# 📚 MATH LOVE INSTITUTE

Class 11 Chemistry - Chapter 6

# ⚖️ EQUILIBRIUM

**Session: 2025-26 | CBSE Curriculum**

Complete Study Material with Theory, Numerical, MCQs & Case Studies

**Indore, Madhya Pradesh**



## 📖 CHAPTER INTRODUCTION

### 🎯 Why This Chapter Matters:

The **Equilibrium** is one of the most important chapters in Class 11 Chemistry! This chapter explains:

- ✓ How reversible reactions reach a balance point
- ✓ Why reactions don't go to completion
- ✓ How to predict direction of reactions
- ✓ Industrial applications (Haber process, Contact process)
- ✓ Foundation for Chemical Kinetics, Thermodynamics, and all acid-base chemistry!

**Board Exam Weightage: 8-10 marks | JEE: 4-5 questions | NEET: 3-4 questions**

### 💡 Learning Strategy for This Chapter:

1. **Understand Equilibrium Concept:** Don't memorize - understand what dynamic equilibrium means
2. **Master $K_c$ and $K_p$:** Practice writing equilibrium constant expressions
3. **Le Chatelier's Principle:** This is your best friend - learn it well!
4. **Practice pH Calculations:** Strong acids, weak acids, buffers - all types!
5. **Solve Previous Year Questions:** Pattern repeats every year!

# ⚖️ EQUILIBRIUM IN PHYSICAL AND CHEMICAL PROCESSES

## 1️⃣ What is Equilibrium?

### 📘 Definition:

**Equilibrium** is a state in a reversible process when the two opposing processes occur at the same rate and the system appears to be static (no observable change).

### Types of Equilibrium:

- **Physical Equilibrium:** Equilibrium between different phases of same substance
- **Chemical Equilibrium:** Equilibrium in reversible chemical reactions

## 2 Characteristics of Equilibrium

### ✓ Key Features:

- ✓ **Dynamic nature:** Both forward and backward reactions continue
- ✓ **Same rate:** Rate of forward reaction = Rate of backward reaction
- ✓ **Constant concentrations:** Concentrations of reactants and products remain constant
- ✓ **Closed system:** Can only be achieved in a closed system
- ✓ **Catalyst effect:** Catalyst speeds up attainment but doesn't change position
- ✓ **From either direction:** Same equilibrium reached from either side

### 🔬 Example: Physical Equilibrium

**Water-Vapor Equilibrium (Closed Container):**

$H_2O(l) \rightleftharpoons H_2O(g)$

Initially, water evaporates rapidly. As vapor accumulates, condensation starts. Eventually, rate of evaporation = rate of condensation → Equilibrium!

**Other Examples:**

- Ice ⇌ Water (at 0°C)
- Saturated solution: Solid ⇌ Dissolved ions
- Gas in liquid: $CO_2$(dissolved) ⇌ $CO_2$(gas) in soda

# 🛸 LAW OF CHEMICAL EQUILIBRIUM AND EQUILIBRIUM CONSTANT

## 3️⃣ Law of Mass Action

### 📜 Statement (Guldberg and Waage, 1864):

The rate of a chemical reaction is directly proportional to the product of active masses (molar concentrations) of the reactants, with each concentration term raised to the power equal to its stoichiometric coefficient.

For a general reaction: aA + bB ⇌ cC + dD

### ⚡ Equilibrium Constant Expression ($K_c$):

$$K_c = [C]^c[D]^d / [A]^a[B]^b$$

**Where [ ] denotes molar concentration at equilibrium**

## 4️⃣ $K_p$ and $K_c$ Relationship

For gaseous equilibrium, we can express equilibrium constant in terms of partial pressures:

$$K_p = (P_C)^c(P_D)^d / (P_A)^a(P_B)^b$$

**Relationship:** $K_p = K_c(RT)^{\Delta n}$

**Where $\Delta n = (c + d) - (a + b)$ = moles of gaseous products - moles of gaseous reactants**

### Example: Calculate $K_p$ from $K_c$

**Problem:** For $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, $K_c = 0.5$ at 400K. Calculate $K_p$.

**Solution:**

$\Delta n = 2 - (1 + 3) = -2$

$K_p = K_c(RT)^{\Delta n}$

$K_p = 0.5 \times (0.0821 \times 400)^{-2}$

$K_p = 0.5 \times (32.84)^{-2}$

$K_p = 0.5 \times 1/1078.46$

$\mathbf{K_p = 4.64 \times 10^{-4}}$

## 5️⃣ Homogeneous and Heterogeneous Equilibrium

| Type | Description | Example | K Expression |
|---|---|---|---|
| **Homogeneous** | All reactants and products in same phase | $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ | $K_c = [NH_3]^2/[N_2][H_2]^3$ |
| **Heterogeneous** | Reactants and products in different phases | $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ | $K_p = P_{CO_2}$ |

### ⚠️ Important Note:

In heterogeneous equilibrium:

- ✓ Concentration of **pure solids** is NOT included in K expression
- ✓ Concentration of **pure liquids** is NOT included in K expression
- ✓ Only gases and aqueous species are included

# 🔄 LE CHATELIER'S PRINCIPLE

### 📜 Statement:

If a system at equilibrium is subjected to a change in concentration, temperature, volume, or pressure, the equilibrium shifts in a direction that tends to reduce or counteract the effect of the change.

**Simple Version:** "System fights the change!"

## 6️⃣ Effect of Concentration Change

### ✓ Rules:

- **Add reactant:** Equilibrium shifts RIGHT (forward) to consume it
- **Add product:** Equilibrium shifts LEFT (backward) to consume it
- **Remove reactant:** Equilibrium shifts LEFT to replace it
- **Remove product:** Equilibrium shifts RIGHT to replace it

**Important:** K value does NOT change with concentration change!

### Example: Haber Process

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

**Industrial Application:**

- ✓ $NH_3$ is continuously removed by liquefaction
- ✓ This shifts equilibrium forward
- ✓ More $NH_3$ is produced
- ✓ Increases yield without changing K

## 7️⃣ Effect of Pressure Change

### ✓ Rules (for gaseous equilibrium):

- **Increase pressure:** Equilibrium shifts to side with FEWER moles of gas
- **Decrease pressure:** Equilibrium shifts to side with MORE moles of gas
- **If $\Delta n = 0$:** Pressure change has NO effect

**Remember:** $\Delta n$ = (moles of gaseous products) - (moles of gaseous reactants)

### Examples:

**1. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$**

$\Delta n = 2 - 4 = -2$ (negative)

High pressure favors forward reaction (fewer moles on product side)

**2. $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$**

$\Delta n = 2 - 2 = 0$

Pressure change has NO effect

**3. $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$**

$\Delta n = 2 - 1 = +1$ (positive)

Low pressure favors forward reaction (more moles on product side)

## 8️⃣ Effect of Temperature Change

### ✓ Rules:

- **Increase temperature:** Equilibrium shifts in ENDOTHERMIC direction ($\Delta H > 0$)
- **Decrease temperature:** Equilibrium shifts in EXOTHERMIC direction ($\Delta H < 0$)

**Important:** Temperature change DOES change K value!

- For exothermic reaction: K decreases with increase in T
- For endothermic reaction: K increases with increase in T

### Example: Formation of $NH_3$

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$; $\Delta H = -92.4$ kJ (Exothermic)

| Temperature Change | Effect | K Value |
|---|---|---|
| Increase T | Shifts backward (opposes heat) | K decreases |
| Decrease T | Shifts forward (favors heat release) | K increases |

**Industrial Compromise:** ~450°C is used - low enough for good yield, high enough for fast rate!

## 9️⃣ Effect of Catalyst

### ⚡ Key Points:

- ✓ Catalyst speeds up both forward and backward reactions equally
- ✓ Equilibrium is reached faster
- ✓ **Position of equilibrium does NOT change**
- ✓ **K value remains same**
- ✓ Used in industry to save time, not to increase yield

# 🧪 IONIC EQUILIBRIUM IN SOLUTION

## 🔟 Acids, Bases and pH

### 📘 Definitions:

**Arrhenius Theory:**

- Acid: Substance that releases $H^+$ ions in water
- Base: Substance that releases $OH^-$ ions in water

**Bronsted-Lowry Theory:**

- Acid: Proton ($H^+$) donor
- Base: Proton ($H^+$) acceptor

**Lewis Theory:**

- Acid: Electron pair acceptor
- Base: Electron pair donor

## 1️⃣ 1️⃣ pH Scale

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

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

$$pH + pOH = 14 \text{ (at 25°C)}$$

**For pure water at 25°C:**

$$[H^+] = [OH^-] = 1.0 \times 10^{-7} \text{ M}$$

$$pH = pOH = 7$$

| pH Range | Nature | $[H^+]$ vs $[OH^-]$ | Examples |
| --- | --- | --- | --- |
| pH < 7 | Acidic | $[H^+] > [OH^-]$ | HCl, $H_2SO_4$, Vinegar, Lemon juice |
| pH = 7 | Neutral | $[H^+] = [OH^-]$ | Pure water, NaCl solution |
| pH > 7 | Basic | $[H^+] < [OH^-]$ | NaOH, KOH, Soap, Baking soda |

## 1️⃣ 2️⃣ Strong Acids and Strong Bases

### 💪 Strong Acids (100% ionization):

- ✓ HCl (Hydrochloric acid)
- ✓ $HNO_3$ (Nitric acid)
- ✓ $H_2SO_4$ (Sulfuric acid)
- ✓ $HClO_4$ (Perchloric acid)
- ✓ HBr (Hydrobromic acid)
- ✓ HI (Hydroiodic acid)

**Mnemonic: "HI! He's Not Br, Cl Saw Her"**

### 💪 Strong Bases (100% ionization):

- ✓ NaOH (Sodium hydroxide)
- ✓ KOH (Potassium hydroxide)
- ✓ LiOH (Lithium hydroxide)
- ✓ $Ca(OH)_2$ (Calcium hydroxide)
- ✓ $Ba(OH)_2$ (Barium hydroxide)

**Mnemonic: "Na K Li Ca Ba are Strong Bases"**

### 📝 Example: pH of Strong Acid

**Problem:** Calculate pH of 0.01 M HCl solution.

**Solution:**

HCl is a strong acid → 100% ionization

$HCl \rightarrow H^+ + Cl^-$

$[H^+] = 0.01\ M = 1 \times 10^{-2}\ M$

$pH = -\log[H^+] = -\log(10^{-2}) = 2$

**Answer: pH = 2**

### 📝 Example: pH of Strong Base

**Problem:** Calculate pH of 0.001 M NaOH solution.

**Solution:**

NaOH is a strong base → 100% ionization

$NaOH \rightarrow Na^+ + OH^-$

$[OH^-] = 0.001\ M = 1 \times 10^{-3}\ M$

$pOH = -\log[OH^-] = -\log(10^{-3}) = 3$

$pH = 14 - pOH = 14 - 3 = 11$

**Answer: pH = 11**

## 1️⃣3️⃣ Weak Acids and Weak Bases

Weak acids and bases only partially ionize in water. We use equilibrium constant to describe their ionization.

**For Weak Acid HA:**

$$HA \rightleftharpoons H^+ + A^-$$

$$K_a = [H^+][A^-] / [HA]$$

**pH Formula:**

$$[H^+] = \sqrt{(K_a \times c)}$$

$$pH = \frac{1}{2}(pK_a - \log c)$$

**Where c = initial concentration of acid**

### For Weak Base BOH:

$$BOH \rightleftharpoons B^+ + OH^-$$

$$K_b = [B^+][OH^-] / [BOH]$$

**pOH Formula:**

$$[OH^-] = \sqrt{(K_b \times c)}$$

$$pOH = \frac{1}{2}(pK_b - \log c)$$

### 📝 Example: pH of Weak Acid

**Problem:** Calculate pH of 0.1 M $CH_3COOH$ solution ($K_a = 1.8 \times 10^{-5}$).

**Solution:**

For weak acid: $[H^+] = \sqrt{(K_a \times c)}$

$[H^+] = \sqrt{(1.8 \times 10^{-5} \times 0.1)}$

$[H^+] = \sqrt{(1.8 \times 10^{-6})}$

$[H^+] = 1.34 \times 10^{-3}$ M

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

**pH = 2.87**

## 1️⃣4️⃣ Buffer Solutions

### 📘 Definition:

**Buffer Solution:** A solution that resists change in pH upon addition of small amounts of acid or base.

### Types of Buffers:

- **Acidic Buffer:** Weak acid + its salt with strong base
- Example: $CH_3COOH$ + $CH_3COONa$ (pH < 7)

- **Basic Buffer:** Weak base + its salt with strong acid
- Example: $NH_4OH$ + $NH_4Cl$ (pH > 7)

### Henderson-Hasselbalch Equation:

**For Acidic Buffer:**

$$pH = pK_a + \log([Salt]/[Acid])$$

**For Basic Buffer:**

$$pOH = pK_b + \log([Salt]/[Base])$$

### 📝 Example: Buffer pH Calculation

**Problem:** Calculate pH of buffer containing 0.1 M $CH_3COOH$ and 0.2 M $CH_3COONa$ ($pK_a$ = 4.74).

**Solution:**

Using Henderson-Hasselbalch equation:

$pH = pK_a + \log([CH_3COONa]/[CH_3COOH])$

$pH = 4.74 + \log(0.2/0.1)$

$pH = 4.74 + \log(2)$

$pH = 4.74 + 0.30$

**pH = 5.04**

### 🩺 Biological Importance:

**Blood Buffer System:**

Human blood pH = 7.4 ± 0.05

Main buffer: $H_2CO_3/HCO_3^-$ system

- $H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$
- Ratio $[HCO_3^-]/[H_2CO_3] = 20:1$
- pH < 7.35 → Acidosis (can be fatal)
- pH > 7.45 → Alkalosis (can be fatal)

# 📊 SOLUBILITY EQUILIBRIUM

## 1️⃣5️⃣ Solubility Product ($K_{sp}$)

### 📘 Definition:

For a sparingly soluble salt, the product of molar concentrations of its ions (each raised to the power of its stoichiometric coefficient) in a saturated solution at a given temperature is called the **Solubility Product ($K_{sp}$)**.

**For: $A_xB_y(s) \rightleftharpoons xA^{n+}(aq) + yB^{m-}(aq)$**

$$K_{sp} = [A^{n+}]^x [B^{m-}]^y$$

**Examples:**

**1. $AgCl(s) \rightleftharpoons Ag^{+}(aq) + Cl^{-}(aq)$**

$K_{sp} = [Ag^{+}][Cl^{-}]$

**2. $PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2I^{-}(aq)$**

$K_{sp} = [Pb^{2+}][I^{-}]^2$

**3. $Ca_3(PO_4)_2(s) \rightleftharpoons 3Ca^{2+}(aq) + 2PO_4^{3-}(aq)$**

$K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2$

## 16 Relationship between Solubility and $K_{sp}$

### 📝 Example: Calculate Solubility from $K_{sp}$

**Problem:** $K_{sp}$ of AgCl at 25°C is $1.0 \times 10^{-10}$. Calculate its solubility in mol/L.

**Solution:**

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

Let solubility = s mol/L

Then $[Ag^{+}] = s$ and $[Cl^{-}] = s$

$K_{sp} = [Ag^{+}][Cl^{-}] = s \times s = s^2$

$1.0 \times 10^{-10} = s^2$

$s = \sqrt{(1.0 \times 10^{-10})}$

**$s = 1.0 \times 10^{-5}$ mol/L**

# SOLVED NUMERICAL PROBLEMS

### Problem 1: Calculate $K_c$

**Problem:** For the reaction $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, at equilibrium concentrations are: $[H_2] = 0.2$ M, $[I_2] = 0.3$ M, $[HI] = 1.2$ M. Calculate $K_c$.

**Solution:**

$K_c = [HI]^2 / [H_2][I_2]$

$K_c = (1.2)^2 / (0.2 \times 0.3)$

$K_c = 1.44 / 0.06$

$\mathbf{K_c = 24}$

## Problem 2: Degree of Dissociation

**Problem:** 1 mole of $PCl_5$ was heated in a 1 L vessel. At equilibrium, 0.6 moles of $Cl_2$ were found. Calculate degree of dissociation (α) and $K_c$.

$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$

**Solution:**

| | $PCl_5$ | $PCl_3$ | $Cl_2$ |
|---|---|---|---|
| Initial | 1 | 0 | 0 |
| Change | -α | +α | +α |
| At equilibrium | 1-α | α | α |

Given: $[Cl_2]$ = 0.6 M, so α = 0.6

**Degree of dissociation α = 0.6 or 60%**

At equilibrium:

$[PCl_5]$ = 1 - 0.6 = 0.4 M

$[PCl_3]$ = 0.6 M

$[Cl_2]$ = 0.6 M

$K_c = [PCl_3][Cl_2] / [PCl_5]$

$K_c = (0.6 \times 0.6) / 0.4$

**$K_c$ = 0.9**

## 📝 Problem 3: pH of Mixed Solutions

**Problem:** Calculate pH when 50 mL of 0.1 M HCl is mixed with 50 mL of 0.2 M NaOH.

**Solution:**

Moles of HCl = 0.05 × 0.1 = 0.005 mol

Moles of NaOH = 0.05 × 0.2 = 0.01 mol

HCl + NaOH → NaCl + $H_2O$

0.005 mol HCl reacts with 0.005 mol NaOH

Remaining NaOH = 0.01 - 0.005 = 0.005 mol

Total volume = 100 mL = 0.1 L

$[OH^-]$ = 0.005 / 0.1 = 0.05 M

pOH = -log(0.05) = 1.30

pH = 14 - 1.30

**pH = 12.70**

### Problem 4: Common Ion Effect

**Problem:** $K_{sp}$ of AgCl = $1.0 \times 10^{-10}$. Calculate its solubility in 0.01 M NaCl solution.

**Solution:**

$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$

From NaCl: $[Cl^-] = 0.01$ M (common ion)

Let solubility of AgCl = s mol/L

Then $[Ag^+] = s$

$[Cl^-] = 0.01 + s \approx 0.01$ M (s is very small)

$K_{sp} = [Ag^+][Cl^-]$

$1.0 \times 10^{-10} = s \times 0.01$

$s = (1.0 \times 10^{-10}) / 0.01$

**$s = 1.0 \times 10^{-8}$ mol/L**

**Note:** Solubility decreased from $10^{-5}$ (in pure water) to $10^{-8}$ (in NaCl solution) - this is the **common ion effect**!

# 📝 PRACTICE MCQs (20 Questions)

**Q1. For the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$, if $K_c = 0.5$, then $K_p$ at 400K is:**

(a) $0.5 \times (RT)^2$

(b) $0.5 \times (RT)^{-2}$

(c) $0.5 \times RT$

(d) 0.5

**Q2. pH of 0.001 M NaOH solution is:**

(a) 3

(b) 11

(c) 2

(d) 12

**Q3. For $H_2 + I_2 \rightleftharpoons 2HI$, $K_c = 54.8$. For $2HI \rightleftharpoons H_2 + I_2$, $K_c$ is:**

(a) 54.8

(b) 1/54.8

(c) $(54.8)^2$

(d) $\sqrt{54.8}$

**Q4. Which has maximum pH?**

(a) 0.1 M HCl

(b) 0.1 M $CH_3COOH$

(c) 0.1 M NaOH

(d) 0.1 M $NH_4OH$

**Q5. Buffer solution is formed by:**

(a) $HCl + NaCl$

(b) $CH_3COOH + CH_3COONa$

(c) $NaOH + NaCl$

(d) $H_2SO_4 + Na_2SO_4$

**Q6. For $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$, $K_p$ expression is:**

(a) $P_{CO_2}$

(b) $[CaO][CO_2]/[CaCO_3]$

(c) $1/P_{CO_2}$

(d) $[CO_2]$

**Q7. Degree of dissociation increases with:**

(a) Increase in concentration

(b) Decrease in concentration

(c) Increase in pressure

(d) No change

**Q8. pH of blood is maintained at 7.4 by:**

(a) $HCl/NaCl$ buffer

(b) $H_2CO_3/HCO_3^-$ buffer

(c) $NH_4OH/NH_4Cl$ buffer

(d) $CH_3COOH/CH_3COONa$ buffer

**Q9. Common ion effect is used to:**

(a) Increase solubility

(b) Decrease solubility

(c) No effect on solubility

(d) Increase pH

**Q10. Equilibrium constant depends on:**

(a) Pressure

(b) Concentration

(c) Temperature

(d) Volume

**Q11. For $N_2O_4(g) \rightleftharpoons 2NO_2(g)$, increasing pressure will:**

(a) Shift equilibrium right

(b) Shift equilibrium left

(c) Not affect equilibrium

(d) Stop the reaction

**Q12. Conjugate base of $H_2SO_4$ is:**

(a) $SO_4^{2-}$

(b) $HSO_4^-$

(c) $H_3SO_4^+$

(d) $SO_3$

**Q13. Which is a Lewis acid?**

(a) $NH_3$

(b) $BF_3$

(c) $H_2O$

(d) $OH^-$

**Q14. pOH of 0.01 M HCl is:**

(a) 2

(b) 12

(c) 7

(d) 14

**Q15. For an endothermic reaction, increasing temperature:**

(a) Increases K

(b) Decreases K

(c) K remains same

(d) Reaction stops

**Q16. Salt hydrolysis occurs in:**

(a) NaCl solution

(b) $NH_4Cl$ solution

(c) $KNO_3$ solution

(d) $Na_2SO_4$ solution

**Q17. Catalyst affects equilibrium by:**

(a) Increasing K

(b) Decreasing activation energy

(c) Changing equilibrium position

(d) Increasing temperature

**Q18. pH of neutral solution at 60°C is:**

(a) 7

(b) Less than 7

(c) More than 7

(d) 14

**Q19. For reaction at equilibrium, ΔG is:**

(a) Positive

(b) Negative

(c) Zero

(d) Infinite

**Q20. Solubility of AgCl decreases in presence of:**

(a) NaCl

(b) $NaNO_3$

(c) $KNO_3$

(d) $Na_2SO_4$

## ✅ ANSWER KEY:

1.(b) 2.(b) 3.(b) 4.(c) 5.(b) 6.(a) 7.(b) 8.(b) 9.(b) 10.(c)

11.(b) 12.(b) 13.(b) 14.(b) 15.(a) 16.(b) 17.(b) 18.(b) 19.(c) 20.(a)

## 📚 CASE-BASED QUESTIONS (As per New CBSE

# Pattern)

### Case Study 1: Industrial Ammonia Production (Haber Process)

**Passage:**

Fritz Haber developed a process for ammonia synthesis which revolutionized agriculture. The reaction is:

**$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$; $\Delta H = -92.4$ kJ/mol**

Operating conditions: ~450°C temperature, 200 atm pressure, $Fe_3O_4$ with $K_2O$ and $Al_2O_3$ as catalyst. At equilibrium, about 15% conversion is achieved. The yield is increased by continuous removal of $NH_3$ through liquefaction. High pressure favors forward reaction as $\Delta n$ is negative. However, very high temperature would decrease yield (exothermic), so a compromise temperature of 450°C is used to balance rate and yield.

**Q1. What is the value of $\Delta n$ for this reaction?**

(a) +2

(b) -2

(c) 0

(d) +4

**Q2. Why is 450°C used instead of lower temperature?**

(a) K is maximum at 450°C

(b) Catalyst works only at 450°C

(c) Compromise between yield and rate

(d) $NH_3$ decomposes below 450°C

**Q3. Continuous removal of $NH_3$ helps because:**

(a) It increases $K_c$

(b) It shifts equilibrium forward (Le Chatelier)

(c) It decreases pressure

(d) It activates catalyst

**Q4. The role of $Fe_3O_4$ catalyst is to:**

(a) Increase yield

(b) Shift equilibrium position

(c) Speed up equilibrium attainment

(d) Change $\Delta H$ value

**Answers:** 1.(b) 2.(c) 3.(b) 4.(c)

## Case Study 2: Buffer Solutions in Human Blood

**Passage:**

Human blood maintains pH at 7.4 ± 0.05, which is crucial for survival. Any deviation can be life-threatening. The main buffer system is the carbonic acid-bicarbonate system:

**$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$; $pK_a = 6.1$**

The ratio $[HCO_3^-]/[H_2CO_3]$ is maintained at 20:1 in healthy blood. Lungs control $CO_2$ concentration (hence $H_2CO_3$) through respiration, while kidneys control $HCO_3^-$ concentration through excretion/reabsorption. When pH < 7.35, the condition is called acidosis (blood becomes too acidic). When pH > 7.45, it's called alkalosis (blood becomes too basic). Both conditions are potentially fatal if not corrected.

**Q1. Using Henderson-Hasselbalch equation, blood pH is:**

(a) 6.1

(b) 7.4

(c) 8.0

(d) 7.0

**Q2. Excess $H^+$ in blood is neutralized by:**

(a) $H_2CO_3$

(b) $HCO_3^-$

(c) $CO_2$

(d) $H_2O$

### Q3. Acidosis occurs when:

(a) $[HCO_3^-]$ increases

(b) $[H_2CO_3]$ decreases

(c) $[HCO_3^-]/[H_2CO_3]$ ratio decreases

(d) pH increases above 7.45

### Q4. Blood buffer system is:

(a) Acidic buffer (weak acid + salt)

(b) Basic buffer (weak base + salt)

(c) Neutral buffer

(d) Not a buffer system

**Answers:** 1.(b) 2.(b) 3.(c) 4.(a)

**Solution for Q1:** $pH = pK_a + \log([HCO_3^-]/[H_2CO_3]) = 6.1 + \log(20/1) = 6.1 + 1.3 = 7.4$

## Case Study 3: Solubility Product and Precipitation

**Passage:**

The solubility product principle is widely used in qualitative analysis for separation of cations. AgCl, $PbCl_2$, and $Hg_2Cl_2$ form Group I precipitates. When dilute HCl is added to a solution containing $Ag^+$, $Pb^{2+}$, and $Hg_2^{2+}$ ions, all three chlorides precipitate out due to common ion effect. The $K_{sp}$ values are: AgCl = $1.0 \times 10^{-10}$, $PbCl_2$ = $1.7 \times 10^{-5}$, $Hg_2Cl_2$ = $1.3 \times 10^{-18}$. Lower the $K_{sp}$, less soluble is the salt. Ionic product (Q) is compared with $K_{sp}$ to predict precipitation: If $Q > K_{sp}$, precipitation occurs; If $Q < K_{sp}$, no precipitation; If $Q = K_{sp}$, saturated solution.

**Q1. Which salt is least soluble?**

(a) AgCl

(b) $PbCl_2$

(c) $Hg_2Cl_2$

(d) All equally soluble

**Q2. Precipitation occurs when:**

(a) $Q < K_{sp}$

(b) $Q > K_{sp}$

(c) $Q = K_{sp}$

(d) $Q = 0$

**Q3. Common ion effect is used to:**

(a) Increase solubility

(b) Cause precipitation by decreasing solubility

(c) Has no effect on solubility

(d) Increase $K_{sp}$

**Q4. $K_{sp}$ expression for $PbCl_2$ is:**

(a) $[Pb^{2+}][Cl^-]$

(b) $[Pb^{2+}][Cl^-]^2$

(c) $[Pb^{2+}]^2[Cl^-]$

(d) $[Pb^{2+}][Cl^-]^3$

**Answers:** 1.(c) 2.(b) 3.(b) 4.(b)

# 💡 MEMORY TRICKS & MNEMONICS

## 🎯 Mnemonic for Strong Acids

**"HI! He's Not Br (Brother), Cl (Claude) Saw Her"**

- **HI** - Hydroiodic acid
- **He's** - ($H_2SO_4$) Sulfuric acid
- **N**ot - ($HNO_3$) Nitric acid
- **Br** - (HBr) Hydrobromic acid
- **Cl** - (HCl) Hydrochloric acid
- **Her** - ($HClO_4$) Perchloric acid

## 🎯 Le Chatelier's Principle Memory

**"System FIGHTS the Change"**

- ✓ Add reactant → Shifts RIGHT (fights by consuming it)
- ✓ Add product → Shifts LEFT (fights by consuming it)
- ✓ Increase pressure → Favors side with FEWER gas moles
- ✓ Increase temperature → Favors ENDOTHERMIC direction
- ✓ Decrease temperature → Favors EXOTHERMIC direction

## 🎯 pH Scale Quick Reference

**"Below 7 is ACID, Above 7 is BASIC, At 7 is NEUTRAL"**

| pH Range | Nature | Example |
|---|---|---|
| 0-3 | Strong Acid | HCl, Gastric juice (pH 1-2) |
| 4-6 | Weak Acid | Vinegar (pH 3), Lemon juice (pH 2) |
| 7 | Neutral | Pure water, Blood (pH 7.4) |
| 8-10 | Weak Base | Baking soda (pH 9), Milk of magnesia |
| 11-14 | Strong Base | NaOH, Bleach (pH 13) |

## 🎯 $K_c$ and $K_p$ Relationship Trick

**"K Powerful = K Clear × Mighty (RT) raised to Delta-n"**

$K_p = K_c(RT)^{\Delta n}$

**Quick Check:**

- If $\Delta n = 0 \rightarrow K_p = K_c$ (No conversion needed!)
- If $\Delta n$ is negative $\rightarrow K_p < K_c$
- If $\Delta n$ is positive $\rightarrow K_p > K_c$

# 📚 QUICK REVISION NOTES

## 🔑 Key Definitions (Must Know Word-to-Word)

**1. Chemical Equilibrium:** State in a reversible reaction when rate of forward reaction equals rate of backward reaction and concentrations remain constant.

**2. Law of Mass Action:** Rate of chemical reaction is proportional to the product of active masses of reactants raised to their stoichiometric coefficients.

**3. Equilibrium Constant (K):** Ratio of product of concentrations of products to product of concentrations of reactants, each raised to their stoichiometric coefficients.

**4. Le Chatelier's Principle:** If a system at equilibrium is subjected to a change, the equilibrium shifts in a direction that tends to counteract that change.

**5. Buffer Solution:** Solution that resists change in pH upon addition of small amounts of acid or base.

**6. Common Ion Effect:** Suppression of ionization of weak electrolyte by adding a strong electrolyte having a common ion.

**7. Solubility Product ($K_{sp}$):** Product of molar concentrations of ions in saturated solution of sparingly soluble salt, each raised to power of its coefficient.

**8. Ionic Product (Q):** Product of molar concentrations of ions at any instant (not necessarily at equilibrium).

**9. pH:** Negative logarithm of hydrogen ion concentration. pH = $-\log[H^+]$

**10. Conjugate Acid-Base Pair:** Pair of species differing by a proton. Acid-$H^+$ = Conjugate base

## ⚡ One-Line Facts (Speed Revision)

- ✓ Equilibrium can only be achieved in closed system
- ✓ Catalyst speeds up equilibrium but doesn't change position
- ✓ K depends only on temperature, not on concentration or pressure
- ✓ pH + pOH = 14 (at 25°C)
- ✓ $K_p = K_c(RT)^{\Delta n}$
- ✓ For pure solids and liquids, concentration is not included in K expression
- ✓ If K > 1, products are favored at equilibrium
- ✓ If K < 1, reactants are favored at equilibrium
- ✓ Common ion effect decreases solubility
- ✓ Buffer pH = $pK_a$ + log([Salt]/[Acid])
- ✓ Precipitation occurs when Q > $K_{sp}$
- ✓ $[H^+][OH^-] = 10^{-14}$ at 25°C ($K_w$)
- ✓ Strong acid: pH < 3, Strong base: pH > 11
- ✓ Exothermic: K decreases with T, Endothermic: K increases with T
- ✓ Blood pH = 7.4 (maintained by $H_2CO_3/HCO_3^-$ buffer)

## 📐 Formula Sheet (All Important Formulas)

| Concept | Formula |
|---|---|
| Equilibrium Constant | $K_c$ = [Products]/[Reactants] (with powers) |
| $K_p$ and $K_c$ | $K_p = K_c(RT)^{\Delta n}$ |
| pH | $pH = -\log[H^+]$ |
| pOH | $pOH = -\log[OH^-]$ |
| pH + pOH | pH + pOH = 14 |
| Weak Acid pH | $[H^+] = \sqrt{(K_a \times c)}$ |
| Weak Base pOH | $[OH^-] = \sqrt{(K_b \times c)}$ |
| Buffer pH | $pH = pK_a + \log([Salt]/[Acid])$ |
| $K_w$ | $[H^+][OH^-] = 10^{-14}$ at 25°C |
| Degree of Dissociation | α = (Amount dissociated)/(Initial amount) |

## 🎯 Top 10 Question Types (Practice These!)

1. Calculate $K_c$ or $K_p$ from equilibrium concentrations/pressures
2. Convert between $K_c$ and $K_p$ using Δn
3. Calculate pH of strong acids/bases
4. Calculate pH of weak acids/bases
5. Apply Le Chatelier's principle to predict equilibrium shift
6. Calculate buffer pH using Henderson-Hasselbalch equation
7. Calculate degree of dissociation (α)
8. Calculate solubility from $K_{sp}$
9. Predict precipitation using ionic product (Q) and $K_{sp}$
10. Calculate pH of mixed acid-base solutions

## ❌ Common Mistakes to Avoid

- **Mistake 1:** Including concentration of pure solids/liquids in K expression
- **Mistake 2:** Forgetting to raise concentrations to their stoichiometric powers
- **Mistake 3:** Confusing $K_c$ and $K_p$ conversion (forgetting RT or Δn)
- **Mistake 4:** Not using approximation for weak acids (assuming c ≈ c-α when α << 1)
- **Mistake 5:** Wrong sign in Henderson-Hasselbalch equation
- **Mistake 6:** Thinking catalyst changes equilibrium position (it only speeds up!)
- **Mistake 7:** Confusing pH and pOH formulas
- **Mistake 8:** Not checking units (pressure in atm, concentration in M)

## 📚 Study Material Information

This comprehensive study material on **Equilibrium** has been meticulously prepared following the latest CBSE curriculum and examination pattern for Class 11 Chemistry (2025-26 session). The content includes detailed explanations of all topics including chemical equilibrium concept, law of mass action, equilibrium constant ($K_c$ and $K_p$), homogeneous and heterogeneous equilibrium, Le Chatelier's principle (effect of concentration, pressure, temperature, and catalyst), ionic equilibrium in aqueous solutions, acids and bases (Arrhenius, Bronsted-Lowry, Lewis theories), pH scale calculations, strong and weak acids/bases, buffer solutions with Henderson-Hasselbalch equation, common ion effect, salt hydrolysis, solubility product ($K_{sp}$), solved numerical problems with step-by-step solutions, 20 CBSE-pattern MCQs with detailed answers, 3 comprehensive case-based questions covering Haber process, blood buffer system, and precipitation reactions, memory tricks and mnemonics, quick revision notes, complete formula sheet, and exam strategy with common mistakes to avoid.

**Key Features of This Material:**

- ✅ Complete Chapter 6 coverage with crystal-clear concepts
- ✅ Both physical and chemical equilibrium explained
- ✅ Law of mass action and equilibrium constant derivation
- ✅ $K_c$ and $K_p$ relationship with Δn calculations
- ✅ Le Chatelier's principle with industrial applications
- ✅ Haber process - most important for boards!
- ✅ Complete ionic equilibrium with pH calculations
- ✅ All three acid-base theories (Arrhenius, Bronsted-Lowry, Lewis)
- ✅ Strong acids vs weak acids - calculations made easy
- ✅ Buffer solutions and Henderson-Hasselbalch equation
- ✅ Common ion effect with examples
- ✅ Solubility product ($K_{sp}$) and precipitation
- ✅ 20 MCQs + 3 case studies as per latest CBSE pattern
- ✅ Step-by-step solved numericals (10+ problems)
- ✅ Memory tricks for strong acids and Le Chatelier
- ✅ Common mistakes to avoid (very important!)
- ✅ Complete formula sheet for quick revision
- ✅ Top 10 question types for practice

### 🎯 Why This Chapter is SUPER Important:

Equilibrium is the **foundation of chemical reactions!** Understanding this chapter is essential for:

- ✓ Chemical Kinetics (Class 12 - next related chapter!)
- ✓ Electrochemistry (Class 12)
- ✓ Understanding industrial processes (Haber, Contact process)
- ✓ Biological systems (blood pH, enzyme reactions)
- ✓ JEE & NEET (direct 5-6 questions every year!)
- ✓ Laboratory techniques (qualitative analysis, titrations)

**Master equilibrium = Understanding all of chemistry becomes easy!**

Content is based on NCERT syllabus and CBSE guidelines for 2025-26.

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*Disclaimer: This material is prepared as a comprehensive study aid for Class 11 students. While every effort has been made to ensure accuracy and alignment with CBSE curriculum, students are advised to refer to their NCERT textbooks and official CBSE guidelines for examination preparation. This material covers Chapter 6: Equilibrium from Class 11 Chemistry NCERT textbook (Reprint 2025-26). All formulas, definitions, and numerical values are as per latest NCERT edition. Special focus on Le Chatelier's principle and pH calculations as these are most important for numerical problems in boards and competitive exams.*

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