EDUVISTA

CLASS 11 • CHEMISTRY

UNIT 6 — EQUILIBRIUM

COMPLETE EXAM-READY NOTES + FORMULA SHEET + NUMERICAL STRATEGY

What this guide gives you

Concepts • Definitions • Formulae • Direction of equilibrium • K_c/K_p • Le Chatelier • Ionic equilibrium • pH • Buffers • K_{sp} • Exam traps

SMART REVISION • CLEAR CONCEPTS • EXAM FOCUS

Prepared by Abhishek Sir
8920696949 • Eduvista Classes

1. CHAPTER MAP — WHAT YOU MUST MASTER

NCERT Unit 6: Equilibrium — from physical equilibrium to ionic equilibrium and solubility.

PHYSICAL EQUILIBRIUM

- Solid ⇌ liquid
- Liquid ⇌ vapour
- Solid ⇌ vapour
- Dissolution of solids
- Dissolution of gases
- Henry's law
- Characteristics of physical equilibrium

CHEMICAL EQUILIBRIUM

- Dynamic equilibrium
- Law of mass action
- Kc and Kp
- Homogeneous & heterogeneous equilibria
- Applications of K
- Reaction quotient Q
- ΔG° and K
- Le Chatelier's principle

IONIC EQUILIBRIUM

- Electrolytes
- Arrhenius / Brønsted–Lowry / Lewis
- Ka, Kb
- Degree of ionisation
- Common-ion effect
- Polyprotic acids
- pH, pOH, Kw
- Salt hydrolysis

ADVANCED EXAM AREAS

- Buffer solutions
- Henderson–Hasselbalch equation
- Solubility equilibria
- Ksp
- Molar solubility
- Ionic product Qsp
- Precipitation
- NCERT-style numericals

EDUVISTA EXAM RULE

Do not study Equilibrium as isolated formulas. First identify the equilibrium, write the balanced equation, then choose the correct expression and finally apply the appropriate comparison or numerical method.

2. PHYSICAL EQUILIBRIUM

Equilibrium in physical processes is dynamic: the opposing processes continue at equal rates.

Core idea

At equilibrium, measurable macroscopic properties remain constant, but microscopic processes continue. For a physical process, the system is in equilibrium when the opposing rates become equal.

Process	Equilibrium	Exam point
Solid–liquid	Ice ⇌ water	At fixed pressure, melting and freezing rates become equal
Liquid–vapour	Liquid ⇌ vapour	Rate of evaporation = rate of condensation
Solid–vapour	$I_2(s) \rightleftharpoons I_2(g)$	Sublimation and deposition continue
Dissolution	Solute ⇌ dissolved state	Saturated solution: dissolution and crystallisation balance
Gas in liquid	Gas ⇌ dissolved gas	Solubility depends on pressure and temperature

General characteristics

- Equilibrium is possible only in a closed system at a given temperature.
- Opposing processes occur at the same rate.
- The condition is dynamic but stable.
- Measurable properties remain constant.
- The equilibrium state is characterised by a constant parameter at a given temperature.

VOCABULARY

Dynamic equilibrium ≠ reaction stops. It means the forward and reverse processes continue simultaneously at equal rates.

3. VAPOUR PRESSURE, BOILING & HENRY'S LAW

High-yield physical-equilibrium relationships.

Liquid–vapour equilibrium

In a closed container, evaporation increases the vapour phase while condensation removes molecules from it. At equilibrium, the rate of evaporation equals the rate of condensation and vapour pressure becomes constant.

VAPOUR PRESSURE

- Depends on temperature.
- At higher temperature, more molecules escape into vapour.
- Boiling occurs when vapour pressure equals external pressure.
- Higher external pressure → higher boiling temperature.

HENRY'S LAW

For a gas dissolved in a liquid, the chapter discusses the dependence of dissolved-gas concentration on pressure and temperature through Henry's law.

Exam focus: know the statement, variables, and qualitative pressure/temperature effect rather than memorising a symbol without context.

APPLICATIONS

Carbon dioxide in soda water • gases dissolved in liquids • boiling at altitude • vapour-pressure reasoning

Quick check

- Why does vapour pressure become constant at equilibrium? Because evaporation and condensation rates become equal.
- What happens when the container volume is suddenly increased? Initially vapour pressure falls; evaporation then increases until equilibrium is restored.

4. CHEMICAL EQUILIBRIUM — DYNAMIC NATURE

The central idea behind the whole chapter.

For a reversible reaction

$$A + B \rightleftharpoons C + D$$

Initially, the forward rate is high and the reverse rate is low. As products accumulate, the forward rate falls and the reverse rate rises. At equilibrium, the two rates become equal and concentrations remain constant.

AT EQUILIBRIUM

- Forward rate = reverse rate
- Reactant and product concentrations are constant
- Reaction does not stop
- Equilibrium can be reached from either direction

HABER-PROCESS IDEA

The chapter uses ammonia synthesis to demonstrate dynamic equilibrium. Starting with reactants or products can lead to the same equilibrium composition under the same conditions.

GRAPH MEMORY

Concentration–time graph: reactants decrease and products increase until both become constant.
Rate–time graph: forward and reverse rates approach the same value.

Must not confuse

- Constant concentration ≠ zero concentration.
- Equal rates ≠ equal concentrations.
- Equilibrium constant tells equilibrium composition tendency, not the speed of reaching equilibrium.

5. LAW OF CHEMICAL EQUILIBRIUM & Kc

Write the balanced equation first; the stoichiometric coefficients become powers.

For the general reaction

$$aA + bB \rightleftharpoons cC + dD$$

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

The chapter presents this as the law of mass action / equilibrium constant expression.

Example

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

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

Change to equation	New equilibrium constant
Reverse reaction	$K' = 1/K$
Multiply equation by n	$K' = K^n$
Divide equation by n	K′ = K^(1/n)

EDUVISTA EXAM TRAP

Never copy K from the original equation after changing the reaction equation. Recalculate the new K using the stoichiometric change.

6. HOMOGENEOUS & HETEROGENEOUS EQUILIBRIA

Know what enters the equilibrium expression and what is omitted.

HOMOGENEOUS

All reacting species are in the same phase.

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

$K_c = [NH_3]^2 / ([N_2][H_2]^3)$

HETEROGENEOUS

More than one phase is present.

Example: $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

Pure solids and pure liquids are omitted from the equilibrium expression.

Why are pure solids/liquids omitted?

Their effective concentration/activity remains constant; therefore it is absorbed into the equilibrium constant.

Heterogeneous example

$Ni(s) + 4CO(g) \rightleftharpoons Ni(CO)_4(g)$

$K_p = p[Ni(CO)_4] / (p[CO])^4$

ONE-LINE RULE

Pure solid / pure liquid → do not write it in K expression. Gas / aqueous species → write it.

7. Kp, Kc & GAS EQUILIBRIA

One of the most important formula-based areas of Unit 6.

For gases

K_p is written using partial pressures instead of molar concentrations.

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

Δn = total moles of gaseous products − total moles of gaseous reactants

Reaction feature	Δn	Relationship
Equal gaseous moles on both sides	0	Kp = Kc
More gaseous moles in products	> 0	Kp = Kc(RT)^Δn
Fewer gaseous moles in products	< 0	Kp = Kc(RT)^Δn

Ideal-gas bridge

$$p = cRT$$

The chapter uses this relation to establish the Kp–Kc relationship.

NUMERICAL CHECKLIST

1) Balance equation → 2) Count gaseous moles → 3) Find Δn → 4) Use consistent R, T and pressure units → 5) Substitute carefully.

8. APPLICATIONS OF EQUILIBRIUM CONSTANT

K tells you where equilibrium lies — not how fast it is reached.

Magnitude of K	Equilibrium mixture
$K_c > 10^3$	Products predominate; reaction proceeds nearly to completion
$K_c < 10^{-3}$	Reactants predominate
10^{-3} to 10^3	Appreciable concentrations of both reactants and products

Important properties

- The equilibrium constant is applicable when equilibrium concentrations have become constant.
- At a fixed temperature, K has a unique value for a particular balanced reaction.
- K is independent of the initial concentrations.
- K for the reverse reaction is the inverse of the forward K.
- K does not tell the rate at which equilibrium is reached.

MEMORY HOOK

K = position/composition tendency; rate = speed. A catalyst can change the speed without changing K or equilibrium composition.

9. REACTION QUOTIENT Q — PREDICT THE DIRECTION

The fastest way to decide whether a mixture will move forward or backward.

For aA + bB ⇌ cC + dD

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

Q uses concentrations at the current time; they do not have to be equilibrium values.

Comparison	Direction	Meaning
Q < K	Forward → products	Products must increase until Q = K
Q > K	Reverse → reactants	Reactants must increase until Q = K
Q = K	No net reaction	Mixture is at equilibrium

NCERT-style mini example

For 2A ⇌ B + C, Kc = 2 × 10^{-3} and [A] = [B] = [C] = 3 × 10^{-4} M:

$$Q = ([B][C])/[A]^2 = 1$$

Since Q > K, the reaction proceeds in the reverse direction.

EXAM TRAP
Do not compare individual concentrations with K. Compare the complete reaction quotient Q with K.

10. CALCULATING EQUILIBRIUM CONCENTRATIONS

Use the stoichiometry-driven ICE method.

Five-step method

- 1. Write and balance the reaction.
- 2. Make an Initial–Change–Equilibrium table.
- 3. Let x represent the concentration change of one species.
- 4. Use stoichiometric coefficients to express every equilibrium concentration in x.
- 5. Substitute into K and solve; reject mathematically possible but chemically impossible roots.

Species	Initial	Change	Equilibrium
A	a	−x	a − x
B	b	−x	b − x
C	0	+x	x
D	0	+x	x

NCERT worked-pattern examples

- $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$: quadratic-equation approach.
- $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$: solve x, then calculate all equilibrium concentrations.
- $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$: use Q to first check whether the given composition is already at equilibrium.

11. ΔG, Q & K — THERMODYNAMIC VIEW

Connect equilibrium with Gibbs energy.

General relationship

$$\Delta G = \Delta G^\circ + RT \ln Q$$

At equilibrium: $\Delta G = 0$ and $Q = K$.

Therefore

$$\Delta G^\circ = -RT \ln K$$

$$K = e^{(-\Delta G^\circ/RT)}$$

ΔG°	K tendency	Interpretation
Negative	K > 1	Products are favoured
Positive	K < 1	Reactants are favoured
Zero	K = 1	Standard-state balance corresponds to K = 1

NCERT NUMERICAL PATTERN

If ΔG° is given in kJ mol^{-1}, convert it to J mol^{-1} before using R = 8.314 J mol^{-1} K^{-1}.

Chapter example

For glucose phosphorylation, $\Delta G^\circ = 13.8$ kJ mol^{-1} at 298 K; the chapter calculates $K_c = 3.81 \times 10^{-3}$.

12. LE CHATELIER'S PRINCIPLE — CONCENTRATION & PRESSURE

When equilibrium is disturbed, the system shifts to counteract the imposed change.

Concentration change

Change	Equilibrium shift
Add reactant	Toward products
Remove reactant	Toward reactants
Add product	Toward reactants
Remove product	Toward products

Pressure change for gaseous equilibrium

- Increase pressure → shift toward the side with fewer moles of gas.
- Decrease pressure → shift toward the side with more moles of gas.
- If the number of gaseous moles is the same on both sides, pressure change does not shift equilibrium.
- Pure solids and liquids do not determine gaseous-mole pressure shifts.

Example

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

Left side = 4 gaseous moles; right side = 2. Increasing pressure favours NH_3 formation.

THINK LIKE THE EQUILIBRIUM

Ask: “Which side reduces the disturbance?” That side is favoured.

13. TEMPERATURE, INERT GAS & CATALYST

Temperature changes equilibrium; a catalyst changes the speed of reaching it.

Temperature

Reaction type	Increase temperature
Endothermic	Favours forward direction
Exothermic	Favours reverse direction

Treat heat as a reactant for an endothermic reaction and as a product for an exothermic reaction.

Inert gas addition

The chapter distinguishes the effect according to the condition under which the gas is added. At constant volume, partial pressures of reacting gases are not changed by simply adding an inert gas, so equilibrium composition is not shifted by that addition.

Catalyst

- Provides a lower-activation-energy pathway.
- Increases forward and reverse reaction rates.
- Allows equilibrium to be reached faster.
- Does not change equilibrium composition.
- Does not change the equilibrium constant.

GOLDEN EXAM LINE

Temperature changes K. Catalyst does not change K.

14. IONIC EQUILIBRIUM & ELECTROLYTES

The chapter now moves from molecular equilibrium to equilibria involving ions.

Electrolytes

Type	Behaviour in water	Example
Strong electrolyte	Almost completely ionised/dissociated	NaCl
Weak electrolyte	Partially ionised/dissociated	CH_3COOH
Non-electrolyte	Does not conduct through ions	Sugar solution

Dissociation vs ionisation

- Dissociation: separation of ions that already exist in the solid, e.g. NaCl in water.
- Ionisation: formation of charged ions from a neutral molecule in solution, e.g. a weak acid.

Why water supports ions

The chapter discusses water's high dielectric constant and hydration of ions as important factors in separating and stabilising ions in aqueous solution.

CONCEPT CHECK

Strong electrolyte does not mean “strong acid” only. Strong acids, strong bases and soluble salts can behave as strong electrolytes.

15. ACIDS & BASES — THREE THEORIES

Know the definition and the limitation/application of each theory.

Theory	Acid	Base
Arrhenius	Produces H^+ in water	Produces OH^- in water
Brønsted–Lowry	Proton donor	Proton acceptor
Lewis	Electron-pair acceptor	Electron-pair donor

Arrhenius

$HX(aq) \rightarrow H^+(aq) + X^-(aq)$; in water the proton is represented through hydronium, H_3O^+.

Limitation: the concept is restricted to aqueous solutions and does not adequately account for bases such as ammonia.

Brønsted–Lowry

Acid = proton donor; base = proton acceptor. A conjugate acid–base pair differs by one proton.

Lewis

$BF_3 + NH_3 \rightarrow BF_3{:}NH_3$. BF_3 accepts an electron pair (Lewis acid); NH_3 donates it (Lewis base).

16. Ka, Kb & DEGREE OF IONISATION

Quantitative treatment of weak acids and weak bases.

Weak acid equilibrium

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

$$K_a = [H_3O^+][A^-] / [HA]$$

Weak base equilibrium

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$

$$K_\beta = [BH^+][OH^-] / [B]$$

Larger value	Meaning
Larger Ka	Stronger acid
Larger Kb	Stronger base
Smaller pKa	Stronger acid
Smaller pKb	Stronger base

Degree of ionisation

α = amount ionised / initial amount

For a weak acid: $K_a = C\alpha^2/(1-\alpha)$; when α is small, $K_a \approx C\alpha^2$.

Hence $\alpha \approx \sqrt{(K_a/C)}$: dilution increases the degree of ionisation.

17. COMMON-ION EFFECT & POLYPROTIC ACIDS

Two frequently tested ideas from ionic equilibrium.

Common-ion effect

Adding an ion already present in a weak-electrolyte equilibrium suppresses ionisation by shifting the equilibrium in the direction that consumes the added ion.

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$

Adding CH_3COO^- suppresses ionisation of CH_3COOH.

Polyprotic acids

$$H_2A \rightleftharpoons H^+ + HA^-$$

$$HA^- \rightleftharpoons H^+ + A^{2-}$$

$$\text{Key order: } K_a1 > K_a2 > K_a3$$

The chapter explains that removing a proton from an increasingly negatively charged species becomes progressively harder due to electrostatic effects.

EXAM TRAP

Do not assume all ionisation steps of a polyprotic acid have equal strength. The successive Ka values decrease.

18. pH, pOH, pKa, pKb & Kw

The formula network you should be able to reproduce without hesitation.

Quantity	Definition
pH	$-\log[H^+]$
pOH	$-\log[OH^-]$
pKa	$-\log Ka$
pKb	$-\log Kb$
pKw	$-\log Kw$

Ionisation of water

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$

$$K_w = [H^+][OH^-]$$

At 298 K: $K_w = 1.0 \times 10^{-14}$ and $pK_w = 14$.

Therefore

$$pH + pOH = pKw$$

$$\text{At } 25\ ^\circ C: pH + pOH = 14.$$

LOG MEMORY
Higher $[H^+]$ → lower pH. Higher Ka → lower pKa. Higher Kb → lower pKb.

19. pH NUMERICAL STRATEGY

Choose the acid/base strength first, then select the correct approximation.

Strong acid

$[H^+] \approx$ concentration of strong monoprotic acid; $pH = -\log[H^+]$.

Strong base

$[OH^-] \approx$ concentration of strong base; $pOH = -\log[OH^-]$; $pH = pKw - pOH$.

Weak acid

$$[H^+] \approx \sqrt{(Ka\ C)}$$

$$pH = \tfrac{1}{2}(pKa - \log C)$$

Weak base

$$[OH^-] \approx \sqrt{(Kb\ C)}$$

$$pOH = \tfrac{1}{2}(pKb - \log C)$$

NUMERICAL CHECK

Use the weak-electrolyte approximation only when the degree of ionisation is small. If the approximation is not justified, solve the equilibrium more exactly.

20. SALT HYDROLYSIS

Predict whether the salt solution is acidic, basic or approximately neutral.

Salt from	Example	Solution tendency
Strong acid + strong base	NaCl	Approximately neutral
Weak acid + strong base	CH_3COONa	Basic
Strong acid + weak base	NH_4Cl	Acidic
Weak acid + weak base	CH_3COONH_4	Depends on Ka and Kb

Weak acid + strong base salt

$$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$$

OH^- formation makes the solution alkaline.

Strong acid + weak base salt

$$NH_4^+ + H_2O \rightleftharpoons NH_4OH + H^+$$

H^+ formation makes the solution acidic.

Weak acid + weak base salt

$$pH = 7 + \tfrac{1}{2}(pKa - pKb)$$

21. BUFFER SOLUTIONS

A buffer resists pH change when small amounts of acid/base are added or on dilution.

Acidic buffer

Weak acid + its salt with a strong base. Example: CH_3COOH + CH_3COONa.

Henderson–Hasselbalch equation

$$\mathbf{pH = pKa + \log([salt]/[acid])}$$

Basic buffer

Weak base + its salt. Example: NH_4OH + NH_4Cl.

$$\mathbf{pOH = pKb + \log([salt]/[base])}$$

$$\mathbf{Then\ pH = pKw - pOH.}$$

Why does it work?

The weak acid/base and its conjugate partner consume added H^+ or OH^-, limiting the pH change.

APPLICATIONS

The chapter highlights the importance of controlled pH in body fluids and in medical, cosmetic, chemical and biochemical formulations.

22. SOLUBILITY EQUILIBRIA & Ksp

The final major section: sparingly soluble salts.

Dissolution equilibrium

$$BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$$

$$Ksp = [Ba^{2+}][SO_4^{2-}]$$

The chapter gives $Ksp(BaSO_4) = 1.1 \times 10^{-10}$ at 298 K.

Molar-solubility pattern

Salt type	Dissolution	Ksp in terms of s
AB	$A^+ + B^-$	s^2
AB_2	$A^{2+} + 2B^-$	$4s^3$
AB_3	$A^{3+} + 3B^-$	$27s^4$

Method

- Write the balanced dissolution equation.
- Let molar solubility = s.
- Write every ion concentration using stoichiometry.
- Substitute into Ksp and solve for s.

23. IONIC PRODUCT, PRECIPITATION & COMMON-ION EFFECT

Use Qsp exactly as you use Qc for equilibrium direction.

Ionic product Qsp

Calculate the product of the current ion concentrations raised to their stoichiometric powers.

Comparison	Result
Qsp < Ksp	Solution is unsaturated; no precipitation
Qsp = Ksp	Saturated equilibrium
Qsp > Ksp	Precipitation occurs

Mixing-solution workflow

- Calculate the new concentrations after mixing (account for dilution).
- Write Qsp for the possible solid.
- Compare Qsp with Ksp.
- Conclude whether precipitation occurs.

Common-ion effect on solubility

Adding a common ion to a sparingly soluble salt equilibrium generally decreases its solubility by shifting the dissolution equilibrium toward the solid.

EXAM TRAP
Never compare raw ion concentration with Ksp. Always calculate Qsp using the correct powers.

24. MASTER FORMULA SHEET — ONE-PAGE REVISION

Screenshot this page mentally before the test.

Area	Formula
Equilibrium constant	Kc = [C]^c[D]^d / [A]^a[B]^b
Gas equilibrium	Kp = Kc(RT)^Δn
Reaction quotient	Qc = [C]^c[D]^d / [A]^a[B]^b
Gibbs energy	ΔG = ΔG° + RT ln Q
At equilibrium	ΔG° = −RT ln K
Acid ionisation	Ka = [H3O+][A−] / [HA]
Base ionisation	Kb = [BH+][OH−] / [B]
pH	pH = −log[H+]
pOH	pOH = −log[OH−]
pKa / pKb	pKa = −log Ka ; pKb = −log Kb
Water	Kw = [H+][OH−]
Water p-scale	pH + pOH = pKw
Weak acid	[H+] ≈ √(Ka C)
Weak base	[OH−] ≈ √(Kb C)
Weak acid pH	pH = ½(pKa − log C)
Weak base pOH	pOH = ½(pKb − log C)
Acidic buffer	pH = pKa + log([salt]/[acid])
Basic buffer	pOH = pKb + log([salt]/[base])
Weak acid + weak base salt	pH = 7 + ½(pKa − pKb)
Solubility product	Ksp = product of ion concentrations^stoichiometric powers

25. MASTER DECISION TREE — WHAT FORMULA DO I USE?

A practical exam-navigation page.

START: What does the question give you?

Equilibrium composition? → write Kc/Kp.

Current, non-equilibrium composition? → calculate Q and compare with K.

Initial concentrations + unknown equilibrium concentrations? → ICE table + K.

ΔG°? → use ΔG° = −RT ln K.

Pressure/temperature/concentration change? → Le Chatelier.

Weak acid/base? → Ka/Kb and pH/pOH relations.

Salt solution? → identify hydrolysis type.

Buffer? → Henderson–Hasselbalch.

Sparingly soluble salt? → Ksp / Qsp.

Three universal exam moves

Move	Question to ask yourself
1. Equation	Is the chemical equation balanced?
2. Expression	Which species actually belongs in K/Q?
3. Direction / number	Am I comparing Q with K, or solving for x?

26. TOP EXAM TRAPS — AVOID THESE MARKS LOSSES

Last-minute error prevention.

- Equilibrium does not mean the reaction has stopped.
- Forward rate = reverse rate; it does not mean reactant concentration = product concentration.
- K uses equilibrium concentrations; Q can use concentrations at any instant.
- $Q < K \rightarrow$ forward; $Q > K \rightarrow$ reverse.
- Pure solids and pure liquids are omitted from heterogeneous equilibrium expressions.
- For Kp–Kc, calculate Δn using gaseous species only.
- Use Kelvin in RT calculations.
- Reverse reaction → K becomes 1/K.
- Multiplying an equation by n → K becomes K^n.
- Catalyst changes rate, not K or equilibrium composition.
- Temperature changes equilibrium and K.
- For pH, remember the logarithm sign: higher $[H^+]$ means lower pH.
- For polyprotic acids: $K_{a1} > K_{a2} > K_{a3}$.
- For precipitation: compare Qsp with Ksp after accounting for dilution.
- For buffers, use the concentration/ratio that matches the stated buffer components.
- For weak-electrolyte approximations, check whether ionisation is small enough.

EDUVISTA SCORE BOOST

In numericals, write the equation → formula → substitution → unit → final statement. This makes your working easy to verify and reduces avoidable errors.

27. NCERT-STYLE NUMERICAL PRACTICE MAP

The uploaded chapter contains a large exercise set. These are the patterns to practise repeatedly.

Pattern	Practice focus
Physical equilibrium	Vapour pressure, volume change, evaporation/condensation
Kc calculation	Given equilibrium concentrations
Kp calculation	Given gas composition / partial pressures
Kp ↔ Kc	Use Δn and temperature
Heterogeneous K	Omit pure solids/liquids
Reaction direction	Calculate Q and compare with K
Equilibrium concentration	ICE table + quadratic
Pressure shift	Compare gaseous mole counts
Temperature shift	Endothermic vs exothermic
Acid/base	Ka, Kb, pH, pOH
Salt hydrolysis	Acidic/basic/neutral tendency
Buffer	Henderson–Hasselbalch
Ksp	Solubility from Ksp
Precipitation	Qsp vs Ksp after mixing
Common ion	Change in ionisation/solubility

FROM THE UPLOADED NCERT CHAPTER

The exercise section runs through problems on vapour pressure, Kc/Kp, equilibrium concentrations, reaction direction, pressure/temperature effects, acid–base calculations, buffers, Ksp and precipitation. Use the exact NCERT exercises as your practice bank after learning the methods above.

28. 60-SECOND REVISION — BEFORE YOU ENTER THE EXAM HALL

Memorise the logic, not just the symbols.

CHEMICAL EQUILIBRIUM

- Dynamic: forward rate = reverse rate.
- K is fixed at a given temperature.
- Q < K → forward.
- Q > K → reverse.
- Q = K → equilibrium.
- Kp = Kc(RT)^Δn.
- Pure solids/liquids omitted.

LE CHATELIER

- Add reactant → products.
- Add product → reactants.
- Higher pressure → fewer gas moles.
- Temperature: favours endothermic direction when increased.
- Catalyst → faster, same equilibrium composition.

IONIC EQUILIBRIUM

- Ka ↑ → acid strength ↑.
- Kb ↑ → base strength ↑.
- pH = −log[H^+].
- Kw = [H^+][OH^-].
- pH + pOH = pKw.
- Ka1 > Ka2 > Ka3.

BUFFERS & Ksp

- Buffer resists pH change.
- pH = pKa + log(salt/acid).
- Ksp describes sparingly soluble salts.
- Qsp > Ksp → precipitation.
- Common ion generally lowers solubility.

FINAL MANTRA

BALANCE → WRITE EXPRESSION → SUBSTITUTE → COMPARE / SOLVE → CHECK UNITS → STATE CONCLUSION.

29. EDUVISTA — STUDY SMARTER, SCORE BETTER

A student-facing closing page designed for WhatsApp/status sharing and enrolment communication.

WHY EDUVISTA?

✓ Concept-first teaching
✓ Exam-oriented revision
✓ Formula + numerical strategy
✓ Clear doubt-solving approach
✓ Structured preparation for school exams

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