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KVS – PGT

Chemistry

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SIERRA
INNOVATIONS

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1 CHAPTER

Equilibrium (Physical + Chemical + Ionic)

Equilibrium in Physical & Chemical Processes

1. **Why Equilibrium is a Foundational Concept**

- Equilibrium is **not a topic**, it is a **way of thinking in chemistry**.

It explains:

- why reactions stop changing macroscopically,
- why some reactions never go to completion,
- why conditions can shift outcomes without changing reactants.

Examiner's core insight:

- Equilibrium describes balance, not inactivity.
- Most conceptual errors arise from treating equilibrium as a **static halt**, which it is not.

2. **Meaning of Equilibrium (Precise & Exam-Ready)**

- **Equilibrium is a state in which macroscopic properties of a system remain constant with time, while microscopic processes continue uninterrupted.**

Key words:

- macroscopic constancy
- microscopic activity

3. **Physical Equilibrium**

3.1 Definition

A physical equilibrium is established in physical processes involving no chemical change, such as:

- phase transitions,
- dissolution,
- vaporization.

3.2 Classic Examples (High-Yield)

(a) Liquid-Vapour Equilibrium

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

At equilibrium:

- rate of evaporation = rate of condensation,
- amount of liquid and vapour remains constant.

(b) Solid-Liquid Equilibrium

$$\text{Ice} \rightleftharpoons \text{Water}$$

- Occurs at melting point under constant pressure.

(c) Solute-Solution Equilibrium

$$\text{Sugar(s)} \rightleftharpoons \text{Sugar(aq)}$$

At saturation:

- rate of dissolution = rate of crystallization.

Exam trap:

- Saturation does not mean maximum dissolution capacity forever-it depends on conditions.

4. Characteristics of Physical Equilibrium

- Established in a closed system.
- Forward and reverse processes occur simultaneously.
- Dynamic at molecular level.
- Attained from either direction.
- Depends on temperature (and pressure where applicable).

5. Chemical Equilibrium

5.1 Definition

A chemical equilibrium is established in a reversible chemical reaction when:

- rate of forward reaction equals rate of reverse reaction,
- concentrations of reactants and products remain constant.

5.2 Representative Example

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

At equilibrium:

- formation of NH_3 continues,
- decomposition of NH_3 continues,
- net composition remains unchanged.

6. Dynamic Nature of Equilibrium (Most Tested Concept)

6.1 What "Dynamic" Means

Dynamic equilibrium means:

- reactions do not stop,
- molecular collisions continue,
- both forward and reverse reactions occur at equal rates.

Golden exam statement:

- *At equilibrium, rate ≠ zero; rate_forward = rate_reverse.*

6.2 Rate Vs Extent (Common Confusion)

- **Rate**: speed of reaction (change per unit time)
- **Extent**: how far reaction proceeds

At equilibrium:

- rate_forward = rate_reverse
- extent of reaction is **maximum under given conditions**, not complete.

7. Graphical Interpretation (Exam-Favourite)

7.1 Concentration vs Time Graph

- Initially: reactant concentration decreases, product increases.
- At equilibrium: concentrations become constant (horizontal lines).

7.2 Rate vs Time Graph

- Forward rate decreases with time.
- Reverse rate increases.
- At equilibrium: both rates intersect and remain equal.

Exam trap:

- Constant concentration does not imply zero rate.

8. Conditions Necessary for Establishing Equilibrium

- Reaction must be reversible.
- System must be closed.
- Temperature must be constant.
- Pressure constant (for gaseous equilibria).

9. Approach to Equilibrium (Important Nuance)

Equilibrium can be attained:

- starting with reactants only,
- starting with products only,
- starting with any mixture.
- Final equilibrium state is **independent of initial path**, but **depends on conditions**.

Examiner's insight:

- This independence leads to the concept of equilibrium constant.

10. Distinction: Physical vs Chemical Equilibrium

Aspect	**Physical Equilibrium**	**Chemical Equilibrium**
Nature	Physical change	Chemical change
Bonds	No bond change	Bond breaking/forming
Example	Ice-water	N_2-NH_3
Equilibrium shift	Mainly T, P	T, P, concentration

11. Common PYQ Traps (Decoded)

- "At equilibrium, reactions stop" → False
- "Equilibrium exists only in chemical reactions" → False
- "Forward and reverse reactions have same rate always" → False
- "Equilibrium can occur in open systems" → False
- "Constant concentration implies completion" → False

12. Exam-Critical Insight Box (Unique Points)

- Equilibrium is a dynamic balance, not static rest.
- Macroscopic constancy hides microscopic activity.
- Rate equality defines equilibrium, not concentration equality.
- Physical equilibrium involves phase changes only.
- Chemical equilibrium requires reversibility.
- Closed system is mandatory.
- Equilibrium state is condition-dependent.
- Path to equilibrium does not affect final state.
- Rate and extent are fundamentally different.
- Graphs are conceptual tools, not numerical data.

13. Why this Module is High-Yield for KVS PGT

- Assertion-reason on dynamic nature
- Graph-based interpretation MCQs
- Rate vs concentration confusion traps
- Physical vs chemical equilibrium comparison
- Teaching-aptitude explanation questions

14. Transition to Next Module

Now that equilibrium as a state is understood, the next question is:

- **How do we express equilibrium quantitatively?**
- This leads to the **Law of Mass Action** and **Equilibrium Constant**, which govern **extent and direction** of reactions.

Law of Mass Action & Equilibrium Constant

1. Why a Law of Equilibrium was Needed

From part-1, we know:

- equilibrium is dynamic,
- concentrations become constant,
- final state is independent of the path.

But chemistry demands:

- a **numerical measure** of equilibrium position,
- a way to **compare different reactions**,
- a criterion to **predict direction of reaction**.
- **This necessity led to the Law of Mass Action.**

2. Law of Mass Action: Statement (Exam-Ready)

- **At a given temperature, the rate of a chemical reaction is directly proportional to the product of the active masses (concentrations) of the reactants, each raised to the power of its stoichiometric coefficient.**
- From this rate concept, the **equilibrium expression** emerges.

3. Equilibrium Constant (K): Core Concept

For a general reversible reaction:

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

The equilibrium constant is:

$$\boxed{K_c = \frac{[\quad [\quad}{[\quad [\quad}}$$

Examiner's golden rule:

- Only species whose concentrations change appear in the expression.

4. Why Pure Solids & Liquids are Excluded (Very High-Yield)

- Concentration of pure solids and liquids remains constant.
- Their effect is already included in the value of K.

Exam trap:

- Students wrongly include solids/liquids in K expressions.

Example:

$$\mathrm{CaCO_3}(s) \rightleftharpoons \mathrm{CaO}(s) + \mathrm{CO_2}(g)$$
$$K_c = [\mathrm{CO_2}]$$

5. Types of Equilibria Based on Phase

5.1 Homogeneous Equilibrium

- All reactants and products are in the **same phase**.

Example:

$$\mathrm{H_2}(g) + \mathrm{I_2}(g) \rightleftharpoons 2\mathrm{HI}(g)$$

5.2 Heterogeneous Equilibrium

- Reactants and/or products are in **different phases**.

Example:

$$\mathrm{NH_4HS}(s) \rightleftharpoons \mathrm{NH_3}(g) + \mathrm{H_2S}(g)$$

- Only gaseous species appear in K.

6. Characteristics of Equilibrium Constant (K)

6.1 Depends Only on Temperature

Independent of:

- initial concentrations,
- pressure (at constant T),
- catalyst.

Examiner's insight:

- Catalyst speeds up attainment of equilibrium but does **not shift equilibrium**.

6.2 Magnitude of K & Extent of Reaction

Value of K	**Interpretation**
$K \gg 1$	Products favoured
$K \approx 1$	Comparable amounts
$K \ll 1$	Reactants favoured

Exam trap:

- Large K does **not** mean fast reaction.

7. Equilibrium Constant for Reversed & Modified Reactions

7.1 Reversed Reaction

$$K_{\text{reverse}} = \frac{1}{K_{\text{forward}}}$$

7.2 Multiplied Equation

If reaction is multiplied by n:

$$K_{\text{new}} = K^n$$

- **Very high-yield PYQ zone**

8. Reaction Quotient (Q): Direction Predictor

8.1 Definition

- **Reaction quotient (Q)** is the ratio of product concentrations to reactant concentrations **at any moment**, not necessarily at equilibrium.
- Expression of Q is **identical** to K, but concentrations are **not equilibrium values**.

8.2 Comparison of Q With K (Classic Exam Area)

Condition	**Prediction**
Q < K	Reaction proceeds forward
Q > K	Reaction proceeds backward
Q = K	System at equilibrium

Golden exam statement:

- *Q decides direction, K decides extent.*

9. KC vs KP (Qualitative, KVS Level)

- K_c: expressed in terms of molar concentrations
- K_p: expressed in terms of partial pressures (gases)

Relation:

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

Where:

$$\Delta n = (\text{moles of gaseous products}) - (\text{moles of gaseous reactants})$$

Exam nuance:

- If Δn = 0 → $K_p = K_c$

10. Units of Equilibrium Constant (Tricky Area)

- K has no units in thermodynamic sense.
- Apparent units may arise from concentration terms but are ignored.

Examiner's phrase:

- Equilibrium constant is dimensionless.

11. Common PYQ Traps (Decoded)

- "K changes with concentration" → False
- "Catalyst increases value of K" → False
- "Large K implies fast reaction" → False
- "Solids must be included in K" → False
- "Q = K always" → False

12. Exam-Critical Insight Box (Unique Points)

- Law of mass action gives equilibrium expression.
- Equilibrium constant depends only on temperature.
- Pure solids and liquids are excluded.
- K indicates extent, not speed.
- Reaction quotient predicts direction.
- Q = K defines equilibrium state.
- Reversing reaction inverts K.
- Multiplying equation raises K to power.
- Δn controls Kp-Kc relation.
- Catalyst does not alter equilibrium constant.

13. Why this Module is High-Yield for KVS PGT

- Expression-writing MCQs
- Assertion-reason on K vs Q
- Solid/liquid exclusion traps
- Direction prediction questions
- Teaching-aptitude explanations

14. Transition to Next Module

Now that equilibrium can be expressed quantitatively, the next step is:

- **How does equilibrium respond when conditions are disturbed?**
- This is answered by **Le Chatelier's Principle**, the most tested applied concept of equilibrium.

Factors Affecting Equilibrium

1. **Why Le Chatelier's Principle is Central to Equilibrium**

   After defining equilibrium and quantifying it via K, chemistry asks:

   - **What happens when an equilibrium system is disturbed?**
   - Industrial chemistry, biological systems, and laboratory control **all rely on this answer**.
   - Le Chatelier's Principle provides a **predictive, qualitative tool**-not a numerical law.

2. **Le Chatelier's Principle: Statement (Exam-Ready)**

   - **When a system at equilibrium is subjected to a change in concentration, pressure, or temperature, the system shifts its equilibrium position in such a way as to oppose the imposed change.**

   Key exam phrase:

   - *The system counteracts the stress.*

3. **Important Pre-Clarifications (Exam-Saving Points)**
   - The principle predicts direction of shift, not rate.
   - It applies only to systems already at equilibrium.
   - “Shift” refers to change in equilibrium position, not movement of substances.

4. **Effect of Change in Concentration**

4.1 GENERAL LOGIC

For a reaction:

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

- Increasing concentration of a reactant → shifts equilibrium forward
- Increasing concentration of a product → shifts equilibrium backward
- Removing a species → equilibrium shifts to replace it

Reason:

- System attempts to reduce the imposed concentration change.

4.2 Important Exam Nuance

- Adding pure solids or liquids → no effect on equilibrium position.
- Only gaseous or dissolved species matter.

Exam trap:

- Adding $CaCO_3(s)$ does not shift $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

5. **Effect of Change in Pressure (Gaseous Equilibria)**

5.1 Pressure Increase

- Equilibrium shifts towards the side with fewer moles of gas.
- If gaseous moles are equal → no effect.

Example:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

- Left: 4 moles gas
- Right: 2 moles gas
- Increased pressure favours **NH_3 formation**

5.2 Pressure Decrease

- Equilibrium shifts towards side with **more moles of gas**.

5.3 Non-Gaseous Systems

- Pressure change has **no effect** if gases are absent.

Examiner’s favourite:

- Pressure affects equilibrium only when $\Delta n(g) \neq 0$.

6. **Effect of Change in Volume**
   - Volume change is the inverse of pressure change.
   - Decrease in volume → increase in pressure
   - Increase in volume → decrease in pressure

Thus:

- ↓ Volume → shift towards fewer gaseous moles
- ↑ Volume → shift towards more gaseous moles

7. Effect of Change in Temperature (Most Critical Area)

7.1 Temperature is Unique

Unlike concentration and pressure:

- Temperature change **alters the value of equilibrium constant (K)**.

Golden exam rule:

- Only *temperature changes K.*

7.2 Treat Heat as a Reactant or Product

- Exothermic reaction: Heat is a product
- Endothermic reaction: Heat is a reactant

7.3 Effect of Temperature Change

Reaction Type	**Increase in T**	**Decrease in T**
Exothermic	Shifts backward	Shifts forward
Endothermic	Shifts forward	Shifts backward

Exam trap:

- Students often reverse this logic.

8. Effect of Catalyst (Very High-Yield Misconception)

- Catalyst **does not shift equilibrium position**.
- It **does not change K**.
- It increases rate of both forward and reverse reactions equally.

Correct statement:

- Catalyst helps the system reach equilibrium faster, not change equilibrium.

9. Limitations Of Le Chatelier's Principle

Le Chatelier's principle:

- is qualitative, not quantitative,
- does not explain why shifts occur at molecular level,
- cannot predict exact composition at equilibrium.

Exam insight:

- It is a **predictive rule**, not a mathematical law.

10. Combined Effects (Advanced Logic)

When multiple factors change simultaneously:

- final shift depends on dominant effect.
- temperature effect usually overrides others due to K change.

Examiner's trick:

- Multiple disturbances in a single question.

11. Common PYQ Traps (Decoded)

- "Increasing pressure always favours products" → False
- "Catalyst shifts equilibrium forward" → False
- "Temperature change never affects K" → False
- "Adding solid reactant shifts equilibrium" → False
- "Volume decrease always increases yield" → False

12. Exam-Critical Insight Box (Unique Points)

- Le Chatelier predicts direction, not extent.
- Only equilibrated systems respond meaningfully.
- Concentration change shifts equilibrium to counteract change.
- Solids and liquids do not affect equilibrium position.
- Pressure affects only gaseous equilibria.
- Shift favours fewer gas moles under high pressure.
- Temperature change alters equilibrium constant.
- Heat behaves like reactant or product.
- Catalyst does not change equilibrium position.
- Le Chatelier's principle is qualitative.

13. Why This Module Is High-Yield For KVS PGT

- Assertion-reason on temperature vs K
- Pressure-volume trap MCQs
- Catalyst misconception questions
- Direction-of-shift prediction problems
- Teaching-aptitude explanation items

14. Transition to Next Module

- So far, equilibrium was treated for **molecular reactions**.

Now chemistry moves to ionic equilibria, where:

- reactions are incomplete,
- equilibrium lies far to one side,
- degree of ionization matters.

This begins with:

- **Strong vs Weak Electrolytes & Ionization Concepts**

Ionic Equilibrium: Acids, Bases & Electrolytes

1. Why Ionic Equilibrium Is Conceptually Different

In molecular equilibrium:

- reactants and products are **distinct chemical species**.

In ionic equilibrium:

- species exist **simultaneously in ionized and unionized forms**.
- equilibrium lies **far to one side**, yet never reaches completion.

Examiner's insight:

- *Ionic equilibrium is equilibrium with very unequal sides.*

2. Electrolytes: Basic Classification

2.1 Electrolytes

- Substances that conduct electricity in aqueous solution or molten state due to formation of ions.

They are classified as:

- Strong electrolytes
- Weak electrolytes

2.2 Non-Electrolytes (Clarity Point)

- Do not ionize in solution.
- Do not conduct electricity.
- Examples: glucose, urea.

Exam trap:

- Solubility does not imply ionization.

3. Strong Electrolytes

3.1 Definition

- **Strong electrolytes** are substances that are **almost completely ionized** in aqueous solution.

Examples:

- Strong acids: HCl, HNO_3, H_2SO_4
- Strong bases: NaOH, KOH
- Most soluble salts

3.2 Ionization Behaviour

- Ionization is nearly 100%.
- Equilibrium lies extremely to the **right.**
- Law of mass action is **not useful** in simple form.

Examiner's phrase:

- Strong electrolytes are treated as fully dissociated at this level.

4. Weak Electrolytes

4.1 Definition

- **Weak electrolytes** are substances that are **partially ionized** in aqueous solution.

Examples:

- Weak acids: CH_3COOH, HF
- Weak bases: NH_4OH

4.2 Ionic Equilibrium in Weak Electrolytes

Example:

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

- Forward and reverse reactions occur.
- Equilibrium constant governs extent of ionization.

Exam insight:

- Weak electrolytes are ideal systems for applying equilibrium laws.

5. Degree Of Ionization (A): Core Parameter

5.1 Definition

- **Degree of ionization (α)** is the fraction of total electrolyte molecules that ionize in solution.

$$\alpha = \frac{\text{number of molecules ionized}}{\text{total number of molecules dissolved}}$$

5.2 Characteristics of α

$$0 < \alpha < 1$$

- For strong electrolytes: $\alpha \approx 1$
- For weak electrolytes: $\alpha \ll 1$

6. Factors Affecting Degree of Ionization

6.1 Nature of Electrolyte

- Strong acid/base → higher α
- Weak acid/base → lower α

6.2 Concentration (Very High-Yield)

- Dilution increases α for weak electrolytes.
- Strong electrolytes show negligible change.

Reason:

- Lower ion concentration reduces reverse reaction probability.

6.3 Temperature

- Increase in temperature generally increases ionization.
- Ionization is usually endothermic.

7. Ostwald's Dilution Law (Conceptual Level)

- For weak electrolytes, **degree of ionization increases on dilution**, and is related to equilibrium constant.

For a weak electrolyte:

$$K = \frac{C\alpha^2}{1-\alpha}$$

At high dilution:

$$\alpha \approx \sqrt{\frac{K}{C}}$$

Examiner's caution:

- Ostwald's law applies **only to weak electrolytes**.

8. Acids & Bases: Ionization View

8.1 ACIDS

- Substances that produce H^+ ions in aqueous solution.
- Strong acids → almost complete ionization
- Weak acids → partial ionization

8.2 Bases

- Substances that produce OH^- ions in aqueous solution.
- Strong bases → complete ionization
- Weak bases → partial ionization

Exam trap:

- Strength depends on **extent of ionization**, not concentration.

9. Acid Strength vs Acid Concentration (Very Important)

Aspect	**Acid Strength**	**Acid Concentration**
Meaning	Degree of ionization	Amount dissolved
Depends on	Nature of acid	Solution preparation
Example	HCl strong	Dilute HCl still strong
Exam confusion	Very common	Avoidable

Golden exam statement:

- *A dilute strong acid is still strong.*

10. Relative Strength of Acids (Conceptual)

Strength depends on:

- bond strength,
- stability of conjugate base,
- polarity of bond.

Example:

$$HCl > HF$$

- Despite HF having higher electronegativity.
- Detailed explanation follows in next part

11. Common PYQ Traps (Decoded)

- "Weak acid means low concentration" → False
- "Strong electrolytes follow Ostwald's law" → False
- "Degree of ionization is constant" → False
- "Ionization means dissolution" → False
- "All electrolytes ionize completely" → False

12. Exam-Critical Insight Box (Unique Points)

- Ionic equilibrium deals with partial ionization.
- Strong electrolytes are nearly fully ionized.
- Weak electrolytes establish true ionic equilibrium.
- Degree of ionization quantifies extent of ionization.
- Dilution increases ionization of weak electrolytes.
- Ostwald's law applies only to weak electrolytes.
- Acid strength depends on ionization, not amount.
- Concentration and strength are independent concepts.
- Temperature generally increases ionization.
- Ionic equilibrium is highly asymmetric equilibrium.

13. Why this Module is High-Yield for KVS PGT

- Assertion-reason on strong vs weak electrolytes
- Degree of ionization vs dilution MCQs
- Acid strength vs concentration traps
- Ostwald's law conceptual questions
- Teaching-aptitude explanations

14. Transition to Next Module

Now that ionization and strength concepts are clear, the next logical questions are:

- How do acids ionize stepwise?
- Why do polybasic acids behave differently?
- How is acidity quantified numerically?
- This leads to:

Ionization of Polybasic Acids & Concept of Ph

1. Why Polybasic Acids Require Separate Treatment

Unlike monoprotic acids, polybasic (polyprotic) acids:

- contain more than one ionizable hydrogen atom,
- ionize stepwise, not simultaneously,
- establish multiple equilibria, each with its own equilibrium constant.

Examiner's insight:

- *Failure to recognize stepwise ionization is the most common error in this area.*

2. Polybasic Acids: Definition & Types

2.1 Definition

- **Polybasic acids** are acids capable of donating **more than one proton (H^+)** per molecule in aqueous solution.

2.2 Classification

Acid Type	Example	Protons Donated
Diprotic	H_2SO_4, H_2CO_3	2
Triprotic	H_3PO_4	3

Exam trap:

- Number of ionizable hydrogens ≠ total hydrogens in molecule.

3. Stepwise Ionization: Fundamental Concept

3.1 Example: Diprotic Acid (H_2A)

$$\mathrm{H_2A} \rightleftharpoons \mathrm{H^+} + \mathrm{HA^-} (K_{a1})$$
$$\mathrm{HA^-} \rightleftharpoons \mathrm{H^+} + \mathrm{A^{2-}} (K_{a2})$$

Each step:

- has its own equilibrium constant,
- proceeds independently,
- has different extent.

4. Successive Ionization Constants (Very High-Yield)

4.1 Order of Magnitude

$$K_{a1} > K_{a2} > K_{a3}$$

4.2 Why this Order Always Holds

After first ionization:

- molecule becomes negatively charged,

removal of next proton faces:

- electrostatic repulsion,
- reduced tendency to release H^+.

Golden exam statement:

- *Each successive ionization is increasingly difficult.*

5. Exam-Critical Consequence

- First ionization contributes maximum H^+.
- Higher ionizations contribute negligibly to total acidity.

Implication:

- For pH calculations of polybasic acids, **Ka_1 dominates**.

6. Special Case: Sulphuric Acid (H_2So_4)

- First ionization: strong (almost complete)
- Second ionization: weak

$$\mathrm{HSO_4^-} \rightleftharpoons \mathrm{H^+} + \mathrm{SO_4^{2-}}$$

Exam trap:

- Treating both ionizations of H_2SO_4 as equally strong is incorrect.

7. Concept of Ph: Need & Origin

Hydrogen ion concentration in solutions ranges from:

- ~1 mol L^{-1} (strong acids),
- to ~10^{-14} mol L^{-1} (strong bases).
- Handling such values is inconvenient.
- Solution: Use a logarithmic scale → pH.

8. Definition of Ph (Exam-Ready)

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

Where:

- $[H^+]$ is molar concentration of hydrogen ions.

Examiner's insight:

- pH is a **measure of acidity**, not acid strength directly.

9. Logarithmic Nature of Ph (Most Misunderstood)

- pH scale is logarithmic, not linear.
- Change of 1 pH unit = 10-fold change in $[H^+]$.

Example:

- pH 3 is **10 times more acidic** than pH 4.

Exam trap:

- Students treat pH as linear scale.

10. pH Scale: Interpretation

pH Value	Nature
< 7	Acidic
= 7	Neutral
> 7	Basic

- At 25°C:
- Pure water: $[H^+] = 10^{-7}$ → pH = 7

Exam nuance:

- Neutral pH changes with temperature.

11. Relation Between Ph & Poh (Qualitative)

$$\mathrm{pH} + \mathrm{pOH} = 14(25^\circ\mathrm{C})$$

Where:

- $\mathrm{pOH} = -\log[OH^-]$

Examiner's caution:

- Value 14 is temperature-dependent.

12. pH Of Weak Acids (Conceptual Logic)

- Weak acid → partial ionization

$[H^+]$ depends on:

- acid strength (Ka),
- concentration.

Key idea:

- Stronger weak acid (higher Ka) → lower pH at same concentration.

13. Common PYQ Traps (Decoded)

- "All ionizable hydrogens ionize equally" → False
- "Ka_2 can exceed Ka_1" → False
- "pH scale is linear" → False
- "pH measures acid concentration" → False
- "Neutral pH is always 7" → False

14. Exam-Critical Insight Box (Unique Points)

- Polybasic acids ionize stepwise.
- Each ionization has its own equilibrium constant.
- Ka_1 is always greater than Ka_2, Ka_3.
- First ionization dominates acidity.

- Electrostatic repulsion reduces successive ionization.
- pH is logarithmic measure of [H^+].
- One pH unit change equals tenfold [H^+] change.
- pH reflects acidity, not acid strength directly.
- Neutral pH depends on temperature.
- Higher Ka means lower pH at same concentration.

15. Why this Module is High-Yield for KVS PGT

- Assertion-reason on Ka order
- Polybasic acid conceptual MCQs
- pH comparison questions
- Logarithmic scale traps
- Teaching-aptitude explanation questions

16. Transition to Next Module

Once pH is understood, the next natural question is:

- **What happens when salts formed from acids and bases dissolve in water?**

This leads to:

- Hydrolysis of salts,
- prediction of solution pH,
- strong-weak parent logic.

Hydrolysis of Salts (Elementary Idea, But Exam-Tricky)

1. **Why Salt Hydrolysis Confuses Students**
    - Salts are often assumed to be neutral because they are “products of neutralization”.
    - This assumption is wrong in most cases.

    Examiner’s core insight:

    - The pH of a salt solution depends on the strengths of its parent acid and base, not on the salt itself.
    - Salt hydrolysis is therefore a logical extension of ionic equilibrium, not a new topic.
2. **What is Hydrolysis of Salts**
    - Hydrolysis of a salt is the reaction of its ions with water to produce either H^+ or OH^- ions, thereby changing the pH of the solution.

    Key point:

    - Only ions of weak acids or weak bases undergo hydrolysis.
    - Ions of strong acids and strong bases do not hydrolyse.
3. **Classification of Salts (Most Important Framework)**
    - Salt behaviour is decided by **parent acid + parent base**.