

Equilibrium — Assignment 2

Class 11 Chemistry · Higher-difficulty challenge set + integrated problems · Answer key included · Narayan Gurukul Academy

Instructions: (1) Attempt all questions; start with the topic-wise challenge set, then attempt the integrated (multi-topic) problems.
(2) Give all answers with correct SI units. $g = 10 \text{ m/s}^2$ unless stated.
(3) A final answer key is provided at the end — try the full set before checking it.

Part A — Topic-wise challenge questions

Reversible Reactions and Equilibrium State

1. Explain the characteristics of the equilibrium state with an example.
2. Distinguish between a reversible and an irreversible reaction with one example each.
3. Why is equilibrium called dynamic?

Equilibrium Constants K_c and K_p

1. Derive the relation between K_p and K_c for a gaseous reaction.
2. Calculate K_p for $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ if $K_c = 2.5 \times 10^{-6}$ at 500 K.
3. Why is the value of the equilibrium constant independent of initial concentrations?

Le Chatelier's Principle

1. State Le Chatelier's principle and explain the effect of a change in temperature on an exothermic equilibrium.
2. Predict the effect of pressure and concentration changes on the ammonia synthesis reaction.
3. How does a catalyst affect the position of equilibrium?

Ionic Equilibrium — Acids and Bases

1. Explain the Bronsted–Lowry concept of acids and bases with an example.
2. Define pH and pOH and derive the relation $\text{pH} + \text{pOH} = 14$.
3. Calculate the pH of 0.001 M HCl.

Ionisation Constants — K_a , K_b and Degree of Dissociation

1. State Ostwald's dilution law and derive the relation $\alpha = \sqrt{(K_a/c)}$.
2. Explain how the degree of dissociation of a weak acid changes on dilution.
3. Derive the relation between K_a , K_b and K_w for a conjugate acid-base pair.

Buffer Solutions and Salt Hydrolysis

1. Define a buffer solution and explain how an acidic buffer resists a change in pH.
2. Derive the Henderson–Hasselbalch equation.
3. Explain salt hydrolysis and state whether the solution is acidic or basic for the four types of salts.

Solubility Product

1. Define the solubility product and derive the relation between solubility and K_{sp} for AB_2 .
2. Explain the common ion effect on the solubility of a sparingly soluble salt.
3. State when a precipitate is formed from a solution.

Part B — Integrated problems (multi-topic)

1. For the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, $K_c = 2.5 \times 10^{-6}$ at 500 K. Calculate K_p . ($R = 0.0831 \text{ L bar K}^{-1} \text{ mol}^{-1}$)
2. Calculate the pH of a 0.1 M acetic acid solution given $K_a = 1.8 \times 10^{-5}$.
3. A buffer is made from 0.1 M acetic acid and 0.1 M sodium acetate ($pK_a = 4.74$). Calculate its pH.
4. The solubility of AgCl is $1.3 \times 10^{-5} \text{ mol L}^{-1}$. Calculate its K_{sp} .
5. The solubility product of AgCl is 1.8×10^{-10} . Calculate its solubility in pure water and in 0.01 M NaCl.

Answer Key

Reversible Reactions and Equilibrium State

At equilibrium the forward and backward rates are equal and concentrations stay constant; the state is dynamic (both reactions continue). Examples: $N_2O_4 \rightleftharpoons 2NO_2$ (reversible); $AgNO_3 + NaCl \rightarrow AgCl$ (irreversible).

Equilibrium Constants K_c and K_p

$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$; $K_p = K_c(RT)^{\Delta n}$. $K_p = 2.5 \times 10^{-6} \times (0.0831 \times 500)^{-2} \approx 1.45 \times 10^{-9}$. K depends only on temperature.

Le Chatelier's Principle

A system at equilibrium shifts to oppose an imposed change. For an exothermic reaction, increasing temperature shifts the equilibrium backwards. For $N_2 + 3H_2 \rightleftharpoons 2NH_3$ increasing pressure shifts it forward. A catalyst does not change the position of equilibrium.

Ionic Equilibrium — Acids and Bases

Bronsted–Lowry: acid is a proton donor, base a proton acceptor (HCl/H_2O , H_2O/NH_3). $pH + pOH = 14$ at 25 °C. pH of 0.001 M $HCl = 3$.

Ionisation Constants — K_a , K_b and Degree of Dissociation

$K_a = \frac{[H^+][A^-]}{[HA]}$; Ostwald: $\alpha = \sqrt{(K_a/c)}$; dilution increases α . $K_a \times K_b = K_w$. pH of 0.1 M acetic acid ≈ 2.87 .

Buffer Solutions and Salt Hydrolysis

A buffer resists pH change; acidic buffer = weak acid + its salt with a strong base. Henderson–Hasselbalch: $pH = pK_a + \log\left(\frac{[\text{salt}]}{[\text{acid}]}\right)$. Salt hydrolysis: strong acid–weak base salt $\rightarrow$ acidic; weak acid–strong base salt $\rightarrow$ basic; weak–weak depends on K_a/K_b ; strong–strong $\rightarrow$ neutral.

Solubility Product

$K_{sp} = [A^+][B^-]$ for AB ; for AB_2 , $K_{sp} = 4s^3$. Precipitate forms when the ionic product exceeds K_{sp} ; the common ion effect decreases solubility.

Integrated problems

1: $\Delta n = -2$; $K_p = K_c(RT)^{-2} \approx 1.45 \times 10^{-9}$. 2: $[H^+] = \sqrt{(K_a \cdot c)} = \sqrt{(1.8 \times 10^{-5})} \approx 1.34 \times 10^{-3}$, $pH \approx 2.87$. 3: $pH = pK_a + \log(1) = 4.74$. 4: $K_{sp} = s^2 = (1.3 \times 10^{-5})^2 = 1.69 \times 10^{-10}$. 5: $s = \sqrt{(1.8 \times 10^{-10})} \approx 1.34 \times 10^{-5} \text{ M}$ in water; in 0.01 M NaCl, $s = K_{sp}/0.01 \approx 1.8 \times 10^{-8} \text{ M}$.