



KVS – PGT

Chemistry

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Volume - 2



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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

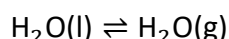
##### 3.1 Definition

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

- phase transitions,
- dissolution,
- vaporization.

##### 3.2 Classic Examples (High-Yield)

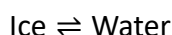
###### (a) Liquid-Vapour Equilibrium



##### At equilibrium:

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

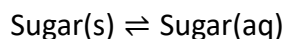
###### (b) Solid-Liquid Equilibrium



- Occurs at melting point under constant pressure.

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### (c) Solute-Solution Equilibrium



**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).

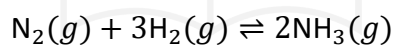
### 5. Chemical Equilibrium

#### 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



**At equilibrium:**

- formation of  $\text{NH}_3$  continues,
- decomposition of  $\text{NH}_3$  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  $\neq$  zero; rate<sub>forward</sub> = rate<sub>reverse</sub>.*

#### 6.2 Rate Vs Extent (Common Confusion)

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

**At equilibrium:**

- rate<sub>forward</sub> = rate<sub>reverse</sub>
- 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.

<b>Law of Mass Action &amp; Equilibrium Constant</b>
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### 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:

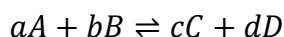
- a **numerical measure** of equilibrium position,
- a way to **compare different reactions**,
- a criterion to **predict direction of reaction**.
- **This necessity led to the Law of Mass Action.**

### 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:



The equilibrium constant is:

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Examiner's golden rule:

- Only species whose concentrations change appear in the expression.

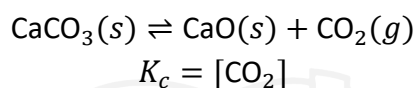
### 4. Why Pure Solids & Liquids are Excluded (Very High-Yield)

- 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:

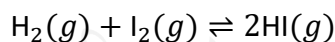


### 5. Types of Equilibria Based on Phase

#### 5.1 Homogeneous Equilibrium

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

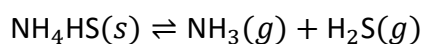
Example:



#### 5.2 Heterogeneous Equilibrium

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

Example:



- 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  $\Delta n = 0 \rightarrow K_p = K_c$

**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.

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## 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.
- $\Delta n$  controls  $K_p$ - $K_c$  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-apptitude 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.*

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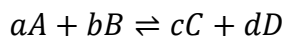
### 3. Important Pre-Clarifications (Exam-Saving Points)

- 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:



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

Reason:

- System attempts to reduce the imposed concentration change.

#### 4.2 Important Exam Nuance

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

Exam trap:

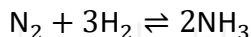
- Adding  $\text{CaCO}_3(\text{s})$  does not shift  $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

### 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  $\rightarrow$  no effect.

Example:



- Left: 4 moles gas
- Right: 2 moles gas
- Increased pressure favours  **$\text{NH}_3$  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  $\Delta n(\text{g}) \neq 0$ .

### 6. Effect of Change in Volume

- Volume change is the inverse of pressure change.
- Decrease in volume  $\rightarrow$  increase in pressure
- Increase in volume  $\rightarrow$  decrease in pressure

Thus:

- $\downarrow$  Volume  $\rightarrow$  shift towards fewer gaseous moles
- $\uparrow$  Volume  $\rightarrow$  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**

<b>Ionic Equilibrium: Acids, Bases &amp; Electrolytes</b>
---

### 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<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>
- 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

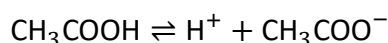
- **Weak electrolytes** are substances that are **partially ionized** in aqueous solution.

**Examples:**

- Weak acids: CH<sub>3</sub>COOH, HF
- Weak bases: NH<sub>4</sub>OH

### 4.2 Ionic Equilibrium in Weak Electrolytes

**Example:**



- 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 ( $\alpha$ )** 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 $\alpha$

$$0 < \alpha < 1$$

- 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  $\rightarrow$  higher  $\alpha$
- Weak acid/base  $\rightarrow$  lower  $\alpha$

### 6.2 Concentration (Very High-Yield)

- Dilution increases  $\alpha$  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  $\rightarrow$  almost complete ionization
- Weak acids  $\rightarrow$  partial ionization

## 8.2 Bases

- Substances that produce  $\text{OH}^-$  ions in aqueous solution.
- Strong bases  $\rightarrow$  complete ionization
- Weak bases  $\rightarrow$  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:

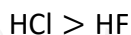
- *A dilute strong acid is still strong.*

## 10. Relative Strength of Acids (Conceptual)

### Strength depends on:

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

### Example:



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

## 11. Common PYQ Traps (Decoded)

- "Weak acid means low concentration"  $\rightarrow$  False
- "Strong electrolytes follow Ostwald's law"  $\rightarrow$  False
- "Degree of ionization is constant"  $\rightarrow$  False
- "Ionization means dissolution"  $\rightarrow$  False
- "All electrolytes ionize completely"  $\rightarrow$  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<sup>+</sup>)** per molecule in aqueous solution.

#### 2.2 Classification

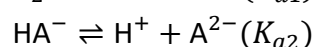
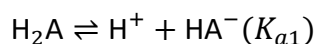
Acid Type	Example	Protons Donated
Diprotic	H <sub>2</sub> SO <sub>4</sub> , H <sub>2</sub> CO <sub>3</sub>	2
Triprotic	H <sub>3</sub> PO <sub>4</sub>	3

**Exam trap:**

- Number of ionizable hydrogens ≠ total hydrogens in molecule.

### 3. Stepwise Ionization: Fundamental Concept

#### 3.1 Example: Diprotic Acid (H<sub>2</sub>A)



**Each step:**

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

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## 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

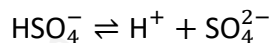
- First ionization contributes maximum  $H^+$ .
- Higher ionizations contribute negligibly to total acidity.

#### Implication:

- For pH calculations of polybasic acids,  **$K_{a1}$  dominates.**

## 6. Special Case: Sulphuric Acid ( $H_2SO_4$ )

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



#### 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:

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

## 8. Definition of Ph (Exam-Ready)

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

#### 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:  $[H^+] = 10^{-7} \rightarrow pH = 7$

**Exam nuance:**

- Neutral pH changes with temperature.

**11. Relation Between Ph & Poh (Qualitative)**

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

**Where:**

- $pOH = -\log[OH^-]$

**Examiner's caution:**

- Value 14 is temperature-dependent.

**12. pH Of Weak Acids (Conceptual Logic)**

- Weak acid  $\rightarrow$  partial ionization

 **$[H^+]$  depends on:**

- acid strength ( $K_a$ ),
- concentration.

**Key idea:**

- Stronger weak acid (higher  $K_a$ )  $\rightarrow$  lower pH at same concentration.

**13. Common PYQ Traps (Decoded)**

- "All ionizable hydrogens ionize equally"  $\rightarrow$  False
- " $K_{a2}$  can exceed  $K_{a1}$ "  $\rightarrow$  False
- "pH scale is linear"  $\rightarrow$  False
- "pH measures acid concentration"  $\rightarrow$  False
- "Neutral pH is always 7"  $\rightarrow$  False

**14. Exam-Critical Insight Box (Unique Points)**

- Polybasic acids ionize stepwise.
- Each ionization has its own equilibrium constant.
- $K_{a1}$  is always greater than  $K_{a2}$ ,  $K_{a3}$ .
- 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  $K_a$  means lower pH at same concentration.

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

- Assertion-reason on  $K_a$  order
- Polybasic acid conceptual MCQs
- pH comparison questions
- Logarithmic scale traps
- Teaching-apptitude 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**.