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CHAPTER NOTES

Equilibrium Class 11 Notes | CBSE Chemistry Chapter 6

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Equilibrium Class 11 Notes | CBSE Chemistry Chapter 6

Equilibrium is Chapter 6 of CBSE Class 11 Chemistry — and one of the highest-scoring, highest-yield chapters of the entire year. It explains *why* reversible reactions seem to "stop" even when reactants are still present, and gives you the tools (Kc, Kp, Q, Le Chatelier's principle, pH, Ksp) to predict and control them. This single chapter unlocks ionic equilibrium, and is a direct prerequisite for Class 12 Electrochemistry.

By the end of these notes you will be able to write equilibrium constant expressions, relate Kc and Kp, use the reaction quotient Q to predict the direction of a reaction, apply Le Chatelier's principle to shift equilibria, calculate pH of acids and bases, design a buffer, and solve solubility-product numericals. This is a high-weightage chapter carrying roughly 6–8 marks in boards and a heavy share of NEET/JEE Physical Chemistry.

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Key Concepts

1. Equilibrium — Physical and Chemical

Equilibrium is the state of a reversible process at which the rate of the forward process equals the rate of the backward process, so the observable properties become constant with time.

It is **dynamic**, not static — both processes keep happening, just at equal rates. Equilibrium is possible only in a **closed system** at constant temperature.

Physical Equilibrium

- **Solid** ⇌ **Liquid** (melting/freezing at the melting point): rate of melting = rate of freezing.
- **Liquid** ⇌ **Vapour** (in a closed vessel): vapour pressure becomes constant.
- **Solid/Gas** ⇌ **Solution** (dissolution of sugar or CO_2 in a sealed bottle): concentration stops changing.

Chemical Equilibrium

In a reversible reaction such as $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, reactant and product concentrations become constant once the forward and backward rates are equal.

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2. Law of Chemical Equilibrium and Equilibrium Constant

The **law of mass action** states that the rate of a reaction is proportional to the product of the molar concentrations of the reactants, each raised to the power of its stoichiometric coefficient.

For a general reaction **aA + bB ⇌ cC + dD**, the equilibrium constant in terms of concentration is:

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

- Kc is constant at a given temperature.
- A large Kc (much greater than 1) means products are favoured; a small Kc (much less than 1) means reactants are favoured.
- Pure solids and pure liquids are **not** included (their activity = 1).

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3. Equilibrium Constant in Gaseous Systems (Kp) and the Kc–Kp Relation

For gaseous reactions it is convenient to express the equilibrium constant in terms of partial pressures. For aA + bB ⇌ cC + dD:

$K_p = (p_C)^c (p_D)^d / (p_A)^a (p_B)^b$

Using the ideal gas equation, Kp and Kc are related by:

Kp = Kc (RT)^Δn

- **Δn = (moles of gaseous products) − (moles of gaseous reactants)**
- If Δn = 0, then **Kp = Kc** (e.g. $H_2 + I_2 \rightleftharpoons 2HI$).
- If Δn is positive, Kp is greater than Kc; if Δn is negative, Kp is less than Kc.
- R = 0.0821 L·atm·K⁻¹·mol⁻¹ when pressures are in atm.

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4. Characteristics of the Equilibrium Constant

- For the **reverse** reaction, K′ = 1/K.
- If a reaction is **multiplied by n**, the new constant is K^n.
- If two reactions are **added**, $K = K_1 \times K_2$.
- K depends only on **temperature** — not on the initial concentrations, pressure, or a catalyst.
- A catalyst speeds up forward and backward rates equally, so it only reaches equilibrium faster — it does not change K.

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5. Reaction Quotient (Q) and Predicting Direction

The **reaction quotient Q** has the same form as Kc but uses concentrations at *any* instant, not just at equilibrium. Comparing Q with K tells you which way the reaction will move.

- **Q less than K:** too many reactants → reaction proceeds **forward** (towards products).
- **Q = K:** system is **at equilibrium**.
- **Q greater than K:** too many products → reaction proceeds **backward** (towards reactants).

Link to thermodynamics: **$\Delta G = \Delta G^\circ + RT \ln Q$**, and at equilibrium $\Delta G = 0$, giving **$\Delta G^\circ = -RT \ln K$**.

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6. Le Chatelier's Principle and Factors Affecting Equilibrium

Le Chatelier's principle: if a system at equilibrium is disturbed by a change in concentration, pressure, or temperature, the equilibrium shifts in the direction that tends to undo the change.

Effect of Each Factor

Change	Equilibrium shifts	Example ($N_2 + 3H_2 \rightleftharpoons 2NH_3$, exothermic)
Add reactant / remove product	Forward (towards products)	Adding N_2 or H_2 makes more NH_3
Increase pressure (decrease volume)	Towards fewer gas moles	Shifts forward (4 mol → 2 mol)
Increase temperature	Towards endothermic direction	Shifts backward (less NH_3)
Add catalyst	No shift	Equilibrium reached faster only
Add inert gas at constant volume	No shift	Partial pressures unchanged

Industrial use: the Haber process (NH_3) uses high pressure and moderate temperature; the Contact process (SO_3) uses similar reasoning.

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7. Ionic Equilibrium — Strong and Weak Electrolytes

Ionic equilibrium is the equilibrium established between unionised molecules and their ions in solution.

- **Strong electrolytes** (NaCl, HCl, NaOH) ionise almost completely — no real equilibrium.
- **Weak electrolytes** (CH_3COOH, NH_4OH) ionise only partially, setting up a genuine ionic equilibrium.

Ostwald's dilution law (for a weak electrolyte, degree of dissociation α): $K_a = C\alpha^2/(1 - \alpha)$, which is approximately $C\alpha^2$ when α is small, so **$\alpha = \sqrt{(K_a/C)}$**. Dilution increases α.

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8. Acids and Bases — Three Concepts

Concept	Acid	Base
Arrhenius	Gives H^+ in water	Gives OH^- in water
Brønsted–Lowry	Proton (H^+) donor	Proton (H^+) acceptor
Lewis	Electron-pair acceptor	Electron-pair donor

In the Brønsted–Lowry view, every acid has a **conjugate base** (formed by losing H^+) and every base has a **conjugate acid**. Example: in $HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$, the pairs HCl/Cl^- and H_2O/H_3O^+ are conjugate pairs. A strong acid has a weak conjugate base.

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9. Ionization of Water, pH Scale and Kw

Water self-ionises: $H_2O \rightleftharpoons H^+ + OH^-$. The **ionic product of water** is:

$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ at 298 K

The **pH** measures acidity: **$pH = -\log[H^+]$**, and similarly **$pOH = -\log[OH^-]$**.

- **$pH + pOH = 14$** at 298 K.
- Neutral: pH = 7; Acidic: pH below 7; Basic: pH above 7.
- For a strong acid, $[H^+]$ = its molarity; for a strong base, $[OH^-]$ = its molarity.

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10. Ionization Constants of Weak Acids and Bases (Ka, Kb)

For a weak acid $HA \rightleftharpoons H^+ + A^-$: **$K_a = [H^+][A^-]/[HA]$**. For a weak base $BOH \rightleftharpoons B^+ + OH^-$: **$K_b = [B^+][OH^-]/[BOH]$**.

- A larger Ka (or Kb) means a stronger acid (or base).
- **$pK_a = -\log K_a$**; a smaller pKa means a stronger acid.
- For a conjugate acid–base pair: **$K_a \times K_b = K_w$**, so **$pK_a + pK_b = 14$**.
- For a weak acid, $[H^+] = \sqrt{(K_a \cdot C)}$, so $pH = \frac{1}{2}(pK_a - \log C)$.

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11. Common Ion Effect and Buffer Solutions

The **common ion effect** is the suppression of the ionisation of a weak electrolyte by adding a strong electrolyte that shares a common ion. For example, adding CH_3COONa to CH_3COOH lowers the H^+ concentration.

A **buffer solution** resists changes in pH on adding small amounts of acid or base.

- **Acidic buffer:** weak acid + its salt ($CH_3COOH + CH_3COONa$). **$pH = pK_a + \log([salt]/[acid])$** (Henderson–Hasselbalch equation).

- **Basic buffer:** weak base + its salt ($NH_4OH + NH_4Cl$). **pOH = pKb + log([salt]/[base])**.

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12. Solubility Product (Ksp) and Salt Hydrolysis

For a sparingly soluble salt $A_xB_y \rightleftharpoons xA^{y+} + yB^{x-}$, the **solubility product** is:

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

- For AB type (e.g. AgCl) with solubility s: **$Ksp = s^2$**.
- For AB_2 type (e.g. CaF_2) with solubility s: **$Ksp = 4s^3$**.
- If **ionic product is greater than Ksp** → precipitation occurs; if **less than Ksp** → more salt dissolves.

Hydrolysis of Salts

Salt type	Example	Nature of solution
Strong acid + strong base	NaCl	Neutral (pH = 7)
Strong acid + weak base	NH_4Cl	Acidic (pH below 7)
Weak acid + strong base	CH_3COONa	Basic (pH above 7)
Weak acid + weak base	CH_3COONH_4	Depends on Ka vs Kb

Weightage in Board & Entrance Exams

Exam	Typical Weightage	Most-Tested Areas
CBSE Board (Class 11)	6–8 marks	Kc/Kp relation, Le Chatelier, pH, buffers, Ksp
JEE Main / Advanced	2–3 questions	Q vs K, ICE-table numericals, Ksp, buffer pH
NEET	2–3 questions	Le Chatelier, pH/pOH, common ion effect, salt hydrolysis

[TABLE: Question-type split — VSA (1 mark): definitions, conjugate pairs, pH; SA (2–3 marks): Kp–Kc, Le Chatelier shifts, buffer/Henderson; LA (5 marks): Ksp numericals, ionic-equilibrium derivations, pH calculations.]

Important Definitions

Term	Definition
Chemical equilibrium	State where forward and backward reaction rates are equal; concentrations stay constant
Equilibrium constant (Kc)	Kc = [products]/[reactants], each raised to its coefficient, at a fixed temperature
Kp–Kc relation	Kp = Kc(RT)^Δn, where Δn = gaseous products – gaseous reactants
Reaction quotient (Q)	Same form as Kc but at any instant; predicts direction by comparing with K
Le Chatelier's principle	A disturbed equilibrium shifts to oppose (undo) the change imposed
Brønsted acid/base	Acid = proton donor; base = proton acceptor
Lewis acid/base	Acid = electron-pair acceptor; base = electron-pair donor
pH	pH = –log[H^+]; measures acidity (pH + pOH = 14 at 298 K)
Kw	Ionic product of water: [H^+][OH^-] = 1.0×10^{-14} at 298 K
Common ion effect	Suppression of a weak electrolyte's ionisation by adding a common ion
Buffer solution	Solution that resists pH change on adding small amounts of acid/base
Solubility product (Ksp)	Product of ionic molar concentrations of a saturated sparingly-soluble salt

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Solved Examples

Example 1

For the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, find the relation between Kp and Kc.

Answer: Δn = 2 – (1 + 3) = –2. So **$Kp = Kc(RT)^{-2}$**, i.e. $Kp = Kc/(RT)^2$.

Example 2

In the reaction $H_2 + I_2 \rightleftharpoons 2HI$, equilibrium concentrations are [H_2] = 0.5 M, [I_2] = 0.5 M, [HI] = 2 M. Find Kc.

Answer: $Kc = [HI]^2/([H_2][I_2]) = (2)^2/(0.5 \times 0.5) = 4/0.25 =$ **16**.

Example 3

Calculate the pH of a 0.001 M HCl solution.

Answer: HCl is a strong acid, so $[H^+] = 0.001 = 10^{-3}$ M. pH = $-\log(10^{-3})$ = **3**.

Example 4

The Ka of acetic acid is 1.8×10^{-5}. Find the pH of a 0.1 M solution.

Answer: $[H^+] = \sqrt{(Ka \cdot C)} = \sqrt{(1.8 \times 10^{-5} \times 0.1)} = \sqrt{(1.8 \times 10^{-6})} = 1.34 \times 10^{-3}$ M. pH = $-\log(1.34 \times 10^{-3})$, which is approximately **2.87**.

Example 5

A buffer is made of 0.2 M CH_3COOH and 0.2 M CH_3COONa (pKa = 4.74). Find its pH.

Answer: pH = pKa + log([salt]/[acid]) = 4.74 + log(0.2/0.2) = 4.74 + 0 = **4.74**.

Example 6

The solubility of AgCl is 1.0×10^{-5} mol/L. Calculate its Ksp.

Answer: $AgCl \rightleftharpoons Ag^+ + Cl^-$, so $Ksp = s^2 = (1.0 \times 10^{-5})^2 = \mathbf{1.0 \times 10^{-10}}$.

Important Questions for Board Exams

1-Mark Questions (VSA)

1. Why is chemical equilibrium called dynamic?
2. Write the conjugate base of H_2SO_4 and the conjugate acid of NH_3.
3. What is the value of Kw at 298 K?
4. What happens to the pH of water as temperature increases? Justify briefly.
5. Does a catalyst change the value of the equilibrium constant? Explain.

2–3-Mark Questions (SA)

1. Derive the relation Kp = Kc(RT)^Δn for a gaseous reaction.
2. State Le Chatelier's principle and apply it to the Haber process ($N_2 + 3H_2 \rightleftharpoons 2NH_3$) for pressure and temperature.
3. What is a buffer solution? Derive the Henderson–Hasselbalch equation for an acidic buffer.
4. Explain the common ion effect with a suitable example and its role in salt analysis.

5-Mark Questions (LA)

1. Distinguish between Arrhenius, Brønsted–Lowry, and Lewis concepts of acids and bases with examples.

2. Define solubility product. Derive the relation between Ksp and solubility for AB and AB_2 type salts, and explain precipitation using the ionic product.
3. Explain salt hydrolysis. Discuss the nature (acidic/basic/neutral) of solutions of NaCl, NH_4Cl, and CH_3COONa with reasons.

Quick Revision Points

- Equilibrium is dynamic: forward rate = backward rate in a closed system at constant T
- Kc = [products]/[reactants] (powers = coefficients); pure solids/liquids excluded
- Kp = Kc(RT)^Δn; Δn = gaseous products – gaseous reactants; if Δn = 0, Kp = Kc
- Reverse reaction: K′ = 1/K; multiply by n: K^n; add reactions: K = $K_1 \times K_2$
- Q below K → forward; Q = K → equilibrium; Q above K → backward; $\Delta G^\circ = -RT \ln K$
- Le Chatelier: system shifts to undo the change in concentration/pressure/temperature
- Acid/base: Arrhenius (H^+/OH^-), Brønsted (proton donor/acceptor), Lewis (electron-pair acceptor/donor)
- pH = $-\log[H^+]$; pH + pOH = 14; Kw = 10^{-14} at 298 K
- Ka × Kb = Kw; pKa + pKb = 14; weak acid $[H^+] = \sqrt{(Ka \cdot C)}$
- Buffer: pH = pKa + log([salt]/[acid]) (Henderson–Hasselbalch)
- Ksp: AB type = s^2; AB_2 type = $4s^3$; ionic product above Ksp → precipitation
- Hydrolysis: SA+SB neutral, SA+WB acidic, WA+SB basic

Next Chapter: Chapter 7 — Redox Reactions