non-electrolytes**.*

Q.4. How are electrolytes classified? [2 Marks]

Ans: The electrolytes are classified into strong and weak electrolytes. This classification is based on their extent of ionization in dilute aqueous solutions.

i. Strong electrolyte: *The electrolytes ionizing completely or almost completely are **strong electrolytes**.*
e.g. Strong acids, strong bases and salts

ii. Weak electrolyte: *The electrolytes which dissociate to a smaller extent in aqueous solution are **weak electrolytes**.*
e.g. Weak acids and weak bases

Q.5. Explain the dissociation of weak electrolytes in water. [2 Marks]

Ans:

i. Weak electrolytes are partially dissociated in dilute aqueous solutions. Hence, an equilibrium exists between the ions and the nonionized molecules.

ii. A double arrow ($\rightleftharpoons$) between the ions and nonionized molecules of the weak electrolyte is used to represent the ionization reaction.

Q.6. Use your brain power *(Textbook Page no. 47)*
Which of the following is a strong electrolyte? [1 Mark]
HF, AgCl, $CuSO_4$, CH_3COONH_4, H_3PO_4

Ans: HF is a weak electrolyte while others are strong electrolytes.

*Q.7. Define degree of dissociation. **[1 Mark]**

Ans: *The **degree of dissociation (α)** of an electrolyte is defined as a fraction of total number of moles of the electrolyte that dissociates into its ions when the equilibrium is attained.*

Q.8. Write a short note on degree of dissociation. **[2 Marks]**

Ans:

i. Definition: *Refer Q.7.*

ii. It is denoted by symbol α and given by

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

iii. Percent dissociation = α × 100

iv. If c is the molar concentration of an electrolyte X^+Y^-, the equilibrium concentration of cation or anion is $(\alpha \times c)$ mol dm^{-3}.
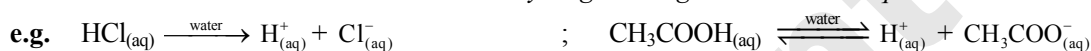
3.3 Acids and bases

*Q.9. What are acids and bases according to Arrhenius theory? **[2 Marks]**

Ans: According to Arrhenius theory, acids and bases are defined as follows:

i. Acid: *An **acid** is a substance which contains hydrogen and gives H$^+$ ions in aqueous solution.*

   e.g. $HCl_{(aq)} \xrightarrow{\text{water}} H^+_{(aq)} + Cl^-_{(aq)}$; $CH_3COOH_{(aq)} \underset{\text{water}}{\rightleftharpoons} H^+_{(aq)} + CH_3COO^-_{(aq)}$

ii. Base: *A **base** is a substance which contains OH group and produces hydroxide ions (OH$^-$ ions) in aqueous solution.*

   e.g. $NaOH_{(aq)} \xrightarrow{\text{water}} Na^+_{(aq)} + OH^-_{(aq)}$; $NH_4OH_{(aq)} \underset{\text{water}}{\rightleftharpoons} NH^+_{4(aq)} + OH^-_{(aq)}$

Note: i. Although, Arrhenius described H$^+$ ions in water as bare ions, we now know that they are hydrated in aqueous solutions and are represented as hydronium ions, H$_3$O$^+$. However, we will conveniently represent them as H$^+$.

ii. Acids and bases are familiar chemical compounds. Acetic acid is found in vinegar, citric acid in lemons, magnesium hydroxide in antacids, ammonia in household cleaning products. The tartaric acid is present in tamarind paste.

iii. Hydrochloric acid, HCl present in the gastric juice is secreted by our stomach and is essential for digestion of food.

Q.10. What are the limitations of Arrhenius theory? **[2 Marks]**

Ans: Limitations of Arrhenius theory:

i. Arrhenius theory accounts for properties of different acids and bases but is applicable only to aqueous solutions.

ii. It does not account for the basicity of NH_3 and Na_2CO_3 which do not have OH group.

Q.11. Give a brief account of Bronsted-Lowry theory of acids and bases. **[2 Marks]**

Ans: Bronsted-Lowry theory: J. N. Bronsted and T. M. Lowry (1923) proposed a more general theory known as the Bronsted-Lowry proton transfer theory.

According to this theory, acids and bases are defined as follows:

Acid: *Acid is a substance that donates a proton (H$^+$) to another substance.*

Base: *Base is a substance that accepts a proton (H$^+$) from another substance.*

   e.g. $HCl + NH_3 \rightleftharpoons NH^+_4 + Cl^-$

   Acid$_1$ Base$_2$ Acid$_2$ Base$_1$
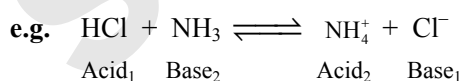
In the above reaction, HCl and NH^+_4 are proton donors and act as acids. The NH_3 and Cl$^-$ are proton acceptors and act as bases. From the above reaction, it is clear that the products of the Bronsted-Lowry acid-base reactions are acids and bases.

Q.12. What is a conjugate base? **[1 Mark]**

Ans: *The base produced by accepting the proton from an acid is the **conjugate base** of that acid.*

   e.g. Cl$^-$ is a conjugate base of acid HCl.

Q.13. What is a conjugate acid? **[1 Mark]**

Ans: *The acid produced when a base accepts a proton is called the **conjugate acid** of that base.*

   e.g. H$_3$O$^+$ is a conjugate acid of base H$_2$O.

Page no. **79** to **9**80are purposely left blank.

To see complete chapter buy **Target Notes** or **Target E-Notes**

Q.26. **Write a short note on the dissociation constant of weak acids and weak bases.** **[3 Marks]**

Ans: **Dissociation constant of weak acids and weak bases:**

i. The dissociation of a weak acid HA in water is expressed as:

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

ii. The equilibrium constant called acid dissociation constant (K_a) for this equilibrium is:

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

iii. Similarly, the dissociation of weak base BOH in water is represented as:

$$BOH_{(aq)} \rightleftharpoons B^+_{(aq)} + OH^-_{(aq)}$$

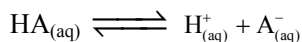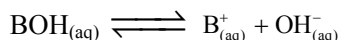iv. The equilibrium constant called base dissociation constant (K_b) for this equilibrium is:

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

v. Thus, *the **dissociation constant** of a weak acid or a weak base is defined as the equilibrium constant for dissociation equilibrium of weak acid or weak base, respectively.*

Q.27. **Derive an expression of Ostwald's dilution law for weak acid.** **[2 Marks] [July 22; July 23]**

Ans:

i. Consider an equilibrium of weak acid HA that exists in solution partly as the undissociated species HA and partly H^+ and A^- ions. Then

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$
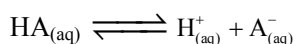
ii. The acid dissociation constant is given as:

$$K_a = \frac{[H^+][A^-]}{[HA]} \qquad \dots(1)$$

iii. Suppose 1 mol of acid HA is initially present in volume V dm^3 of the solution. At equilibrium, the fraction dissociated would be α, where α is degree of dissociation of the acid. The fraction of an acid that remains undissociated would be $(1 - \alpha)$.

	$HA_{(aq)}$ $\rightleftharpoons$ $H^+_{(aq)}$	+	$A^-_{(aq)}$
Amount present at equilibrium (mol)	$(1 - \alpha)$	α	α
Concentration at equilibrium (mol dm^{-3})	$\dfrac{1 - \alpha}{V}$	$\dfrac{\alpha}{V}$	$\dfrac{\alpha}{V}$

iv. Thus, at equilibrium $[HA] = \dfrac{1 - \alpha}{V}$ mol dm^{-3},

$$[H^+] = [A^-] = \frac{\alpha}{V} \text{ mol } dm^{-3}$$

v. Substituting these in equation (1),

$$K_a = \frac{(\alpha/V)(\alpha/V)}{(1-\alpha)/V} = \frac{\alpha^2}{(1-\alpha)V} \qquad \dots(2)$$

vi. If c is the initial concentration of an acid in mol dm^{-3} and V is the volume in dm^3 mol^{-1} then c = 1/V. Replacing 1/V in equation (2) by c, we get

$$K_a = \frac{\alpha^2 c}{1 - \alpha} \qquad \dots(3)$$

vii. For the weak acid HA, α is very small, or $(1 - \alpha) \cong 1$.
With this equation (2) and (3) becomes:

$$K_a = \alpha^2/V \text{ and } K_a = \alpha^2 c \qquad \dots(4)$$

$$\alpha = \frac{\sqrt{K_a}}{c} \text{ or } \alpha = \sqrt{K_a . V} \qquad \dots(5)$$

The equation (5) implies that the degree of dissociation of a weak acid is inversely proportional to the square root of its concentration or directly proportional to the square root of volume of the solution containing 1 mol of the weak acid.

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To see complete chapter buy **Target Notes** or **Target E-Notes**

v. Substituting these concentrations in equation (1),

$$K_b = \frac{(\alpha / V)(\alpha / V)}{(1 - \alpha) / V} = \frac{\alpha^2}{(1 - \alpha)V} \quad \ldots(2)$$

vi. If c is the initial concentration of a base in mol dm^{-3} and V is the volume in dm^3 mol^{-1} then c = 1/V. Replacing 1/V in equation (2) by c, we get

$$K_b = \frac{\alpha^2 c}{1 - \alpha} \quad \ldots(3)$$

vii. For the weak base, α is very small, or $(1 - \alpha) \cong 1$.
With this equation (2) and (3) becomes:
$K_b = \alpha^2/V$ and $K_b = \alpha^2 c$

$$\alpha = \sqrt{K_b \cdot V} \quad \text{or} \quad \alpha = \sqrt{\frac{K_b}{c}} \quad \ldots(4)$$

The degree of dissociation of a weak base is inversely proportional to square root of its concentration and is directly proportional to square root of volume of the solution containing 1 mol of weak base.

*Q.30. **Derive the relationship between degree of dissociation and dissociation constant in weak electrolytes.**

[4 Marks]

Ans: *Refer Q.27 or Q.29.*

Solved Examples

+Q.31. **A weak monobasic acid is 0.05% dissociated in 0.02 M solution. Calculate dissociation constant of the acid.** *(Problem 3.1 of Textbook page no. 50)* **[3 Marks]**

Solution:
Given: Percent dissociation = 0.05%,   Concentration (c) = 0.02 M
To find: Dissociation constant of acid (K_a)
Formulae: i.   Percent dissociation = $\alpha \times 100$ ii. $K_a = \alpha^2 c$
Calculation: Using formula (i),

$$\alpha = \frac{\text{Percent dissociation}}{100} = \frac{0.05}{100} = 5 \times 10^{-4}$$

$c = 0.02 \text{ M} = 2 \times 10^{-2} \text{ M}$
Using formula (ii),
$K_a = (5 \times 10^{-4})^2 \times 2 \times 10^{-2}$
$\quad = 25 \times 10^{-8} \times 2 \times 10^{-2} = 50 \times 10^{-10} = \mathbf{5 \times 10^{-9}}$

Ans: Dissociation constant of acid (K_a) is $\mathbf{5 \times 10^{-9}}$.

*Q.32. **Acetic acid is 5% ionised in its decimolar solution. Calculate the dissociation constant of acid.**

[2 Marks]

Solution:
Given: Percent dissociation = 5%,   Concentration (c) = 1 decimolar
To find: Dissociation constant of acid (K_a)
Formulae: i.   Percent dissociation = $\alpha \times 100$ ii. $K_a = \alpha^2 c$
Calculation: Using formula (i),

$$\alpha = \frac{\text{Percent dissociation}}{100} = \frac{5}{100} = 0.05$$

$c = 1 \text{ decimolar} = 0.1 \text{ M}$
Using formula (ii),
$K_a = (0.05)^2 \times (0.1) = \mathbf{2.5 \times 10^{-4}}$

Ans: Dissociation constant of acid is $\mathbf{2.5 \times 10^{-4}}$.

Q.33. **The dissociation constant of a weak monobasic acid is 3.5×10^{-8}. Calculate its degree of dissociation in 0.05 M solution.** **[2 Marks]**

Solution:
Given: Dissociation constant (K_a) = 3.5×10^{-8}, Concentration (c) = 0.05 M
To find: Degree of dissociation (α)

Formula: $K_a = \alpha^2 c$

Calculation: Using formula,

$$\therefore \quad \alpha = \sqrt{\frac{K_a}{c}}$$

$$= \sqrt{\frac{3.5 \times 10^{-8}}{0.05}}$$

$$= \sqrt{\frac{3.5}{5} \times 10^{-6}} = \sqrt{0.7} \times 10^{-3}$$

$$= 0.8366 \times 10^{-3} \qquad \textit{(Using log table)}$$

$$= \mathbf{8.366 \times 10^{-4}}$$

Calculation using log table:

$\sqrt{0.7}$ = Antilog$_{10}$ [log$_{10}$ $\sqrt{0.7}$]

$\quad$ = Antilog$_{10}$[$\frac{1}{2} \times$ log$_{10}$ 0.7]

$\quad$ = Antilog$_{10}$ [$\frac{1}{2} \times$ $\bar{1}.8451$]

$\quad$ = Antilog$_{10}$ [$\frac{1}{2} \times$ ($\bar{2} \times 1.8451$)]

$\quad$ = Antilog$_{10}$ [$\bar{1}.9225$] = **0.8366**

Ans: Degree of dissociation for 0.05 M weak monobasic acid is $\mathbf{8.366 \times 10^{-4}}$.

+Q.34. The dissociation constant of NH$_4$OH is 1.8×10^{-5}. Calculate its degree of dissociation in 0.01 M solution. *(Problem 3.2 of Textbook page no. 51)* [2 Marks]

Solution:

Given: Dissociation constant (K_b) = 1.8×10^{-5}, Concentration (c) = 0.01 M

To find: Degree of dissociation (α)

Formula: $K_b = \alpha^2 c$

Calculation: c = 0.01 M = 1×10^{-2} M

Using formula,

$K_b = \alpha^2 c$

$$\alpha = \sqrt{\frac{K_b}{c}}$$

Hence, $\alpha = \sqrt{\frac{1.8 \times 10^{-5}}{1 \times 10^{-2}}} = \sqrt{1.8 \times 10^{-3}} = \sqrt{18 \times 10^{-4}}$

$$= 4.242 \times 10^{-2} = \mathbf{0.04242}$$

Ans: Degree of dissociation (α) of 0.01 M NH$_4$OH is **0.04242**.

*Q.35. Dissociation constant of acetic acid is 1.8×10^{-5}. Calculate percent dissociation of acetic acid in 0.01 M solution. [2 Marks]

Solution:

Given: Dissociation constant (K_a) = 1.8×10^{-5}, Concentration (c) = 0.01 M

To find: Percent dissociation

Formulae: i. $K_a = \alpha^2 c$ $\qquad$ ii. Percent dissociation = $\alpha \times 100$

Calculation: c = 0.01 M = 1×10^{-2} M

Using formula (i),

$$\therefore \quad \alpha = \sqrt{\frac{K_a}{c}}$$

$$= \sqrt{\frac{1.8 \times 10^{-5}}{1 \times 10^{-2}}} = \sqrt{1.8 \times 10^{-3}} = \sqrt{18 \times 10^{-4}}$$

$$= 4.242 \times 10^{-2}$$

Using formula (ii),

Percent dissociation = $\alpha \times 100$

$$= 4.242 \times 10^{-2} \times 100 = \mathbf{4.242\%}$$

Ans: Percent dissociation of 0.01 M acetic acid solution is **4.242%**.

+Q.36. A weak monobasic acid is 12% dissociated in 0.05 M solution. What is percent dissociation in 0.15 M solution? *(Problem 3.3 of Textbook page no. 51)* [2 Marks]

Solution:

Given: Percent dissociation = 12%, Concentration (c_1) = 0.05 M, Concentration (c_2) = 0.15 M

To find: Percent dissociation in 0.15 M solution

Formulae: i. $K_a = \alpha^2 c$ $\qquad$ ii. Percent dissociation = $\alpha \times 100$

Calculation: If α_1 and α_2 are the values of degree of dissociation at two different concentrations c_1 and c_2 respectively, then

$K_a = \alpha_1^2 c_1 = \alpha_2^2 c_2$

∴ $\alpha_1^2 c_1 = \alpha_2^2 c_2$

$\alpha_1 = \dfrac{12}{100} = 0.12$

∴ $(0.12)^2 \times 0.05 = \alpha_2^2 \times 0.15$

$\alpha_2^2 = \dfrac{(0.12)^2 \times 0.05}{0.15} = 0.0048$

Hence, $\alpha_2 = 0.0693$

∴ Percent dissociation in 0.15 M solution $= 0.0693 \times 100 = $ **6.93%**

Ans: Percent dissociation in 0.15 M solution is **6.93%**.

+Q.37. Calculate $[H_3O^+]$ in 0.1 mol dm^{-3} solution of acetic acid. [Given: K_a (CH$_3$COOH) = 1.8 × 10^{-5}]
(Problem 3.4 of Textbook page no. 51) [2 Marks]

Solution:

Given: Dissociation constant (K_a) = 1.8 × 10^{-5}, Concentration (c) = 0.1 mol dm^{-3}
To find: Concentration of H_3O^+
Formula: $K_a = \alpha^2 c$
Calculation: Let α be the degree of dissociation.
Concentrations of various species involved at equilibrium are as follows:

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$
   (1- α)c αc αc
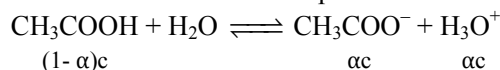
Using formula,

∴ $\alpha = \sqrt{\dfrac{K_a}{c}} = \sqrt{\dfrac{1.8 \times 10^{-5}}{0.1}} = 1.34 \times 10^{-2}$

$[H_3O^+] = \alpha \times c = 1.34 \times 10^{-2} \times 0.1$
          $= $ **1.34 × 10^{-3} mol dm^{-3}**

Ans: Concentration of H_3O^+ in 0.1 mol dm^{-3} solution of acetic acid is **1.34 × 10^{-3} mol dm^{-3}**.

3.5   Autoionization of water

Q.38. Derive an expression for ionic product of water. [2 Marks]
Ans:
i.   Pure water ionizes to a very small extent. The ionization equilibrium of water is represented as,

$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$
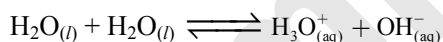
ii.   The equilibrium constant (K) for the ionization of water is given by

$K = \dfrac{[H_3O^+][OH^-]}{[H_2O]^2}$ …(1)

or $K[H_2O]^2 = [H_3O^+][OH^-]$ …(2)

iii. A majority of H_2O molecules are undissociated, consequently concentration of water $[H_2O]$ can be treated as constant. Then,
$[H_2O]^2 = K'$.

iv.   Substituting this in equation (2), we get,
$K \times K' = [H_3O^+][OH]$ …(3)
$K_w = [H_3O^+][OH^-]$
where $K_w = KK'$ is called **ionic product of water**.

v.   *The product of molar concentrations of hydronium (or hydrogen) ions and hydroxyl ions at equilibrium in pure water at the given temperature is called **ionic product of water**.*

Q.39. What is the value of K_w at 298 K? [1 Mark]
Ans: In pure water, H_3O^+ ion concentration always equals the concentration of OH^- ion.
Thus, at 298 K, this concentration is found to be 1.0×10^{-7} mol/L.

∴ $K_w = [H_3O^+][OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7})$

$K_w = 1.0 \times 10^{-14}$

Q.40. Internet my friend *(Textbook Page no. 51)*
Find out the values of ionic product K_w of water at various temperatures:
237 K, 283 K, 293 K, 303 K, 313 K, 323 K

Ans: The values of K_w at various temperatures are:

Temperature	Value of K_w
273 K	0.114×10^{-14}
283 K	0.292×10^{-14}
293 K	0.681×10^{-14}
303 K	1.47×10^{-14}
313 K	2.92×10^{-14}
323 K	5.47×10^{-14}

Note: K_w of water increases as the temperature increases.

3.6 pH scale

Q.41. What is pH and pOH? [2 Marks]
Ans:
i. *The **pH** of a solution is defined as the negative logarithm to the base 10, of the concentration of H^+ ions in solution in mol dm^{-3}.*
   pH is expressed mathematically as
   $pH = -\log_{10}[H^+]$ or $pH = -\log_{10}[H_3O^+]$
ii. Similarly, ***pOH*** *of a solution can be defined as the negative logarithm to the base 10, of the molar concentration of OH^- ions in solution.*
   $pOH = -\log_{10}[OH^-]$

*Q.42. **Derive the relation pH + pOH = 14.**

                    OR

Derive the relationship between pH and pOH. [2 Marks] [Mar 22]

Ans: Relationship between pH and pOH:
The ionic product of water is given as:
$K_w = [H_3O^+][OH^-]$
Now, $K_w = 1 \times 10^{-14}$ at 298 K
Thus, $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$
Taking logarithm of both the sides, we write
$\log_{10}[H_3O^+] + \log_{10}[OH^-] = -14$
$-\log_{10}[H_3O^+] + \{-\log_{10}[OH^-]\} = 14$
Now, $pH = -\log_{10}[H_3O^+]$ and $pOH = -\log_{10}[OH^-]$
∴   **pH + pOH = 14**

*Q.43. **Define pH and pOH. Derive the relationship between pH and pOH.** [4 Marks]
Ans: *Refer Q.41 and Q.42.*

Q.44. Explain the pH value of neutral, acidic and basic solutions. [3 Marks]
Ans:
i. **Neutral solution:** For pure water or any aqueous neutral solution at 298 K,
   $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M
   Hence, $pH = -\log_{10}[H_3O^+] = -\log_{10}[1 \times 10^{-7}] = $ **7**
ii. **Acidic solution:** In acidic solution, there is excess of H_3O^+ ions, or $[H_3O^+] > [OH^-]$
   Hence, $[H_3O^+] > 1 \times 10^{-7}$ and **pH < 7**
iii. **Basic solution:** In basic solution, there is excess of OH^- ions, or $[H_3O^+] < [OH^-]$
   Hence, $[H_3O^+] < 1.0 \times 10^{-7}$ and **pH > 7**

Q.45. Write a short note on pH scale. [2 Marks]
Ans: pH Scale:
i. Instead of writing concentration of H_3O^+ ions in mol dm^{-3}, sometimes it is convenient to express it on the logarithmic scale. This is known as pH scale.
ii. $pH = -\log_{10}[H_3O^+]$

iii. Based on pH value, solutions are classified as acidic (pH < 7), basic (pH > 7) and neutral solution (pH = 7).

iv. pH scale is shown in the following figure:

pH Scale
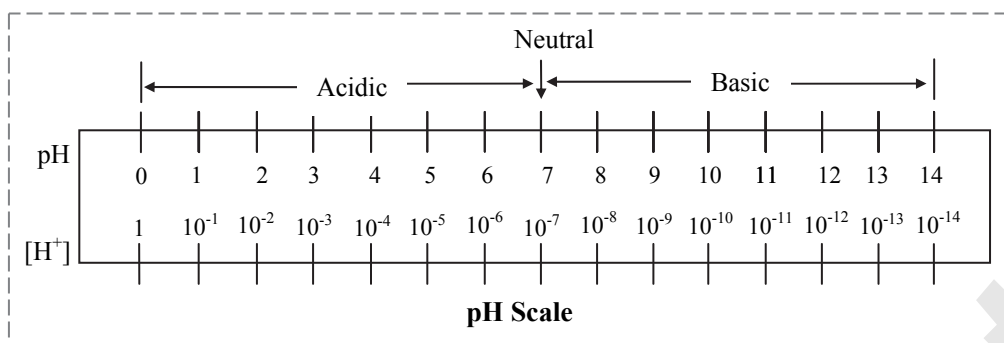
Q.46. Use your brain power *(Textbook Page no. 53)*

i. Suppose that pH of monobasic and dibasic acid is the same. Does this mean that the molar concentrations of both acids are identical? **[2 Marks]**

ii. How pH of pure water vary with temperature? Explain. **[2 Marks]**

Ans:

i. No, it does not mean that the molar concentrations of both acids are identical.

Consider a strong monobasic acid (HA) and a strong dibasic acid (H_2A), both having the same molar concentration (c).

They dissociate as given below:

$HA \longrightarrow H^+ + A^-$; $H_2A \longrightarrow 2H^+ + A^{2-}$
   c   c   c   c   2c   c

where c is molar concentration of the acid

It is clear that dibasic acid gives twice the number of H^+ ions as compared to monobasic acid.

So, if both the acids have same pH, then the molar concentration of dibasic acid would be half that of monobasic acid.

ii. The pH of pure water decreases as the temperature increases.

pH is a measure of H_3O^+ ions in a solution. As concentration of H_3O^+ ions increase, pH decreases.

When water (H_2O) dissociates, it forms H_3O^+ and OH^- ions.

$$2H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$
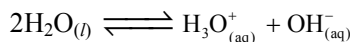
The equilibrium constant for this reaction is known as K_w.

When temperature increases, the value of K_w increases. That is, there are more (H_3O^+) ions in water. As a result, pH of water decreases.

Note: pH decreases with increase in temperature. However, pure water will remain neutral as there is always the same concentration of each of H_3O^+ and OH^- ions.

Solved Examples

+Q.47. Calculate pH and pOH of 0.01 M HCl solution. *(Problem 3.5 of Textbook page no. 52)* **[2 Marks]**

Solution:

Given: Concentration of HCl solution = 0.01 M

To find: pH and pOH

Formulae: i. $pH = -\log_{10}[H_3O^+]$ ii. $pH + pOH = 14$

Calculation: HCl is a strong acid. It dissociates almost completely in water as:

$HCl_{(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$
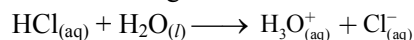
Hence, $[H_3O^+] = c = 0.01 \text{ M} = 1 \times 10^{-2}$ M

From formula (i),

$pH = -\log_{10}[H_3O^+] = -\log_{10}[1 \times 10^{-2}] = $ **2**

From formula (ii),

$pH + pOH = 14$

∴ $pOH = 14 - pH = 14 - 2 = $ **12**

Ans: The pH and pOH of 0.01 M HCl solution are **2** and **12** respectively.

Page no. **87** to **94** are purposely left blank.

To see complete chapter buy **Target Notes** or **Target E-Notes**

Q.70. Explain hydrolysis of salt of weak acid and weak base for which $K_a = K_b$. [2 Marks]

Ans:

i. CH_3COONH_4 is a salt of weak acid, CH_3COOH ($K_a = 1.8 \times 10^{-5}$) and weak base, NH_4OH ($K_b = 1.8 \times 10^{-5}$). When the salt CH_3COONH_4 is dissolved in water, it undergoes hydrolysis:

$$CH_3COO^-_{(aq)} + NH^+_{4(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + NH_4OH_{(aq)}$$

$$\text{(weak acid)} \qquad \text{(weak base)}$$

ii. The ions of the salt react with water as

$$CH_3COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)} \qquad \dots(1)$$

$$NH^{4+}_{(aq)} + 2H_2O_{(l)} \rightleftharpoons NH_4OH_{(aq)} + H_3O^+_{(aq)} \qquad \dots(2)$$

iii. As $K_a = K_b$, the relative strength of acid and base produced in hydrolysis is the same.

iv. Therefore, the solution is neutral. Hydrolysis of NH_4^+ produces as many H_3O^+ ions as that of CH_3COO^- produces OH^- ions.

*Q.71. What is meant by hydrolysis? A solution of CH_3COONH_4 is neutral. Why? [3 Marks]

Ans: *Refer Q.58. (vi) and Q.70.*

3.8 Buffer solutions

Q.72. Define buffer solution. [1 Mark]

Ans: *Buffer solution* is defined as a solution which resists drastic changes in pH when a small amount of strong acid or strong base or water is added to it.

Q.73. What are the types of buffer solutions? [1 Mark]

Ans: Types of buffer solutions:

There are two types of buffer solutions:

i. Acidic buffer maintains an acidic pH.   ii. Basic buffer maintains an alkaline pH.

Q.74. i. Define: Acidic buffer solution [1 Mark] [Mar 23]

    ii. Give Henderson Hasselbalch equation to calculate pH of acidic buffer solution. [1 Mark]

Ans: Acidic buffer solution:

i. *A solution containing a weak acid and its salts with strong base is called an **acidic buffer solution**.*

    e.g. A solution containing weak acid such as CH_3COOH and its salt such as CH_3COONa is an acidic buffer solution.

ii. pH of acidic buffer is given by Henderson Hasselbalch equation.

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

where $pK_a = -\log_{10}K_a$ and K_a is the dissociation constant of the acid.

Q.75. What is a basic buffer solution? Give Henderson Hasselbalch equation to calculate pOH of basic buffer solution. [2 Marks]

Ans: Basic buffer solution:

i. *A solution containing a weak base and its salt with strong acid is the **basic buffer solution**.*

    e.g. A solution containing a weak base such as NH_4OH and its salt such as NH_4Cl is a basic buffer solution.

ii. The pOH of basic buffer is given by Henderson Hasselbalch equation.

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

where $pK_b = -\log_{10}K_b$ and K_b is the dissociation constant for the base.

*Q.76. How are basic buffer solutions prepared? [1 Mark]

Ans: Basic buffer solutions are prepared by mixing aqueous solutions of a weak base and its salt with strong acid.

*Q.77. Classify the following buffers into different types: [2 Marks]

i. $CH_3COOH + CH_3COONa$ ii. $NH_4OH + NH_4Cl$

iii. Sodium benzoate + benzoic acid   iv. $Cu(OH)_2 + CuCl_2$

Ans:

	Buffer	Type
i.	$CH_3COOH + CH_3COONa$	Acidic buffer
ii.	$NH_4OH + NH_4Cl$	Basic buffer
iii.	Sodium benzoate + benzoic acid	Acidic buffer
iv.	$Cu(OH)_2 + CuCl_2$	Basic buffer

Reading between the lines

*Weak acid + Its salt of strong base $\Rightarrow$ **Acidic** buffer*

*Weak base + Its salt of strong acid $\Rightarrow$ **Basic** buffer*

i. *CH_3COOH is weak acid and CH_3COONa is its salt of strong base $\Rightarrow$ Acidic buffer*

ii. *NH_4OH is weak base and NH_4Cl is its salt of strong acid $\Rightarrow$ Basic buffer*

iii. *Benzoic acid is weak acid and sodium benzoate is its salt of strong base $\Rightarrow$ Acidic buffer*

iv. *$Cu(OH)_2$ is weak base and $CuCl_2$ is its salt of strong acid $\Rightarrow$ Basic buffer*

Q.78. Explain buffer action with the help of sodium acetate-acetic acid buffer. [2 Marks]

Ans: **Buffer action:**

i. Consider sodium acetate-acetic acid buffer. Here sodium acetate is a strong electrolyte which dissociates completely in water producing large concentration of CH_3COO^- as follows:

$$CH_3COONa_{(aq)} \longrightarrow CH_3COO^-_{(aq)} + Na^+_{(aq)}$$

ii. On the other hand since the acetic acid is a weak acid, the concentration of undissociated CH_3COOH molecules is usually high.

iii. If a strong acid is added to this solution, the added H^+ ions will be consumed by the conjugate base CH_3COO^- present in large concentration.

iv. Similarly, if small amount of base is added, the added OH^- ions will be neutralized by the large concentration of acetic acid as shown in the following reactions:

$$CH_3COO^-_{(aq)} + H^+_{(aq)} \longrightarrow CH_3COOH_{(aq)}$$

(large concentration) (added acid)

$$CH_3COOH_{(aq)} + OH^-_{(aq)} \longrightarrow CH_3COO^-_{(aq)} + H_2O_{(l)}$$

(large concentration) (added acid)

v. The acid or base added thus cannot change the $[H^+]$ or $[OH^-]$ concentrations and the pH of the buffer remains unchanged.

vi. Dilution does not have any effect on pH of buffer. This is because the concentration ratio term in Henderson Hasselbalch equation remains the same. The dilution does not change this ratio.

Q.79. What are the properties of buffer solutions? [2 Marks]

Ans: **Properties of buffer solution:**

i. 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.

ii. The pH of a buffer solution is independent of the volume of the solution. Hence, the dilution of a buffer solution will not change its pH.

iii. The pH of a buffer solution does not change even if it is kept for a long time.

*Q.80. Write one property of a buffer solution. [1 Mark]

Ans: *Refer Q.79.*

Q.81. **Can you tell?** *(Textbook Page no. 56)*

It is enough to add a few mL of a buffer solution to maintain its pH. Which property of buffer is used here? [1 Mark]

Ans: Yes, it is enough to add a few mL of buffer solution to maintain the pH of a solution. This is because the pH of a buffer solution is independent of the volume of the solution.

Q.82. Give the applications of buffer solution. **[2 Marks]**

Ans: Applications of buffer solution:

Buffer solution finds extensive applications in a variety of fields. Some of its applications are given.

i. **In biochemical system:** pH of blood in our body is maintained at 7.36-7.42 due to (HCO_3^- + H_2CO_3) buffer. A mere change of 0.2 pH units can cause death. The saline solution used for intravenous injection must contain buffer system to maintain the proper pH of the blood.

ii. **Agriculture:** The soils get buffered due to presence of salts such as carbonate, bicarbonate, phosphates and organic acids. The choice of fertilizers depends upon pH of soil.

iii. **Industry:** Buffers play an important role in paper, dye, ink, paint and drug industries.

iv. **Medicine:** Penicillin preparations are stabilized by addition of sodium citrate as buffer. When citric acid is added to milk of magnesia [Mg(OH)$_2$], magnesium citrate is formed, which is a buffer.

v. **Analytical chemistry:** In qualitative analysis, a pH of 8 to 10 is required for precipitation of cations IIIA group. It is maintained with the use of (NH$_4$OH + NH$_4$Cl) buffer.

(Any two applications)

Solved Examples

+**Q.83. Calculate the pH of buffer solution containing 0.05 mol NaF per litre and 0.015 mol HF per litre.** **[K$_a$ = 7.2 × 10^{-4} for HF]** *(Problem 3.9 of Textbook page no. 57)* **[3 Marks]**

Solution:

Given: [Salt] = 0.05 M, [Acid] = 0.015 M, K_a = 7.2 × 10^{-4} for HF

To find: pH of the buffer solution

Formula: $pH = pK_a + \log_{10} \dfrac{[salt]}{[acid]}$

Calculation: The pH of acidic buffer is given by Henderson-Hasselbalch equation:

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

$pK_a = -\log_{10}K_a = -\log_{10} 7.2 \times 10^{-4}$
$= 4 - \log_{10} 7.2 = 4 - 0.8573 = 3.1427$

Substitution in the Henderson-Hasselbalch equation gives

$pH = 3.1427 + \log_{10} \dfrac{0.05}{0.015} = 3.1427 + \log_{10} 3.33$

$= 3.1427 + 0.5224 = 3.6651 \approx$ **3.67**

Ans: The pH of the given buffer solution is **3.67**.

+**Q.84. Calculate the pH of buffer solution composed of 0.1 M weak base BOH and 0.2 M of its salt BA.** **[K$_b$ = 1.8 × 10^{-5} for the weak base]** *(Problem 3.10 of Textbook page no. 57)* **[3 Marks]**

Solution:

Given: [Base] = 0.1 M, [Salt] = 0.2 M, K_b = 1.8 × 10^{-5} for the weak base

To find: pH of the buffer solution

Formulae: i. $pOH = pK_b + \log_{10} \dfrac{[salt]}{[base]}$ ii. $pH + pOH = 14$

Calculation: pOH of basic buffer is given by Henderson-Hasselbalch equation:

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

$pK_b = -\log_{10}K_b$
$= -\log_{10} (1.8 \times 10^{-5}) = 5 - \log_{10} 1.8$
$= 5 - 0.2553 = 4.7447$

Substitution in the Henderson-Hasselbalch equation gives

$pOH = 4.7447 + \log_{10} \dfrac{0.2}{0.1} = 4.7447 + \log_{10} 2$

$= 4.7447 + 0.3010 = 5.0457$

From formula (ii),

$pH = 14 - pOH = 14 - 5.0457$
$= 8.9543 \approx$ **8.95**

Ans: The pH of the given buffer solution is **8.95**.

Q.85. Calculate the pH of buffer solution composed of 0.01 M weak base BOH and 0.02 M of its salt BA.
   [$K_b = 1.8 \times 10^{-5}$ for weak base]   **[3 Marks] [Mar 22]**
Solution:
Given: [Base] = 0.01 M, [Salt] = 0.02 M, $K_b = 1.8 \times 10^{-5}$ for the weak base
To find: pH of the buffer solution

Formulae: i. $pOH = pK_b + \log_{10} \dfrac{[salt]}{[base]}$ ii. $pH + pOH = 14$

Calculation: pOH of basic buffer is given by Henderson-Hasselbalch equation:

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

$$pK_b = -\log_{10}K_b$$
$$= -\log_{10}(1.8 \times 10^{-5}) = 5 - \log_{10} 1.8$$
$$= 5 - 0.2553 = 4.7447$$

Substitution in the Henderson-Hasselbalch equation gives

$$pOH = 4.7447 + \log_{10} \frac{0.02}{0.01} = 4.7447 + \log_{10} 2$$
$$= 4.7447 + 0.3010 = 5.0457$$

From formula (ii),
$$pH = 14 - pOH = 14 - 5.0457$$
$$= 8.9543 \approx \textbf{8.95}$$

Ans: The pH of the given buffer solution is **8.95**.

3.9	Solubility product

Q.86. Explain solubility product with suitable example.   **[3 Marks]**
Ans:
i.   When powdered sparingly soluble salt such as AgCl is put into water and stirred vigorously, a very small amount of AgCl dissolves in water to form its saturated solution. Most of the salt remains undissolved. Thus, solid AgCl is in contact with its saturated solution.
ii.   AgCl is a strong electrolyte. Hence, the quantity of AgCl that dissolves in water dissociates completely into its constituent ions, Ag^+ and Cl^-.
iii.   A dynamic equilibrium exists between undissolved solid AgCl and the dissolved ions, Ag^+ and Cl^-, in the saturated solution. This equilibrium is called solubility equilibrium.
iv.   Solubility equilibrium is represented as:

$$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$$
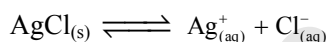
v.   The expression for its equilibrium constant is:

$$K = \frac{[Ag^+][Cl^-]}{[AgCl]} \qquad …(1)$$

vi.   The concentration of undissolved solid AgCl is constant. So, we can write
        [AgCl] = constant = K'
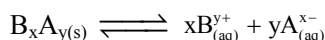vii.   Substituting in equation (1) we have

$$K = \frac{[Ag^+][Cl^-]}{K'}$$

        $K \times K' = [Ag^+][Cl^-]$
viii.   The product of $K \times K'$ is another constant and is called solubility product. It is the product of concentrations of ions in a saturated solution. It is denoted by K_{sp}.
        $K_{sp} = [Ag^+][Cl^-]$
ix.   For the general salt, the solubility equilibrium is:

$$B_xA_{y(s)} \rightleftharpoons xB^{y+}_{(aq)} + yA^{x-}_{(aq)}$$

        The solubility product is
        $K_{sp} = [B^{y+}]^x [A^{x-}]^y$ …(2)
x.   Thus, *in the saturated solution of sparingly soluble salt, the product of equilibrium concentrations of the constituent ions raised to the power equal to their respective coefficients in the balanced equilibrium expression at a given temperature is called **solubility product**.*

Q.87. Complete the following table: [2 Marks]

Salt	Solubility equilibrium	K_{sp} Expression
$BaSO_4$	-------	-------
Bi_2S_3	-------	$K_{(sp)} = [Bi^{3+}]^2[S^{2-}]^3$
CaF_2	$CaF_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2F^-_{(aq)}$	-------

Ans:

Salt	Solubility equilibrium	K_{sp} Expression
$BaSO_4$	$\mathbf{BaSO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + SO^{2-}_{4(aq)}}$	$\mathbf{K_{sp} = [Ba^{2+}][SO^{2-}_4]}$
Bi_2S_3	$\mathbf{Bi_2S_{3(s)} \rightleftharpoons 2Bi^{3+}_{(aq)} + 3S^{2-}_{(aq)}}$	$K_{(sp)} = [Bi^{3+}]^2[S^{2-}]^3$
CaF_2	$CaF_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2F^-_{(aq)}$	$\mathbf{K_{(sp)} = [Ca^{2+}][F^-]^2}$

Q.88. Write the K_{sp} expression for $Ca_3(PO_4)_2$. [1 Mark]

Ans: $Ca_3(PO_4)_{2(s)} \rightleftharpoons 3Ca^{2+}_{(aq)} + 2PO^{3-}_{4(aq)}$

$K_{sp} = [Ca^{2+}]^3[PO^{3-}_4]^2$
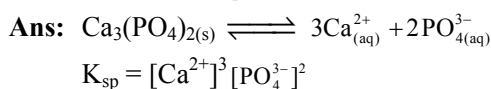
Q.89. Define the term: Solubility. [1 Mark]

Ans: *The **solubility** of a compound is the amount in grams that dissolves per unit volume (which may be 100 mL or 1 L of its saturated solution).*

Q.90. What is molar solubility? [1 Mark]

Ans: Molar solubility:

i. *The number of moles of a compound that dissolve to give one litre of saturated solution is called its **molar solubility**.*

ii. Molar solubility (mol/L) = $\dfrac{\text{Solubility in g/L}}{\text{Molar mass in g/mol}}$

Q.91. Derive the relationship between solubility and solubility product for a general salt B_xA_y. [2 Marks]

Ans:

i. Consider the solubility equilibrium for salt B_xA_y.

$B_xA_{y(s)} \rightleftharpoons xB^{y+}_{(aq)} + yA^{x-}_{(aq)}$
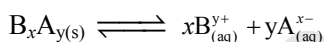
The solubility product is given as:
$K_{sp} = [B^{y+}]^x[A^{x-}]^y$ …(1)

ii. If S is the molar solubility of the salt, the equilibrium concentrations of the ions in the saturated solution will be
$[B^{y+}] = xS$ mol/L
$[A^{x-}] = xS$ mol/L

iii. Substituting these values in equation (1),
$K_{sp} = [xS]^x[yS]^y = x^x y^y S^{x+y}$

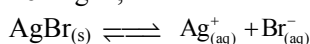**Q.92. What is the relationship between molar solubility and solubility product for the following salts?**
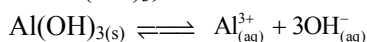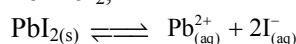
[1 Mark Each]

i. **AgBr**   ii. **PbI_2**   iii. **$Al(OH)_3$**

Ans:

i. For AgBr,
$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$
Here, $x = 1$, $y = 1$
∴ $K_{sp} = x^x y^y S^{x+y} = (1)^1(1)^1 S^{1+1} = S^2$

ii. For PbI_2,
$PbI_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2I^-_{(aq)}$
$x = 1$, $y = 2$
∴ $K_{sp} = x^x y^y S^{x+y} = (1)^1(2)^2 S^{1+2} = 4S^3$

iii. For $Al(OH)_3$,
$Al(OH)_{3(s)} \rightleftharpoons Al^{3+}_{(aq)} + 3OH^-_{(aq)}$

$x = 1$, $y = 3$
$K_{sp} = x^x y^y S^{x+y} = (1)^1(3)^3 S^{1+3} = 27S^4$

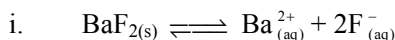**Q.93. Write the relationship between molar solubility and solubility product for the following compounds:**

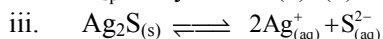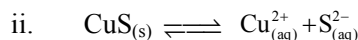**[1 Mark Each]**

i.   BaF₂   **ii.   CuS**   **iii.   Ag₂S**

Ans:

i. $BaF_{2(s)} \rightleftharpoons Ba^{2+}_{(aq)} + 2F^-_{(aq)}$

$x = 1$, $y = 2$
$K_{sp} = x^x y^y S^{x+y} = (1)^1 (2)^2 S^{1+2} = 4S^3$

ii. $CuS_{(s)} \rightleftharpoons Cu^{2+}_{(aq)} + S^{2-}_{(aq)}$

$x = 1$, $y = 1$
$K_{sp} = x^x y^y S^{x+y} = (1)^1 (1)^1 S^{1+1} = S^2$

iii. $Ag_2S_{(s)} \rightleftharpoons 2Ag^+_{(aq)} + S^{2-}_{(aq)}$

$x = 2$, $y = 1$
$K_{sp} = x^x y^y S^{x+y} = (2)^2 (1)^1 S^{2+1} = 4S^3$

Q.94. Use your brain power *(Textbook Page no. 60)*
What is the relationship between molar solubility and solubility product for salts given below?
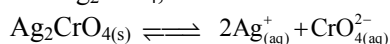
[1 Mark Each]

i.   Ag₂CrO₄   **ii.   Ca₃(PO₄)₂**   **iii.   Cr(OH)₃**

Ans:

i.   For Ag₂CrO₄,
$Ag_2CrO_{4(s)} \rightleftharpoons 2Ag^+_{(aq)} + CrO^{2-}_{4(aq)}$

Here, $x = 2$, $y = 1$

∴ $K_{sp} = x^x y^y S^{x+y} = (2)^2 (1)^1 S^{2+1} = 4S^3$

ii.   For Ca₃(PO₄)₂,
$Ca_3(PO_4)_{2(s)} \rightleftharpoons 3Ca^{2+}_{(aq)} + 2PO^{3-}_{4(aq)}$

Here, $x = 3$, $y = 2$
$K_{sp} = x^x y^y S^{x+y} = (3)^3 (2)^2 S^{3+2} = 27 \times 4 \times S^5 = 108S^5$

iii.   For Cr(OH)₃,
$Cr(OH)_{3(s)} \rightleftharpoons Cr^{3+}_{(aq)} + 3OH^-_{(aq)}$

Here, $x = 1$, $y = 3$

∴ $K_{sp} = x^x y^y S^{x+y} = (1)^2 (3)^3 S^{1+3} = 1 \times 27 \times S^4 = 27S^4$

*Q.95. Explain the relation between ionic product and solubility product to predict whether a precipitate will form when two solutions are mixed?   [2 Marks]

Ans: Condition of precipitation:
Ionic product (IP) of an electrolyte is defined in the same way as solubility product (K_{sp}). The only difference is that the ionic product expression contains concentration of ions under any condition whereas expression of K_{sp} contains only equilibrium concentrations. If,

i.   IP = K_{sp}; the solution is saturated and solubility equilibrium exists.
ii.   IP > K_{sp}; the solution is supersaturated and hence precipitation of the compound will occur.
iii.   If IP < K_{sp}, the solution is unsaturated and precipitation will not occur.

Solved Examples

+Q.96. A solution is prepared by mixing equal volumes of 0.1 M MgCl₂ and 0.3 M Na₂C₂O₄ at 293 K. Would MgC₂O₄ precipitate out? K$_{sp}$ of MgC₂O₄ at 293 K is 8.56 × 10⁻⁵. *(Problem 3.11 of Textbook page no. 59)*

[3 Marks]

Solution:

Given: Concentration of MgCl₂ solution = 0.1 M, Concentration of Na₂C₂O₄ solution = 0.3 M,
                  Volumes of both solutions mixed are equal.
 K_{sp} of MgC₂O₄ = 8.56 × 10⁻⁵

To find: Whether MgC₂O₄ will precipitate out or not

Calculation: When solution is prepared by mixing equal volumes, volume gets doubled and hence effective concentration of ions would be half of initial concentration.

$[Mg^{2+}] = \dfrac{0.1}{2} = 0.05$ mol/L

$[C_2O_4^{2-}] = \dfrac{0.3}{2}$ M $= 0.15$ mol/L

These ions would react to form sparingly soluble salt MgC₂O₄ in accordance with reaction
$Mg^{2+}_{(aq)} + C_2O^{2-}_{4(aq)} \rightleftharpoons MgC_2O_{4(s)}$
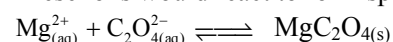
Page no. **99** to**101** are purposely left blank.

To see complete chapter buy **Target Notes** or **Target E-Notes**

*Q.102. The dissociation of H_2S is suppressed in the presence of HCl. Name the phenomenon. [1 Mark]

Ans: The phenomenon due to which dissociation of H_2S is suppressed in the presence of HCl is known as common ion effect.

┌─────────────────────────────┐
│ **Reading between the lines** │
└─────────────────────────────┘

Both H_2S and HCl produce common ion (H^+) on dissociation.
The H^+ ions combine with S^{2-} ions to produce unionized H_2S. Thus, the dissociation of H_2S is suppressed in the presence of HCl.

*Q.103. Dissociation of HCN is suppressed by the addition of HCl. Explain. [2 Marks]

Ans:

i. HCN and HCl both dissociate to produce H^+ ions which are common to both.

ii. HCN is a weak electrolyte. It dissociates to a little extent.

$$HCN_{(aq)} \rightleftharpoons H^+_{(aq)} + CN^-$$

iii. HCl is a strong electrolyte. It undergoes complete dissociation.

$$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

Both HCN and HCl provide H^+ ions.

iv. The concentration of H^+ ions in the solution increases due to the complete dissociation of HCl.

v. According to Le-Chatelier's principle, the effect of the stress (the addition of H^+ ions from HCl) applied to the ionization equilibrium of HCN is reduced by shifting the equilibrium in the backward direction.

vi. H^+ ions combines with CN^- ions to produce unionized HCN. Thus, the dissociation of HCN is suppressed by the addition of HCl.

*Q.104. Solubility of a sparingly soluble salt get affected in presence of a soluble salt having one common ion. Explain. [2 Marks]

Ans:

i. The presence of a common ion affects the solubility of a sparingly soluble salt.

ii. Consider, the solubility equilibrium of AgCl,

$$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$$

The solubility product of AgCl is
$$K_{sp} = [Ag^+][Cl^-]$$

iii. Suppose $AgNO_3$ is added to the saturated solution of AgCl. The salt $AgNO_3$ being a strong electrolyte dissociates completely in the solution.

$$AgNO_{3(aq)} \longrightarrow Ag^+_{(aq)} + NO^-_{3(aq)}$$

iv. The dissociation of AgCl and $AgNO_3$ produce a common Ag^+ ion. The concentration of Ag^+ ion in the solution increases owing to complete dissociation of $AgNO_3$.

v. According to Le-Chatelier's principle, the addition of Ag^+ ions from $AgNO_3$ to the solution of AgCl shifts the solubility equilibrium of AgCl from right to left. The reverse reaction in which AgCl precipitates, is favoured until the solubility equilibrium is re-established.

vi. However, the value of K_{sp} remains the same since it is an equilibrium constant. Thus, the solubility of a sparingly soluble compound decreases with the presence of a common ion in solution.

*Q.105. Sulphides of cation of group II are precipitated in acidic solution (H_2S + HCl) whereas sulphides of cations of group IIIB are precipitated in ammoniacal solution of H_2S.
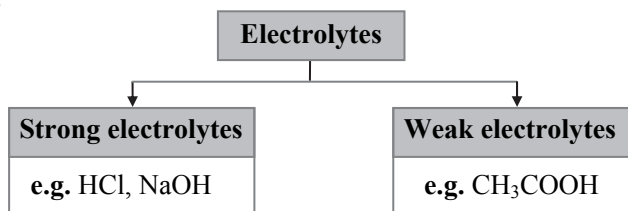Comment on the relative values of solubility product of sulphides of these. [3 Marks]

Ans:

i. Group II and group IIIB cations are precipitated as their sulphides. However, the solubility product of sulphides of group II cations is lower than group IIIB cations.

ii. Therefore, for the precipitation of cations of group II only small concentration of sulphide ion is required. This is achieved by passing H_2S gas in presence of strong electrolyte HCl, which has a common ion (H^+) with H_2S. Due to common ion effect, the dissociation of H_2S is suppressed and thus, the concentration of S^{2-} ions decreases. This results only in the precipitation of sulphides of group II while sulphides of higher group remain in solution as they require higher concentration of S^{2-} ions for precipitation.

iii. This higher concentration of S^{2-} ions is achieved by using ammoniacal solution of H_2S, where the H^+ ions from H_2S are removed by OH^- ions in solution thereby increasing the dissociation of H_2S.
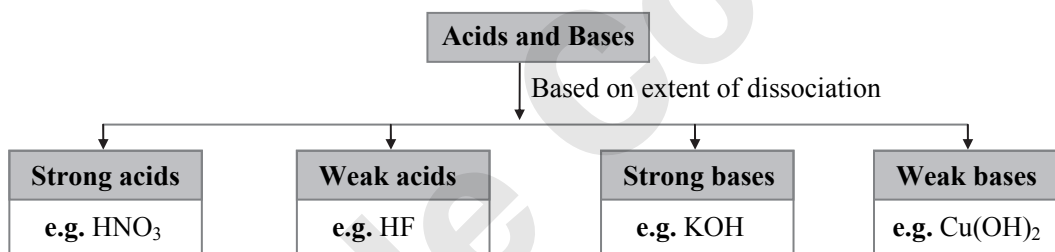
Quick Review

➢ **Types of electrolytes:**

```
                    ┌─────────────────┐
                    │   Electrolytes  │
                    └─────────────────┘
                ┌───────────┴───────────┐
     ┌────────────────────┐    ┌────────────────────┐
     │ Strong electrolytes │    │  Weak electrolytes │
     │   e.g. HCl, NaOH    │    │   e.g. CH₃COOH     │
     └────────────────────┘    └────────────────────┘
```

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

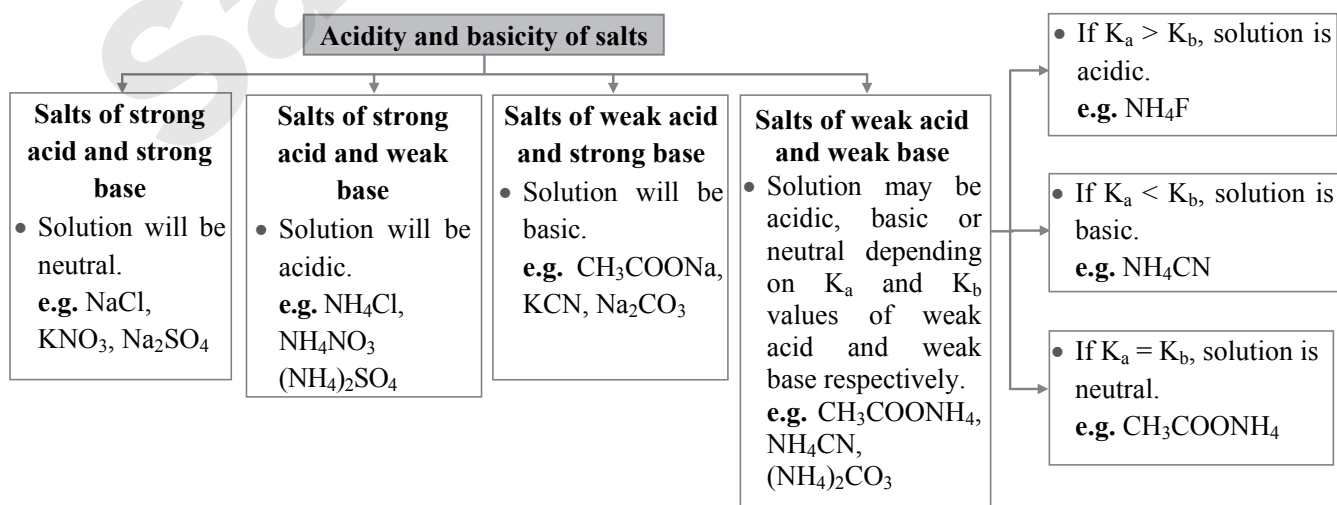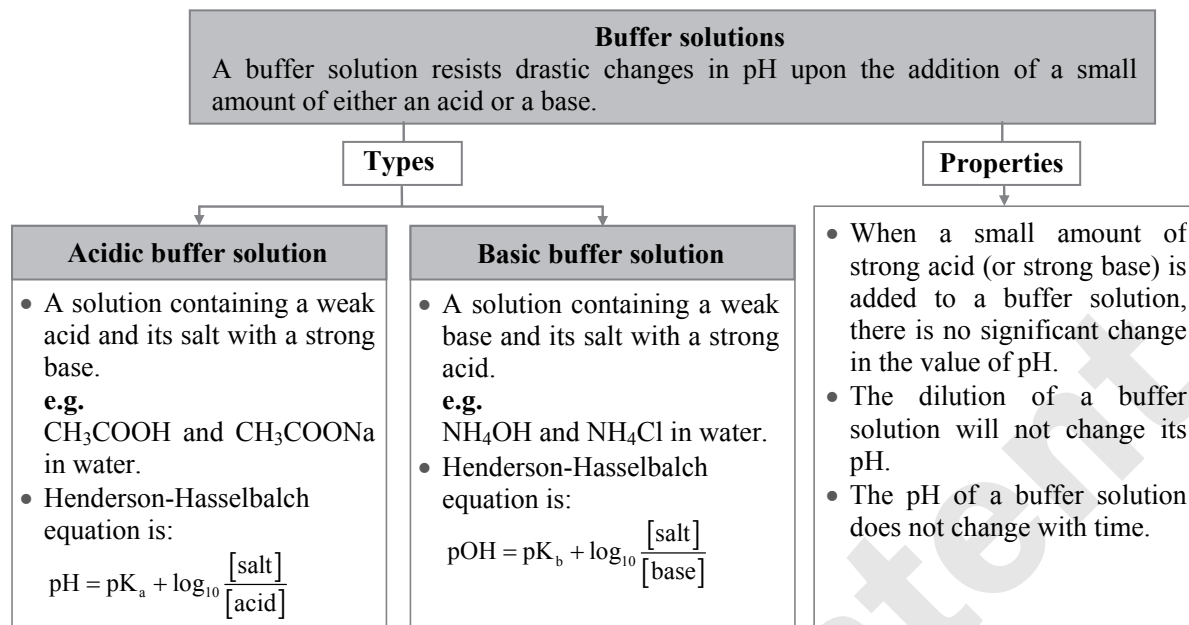| Theory	Acid	Base
Arrhenius theory	A substance that contains hydrogen and produces H^+ ions in aqueous solution.	A substance that contains OH group and produces OH^- ions in aqueous solution.
Bronsted-Lowry theory	Any substance that can donate a proton (H^+) i.e., proton donor.	Any substance that can 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│ │Weak acids │ │Strong bases│ │Weak bases │
│ e.g. HNO₃ │ │ e.g. HF   │ │ e.g. KOH  │ │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	Salts of strong acid and weak base	Salts of weak acid and strong base	Salts of weak acid and weak base
• Solution will be neutral. **e.g.** NaCl, KNO₃, Na₂SO₄	• Solution will be acidic. **e.g.** NH₄Cl, NH₄NO₃ (NH₄)₂SO₄	• Solution will be basic. **e.g.** CH₃COONa, KCN, Na₂CO₃	• Solution may be acidic, basic or neutral depending on K_a and K_b values of weak acid and weak base respectively. **e.g.** CH₃COONH₄, NH₄CN, (NH₄)₂CO₃

- If $K_a > K_b$, solution is acidic. **e.g.** NH₄F
- If $K_a < K_b$, solution is basic. **e.g.** NH₄CN
- If $K_a = K_b$, solution is neutral. **e.g.** CH₃COONH₄

➤ **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	**Basic 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: $$pH = pK_a + \log_{10}\frac{[salt]}{[acid]}$$	• 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 solubility
AB (1 : 1 type salt)	AgCl	$AgCl \rightleftharpoons \underset{S}{Ag^+} + \underset{S}{Cl^-}$	$K_{sp} = [Ag^+][Cl^-]$ $\mathbf{K_{sp} = S^2}$	$S = \sqrt{K_{sp}}$
AB_2 (1 : 2 type salt)	$PbCl_2$	$PbCl_2 \rightleftharpoons \underset{S}{Pb^{2+}} + \underset{2S}{2Cl^-}$	$K_{sp} = [Pb^{2+}][Cl^-]^2$ $K_{sp} = [S][2S]^2$ $\mathbf{K_{sp} = 4S^3}$	$S = \sqrt[3]{\dfrac{K_{sp}}{4}}$
A_2B (2 : 1 type salt)	Ag_2CrO_4	$Ag_2CrO_4 \rightleftharpoons \underset{2S}{2Ag^+} + \underset{S}{CrO_4^{2-}}$	$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$ $K_{sp} = [2S]^2[S]$ $\mathbf{K_{sp} = 4S^3}$	$S = \sqrt[3]{\dfrac{K_{sp}}{4}}$
AB_3 (1 : 3 type salt)	$AlCl_3$	$AlCl_3 \rightleftharpoons \underset{S}{Al^{3+}} + \underset{3S}{3Cl^-}$	$K_{sp} = [Al^{3+}][3Cl^-]$ $K_{sp} = [S][3S]^3$ $\mathbf{K_{sp} = 27S^4}$	$S = \sqrt[4]{\dfrac{K_{sp}}{27}}$
A_2B_3 (2 : 3 type salt)	As_2S_3	$As_2S_3 \rightleftharpoons \underset{2S}{2As^{3+}} + \underset{3S}{3S^{2-}}$	$K_{sp} = [As^{3+}]^2[S^{2-}]^3$ $K_{sp} = [2S]^2[3S]^3$ $K_{sp} = 4S^2 \times 27S^3$ $\mathbf{K_{sp} = 108S^5}$	$S = \sqrt[5]{\dfrac{K_{sp}}{108}}$

➤ **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

Important Formulae

1. Degree of dissociation (α):

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

2. Ostwald's dilution law:

$$\alpha \propto \frac{1}{\sqrt{C}}$$

OR

$$\alpha \propto \sqrt{V}$$

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

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

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

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

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

$K_b = \alpha^2/V$ 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:

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

Basic buffer:

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

10. **Solubility product (K_{sp}):**
$K_{sp} = [B^{y+}]^x [A^{x-}]^y$

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

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

12. **Relation between K_{sp} and S:**
$K_{sp} = x^x y^y S^{x+y}$

Exercise

3.1 Introduction

1. What is ionic equilibrium? **[1 Mark]**
Ans: *Refer Q.1. (Definition).*

3.2 Types of electrolyte

2. Define strong electrolytes. **[1 Mark]**
Ans: *Refer Q.4. (i)*

3. What is degree of dissociation? Give its formula. **[2 Marks]**

Ans: *Refer Q.7 and Q.8. (ii)*

3.3 Acids and bases

4. Define acids and bases according to the Arrhenius theory. Give suitable examples. **[2 Marks]**
Ans: *Refer Q.9.*

5. What are acids and bases according to Bronsted-Lowry theory? Give an example. **[2 Marks]**
Ans: *Refer Q.11.*

6. Explain conjugate acid-base pair with a suitable example. **[2 Marks]**
Ans: *Refer Q.14.*

7. What are acids and bases according to Lewis theory? Give an example. **[2 Marks]**
Ans: *Refer Q.17. (Definitions and any one example)*

8. $AlCl_3$ is a Lewis acid. Explain. **[2 Marks]**
Ans: *Refer Q.19. (i)*

3.4 Ionization of acids and bases

9. Give two examples of weak bases. **[1 Mark]**
Ans: *Refer Q.23. (v)*

10. Give two examples of strong acids. **[1 Mark]**
Ans: *Refer Q.24. (i)*

11. Define dissociation constant of a weak acid or weak base. **[1 Mark]**
Ans: *Refer Q.26. (v)*

12. The dissociation constant of NH_4OH is 1.8×10^{-5}. Calculate its degree of dissociation in 0.05 M solution. **[2 Marks]**
Ans: 0.01897

13. The dissociation constant of weak monobasic acid is 3.2×10^{-4}. Calculate its degree of dissociation in 0.02 M solution. **[2 Marks]**
Ans: 0.1265

14. A weak monobasic acid is 10% dissociated in 0.05 M solution. What is percent dissociation in 0.15 M solution? **[3 Marks][Mar 23]**
Ans: 5.77%

15. Calculate $[H_3O^+]$ in 0.3 mol dm^{-3} solution of acetic acid.
[Given**:** K_a (CH_3COOH) = 1.8×10^{-5}]
[2 Marks]

Ans: 2.324×10^{-3} mol dm^{-3}

3.5 Autoionization of water

16. Define ionic product of water. **[1 Mark]**
Ans: *Refer Q.38. (v)*

17. Derive the expression: $K_w = [H_3O^+][OH^-]$. **[2 Marks]**
Ans: *Refer Q.38.*

3.6 pH scale

18. Define pH and pOH. **[2 Marks]**
Ans: *Refer Q.41. (i) and (ii)*

19. Calculate pH of 0.02 M sulphuric acid.
[2 Marks]

Ans: 1.3979

20. The concentration of hydrogen ion in a sample of soft drink is 3.8×10^{-3} M. What is its pH?
[2 Marks]

Ans: 2.42

21. pH of a solution is 5.68. Calculate the concentration of H_3O^+ ion. **[2 Marks]**
Ans: 2.089×10^{-6} M

22. A weak monobasic acid is 0.05 % dissociated in 0.2 M solution. Calculate the pH of the solution.
[2 Marks]

Ans: 4

23. pH of a weak monobasic acid is 2.52 in its 0.02 M solution. Calculate its dissociation constant. **[3 Marks]**

Ans: 4.56×10^{-4}

3.7 Hydrolysis of salts

24. Give four examples of salts derived from strong acid and strong base. **[2 Marks]**
Ans: *Refer Q.57. (i)*

25. Explain hydrolysis of a salt of strong acid and strong base. **[2 Marks]**
Ans: *Refer Q.59.*

26. Why does aq. $CuSO_4$ solution turn blue litmus red? **[2 Marks] [July 22]**
Ans: *Refer Q.61.*

27. To get clear solution of $CuSO_4$, the addition of H_2SO_4 would be required. Give reason.
[2 Marks]

Ans: *Refer Q.62.*

3.8 Buffer solutions

28. Define acidic buffer solution. **[1 Mark]**
Ans: *Refer Q.74. (i)*

29. Give two properties of buffer solution.
[2 Marks]
Ans: *Refer Q.79.*

30. Write any four applications of buffer solution.
[2 Marks][July 23]
Ans: *Refer Q.82. (Any four)*

31. Find the pH of buffer solution if it contains 0.06 mol NaF per litre and 0.02 mol HF per litre. [$K_a = 7.2 \times 10^{-4}$ for HF] **[3 Marks]**
Ans: 3.62

3.9 Solubility product

32. Define solubility product. **[1 Mark]**
Ans: *Refer Q.86. (x)*

33. Define molar solubility. **[1 Mark]**
Ans: *Refer Q.90. (i)*

34. Write the relationship between solubility and solubility product for PbI_2. **[1 Mark] [Mar 23]**
Ans: *Refer Q.92. (ii)*

35. A solution is prepared by mixing equal volumes of 0.2 M $MgCl_2$ and 0.5 M $Na_2C_2O_4$ at 293 K. Would MgC_2O_4 precipitate out? K_{sp} of MgC_2O_4 at 293 K is 8.56×10^{-5}. **[3 Marks]**
Ans: Yes, MgC_2O_4 **will precipitate** out from the solution.

36. Solubility product of BaF_2 is 1.7×10^{-6}. Estimate its molar solubility. **[2 Marks]**
Ans: 7.518×10^{-3} mol dm^{-3}

3.10 Common ion effect

37. Write a short note on common ion effect.
[2 Marks]
Ans: *Refer Q.100.*

38. Solubility of AgCl decreases by the addition of $AgNO_3$. Explain. **[2 Marks]**
Ans: *Refer Q.104.*

Multiple Choice Questions

[1 Mark Each]

*1. The conjugate base of $[Zn(H_2O)_4]^{2+}$ is _____
   (A) $[Zn(H_2O)_4]^{2-}$ NH_3
   (B) $[Zn(H_2O)_3]^{2-}$
   (C) $[Zn(H_2O)_3OH]^+$
   (D) $[Zn(H_2O)H]^{3+}$

2. Which of the following fluoro compounds is most likely to behave as a Lewis base?
   (A) SiF_4 (B) BF_3
   (C) PF_3 (D) CF_4

3. The value of $[H_3O^+]$ in mol lit^{-1} of 0.001 M acetic acid solution ($K_a = 1.8 \times 10^{-5}$) is _____. **[Mar 22]**
   (A) 1.34×10^{-1}
   (B) 1.34×10^{-2}
   (C) 1.34×10^{-3}
   (D) 1.34×10^{-4}

*4. For pH > 7, the hydronium ion concentration would be _____.
   (A) 10^{-7} M   (B) $< 10^{-7}$ M
   (C) $> 10^{-7}$ M   (D) $\geq 10^{-7}$ M

5. The pH of weak monoacidic base is 11.2, its OH⁻ ion concentration is: **[Mar 23]**
   (A) 1.585×10^{-3} mol dm⁻³
   (B) 3.010×10^{-11} mol dm⁻³
   (C) 3.010×10^{-3} mol dm⁻³
   (D) 1.585×10^{-11} mol dm⁻³

*6. The pH of 10^{-8} M of HCl is _____.
   (A) 8   (B) 7
   (C) less than 7   (D) greater than 7

7. What is the pH of millimolar solution of ammonium hydroxide which is 20% dissociated?
   (A) 3.699   (B) 10.301
   (C) 4.691   (D) 9.301

8. Aqueous solution of which of the following will be basic?
   (A) NH_4Cl (B) $FeCl_3$
   (C) $CuSO_4$ (D) CH_3COONa

9. Which of the following salts will give highest pH in water?
   (A) KCl   (B) NaCl
   (C) Na_2CO_3 (D) $CuSO_4$

*10. Which of the following solution will have pH value equal to 1.0?
   (A) 50 mL of 0.1 M HCl + 50 mL of 0.1 M NaOH
   (B) 60 mL of 0.1 M HCl + 40 mL of 0.1 M NaOH
   (C) 20 mL of 0.1 M HCl + 80 mL of 0.1 M NaOH
   (D) 75 mL of 0.2 M HCl + 25 mL of 0.2 M NaOH

*11. Which of the following is a buffer solution?
   (A) CH_3COONa + NaCl in water
   (B) CH_3COOH + HCl in water
   (C) CH_3COOH + CH_3COONa in water
   (D) HCl + NH_4Cl in water

*12. Blood in human body is highly buffered at pH of _____.
   (A) 7.4   (B) 7.0
   (C) 6.9   (D) 8.1

13. Penicillin preparation are stabilized by addition of _____ as buffer.
   (A) sodium citrate
   (B) sodium carbonate
   (C) sodium benzoate
   (D) sodium acetate

14. A buffer solution contains 0.1 M of acetic acid and 0.1 M of sodium acetate. What will be its pH, if pK_a of acetic acid is 4.75?
   (A) 4.00   (B) 4.75
   (C) 5.00   (D) 5.25

15. What is the solubility product (K_{sp}) of $BaSO_4$ in pure water?
   [S = molar solubility]
   (A) $4 S^3$ (B) S^2
   (C) $27 S^4$ (D) $108 S^5$

16. What is the solubility product (K_{sp}) of calcium phosphate in pure water?
   [S = molar solubility]
   (A) $108 S^5$ (B) $72 S^3$
   (C) $6 S^5$ (D) $121 S^2$

*17. The solubility product of a sparingly soluble salt AX is 5.2×10^{-13}. Its solubility in mol dm⁻³ is _____.
   (A) 7.2×10^{-7} (B) 1.35×10^{-4}
   (C) 7.2×10^{-8} (D) 13.5×10^{-8}

18. Concentration of the Ag^+ ions in a saturated solution of $Ag_2C_2O_4$ is 2.2×10^{-4} mol L⁻¹. Solubility product of $Ag_2C_2O_4$ is _____.
   (A) 2.42×10^{-8}
   (B) 2.66×10^{-12}
   (C) 4.25×10^{-11}
   (D) 5.3×10^{-12}

19. Solubility of AgCl is least in _____.
   (A) 0.1 M $BaCl_2$
   (B) 0.1 M $AlCl_3$
   (C) 0.1 M NaCl
   (D) pure water

Answers to Multiple Choice Questions

1.	(C)	2.	(C)	3.	(D)	4.	(B)
5.	(A)	6.	(C)	7.	(B)	8.	(D)
9.	(C)	10.	(D)	11.	(C)	12.	(A)
13.	(A)	14.	(B)	15.	(B)	16.	(A)
17.	(A)	18.	(D)	19.	(B)		

Hints to Multiple Choice Questions

3. $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$
   $(1-\alpha)c \qquad\qquad \alpha c \qquad \alpha c$

   $\alpha = \sqrt{\dfrac{K_a}{c}} = \sqrt{\dfrac{1.8 \times 10^{-5}}{0.001}} = 0.134$

   $[H_3O^+] = \alpha \times c = 0.134 \times 0.001$
   $\qquad\qquad = 1.34 \times 10^{-4}$ mol L⁻¹

5. 1.585×10^{-3} mol dm⁻³
   pOH = 14 − pH = 14 − 11.2 = 2.8
   pOH = $-\log_{10}[OH^-]$
   ∴ $\log_{10}[OH^-] = -$ pOH = $-2.8 = -2 - 0.8 - 1 + 1$
   $\qquad\qquad\qquad = -3 + 0.2 = \bar{3}.2$

   $[OH^-] = $ Antilog₁₀ $\bar{3}.2 = 1.585 \times 10^{-3}$ mol dm⁻³