

PARISHRAM



2026

CHEMISTRY

Lecture 01

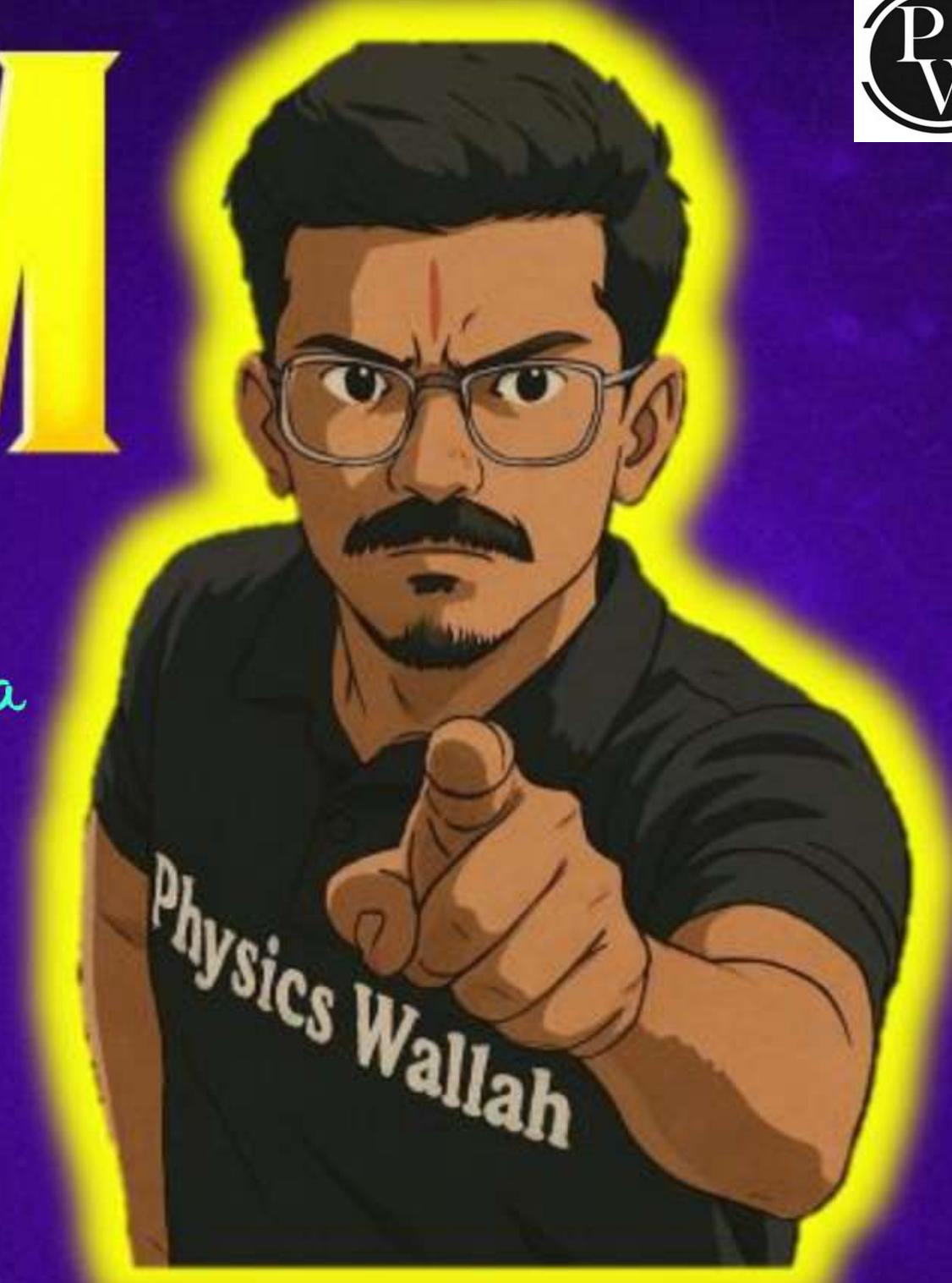
ELECTROCHEMISTRY ✓

Introduction to Electrochemistry

Bharat Mata
Ki Jai

Physics Wallah

BY – PRIYA-PUTRA-SUNIL



TOPICS TO BE COVERED

- (i) Essential Basics for this Chapter ✓
- (ii) Introduction to Electrochemistry ✓



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BYSUNILBHAIIYA





ESSENTIAL CHEMISTRY BASICS FOR THIS CHAPTER

Aye Bhaiya ❤

A chemical reaction in which **reduction** and **oxidation** take place simultaneously.

↳ (पक्ष साथ हीना)

Redox

(i) On Basis of Gain/Loss of Hydrogen and Oxygen or Any Electropositive and Electronegative Element

(iii) On Basis of Oxidation Number

(ii) On Basis of Loss/gain of Electron(s)

ADDITION/REMOVAL OF HYDROGEN OR OXYGEN

Oxidation ✓

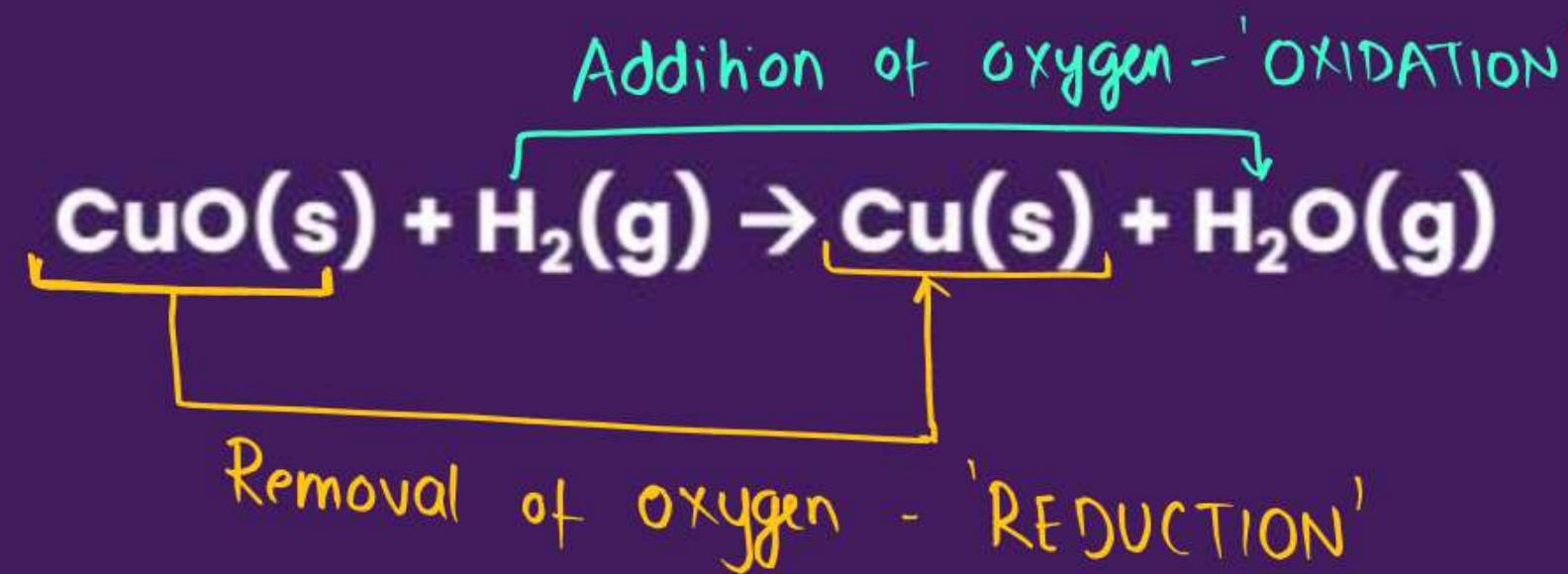
- Addition of oxygen or any electronegative element
- Removal of hydrogen or any electropositive element
- Both

Reduction ✓

- Addition of hydrogen or any electropositive element
- Removal of oxygen or any electronegative element
- Both

→ Electronegative Element: Tendency to attract or gain electron(s)
→ Electropositive Element: Tendency to lose electron(s)

REDOX REACTION - EXAMPLE



OXIDISING AND REDUCING AGENT



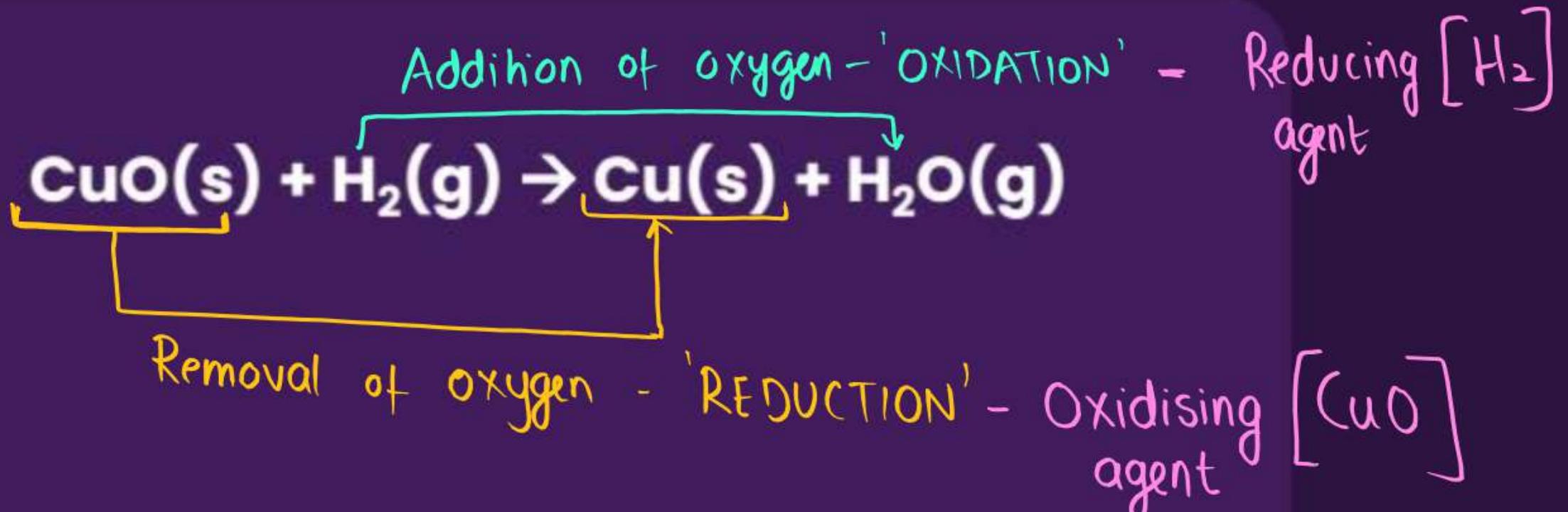
Oxidising Agent/Oxidant

- Oxidises others and itself gets reduced
- The one which shows reduction is the oxidizing agent.
oxidant.

Reducing Agent/Reductant

- Reduces others and itself gets oxidised
- The one which shows oxidation is the reducing agent.
or reductant

REDOX REACTION - EXAMPLE



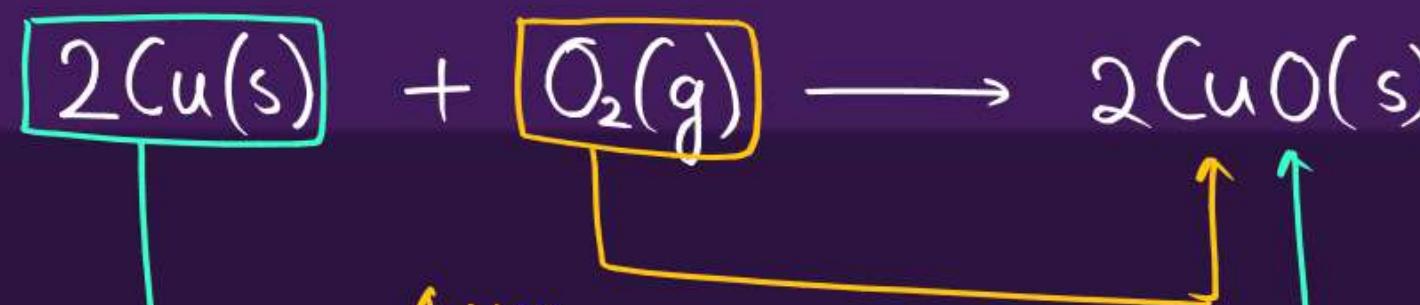
GIVE A THOUGHT



The below reaction is an example of a redox reaction.



- A. Yes
- B. No



Addition of electropositive element - REDUCTION - Oxidising agent [O₂]

Addition of oxygen or addition of E.N. element - OXIDATION

Reducing agent [Cu]

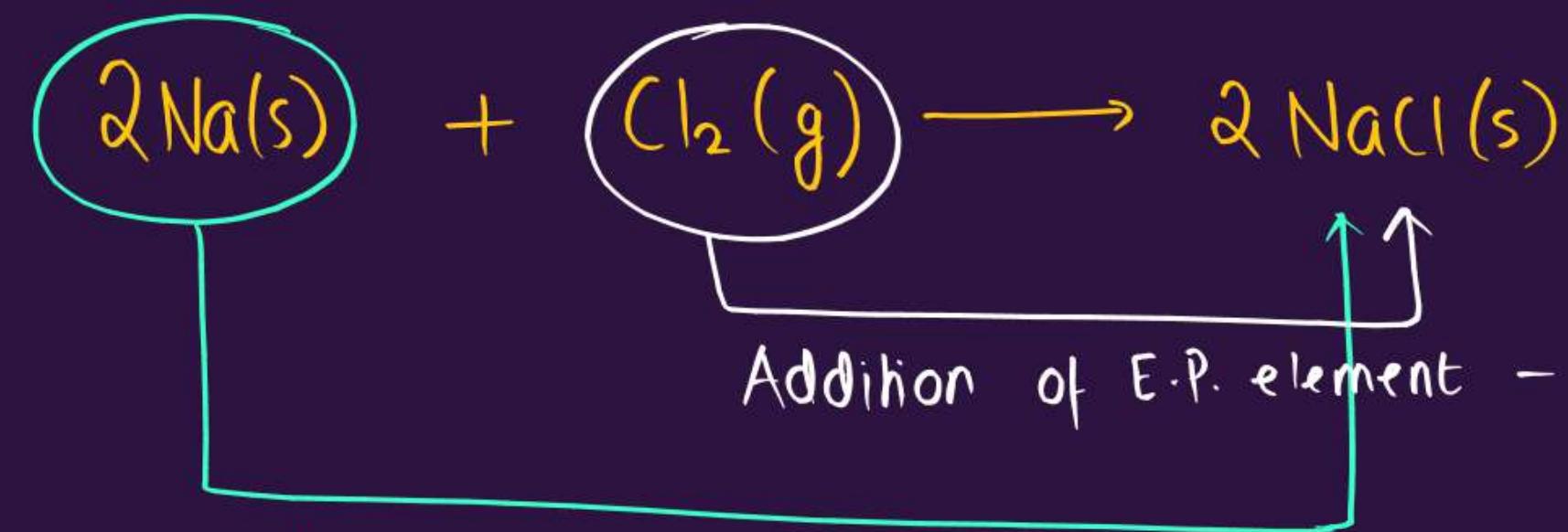
**SAMAJ AAYA TOH
LIKH DO.**

AYE BHAIYA ✓



ELECTRON TRANSFER CONCEPT – REDUCTION AND OXIDATION

- ✓ Oxidation is a process in which an atom or an ion loses one or more electrons. This is why it is also called **de-electronation**.
- **OIL:** → Oxidation is loss of electron(s).
- ✓ Reduction is a process in which an atom or an ion gains one or more electrons. This is why it is also called **electronation**.
- **RIG:** → Reduction is gain of electron(s).



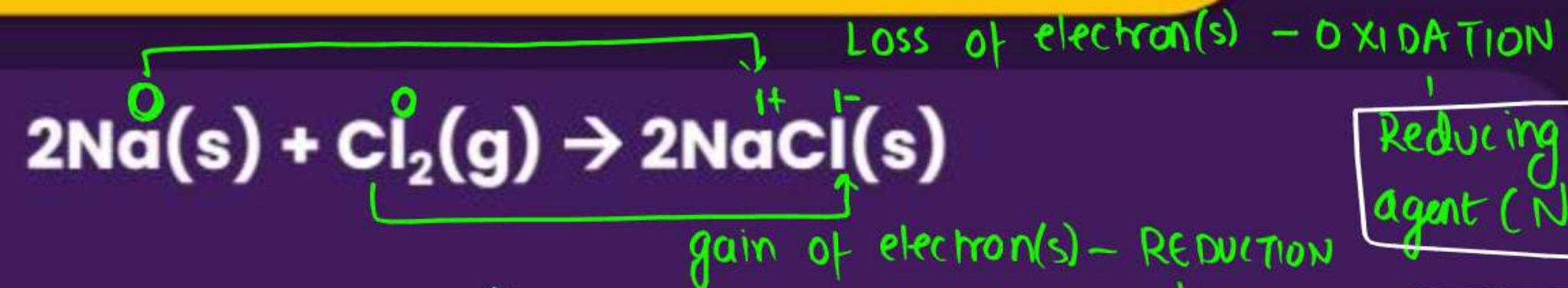
Addition of E.P. element - REDUCTION - Oxidising agent (Cl_2)

Addition of E.N. element - OXIDATION - REDUCING AGENT (Na)

$\text{Na} \rightarrow \text{metal} \rightarrow e^- \text{ donor}$

$\text{Cl} \rightarrow \text{non-metal} \rightarrow e^- \text{ acceptor}$

REDOX REACTION - EXAMPLE



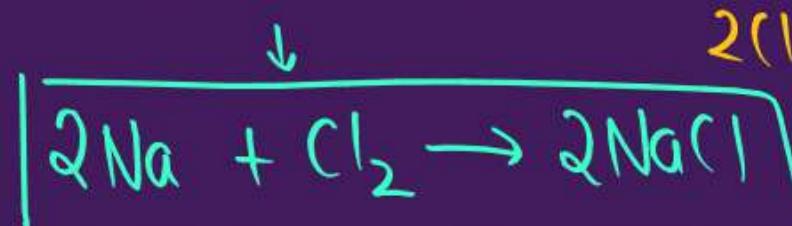
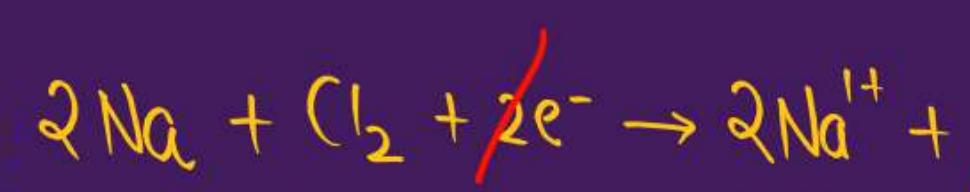
Oxidation half rxn



Reduction half rxn $\rightarrow \text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^{\text{I}-}$

from ① & ②

Complete redox rxn



Reducing agent (Na)

Oxidising agent (Cl_2)

• There is never an imbalance in the number of electron(s) lost/gained during a chemical reaction. The electron(s) lost by one species will be equal to the electron(s) gained by another species.

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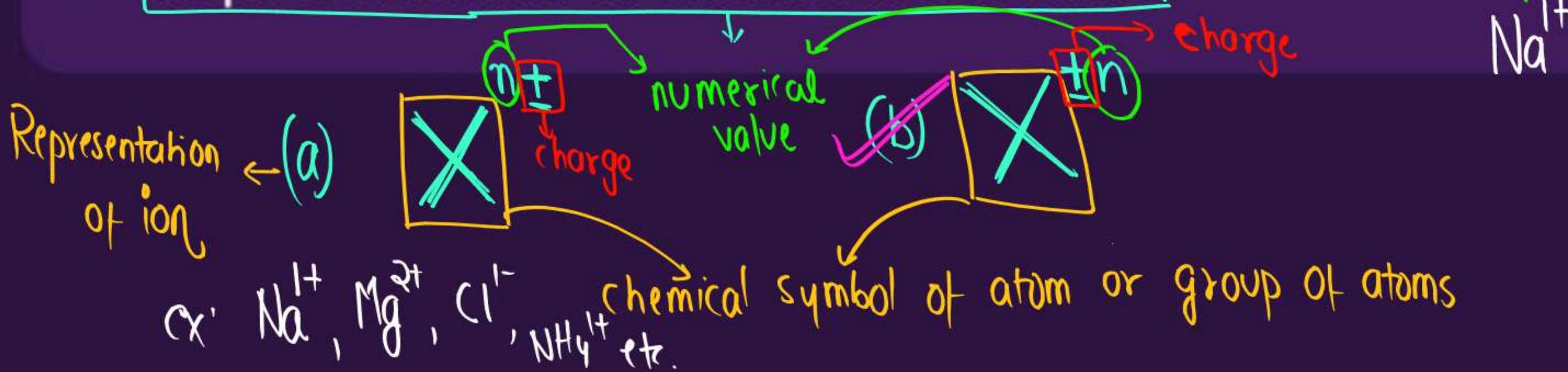


OXIDATION NUMBER/OXIDATION STATE

- [electron(s) lost | gained during formation of that species]
- It is a real or imaginary charge developed on an atom when it goes from its most stable elemental free state to the combined state by complete loss/gain of electrons.

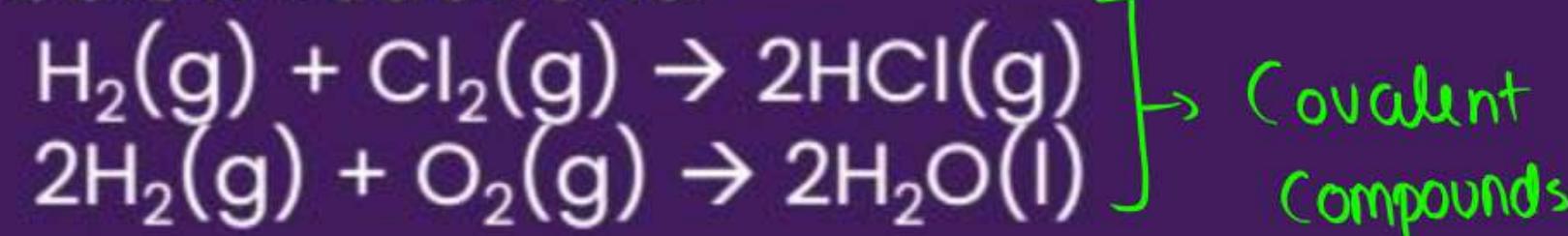
- Oxidation number can be **+ve, -ve** and **0**.

- Representation of oxidation number:



OXIDATION NUMBER/OXIDATION STATE

- Let's observe the below reactions:



- (In order to keep track of electron shifts in above chemical reactions involving formation of covalent compounds, a more practical method of using **oxidation number** has been developed.)
- In this method, (it is always assumed that there is a complete transfer of electron from a less electronegative atom to a more electronegative atom.)
- Oxidation number is also called **oxidation state**.

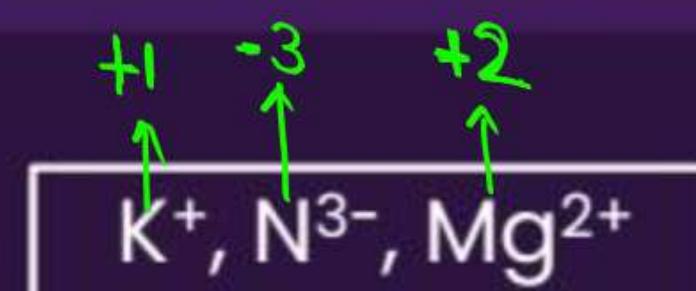
OXIDATION NUMBER/OXIDATION STATE

- Elements: 0
- Group 1A: Always +1
- Group 2A: Always +2
- Halogens: Usually -1, positive with oxygen and fluorine
- Monoatomic ion: Ion charge
- Hydrogen: +1 with non-metals and -1 with metals
- Oxygen: Usually -2. -1 in peroxide and -1/2 in superoxide
- Fluorine: Always -1.
- Sum of oxidation number of a neutral compound: 0
- Sum of oxidation number for a polyatomic ion: ion charge



OXIDATION NUMBER/OXIDATION STATE

- Elements: 0
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- Halogens: Usually -1, positive with oxygen and fluorine
- Sum of oxidation number of a neutral compound: 0
- Sum of oxidation number for a polyatomic ion: ion charge
- **Monoatomic ion: Ion charge**
 - one atom
 - electrical charge
- Hydrogen: +1 with non-metals and -1 with metals
- Oxygen: Usually -2. -1 in peroxide and -1/2 in superoxide
- Fluorine: Always -1.

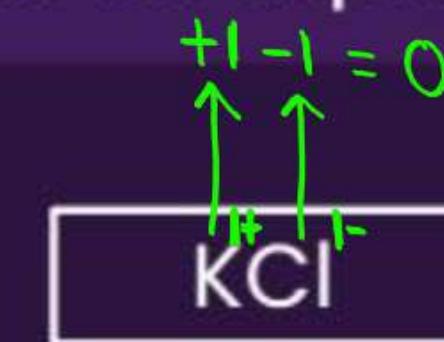


OXIDATION NUMBER/OXIDATION STATE

- Elements: 0
 - Group 1A: Always +1
 - Group 2A: Always +2
 - Halogens: Usually -1,
positive with oxygen
and fluorine
- F, Cl,
Br, I

- Sum of oxidation number of a neutral compound: 0
- Sum of oxidation number for a polyatomic ion: ion charge

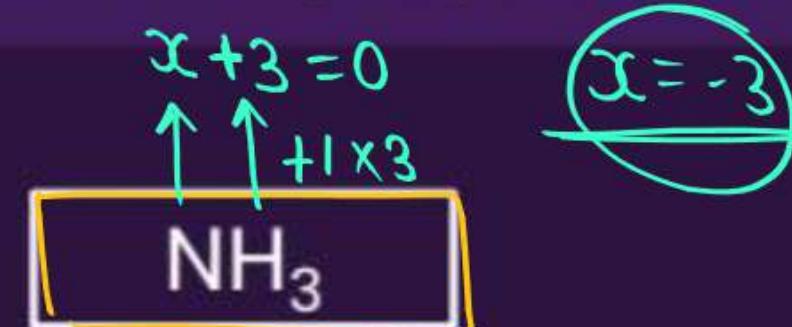
[Ionic / Electrovalent Compounds
are electrically neutral]



OXIDATION NUMBER/OXIDATION STATE

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 - Fluorine: Always -1.
-
- Sum of oxidation number of a neutral compound: 0
 - Sum of oxidation number for a polyatomic ion: ion charge

Ammonia | Nitrogen trihydride



OXIDATION NUMBER/OXIDATION STATE

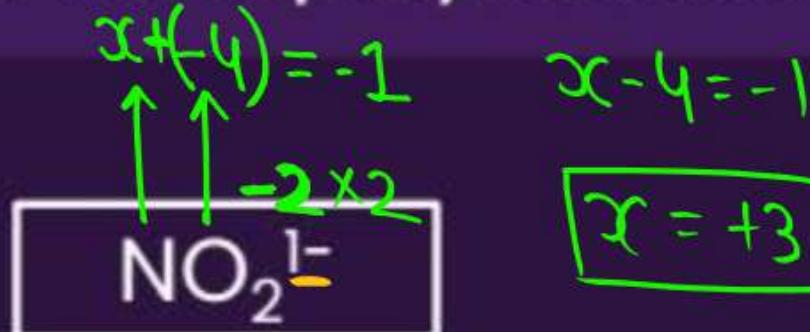
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- Monoatomic ion: Ion charge
- Hydrogen: +1 with non-metals and -1 with metals
- Oxygen: Usually -2. -1 in peroxide and $-1/2$ in superoxide
- Fluorine: Always -1.

- Sum of oxidation number of a neutral compound: 0

- [Sum of oxidation number for a polyatomic ion: ion charge]

Polyatomic Ion: Two or more than two atoms with an overall charge

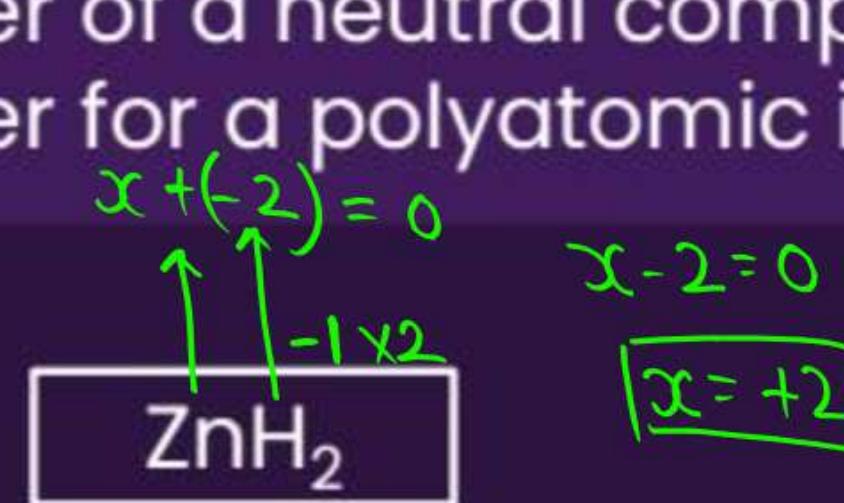
two or more than two atoms overall charge



OXIDATION NUMBER/OXIDATION STATE

- Elements: 0
 - Group 1A: Always +1
 - Group 2A: Always +2
 - Halogens: Usually -1, positive with oxygen and fluorine
 - Monoatomic ion: Ion charge
 - Hydrogen: +1 with non-metals and -1 with metals ✓
 - Oxygen: Usually -2. -1 in peroxide and -1/2 in superoxide
 - Fluorine: Always -1.
- [Sum of oxidation number of a neutral compound: 0]
- Sum of oxidation number for a polyatomic ion: ion charge

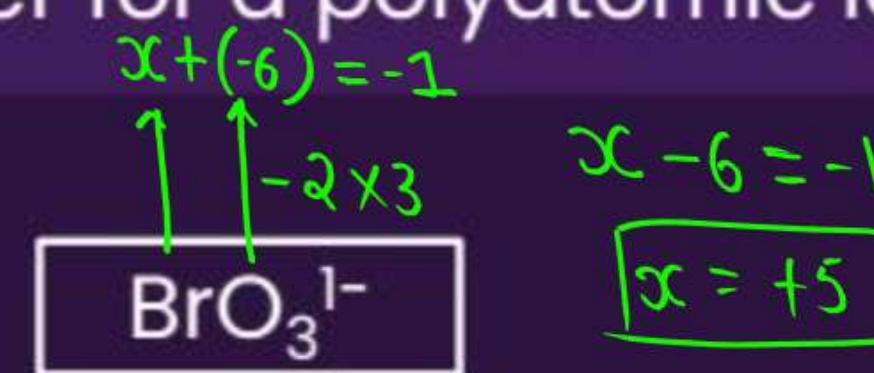
[Zinc hydride]



OXIDATION NUMBER/OXIDATION STATE

- Elements: 0
 - Group 1A: Always +1
 - Group 2A: Always +2
 - Halogens: Usually -1,
(positive with oxygen and fluorine)
F, Cl, Br, I
 - Monoatomic ion: Ion charge
 - Hydrogen: +1 with non-metals and -1 with metals
 - Oxygen: Usually -2. -1 in peroxide and -1/2 in superoxide
 - Fluorine: Always -1.
-
- Sum of oxidation number of a neutral compound: 0
 - Sum of oxidation number for a polyatomic ion: ion charge

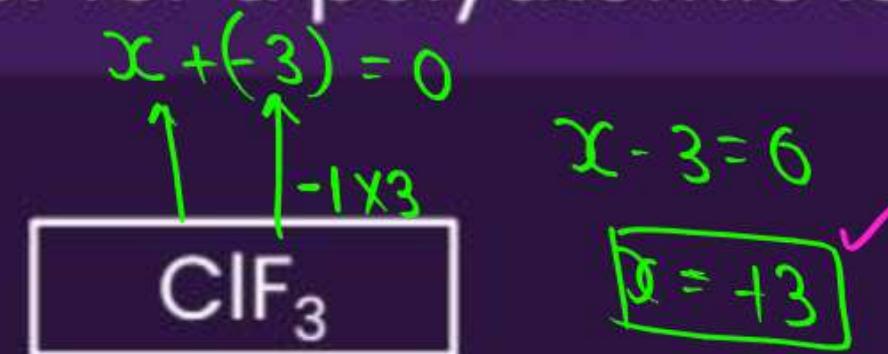
'Bromate Ion'

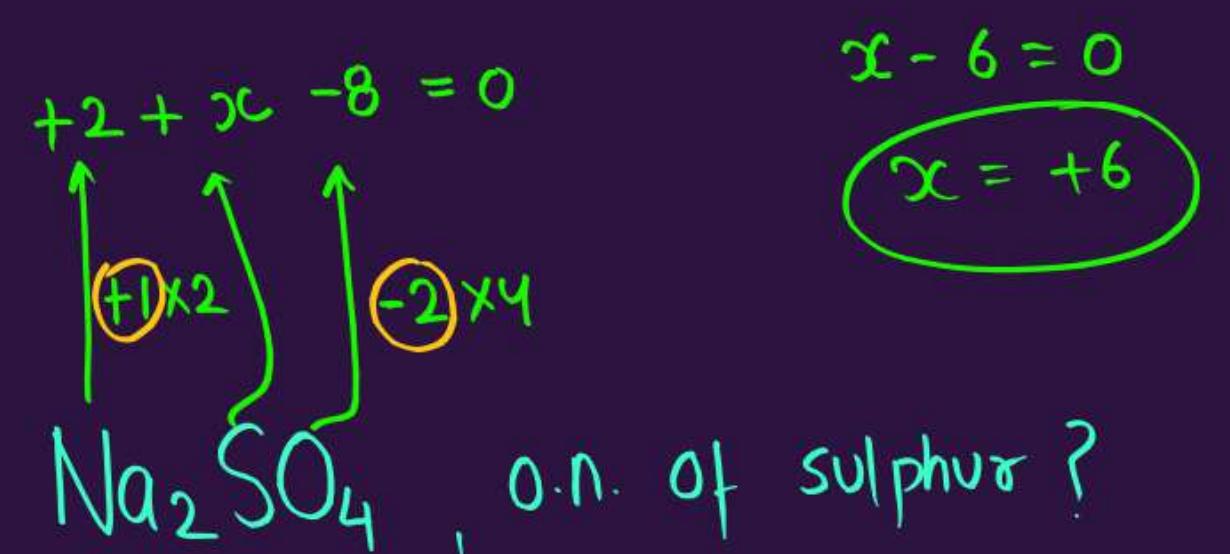


OXIDATION NUMBER/OXIDATION STATE

- Elements: 0
 - Group 1A: Always +1
 - Group 2A: Always +2
 - Halogens: Usually -1,
*f, Cl, (positive with oxygen)
B, O, I → and fluorine)*
 - Monoatomic ion: Ion charge
 - Hydrogen: +1 with non-metals and -1 with metals
 - Oxygen: Usually -2. -1 in peroxide and $-1/2$ in superoxide
 - Fluorine: Always -1.
-
- Sum of oxidation number of a neutral compound: 0
 - Sum of oxidation number for a polyatomic ion: ion charge

[Chlorine trifluoride]





$$x - 6 = 0$$

$$x = +6$$

O.N. of elements in Na_2SO_4



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LIKH DO.

AYE BHAIYA ✓



OXIDATION NUMBER – REDUCTION AND OXIDATION

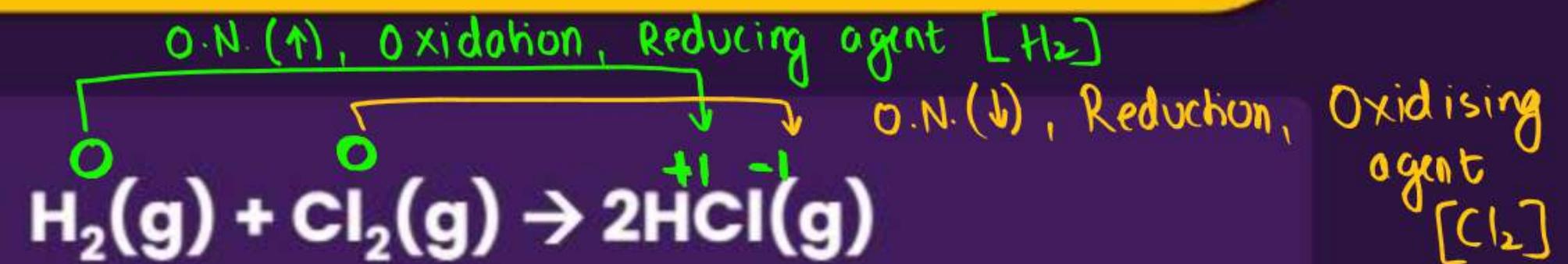
Loss of electron(s) → OXIDATION

gain of electron(s) – Reduction

- (i) **Oxidation:** An increase in oxidation number, which occurs when an element or ion loses electrons. O.N. (\uparrow), Oxidation
- (ii) **Reduction:** A decrease in oxidation number, which occurs when an element or ion gains electrons. O.N. (\downarrow), reduction
- (iii) **Oxidising Agent:** It oxidises other substances (loss of electrons on others) and increases their oxidation number; itself gets reduced (accepts electrons) and decreases its oxidation number.
- (iv) **Reducing Agent:** It reduces other substances (gain of electrons on others) and decreases their oxidation number; itself gets oxidised (loses electrons) and increases its oxidation number.



REDOX REACTION - EXAMPLE



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LIKH DO.

AYE BHAIYA ✓





INTRODUCTION TO ELECTROCHEMISTRY

INTRODUCTION TO ELECTROCHEMISTRY

- It is a branch of physical chemistry that deals with the interconversion of chemical and electrical energy.

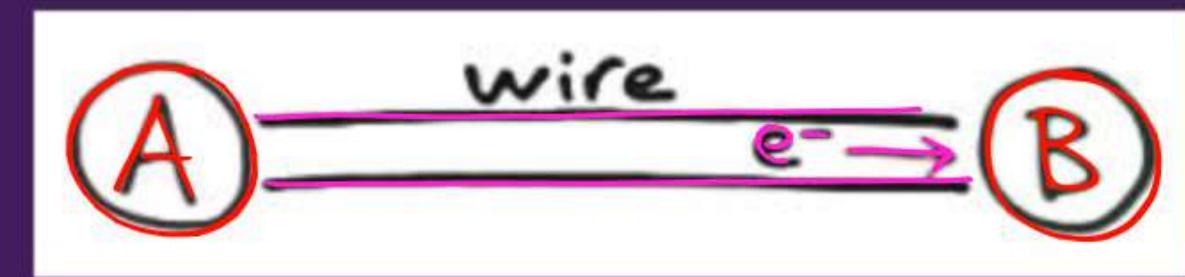
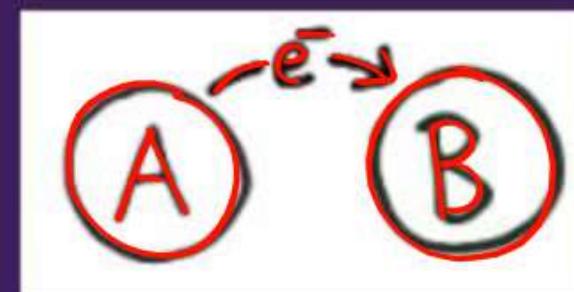
(अपनाए दैन वाला)

- Certain chemical reactions create electricity (flow of electrons)

Spontaneous

Redox

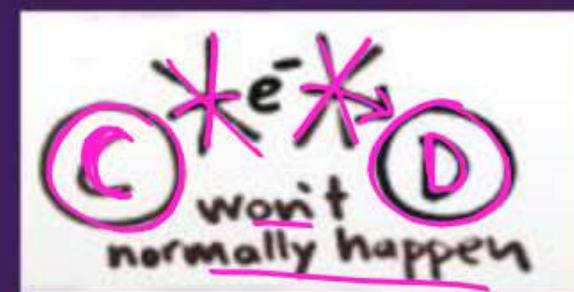
Reaction



- Electricity can make certain chemical reactions happen that wouldn't happen otherwise.

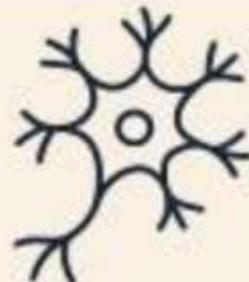
Non-spontaneous

Redox Reaction



IMPORTANCE & SCOPE OF ELECTROCHEMISTRY

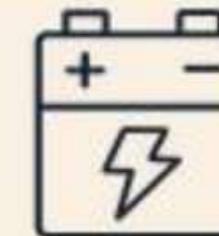
IV



NERVE SIGNALS

Transmission of nerve impulses involves electrochemical processes

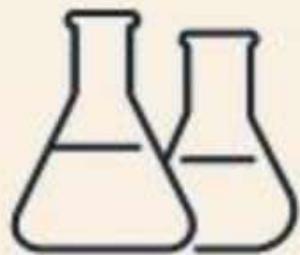
III



BATTERIES AND FUEL CELLS

Convert chemical energy to electrical energy, powering various devices

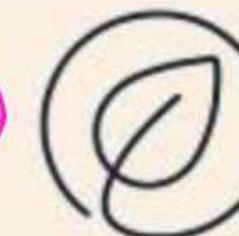
I



PRODUCTION OF CHEMICALS

Metals, NaOH, Cl₂, F₂, etc. are produced using electrochemical methods

II



ECO-FRIENDLY METHODS

Electrochemical processes are often energy-efficient and less polluting

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LIKH DO.

AYE BHAIYA ✓



CONCEPT POLISH – HOMEWORK



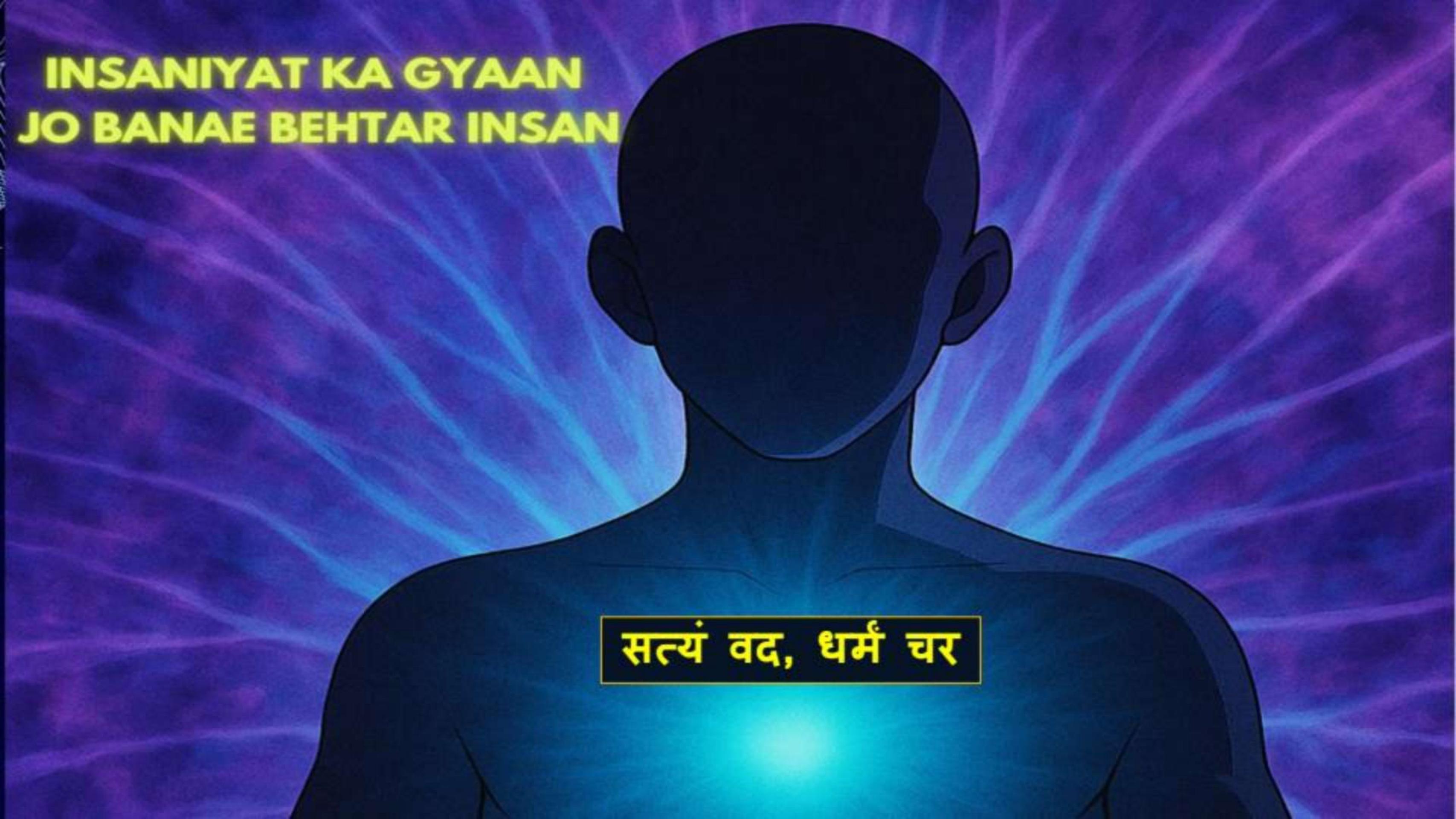
GIVE A THOUGHT

Identify which substance is reduced and which is oxidized in the below reaction.



H.W.

**INSANIYAT KA GYAAN
JO BANAE BEHTAR INSAN**



सत्यं वद, धर्मं चर

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SUNIL BHAIYA IS ALWAYS THERE FOR YOU.

#sbsathhai (✓)

#pwsathhai (✓)

Thank
You



PARISHRAM



2026

CHEMISTRY

Lecture 02

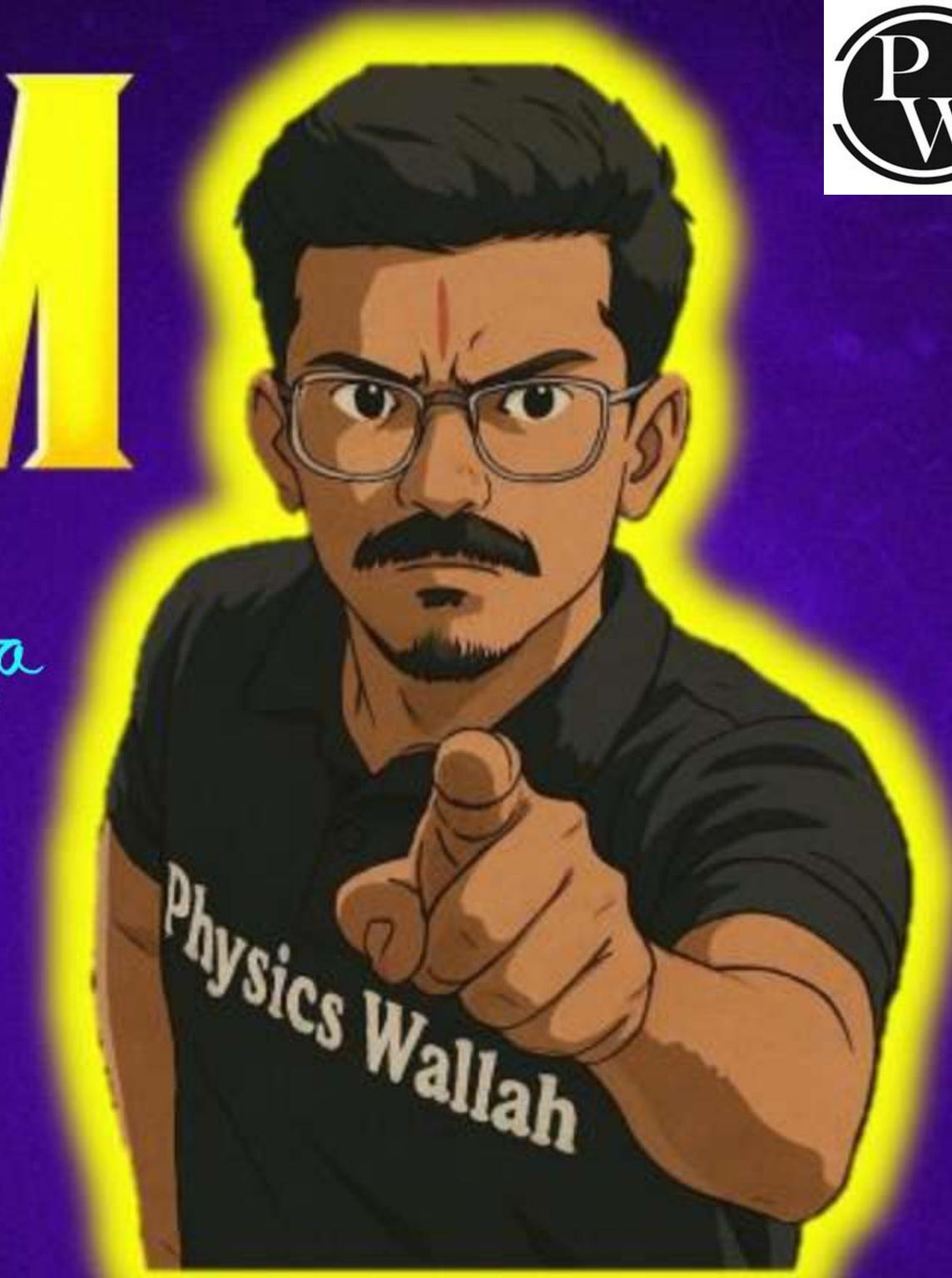
ELECTROCHEMISTRY

Galvanic Cell

Bharat Mata
Ki Jai ❤

Physics Wallah

BY – PRIYA-PUTRA-SUNIL



TOPICS TO BE COVERED

- (i) Galvanic Cell ✓
- (ii) Cell Notation ✓



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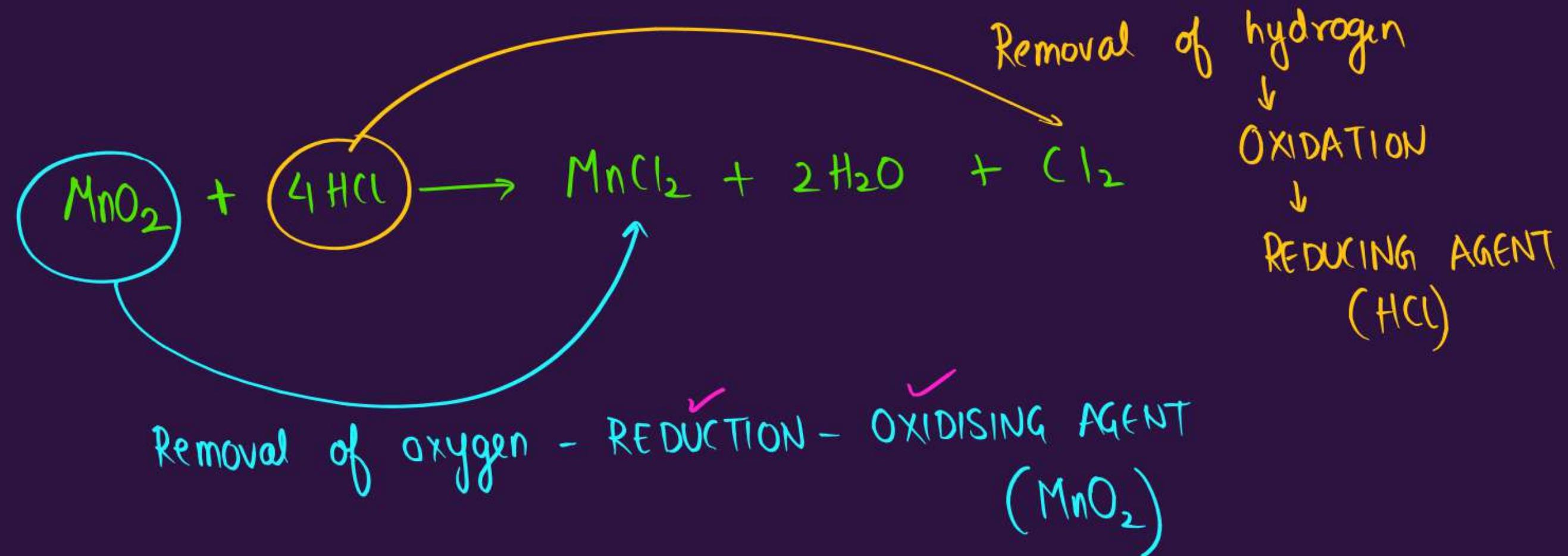
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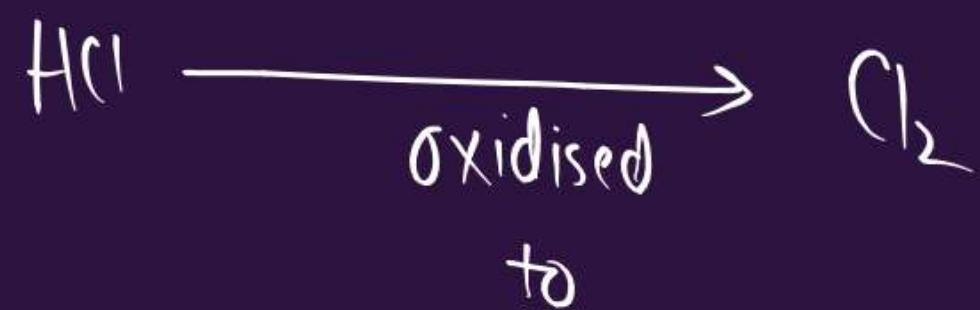
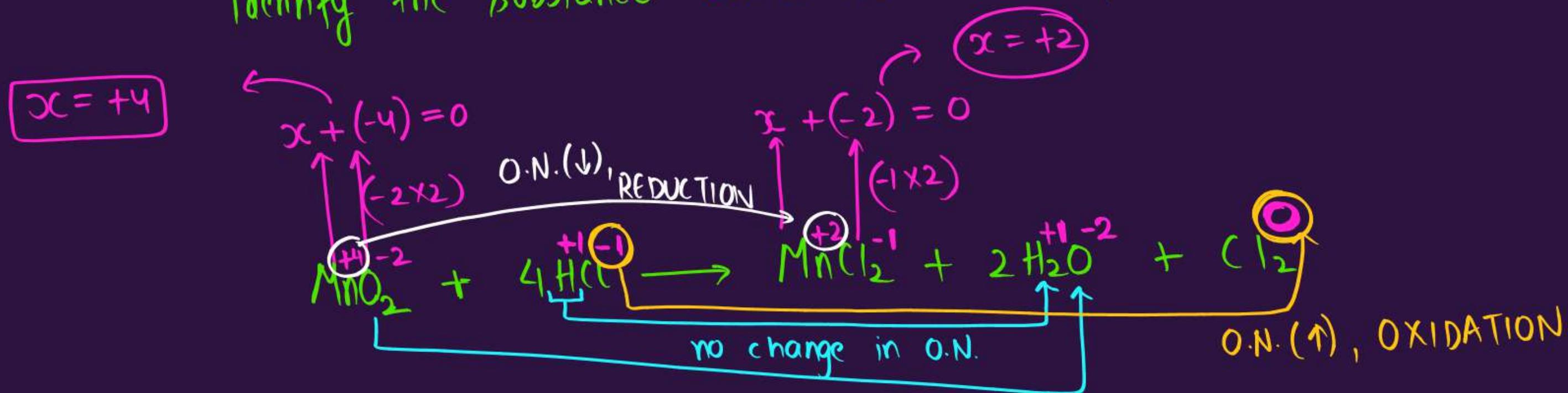
CONCEPT POLISH - HOMEWORK DISCUSSION



Identify the substance which is reduced & oxidised in:



Identify the substance which is reduced & oxidised in:



✓
GALVANIC CELL

aye
bhanya

$$x + (-8) = -2$$

$$\uparrow \uparrow (-2 \times 4)$$

$$x = +6$$

$$SO_4^{2-}$$

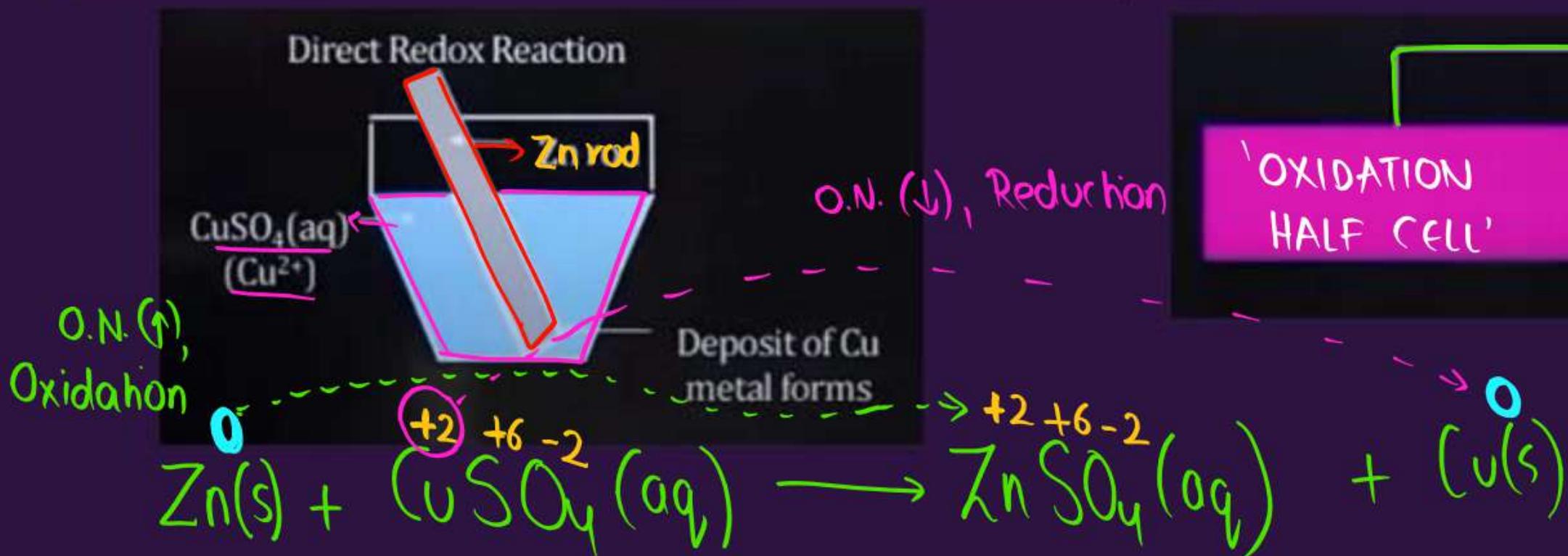
DIRECT AND INDIRECT REDOX REACTIONS

DIRECT REDOX REACTIONS

In this oxidation and reduction are carried out in the same vessel.

INDIRECT REDOX REACTIONS

In this oxidation and reduction are carried out in the separate vessel.



TYPES OF ELECTROCHEMICAL CELL

(C-I)

A device in which interconversion of electrical and chemical energy takes place is called an electrochemical cell.

(C-II)

GALVANIC OR VOLTAIC CELL

(i) Chemical Energy → Electrical Energy

(ii) In this indirect redox reaction takes place.

(iii) SPONTANEOUS Rxn

ELECTROLYTIC CELL

(i) Electrical Energy → Chemical Energy

(ii) In this direct redox reaction takes place.

(iii) NON SPONTANEOUS Rxn

LET'S PRACTICE



QUESTION

Which of the following is correct? ✓

[C.E.: Chemical Energy and E.E.: Electrical Energy]

A

Galvanic Cell: [Indirect Redox Reaction] ✓ [C.E. → E.E.] ✓

B

Galvanic Cell: Direct ✗ Redox Reaction, E.E. → C.E.

C

Electrolytic Cell: Indirect Redox Reaction, E.E. → C.E. ✗

D

Electrolytic Cell: Direct Redox Reaction, C.E. → E.E. ✓ ✗

GENERAL FEATURES OF GALVANIC/VOLTAIC CELL

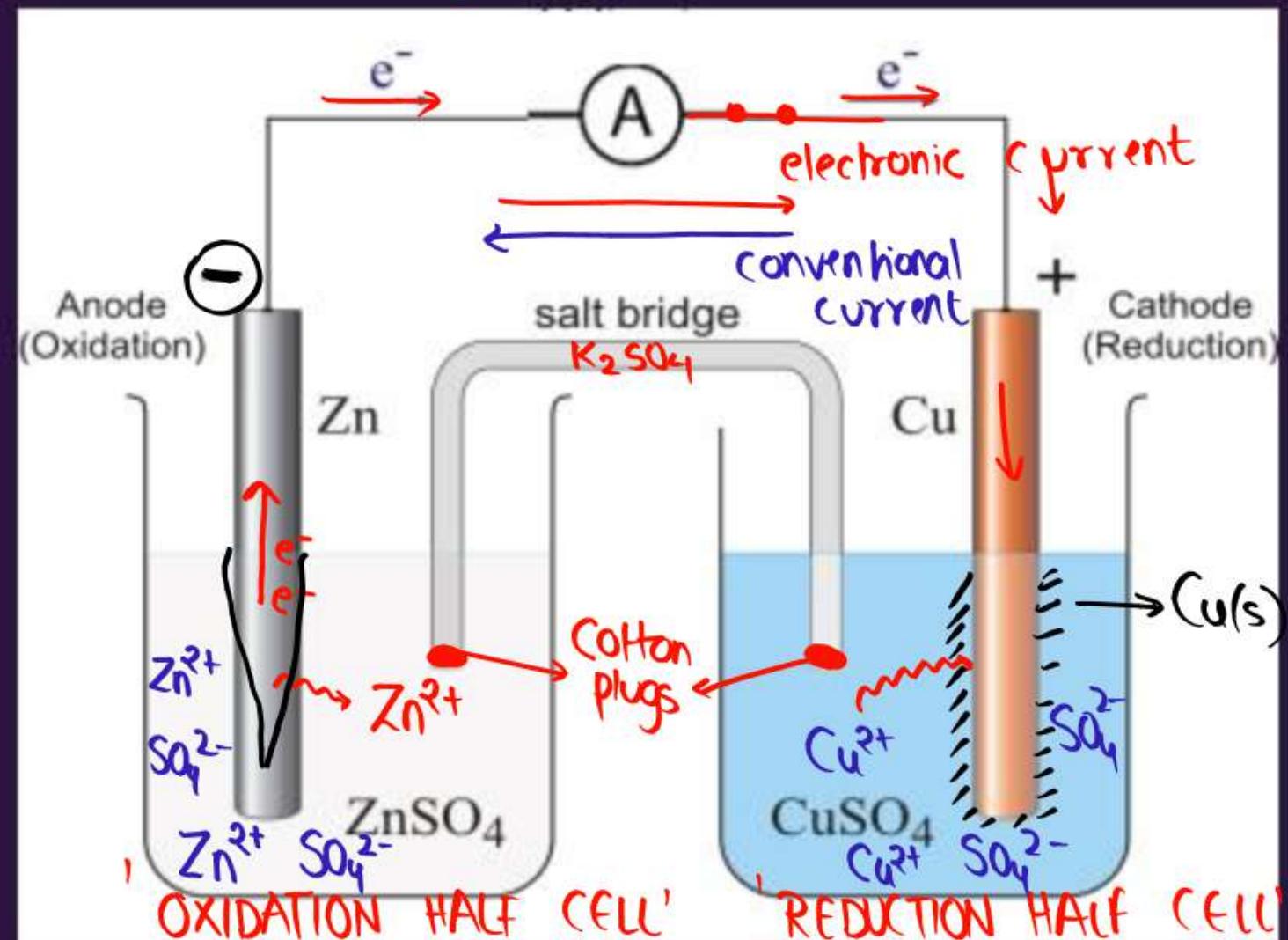
- ✓ It is a type of electrochemical cell that use spontaneous chemical reaction (indirect redox reaction) to create electrical energy.
- Oxidation and reduction happens in separate vessels.
Oxidation → [Oxidation Half Cell]; Reduction → [Reduction Half Cell]
- Each half cell consists of an **electrode** dipped in aqueous solution of electrolyte.
 - An electrolyte is a chemical substance that breaks into ions in aqueous or molten form and conducts electricity.
 - An electrode is a conductor through which electric current enters or leaves an electrolyte in an electrochemical cell (like a battery, galvanic cell, or electrolytic cell).

GENERAL FEATURES OF GALVANIC/VOLTAIC CELL

- ✓ Anode (-ve electrode) -> Oxidation; Cathode (+ve electrode)
→ Reduction [gain of electron(s)] loss of electron(s)
- Two half cells are connected by a copper wire through an ammeter and a switch (*External circuit*).
- The electrolytes of two half cells are connected to each other by a salt bridge (*Internal circuit*).
- Salt bridge consists of an inert electrolyte such as KCl, K_2SO_4 , NaCl etc. in agar-agar.
(jelly-like substance) ↓
 that don't react with electrolyte or electrodes of any half cell

DANIELL CELL

↳ Type of Galvanic / Voltic Cell



- Switch OFF: No deflection in ammeter
- Switch ON: Deflection in ammeter

Let's investigate this!

Zn is more reactive than Cu & loses electron(s) easily.

✓ Anode - Oxidation - Loss of Electrons

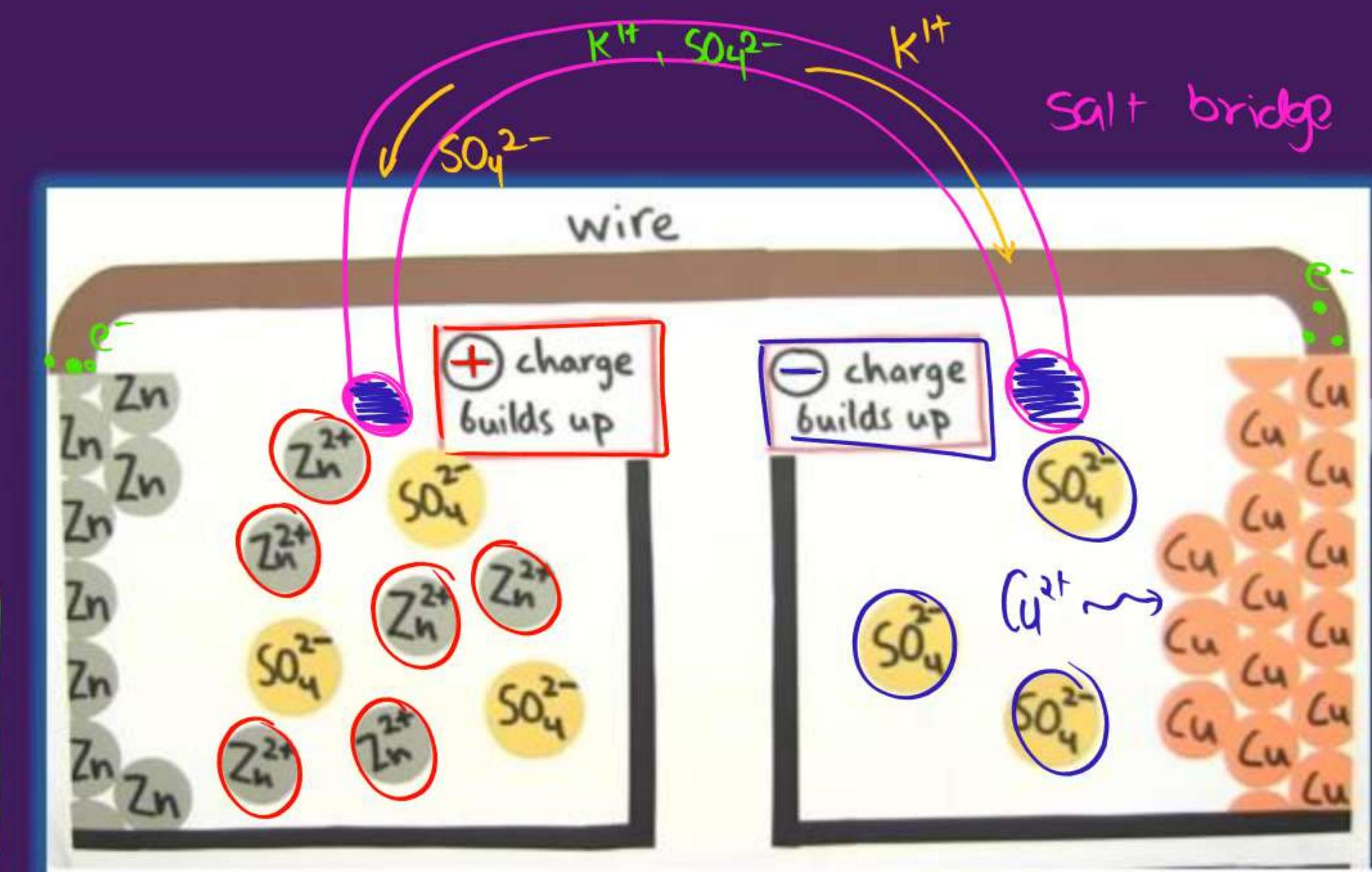


Cathode - Reduction - Gain of Electrons



FUNCTIONS OF A SALT BRIDGE

- With time anode becomes thinner and cathode becomes thicker.
- As +ve charge builds-up in oxidation half cell it won't allow -vely charged particles (electrons) to move to cathode.
- Increase in -ve charge in reduction half cell won't allow $\text{Cu}^{2+}(\text{aq})$ to accept electron(s).
- Salt Bridge:** It helps to maintain the electrical neutrality in two half cells due to movement of ions which ensures the flow of current in the external circuit.



SUMMARY TABLE

Summary Table:

Part	Electrode	Reaction	Charge
Zinc rod	Anode	Oxidation loss of electron(s)	Negative ✓
Copper rod	Cathode	Reduction gain of electron(s)	Positive ✓
Wire	—	Electron flow →	$\text{Zn} \rightarrow \text{Cu}$
Salt Bridge	—	Ion flow ✓	Maintains neutrality

**SAMAJ AAYA TOH
LIKH DO.
AYE BHAIYA**





A large, solid yellow oval shape centered on a dark purple background, serving as a title bar.

CELL NOTATION

CELL NOTATION

C-I

- A shorthand abbreviation of the chemical reactions in a voltaic cell or galvanic cell.
- Left: Oxidation (Anode)
- Right: Reduction (Cathode)
- An oxidising and reducing agent which appear on opposite sides of a half-equation constitute a redox couple.

ex:

Anode - Oxidation - Loss of Electrons

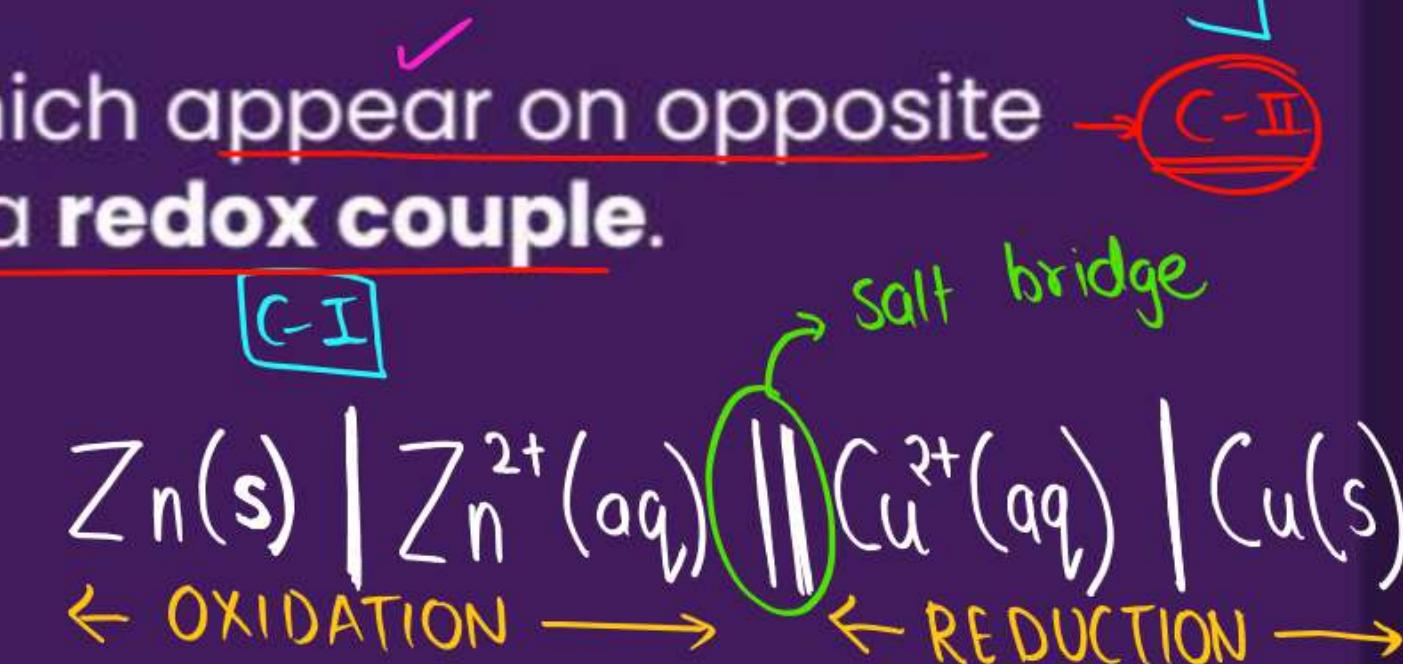


C-I

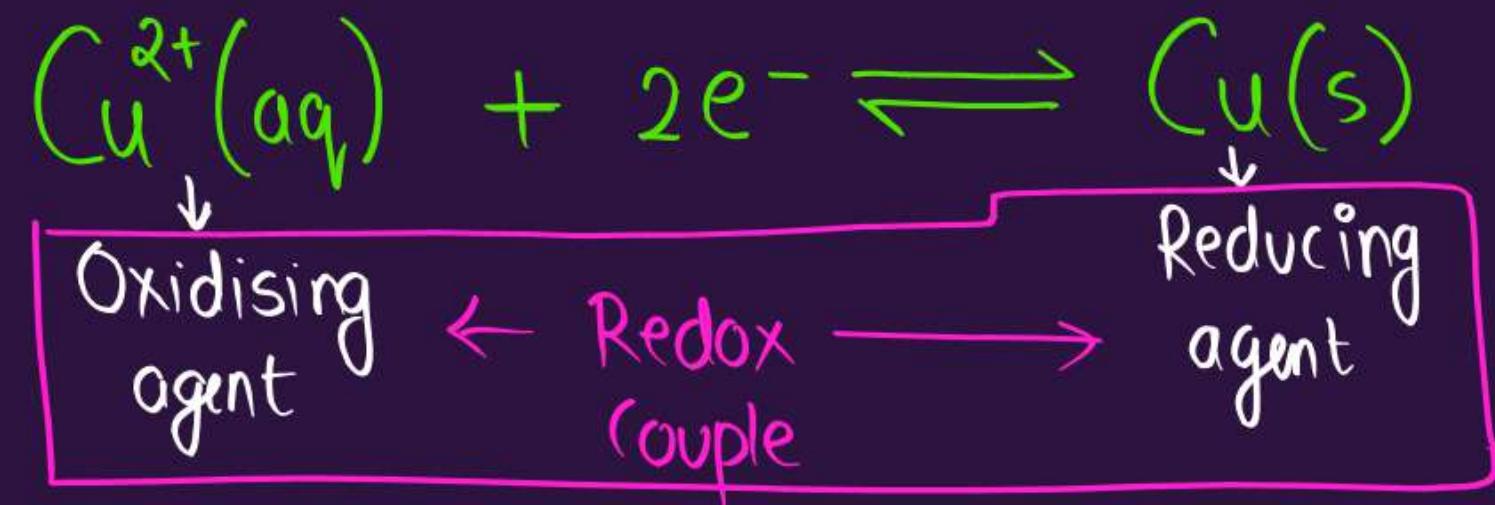
Cathode - Reduction - Gain of Electrons



C-I



(C-II)

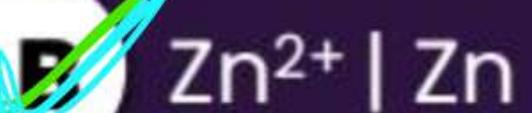


LET'S PRACTICE



QUESTION

Which of the following is a correct redox couple?



SAMAJ AAYA TOH
LIKH DO.

AYE BHAIYA ✓



CONCEPT POLISH – HOMEWORK

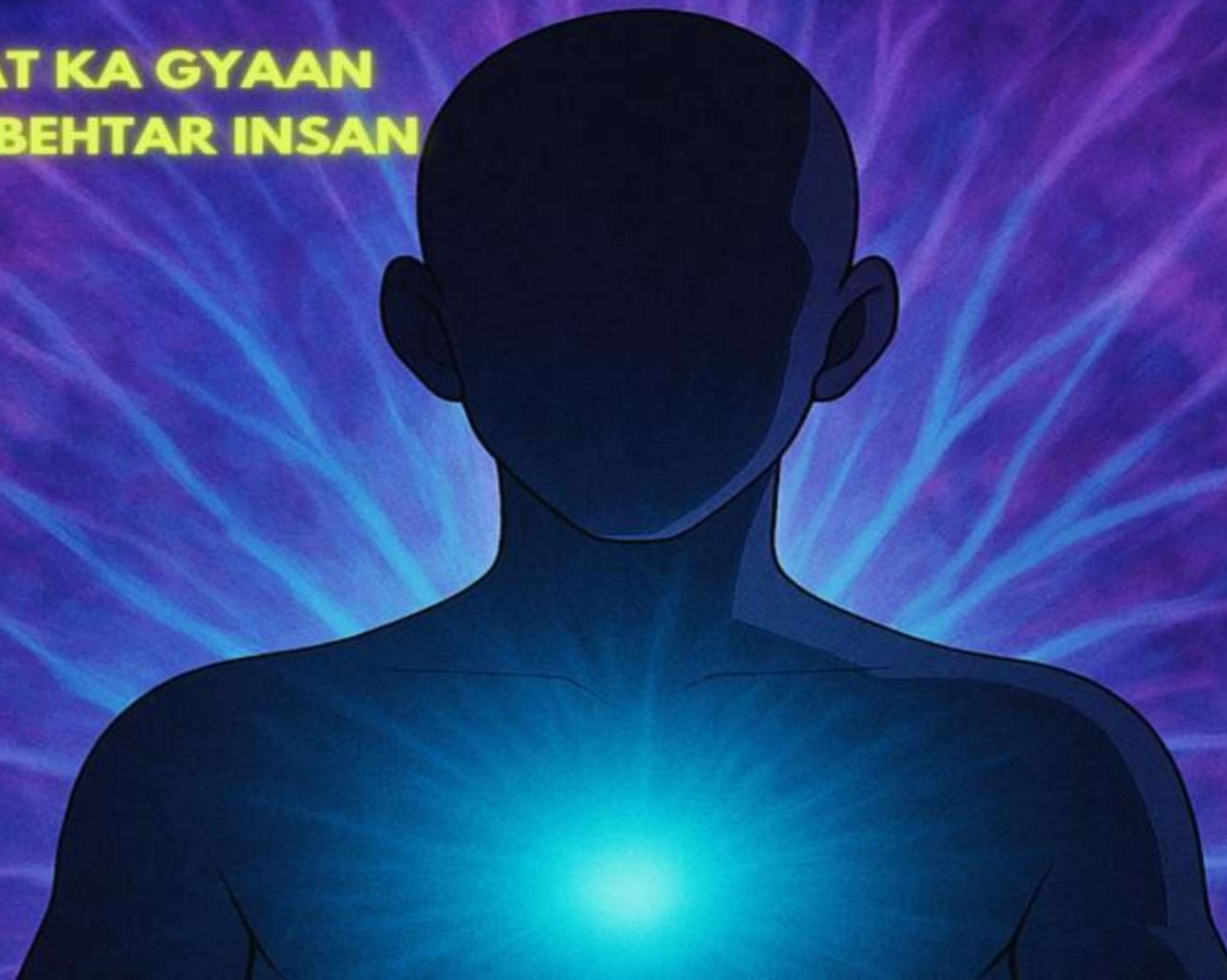


QUESTION

Do a research on:

‘Sometimes, both the electrodes dip in the same electrolyte solution and in such cases we do not require a salt bridge’

**INSANIYAT KA GYAAN
JO BANAE BEHTAR INSAN**



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SUNIL BHAIYA IS ALWAYS THERE FOR YOU.

#sbsathhai ✓

#pwsathhai ✓

Thank
You

PARISHRAM



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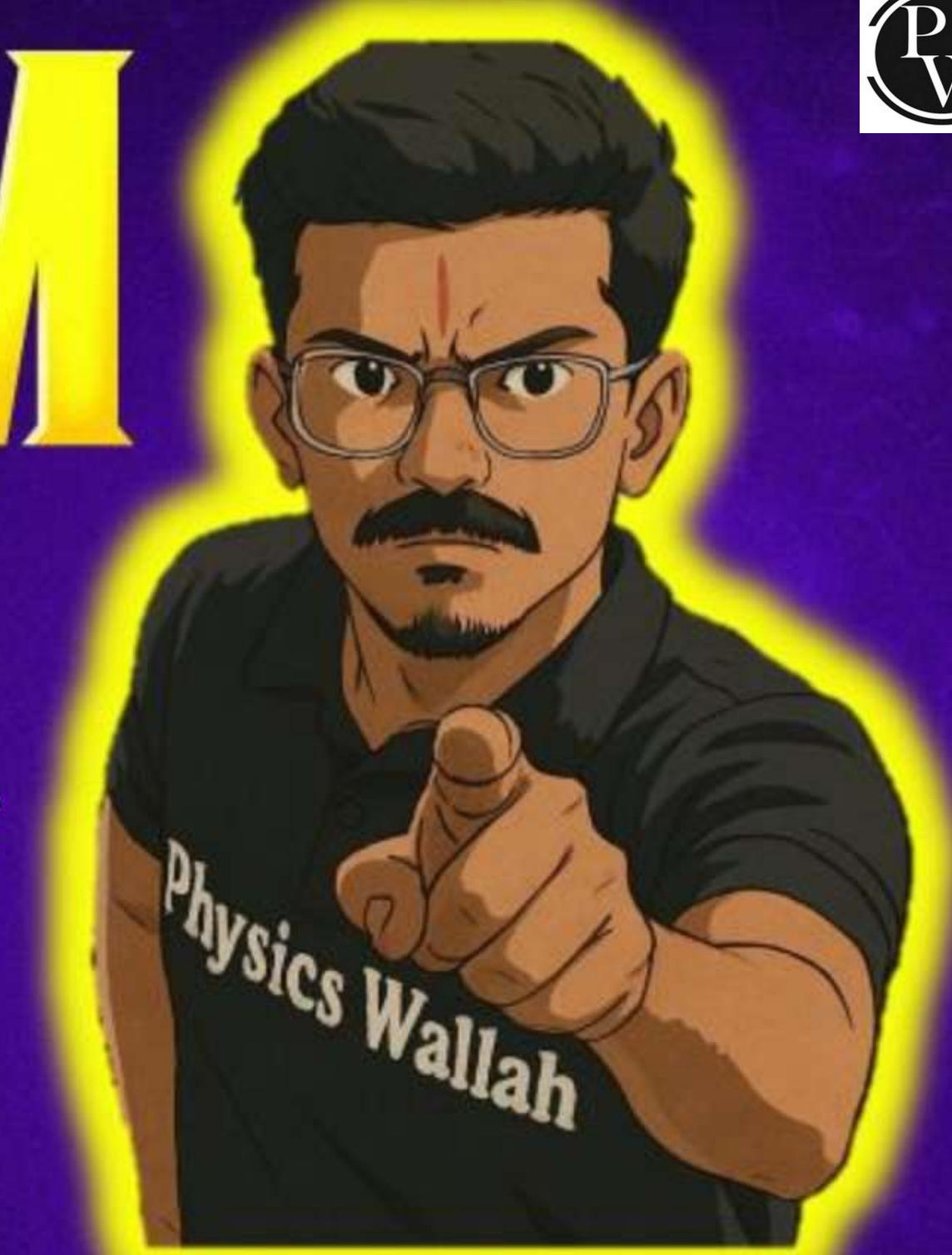
CHEMISTRY

Lecture 03

ELECTROCHEMISTRY

Doubts, Electrode Potential and
Its Measurement

Bharat Mata
Ki Jai ♡



BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

- (i) Doubts (✓)
- (ii) Electrode Potential and Its Measurement (✓)



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CONCEPT POLISH – HOMEWORK DISCUSSION



QUESTION

Here in this case, flow of electrons won't get obstructed because of charge built-up. Hence, no salt bridge is required.

Do a research on:

'Sometimes, both the electrodes dip in the same electrolyte solution and in such cases we do not require a salt bridge'

This is possible! ✓

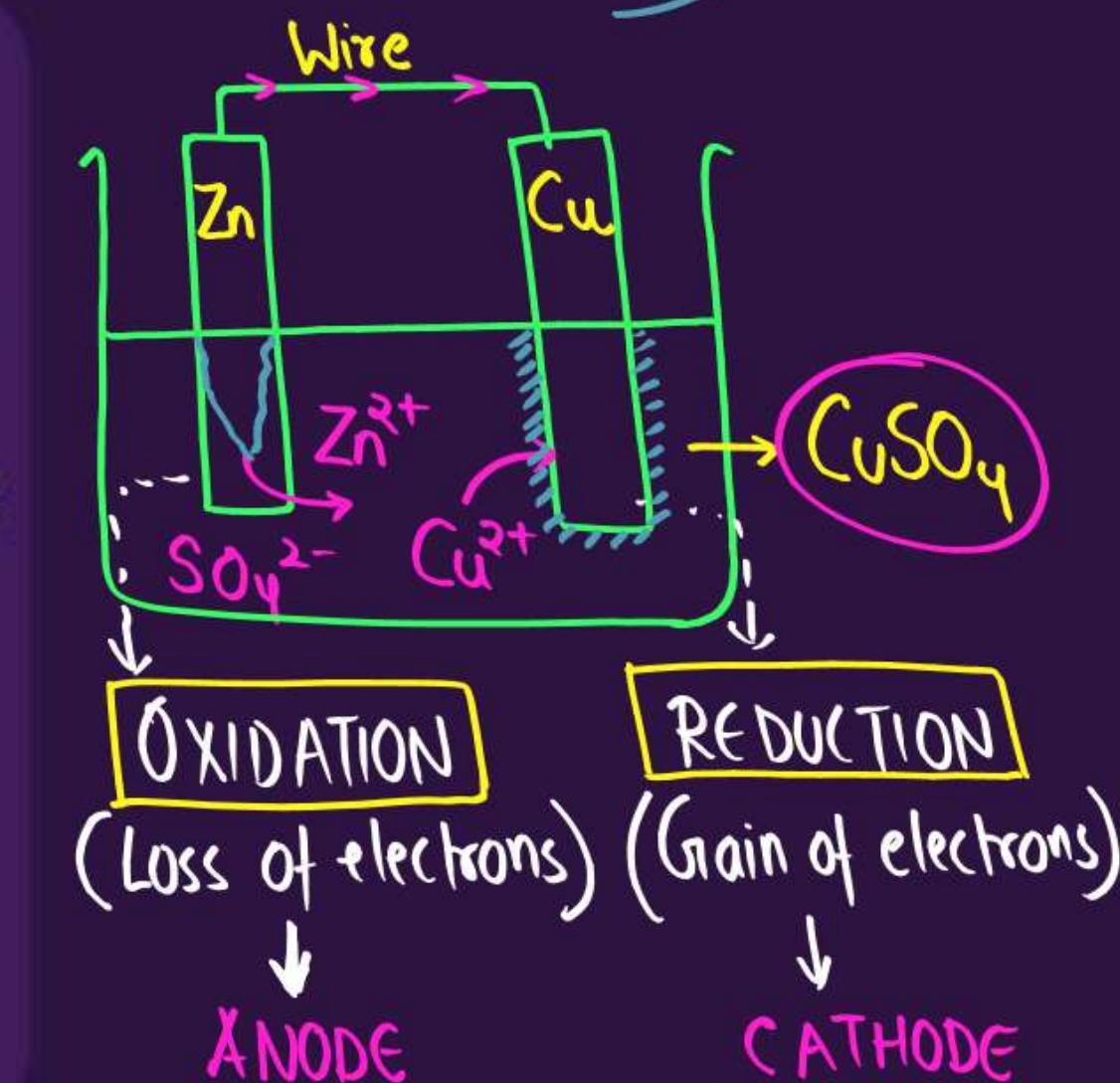
→ Spontaneous Rxn

→ Direct Redox Rxn

- Zn and Cu rods dipped in CuSO₄ solution.
- Both the rods are connected through a wire.

→ Zn will lose electrons to form Zn²⁺ and moves into the solution. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

→ These electrons will travel through the wire and Cu²⁺ from solution will gain these electrons at Cu rod $(Cu^{2+}(aq) + 2e^- \rightarrow Cu(s))$



**SAMAJ AAYA TOH
LIKH DO.
AYE BHAIYA**





DOUBTS ✓

DOUBTS

Bhaiya ek baar please animation ke through dikha dijiye
Galvanic Cell ki working dikha dijiye.



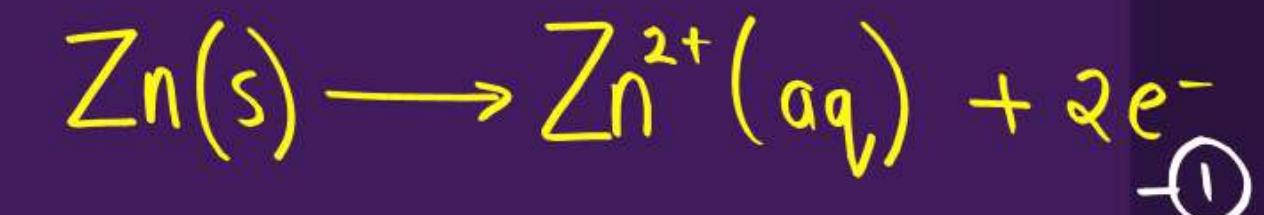
'Daniel Cell'



DOUBTS

Bhaiya ek baar Daniel cell (Zn-Cu cell) ki overall equation bata dijiye.

Anode - oxidation [loss of electrons]



Cathode - reduction [gain of electrons]



① + ②



Overall eqⁿ

DOUBTS

Bhaiya yahan par copper aur zinc ki jagah graphite or platinum' ka electrode hota toh kya woh bhi consume hote hai ya reactions mein participate krte hai?

‘Graphite (form of C) and Platinum’ are inert electrodes which do not involve themselves in the chemical reactions, (i.e. they don’t undergo oxidation or reduction). Their job is to provide a surface for electron transfer when no reactive metal is available.

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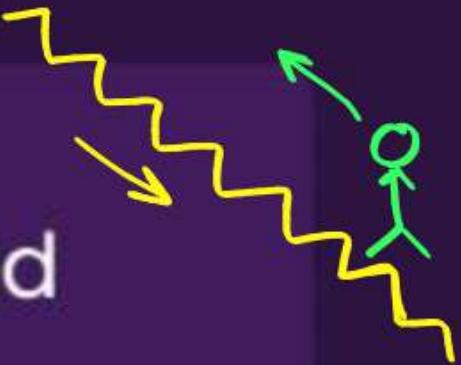


ELECTRODE POTENTIAL AND ITS MEASUREMENT

ELECTRODE POTENTIAL (E)

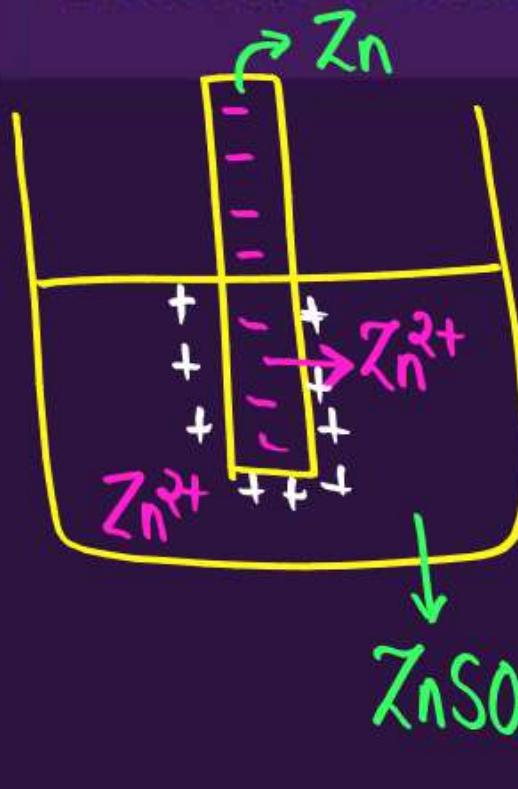
C-I

The potential difference between electrode (metal) and electrolyte (ions of same metal) under equilibrium is called electrode potential.



C-II

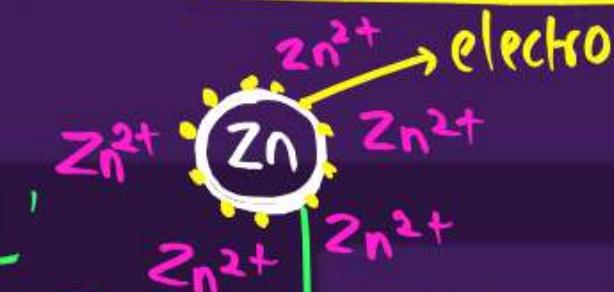
Metal Tendency - Oxidise



at equilibrium

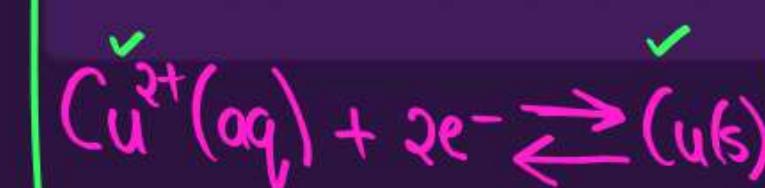


Dominant Rxn (More favoured)

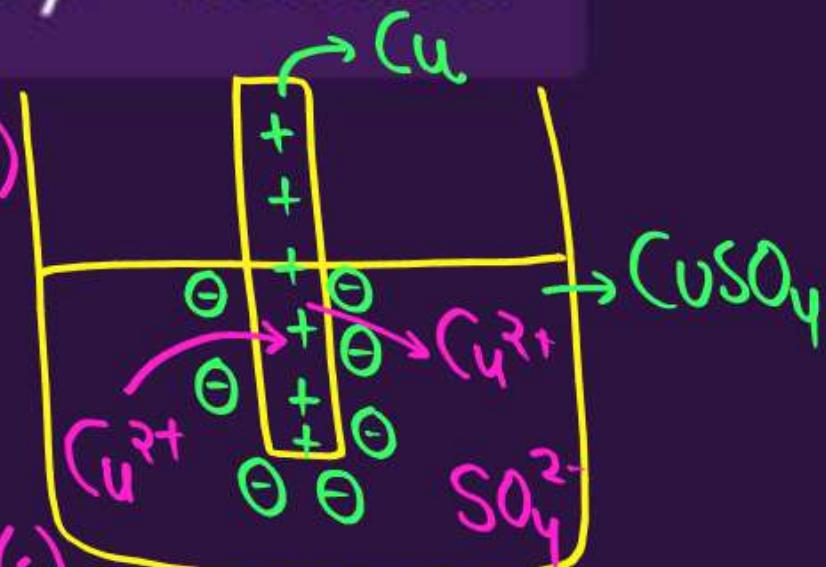
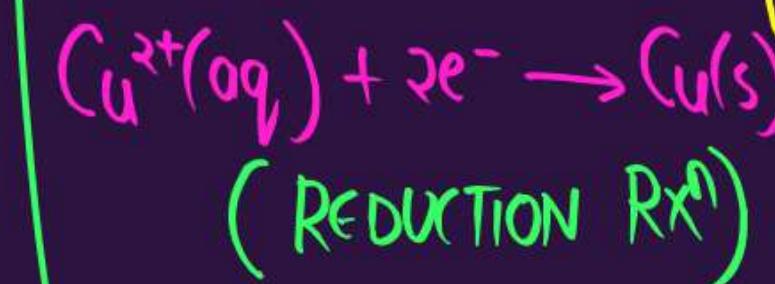


'REDUCTION POTENTIAL'

Metal Tendency - Reduce

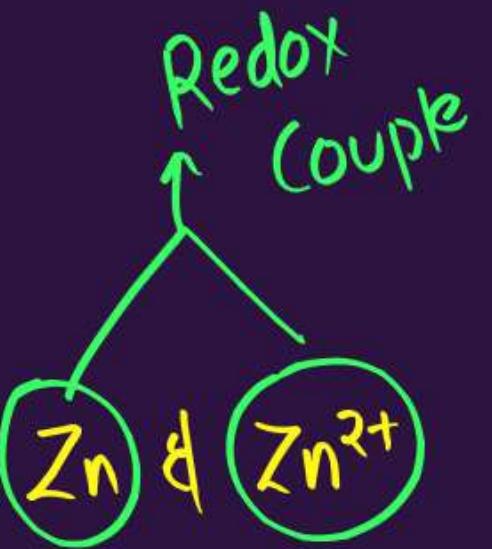


Dominant Rxn
(more favoured rxn)

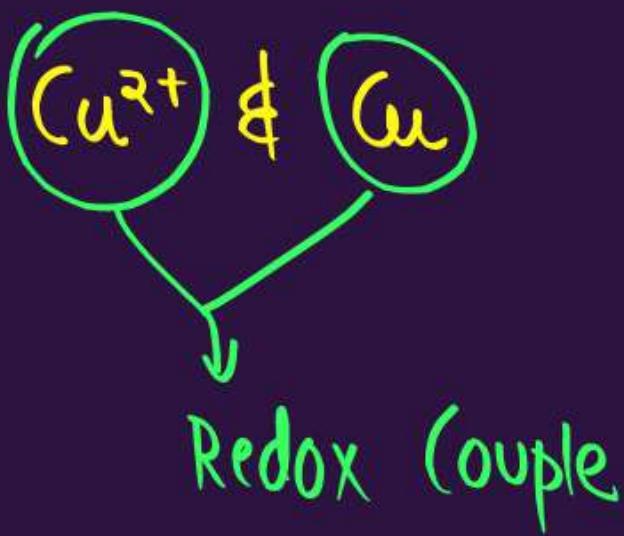


'Imp. to Remember'

I Electrode potential \rightarrow P.D. between



II \rightarrow P.D. between



GIVE A THOUGHT



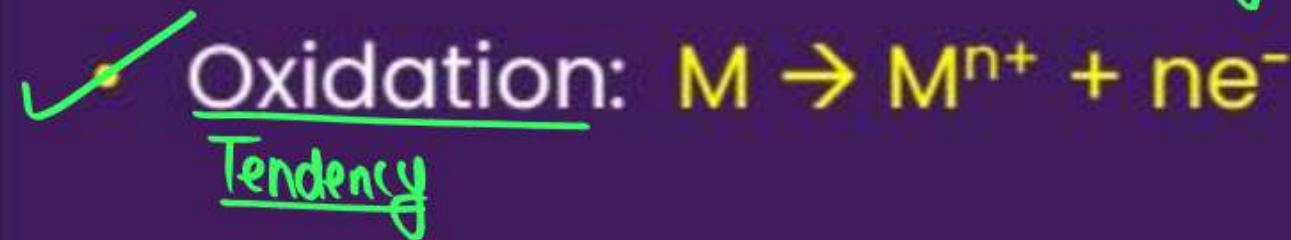
We can't see any visible change in the electrode or the electrolyte because of the equilibrium attained.

- A. YES
- B. NO

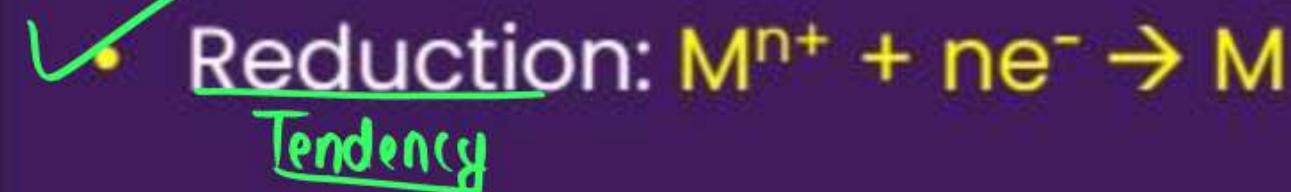
STANDARD ELECTRODE POTENTIAL (E°)

- ✓ Standard electrode potential is the potential difference between electrode (metal) and electrolyte (ions of same metal) measured under equilibrium at standard conditions (25 °C, 1 M concentration, 1 atm pressure)

↳ 'electrolyte'



$$\text{S.R.P.} = -\text{S.O.P.} \quad **$$



- ✓ According to IUPAC convention, standard reduction potentials are now called standard electrode potentials.

GIVE A THOUGHT



or electrode potential

Can we measure the standard electrode potential of a half cell?

- A. YES
- B. NO

GIVE A THOUGHT



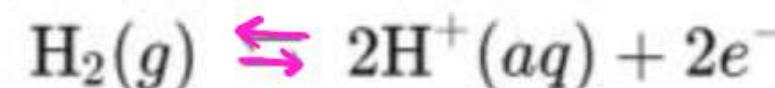
Can we measure the standard electrode potential of a half cell?

- A. YES
- B. NO

No, we can only measure the difference between the potential of two half cells and not standard electrode potential of a half cell.

STANDARD HYDROGEN ELECTRODE (SHE)

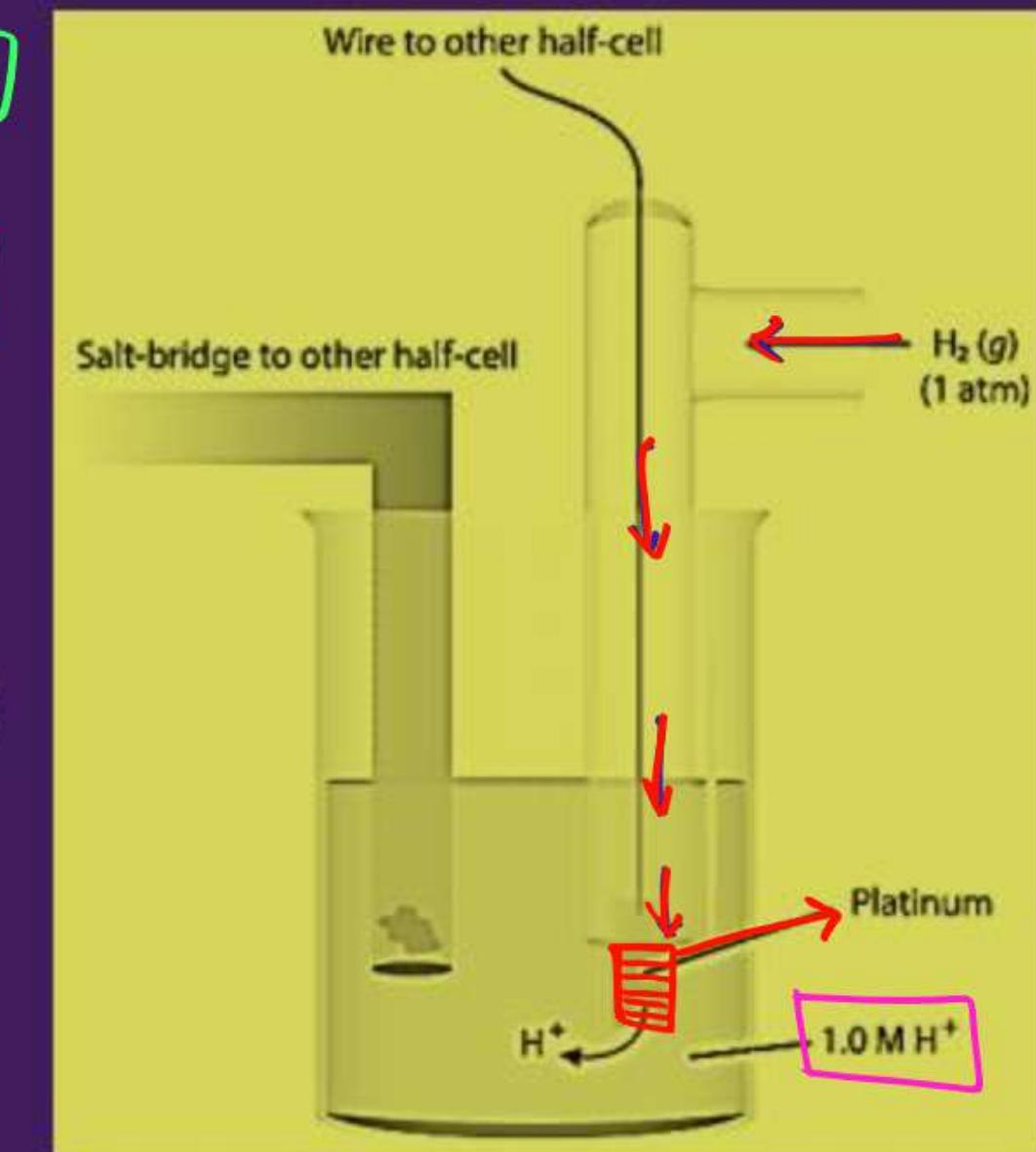
- Electrode: **Platinum (Pt)** – inert surface
- Electrolyte: **1 M H⁺ ions** (like HCl)
- Gas bubbled: **H₂ gas at 1 atm, T = 25 °C**



* The standard electrode potential of hydrogen electrode is considered to be **0 V** because it is used to measure standard electrode potential (E°) of other metals.

Standard reduction potential

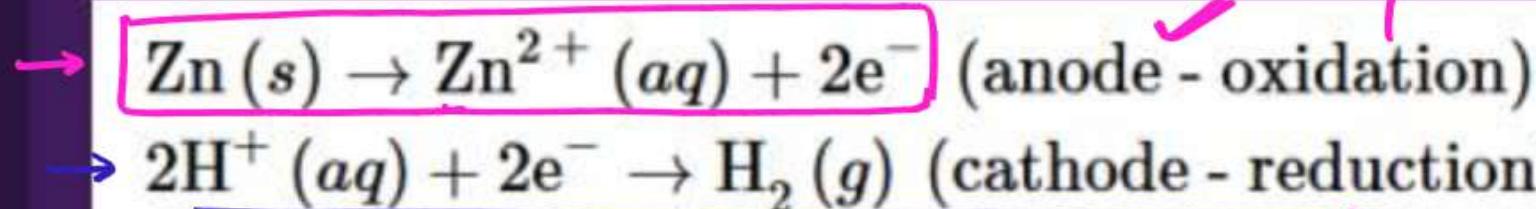
$$E^\circ(\text{H}^+ \mid \text{H}_2) = 0 \text{ V}$$



Reduction potential → Tendency to accept electron(s).

STANDARD HYDROGEN ELECTRODE (SHE)

- When SHE acts as cathode



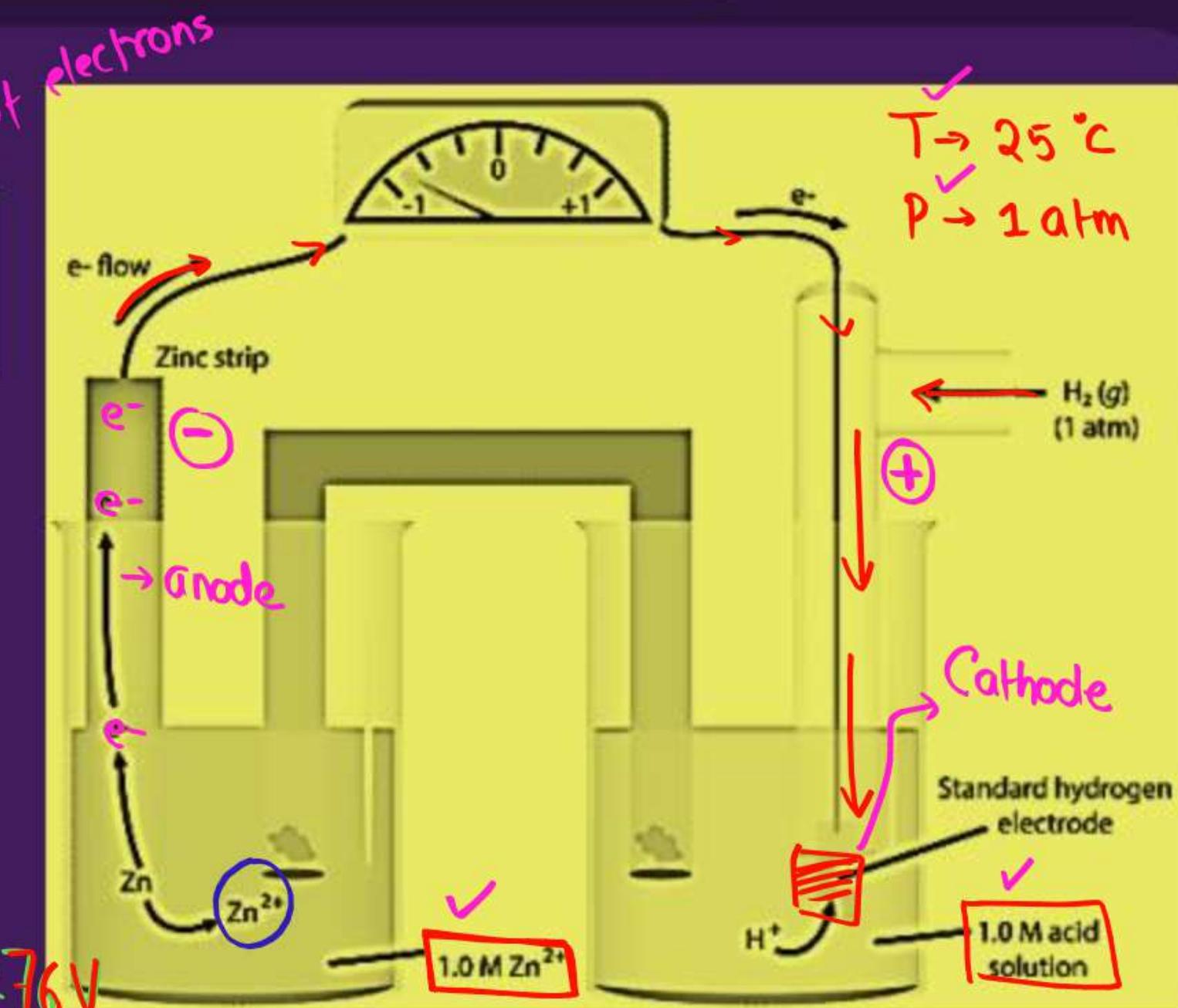
CONCLUSION

- $\boxed{\text{Zn}^{2+} \text{ has lower reduction potential than } \text{H}^+}$
- So $\boxed{\text{Zn}}$ gets oxidised
- SHE side: $\boxed{\text{H}^+ \text{ gets reduced}}$

Standard Electrode Potential
OR

Standard Reduction Potential

$$E^\ominus (\text{Zn}^{2+} | \text{Zn}) = -0.76 \text{ V}$$

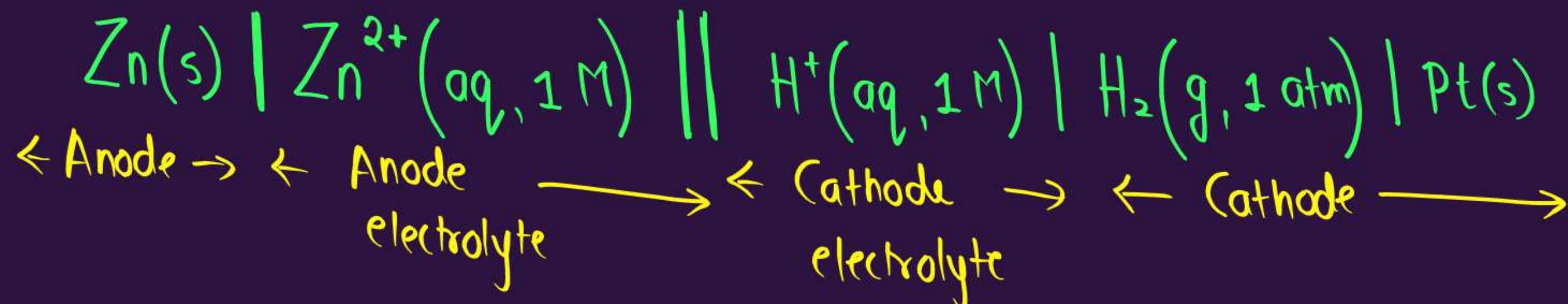




GIVE A THOUGHT



Write the cell notation for the above chemical reaction if the format looks like this: ↓



STANDARD HYDROGEN ELECTRODE (SHE)

- When SHE acts as (anode)
 - $\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$ (anode) \rightarrow oxidation
 - $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ (cathode) \downarrow reduction

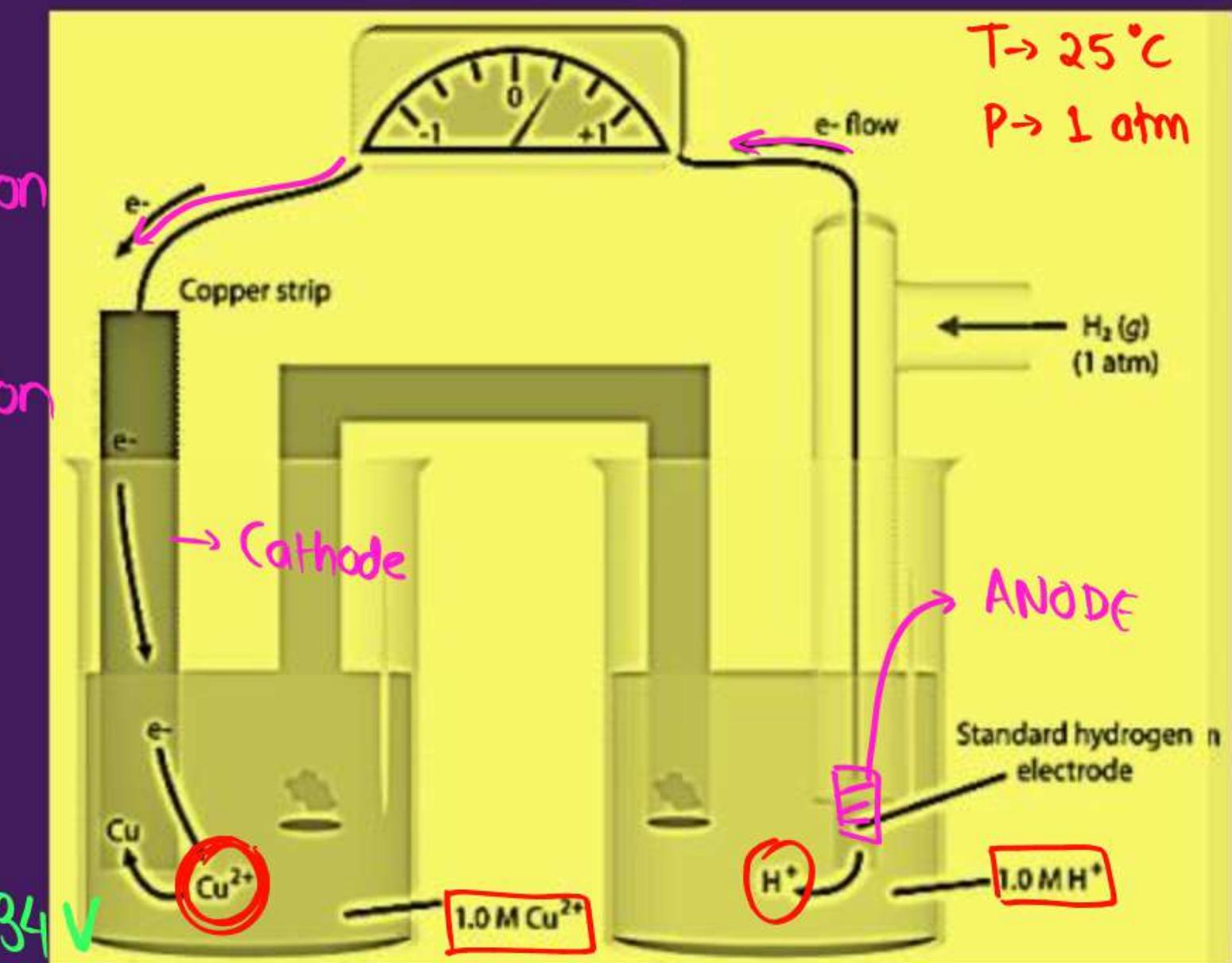
CONCLUSION

- [Cu²⁺ has higher reduction potential than H⁺]
 - So Cu²⁺ gets reduced
 - SHE side: **H₂ gets oxidized**

Standard Electrode Potential
OR

Standard Reduction Potential

$$E^\circ (\text{Cu}^{2+}/\text{Cu}) = +0.34 \text{ V}$$

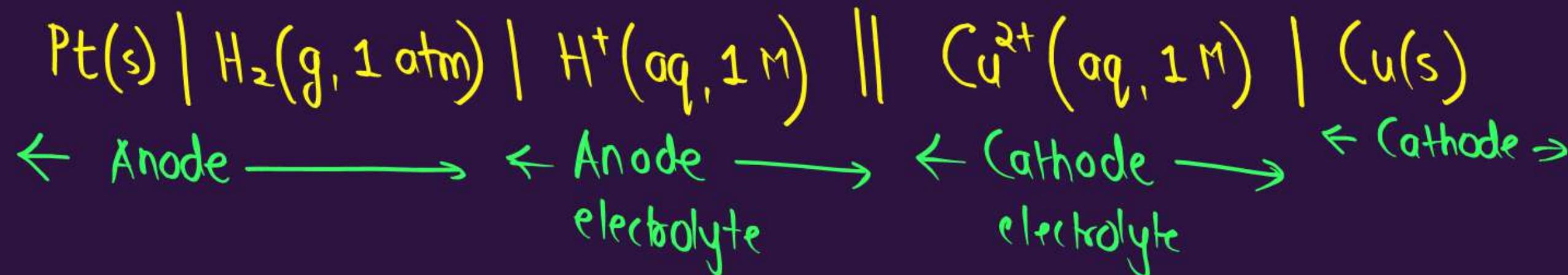




GIVE A THOUGHT



Write the cell notation for the above chemical reaction if the format looks like this:



SAMAJ AAYA TOH
LIKH DO.

AYE BHAIYA ✓

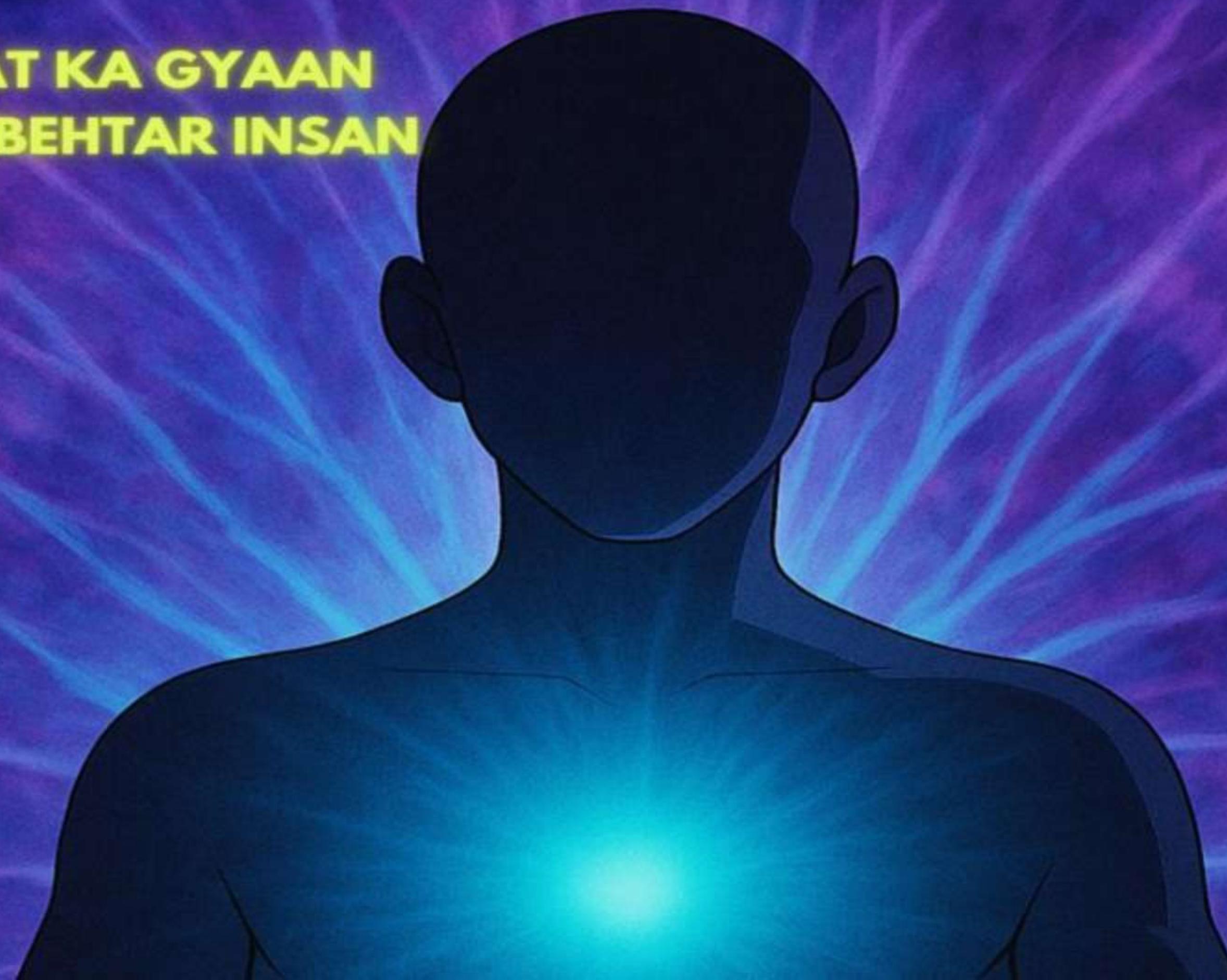


CONCEPT POLISH – HOMEWORK

NO HOMEWORK!



**INSANIYAT KA GYAAN
JO BANAE BEHTAR INSAN**



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SUNIL BHAIYA IS ALWAYS THERE FOR YOU.

#sbsathhai (✓)
#pwsathhai (✓)

Thank
You



PARISHRAM



2026

CHEMISTRY

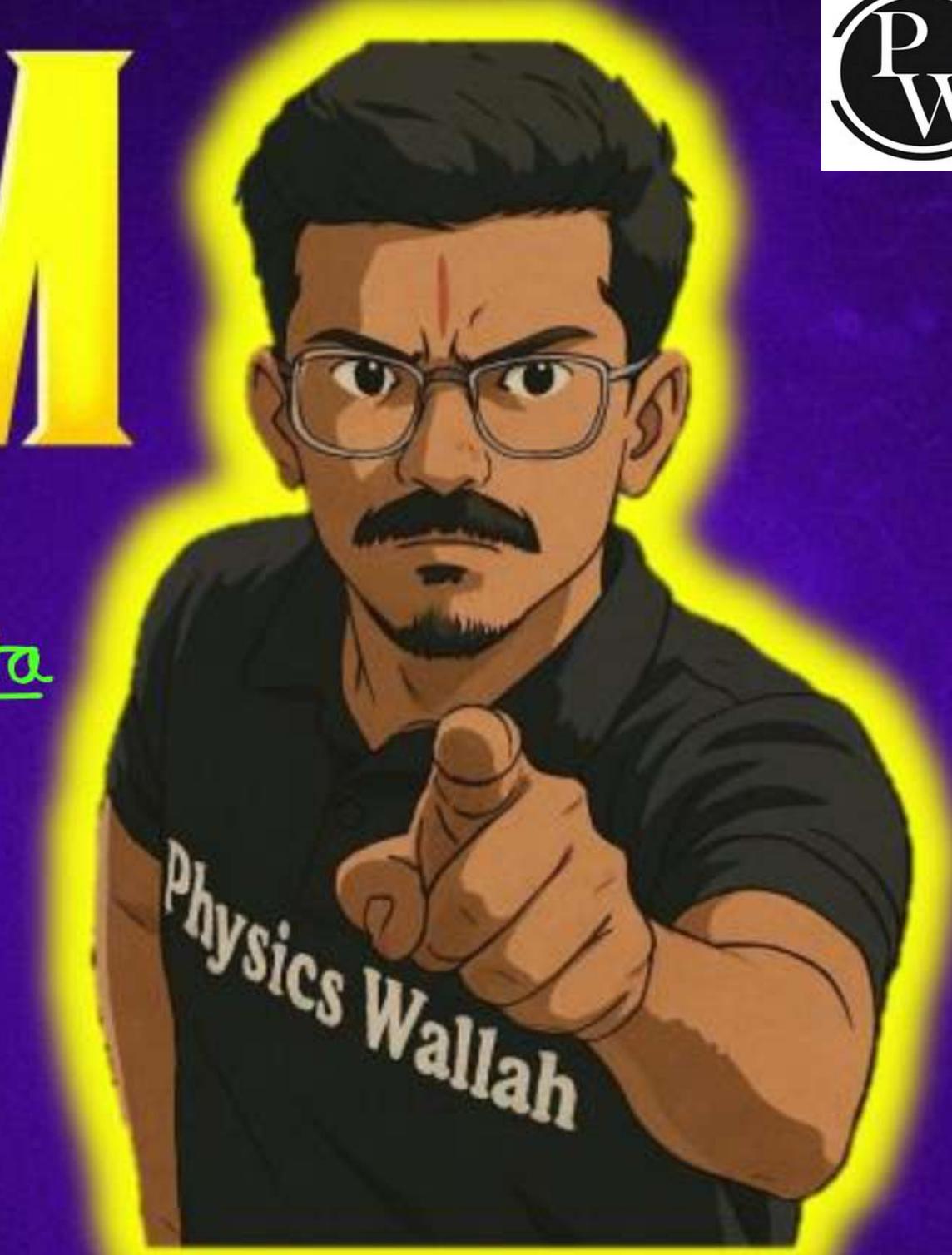
Lecture 04

ELECTROCHEMISTRY

6 Recap, Electrochemical Series
and Nernst Equation

Bharat Mata
Ki Jai ♡

Physics Wallah



BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

- (i) Easy Hinglish Recap (✓)
- (ii) Standard EMF or Electromotive Force (✓)
- (iii) Electrochemical Series and Interpretations From It (✓)
→ (NCERT Decoder)
- (iv) Nernst Equation (✓)



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EASY HINGLISH RECAP



EASY HINGLISH RECAP

- (i) **Electrode potential (E)**: Jab koi metal apne metal solution mein dip kiya jaata hai, toh metal electrode ki tendency ya toh electron lose krne ki hoti hai ya gain krne ki jiski vajah se uske aur uske solution ke beech opposite charges ki ek layer ban jaati hai jo voltage (potential difference) banta hai, usse electrode potential kehte hain.
- (ii) **Oxidation Potential**: Kisi bhi species ke electron lose karne ki tendency. Jitna zyada oxidation potential, utni zyada uski electrons dene ki power.
- (iii) **Reduction Potential**: Kisi bhi species ke electron gain karne ki tendency. Jitna zyada reduction potential, utni zyada uski electrons lene ki power.

EASY HINGLISH RECAP

✓ **(iv) Standard Electrode Potential (E°):** Electrode potential ko jab standard conditions (1 M electrolyte solution, 298 K temperature, 1 atm pressure) par measure kia jae with respect to standard hydrogen electrode. YAANI **potential of a single electrode relative to Standard Hydrogen Electrode (SHE)**.

IMPORTANT: We cannot measure the absolute electrode potential of a single half-cell. We can only measure the potential difference between two electrodes of two half cells. → one half cell → 'SHE'

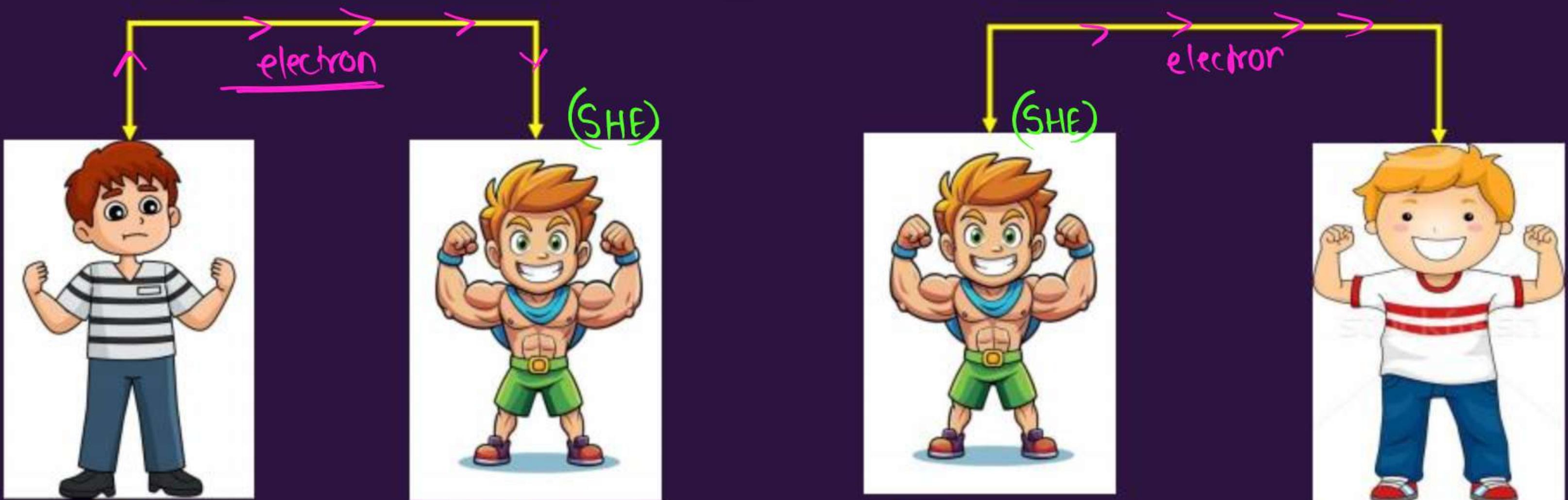
✓ **(v) Standard Oxidation Potential:** Kisi electrode ka oxidation potential hota hai jab usse 1 M electrolyte solution, 298 K temperature aur 1 atm pressure jaise standard conditions me measure kiya jaye, standard hydrogen electrode ke comparison me.

EASY HINGLISH RECAP

(vi) Standard Reduction Potential: Kisi electrode ka reduction potential hota hai jab usse 1 M electrolyte solution, 298 K temperature aur 1 atm pressure jaise standard conditions me measure kiya jaye, standard hydrogen electrode ke comparison me. According to IUPAC convention, standard reduction potentials are now called standard electrode potentials.

(vii) Standard Hydrogen Electrode (SHE): Isme platinum electrode hota hai, H_2 gas 1 atm par hoti hai aur 1 M H^+ solution mein dip kiya jaata hai. Yeh ek reference electrode hai jiska potential 0 V maana jaata hai aur isko reference maankar bakki metals ka standard electrode potential or standard reduction potential measure kia jaata hai.

ANALOGY FOR MEASUREMENT OF STANDARD ELECTRODE POTENTIAL



Zinc

$$E^\circ(Zn^{2+}/Zn) = -0.76V$$

Hydrogen

$$E^\circ(H^+/H_2) = 0V$$

Hydrogen

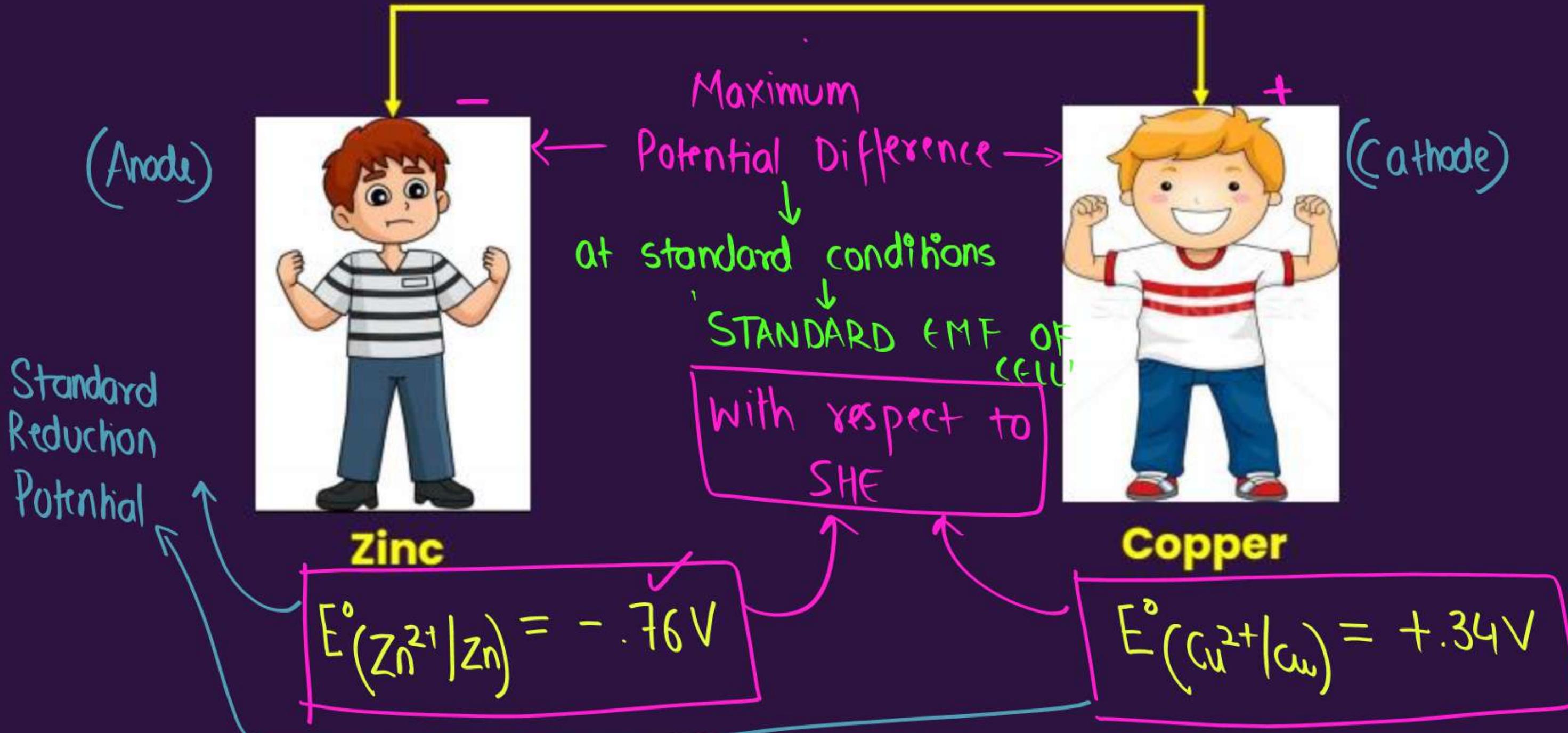
$$E^\circ(Cu^{2+}/Cu) = +0.34V$$

**SAMAJ AAYA TOH
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AYE BHAIYA**



STANDARD EMF OR CELL **ELECTROMOTIVE FORCE**

WHAT IF ZINC FIGHTS WITH COPPER?



STANDARD EMF OR CELL ELECTROMOTIVE FORCE

(E°_{cell})

34
76
1.10

- The standard EMF (electromotive force) of a cell is the maximum potential difference between the two electrodes of an electrochemical cell when no current is flowing through the circuit.
- The cells are at standard conditions (**25 °C, 1 M concentration, 1 atm pressure**)

$$\Rightarrow E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

[max^m possible potential difference $\rightarrow E^\circ_{\text{cell}}$]

ex: In a Daniel cell, $= +.34 - (-.76)$

$$E^\circ_{\text{cell}} = 1.10 \text{ V}$$

MORE TO KNOW

$$E^\circ_{\text{cell}} = E^\circ_{\text{O.P. (anode)}} + E^\circ_{\text{R.P. (cathode)}}$$

$$\begin{aligned} E^\circ_{\text{O.P. (anode)}} &= -E^\circ_{\text{R.P. (anode)}} \\ &= -E^\circ_{\text{R.P. (anode)}} + E^\circ_{\text{R.P. (cathode)}} \\ &= E^\circ_{\text{R.P. (cathode)}} - E^\circ_{\text{R.P. (anode)}} \end{aligned}$$

OR

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$
$$E^\circ_{\text{cell}} = E^\circ_{\text{right}} - E^\circ_{\text{left}}$$

aye
bhaiya

LET'S PRACTICE



QUESTION



In a Daniel cell (Zn-Cu cell), what will be the standard cell EMF if standard reduction potential of Zn²⁺ is -.76V while that of Cu²⁺ is +.34 V.

Standard Reduction Potential

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

$$= +.34 \text{ V} - (-.76 \text{ V}) \\ = +.34 \text{ V} + .76 \text{ V}$$

$$\boxed{E^{\circ}_{\text{cell}} = +1.10 \text{ V}}$$

$$E^{\circ}(\text{Zn}^{2+}/\text{Zn}) = -.76 \text{ V}$$

↓
ANODE'

$$E^{\circ}(\text{Cu}^{2+}/\text{Cu}) = +.34 \text{ V}$$

↓
Cathode

QUESTION

In a galvanic cell, what will be the standard cell EMF if standard reduction potential of Ni^{2+} is $-.25\text{V}$ while that of Ag^+ is $.80\text{ V}$.

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$= .80\text{ V} - (-.25\text{ V})$$

$$= .80\text{ V} + .25\text{ V}$$

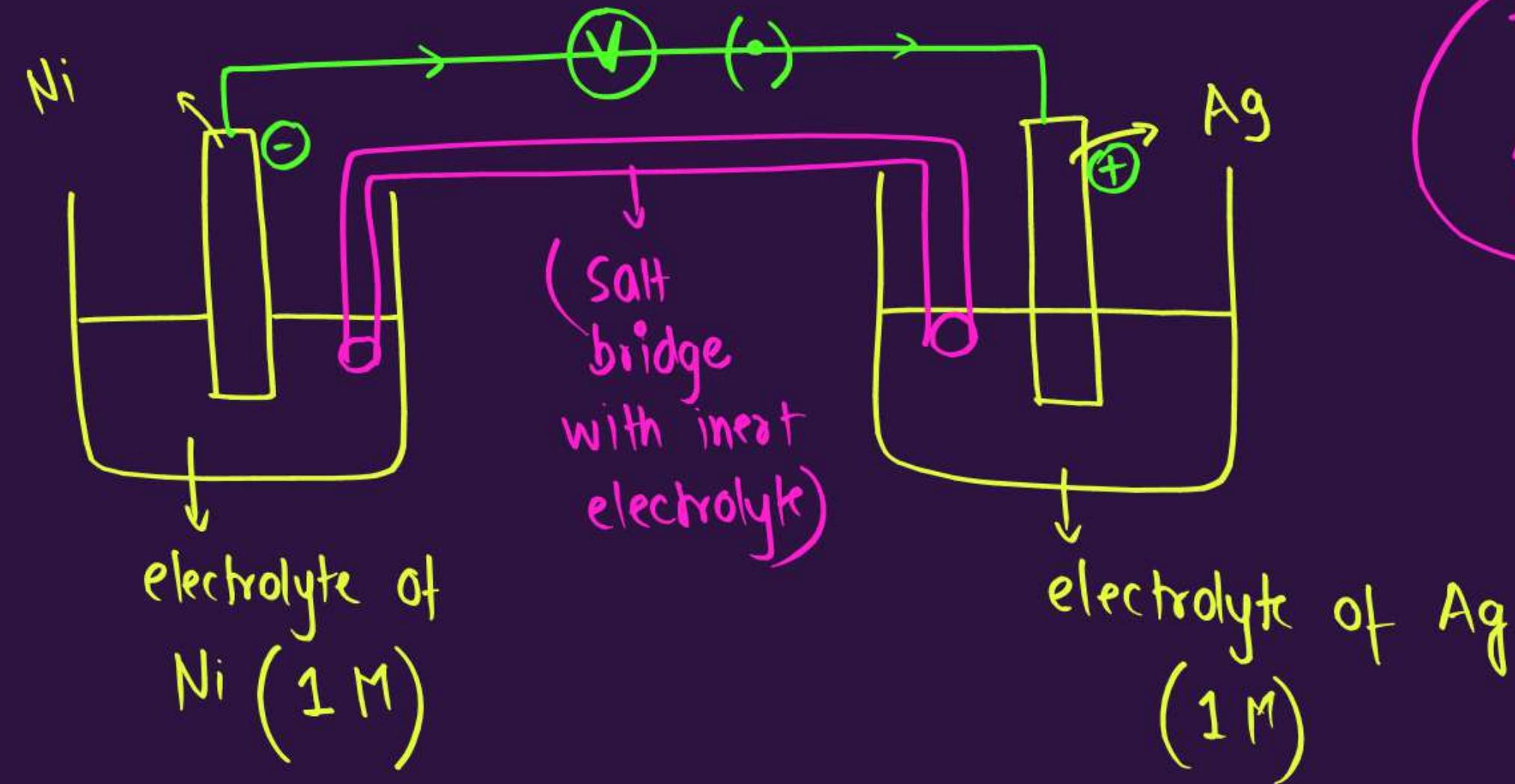
$$\boxed{E^\circ_{\text{cell}} = +1.05\text{ V}}$$

$$E^\circ(\text{Ni}^{2+}/\text{Ni}) = -.25\text{ V}$$

\downarrow
Anode

$$E^\circ(\text{Ag}^+/\text{Ag}) = +.80\text{ V}$$

\downarrow
Cathode

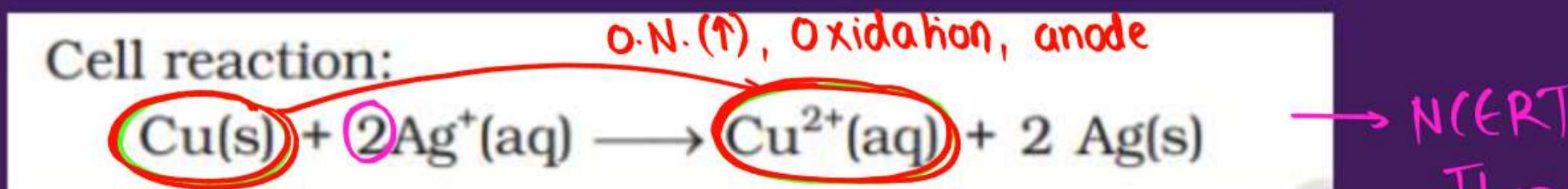


L O A N → Negative

Left ↙ Oxidation ↘ Anode

QUESTION

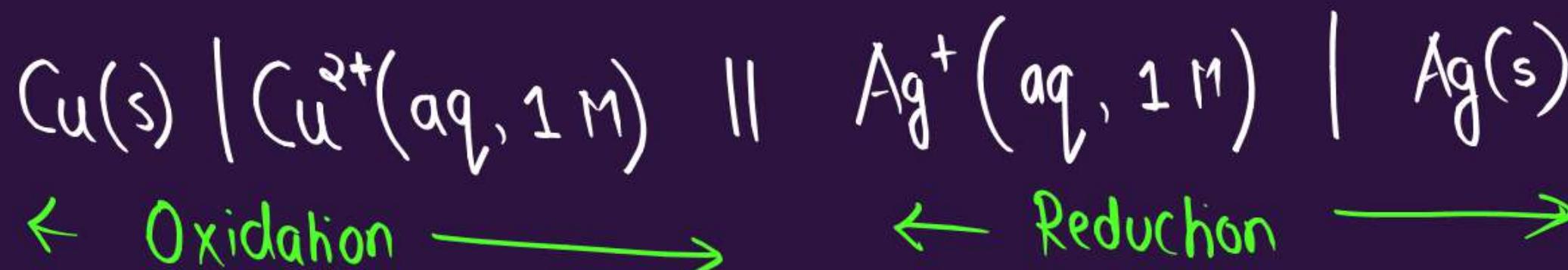
For the below cell reaction, write the cell notation and expression of standard EMF of the cell.



(Cell Notation / Cell Expression)

Anode | Anode electrolyte || Cathode electrolyte | Cathode

Theory

$$\begin{aligned} & \left[\text{Cu(s)} \longrightarrow \text{Cu}^{2+}(\text{aq}) + 2e^- \right] \rightarrow \text{'ANODE'} \\ & \left[2\text{Ag}^+(\text{aq}) + 2e^- \rightarrow 2 \text{Ag(s)} \right] \rightarrow \text{'CATHODE'} \end{aligned}$$


(Standard EMF of cell)

$$\hookrightarrow E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

Standard Reduction Potential

$$E^\circ_{\text{cell}} = E^\circ(\text{Ag}^+/\text{Ag}) - E^\circ(\text{Cu}^{2+}/\text{Cu})$$

NCERT CORNER

 (E_{cell}) (E°_{cell}) 

Cell potential is called cell electromotive force (emf) of the cell when no current is drawn through the cell.

- A. YES
- B. NO

CELL POTENTIAL

Feature	Standard EMF (E°_{cell})	Cell Potential (E_{cell})
Conditions	Standard (1 M, 1 atm, 25°C)	Non-standard / actual
Symbol	E°_{cell}	E_{cell}
Fixed or variable?	Fixed (for given cell)	Variable
Uses	Theoretical, comparison	Practical, real-time
Formula	$E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$	From Nernst equation

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AYE BHAIYA ✓



ELECTROCHEMICAL SERIES AND INTERPRETATIONS FROM IT

↳ (ncert decoder through questions)

Reaction (Oxidised form + ne ⁻)	→ Reduced form)
$F_2(g) + 2e^-$	$\rightarrow 2F^-$
$Co^{3+} + e^-$	$\rightarrow Co^{2+}$
$H_2O_2 + 2H^+ + 2e^-$	$\rightarrow 2H_2O$
$MnO_4^- + 8H^+ + 5e^-$	$\rightarrow Mn^{2+} + 4H_2O$
$Au^{3+} + 3e^-$	$\rightarrow Au(s)$
$Cl_2(g) + 2e^-$	$\rightarrow 2Cl^-$
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightarrow 2Cr^{3+} + 7H_2O$
$O_2(g) + 4H^+ + 4e^-$	$\rightarrow 2H_2O$
$MnO_2(s) + 4H^+ + 2e^-$	$\rightarrow Mn^{2+} + 2H_2O$
$Br_2 + 2e^-$	$\rightarrow 2Br^-$
$NO_3^- + 4H^+ + 3e^-$	$\rightarrow NO(g) + 2H_2O$
$2Hg^{2+} + 2e^-$	$\rightarrow Hg_2^{2+}$
$Ag^+ + e^-$	$\rightarrow Ag(s)$
$Fe^{3+} + e^-$	$\rightarrow Fe^{2+}$
$O_2(g) + 2H^+ + 2e^-$	$\rightarrow H_2O_2$
$I_2(s) + 2e^-$	$\rightarrow 2I^-$
$Cu^+ + e^-$	$\rightarrow Cu(s)$
$Cu^{2+} + 2e^-$	$\rightarrow Cu(s)$
$AgCl(s) + e^-$	$\rightarrow Ag(s) + Cl^-$
$AgBr(s) + e^-$	$\rightarrow Ag(s) + Br^-$
$2H^+ + 2e^-$	$\rightarrow H_2(g)$

↑ Reducing agent strength (↓)

	E° / V
+	2.87
+	1.81
+	1.78
+	1.51
+	1.40
+	1.36
+	1.33
+	1.23
+	1.23
+	1.09
+	0.97
+	0.92
+	0.80
+	0.77
+	0.68
+	0.54
+	0.52
+	0.34
+	0.22
+	0.10
	0.00

Standard Reduction Potential / Standard Electrode Potential

(i)

- Positive E° [High tendency to get reduced (electrons lena)]

(ii)

Reducing Agent: The one that reduces others (gain of electron on other species because it loses electrons).

- A positive E° means that the reduced form is a weaker reducing agent than H₂

NOTE: Standard electrode potential (E°) ki har value ko Standard Hydrogen Electrode (SHE) ke reference mein measure kia gaya hai.

Reaction (Oxidised form + ne ⁻)	→ Reduced form)		E° / V
$2\text{H}^+ + 2\text{e}^-$	→ H ₂ (g) ✓	Increase	0.00
Pb ²⁺ + 2e ⁻	→ Pb(s)		-0.13
Sn ²⁺ + 2e ⁻	→ Sn(s)		-0.14
Ni ²⁺ + 2e ⁻	→ Ni(s)		-0.25
Fe ²⁺ + 2e ⁻	→ Fe(s)		-0.44
Cr ³⁺ + 3e ⁻	→ Cr(s)		-0.74
Zn ²⁺ + 2e ⁻	→ Zn(s) ✓		-0.76
2H ₂ O + 2e ⁻	→ H ₂ (g) + 2OH ⁻	Strength of Reducing agent (↑)	-0.83
Al ³⁺ + 3e ⁻	→ Al(s)		-1.66
Mg ²⁺ + 2e ⁻	→ Mg(s)		-2.36
Na ⁺ + e ⁻	→ Na(s)		-2.71
Ca ²⁺ + 2e ⁻	→ Ca(s)		-2.87
K ⁺ + e ⁻	→ K(s)		-2.93
Li ⁺ + e ⁻	→ Li(s)		-3.05

Weakest oxidising agent

Strongest reducing agent

NOTE: Standard electrode potential (E°) ki har value ko Standard Hydrogen Electrode (SHE) ke reference mein measure kia gaya hai.

(Standard Reduction Potential)

C-I

- **Negative E°:** Low tendency to get reduced.

- Reducing Agent: The one that reduces others (gain of electron on other species because it loses electrons).

- A negative E° means that the reduced form is a stronger reducing agent than H₂.

NCERT CORNER (E°)

If the standard electrode potential of an electrode is greater than zero then its reduced form is more stable compared to hydrogen gas.

- A. YES
- B. NO

$(E^\circ > 0, E^\circ \text{ is +ve})$

NCERT CORNER

If the standard electrode potential is negative then hydrogen gas is more stable than the reduced form of the species.

- A. YES
- B. NO

\downarrow
 $(E^\circ < 0, E^\circ \text{ is -ve})$

NCERT CORNER

Fluorine gas (F_2) has the maximum tendency to get reduced to fluoride ions (F^-) and therefore fluorine gas is the strongest oxidising agent and fluoride ion is the weakest reducing agent.)

- A. YES
- B. NO



NCERT CORNER

(E° is -ve)

Lithium has the lowest electrode potential indicating that (lithium ion is the weakest oxidising agent) while lithium metal is the most powerful reducing agent in an aqueous solution.

- A. YES
B. NO



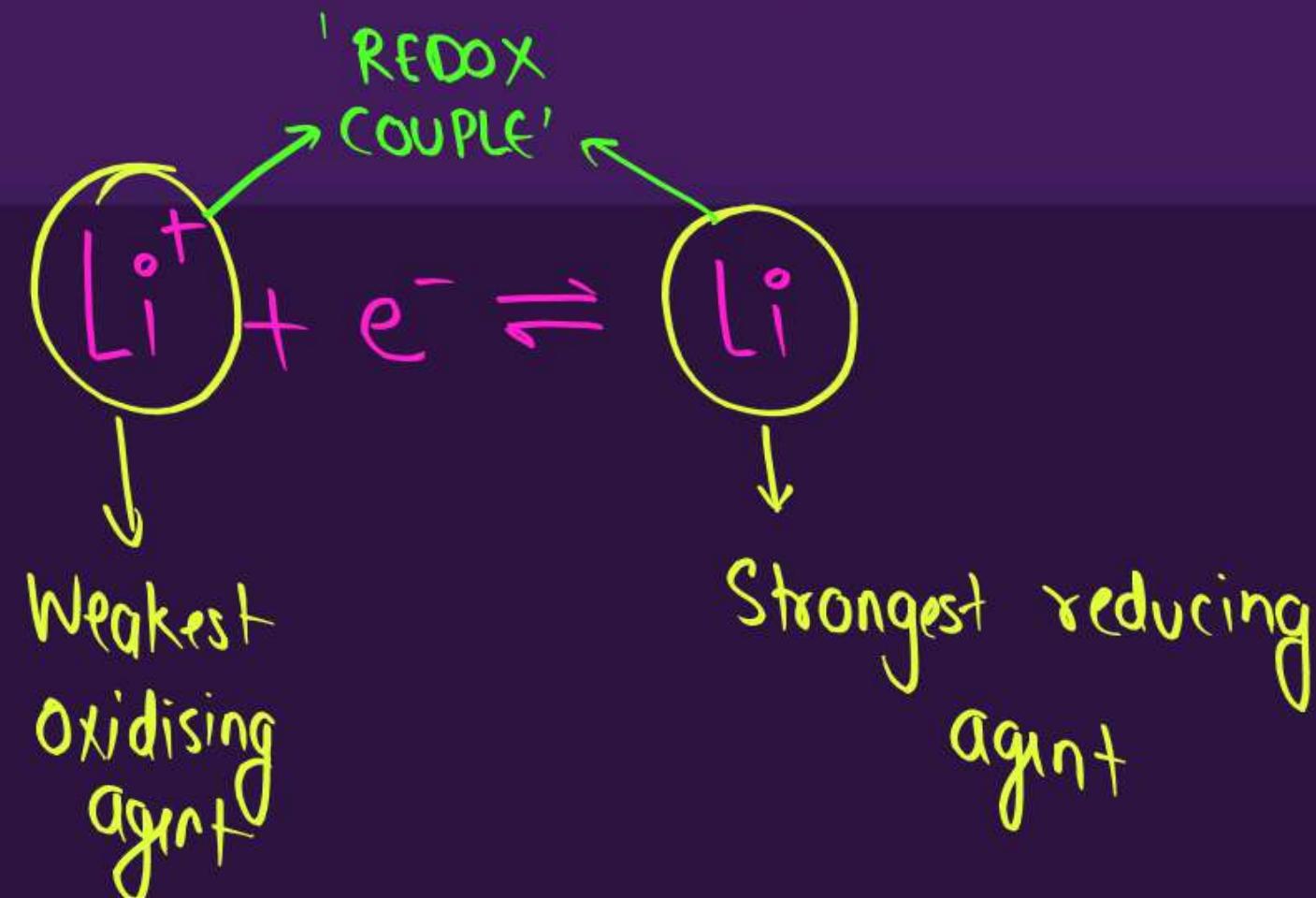
$$E^\circ(\text{Li}^+/\text{Li}) = \text{most - ve value}$$

GIVE A THOUGHT



In a redox couple, if one species is a strong oxidizing agent then the other one is a weak reducing agent and vice versa.

- A. YES
B. NO



**SAMAJ AAYA TOH
LIKH DO.**

AYE BHAIYA



✓ NEED: Nernst Equation helps calculate [cell EMF] at any non-standard condition of concentration/pressure/temperature $\rightarrow E_{cell}$

NERNST EQUATION

Nurse (x)

ΔG IN ELECTROCHEMICAL CELL

(change in Gibbs free energy)

From knowledge of thermodynamics let's understand ΔG in an electrochemical cell,

⇒ Total energy released (ΔH) = Part of energy used to move electrons or electrical work (ΔG) + Part goes as random motion ($T\Delta S$)

$$\Delta H = [\text{Electrical Work } (\Delta G)] + [\text{Heat Loss } (T\Delta S)]$$

$$\Delta G = \Delta H - T\Delta S$$

$$\Delta G = -ve$$

Spontaneous Rxn → Happens in electrochemical cell

ΔG IN ELECTROCHEMICAL CELL

Electrical work (ΔG)

✓(i) Work done (W) = $q \times V$ -①

Total charge transferred

✓(i) Charge of one electron:

$$e = 1.602 \times 10^{-19} \text{ Coulombs}$$

Charge of 1 mole of electrons:

$$= (6.022 \times 10^{23}) \times (1.602 \times 10^{-19}) = 96,485 \text{ C/mol} \approx 96500 \text{ C/mol}$$

(Also called Faraday's constant) (F)

If 'n' moles of electron transfers in a reaction then:

$$q = nF$$

1 mole $e^- = F$
 'n' mole $e^- = 'x'$

$$'x' = nF$$

ΔG IN ELECTROCHEMICAL CELL

(iii) Now, Work done (W) = $q \times V$ = $nF \times E_{\text{cell}}$

But for spontaneous reactions ΔG is -ve

(iv) Hence,

$$\boxed{\Delta G = -nFE_{\text{cell}}}$$

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LIKH DO.
AYE BHAIYA**



DERIVATION OF NERNST EQUATION

(Class "Thermodynamics")

- **STEP I:** The relationship between the Gibbs free energy change (ΔG) and the reaction quotient (Q) at any point in the reaction is:

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



$$Q = \frac{[C]^c [D]^d}{[A]^a [B]^b} \rightarrow \text{conc. of products}$$

$\rightarrow \text{conc. of reactants}$

Where:

- ΔG = Change in Gibbs Free Energy at non-standard conditions
- ΔG° = Change in Gibbs Free Energy at standard conditions
- R = Gas constant (8.314 J/mol·K)
- T = Temperature in Kelvin
- Q = Reaction quotient

The reaction quotient (denoted as Q) is a ratio of the concentrations (or partial pressures) of the products and reactants at any point in time during a chemical reaction.

Predicts the direction of
to attain equilibrium

DERIVATION OF NERNST EQUATION

- ✓ **STEP II:** Link ΔG to Electrochemical Work:

$$\Delta G = -nFE_{\text{cell}}$$

and

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ$$

Substituting these values in equation 1 we get:

$$-nFE_{\text{cell}} = -nFE_{\text{cell}}^\circ + RT \ln Q$$

- ✓ **STEP III:** Divide the whole equation by $-nF$

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{RT}{nF} \ln Q \quad \text{---(1)}$$

DERIVATION OF NERNST EQUATION

- ✓ **STEP IV:** Convert to Base 10 Log

$$\ln Q = 2.303 \log_{10} Q$$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{2.303 RT}{nF} \log Q \quad -②$$

- ✓ **STEP V:** Plug in constants at T = 298 K, R = 8.314 J/mol·K, F = 96500 C/mol

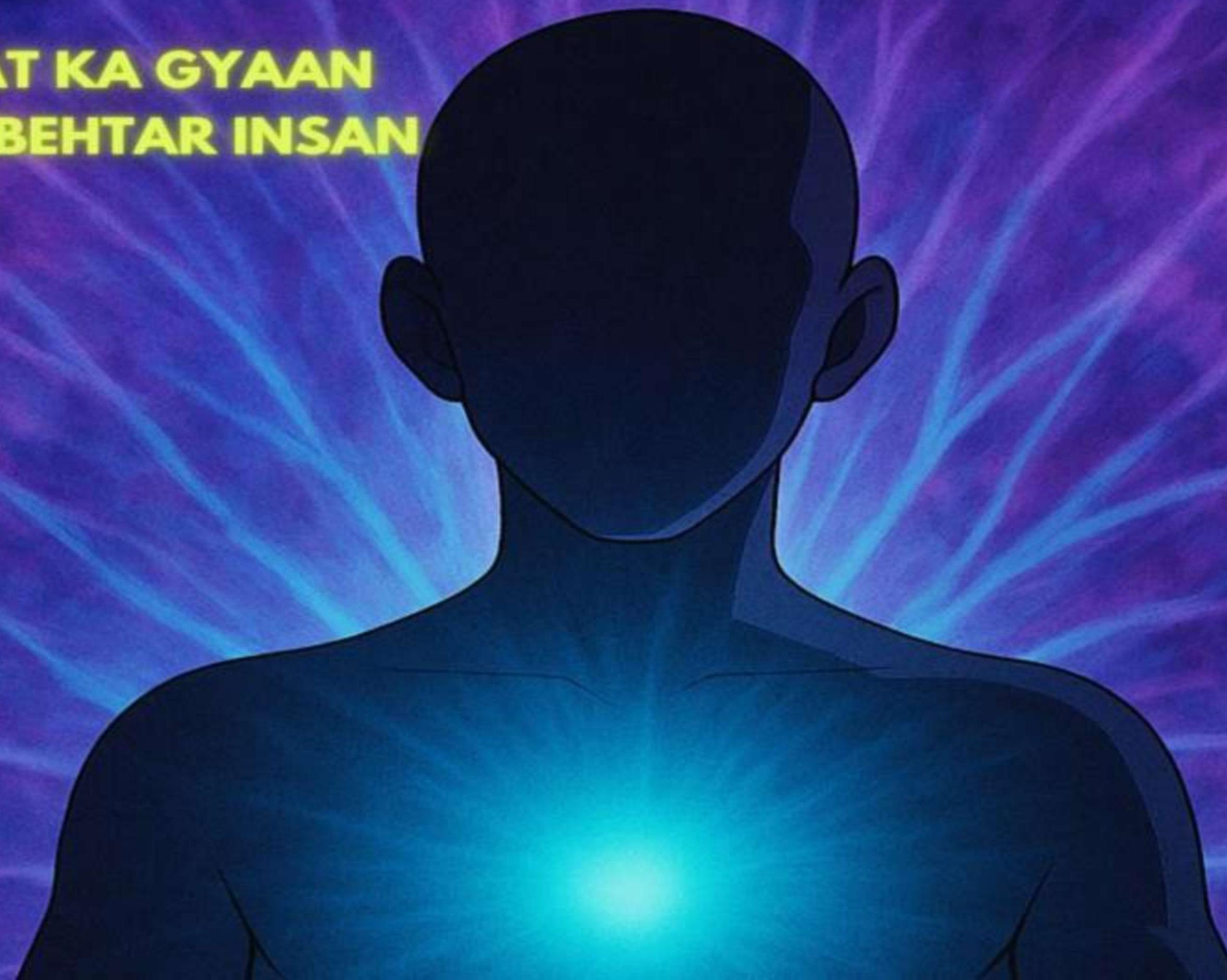
MOST USED

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059 \log Q}{n} \quad -③$$

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CONCEPT POLISH – HOMEWORK

↓
[NO H.W.]





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#sbsathhai (✓)

#pwsathhai (✓)

Thank
You

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2026

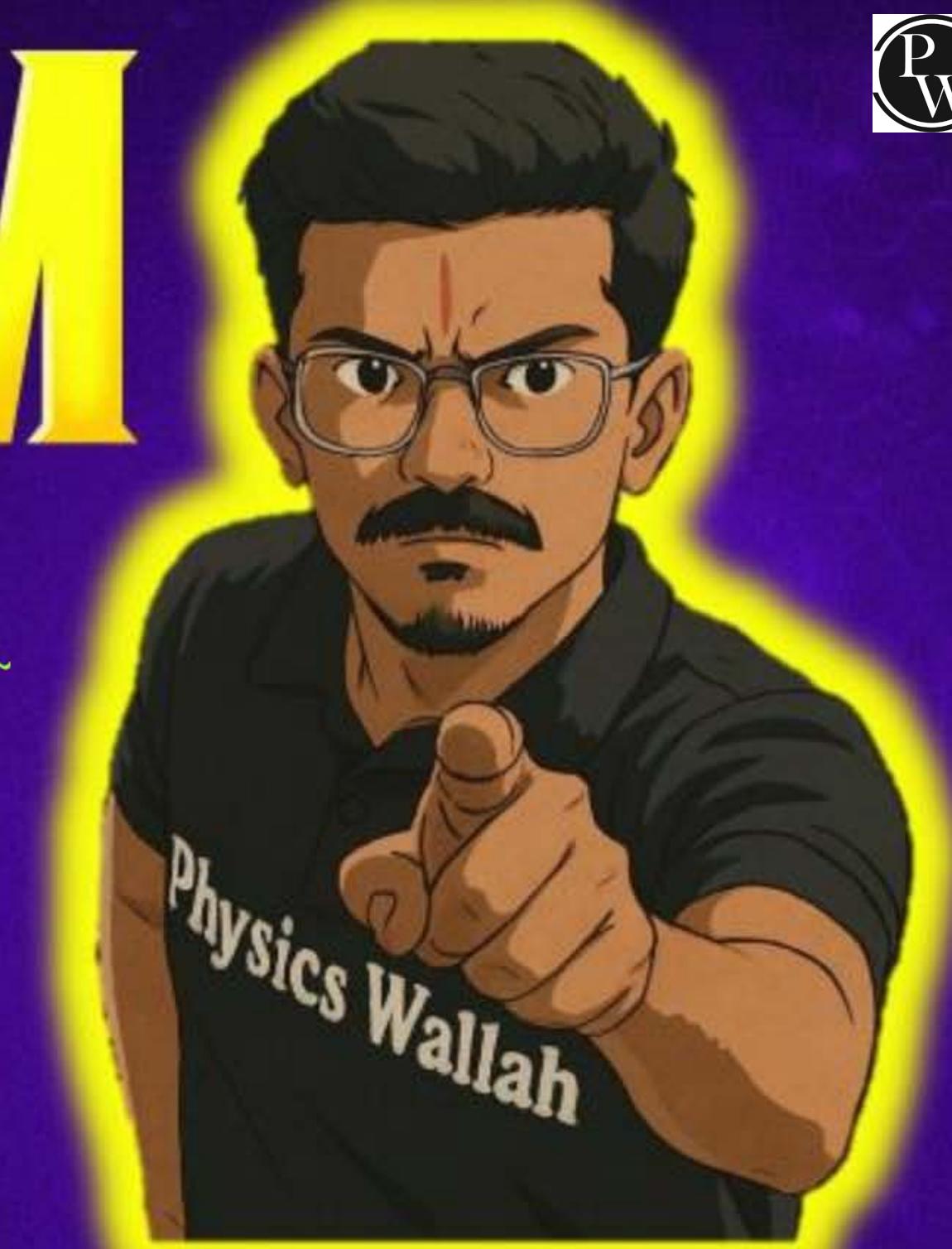
CHEMISTRY

Lecture 05

ELECTROCHEMISTRY

Numerical Practice and Conductance
of Electrolytic Solutions

Bharat Mata
Ki Jai ♡



BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

- (i) Numerical Practice (✓)
- (ii) Conductance of Electrolytic Solutions (✓)

Nernst eqⁿ





NEWS WALLAH

After completion of this chapter, Organic and Physical Chemistry will run parallelly.
We consider your P.O.V.

NUMERICAL PRACTICE

NCERT SOLVED EXAMPLE 2.3

The standard electrode potential for Daniell cell is 1.1V. Calculate the standard Gibbs energy for the reaction:



$$\Delta G^\circ = - nFE^\circ_{\text{cell}}$$

$$F \rightarrow 96,500 \frac{\text{C}}{\text{mol}}$$

$$n = 2$$

$$E^\circ_{\text{cell}} = 1.1 \text{ V}$$

$$= - 2 \times 96500 \times \frac{1.1}{100}$$

$$193000$$

$$= - 9650 (20 + 2)$$

$$\frac{19300}{212300}$$

$$= - [193000 + 19300]$$

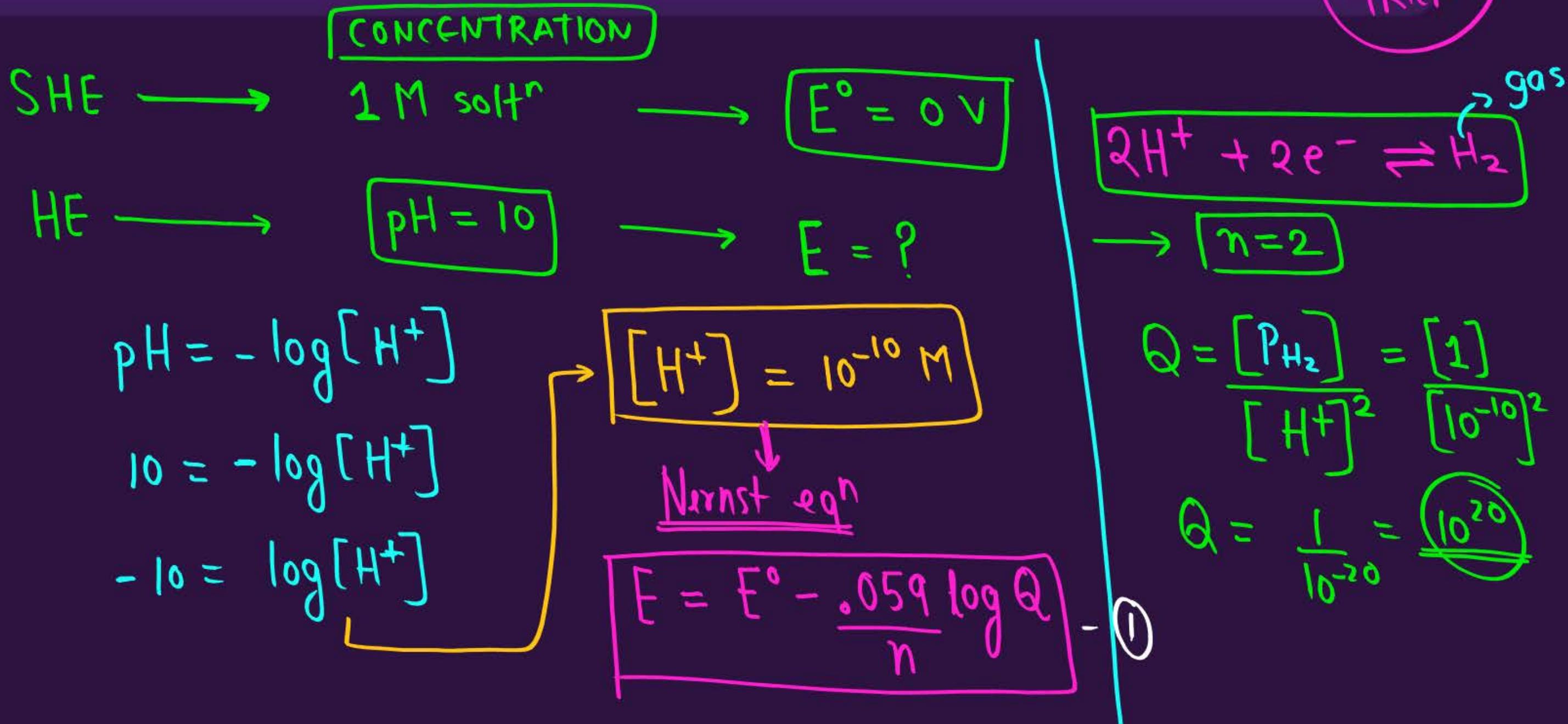
$$= - 212300 \frac{\text{C} \times \text{V}}{\text{mol}} \xrightarrow{-212,300 \frac{\text{J}}{\text{mol}}} \boxed{-212.3 \frac{\text{kJ}}{\text{mol}}}$$

NCERT INTEXT 2.4

$$\log K = x \downarrow \\ K = 10^x$$

Calculate the potential of hydrogen electrode in contact with a solution whose pH is 10.

→ SHORT TRICK



Put values in ①

$$E = 0 - \frac{.059}{2} \log_{10}(10^{20})$$

$$= 0 - \frac{.059}{2} \times 20 \cancel{\log_{10}}^1$$

$$= 0 - \frac{.059 \times 20}{2}$$

$$= -.059 \times 10$$

$$E = -.59 \text{ V}$$

$$\log_{10} a^b = b \log_{10} a$$

~~$\log_{10} 10 = 1$~~

SHORT-TRICK

pH → Only value

Calculate E

$$E = -.059 \times \text{pH}$$

QUESTION

Calculate the potential of hydrogen electrode in contact with a solution whose pH = 3

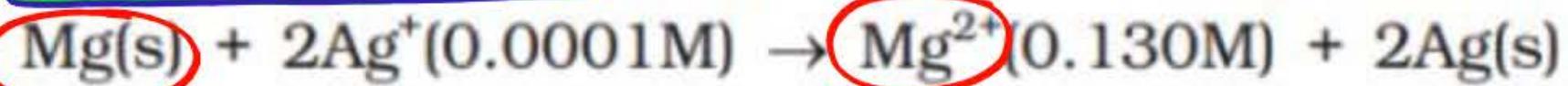
$$E = -0.059 \times \text{pH}$$

(for hydrogen electrode)

$$E = -0.059 \times 3 \text{ V}$$

NCERT SOLVED EXAMPLE 2.1

Represent the cell in which the following reaction takes place



Calculate its E_{cell} if $E_{\text{cell}}^{\circ} = 3.17 \text{ V}$.

Cell Notation | Cell Representation

Anode | Anode electrolyte || Cathode electrolyte | Cathode
 ← OXIDATION → ← REDUCTION →



Nernst eqn

$$\{_{\text{cell}} = \{^{\circ}_{\text{cell}} - \frac{.059}{n} \log Q$$

$$= 3.17 - \frac{.059}{2} \log \frac{[\text{Mg}^{2+}][\text{Ag}]^2}{[\text{Mg}] [\text{Ag}^+]^2}$$

$$= 3.17 - \frac{.059}{2} \log \frac{(.130)}{(.0001)^2}$$

$$= 3.17 - \frac{.059}{2} \times 7$$

$$= 3.17 - (.0295 \times 7)$$

$$= 3.17 - .2065 = 2.9635 \checkmark$$

~~pure solid
pure liquid \rightarrow unity
concentration~~

$$\frac{(130 \times 10^8)^5}{(00000001 \times 10^6)}$$

$$\Rightarrow 130 \times 10^5$$

$$\Rightarrow 13 \times 10^6$$

$$\Rightarrow 1.3 \times 10^7$$

$$\Rightarrow \log(1.3 \times 10^7)$$

$$\Rightarrow \underline{\log(1.3)} + \log 10^7$$

$$\approx 0 + 7 \log_{10} 10$$

$$\log_{10} 1 = 0$$

EQUILIBRIUM CONSTANT FROM NERNST EQUATION

- At equilibrium, [cell stops working and so, $E_{\text{cell}} = 0$]
- Also, Q (Reaction Quotient) = K (Equilibrium Constant)
- So, Nernst Equation will be:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.059}{n} \log Q$$

at equilibrium,

$$0 = E^{\circ}_{\text{cell}} - \frac{0.059}{n} \log K_c$$

$$E^{\circ}_{\text{cell}} = \frac{0.059}{n} \log K_c \Rightarrow$$

$$\frac{E^{\circ}_{\text{cell}} \times n}{0.059} = \log K_c$$

K_c
↓
equilibrium constant
in terms of concentration

$$\frac{E_{\text{cell}}^{\circ} \times n}{0.059} = \log K_c \quad \textcircled{1}$$

$$x = \log K$$
$$10^x = K$$

$$10^{\frac{E_{\text{cell}}^{\circ} \times n}{0.059}} = K_c \quad \textcircled{2}$$

① & ②

both are imp.

NCERT SOLVED EXAMPLE 2.2

Calculate the equilibrium constant of the reaction:



$$E_{\text{cell}}^{\circ} = 0.46 \text{ V}$$

$$10^{\frac{E_{\text{cell}}^{\circ} \times n}{0.059}} = K_c$$

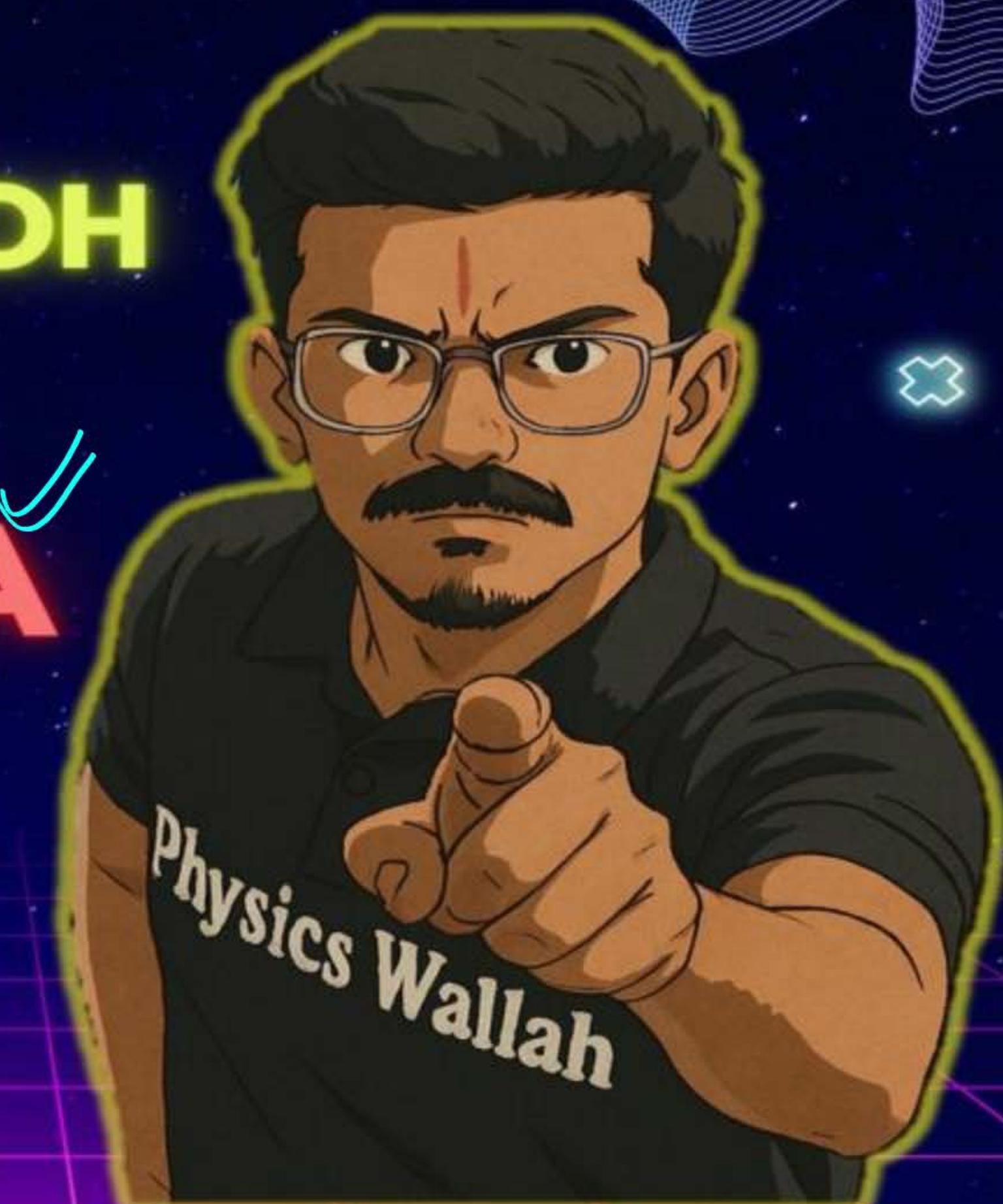
$$\frac{10^{\frac{(0.46 \times 2)}{0.059}}}{10^{\frac{(0.92 \times 1)}{0.059}}} = K_c$$

$$\frac{10^{\frac{0.920}{0.059}}}{10^{\frac{0.920}{0.059}}} = K_c$$

$$10^{\frac{0.920}{0.059}} = K_c$$

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LIKH DO.**

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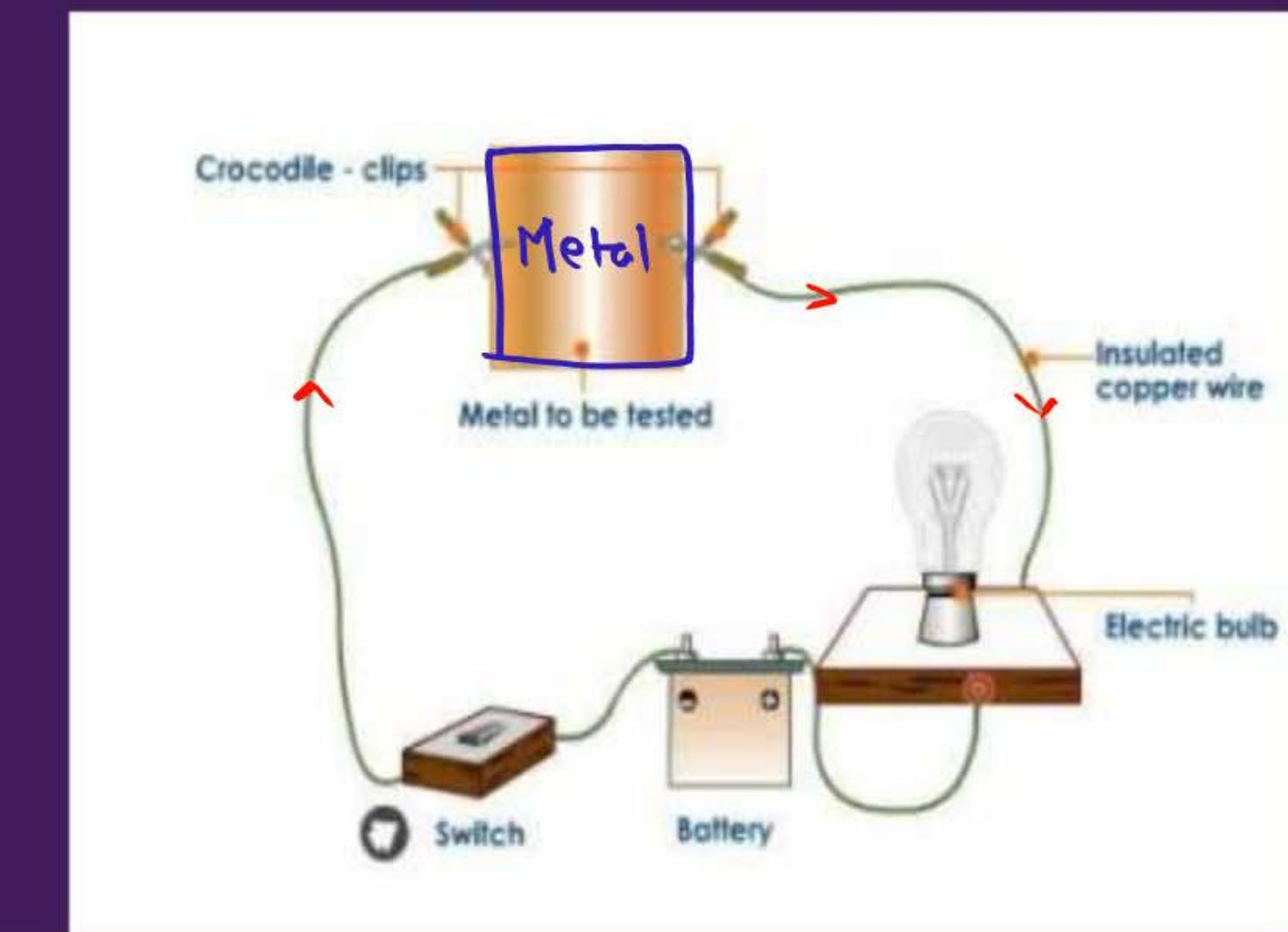
✓

CONDUCTANCE OF ELECTROLYTIC SOLUTIONS

GIVE A THOUGHT

In the below example the flow of current is due to the 'movement of free electrons'. It is an example of metallic conductor.

- A. True
- B. False

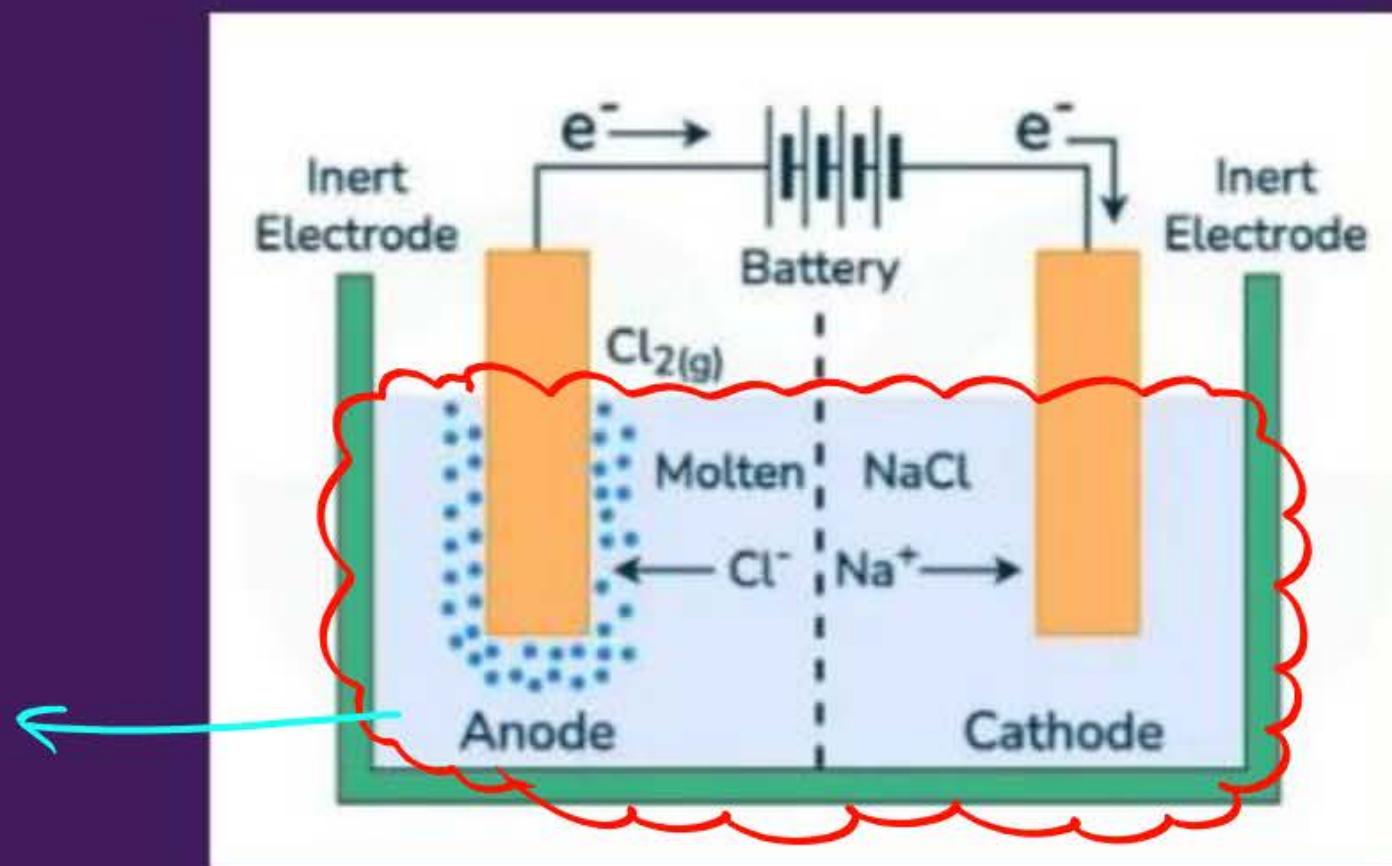


GIVE A THOUGHT

In the below example the flow of current is due to the movement of ions. It is not an example of metallic conductor.

- A. True
- B. False

Conducts electricity
Electrolytic solution

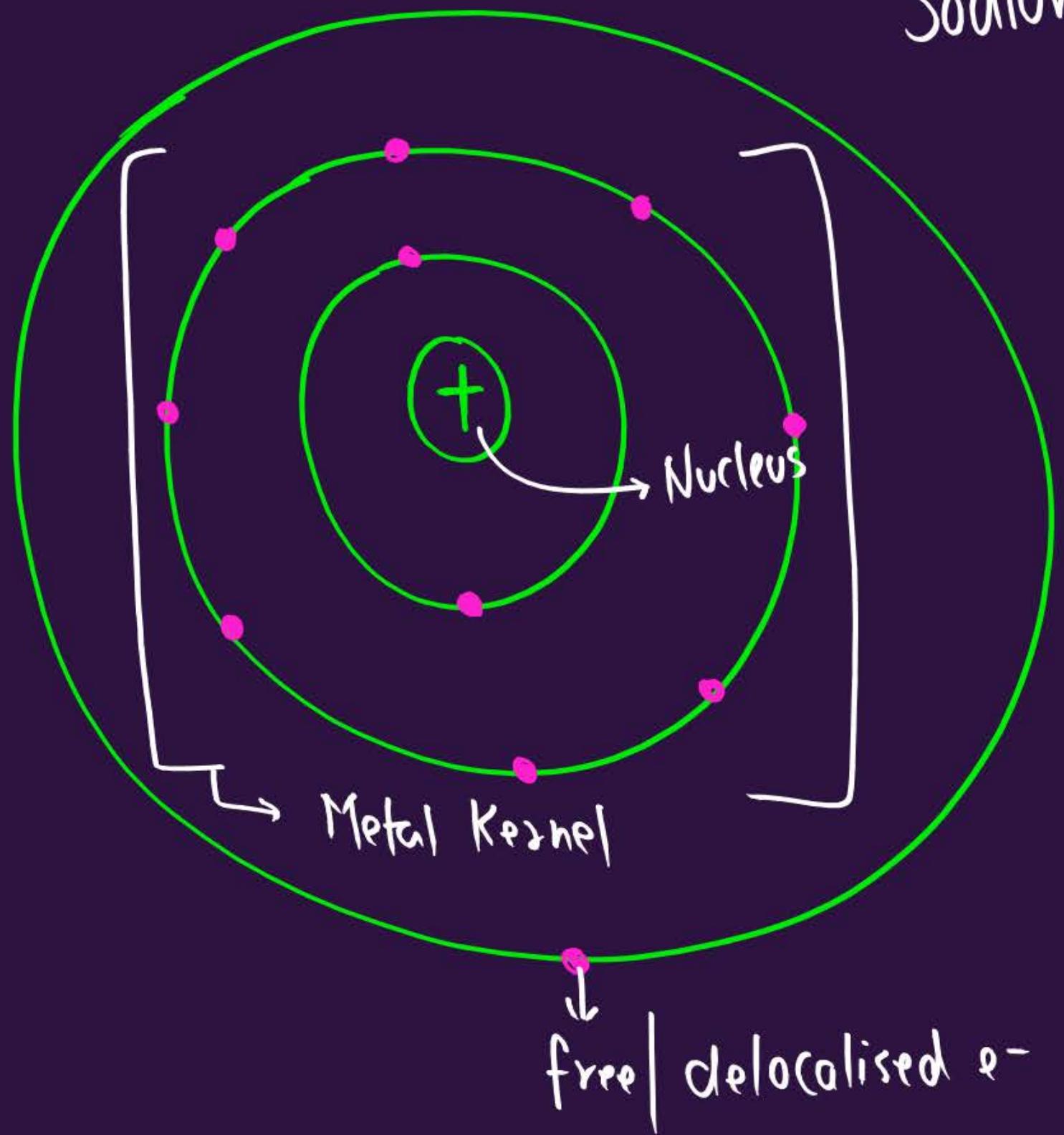


METALLIC VS ELECTROLYTIC CONDUCTOR

Feature / Basis	Metallic Conductor ✓	Electrolytic Conductor ✓	Anion ↑
(i) Current flows due to	Free electrons Delocalised electrons	Movement of positive and negative ions	Cation ↑
(ii) Involves chemical decomposition?	No chemical change	Yes, electrolysis occurs (chemical decomposition)	
(iii) Reason for resistance	Collision of free electrons with vibrating metal kernels	<p>Viscosity ↑</p> <p>flow of ions ↓</p> <ul style="list-style-type: none"> → Size of ion → Viscosity of medium → Interionic attraction → Solvation (hydration) <p>larger the size of ion movement will be slower</p>	
(iv) Effect of temperature on resistance	 Increases with temperature (more lattice vibrations)	Decreases with temperature (ion mobility increases)	

Sodium (Na) →

K	L	M
2	8	1



CONCEPT CLARITY



- In metallic conductors:

Temperature increase → Metal ions (kernels) vibrate more → More collisions with electrons → More opposition to flow → More resistance

- In electrolytic conductors:

Temperature increase → Viscosity of solution decreases → Ions move more freely → Less opposition to flow → Resistance decreases

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IMPORTANT TERMS TO KNOW

(i) **Resistance (R)**: It is the opposition offered by a material to the flow of electric current. Its S.I. unit is ohm (Ω).

The electrical resistance of any object is directly proportional to its length, l, and inversely proportional to its area of cross section, A.

$$R \propto l$$

$$R \propto \frac{l}{A}$$

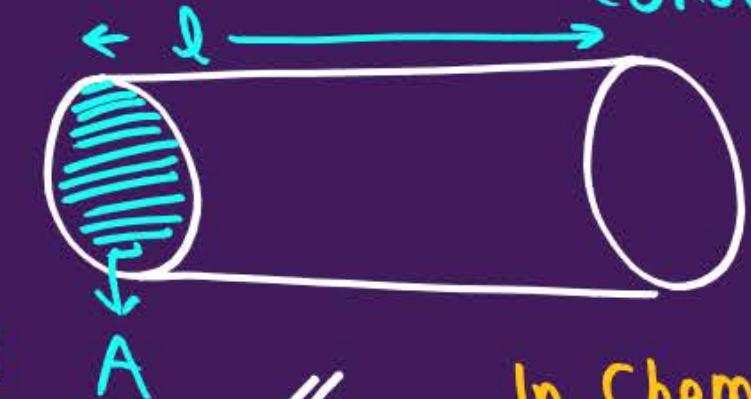
$$R \propto \frac{l}{A}$$

$$R = \rho \frac{l}{A}$$

$$\rho (\text{rho})$$

(resistivity) or
specific resistance

$$\rho = \frac{R \times A}{l} = \frac{\text{ohm} \times \text{m}^2}{\text{m}} = \boxed{\text{Ohm} \cdot \text{m}}$$



In Chemistry,
 $\boxed{\text{Ohm} \cdot \text{cm}}$
most often

IMPORTANT TERMS TO KNOW

Resistance and Resistivity in Terms of Electrolytic Conductor

$$R \propto l$$

$$R \propto \frac{l}{A}$$

$$R \propto \frac{l}{A}$$

$$R = \rho \frac{l}{A}$$

$$\rho = \frac{R \times A}{l}$$

Volume of electrolyte

$$= A \times L$$

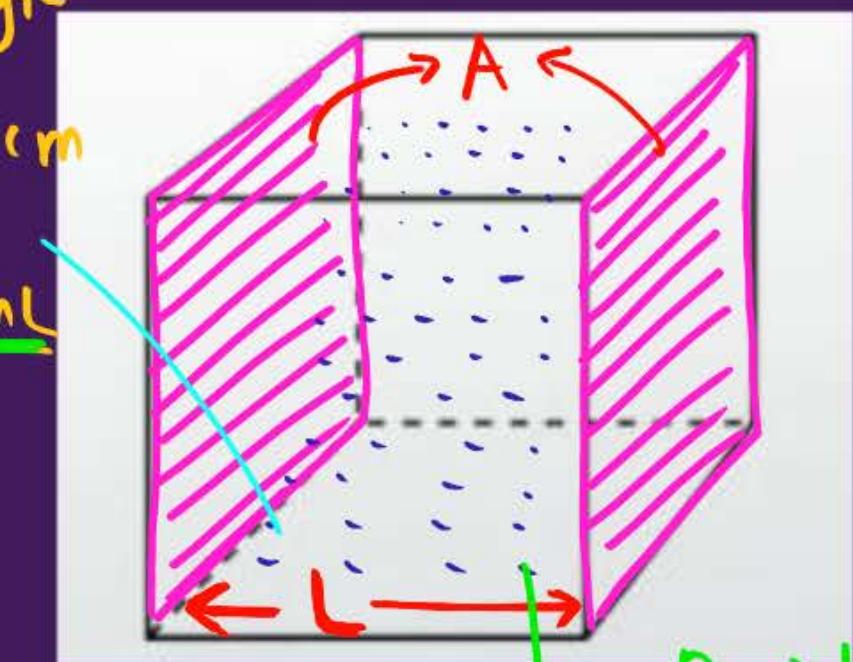
if $A = 1 \text{ cm}^2$ & $L = 1 \text{ cm}$

$$V = 1 \text{ cm}^3 \text{ or}$$

$$A = 1 \text{ cm}^2$$

$$L = 1 \text{ cm}$$

$$\rho = R$$



Resistance of
electrolytic
soln

Resistivity is the resistance offered to all the ions of that
electrolyte present in 1 mL or 1 cm³ of solution.

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IMPORTANT TERMS TO KNOW

(ii) **Conductance (G)**: It is the ease with which electric current flows through a conductor. It is reciprocal of resistance.

$$G = \frac{1}{R} = \frac{1}{\Omega} = \text{ohm}^{-1}$$

Its S.I. unit is **Siemens (S)** or **ohm⁻¹** (also known as mho) or **(Ω⁻¹)**

$$G = \frac{1}{R} = \frac{A}{\rho L} = \frac{1}{\rho} \cdot \frac{A}{L} = K \cdot \frac{A}{L}$$

Conductivity or specific conductance (k) =

$$\left[\frac{G \times L}{A} \right] = \left[\frac{1}{R} \cdot \frac{1}{A} \right]$$

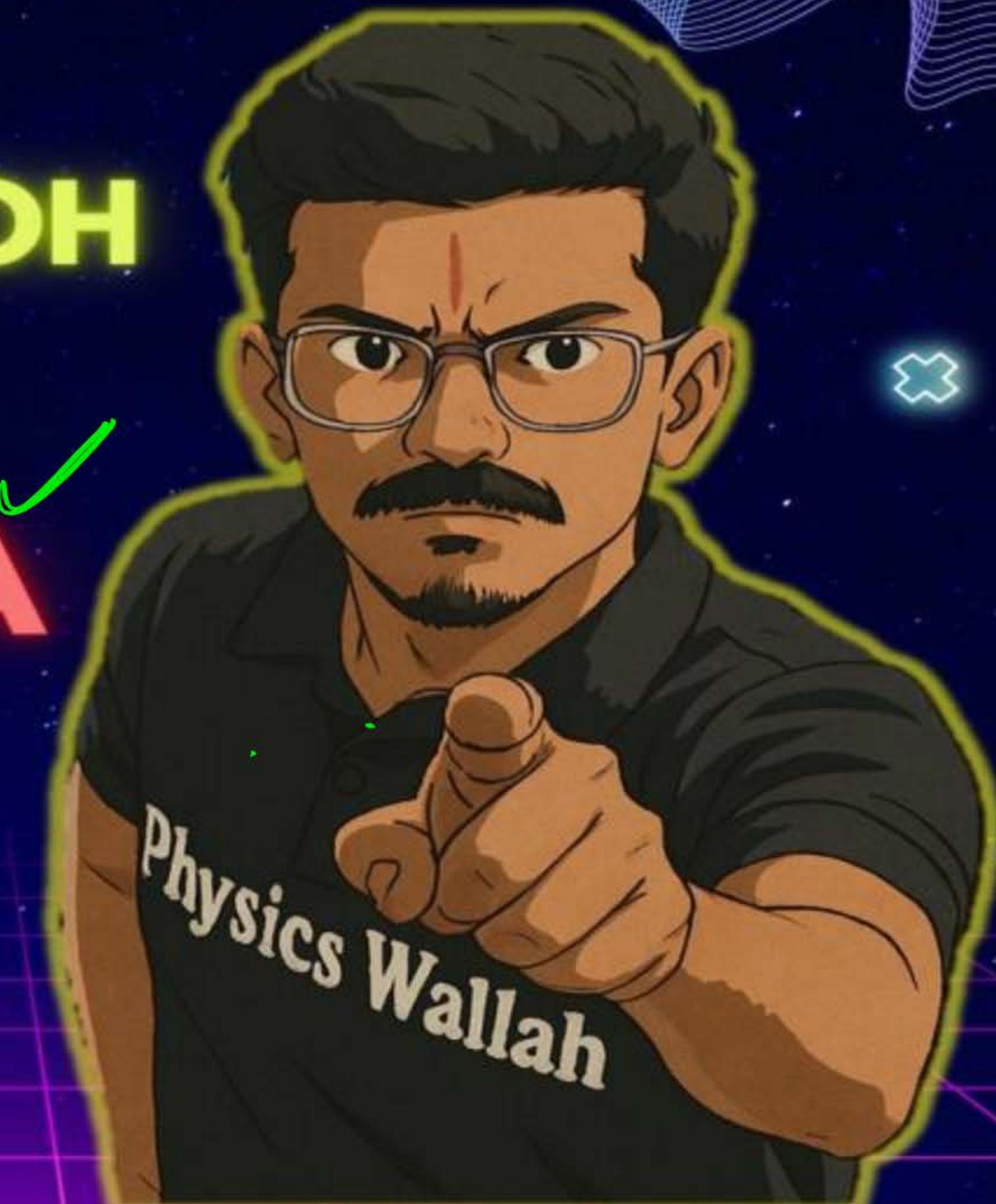
$K = \frac{1}{\rho}$ *

K (kappa)

Conductivity is the conductance due to all the ions of that electrolyte present in 1 mL or 1 cm³ of solution.

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CONCEPT POLISH – HOMEWORK

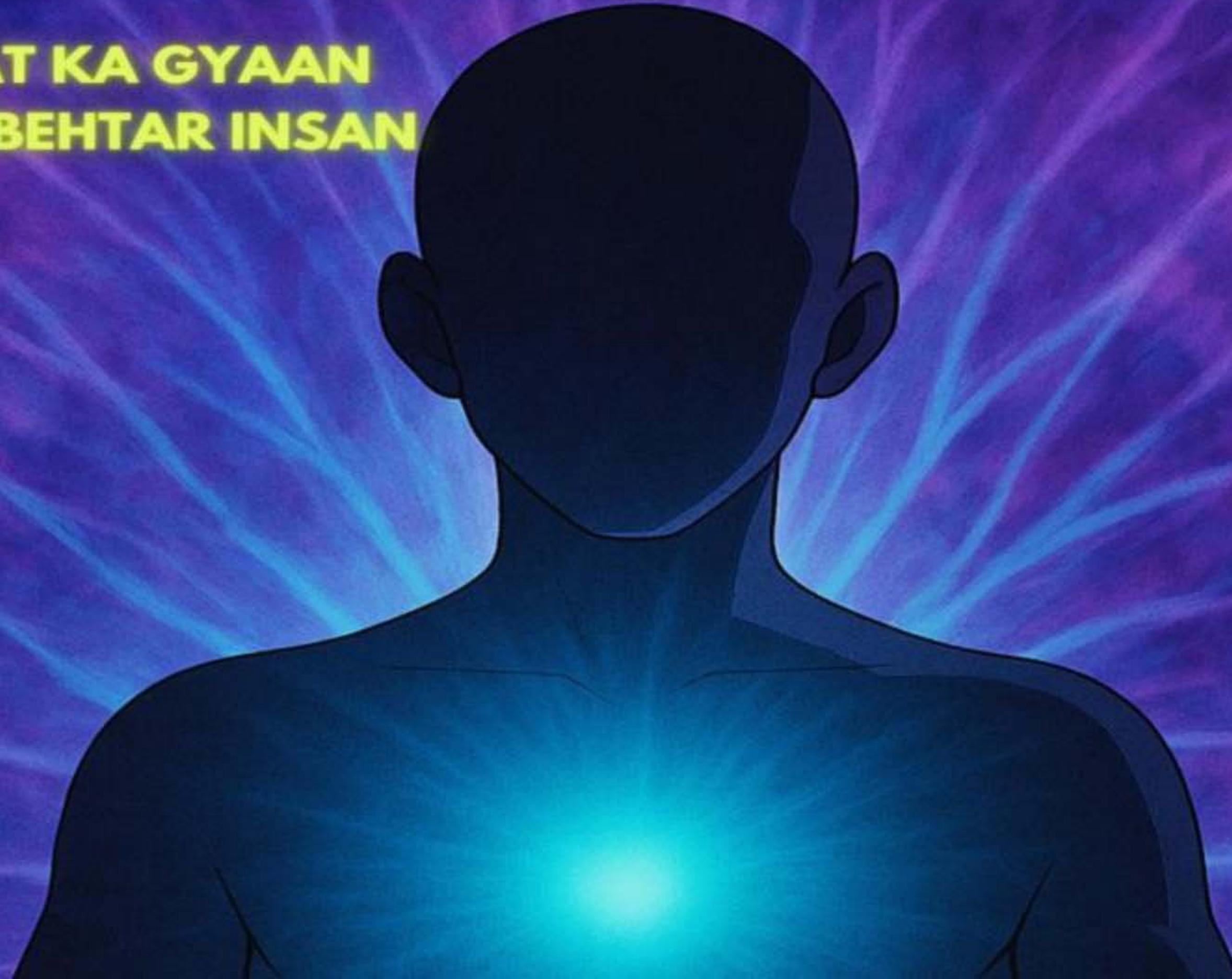


GIVE A THOUGHT

The S.I. unit of specific conductance or conductivity is:

- A. $\text{ohm}^{-1} \cdot \text{m}^{-1}$
- B. $\text{S} \cdot \text{m}^{-1}$
- C. Both A and B

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#sbsathhai(✓)

#pwsathhai (✓)

**Thank
You**



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Chemistry

Basic Maths
'Log' ✓

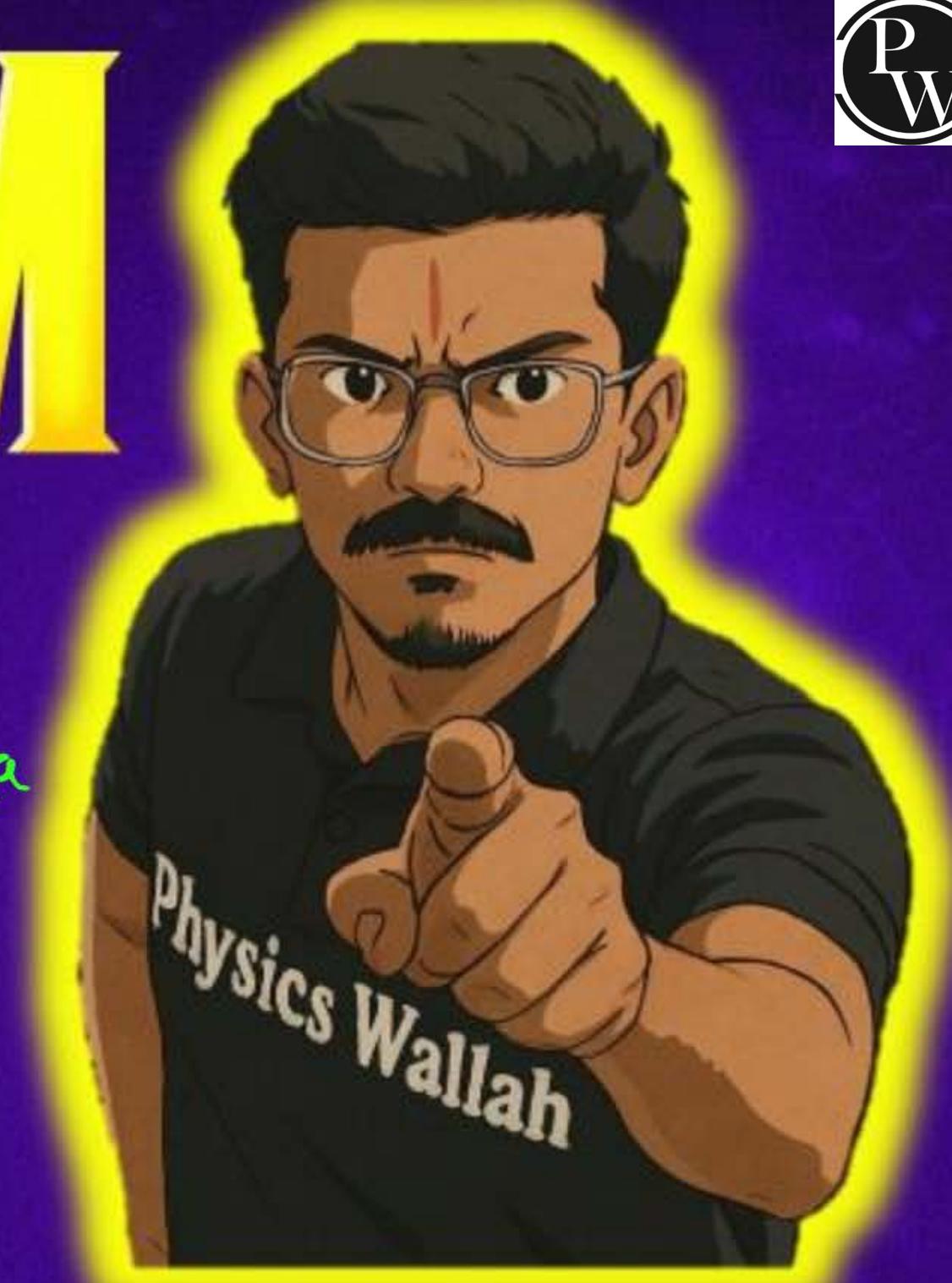
Lecture 06

ELECTROCHEMISTRY

Measurement of Conductivity and
Molar Conductivity

Bharat Mata
Ki Jai ❤

Physics Wallah



BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

- (i) Basic Maths - Logarithm (✓)
- (ii) Measurement of Conductivity of Electrolytic Solutions and Molar Conductivity (✓)



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' Gaana Sunenge Aur

Revise Karenge '



Mood fresh Hoga

CONCEPT POLISH

- HOMEWORK

DISCUSSION



GIVE A THOUGHT

The S.I. unit of specific conductance or 'conductivity' is:

- A. $\text{ohm}^{-1} \cdot \text{m}^{-1}$ (✓)
- B. $\text{S} \cdot \text{m}^{-1}$ (✓)
- C. Both A and B

$$K = \frac{1}{\rho} = \frac{1}{\text{Ohm} \cdot \text{m}} = \boxed{\text{ohm}^{-1} \cdot \text{m}^{-1}}$$

\downarrow

$$\boxed{\text{S} \cdot \text{m}^{-1}}$$

$$R = \rho \frac{l}{A}$$

$$\frac{R \times A}{l} = \rho$$

$$\frac{\text{ohm} \times \text{m}^2}{\text{m}} = \rho$$

$$\boxed{\text{ohm} \cdot \text{m} = \rho}$$

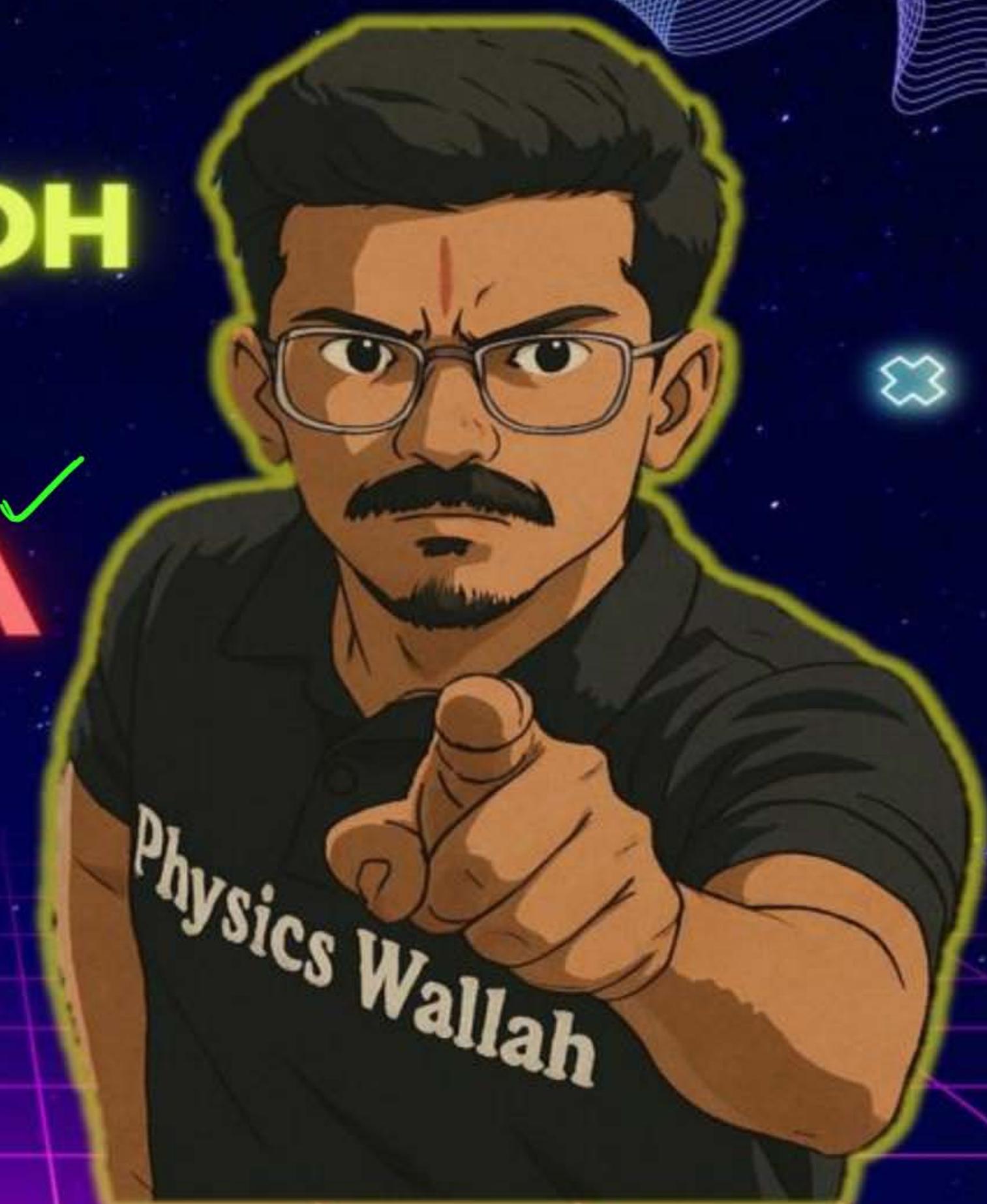
Length & Area \rightarrow 'cm'

$$\rho = \text{ohm} \cdot \text{cm}$$

$$k = \text{S} \cdot \text{cm}^{-1}$$

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BASICS MATHS - LOGÁRITHM

(Utno ↓ pdhenge, jitna kaam auega)

Common

LOGARITHM

- Logarithm is the power to which a base number must be raised to get another number.

- Example: $10^3 = 1000$, $2^2 = 4$, $5^3 = 125$

base \downarrow exponent

$3 = \log_{10} 1000$

$2 = \log_2 4$

$3 = \log_5 125$

The diagram illustrates the inverse relationship between exponents and logarithms. It shows three examples: $10^3 = 1000$, $2^2 = 4$, and $5^3 = 125$. Handwritten annotations explain each: 'base' points to the 10 in 10^3 , 'exponent' points to the 3, and a green arrow labeled 'down' points from the base to the exponent. Brackets group the first two examples: $3 = \log_{10} 1000$ and $2 = \log_2 4$. A bracket also groups the last two examples: $3 = \log_5 125$.

$$3 = \log_5 125$$

IMPORTANT TO REMEMBER

Expression	Base	(ln chemistry)	Value
$\log_{10} 2$	Base 10		≈ 0.3010
$\log_{10} 3$	Base 10		≈ 0.4771
$\log_{10} 5$	Base 10		≈ 0.6990
$\log_{10} 10$	Base 10		= 1
$\log_{10} 1$	Base 10		= 0

PROPERTIES OF LOGARITHM

①

$$\log_a a = \checkmark$$

$$\log_{10} 10 = \textcircled{1}$$

②

$$\log(a^n) = \underline{n} \times \underline{\log a}$$

$$\log_{\textcolor{teal}{10}}(10^{20})$$

$$20 \times \log_{10} 10 = \textcircled{20}$$

③

$$\log(ab) = \overbrace{\log a} + \overbrace{\log b}$$

$$\log(\underbrace{1.3 \times 10^7}) \rightarrow \log(1.3) + \log(10^7)$$

$$\approx 0 + (7 \times \log 10)$$

$$\approx 0 + 7$$

$$\approx \textcircled{7}$$

④

$$\log\left(\frac{a}{b}\right) = \underline{\log a} - \underline{\log b}$$

LET'S PRACTICE

$$\text{pH} = 10$$

$$\text{pH} = -\log [\text{H}^+]$$

$$10 = -\log [\text{H}^+]$$

① $-10 = \log_{10} [\text{H}^+]$

$$10^{-10} \text{ M} = [\text{H}^+]$$

$$3 = \log_{10} 1000$$

$$[10^3 = 1000]$$

② $\log_{10}(100) = 2$

③ $\log_{10}(1000) = 3$

④ $\log_{10}(1) = 0$

⑤ $\log_{10}(10) = 1$

⑥ $\log_{10}(0.1) = \log_{10}\left(\frac{1}{10}\right) = \log_{10}\left(\frac{1}{10}\right) = \log_{10}1 - \log_{10}10 = 0 - 1 = -1$

$$\log_{10}(100) = \log_{10}(10^2) = 2 \times \log_{10}10 = 2 \times 1 = 2$$

$$\log_{10}(1000) = \log_{10}(10^3) = 3 \times \log_{10}10 = 3 \times 1 = 3$$

LET'S PRACTICE

①
 \log_{10}

VII

$$\log(2 \times 5) = \underline{\log 2} + \underline{\log 5} = .3010 + .6990 = \boxed{1.0000}$$

VIII

$$\log\left(\frac{100}{10}\right) = \log 100 - \log 10 = 2 - 1 = \textcircled{1}$$

IX

$$\log(3^2) = 2 \log 3 = 2 \times .4771 = \underline{.9542}$$

X

$$\log\left(\frac{1}{1000}\right) = \log 1 - \log 1000 \Rightarrow 0 - \log_{10} 10^3 \Rightarrow 0 - 3 \log 10 \\ \Rightarrow \textcircled{-3}$$

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(**MEASUREMENT OF**
CONDUCTIVITY OF
ELECTROLYTIC SOLUTIONS)

SHORT RECAP

$$\rho = \frac{R \times A}{l}$$

$$\boxed{\rho = R}$$

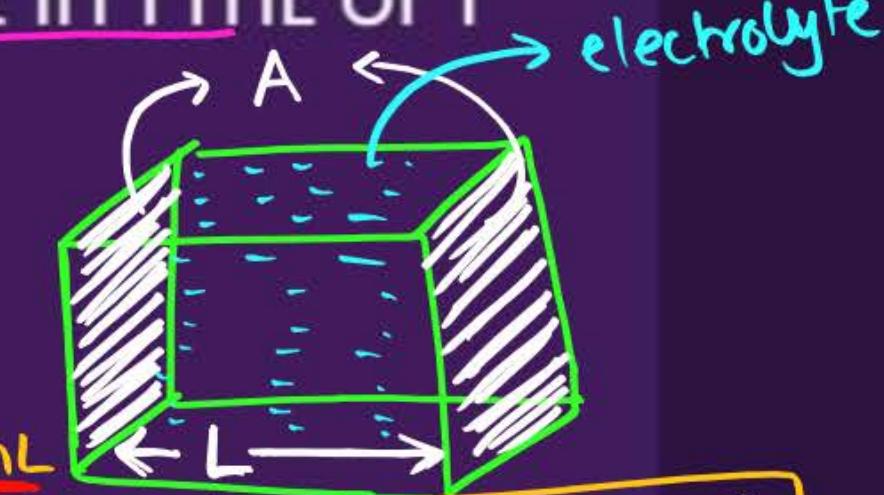
(ρ)

- Resistivity is the resistance (obstruction in flow of current) offered to all the ions of that electrolyte present in 1 mL or 1 cm³ of solution.

$$L \rightarrow 1\text{ cm}$$

$$A \rightarrow 1\text{ cm}^2$$

Volume of electrolyte
 $= A \times l = \underline{\underline{1\text{ cm}^3}}$
 or $\underline{\underline{1\text{ mL}}}$



- Resistivity and conductivity are related as:

(κ)

- Conductivity is the conductance (ease in flow of current) due to all the ions of that electrolyte present in 1 mL or 1 cm³ of solution.

$$\boxed{\kappa = \frac{1}{\rho}}$$

$$\boxed{\kappa = \frac{G \times l}{A} = \frac{1 \times l}{R \times A}}$$

FACTORS AFFECTING CONDUCTIVITY AND RESISTIVITY OF AN ELECTROLYTE

'Resistivity'

I Nature of Electrolyte

Strong
Weak

Strong electrolytes (like HCl, NaCl) give more ions → higher conductivity → lower resistivity

II Concentration of solution

More concentration = more ions = higher conductivity (up to a point)
→ lower resistivity

III Temperature

Higher temperature → more ion movement → higher conductivity → lower resistivity

IV Mobility of Ions

Smaller, faster ions (like H^+ , OH^-) increase conductivity → lower resistivity

NCERT CORNER



Material	Conductivity/ $S \cdot m^{-1}$	Material	Conductivity/ $S \cdot m^{-1}$
Conductors		Aqueous Solutions	
Sodium	2.1×10^3	Pure water	3.5×10^{-5}
Copper	5.9×10^3	0.1 M HCl	3.91
Silver	6.2×10^3	0.01M KCl	0.14
Gold	4.5×10^3	0.01M NaCl	0.12
Iron	1.0×10^3	0.1 M HAc	0.047
Graphite	1.2×10^3	0.01M HAc	0.016

non-metallic conductor

FUN FACT: Superconductors are materials that show zero resistivity or infinite conductivity when cooled below a certain temperature. \rightarrow critical temp.

- Metals and their alloys have very large conductivity and are known as conductors.
- Certain non-metals like carbon-black, graphite and some organic polymers (polyacetylene, polyaniline, polypyrrole and polythiophene) are also electronically conducting.

NCERT CORNER



✓ Insulators

Glass
Teflon

1.0×10^{-16}
 1.0×10^{-18}

✓ Semiconductors

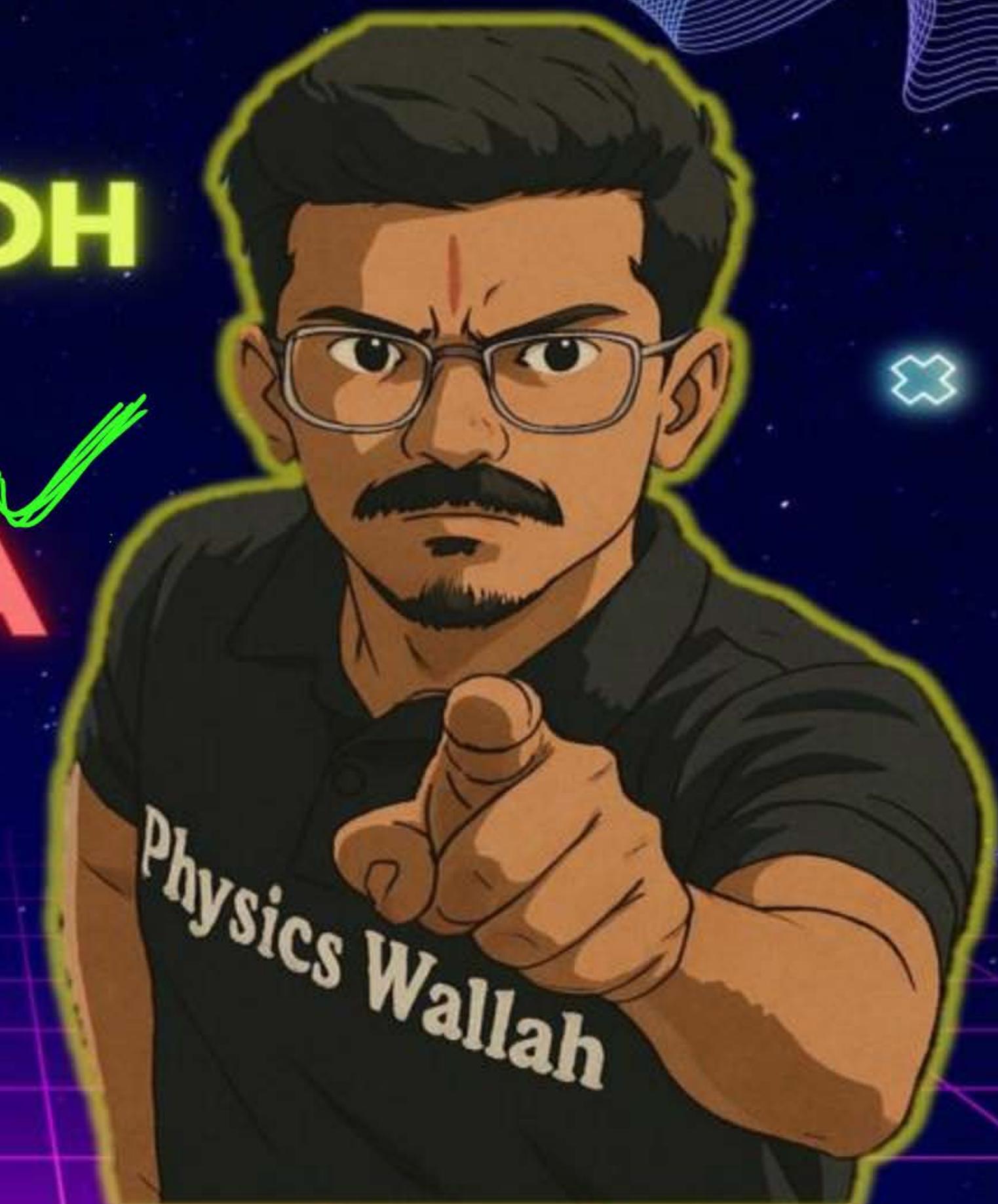
CuO
Si
Ge

1×10^{-7}
 1.5×10^{-2}
2.0

- Substances like glass, ceramics, etc., having very low conductivity are known as **insulators**.
- Substances like silicon, germanium, doped silicon and gallium arsenide having conductivity between conductors and insulators are called **semiconductors**.

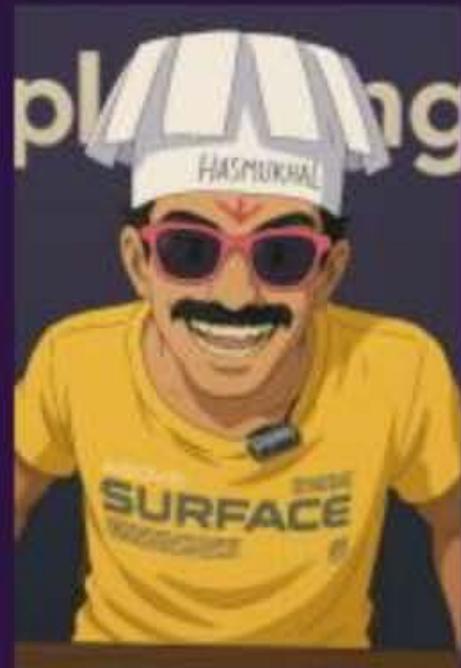
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PROBLEM AND SOLUTION

$$K = \frac{G \times l}{A} = \frac{1}{R} \times \frac{l}{A}$$



'Hasmukhhal'

(alternating current)

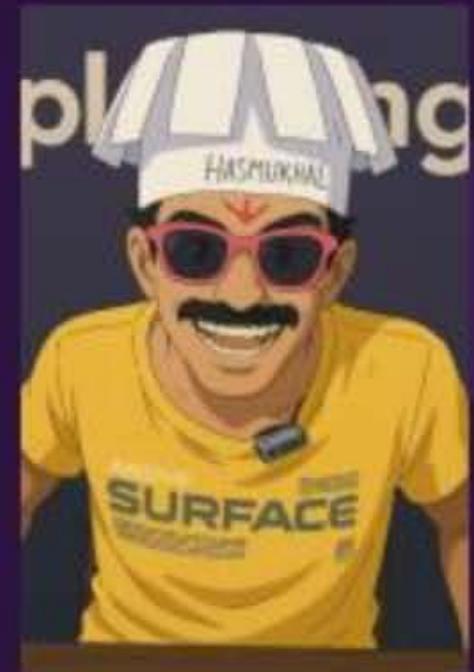
Solutions:

- ✓(i) AC is used instead of DC to avoid electrolysis of electrolyte.
- ✓(ii) A specially designed vessel called conductivity cell is used which gets connected to Wheatstone bridge.

(direct current)



PROBLEM AND SOLUTION



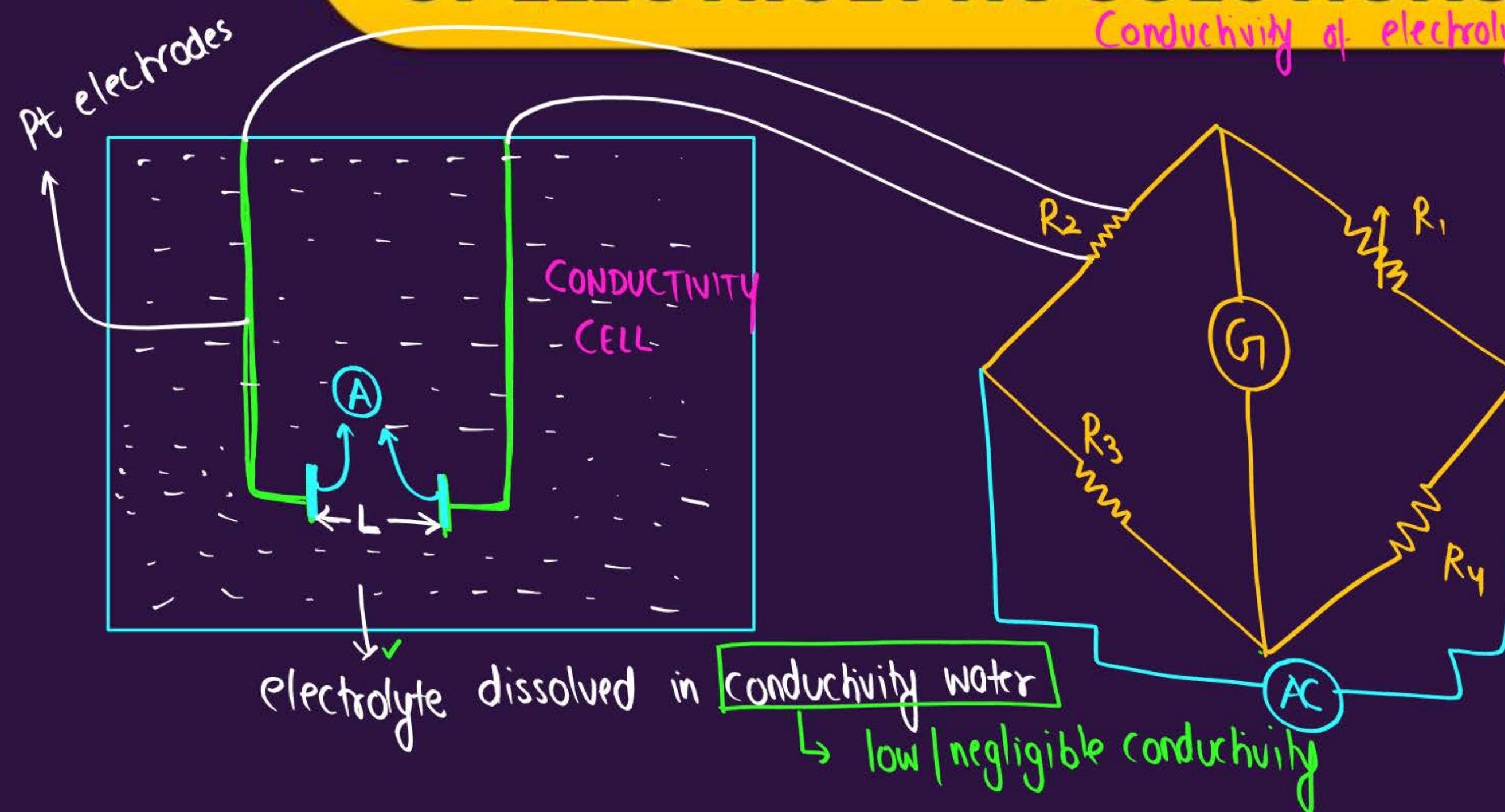
How no electrolysis happens when we use an AC source?

In AC, the direction of current keeps on changing due to which no deposition of ions takes place on electrodes. Also, the number of ions remains the same in the electrolyte and hence, conductivity of electrolyte remains constant.



MEASUREMENT OF CONDUCTIVITY OF ELECTROLYTIC SOLUTIONS

$$\text{Conductivity of electrolyte } (\kappa) = \frac{1}{R} \cdot \frac{A}{l}$$



$R_1, R_3 \downarrow$ known resistance
 $R_2 \rightarrow$ resistance of electrolyte to be measured.
 Balanced condition of Wheatstone bridge

→ No current flows through galvanometer

⇒

$$\frac{R_2}{R_1} = \frac{R_3}{R_4}$$

⇒

$$R_2 = \frac{R_1 \times R_3}{R_4}$$

calculated using cell & wheatstone bridge
 cell constant

$$k = \frac{l}{R} \times \frac{A}{l}$$

Can you (directly measure) $\left(\frac{l}{A}\right)$ of a cell?

- A Yes
- B No

IMPORTANT TERMS TO KNOW

$$\frac{l}{A} = G^*$$

(i) **Cell Constant (G^*)**: It is the ratio of the distance between the electrodes to the area of cross-section of the electrodes in a conductivity cell. Its unit is cm^{-1} .

$$\frac{\text{cm}}{\text{cm}^2} = \text{cm}^{-1}$$

The quantity l/A is called cell constant denoted by the symbol, G^* . It depends on the distance between the electrodes and their area of cross-section and has the dimension of length^{-1} and can be calculated if we know l and A . (Measurement of l and A is not only inconvenient but also unreliable.) The cell constant is usually determined by measuring the resistance of the cell containing a solution whose conductivity is already known. For this purpose, we generally use KCl solutions whose conductivity is known accurately at various concentrations (Table 2.3) and at different temperatures. The cell constant, G^* , is then given by the equation:

$$G^* = \frac{l}{A} = R K \quad (2.18)$$

$$K = \frac{1}{R} \times \frac{l}{A}$$

$$K \times R = G^*$$

*measured
using
standard
conditions*

Step Ist

We will measure $\frac{l}{A}$ or G^* of a conductivity cell by putting KCl in conductivity cell [value of K of KCl is already known] & we will measure its Resistance (R) by connecting it to Wheatstone bridge

$$R \times K = G^* = \left[\frac{1}{A} \right]$$

cell constant we have

Step IInd

Now, clean conductivity cell & put electrolyte whose K (conductivity) needs to be measured. This will be connected to wheatstone bridge

& R of electrolyte (A_1) is calculated.

$$K \text{ (of electrolyte } A_1) = \left(\frac{1}{R} \right) \times \left(\frac{l}{A} \right)$$

Step 2

Step 1

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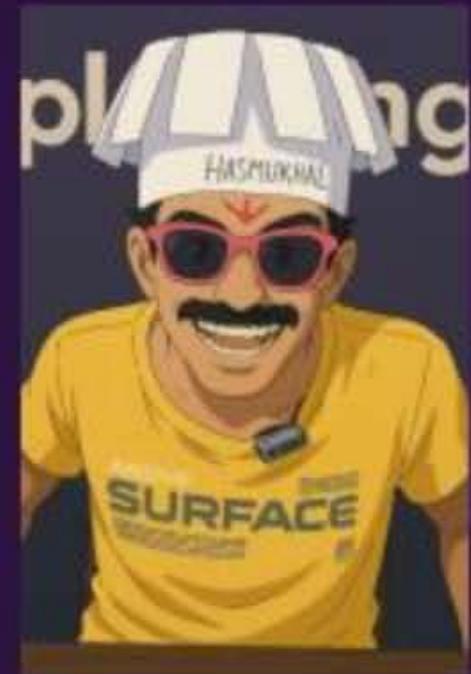
AYE BHAIYA





INTRODUCTION TO MOLAR **CONDUCTIVITY**

PROBLEM AND SOLUTION



Hi! Maan lijiye meine **1 mole NaCl ko 1 L and 10 L water mein daala.** Toh conductivity alag aaegi. Par **same substance ke 1 mole ka performance alag hai?** Toh agar 1 mole ka exact performance of electrolyte?

Toh badia hena, iska solution **yehi hai ki ab ek aesi term Sunil Bhaiya se pdhte hai jo seedha 1 mole ki takat batae.**



IMPORTANT TERMS TO KNOW

(ii) Molar Conductivity (Λ_m): It is the conductivity of 1 mole of electrolyte in a given volume of solution.

(molarity) → Concentration of a soln = ' C ' M = ' C ' $\frac{\text{mol}}{\text{L}}$

' C ' mol are present in 'x' L → $1 \text{ L} = 1000 \text{ mL} | 1000 \text{ cm}^3$
 $\rightarrow \underline{1 \text{ mL} | 1 \text{ cm}^3}$

$$\Rightarrow x \times 1000 = C$$

$$\Rightarrow \boxed{x = \frac{C}{1000}}$$

$$\Lambda_m \times \frac{C}{1000} = K$$

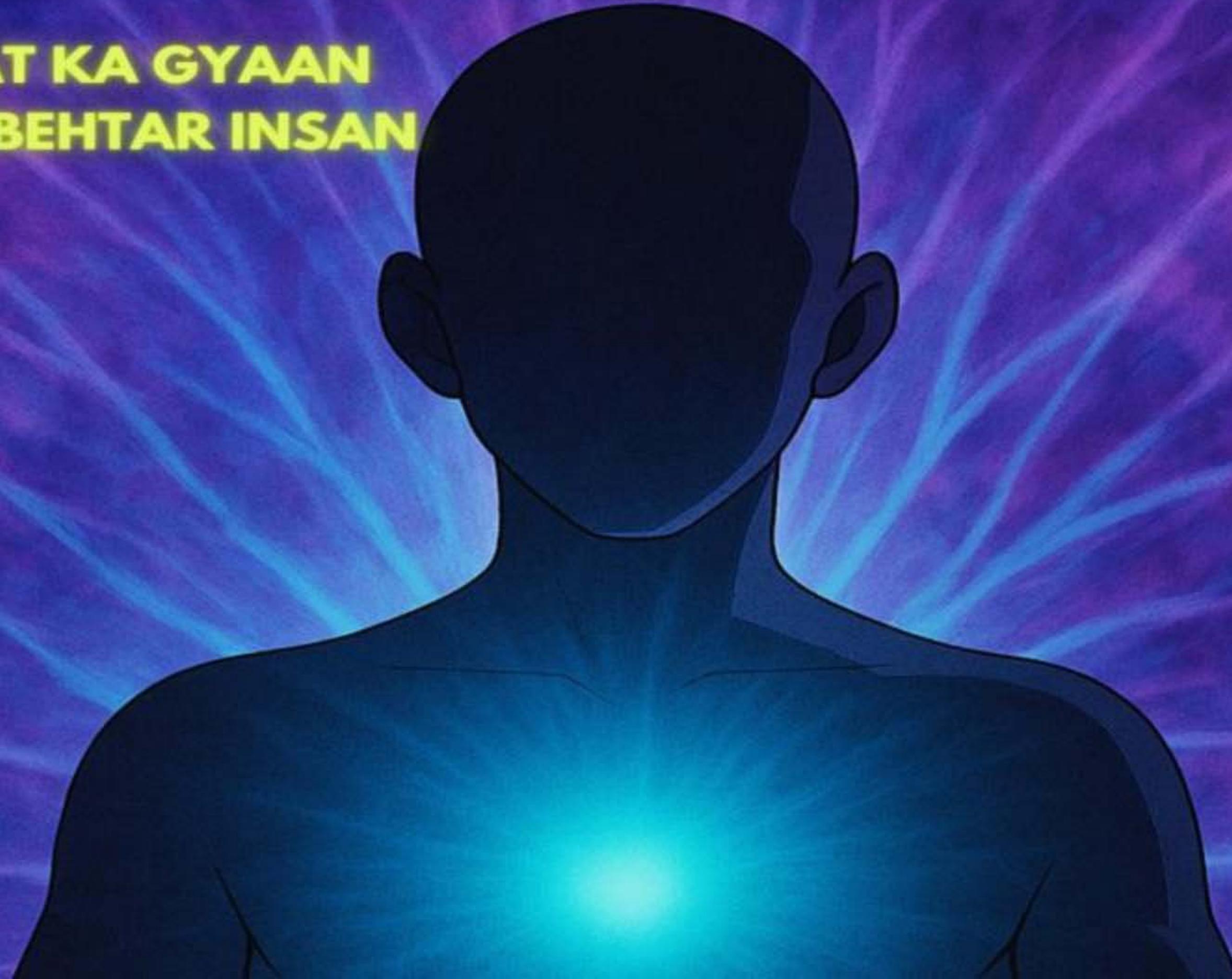
$$\Lambda_m = \frac{K \times 1000}{C}$$

CONCEPT POLISH – HOMEWORK



(Unit of $\lambda_m \rightarrow$ calculate)

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JO BANAE BEHTAR INSAN**



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#pwsathhai (✓)

Thank
You

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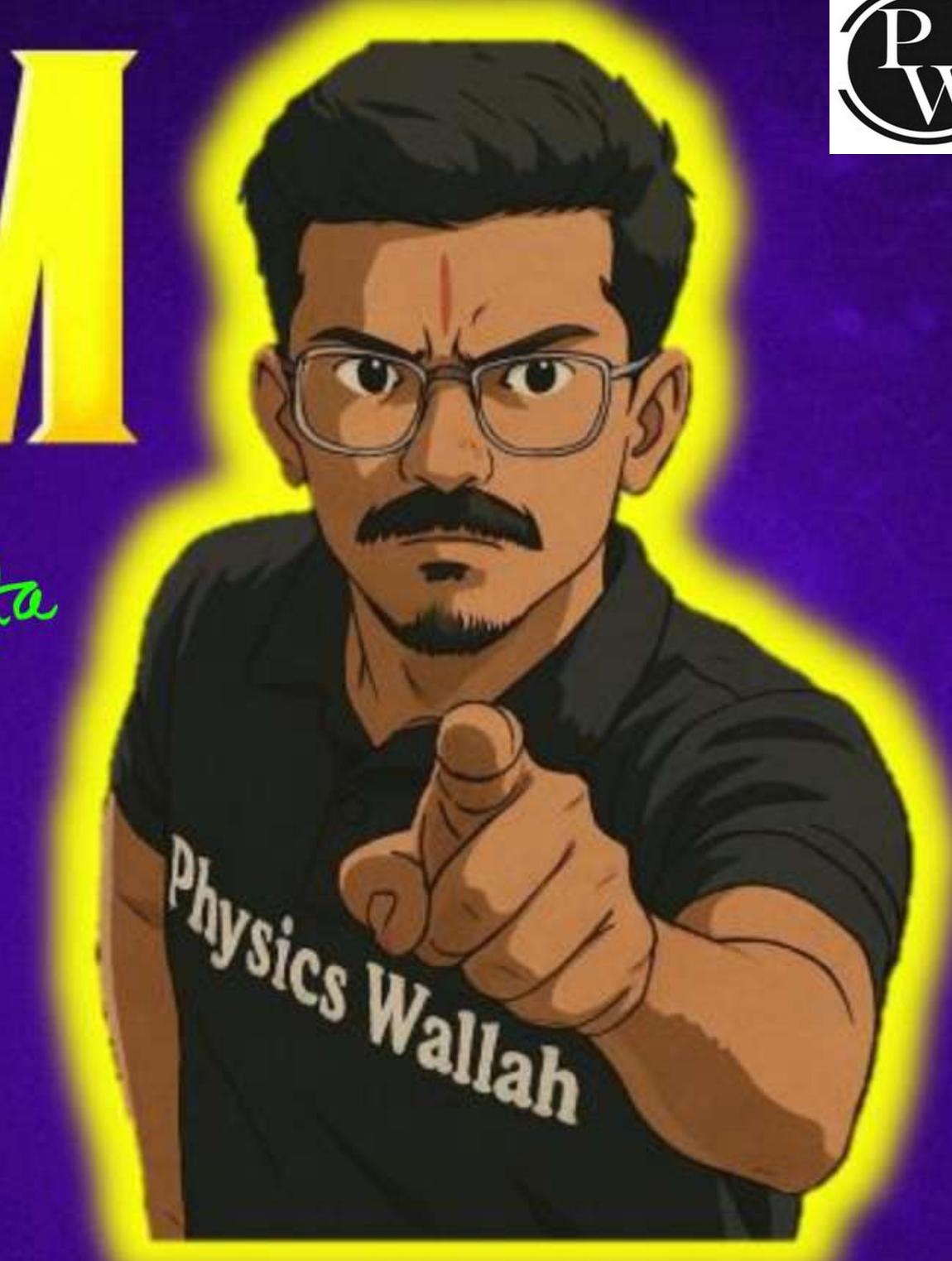
CHEMISTRY

Lecture 07

ELECTROCHEMISTRY

Numerical Practice, Variation of
Conductivity and Molar Conductivity

Bharat Mata
Ki Jai O



BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

- (i) Numerical Practice ✓**
- (ii) Easy Recap – Molar Conductivity ✓**
- (iii) Variation of Conductivity, Molar Conductivity ✓**



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What is the unit of molar conductivity (Λ_m)?

$$\Lambda_m = \frac{K \times 1000}{C}$$

↑ conductivity
↓ molarity

$$= \frac{S \cdot \text{cm}^{-1}}{\text{mol} \cdot \text{cm}^{-3}}$$

$\text{mol} \cdot \text{L}^{-1}$
↓
mol $\cdot \text{cm}^{-3}$

$$= \boxed{S \cdot \text{cm}^2 \cdot \text{mol}^{-1}}$$



NUMERICAL PRACTICE

NCERT EXERCISE 2.8

The conductivity of 0.20 M solution of KCl at 298 K is 0.0248 S cm⁻¹. Calculate its molar conductivity.

Given: M or C = .20 M

To find: $\Lambda_m = ?$

$$K = .0248 \text{ S} \cdot \text{cm}^{-1}$$

$$\Lambda_m = \frac{K \times 1000}{C}$$

$$= \frac{0.0248 \times 1000 \times 10^2}{0.20 \times 10^4} = \frac{124}{20} = 124 \text{ S} \cdot \text{cm}^2 \cdot \text{mol}^{-1}$$

NCERT EXERCISE 2.9

$$G^* = \frac{1}{A} = \frac{\text{cm}}{\text{cm}^2} = (\text{m}^1 \cdot \text{m}^{-2}) = \text{m}^{-1}$$

The resistance of a conductivity cell containing 0.001M KCl solution at 298 K is 1500 Ω . What is the cell constant if conductivity of 0.001M KCl solution at 298 K is $0.146 \times 10^{-3} \text{ S cm}^{-1}$.

2 3
 . 1 4 6
 1 5

 1 7 3 0
 1 4 6 0

 2 1 9 0

Given: $R = 1500 \Omega$

To find: $G^* = ?$

$$K = 0.146 \times 10^{-3} \text{ S.cm}^{-1}$$

$$K = \frac{G^* \times l}{A}$$

$$K = \frac{l \times G^*}{R \times A}$$

$$K = \frac{l \times G^*}{R} \Rightarrow G^* = K \times R$$

$$G^* = 0.146 \times 10^{-3} \times 15 \times 10^2$$

$$= 0.146 \times 15 \times 10^{-1}$$

$$= \frac{0.146 \times 15}{10} = \frac{2.190}{10}$$

$$= 0.219 \text{ cm}^{-1}$$

NCERT SOLVED EXAMPLE 2.4

$$S = \Omega^{-1}$$

Resistance of a conductivity cell filled with 0.1 mol L^{-1} KCl solution is 100Ω . If the resistance of the same cell when filled with 0.02 mol L^{-1} KCl solution is 520Ω , calculate the conductivity and molar conductivity of 0.02 mol L^{-1} KCl solution. The conductivity of 0.1 mol L^{-1} KCl solution is 1.29 S/m .

Given: $R_1 = 100 \Omega$

$\cdot 1 \text{ mol/L KCl}$

$K = 1.29 \text{ S} \cdot \text{m}^{-1}$

↓

$$G^* = K \times R$$

$$= \frac{1.29 \text{ S} \cdot \text{m}^{-1} \times 100 \Omega}{100}$$

$$= 1.29 \text{ m}^{-1} = \frac{1.29}{1\text{m}} = \frac{1.29}{100 \text{ cm}} = 1.29 \text{ cm}^{-1}$$

$R_2 = 520 \Omega$

$K = ?$ $\Lambda_m = ?$

$\cdot 0.02 \text{ mol/L KCl}$

$$K = G \times \frac{l}{A}$$

$$= \frac{1}{R} \times \frac{l}{A} = \frac{1}{R} \times G^*$$

need to find

Put value of G^* given in ①

$$K = \frac{1}{520} \times \frac{1.29}{100}$$

$$= \frac{1.29}{520} \times \frac{1}{100}$$

$$K = \frac{0.24 \times 1}{100}$$

$$K = 0.24 \times 10^{-2} \text{ S} \cdot \text{cm}^{-1}$$

$$\Lambda_m = \frac{K \times 1000}{C}$$

$$= \frac{0.24 \times 10^{-2} \times 1000}{\cancel{C}}$$

$$= \frac{0.24 \times 10^1}{0.02 \times \cancel{10^{-2}}} \times 10^2 = \frac{240}{\cancel{2}} =$$

120 $\text{S} \cdot \text{cm}^2 \cdot \text{mol}^{-1}$

$$\begin{array}{r}
 \cdot 24 \\
 \hline
 520 \quad \boxed{1290} \\
 \hline
 1040 \\
 \hline
 2500 \\
 \hline
 \underline{2080}
 \end{array}$$

NCERT SOLVED EXAMPLE 2.5

molarity



The electrical resistance of a column of 0.05 mol L^{-1} NaOH solution of diameter 1 cm and length 50 cm is $5.55 \times 10^3 \text{ ohm}$. Calculate its resistivity, conductivity and molar conductivity.

Given: $R = 5.55 \times 10^3 \Omega$

$d = 1 \text{ cm}$, $L = 50 \text{ cm}$, C or M = .05 M

$$\begin{aligned}
 \textcircled{1} \quad \text{Resistivity } (\rho) &= \frac{R \times A}{l} = \frac{\cancel{5.55} \times 10^3 \times 6785}{\cancel{50} \times 10^2 \times \cancel{1000}} \\
 &= \frac{111 \times 785}{10^3} \\
 &= \frac{87135}{1000} = 87.135 \Omega \cdot \text{cm}
 \end{aligned}$$

To find: $\rho, K\epsilon, \Lambda_m$

$$\begin{aligned}
 A &= \pi r^2 \\
 &= 3.14 \times \left(\frac{1}{2}\right)^2 \\
 &= \cancel{3.14} \times \frac{1}{4} \\
 &= \frac{1.57}{2} \\
 &= \underline{\underline{.785 \text{ cm}^2}}
 \end{aligned}$$

$$\textcircled{2} \quad K = \frac{1}{\beta} = \frac{1}{87.135} \approx \frac{1}{87} \quad \approx .011 \text{ S} \cdot \text{cm}^{-1}$$

$$\textcircled{3} \quad \lambda_m = \frac{K \times 1000}{c} \\ \approx \frac{.011 \times 1000}{.05 \times 1000} \\ \approx \frac{11 \times 10^2}{605} \\ \approx \frac{11 \times 10^2}{5} \quad \approx \boxed{220 \text{ S} \cdot (\text{m}^2 \cdot \text{mol})^{-1}}$$

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EASY RECAP – MOLAR CONDUCTIVITY (Λ_m)

MOLAR CONDUCTIVITY

(no. of moles
of solute in 1 L
soln)

(i) Molar Conductivity (Λ_m) in Terms of Molarity (M or C): It is the conductance of all the ions produced by 1 mole of electrolyte in a given volume of solution.

mole (mol)

'C' mol of electrolyte are present in \rightarrow 1 L soln

\hookrightarrow 'C' mol of electrolyte are present in \rightarrow 1000 mL or 1000 cm³ soln

(I)

'x' mol of electrolyte \rightarrow 1 mL or 1 cm³

$$x \times 1000 = C$$

$$\boxed{x = \frac{C \text{ mol}}{1000}}$$

(II)

$\frac{C}{1000}$ mol $\xrightarrow{\text{Conductance}}$ K

1 mol $\xrightarrow{\text{Conductance}}$

K

Λ_m

$$\Lambda_m \times \frac{C}{1000} = K$$

$$\boxed{\Lambda_m = \frac{K \times 1000}{C}}$$

MISCONCEPTION REMOVER

one bhaiya ☺

Property	Conductivity (κ)	Molar Conductivity (Λ_m)
Definition	Conductance of <u>1 cm³</u> of solution <u>(any concentration)</u>	Conductance of <u>all ions produced by 1 mole of electrolyte</u> → <u>concentration is fixed</u>
Depends on	Volume = 1 cm³ (geometry fixed)	Volume = V litres (variable, enough to contain 1 mole)
Symbol	κ	Λ_m
Unit	S cm⁻¹	S cm²mol⁻¹

MOLAR CONDUCTIVITY

(ii) Molar Conductivity (Λ_m) in Terms of Electrode Geometry:

Molar conductivity of a solution at a given concentration is the conductance of the volume V of solution containing one mole of electrolyte kept between two electrodes with area of cross section A and distance of unit length.

$$K = \frac{G \times l}{A}$$

$$G = \frac{K \times A}{l}$$

$$l = 1 \text{ cm}$$

$$G = K \cdot A$$

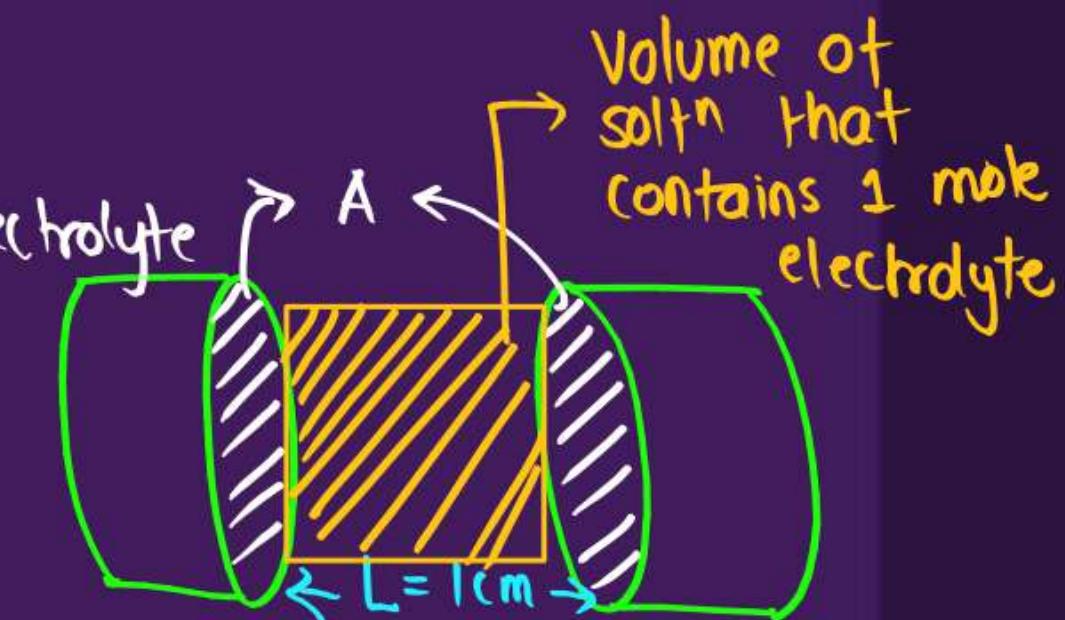
G is conductance
 Λ_m is conductance of 1 mole of electrolyte

for 1 mole electrolyte,

$$\Lambda_m = G = K \cdot A \quad (1)$$

$$\text{Volume} = A \times l = A \times 1$$

Now, (1) will be $\Lambda_m = G = KV$



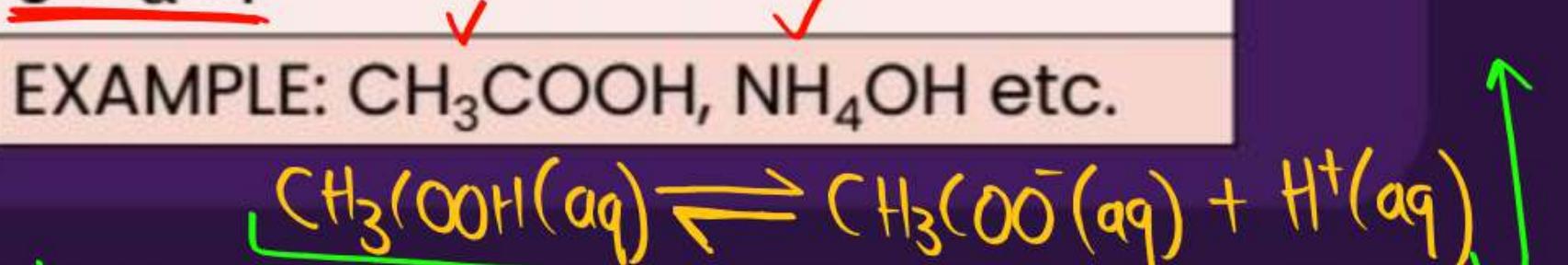
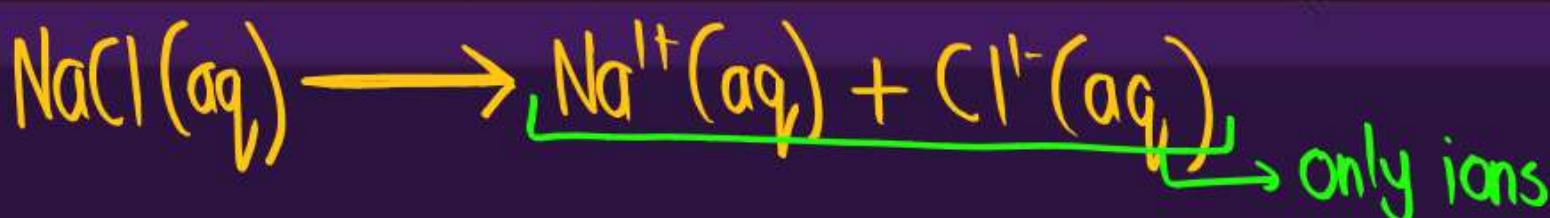
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**[VARIATION OF CONDUCTIVITY,
MOLAR CONDUCTIVITY AND
KOHLRAUSCH LAW]**

CONCEPT RECAP

STRONG ELECTROLYTE	WEAK ELECTROLYTE
(i) They <u>completely dissociate</u> into <u>ions in aqueous or molten state</u> .	(i) They <u>partially dissociate into</u> ions in aqueous or molten state.
(ii) Contains <u>only ions</u> .	(ii) Contains <u>both ions and molecules</u> .
(iii) Strong acids, strong bases and salts	(iii) Weak acids and bases
(iv) Degree of dissociation (α) is <u>1</u> .	(iv) Degree of dissociation (α) is: <u>$0 < \alpha < 1$</u>
EXAMPLE: KCl, HCl, MgSO ₄ etc.	EXAMPLE: CH ₃ COOH, NH ₄ OH etc.



molecules
↔ ions both

Dilution → adding more water

GIVE A THOUGHT

In B, no. of ions
in 1 cm^3 / 1 mL (\downarrow)
Hence, conductivity (K) \downarrow



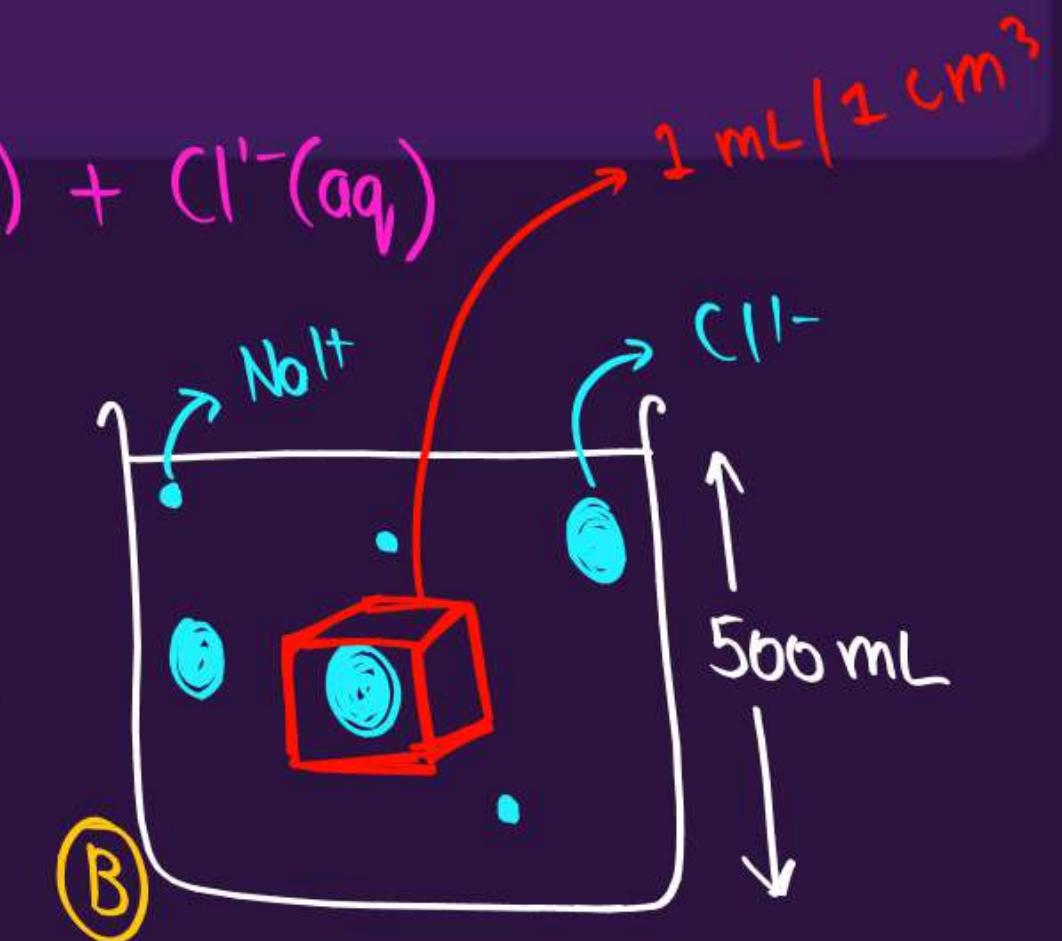
What happens to conductivity of a strong electrolyte solution on dilution?

- A. Conductivity increases
- B. Conductivity decreases

Conductance due
to ions present
in this volume



Dilution
Adding more water

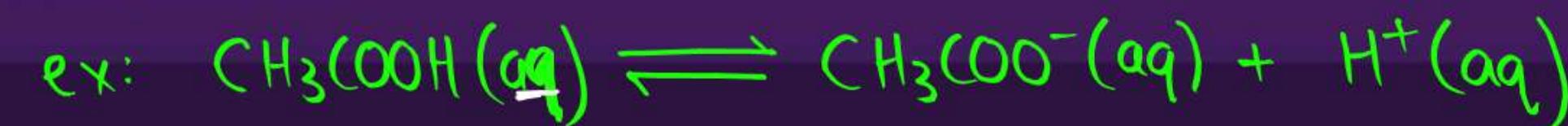


GIVE A THOUGHT

What happens to conductivity of a weak electrolyte solution on dilution?

- A. Conductivity increases
- B. Conductivity decreases

CONCEPTUAL
Twist



Dilution of weak electrolyte → Equilibrium disturbs

According Le Chatlier principle, dissociation into ions will (\uparrow) to restore equilibrium

So, still in diluted solution

there is less no. of ions present in $1\text{ mL}/1\text{ cm}^3$ region as compare to initial soln

BUT, Volume of soln also (\uparrow)

GIVE A THOUGHT

Suppose a strong and a weak electrolyte solution gets diluted then molar conductivity will be:

- A. Increased (Λ_m)
- B. Decreased

On dilution, volume of soln (\uparrow)

Strong electrolyte

$$\boxed{\Lambda_m = K \times V}$$

decrease increase

Weak electrolyte

$$\boxed{\Lambda_m = K \times V}$$

decrease increase

Effect of (\uparrow) in volume of soln $>$ effect of (\downarrow) in K

Λ_m that will (\uparrow)

ANOTHER WAY OF UNDERSTANDING

On
Dilution, Volume of
solution (\uparrow)

$$C \text{ or } M = \frac{n}{V}$$

$$V (\uparrow), C \downarrow$$

1. Concentration (c) (\downarrow)

- Water add karne se solution mein electrolyte zyada dilute ho jaata hai.

2. Ion-ion interactions (\downarrow) force of attraction btw. ions (\downarrow)

- Ions ab ek dusre se door door ho jaate hain \rightarrow unki mobility badh jaati hai

3. Conductivity (κ) (\downarrow)

($1\text{ cm}^3/\text{s m}\text{Cl}$)

- Ions per volume kam ho jaate hain \rightarrow overall conductivity kam hoti hai

4. BUT:

$$\Lambda_m = K \times V$$

or

$$\Lambda_m = \frac{\kappa}{c}$$

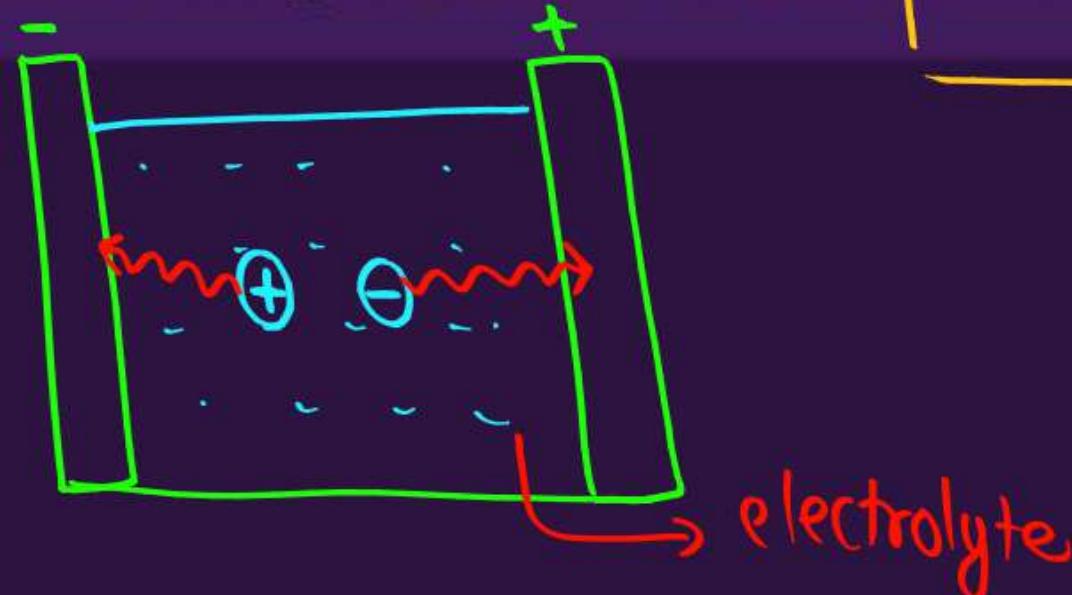
- Yahaan pe k thoda kam hota hai, par c aur bhi zyada kam hota hai

- Isliye Λ_m ka net effect $\rightarrow \Lambda_m$ increases on dilution

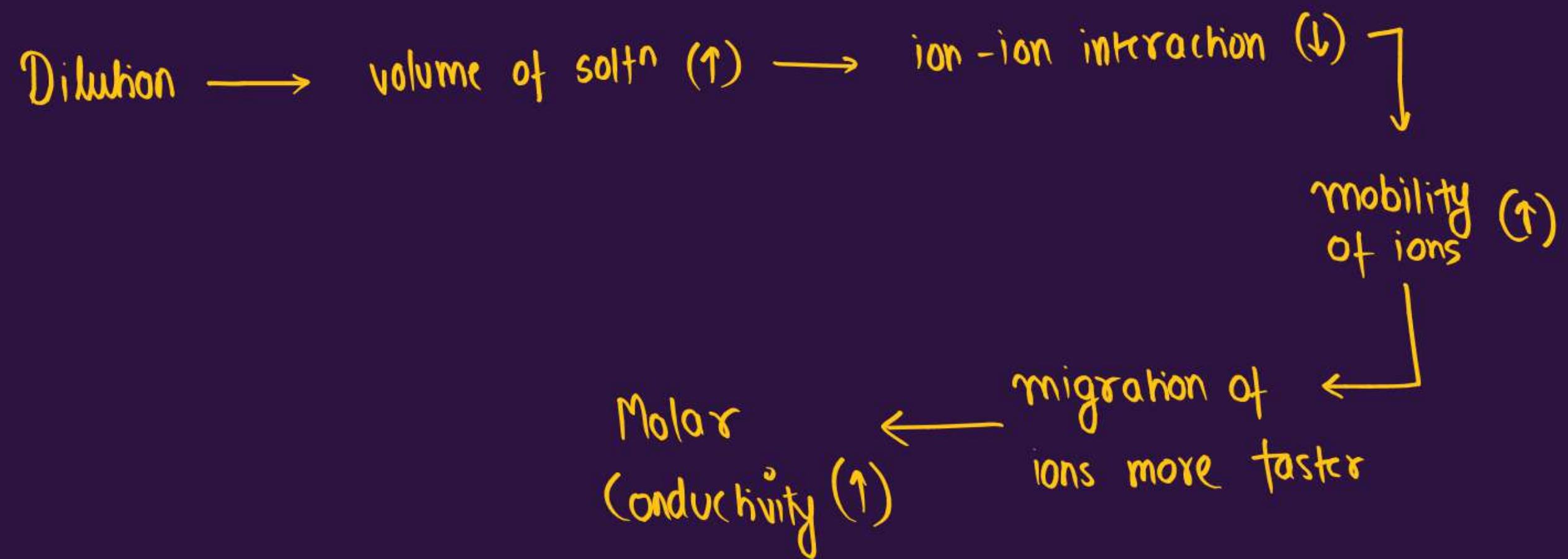
Using this
will also tell
about Λ_m

WHAT IS MIGRATION?

- Movement of ions under the influence of electric field. This movement of ions is what carries the electric current through the solution.
- Mobility of ions = how easily they migrate
- Faster and easy the migration, more will be the molar conductivity (Λ_m).



Way - 03



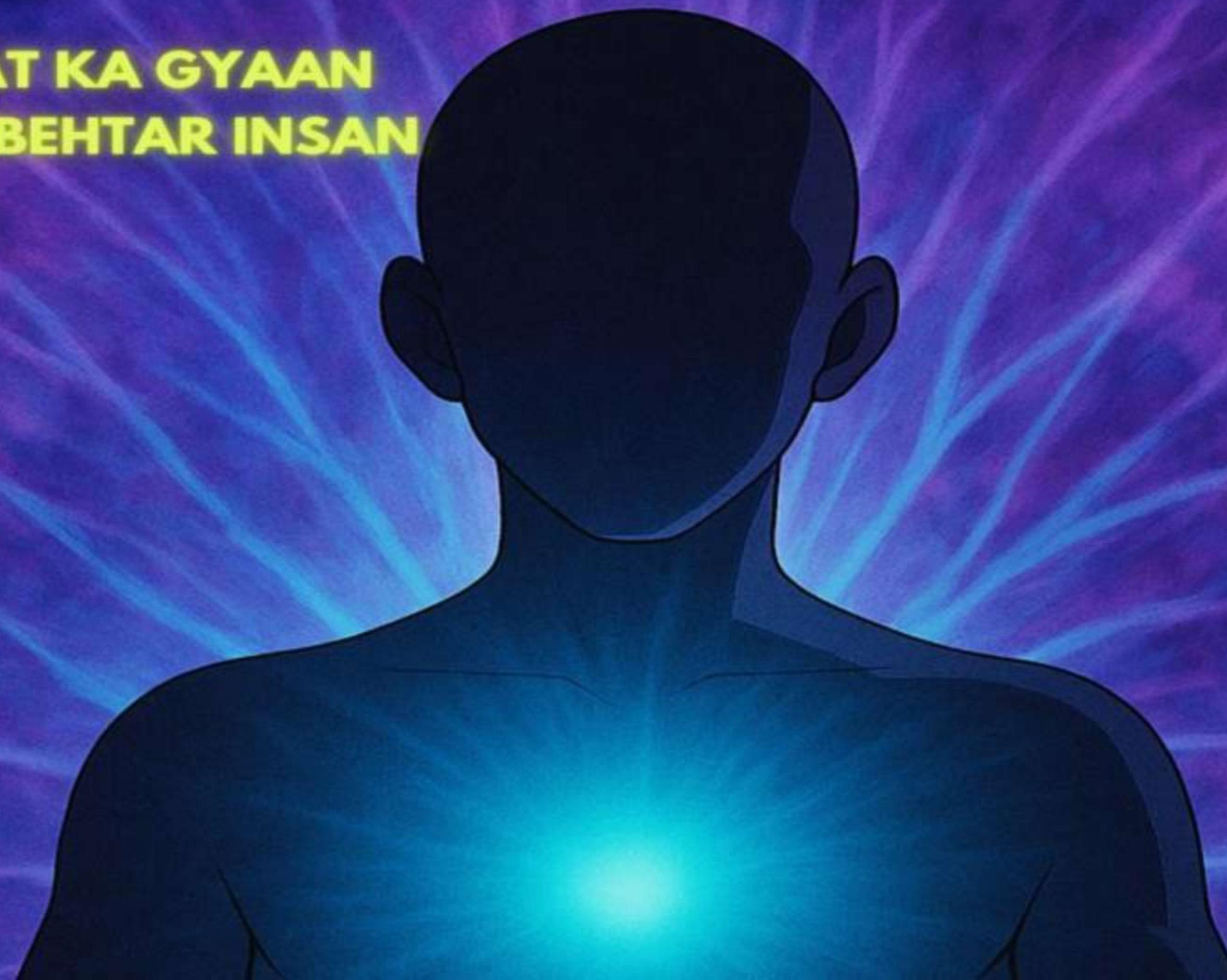
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Physics Wallah



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#pwsathhai (✓)



**Thank
You**



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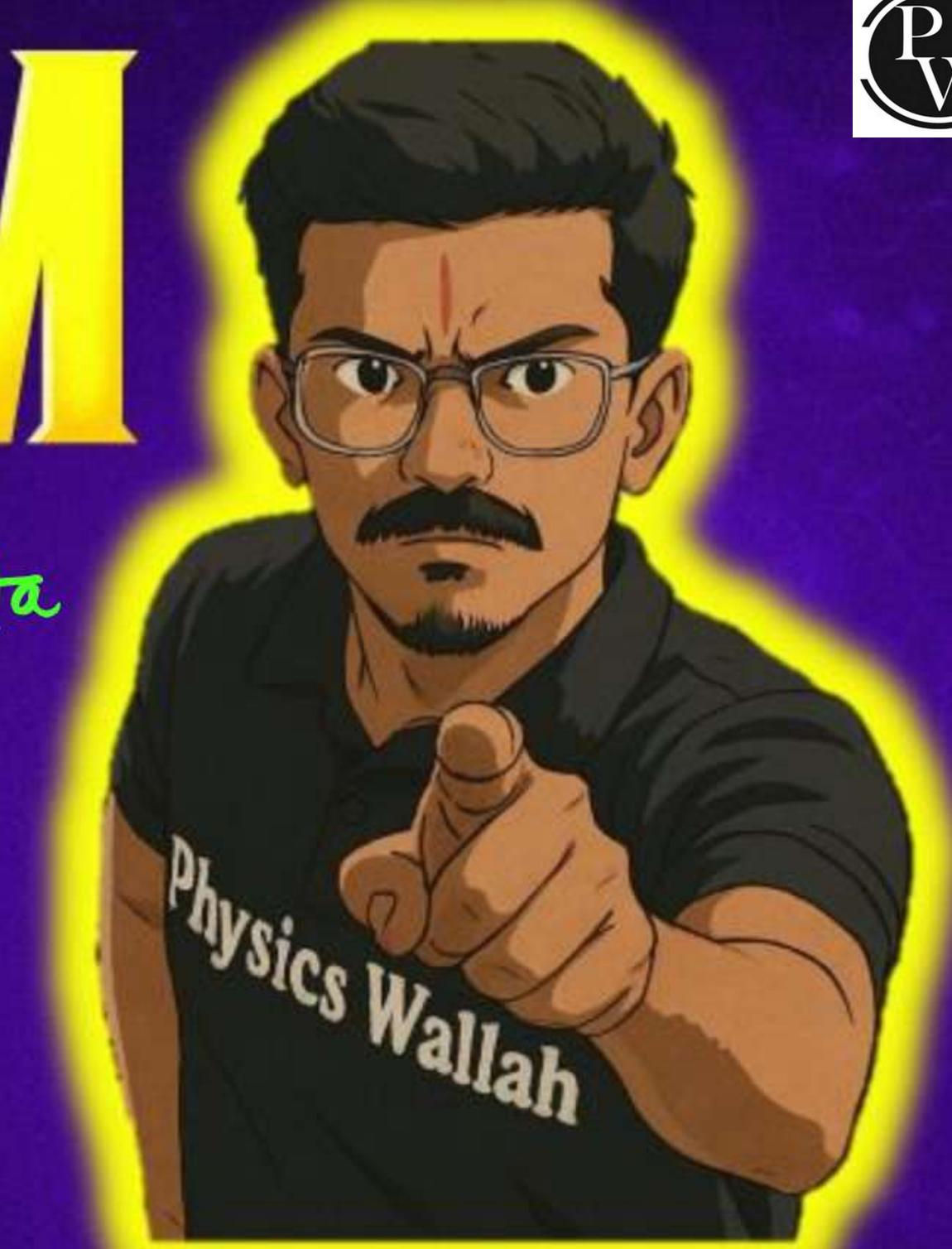
CHEMISTRY

Lecture 08

ELECTROCHEMISTRY

PYQ Practice, Kohlrausch's Law and Its
Applications

Bharat Mata
Ki Jai ❤



BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

- (i) Limiting Molar Conductivity,
Kohlrausch Law and Its
Applications (✓)
- (ii) CBSE Previous Year Questions (✓)



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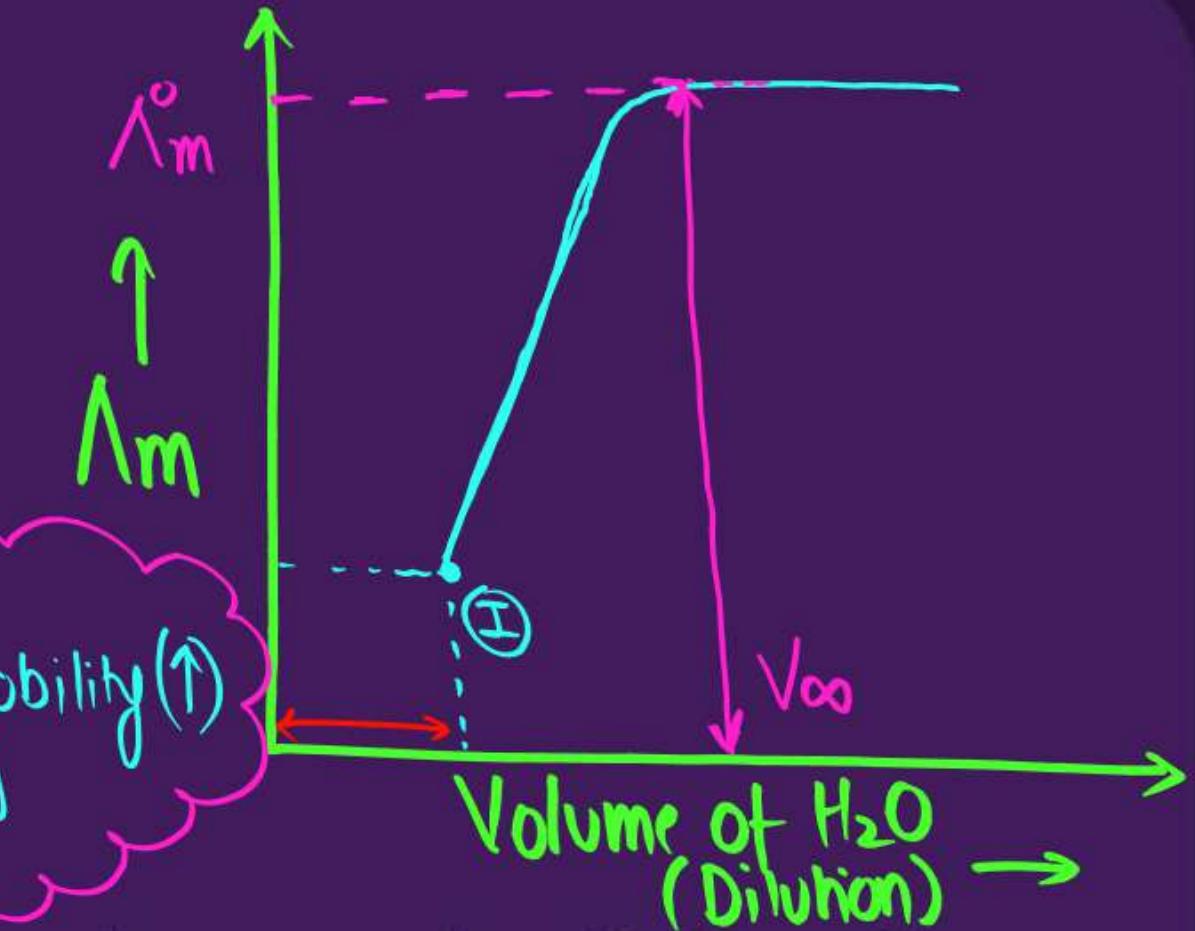
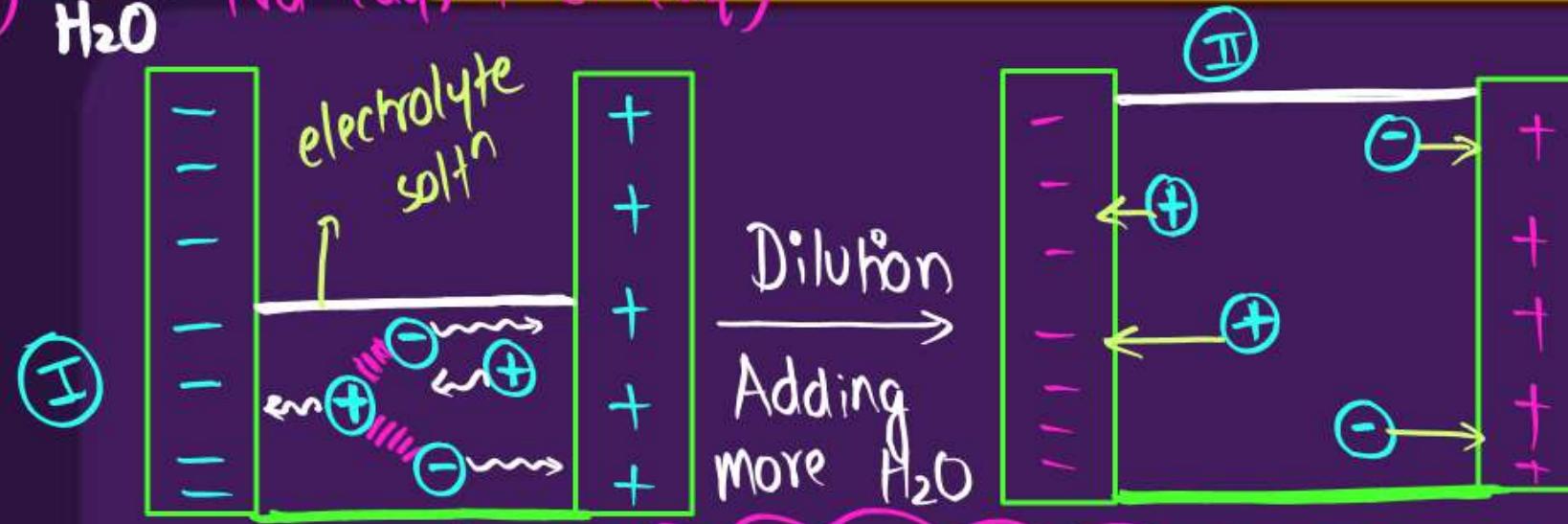
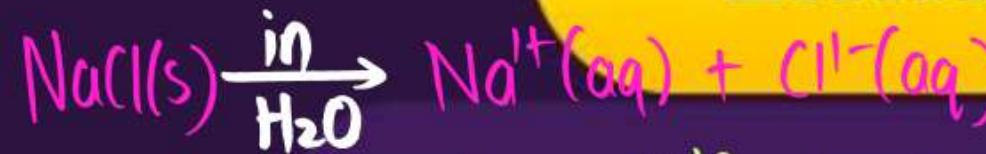


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(**LIMITING MOLAR
CONDUCTIVITY, KOHLRAUSCH
LAW AND ITS APPLICATIONS**)

LIMITING MOLAR CONDUCTIVITY

 Λ_m°


CORE CONCEPT

Dilution (\uparrow), Ion-Ion Attraction (\downarrow), Ion mobility (\uparrow)
 Molar conductivity (\uparrow) \leftarrow Migration speed \uparrow

- Limiting molar conductivity is the molar conductivity of an electrolyte at infinite dilution, i.e., when concentration approaches zero and ions are far apart with no interaction between them.
- Its S.I. unit is $\text{S}\cdot\text{cm}^2\cdot\text{mol}^{-1}$

$$C = \frac{n_{\text{solute}}}{V_{\text{solution}}} \xrightarrow{\text{Dilution}} V_{\text{soln}}^{\infty} (\uparrow), C \rightarrow 0$$

LET'S REVISE

Property	At limiting molar conductivity (Λ_m°)	At Infinite Dilution
Ion-ion interaction		zero
Ions		Move freely
Dissociation		Complete (even for weak electrolytes)
Molar conductivity (Λ_m)		Maximum possible → called Λ_m°
Concentration	$C \neq 0$ $C \rightarrow 0$	Not zero, but tends towards zero

GIVE A THOUGHT

Can we say that limiting molar conductivity is the maximum molar conductivity of an electrolyte solution?

- A. Yes
- B. No

KOHLRAUSCH LAW

- It is also called the law of independent migration of ions.
- The law states that limiting molar conductivity of an electrolyte can be represented as the sum of the individual contributions of the anion and cation of the electrolyte.

$$\Lambda_m^{\circ}(A_x B_y) = x \lambda_{A^+}^{\circ} + y \lambda_{B^-}^{\circ}$$

Ex:

$$\Lambda_m^{\circ}(NaCl) = 1 \lambda_{Na^+}^{\circ} + 1 \lambda_{Cl^-}^{\circ}$$

$$\Lambda_m^{\circ}(AlCl_3) = 1 \lambda_{Al^{3+}}^{\circ} + 3 \lambda_{Cl^-}^{\circ}$$

At infinite dilute
there is zero force
of attraction b/w
(cation & anion)
(cation & anion independently
migrate to oppositely
charged electrodes)

KOHLRAUSCH LAW

- For strong electrolytes he observed that:

$$\Lambda_m = \Lambda_m^0 - A\sqrt{c}$$

↓
 y ↓ c + m ↓ x

Graph of Λ_m vs \sqrt{c} is a straight line

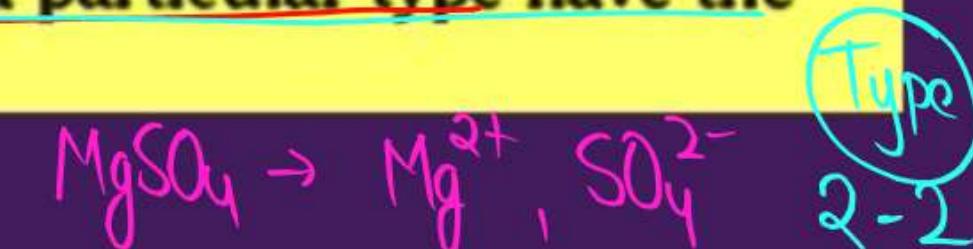
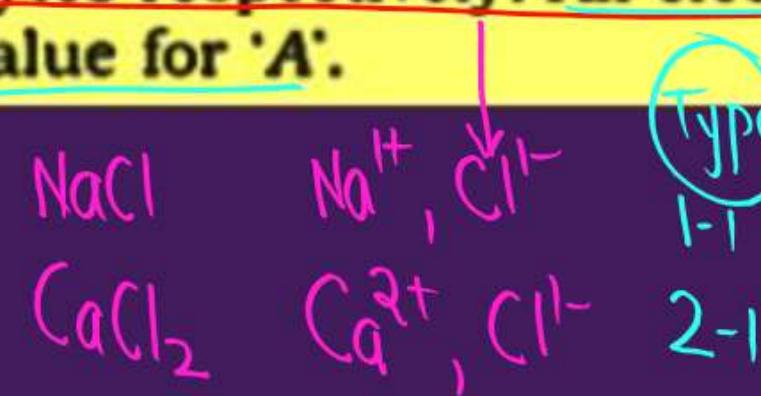
Intercept = Λ_m^0

$m \rightarrow$ (Slope) = $-A$ (depends on ion type)

$\Lambda_m \rightarrow$ molar conductivity
 $c \rightarrow$ concentration of electrolyte
 $\Lambda_m^0 \rightarrow$ limiting molar conductivity

(NCERT)

It can be seen that if we plot (Fig. 2.6) Λ_m against $c^{1/2}$, we obtain a straight line with intercept equal to Λ_m^0 and slope equal to ' $-A$ '. The value of the constant ' A ' for a given solvent and temperature depends on the type of electrolyte i.e., the charges on the cation and anion produced on the dissociation of the electrolyte in the solution. Thus, NaCl, CaCl₂, MgSO₄, are known as 1-1, 2-1 and 2-2 electrolytes respectively. All electrolytes of a particular type have the same value for ' A '.



The value of A for NaCl is 'd'. Which among the following electrolyte has same value of A?



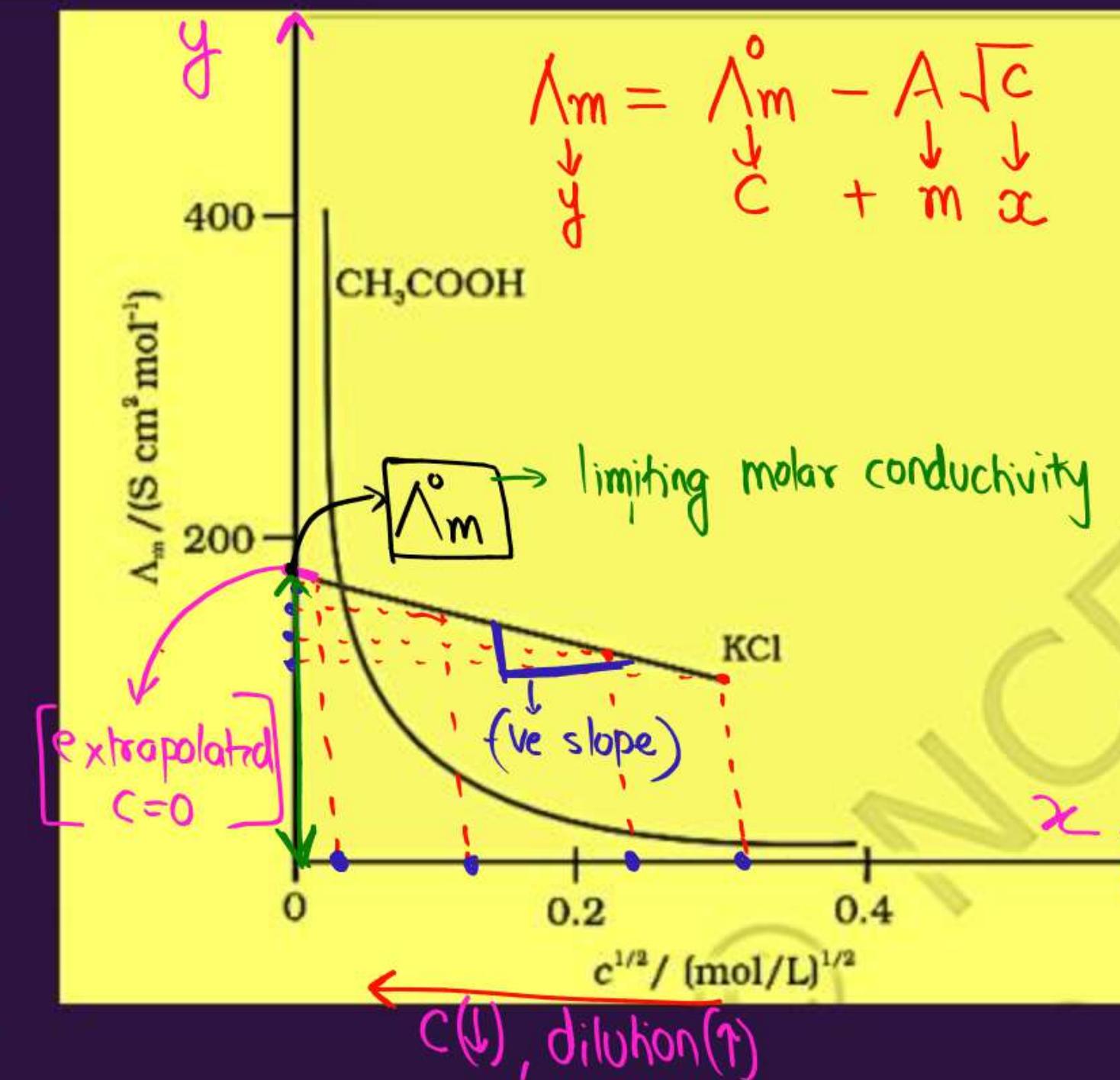
A	Type	'd'
NaCl	$\text{Na}^{1+}, \text{Cl}^{1-}$	1-1
MgCl_2	$\text{Mg}^{2+}, \text{Cl}^{1-}$	2-1
CaCl_2	$\text{Ca}^{2+}, \text{Cl}^{1-}$	2-1
CaSO_4	$\text{Ca}^{2+}, \text{SO}_4^{2-}$	2-2
KCl	$\text{K}^{1+}, \text{Cl}^{1-}$	1-1

GIVE A THOUGHT

How limiting molar conductivity will be calculated for strong electrolyte like KCl?

Steps:

- ✓ 1. Measure molar conductivity (Λ_m) at different concentrations
- ✓ 2. Plot graph of Λ_m vs \sqrt{c}
- ✓ 3. Extrapolate the straight line to $\sqrt{c} = 0$
4. The intercept gives Λ_m^0



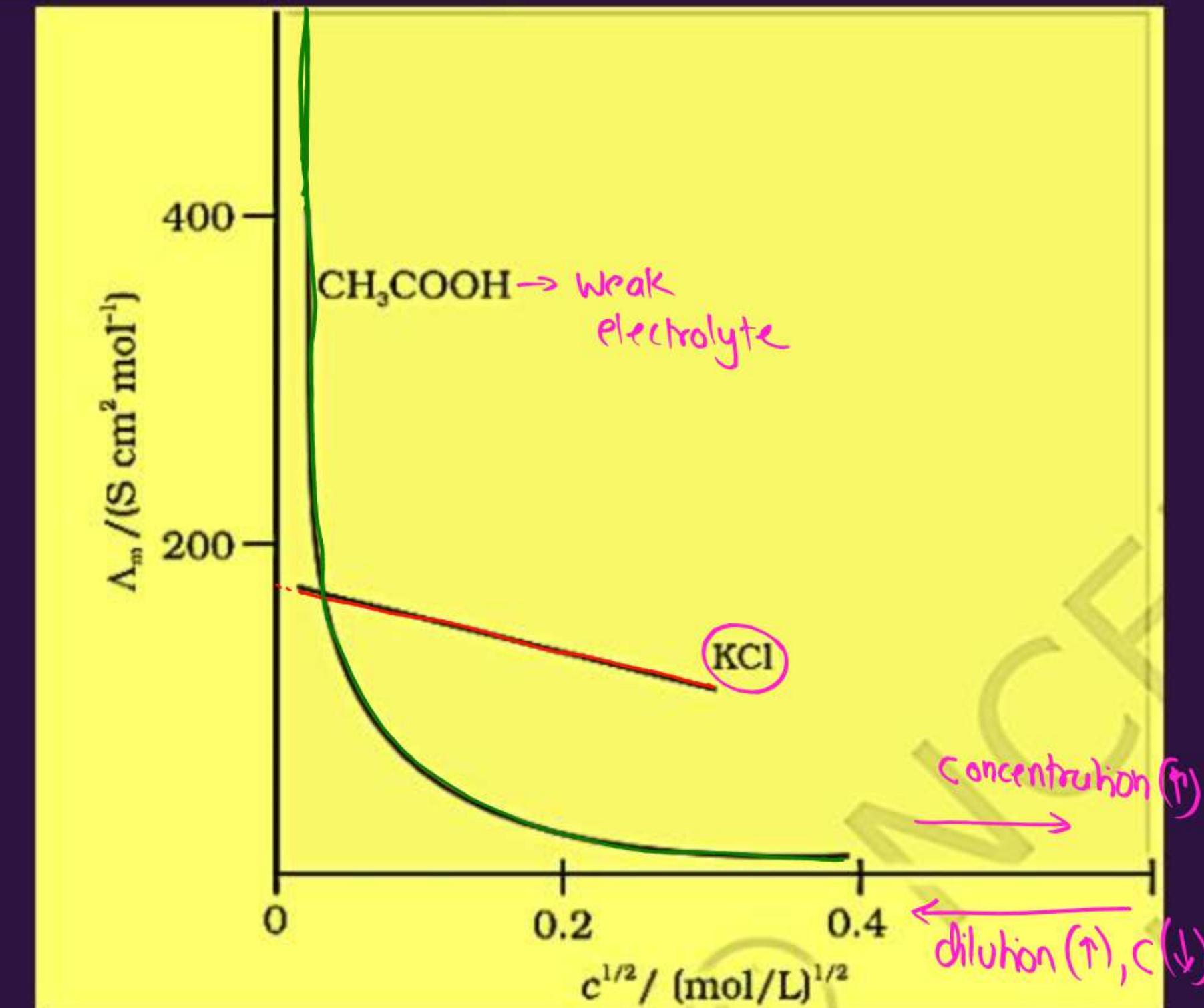
MISCONCEPTION REMOVER

Jab graph ka intercept Λ_m^0 hota hai at $\sqrt{c} = 0$,
toh wo theoretical extrapolated point hota hai —

Actual concentration zero nahi hoti, bas zero ke kareeb hoti hai $(c \rightarrow 0)$

GIVE A THOUGHT

In the given graph, molar conductivity of strong electrolyte increases linearly on dilution but for weak electrolyte it sharply increases non-linearly. Why?



GIVE A THOUGHT



- Strong electrolytes are fully dissociated at all concentrations.



On dilution: → Ion-ion attraction (↓)

- Ions get freedom to move faster → mobility increases (↑), migration speed (↑)
- So Λ_m increases



- Weak electrolytes are only partially dissociated.



- On dilution:

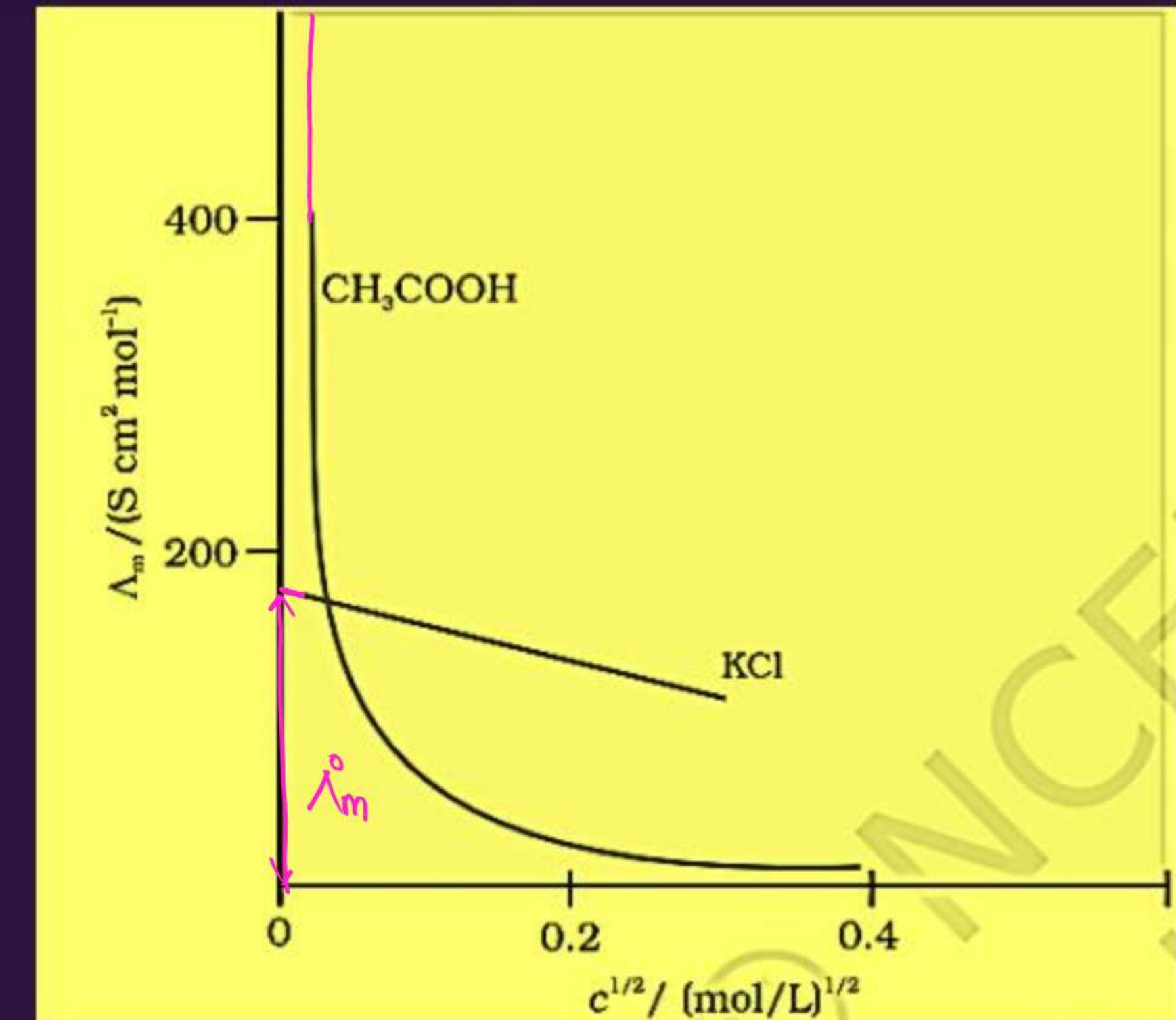
Dissociation increases

- More ions are produced + mobility also increases

- So Λ_m sharply increases (not linearly)

GIVE A THOUGHT

Can we extrapolate the limiting molar conductivity of CH_3COOH – weak electrolyte using this graph?



APPLICATIONS OF KOHLRAUSCH'S LAW

- Calculation of limiting molar conductivity of weak electrolytes:** We will use Kohlrausch law of strong electrolytes to calculate limiting molar conductivity of weak electrolytes because the same cannot be extrapolated through the graph.

$$\lambda_m^{\circ}(\text{CH}_3\text{COOH}) = \lambda_{\text{H}^+}^{\circ} + \lambda_{\text{CH}_3\text{COO}^-}^{\circ}$$

$$\lambda_{\text{CH}_3\text{COO}^-}^{\circ} + \lambda_{\text{H}^+}^{\circ}$$

$$\lambda_m^{\circ}(\text{CH}_3\text{COONa}) + \lambda_{\text{HCl}}^{\circ} - \lambda_{\text{NaCl}}^{\circ}$$

$$\lambda_{\text{CH}_3\text{COO}^-}^{\circ} + \cancel{\lambda_{\text{Na}^+}^{\circ}} + \lambda_{\text{H}^+}^{\circ} + \cancel{\lambda_{\text{Cl}^-}^{\circ}} - \cancel{\lambda_{\text{Na}^+}^{\circ}} - \cancel{\lambda_{\text{Cl}^-}^{\circ}}$$

LET'S PRACTICE



QUESTION

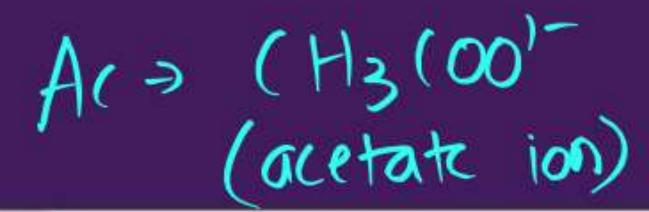
Calculate the limiting molar conductivity of AgI (weak electrolyte) if limiting molar conductivity of strong electrolytes like NaI, NaNO₃ and AgNO₃ are x, y and z respectively.

$$\Lambda_m^{\circ}(\text{AgI}) = 1\lambda_{\text{Ag}^+}^{\circ} + 1\lambda_{\text{I}^-}^{\circ}$$

$$\Lambda_m^{\circ}(\text{AgNO}_3) + \Lambda_m^{\circ}(\text{NaI}) - \Lambda_m^{\circ}(\text{NaNO}_3)$$

$$z + x - y$$

NCERT SOLVED EXAMPLE 2.8



CBSE PYQ



Λ_m° for NaCl, HCl and NaAc are 126.4, 425.9 and 91.0 S cm² mol⁻¹ respectively. Calculate Λ° for HAc

$$\Lambda_m^{\circ}(\text{CH}_3\text{COOH}) = 1\lambda^{\circ}_{\text{CH}_3\text{COO}^-} + 1\lambda^{\circ}_{\text{H}^+}$$



$$\Lambda_m^{\circ}(\text{CH}_3\text{COONa}) + \Lambda_m^{\circ}(\text{HCl}) - \Lambda_m^{\circ}(\text{NaCl})$$

$$= 91 + 425.9 - 126.4$$

$$= \boxed{390.5 \text{ S.cm}^2 \cdot \text{mol}^{-1}}$$

$$\begin{array}{r}
 425.9 \\
 91.0 \\
 \hline
 516.9 \\
 -126.4 \\
 \hline
 390.5
 \end{array}$$

LET'S REVISE

- Calculation of limiting molar conductivity of strong and weak electrolyte)

Type of Electrolyte	How to Find Λ_m^0
Strong	From graph of Λ_m vs \sqrt{c} (extrapolation) $c=0 \rightarrow \Lambda_m^0$ (intercept)
Weak	Using Kohlrausch's Law (indirect calculation) values of strong electrolyte will help to calculate Λ_m^0 of weak electrolyte

PREREQUISITES



Degree of Dissociation (α): It indicates how much of the substance has broken down into its constituent ions or molecules relative to its original amount.

$$\alpha = \frac{\text{Number of moles dissociated}}{\text{Initial number of moles}}$$

$\alpha = 1 \rightarrow$ strong electrolytes

$0 < \alpha < 1 \rightarrow$ weak electrolyte



Ostwald Dilution Law: According to Ostwald Dilution Law, at infinite dilution the degree of dissociation of weak electrolytes is 1.

(α)

(weak acids & weak bases)

APPLICATIONS OF KOHLRAUSCH'S LAW

- Degree of dissociation of weak electrolytes:

Degree of dissociation (α)

Molar Conductivity

$$1 \text{ (infinite dilution)} \rightarrow \Lambda_m^{\circ}$$

$$\alpha \left(\text{at any point before infinite dilution} \right) \rightarrow \Lambda_m$$

$$\Rightarrow \alpha \times \Lambda_m^{\circ} = \Lambda_m \Rightarrow \boxed{\alpha = \frac{\Lambda_m}{\Lambda_m^{\circ}}}$$

NCERT HORROR

But we know that for a weak electrolyte like acetic acid (Class XI, Unit 7),

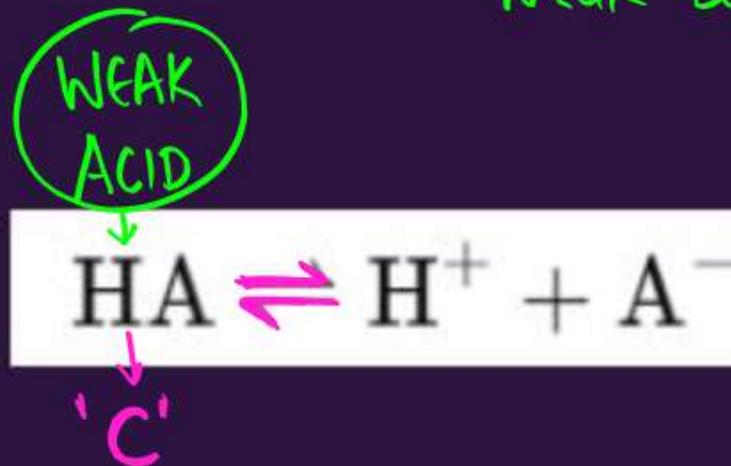
$$K_a = \frac{c\alpha^2}{(1 - \alpha)} = \frac{cA_m^2}{A_m^{\circ 2} \left(1 - \frac{A_m}{A_m^{\circ}}\right)} = \frac{cA_m^2}{A_m^{\circ} (A_m^{\circ} - A_m)} \quad (2.27)$$

Class XI
Ionic
Equilibrium



Weak electrolyte → Weak Acid & Weak Base

PREREQUISITES



$$\text{D.o.d.}(\alpha) = \frac{\alpha}{f} = \boxed{\alpha}$$

Species	Initial Concentration	Change in Concentration	Equilibrium Concentration
HA	c	$-c\alpha$	$c(1 - \alpha)$
H^+	0	$+c\alpha$	$c\alpha$
A^-	0	$+c\alpha$	$c\alpha$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

$$K_a = \frac{(c\alpha)(c\alpha)}{c(1-\alpha)}$$

\Downarrow

$$\frac{c^2\alpha^2}{c(1-\alpha)}$$

$$K_a = \frac{c\alpha^2}{1 - \alpha}$$

(equilibrium
constant)

K_a is the acid dissociation constant or ionization constant of acid which is a quantitative measure of strength of acid.

PREREQUISITES

Weak base



Species	Initial Concentration	Change in Concentration	Equilibrium Concentration
HA	c	$-c\alpha$	$c(1 - \alpha)$
H^+	0	$+c\alpha$	$c\alpha$
A^-	0	$+c\alpha$	$c\alpha$

$$K_b = \frac{[\text{M}^+][\text{OH}^-]}{[\text{MOH}]}$$

$$K_b = \frac{(c\alpha)(c\alpha)}{c(1 - \alpha)}$$

$$K_b = \frac{c\alpha^2}{1 - \alpha}$$

K_b is the base dissociation constant or ionization constant of base which is a quantitative measure of strength of base.

APPLICATIONS OF KOHLRAUSCH'S LAW

- **Equilibrium Constant and Molar Conductivity:**

$$K_a \text{ or } K_b = \frac{C\alpha^2}{1-\alpha}$$

Weak electrolytes (weak acids & weak bases)

'out-of-syllabus'

final
version on
slide 23

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

$$K_a = C \left(\frac{\Lambda_m}{\Lambda_m^0} \right)^2 \quad \left[\begin{array}{l} \oplus \\ \ominus \end{array} \right]$$

$$1 - \left(\frac{\Lambda_m}{\Lambda_m^0} \right)$$

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

Divide by α in num. & den.

$$\Rightarrow K_a = \frac{C\alpha}{1-1}$$

$$\Rightarrow \frac{1}{\alpha} = \frac{C\alpha}{K_a}$$

$$\frac{1}{\alpha} = 1 + \frac{C\alpha}{K_a}$$

$$\left[\alpha = \frac{\Lambda_m}{\Lambda_m^0} \right]$$

Out-of-syllables

$$\frac{1}{\left(\frac{\Delta m}{\lambda^o m}\right)} = 1 + \frac{C \frac{\Delta m}{\lambda^o m}}{K_a}$$

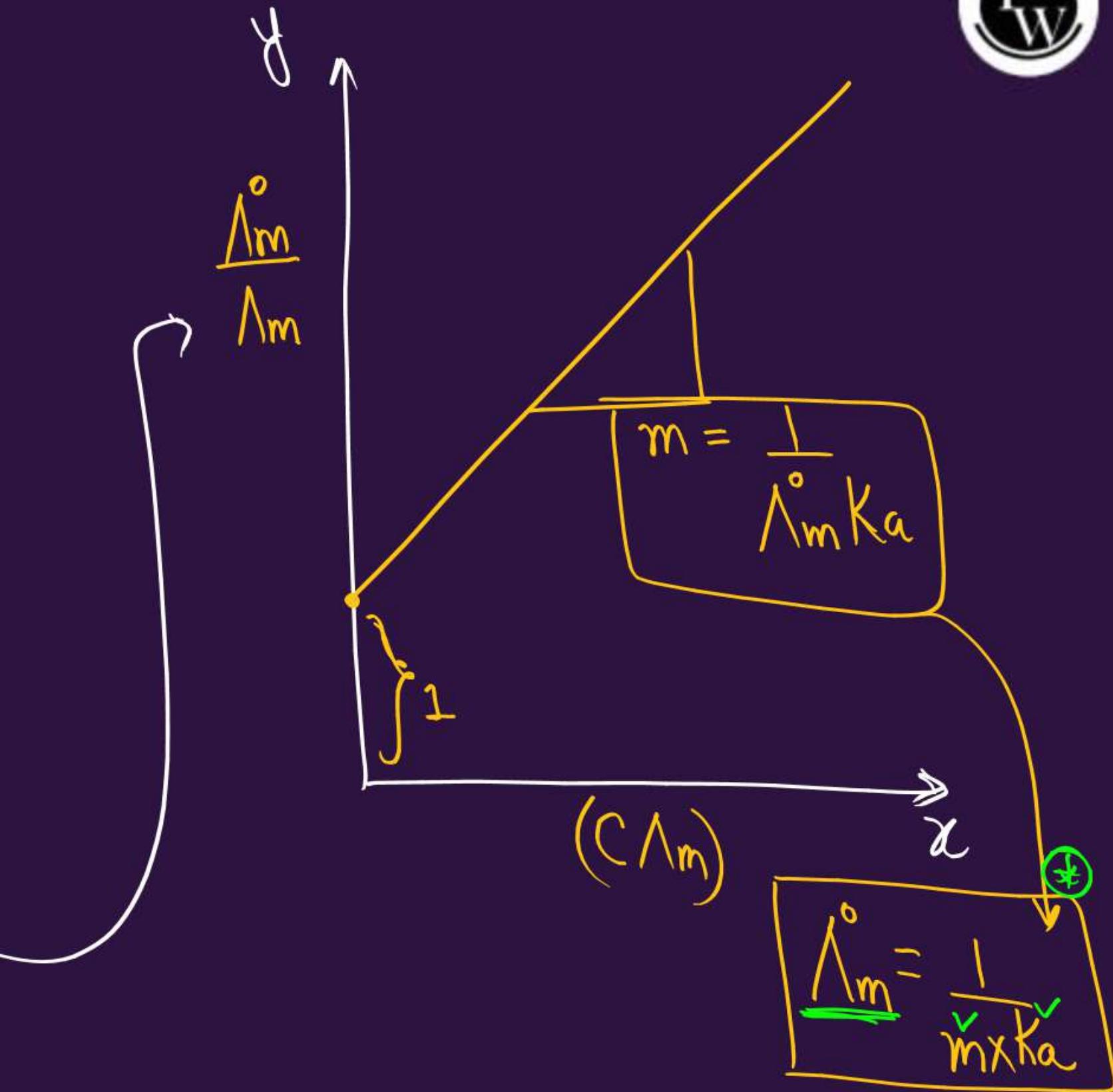
$$\Rightarrow \frac{\lambda^o m}{\Delta m} = 1 + \frac{C \lambda m}{\lambda^o m K_a}$$

*

$$\Rightarrow \boxed{\frac{\lambda^o m}{\Delta m} = 1 + \frac{1}{C} \left(\frac{\lambda m}{\lambda^o m K_a} \right)}$$

$\downarrow \quad \downarrow \quad \downarrow \quad \downarrow$

$y \quad C \quad m \quad x$



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LIKH DO.

AYE BHAIYA





CBSE Previous Year Questions

Why does the conductivity of a solution decreases on dilution?

Conductivity reduces on dilution because number of ions present in 1 cm^3 or 1 mL of solution decreases.

Define limiting molar conductivity. Why there is a steep rise in the molar conductivity of a weak electrolyte on dilution?

(Λ_m)

- Limiting molar conductivity is the molar conductivity of an electrolyte at infinite dilution, i.e., when concentration approaches zero and ions are far apart with no interaction between them. (It is the maximum molar conductivity of an electrolyte solution which is attained at infinite dilution)
 - Weak electrolytes are only partially dissociated.
 - On dilution:
 - Dissociation increases
 - More ions are produced + mobility also increases
 - So Λ_m sharply increases (not linearly)

How can you determine limiting molar conductivity for strong and weak electrolyte?

Type of Electrolyte

On slide 20

- Strong
- Weak

How to Find Λ_m^0

From graph of Λ_m vs \sqrt{c} (extrapolation)

Using Kohlrausch's Law (indirect calculation)

The conductivity of a .01 M solution of acetic acid at 298 K is $1.65 \times 10^{-4} \text{ S cm}^{-1}$. Calculate molar conductivity of the solution.

$$C = .01 \text{ M}$$

$$K = 1.65 \times 10^{-4} \text{ S cm}^{-1}$$

$$\Lambda_m = ?$$

$$\Lambda_m = \frac{K \times 1000}{C} = \frac{1.65 \times 10^{-4} \times 1000}{\frac{.01}{10^2}}$$

$$\Rightarrow \frac{1.65 \times 10^{-4} \times 10^3 \times 10^2}{1}$$

$$\Rightarrow 1.65 \times 10^1$$

$$\Rightarrow 16.5 \text{ S cm}^2 \cdot \text{mol}^{-1}$$

(a) State Kohlrausch law of independent migration of ions. Write an expression for the molar conductivity of acetic acid at infinite dilution according to Kohlrausch law.

[All India(C)]

Slide 8

Slide 16

(b) Calculate Λ_m° for acetic acid.

Given that Λ_m° (HCl) = 426 S cm² mol⁻¹

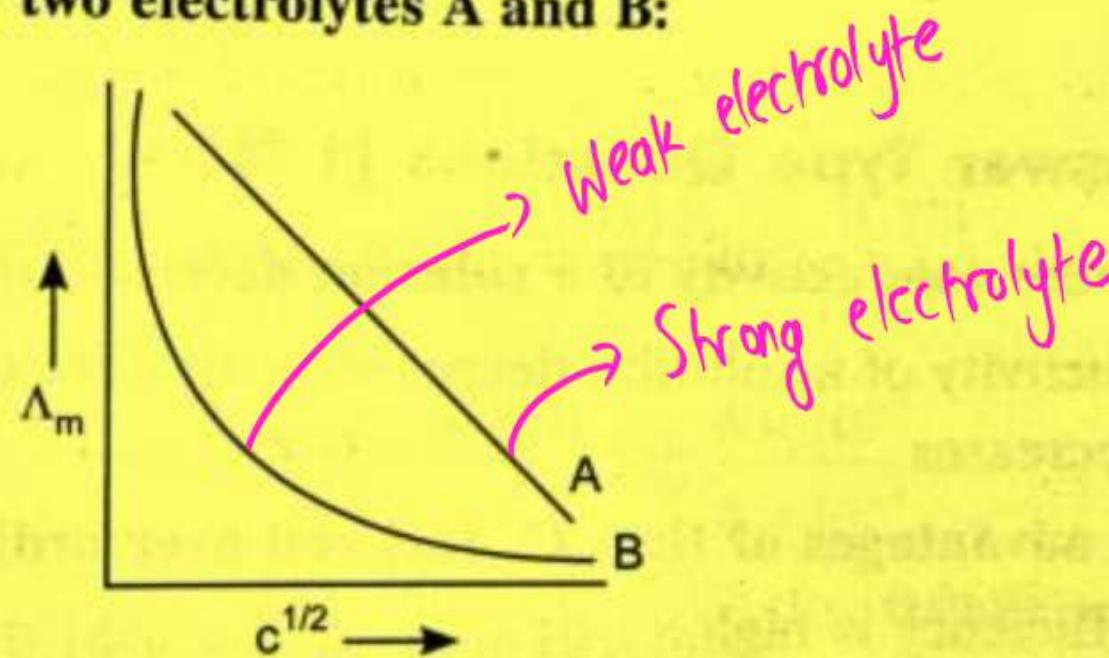
Λ_m° (NaCl) = 126 S cm² mol⁻¹

Λ_m° (CH₃COONa) = 91 S cm² mol⁻¹

[Delhi]

Slide 19

In the plot of molar conductivity (Λ_m) vs square root of concentration ($c^{1/2}$), following curves are obtained for two electrolytes A and B:

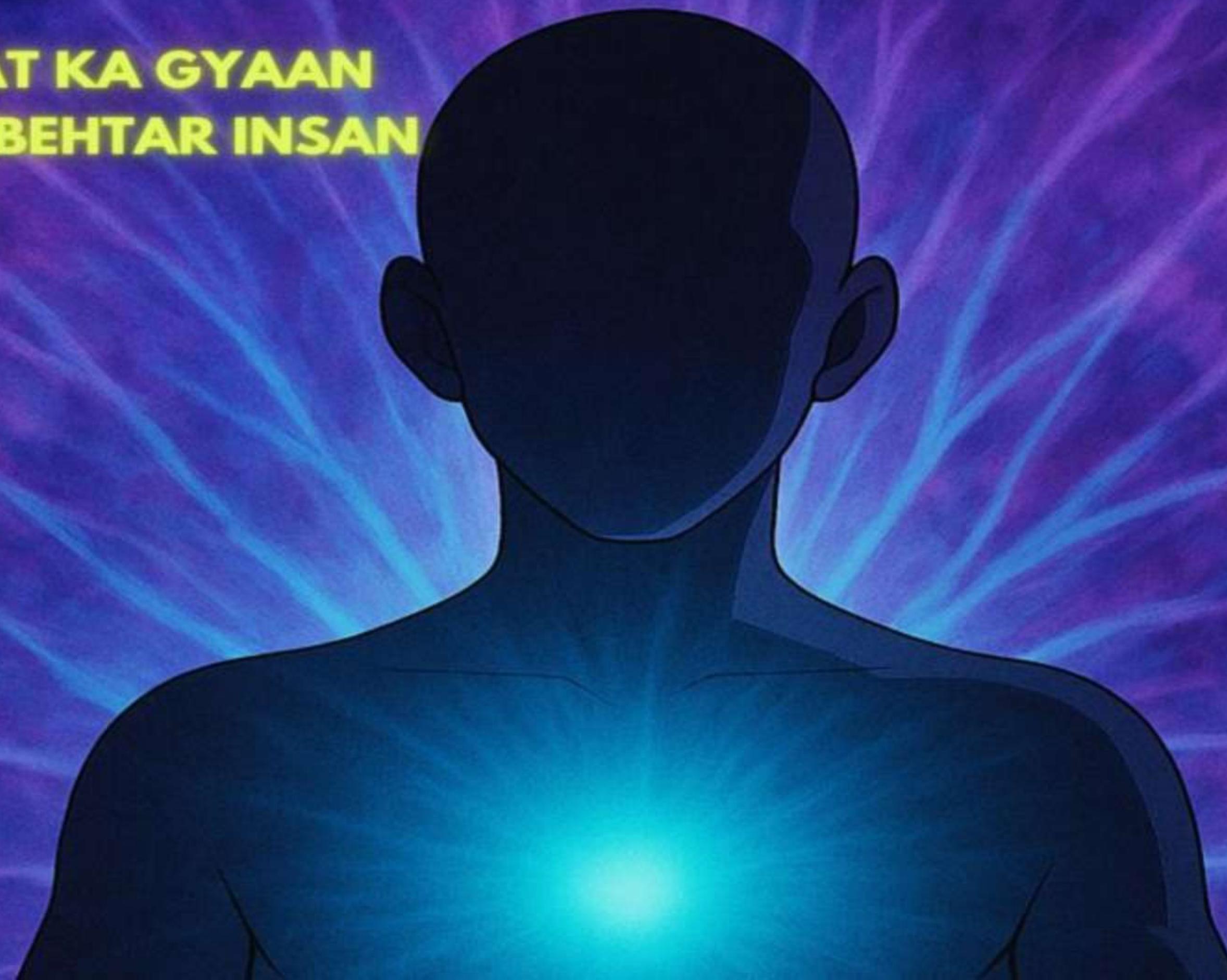


Answer the following:

- (i) Predict the nature of electrolytes A and B.
- (ii) What happens on extrapolation of Λ_m to concentration approaching zero for electrolytes A and B?

[Delhi]

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CONCEPT POLISH – HOMEWORK



NCERT SOLVED EXAMPLE 2.9

The conductivity of $0.001028 \text{ mol L}^{-1}$ acetic acid is $4.95 \times 10^{-5} \text{ S cm}^{-1}$. Calculate its dissociation constant if Λ_m^0 for acetic acid is $390.5 \text{ S cm}^2 \text{ mol}^{-1}$.



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#sbsathhai ✓

#pwsathhai ✓

Thank
You

PARISHRAM



2026

CHEMISTRY

Lecture 09

ELECTROCHEMISTRY

Introduction to Electrolysis and
Faraday's Law

Bharat Mata
Ki Jai ♥



BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

- (i) Introduction to Electrolysis (✓)
- (ii) Basic Electrolytic Process – Class 10 Chemistry (✓)
- (iii) Quantitative Aspects of Electrolysis – Faraday's Law (✓)
 - ↳(PQs & Numerical)



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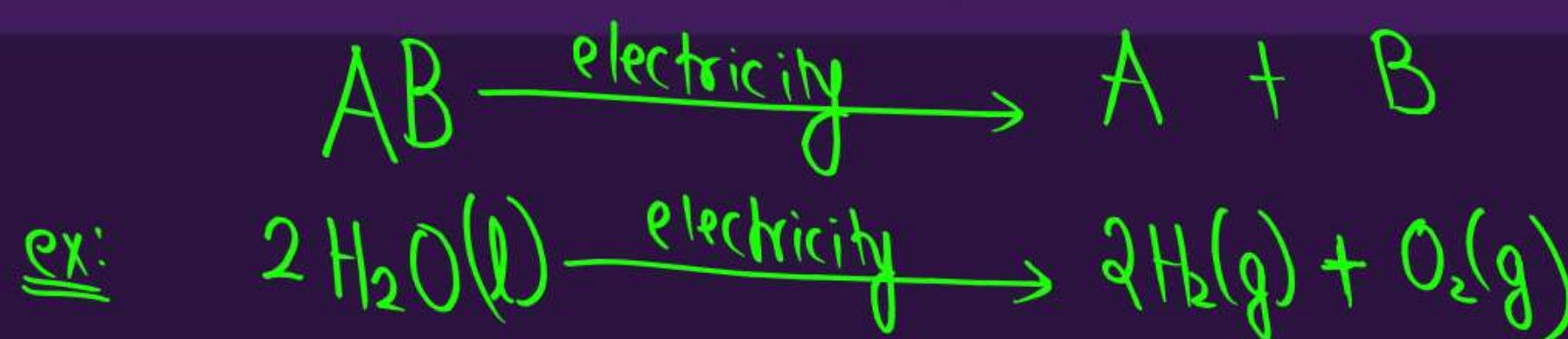


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INTRODUCTION TO ELECTROLYSIS

MEANING OF ELECTROLYSIS

- The word electrolysis is made up of two words 'electro' meaning flow of electrons or electricity and 'lysis' meaning breakdown or separating, i.e., bringing about a (chemical) change in a substance by passage of electricity.
- It is a type of decomposition reaction in which energy is supplied to a compound in the form of electricity. It is also called electrolytic decomposition.



GIVE A THOUGHT

Do all substances undergo a chemical change when electric current passes through them? or chemical rxn

- A. Yes
- B. No

GIVE A THOUGHT

Do all substances undergo a chemical change when electric current passes through them?

- A. Yes
- B. No

No!

- Metals, alloys and some non-metals like graphite (form of carbon) conducts electricity because of presence of free/delocalized electrons.
- ✓ It is not a chemical change but a physical property.

GIVE A THOUGHT

(Ionic compounds) or (electrovalent, ^{Compounds} conduct electricity in aqueous or molten form but not in solid form.

- A. Yes
- B. No

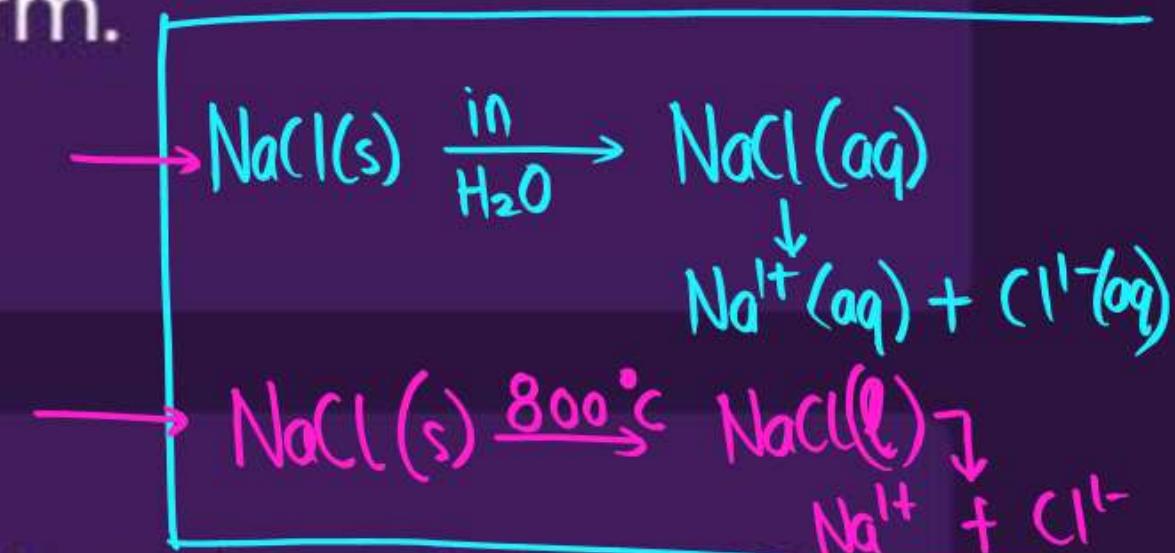
GIVE A THOUGHT

(Ionic compounds) or (electrovalent compounds) conduct electricity in aqueous or molten form but not in solid form.

- A. Yes
- B. No

Yes!

ex:



Ions are free to move in the aqueous or molten form but not in the solid form which is why they don't conduct electricity in the solid form.

GIVE A THOUGHT

Polar covalent compounds like HCl(aq) conducts electricity.

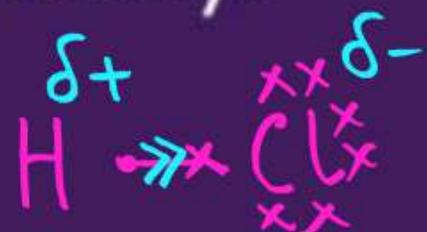
- A. Yes
- B. No

↓
(Hydrochloric acid)

GIVE A THOUGHT

Polar covalent compounds like HCl(aq) conducts electricity.

- A. Yes
- B. No



Yes!

They dissociate into ions in aqueous form which conducts electricity.

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LIKH DO.
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MEANING OF ELECTROLYTE

- Electrolytes are compounds which either in aqueous solution or in molten state allow electric current to pass through them by breaking into ions.
- Electrolytes can be electrovalent compounds, composed of cation and anion, or polar covalent compounds, that form ions in water.
- Faraday classifies electrolytes into two types: (i) **Strong electrolyte** (ii) **Weak electrolyte**

TYPES OF ELECTROLYTE

STRONG ELECTROLYTE	WEAK ELECTROLYTE
(i) They completely dissociate into ions in aqueous or molten state.	(i) They partially dissociate into ions in aqueous or molten state.
(ii) Contains only ions.	(ii) Contains both ions and molecules.
(iii) Strong acids, strong bases and salts	(iii) Weak acids and bases
(iii) Degree of dissociation (α) is 1.	(iii) Degree of dissociation (α) is: $0 < \alpha < 1$
EXAMPLE: KCl, HCl, MgSO ₄ etc.	EXAMPLE: CH ₃ COOH, NH ₄ OH etc.

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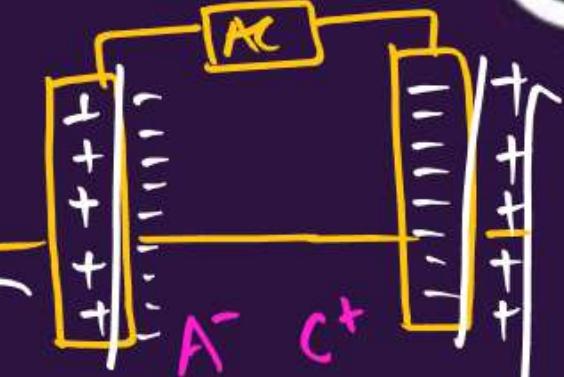


BASIC ELECTROLYTIC PROCESS –

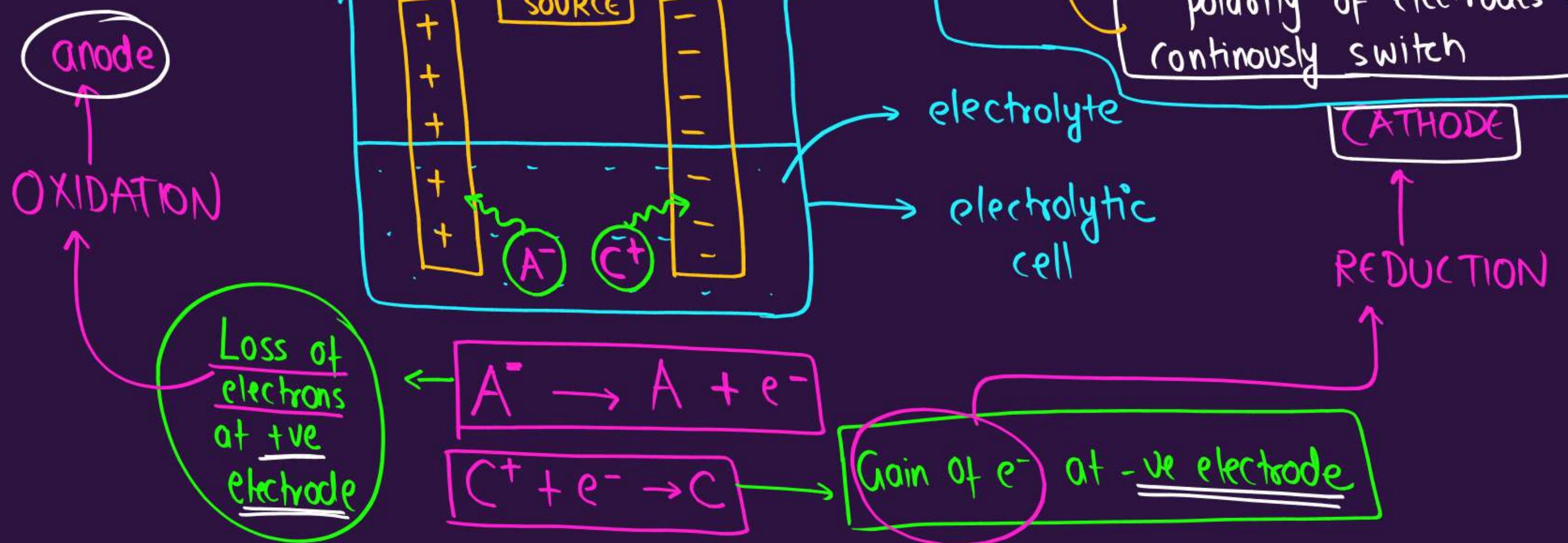
[CLASS 10 CHEMISTRY]

↳ (metals & non-metals)

Why AC. Power
is not used?

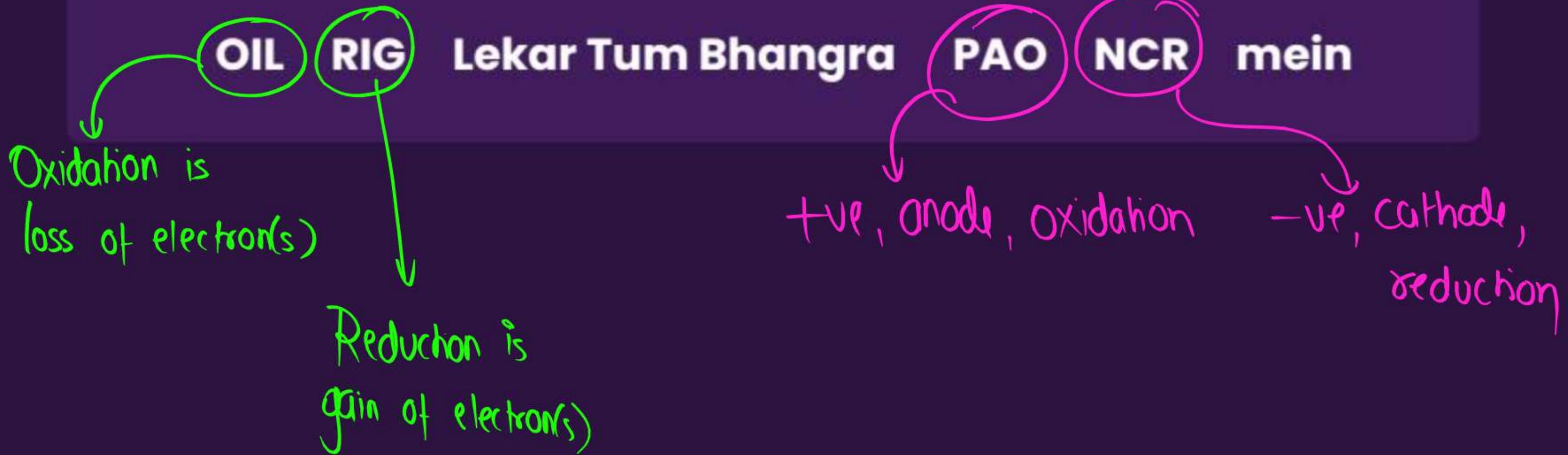


No electrolysis
polarity of electrodes will
continuously switch



TRICK TO REMEMBER

SUNIL DANCE
CLASSES



GIVE A THOUGHT

The power source for electrolysis to happen should always be DC.

- A. Yes
- B. No

GIVE A THOUGHT

The power source for electrolysis to happen should always be DC.

- A. Yes
- B. No

Yes!

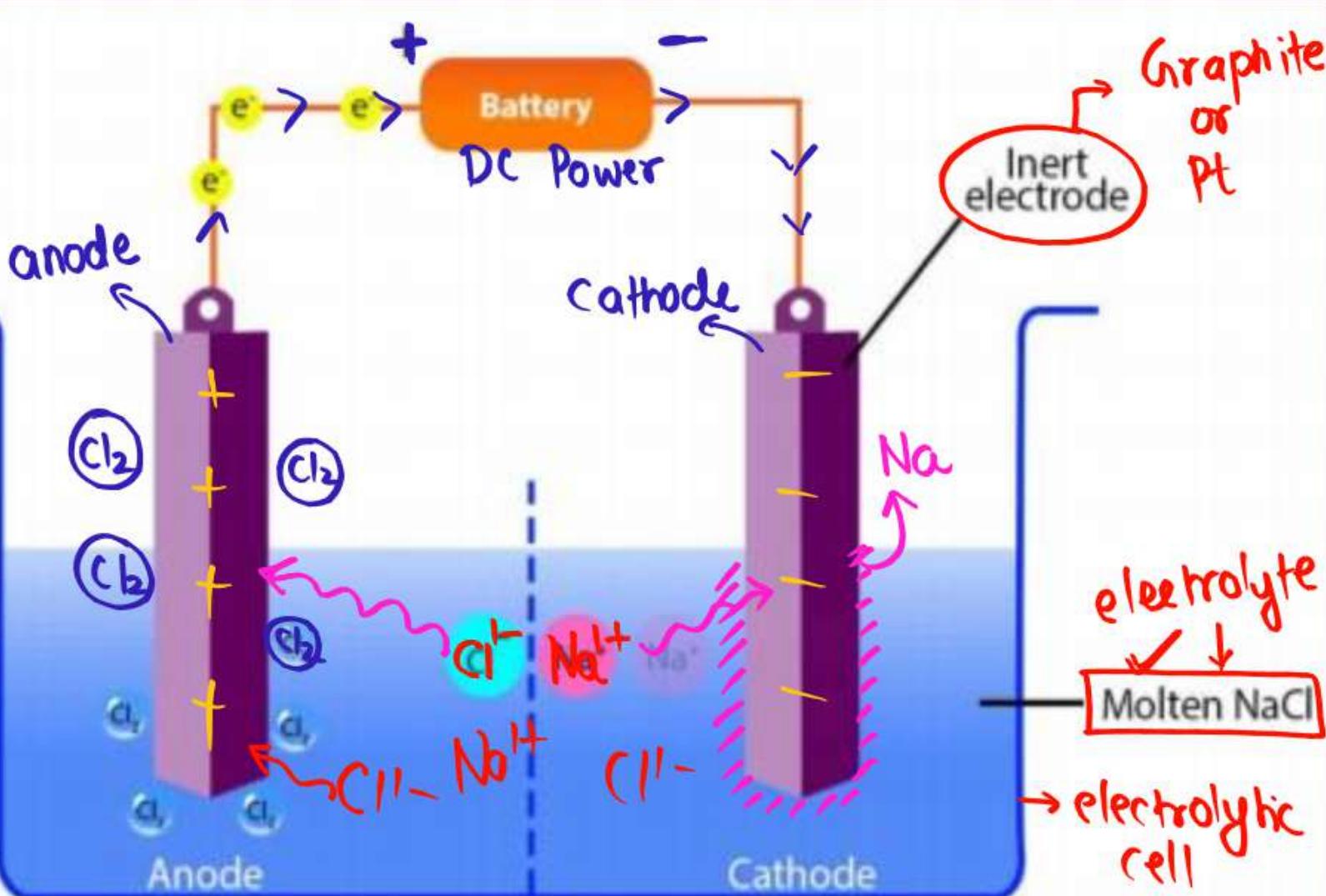
(Explanation on slide 17)

Electrolysis requires a direct current (DC) source because it provides constant polarity to electrodes, allowing ions to move in a fixed direction.

Alternating current (AC) cannot be used, as the changing polarity of electrodes prevents proper deposition of substances at the electrodes.

ELECTROLYTIC REDUCTION OF NaCl

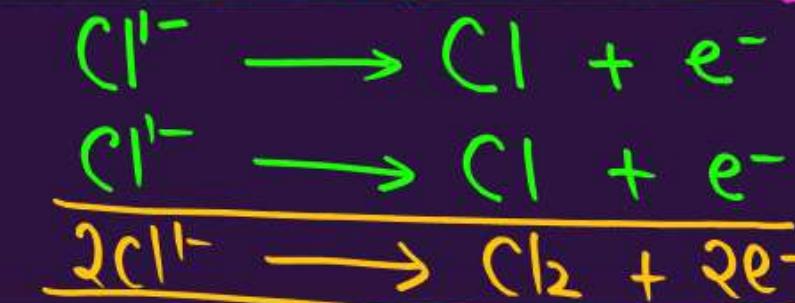
↳ (used for oxides & chlorides of K, Na, Co, Mg & Al to get pure metal)



- Molten/Fused NaCl: Strong Electrolyte
- $\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$
- Inert Electrode: Graphite or Platinum

$\text{Cl}^- \rightarrow$ Nascent chlorine
highly reactive

Reaction at anode (Oxidation: Loss of electrons)



Reaction at cathode (Reduction: Gain of electrons)



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**(QUANTITATIVE ASPECTS OF
ELECTROLYSIS – FARADAY'S
LAW)**

FARADAY'S FIRST LAW OF ELECTROLYSIS

The mass (m) of a substance deposited or liberated at an electrode is directly proportional to the electric charge (Q) passed through the electrolyte.

$$m \propto Q$$

$$m = ZQ ; (Q = IT)$$

$$m = ZIT$$

(electrochemical equivalent)

$$Z = \frac{\text{Eq. wt}}{\text{Faraday's constant}} = \frac{E}{F}$$

charge on 1 mol e⁻

≈ 96500 C/mol

CONCEPT RECAP – EQUIVALENT WEIGHT OF IONS

Equivalent weight of an ion is the mass of the ion in grams which will combine with or displace 1 mole of monovalent ions (like H^+ or e^-), i.e., carry or neutralize one mole of charge

$$\text{Equivalent Weight of Ion} = \frac{\text{Molar Mass}}{\text{Valency Factor}} \quad (\text{where Valency Factor} = {}^6\text{Magnitude of charge on ion})$$

Ion	Valency Factor (n)	Equivalent Weight (E) $\left(\frac{\text{g}}{\text{mol}}\right)$
Al^{3+}	3	$Al^{3+} \rightarrow 27 \frac{\text{g}}{\text{mol}}$ $\frac{27}{3} = 9 \frac{\text{g}}{\text{mol}}$
Na^{1+}	1	$Na^{1+} \rightarrow 23 \frac{\text{g}}{\text{mol}}$ $\frac{23}{1} = 23 \frac{\text{g}}{\text{mol}}$
Cl^{1-}	1	$Cl^{1-} \rightarrow 35.5 \frac{\text{g}}{\text{mol}}$ $\frac{35.5}{1} = 35.5 \frac{\text{g}}{\text{mol}}$

What is eq. wt. of Mg^{2+} ?

- (A) 24
- (B) 12
- (C) 48
- (D) 13

$$\text{Eq. wt. of } Mg^{2+} = \frac{\text{Molar mass of } Mg^{2+}}{\text{Valency factor}}$$

$$= \frac{\cancel{12} \text{ g/mol}}{\cancel{2}}$$
$$= 12 \frac{\text{g}}{\text{mol}}$$

What is unit of Z (electrochemical equivalent)?

① $Z = \frac{E}{F} = \frac{\frac{g}{mol}}{\frac{C}{mol}} = \cancel{\frac{g}{mol}} \times \cancel{\frac{mol}{C}}$

$= \frac{g}{C}$

② $m = Z \times Q$

$$\begin{aligned} &= \frac{g}{C} \times C \\ &= \underline{\underline{g}} \end{aligned}$$

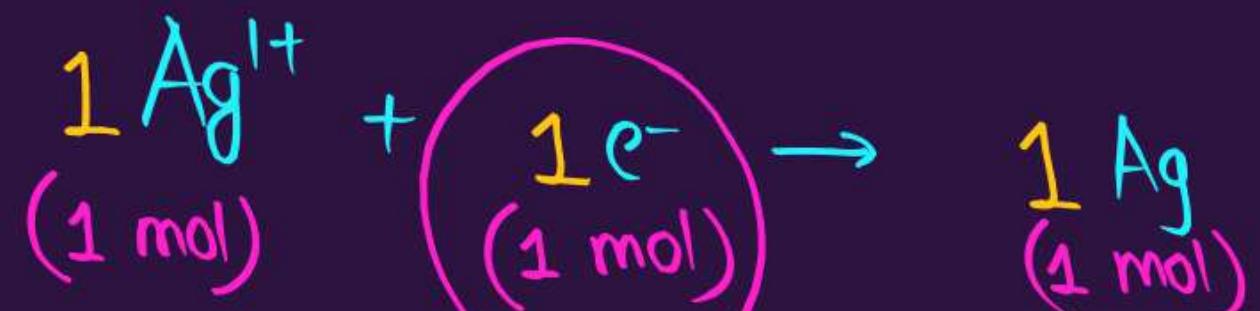
LET'S PRACTICE



How much charge in Faraday is required for the reduction of 1 mol of Ag^+ to Ag?



↓ On molar level

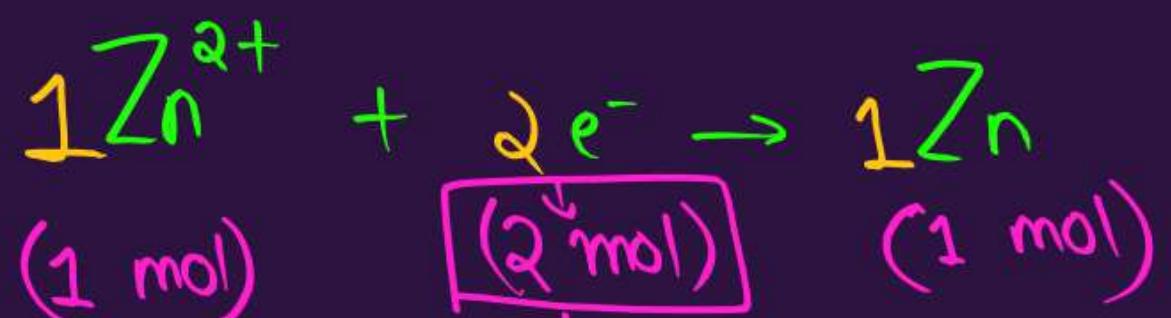


$$1F = 96,500 \frac{\text{C}}{\text{mol}}$$

How much charge in Faraday is required for the reduction of 1 mol of Zn²⁺ to Zn?



↓ On molar level

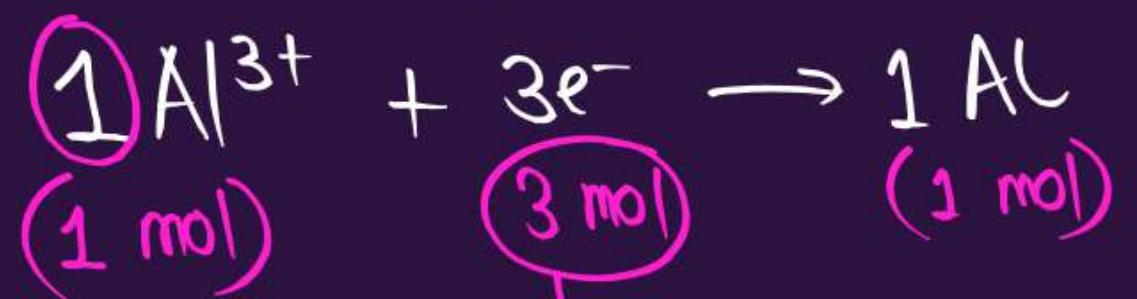


$$Q_F = 2 \times 96500 \frac{\text{C}}{\text{mol}}$$

How much charge in Faraday is required for the reduction of 1 mol of Al³⁺ to Al?



↓ at molar level



$$3F = 3 \times 96500 \frac{\text{C}}{\text{mol}}$$

QUESTION



Find the amount of copper deposited when 5 A current is used for 3600 seconds to electrolyte (CuSO_4 solution).

Atomic mass of Cu is 63.5 g



Formula Approach

$$m = ZIT$$

$$= \frac{E}{F} \times I \times T$$

$$= \frac{63.5}{2} \times 5 \times 3600$$

$$= \frac{63.5 \times 1}{2} \times 5 \times 3600$$

→ m = ?

$$m = ? , \quad I = 5 \text{ A} , \quad t = 3600 \text{ s} \\ F = 96500 \frac{\text{C}}{\text{mol}}$$

$$\text{Eq. wt of } \text{Cu}^{2+} (\text{E}) = \frac{\text{Molar mass}}{\text{Valency factor}}$$

$$= \frac{63.5 \text{ g}}{\text{mol}}$$

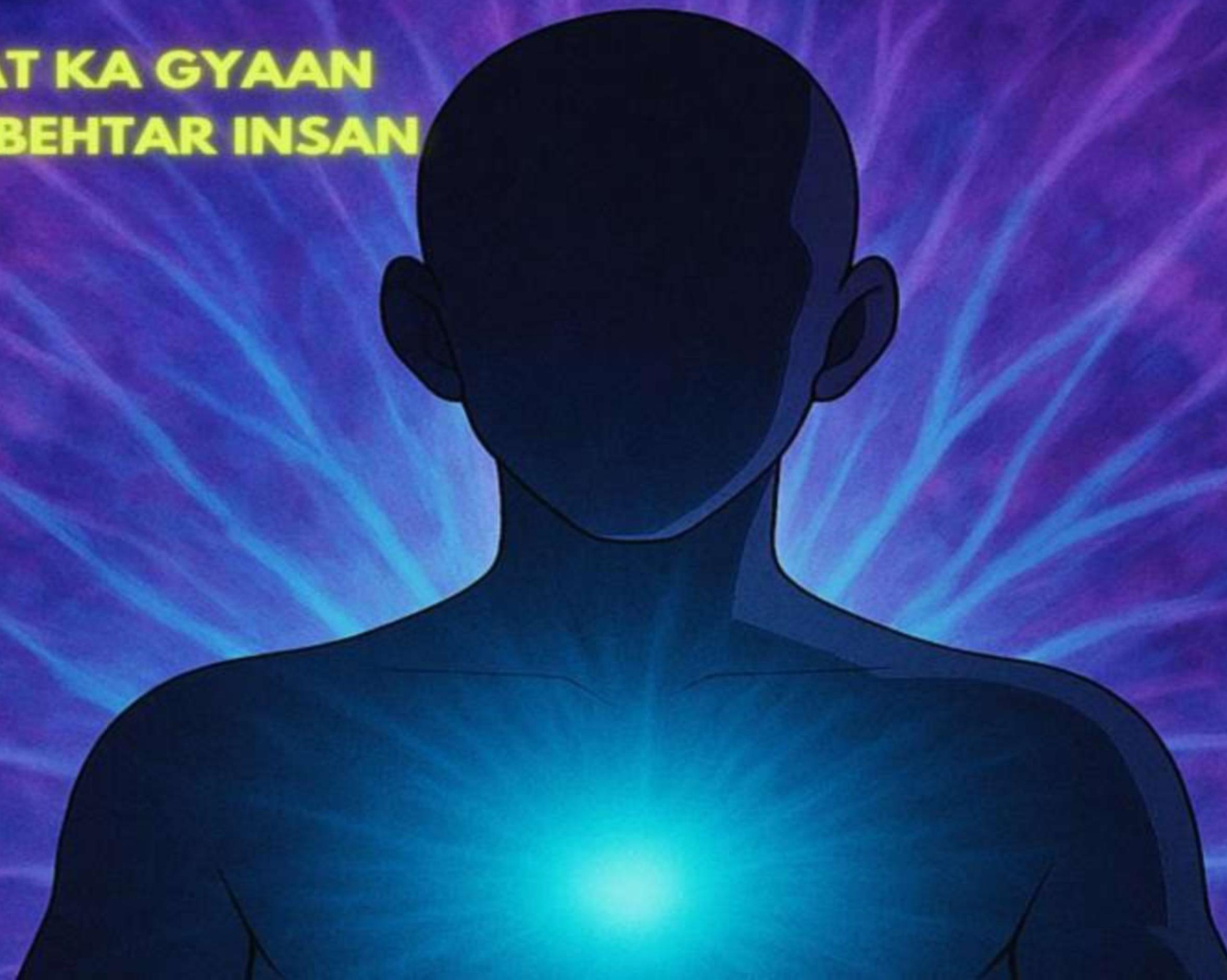
$$= \frac{63.5 \times 18}{193} \approx \frac{63 \times 18}{200} \approx \frac{54}{10} \approx 5.4 \text{ g}$$

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CONCEPT POLISH – HOMEWORK



QUESTION

Find the amount of copper deposited when 5 A current is used for 3600 seconds to electrolyte (CuSO_4 solution).

Atomic mass of Cu is 63.5 g

Concept Approach - 'UNITARY METHOD' → 'TRY'

Calculate the time to deposit 1.5 g of silver at cathode when a current of 1.5 A was passed through the solution of AgNO_3 .

Molar mass of Ag is 108 g/mol

TRY

CBSE PYQ

Calculate the time to deposit 1.17 g of Ni at cathode when a current of 5 A was passed through the solution of $\text{Ni}(\text{NO}_3)_2$.

Molar mass of Ag is 58.5 g/mol

TRY



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#sbsathhai (✓)

#pwsathhai (✓)

Thank
You

PARISHRAM



2026

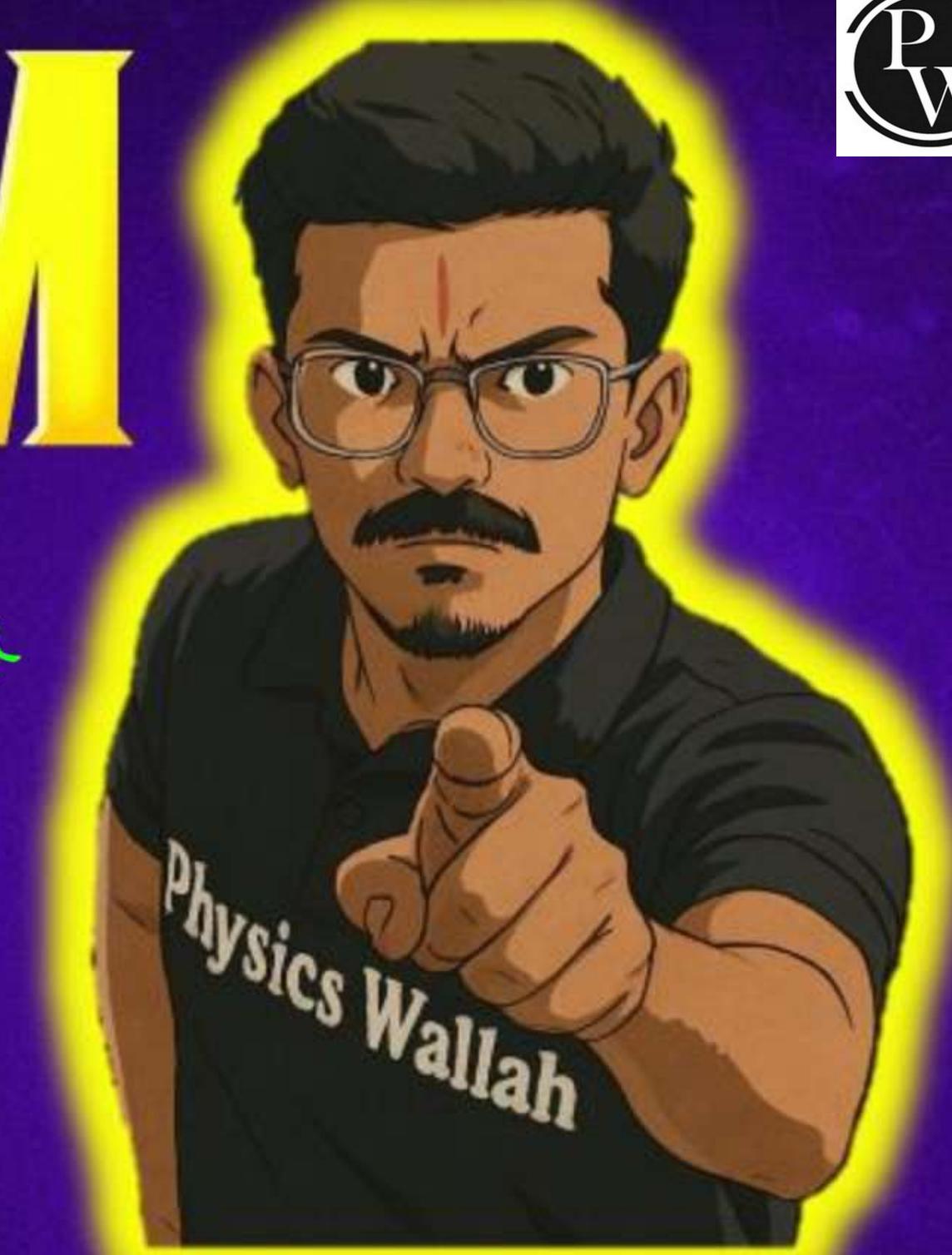
CHEMISTRY

Lecture 10

ELECTROCHEMISTRY

Faraday's Laws (Contd.) and Products
of Electrolysis

Bharat Mata
Ki Jai O



BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

- (i) Faraday's Law (Contd.) ✓
- (ii) Electrolysis Par Soch Vichar ✓
- (iii) Qualitative Analysis - Products of Electrolysis ✓



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CONCEPT POLISH – HOMEWORK DISCUSSION



QUESTION

Find the amount of copper deposited when 5 A current is used for 3600 seconds to electrolyte (CuSO_4 solution).

Atomic mass of Cu is 63.5 g u

$$Q = I \times T \\ = 5 \times 3600 \text{ C}$$

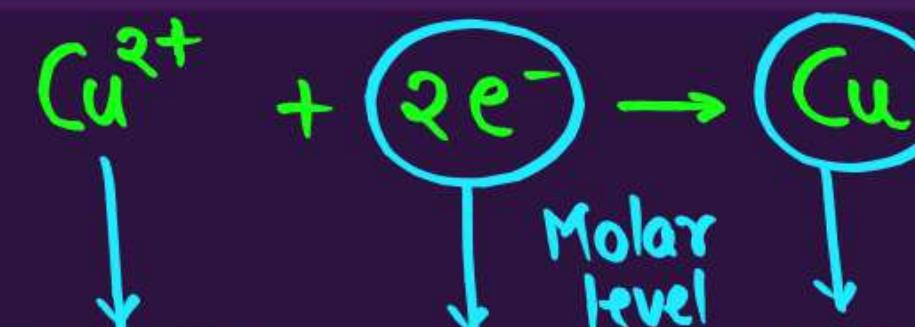
✓
Concept Approach → (Unitary method)

$$2 \times 96500 \text{ C} = 63.5 \text{ g}$$

$$5 \times 3600 \text{ C} = 'x'$$

$$\Rightarrow x \times 2 \times 96500 = 63.5 \times 5 \times 3600$$

$$\Rightarrow x = \frac{63.5 \times 5 \times 3600}{2 \times 96500} \approx \frac{63.5 \times 18}{200} \approx 5.4 \text{ g}$$



$$2 F = 63.5 \frac{\text{g}}{\text{mol}}$$

$$2 \times 96500 \frac{\text{C}}{\text{mol}} = 63.5 \frac{\text{g}}{\text{mol}}$$

CBSE PYQ ✓

Calculate the time to deposit 1.5 g of silver at cathode when a current of 1.5 A was passed through the solution of AgNO_3 .

Molar mass of Ag is 108 g/mol

Formula Approach $m = ZIT \Rightarrow m = \frac{E}{F} \times I \times T \Rightarrow T = \frac{m \times F}{E \times I}$

$$T = \frac{1.5 \times 96500}{E \times 1.5}$$

$$= \frac{96500}{108} \approx \frac{96500}{110}$$

$$\approx 877.3 \text{ s}$$

$$E = \frac{\text{Molar Mass}}{\text{Valency factor}}$$



$$= \frac{108}{1}$$

$$= 108 \text{ g/mol}$$

CBSE PYQ

Calculate the time to deposit 1.17 g of Ni at cathode when a current of 5 A was passed through the solution of $\text{Ni}(\text{NO}_3)_2$.

Molar mass of ~~Ag~~^{Ni} is 58.5 g/mol

Formula Approach

$$\begin{aligned} T &= \frac{m \times F}{E \times I} = \frac{1.17 \times 96500}{\frac{58.5}{2} \times 5} \\ &= \frac{1.17 \times 96500 \times 2 \times 10^3}{58.5 \times 5 \times 10^3} \\ &= 193 \times 4 \\ &= 772 \text{ s} \end{aligned}$$



$$E = \frac{58.5}{2}$$

SAMAJ AAYA TOH
LIKH DO.

AYE BHAIYA ✓





A large, yellow, rounded rectangular button centered on a dark purple background. It contains the title text.

FARADAY'S LAWS (CONTD.)

FARADAY'S SECOND LAW OF ELECTROLYSIS

$$Q = IT$$

→ same $Q \rightarrow$ (same current for some time)

When the same quantity of electricity is passed through
different electrolytic solutions connected in series, the
masses of substances liberated or deposited at the
electrodes are directly proportional to their equivalent
weights.)

$$m = ZIT$$

$$m = \frac{E}{F} IT$$

→ constant

$$m \propto E$$

$$m_1 = \frac{E_1}{F} IT$$

$$\frac{m_1}{E_1} = \frac{IT}{F} \quad \textcircled{1}$$

$$m_2 = \frac{E_2}{F} IT$$

$$\frac{m_2}{E_2} = \frac{IT}{F} \quad \textcircled{2}$$

From $\textcircled{1} \& \textcircled{2}$

$$\frac{m_1}{E_1} = \frac{m_2}{E_2}$$

GIVE A THOUGHT



same current will flow

Suppose 3 electrolytic solutions AgNO_3 , CuSO_4 and $\text{Al}(\text{NO}_3)_3$ are connected in series and current passes through all of them for same time 't' seconds. In which case, more mass will be deposited if the molar masses in g/mol are $\text{Ag} = 108$, $\text{Cu} = 63.5$, $\text{Al} = 27$.

$$m = ZIT$$

$$m = E \times IT$$

$$m \propto E$$

(jiska E(eq.wt) jyda uska mass(m) jyda)

$$\text{Eq.wt}(\text{Ag}^{2+}) = \frac{108}{1} = 108 \text{ g/mol}$$

$$\text{Eq.wt.}(\text{Cu}^{2+}) = \frac{63.5}{2} = 31.75 \text{ g/mol}$$

$$\text{Eq.wt.}(\text{Al}^{3+}) = \frac{27}{3} = 9 \text{ g/mol}$$

LET'S PRACTICE



QUESTION

Same current is passed through solutions of AgNO_3 and CuSO_4 .
If 10.8 g of silver (Ag) is deposited, how much copper (Cu) will be deposited?

Molar mass of Ag is 108 g/mol and 63.5 g/mol

$$\frac{m_1}{E_1} = \frac{m_2}{E_2}$$

$$\frac{10.8}{108 \times 10} = \frac{x}{31.75}$$

$$\Rightarrow \frac{31.75}{10} = x$$

$$\Rightarrow \boxed{3.175 \text{ g}}$$

$$\text{Eq.wt}(\text{Ag}^{+}) = \frac{108}{1} = 108 \text{ g/mol}$$

$$\text{Eq.wt}(\text{Cu}^{2+}) = \frac{63.5}{2} = 31.75 \text{ g/mol}$$

**SAMAJ AAYA TOH
LIKH DO.
AYE BHAIYA**





**ELECTROLYSIS PAR SOCH-
VICHAR**

ELECTROLYSIS PAR SOCH-VICHAR

Electrolysis is a non-spontaneous reaction. /non-spontaneous process.

- A. Yes
- B. No



This doesn't happens on its own.



ELECTROLYSIS PAR SOCH-VICHAR

In an electrolytic cell, electrical energy converts into chemical energy.

- A. Yes
- B. No

ELECTROLYSIS PAR SOCH-VICHAR

$\Delta G > 0$ for non-spontaneous reactions.

- A. Yes
- B. No

$\Delta G > 0$, $\Delta G \rightarrow +ve$, Non-spontaneous Rxn

$\Delta G < 0$, $\Delta G \rightarrow -ve$, Spontaneous Rxn

ELECTROLYSIS PAR SOCH-VICHAR

If I am charging a phone, $\Delta G > 0$ but after removing the charger and I am using phone, the battery gets discharged and during discharging $\Delta G < 0 \rightarrow \Delta G \rightarrow -Vc$

- A. Yes
- B. No

ELECTROLYSIS PAR SOCH-VICHAR

- ✓ An electrolytic cell **requires external energy** (usually from a battery) to drive a non-spontaneous chemical reaction.
- ✓ Since the reaction is **non-spontaneous**, it does **not occur on its own** – energy is **supplied to the system** to force the reaction to proceed.
- ✓ For non-spontaneous reactions, by definition: $\Delta G > 0$

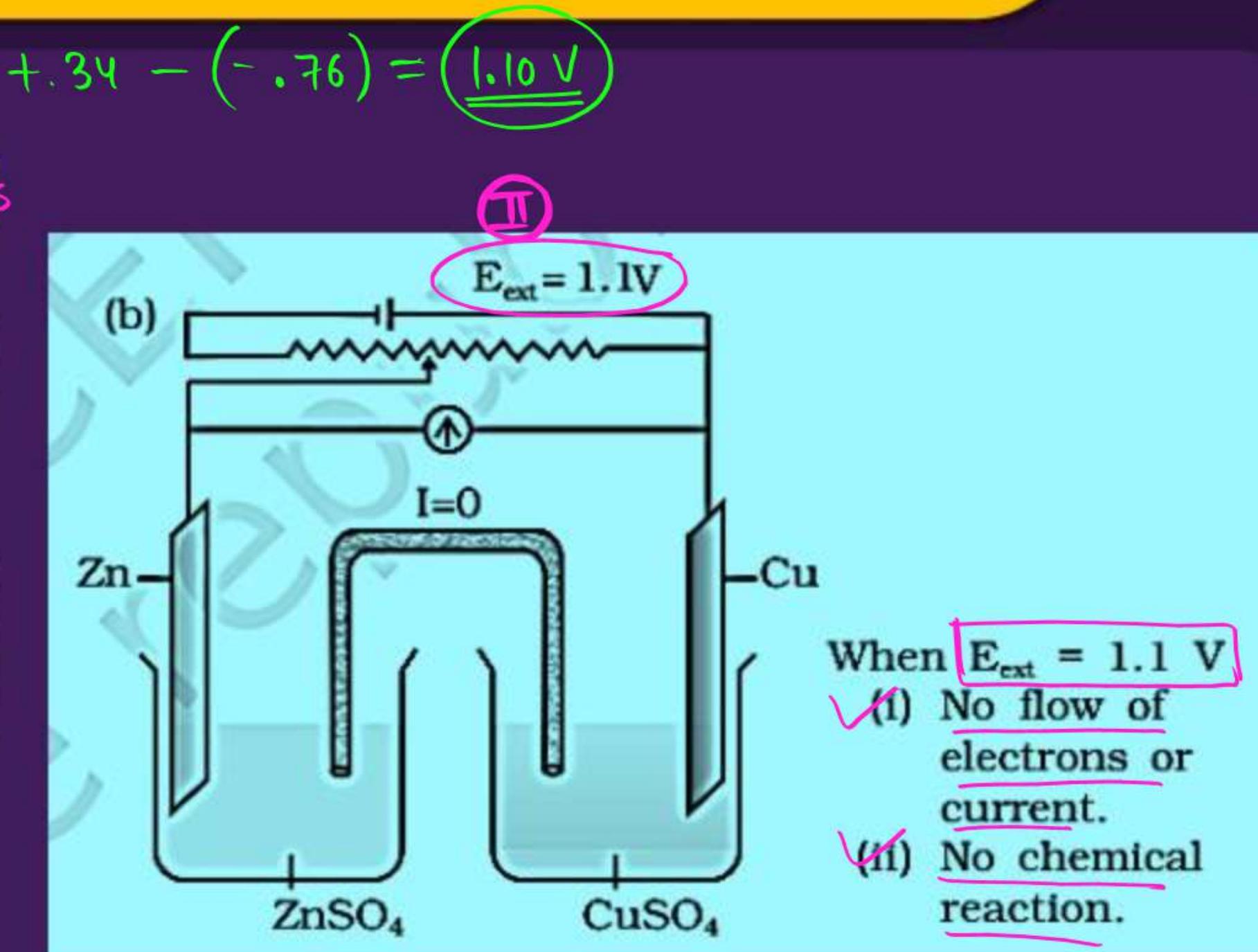
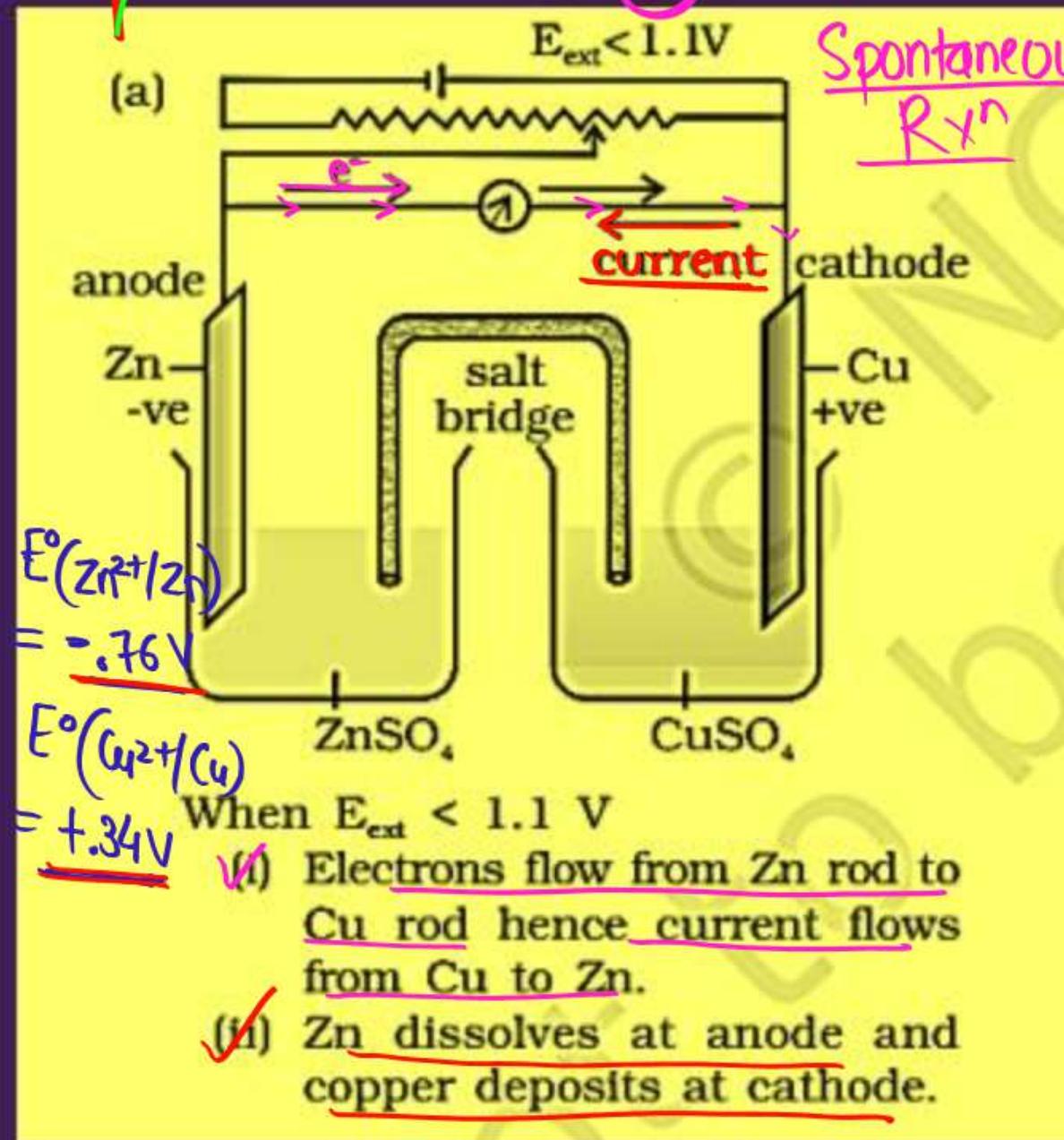
Student (X) Aashik (✓)

Daniel Cell

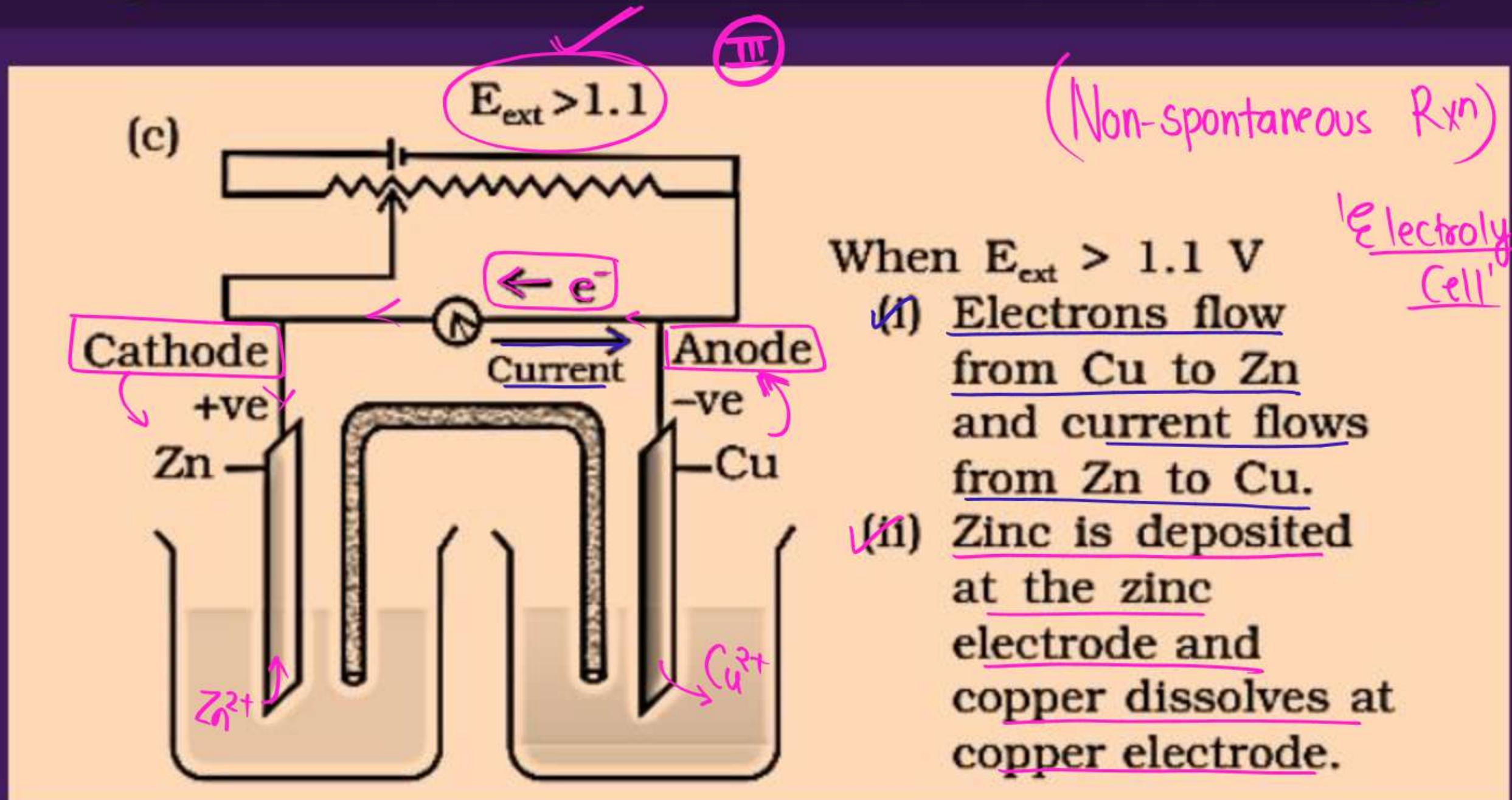
(Type of Galvanic Cell)

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = +.34 - (-.76) = \underline{\underline{1.10 \text{ V}}}$$

(I)



ELECTROLYSIS PAR SOCH-VICHAR



SUMMARY

Daniel Cell (Galvanic Cell)

External Voltage (E_{ext})	Electron Flow	Current Flow	Oxidation	Reduction	Reaction
$E_{ext} < 1.1 \text{ V}$	$\text{Zn} \rightarrow \text{Cu}$	$\text{Cu} \rightarrow \text{Zn}$	$\text{Zn} \text{ (anode)} \rightarrow \text{Zn}^{2+}$	$\text{Cu}^{2+} \rightarrow \text{Cu} \text{ (cathode)}$	<u>Zn dissolves, Cu deposits</u>
$E_{ext} = 1.1 \text{ V}$	No flow ✓	No flow ✓	None ✓	None ✓	No reaction ✓
$E_{ext} > 1.1 \text{ V}$	$\text{Cu} \rightarrow \text{Zn}$	$\text{Zn} \rightarrow \text{Cu}$	$\text{Cu} \text{ (anode)} \rightarrow \text{Cu}^{2+}$	$\text{Zn}^{2+} \rightarrow \text{Zn} \text{ (cathode)}$	<u>Cu dissolves, Zn deposits</u>

Electrolytic Cell

PYQS' WALLAH



What happens when applied external potential becomes greater than E_{cell}° of electrochemical cell?

When $E_{ext} > E_{cell}^{\circ}$, then external voltage overpowers the natural tendency of the cell reaction. It reverses the direction of electron flow as previous **cathode becomes anode**, and **anode becomes cathode**.

The cell now behaves as an electrolytic cell.

SAMAJ AAYA TOH
LIKH DO.

AYE BHAIYA





QUALITATIVE ANALYSIS – **PRODUCTS OF ELECTROLYSIS**

PRODUCTS OF ELECTROLYSIS

The products formed during electrolysis of an electrolyte depends on:

- (I) • Nature of Electrolyte (Molten or Aqueous)
*पिघला हुआ है।
पानी में घुला हुआ है।*
- (II) • Nature of Electrode (Active electrode or Inert electrode)

The electrodes that participate in the redox reactions and get oxidized/reduced.

(Their mass may change due to deposition or dissolution.)

Ex: Cu, Zn etc.

The electrodes that don't participate in the redox reactions and act as a surface of electron transfer only. ex: Pt & Graphite

PRODUCTS OF ELECTROLYSIS

The products formed during electrolysis of an electrolyte depends on:

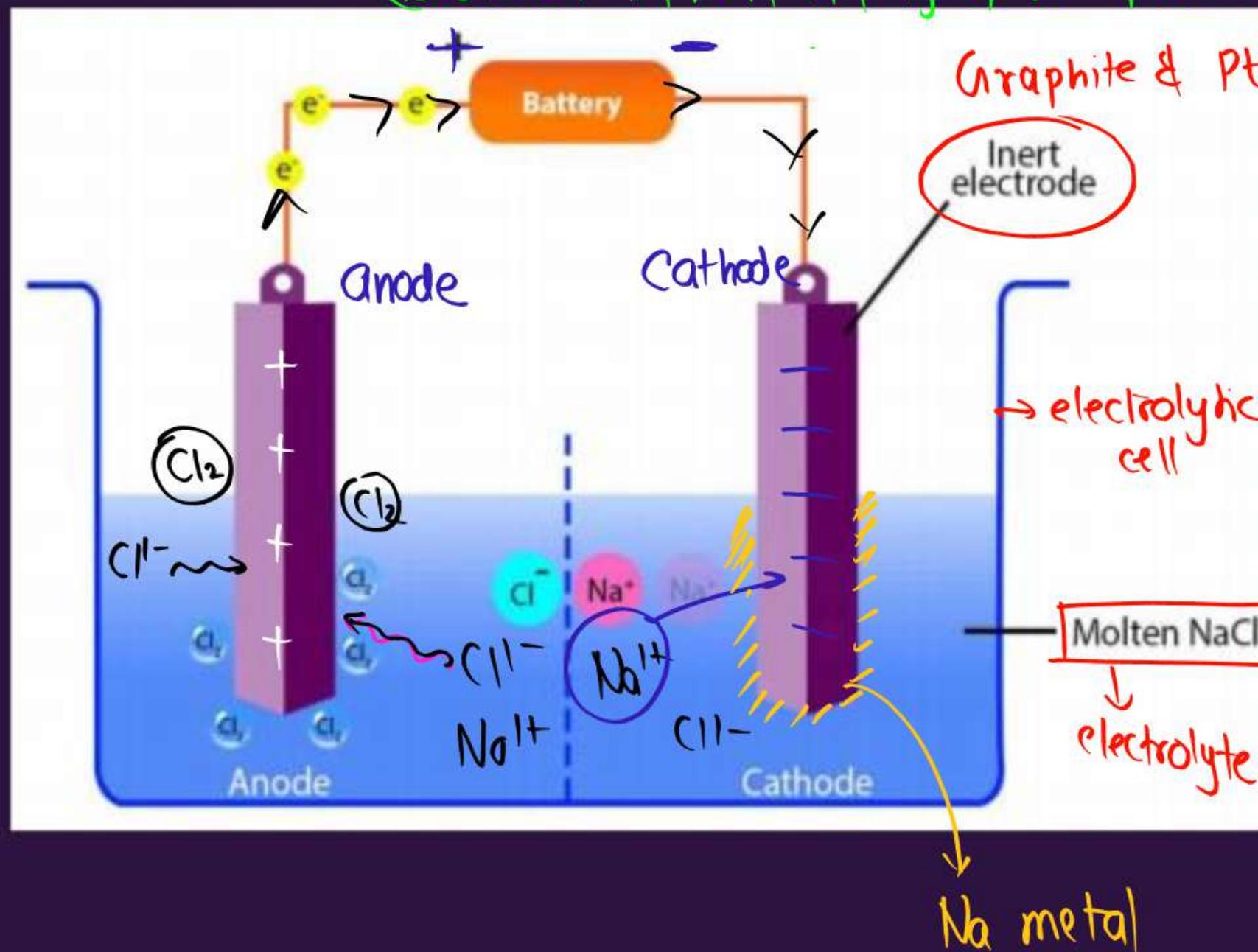
- III • Position in electrochemical series
- IV • Concentration of ions ✓
- I • Overpotential (extra voltage required above the theoretical value to make a redox reaction actually happen on an electrode in real conditions.)

Which 1 cation will move to cathode & which 1 anion will move to anode?
More than 1 cation is present or more than 1 anion is present.
Depends on III, IV & I

↳ because these rxn have a kinetic hindrance
↓
VERY SLOW RXN

ELECTROLYSIS OF [MOLTEN NaCl] - Recap

(to extract K, Na, Ca, Mg & Al from their oxides & chlorides)



- Molten/Fused NaCl: Strong Electrolyte
- $\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$
- Inert Electrode: Graphite or Platinum

Reaction (Oxidised form + ne ⁻	→ Reduced form)	E^\ominus / V
$F_2(g) + 2e^-$	$\rightarrow 2F^-$	+2.87
$Co^{3+} + e^-$	$\rightarrow Co^{2+}$	+1.81
$H_2O_2 + 2H^+ + 2e^-$	$\rightarrow 2H_2O$	+1.78
$MnO_4^- + 8H^+ + 5e^-$	$\rightarrow Mn^{2+} + 4H_2O$	+1.51
$Au^{3+} + 3e^-$	$\rightarrow Au(s)$	+1.40
$Cl_2(g) + 2e^-$	$\rightarrow 2Cl^-$	+1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightarrow 2Cr^{3+} + 7H_2O$	+1.33
$O_2(g) + 4H^+ + 4e^-$	$\rightarrow 2H_2O$	+1.23
$MnO_2(s) + 4H^+ + 2e^-$	$\rightarrow Mn^{2+} + 2H_2O$	+1.23
$Br_2 + 2e^-$	$\rightarrow 2Br^-$	+1.09
$NO_3^- + 4H^+ + 3e^-$	$\rightarrow NO(g) + 2H_2O$	+0.97
$2Hg^{2+} + 2e^-$	$\rightarrow Hg_2^{2+}$	+0.92
$Ag^+ + e^-$	$\rightarrow Ag(s)$	+0.80
$Fe^{3+} + e^-$	$\rightarrow Fe^{2+}$	+0.77
$O_2(g) + 2H^+ + 2e^-$	$\rightarrow H_2O_2$	+0.68
$I_2(s) + 2e^-$	$\rightarrow 2I^-$	+0.54
$Cu^+ + e^-$ ✓	$\rightarrow Cu(s)$	+0.52
$Cu^{2+} + 2e^-$	$\rightarrow Cu(s)$	+0.34
$AgCl(s) + e^-$	$\rightarrow Ag(s) + Cl^-$	+0.22
$AgBr(s) + e^-$	$\rightarrow Ag(s) + Br^-$	+0.10
$2H^+ + 2e^-$	$\rightarrow H_2(g)$	0.00

(Standard reduction potential / Standard electrode potential)

Increasing strength of reducing agent

Recap ↴

- **Positive E°:** High tendency to get reduced (electrons lena)
- Reducing Agent: The one that reduces others (gain of electron on other species because it loses electrons).
- A positive E° means that the reduced form is a weaker reducing agent than H₂

NOTE: Standard electrode potential (E°) ki har value ko Standard Hydrogen Electrode (SHE) ke reference mein measure kia gaya hai.

Reaction (Oxidised form + $n e^-$	\rightarrow Reduced form)		E^\ominus / V
$2H^\circ + 2e^-$	$\rightarrow H_2(g)$	Increase ↓	0.00
$Pb^{2+} + 2e^-$	$\rightarrow Pb(s)$		-0.13
$Sn^{2+} + 2e^-$	$\rightarrow Sn(s)$		-0.14
$Ni^{2+} + 2e^-$	$\rightarrow Ni(s)$		-0.25
$Fe^{2+} + 2e^-$	$\rightarrow Fe(s)$		-0.44
$Cr^{3+} + 3e^-$	$\rightarrow Cr(s)$		-0.74
$Zn^{2+} + 2e^-$	$\rightarrow Zn(s)$		-0.76
$2H_2O + 2e^-$	$\rightarrow H_2(g) + 2OH^-$		-0.83
$Al^{3+} + 3e^-$	$\rightarrow Al(s)$		-1.66
$Mg^{2+} + 2e^-$	$\rightarrow Mg(s)$		-2.36
$Na^\circ + e^-$	$\rightarrow Na(s)$		-2.71
$Ca^{2+} + 2e^-$	$\rightarrow Ca(s)$		-2.87
$K^\circ + e^-$	$\rightarrow K(s)$		-2.93
$Li^\circ + e^-$	$\rightarrow Li(s)$		-3.05

- **Negative E^\ominus :** Low tendency to get reduced.

- Reducing Agent: The one that reduces others (gain of electron on other species because it loses electrons).
- A negative E^\ominus means that the reduced form is a stronger reducing agent than H_2 .

NOTE: Standard electrode potential (E^\ominus) ki har value ko Standard Hydrogen Electrode (SHE) ke reference mein measure kia gaya hai.

IMPORTANT TO UNDERSTAND

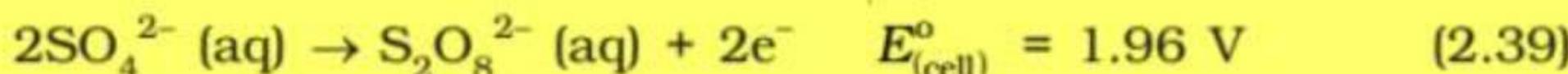
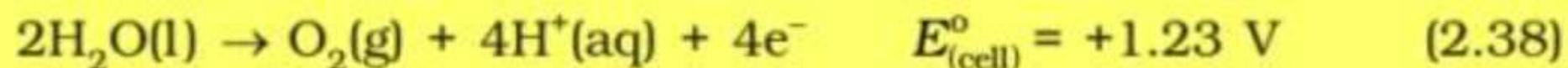
Electrode	Reaction Type	E° Zyada →	E° Kam →
Cathode	Reduction <i>gain of electron(s)</i>	<input checked="" type="checkbox"/> Preferred	<input checked="" type="checkbox"/> Less preferred
Anode	Oxidation <i>loss of electron(s)</i>	<input checked="" type="checkbox"/> Not preferred	<input checked="" type="checkbox"/> Preferred

Out of Cu^{1+} & Cu^{2+} , which will get reduced at cathode on the basis of position in electrochemical series.



NCERT HORROR

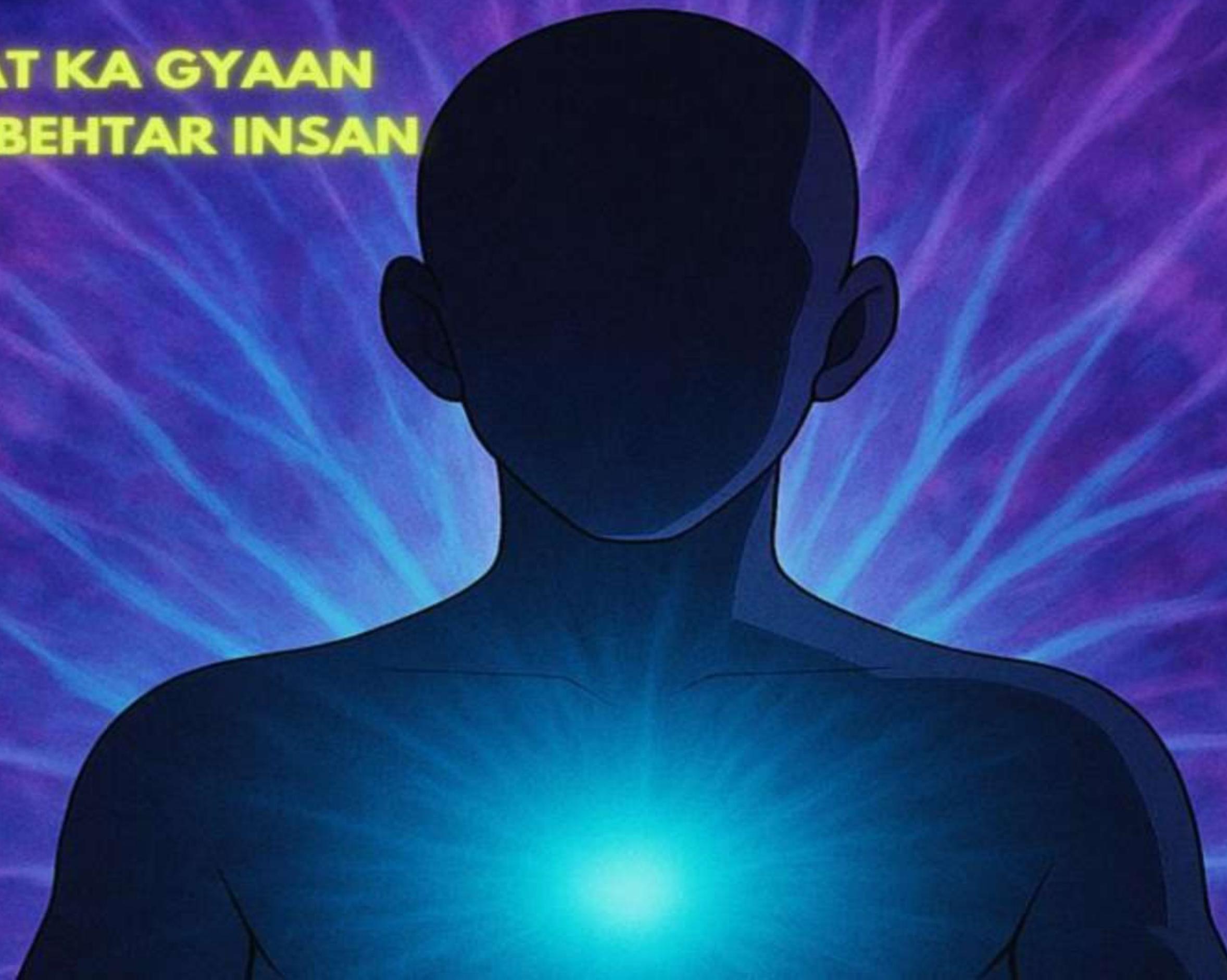
The standard electrode potentials are replaced by electrode potentials given by Nernst equation (Eq. 2.8) to take into account the concentration effects. During the electrolysis of sulphuric acid, the following processes are possible at the anode:



For dilute sulphuric acid, reaction (2.38) is preferred but at higher concentrations of H_2SO_4 , reaction (2.39) is preferred.



**INSANIYAT KA GYAAN
JO BANAE BEHTAR INSAN**



SUNIL BHAIYA

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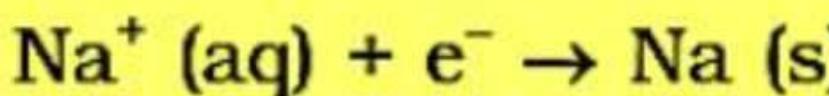
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CONCEPT POLISH – HOMEWORK

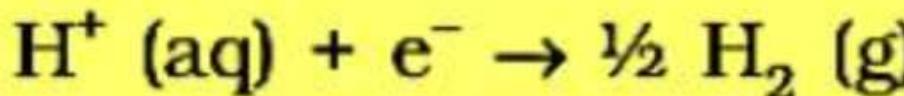


ELECTROLYSIS OF AQUEOUS NaCl

- At the cathode there is competition between the following reduction reactions:



$$E_{(\text{cell})}^\circ = - 2.71 \text{ V}$$



$$E_{(\text{cell})}^\circ = 0.00 \text{ V}$$

The reaction with higher value of E° is preferred and therefore, the reaction at the cathode during electrolysis is:



but $\text{H}^+ (\text{aq})$ is produced by the dissociation of H_2O , i.e.,

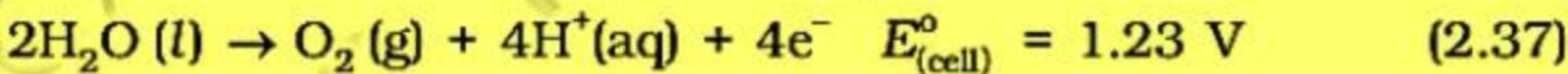


Therefore, the net reaction at the cathode may be written as the sum of (2.33) and (2.34) and we have



ELECTROLYSIS OF AQUEOUS NaCl

At the anode the following oxidation reactions are possible:



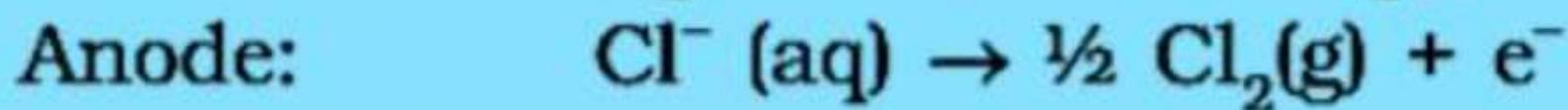
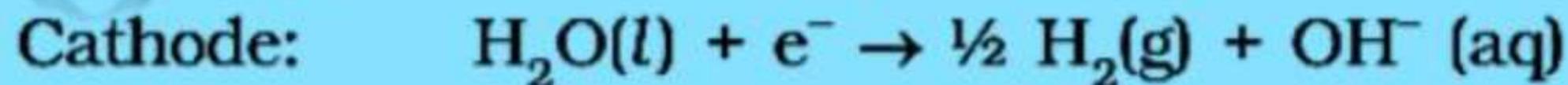
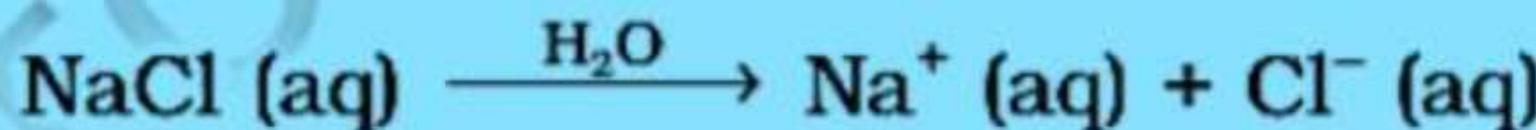
The reaction at anode with lower value of E° is preferred and therefore water should get oxidised in preference to Cl^- .

But:

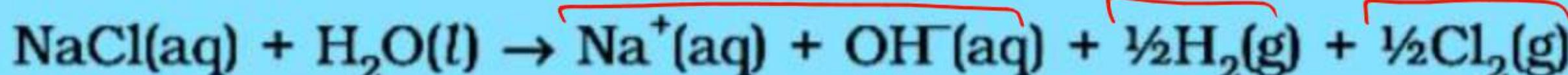
- But Cl^- is in **high concentration (brine)**
- **Overpotential for O_2 evolution is high**
- Hence, **Cl^- is discharged preferentially**

ELECTROLYSIS OF [AQUEOUS NaCl]

The net reactions will be:



Net reaction:



GIVE A THOUGHT



In the electrolysis of $\text{NaCl}(\text{aq})$, the pH of solution will:

- A. INCREASE
- B. DECREASE



SUNIL BHAIYA IS ALWAYS THERE FOR YOU.

#sbsathhai (✓)

#pwsathhai (✓)

Thank
You



PARISHRAM



2026

CHEMISTRY

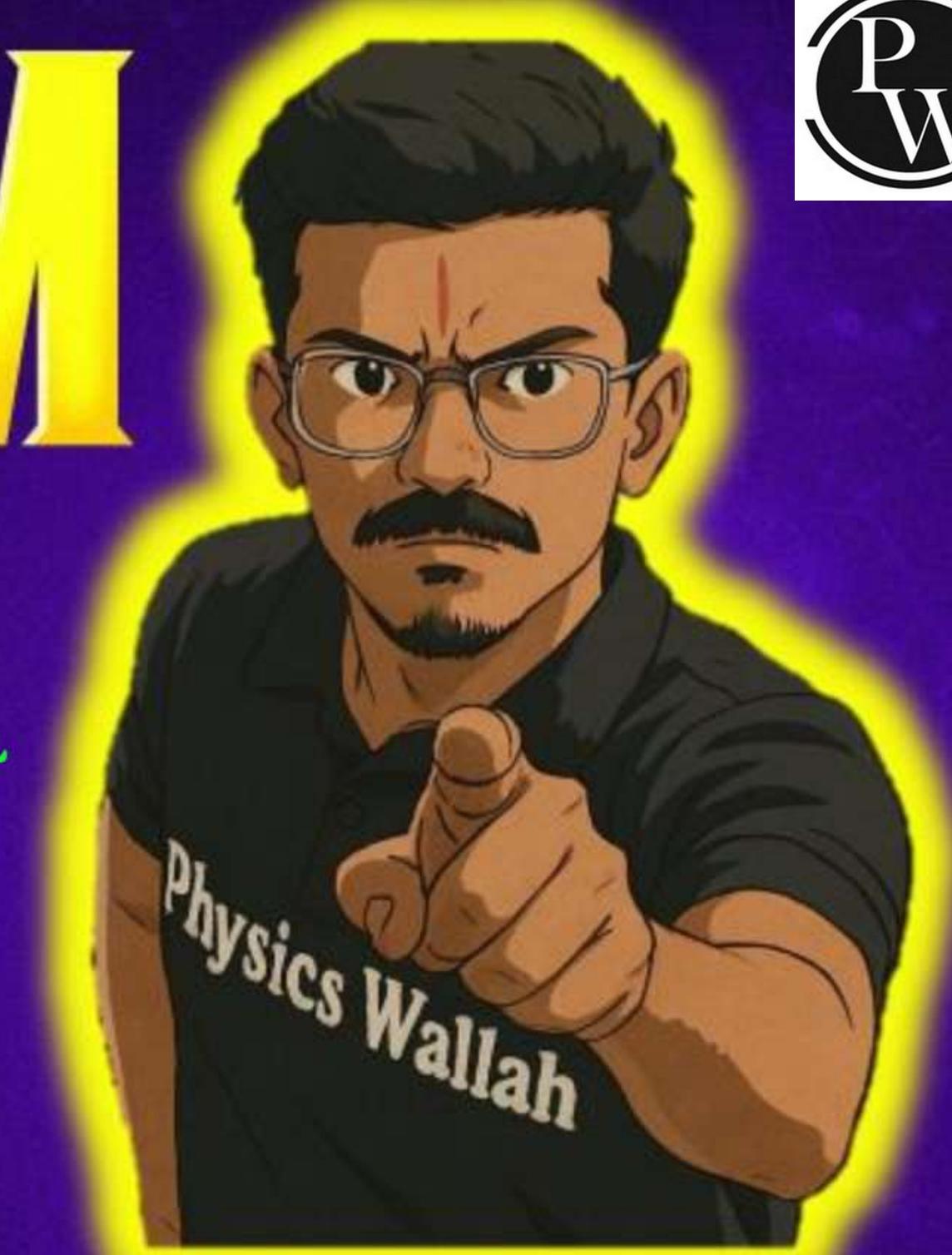
Lecture 11

ELECTROCHEMISTRY

Products of Electrolysis (Contd.) and
Question Practice (PYQs)

Bharat Mata
Ki Jai 🙏

Physics Wallah



BY – PRIYA-PUTRA-SUNIL

TOPICS TO BE COVERED

(i) Qualitative Analysis - Products of Electrolysis (Contd.)

(ii) PYQs' and Practice Problems (✓)



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NEWS WALLAH

- Last lecture of Electrochemistry will be based on Batteries, Fuel Cells and Corrosion
- It will be provided on Monday.
- [From next week, Physical and Organic Chemistry will run parallelly in the batch.]



QUALITATIVE ANALYSIS – PRODUCTS OF ELECTROLYSIS (CONTD.)

PRODUCTS OF ELECTROLYSIS

Recap

The products formed during electrolysis of an electrolyte depends on:

I ✓ Nature of Electrolyte (**Molten or Aqueous**)

II ✓ Nature of Electrode (**Active electrode or Inert electrode**)

The electrodes that participate in the redox reactions and get oxidized/reduced.

Their mass may change due to deposition or dissolution.

The electrodes that don't participate in the redox reactions and act as a surface of electron transfer only.

PRODUCTS OF ELECTROLYSIS

Recap

The products formed during electrolysis of an electrolyte depends on:

- ✓ Position in electrochemical series
 - ✓ Concentration of ions
 - ✓ Overpotential (extra voltage required above the theoretical value to make a redox reaction actually happen on an electrode in real conditions.)
- multiple cations & anions in electrolyte
which one will discharge at cathode or anode
- II, III, IV, V

Reaction (Oxidised form + ne ⁻	→ Reduced form)	E° / V
F ₂ (g) + 2e ⁻	→ 2F ⁻	2.87
Co ³⁺ + e ⁻	→ Co ²⁺	1.81
H ₂ O ₂ + 2H ⁺ + 2e ⁻	→ 2H ₂ O	1.78
MnO ₄ ⁻ + 8H ⁺ + 5e ⁻	→ Mn ²⁺ + 4H ₂ O	1.51
Au ³⁺ + 3e ⁻	→ Au(s)	1.40
Cl ₂ (g) + 2e ⁻	→ 2Cl ⁻	1.36
Cr ₂ O ₇ ²⁻ + 14H ⁺ + 6e ⁻	→ 2Cr ³⁺ + 7H ₂ O	1.33
O ₂ (g) + 4H ⁺ + 4e ⁻	→ 2H ₂ O	1.23
MnO ₂ (s) + 4H ⁺ + 2e ⁻	→ Mn ²⁺ + 2H ₂ O	1.23
Br ₂ + 2e ⁻	→ 2Br ⁻	1.09
NO ₃ ⁻ + 4H ⁺ + 3e ⁻	→ NO(g) + 2H ₂ O	0.97
2Hg ²⁺ + 2e ⁻	→ Hg ₂ ²⁺	0.92
Ag ⁺ + e ⁻	→ Ag(s)	0.80
Fe ³⁺ + e ⁻	→ Fe ²⁺	0.77
O ₂ (g) + 2H ⁺ + 2e ⁻	→ H ₂ O ₂	0.68
I ₂ (s) + 2e ⁻	→ 2I ⁻	0.54
Cu ⁺ + e ⁻	→ Cu(s)	0.52
Cu ²⁺ + 2e ⁻	→ Cu(s)	0.34
AgCl(s) + e ⁻	→ Ag(s) + Cl ⁻	0.22
AgBr(s) + e ⁻	→ Ag(s) + Br ⁻	0.10
2H ⁺ + 2e ⁻	→ H ₂ (g)	0.00

Increasing strength of reducing agent

- **Positive E°:** High tendency to get reduced (electrons lena)
- Reducing Agent: The one that reduces others (gain of electron on other species because it loses electrons).
- A positive E° means that the reduced form is a weaker reducing agent than H₂

NOTE: Standard electrode potential (E°) ki har value ko Standard Hydrogen Electrode (SHE) ke reference mein measure kia gaya hai.

Reaction (Oxidised form + ne ⁻	→ Reduced form)		E° / V
2H ⁺ + 2e ⁻	→ H ₂ (g)	Increas-	0.00
Pb ²⁺ + 2e ⁻	→ Pb(s)		-0.13
Sn ²⁺ + 2e ⁻	→ Sn(s)		-0.14
Ni ²⁺ + 2e ⁻	→ Ni(s)		-0.25
Fe ²⁺ + 2e ⁻	→ Fe(s)		-0.44
Cr ³⁺ + 3e ⁻	→ Cr(s)		-0.74
Zn ²⁺ + 2e ⁻	→ Zn(s)		-0.76
2H ₂ O + 2e ⁻	→ H ₂ (g) + 2OH ⁻		-0.83
Al ³⁺ + 3e ⁻	→ Al(s)		-1.66
Mg ²⁺ + 2e ⁻	→ Mg(s)		-2.36
Na ⁺ + e ⁻	→ Na(s)		-2.71
Ca ²⁺ + 2e ⁻	→ Ca(s)		-2.87
K ⁺ + e ⁻	→ K(s)		-2.93
Li ⁺ + e ⁻	→ Li(s)		-3.05

- **Negative E°:** Low tendency to get reduced.

- Reducing Agent: The one that reduces others (gain of electron on other species because it loses electrons).
- A negative E° means that the reduced form is a stronger reducing agent than H₂

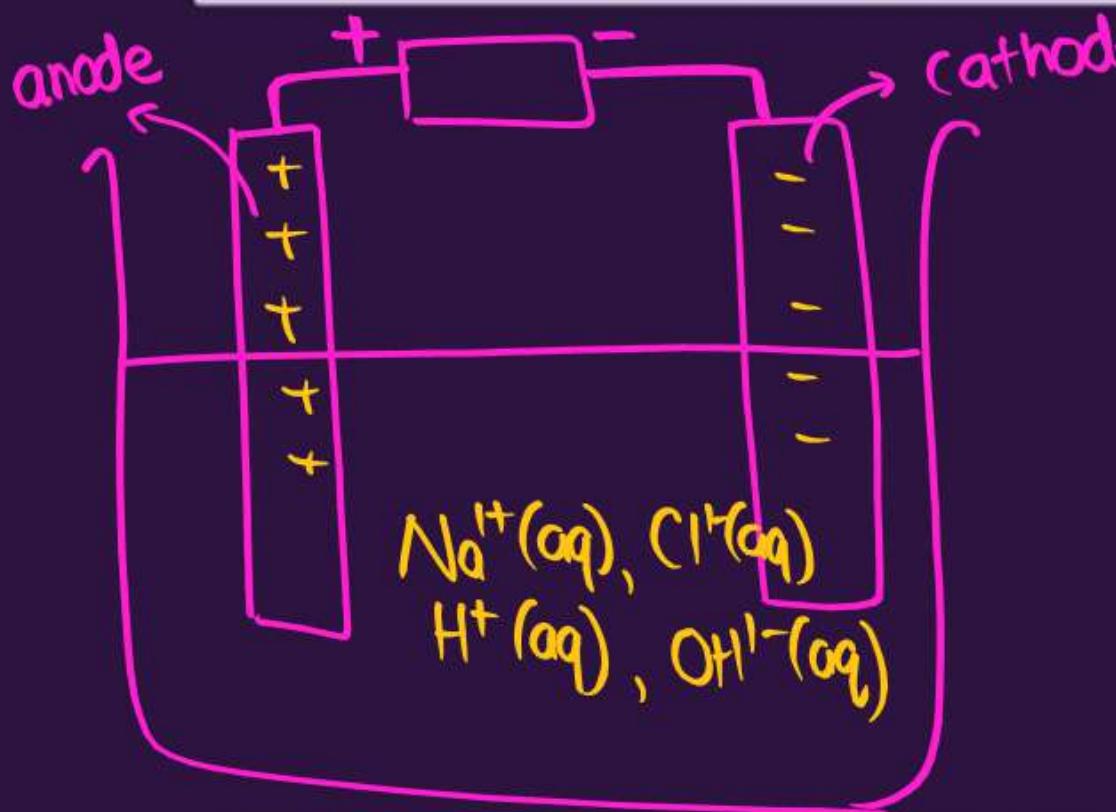
NOTE: Standard electrode potential (E°) ki har value ko Standard Hydrogen Electrode (SHE) ke reference mein measure kia gaya hai.

IMPORTANT TO UNDERSTAND

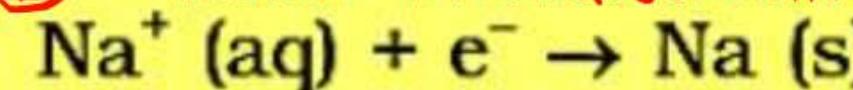
Electrode	Reaction Type	E° Zyada →	E° Kam →
Cathode ✓	Reduction ✓ <i>(gain of electron(s))</i>	<input checked="" type="checkbox"/> Preferred	<input checked="" type="checkbox"/> Less preferred
Anode ✓	Oxidation ✓ <i>Loss of electron(s)</i>	<input checked="" type="checkbox"/> Not preferred	<input checked="" type="checkbox"/> Preferred

ELECTROLYSIS OF AQUEOUS NaCl

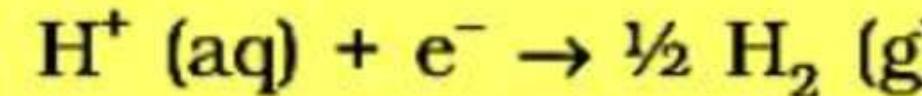
- At the cathode there is competition between the following reduction reactions: ① [Anode & cathode \rightarrow Inert electrode]



② Position in electrochemical series



$$E_{\text{(cell)}}^{\circ} = -2.71 \text{ V}$$



$$E_{\text{(cell)}}^{\circ} = 0.00 \text{ V}$$

The reaction with higher value of E° is preferred and therefore, the reaction at the cathode during electrolysis is:



but $\text{H}^+(\text{aq})$ is produced by the dissociation of H_2O , i.e.,



mostly H_2O molecules

Therefore, the net reaction at the cathode may be written as the sum of (2.33) and (2.34) and we have



(iii) Concentrations of ions: Conc. of Na^+ (aq) >>> Conc. of H^+ (aq)

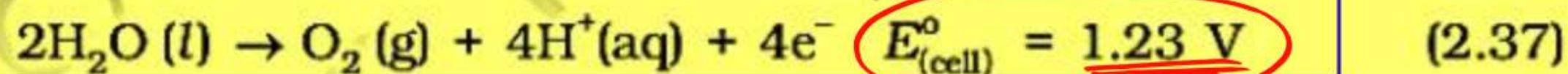
(iv) Overpotential: formation of H_2 gas is a slow rxn.

(iii) & (iv) $\rightarrow \text{Na}^+$ (aq) should get preference but as there is a huge SRP difference. Hence, H^+ (aq) or H_2O will get reduced.

ELECTROLYSIS OF AQUEOUS NaCl

Cl^-
 OH^-
 from water

At the anode the following oxidation reactions are possible:



① Inert electrode
 ↓
 No effect

The reaction at anode with lower value of E° is preferred and therefore water should get oxidised in preference to Cl^- .

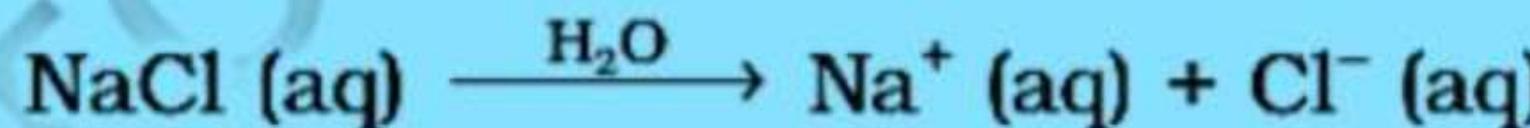
But:

- But Cl^- is in **high concentration (brine)**
- **Overpotential for O_2 evolution is high**
- Hence, **Cl^- is discharged preferentially**

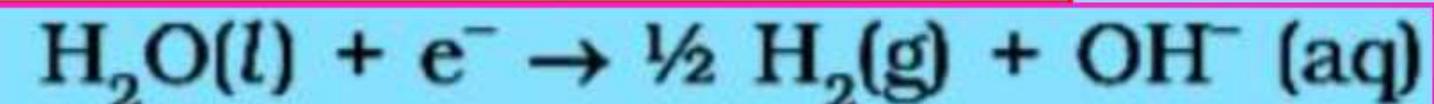
Because there is not a huge difference in SRP values so, ③ & ④ overpowers.

ELECTROLYSIS OF AQUEOUS NaCl

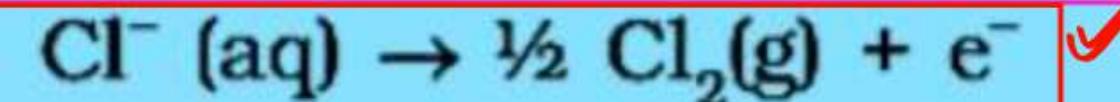
The net reactions will be:



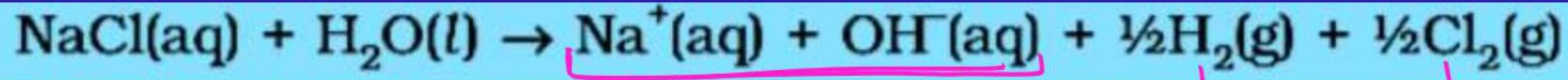
Cathode:



Anode:



Net reaction:



↓
in solution

↓
cathode

↓
anode

GIVE A THOUGHT



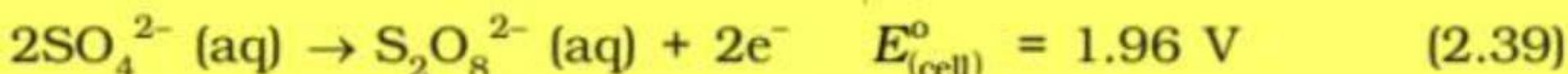
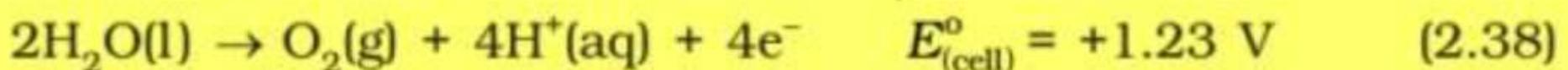
In the electrolysis of $\text{NaCl}(\text{aq})$, the pH of solution will:

- A. INCREASE
- B. DECREASE

$$\text{pH} > 7$$

NCERT HORROR

The standard electrode potentials are replaced by electrode potentials given by Nernst equation (Eq. 2.8) to take into account the concentration effects. During the electrolysis of sulphuric acid, the following processes are possible at the anode:



For dilute sulphuric acid, reaction (2.38) is preferred but at higher concentrations of H_2SO_4 , reaction (2.39) is preferred.

→ LOGIC

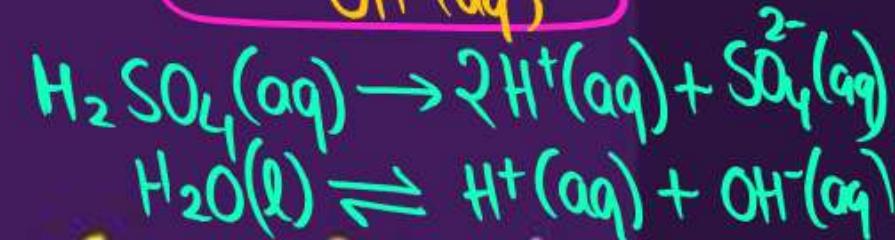
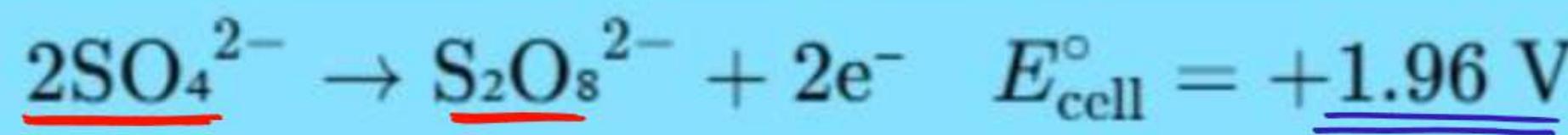
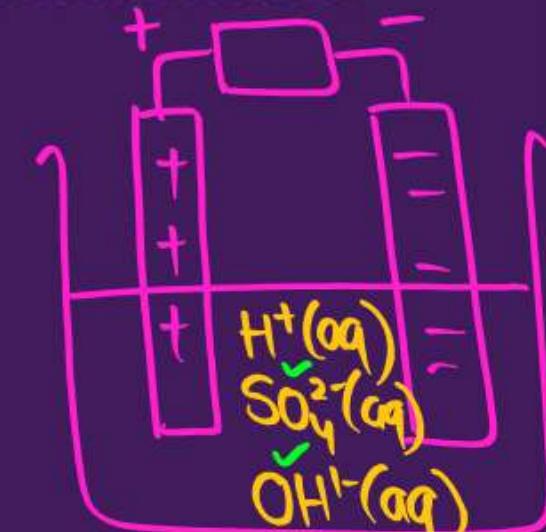
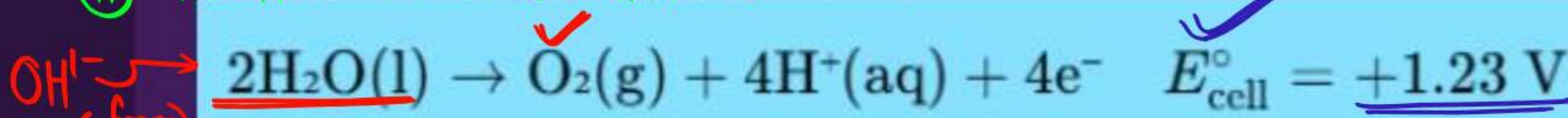


EFFECT OF CONCENTRATION ON PRODUCTS OF ELECTROLYSIS

- During electrolysis of sulphuric acid (H_2SO_4), oxidation occurs at the anode. But two different oxidation reactions are possible!

① Inert electrodes
↓
no effect

② Position in electrochemical series



The reaction at anode with **lower value of E° is preferred** and therefore **water should get oxidised** in preference to SO_4^{2-} . But **concentration and overpotential** plays a vital role here.

(III)

(IV)

EFFECT OF CONCENTRATION ON PRODUCTS OF ELECTROLYSIS

- But in real, **concentration changes**, so we use the **Nernst Equation** to adjust E values based on real conditions. Here, concentration of electrolyte is not always 1 M.

Concentration	Preferred Reaction	Why?
Dilute H_2SO_4	Reaction (2.38)	More water molecules → water gets oxidized easily (lower E°)
Concentrated H_2SO_4	Reaction (2.39)	More SO_4^{2-} ions → peroxodisulfate formation dominates despite higher E°

- ✓ SO_4^{2-} is in **high concentration**
- ✓ **Overpotential for O_2 evolution is high**
- Hence (**SO_4^{2-} is discharged preferentially**)

factor 2 & 3
 in favour of
 H_2O
 ✓
 ✓

Overpowers
 factor 2 ← in favour of
 SO_4^{2-}
 ✓ factor 3 & 4

Standard Temperature and Pressure (STP):

- From 1982: Temperature (~~0°C~~), Pressure ~~1 atm~~
- From 1982: Temperature (~~0°C~~), Pressure ~~1 bar~~

aye
bhaiya ❤

PYQS' AND PRACTICE PROBLEMS

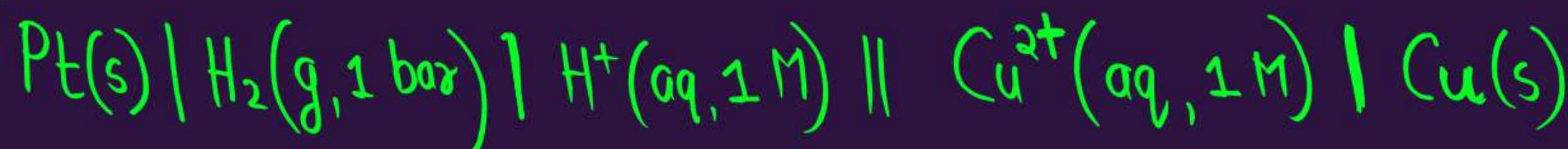
Which cell will measure standard electrode potential of copper electrode?

- (i) Pt (s) | H₂ (g, 0.1 bar) | H⁺ (aq., 1 M) || Cu²⁺ (aq., 1 M) | Cu
- (ii) Pt(s) | H₂ (g, 1 bar) | H⁺ (aq., 1 M) || Cu²⁺ (aq., 2 M) | Cu
- (iii) Pt(s) | H₂ (g, 1 bar) | H⁺ (aq., 1 M) || Cu²⁺ (aq., 1 M) | Cu
- (iv) Pt(s) | H₂ (g, 1 bar) | H⁺ (aq., 0.1 M) || Cu²⁺ (aq., 1 M) | Cu

← Oxidation → ← Reduction →

✓ Anode | Anode electrolyte || Cathode electrolyte | Cathode

T = 25 °C



The difference between the electrode potentials of two electrodes when no current is drawn through the cell is called _____.

- (i) Cell potential
- (ii) Cell emf / Standard emf
- (iii) Potential difference
- (iv) Cell voltage

Match the items of Column I and Column II.

Column I

(i) κ

(ii) Λ_m

(iii) α

(iv) Q

$$\Lambda_m = \frac{K}{C}$$

or

$$\Lambda_m = \frac{K \times 1000}{C}$$

Column II

(a) $I \times t$

(b) Λ_m / Λ_m^0

(c) $\frac{\kappa}{c}$

(d) $\frac{G^*}{R}$

$$G_1 = \frac{1}{R} = \frac{A}{\beta l} = \frac{KA}{l}$$



$$\left(\beta = \frac{1}{K}\right)$$

$$\left(K = \frac{1}{\beta}\right)$$

$$\frac{G_1 \times l}{A} = K$$

$$\frac{l \times \frac{l}{R}}{A} = K$$

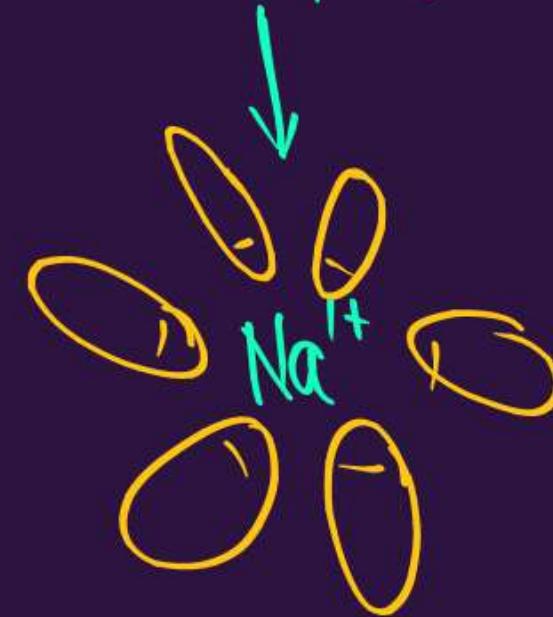
$$\frac{l \times G^*}{R} = K$$

Cell
Constant

Which of the statements about solutions of electrolytes is **not** correct?

- (i) Conductivity of solution depends upon size of ions. (✓)
- (ii) Conductivity depends upon viscosity of solution. (✗)
- ~~(iii)~~ Conductivity does not depend upon solvation of ions present in solution. (✗)
- (iv) Conductivity of solution increases with temperature. (✓)

Solvation | Hydration of ions



Using the data given below find out the **strongest reducing agent.**

$$E^\ominus_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}} = 1.33\text{V} \quad E^\ominus_{\text{Cl}_2/\text{Cl}^-} = 1.36\text{V}$$

(One that donates e^-)

$$E^\ominus_{\text{MnO}_4^-/\text{Mn}^{2+}} = 1.51\text{V} \quad E^\ominus_{\text{Cr}^{3+}/\text{Cr}} = -0.74\text{V}$$

Oxidised form + $ne^- \rightarrow$ Reduced form



- (i) Cl^-
- (ii) Cr
- (iii) Cr^{3+}
- (iv) Mn^{2+}

above question

Use the data given in Q.8 and find out the most stable ion in its reduced form.

- (i) Cl^-
- (ii) Cr^{3+}
- (iii) Cr
- (iv) Mn^{2+}

Oxidised form + $\text{he}^- \rightarrow$ Reduced form

more value of E° , it means high tendency to get reduced \rightarrow Reduced form is more stable.

The quantity of charge required to obtain (one mole) of aluminium from Al_2O_3 is _____.

- (i) 1F
- (ii) 6F
- (iii) 3F
- (iv) 2F

electrolyte



3F

NCERT EXEMPLAR

Λ_m^0 (NH₄OH) is equal to _____.

(i) Λ_m^0 (NH₄OH) + Λ_m^0 (NH₄Cl) - Λ_m^0 (HCl)

(ii) $\underline{\Lambda_m^0}$ (NH₄Cl) + $\underline{\Lambda_m^0}$ (NaOH) - $\underline{\Lambda_m^0}$ (NaCl)

(iii) $\underline{\Lambda_m^0}$ (NH₄Cl) + Λ_m^0 (NaCl) - Λ_m^0 (NaOH) X

(iv) Λ_m^0 (NaOH) + Λ_m^0 (NaCl) - Λ_m^0 (NH₄Cl)

NCERT EXEMPLAR

In the electrolysis of aqueous sodium chloride solution which of the half cell reaction will occur at anode?

- (i) $\text{Na}^+ (\text{aq}) + \text{e}^- \longrightarrow \text{Na} (\text{s}); E_{\text{Cell}}^\ominus = -2.71\text{V}$
- (ii) $2\text{H}_2\text{O} (\text{l}) \longrightarrow \text{O}_2 (\text{g}) + 4\text{H}^+ (\text{aq}) + 4\text{e}^-; E_{\text{Cell}}^\ominus = 1.23\text{V}$
- (iii) $\text{H}^+ (\text{aq}) + \text{e}^- \longrightarrow \frac{1}{2} \text{H}_2 (\text{g}); E_{\text{Cell}}^\ominus = 0.00\text{ V}$
- (iv) $\text{Cl}^- (\text{aq}) \longrightarrow \frac{1}{2} \text{Cl}_2 (\text{g}) + \text{e}^-; E_{\text{Cell}}^\ominus = 1.36\text{ V}$

Conductivity κ , is equal to _____.

(i) $\frac{1}{R} \left(\frac{l}{A} \right) \rightarrow G^* \rightarrow \text{cell constant}$

(ii) $\frac{G^*}{R}$

(iii) $\Lambda_m \times$

(iv) $\frac{l}{A} \times$

For the given cell, $Mg | Mg^{2+} \parallel Cu^{2+} | Cu$

(i) Mg is cathode \times

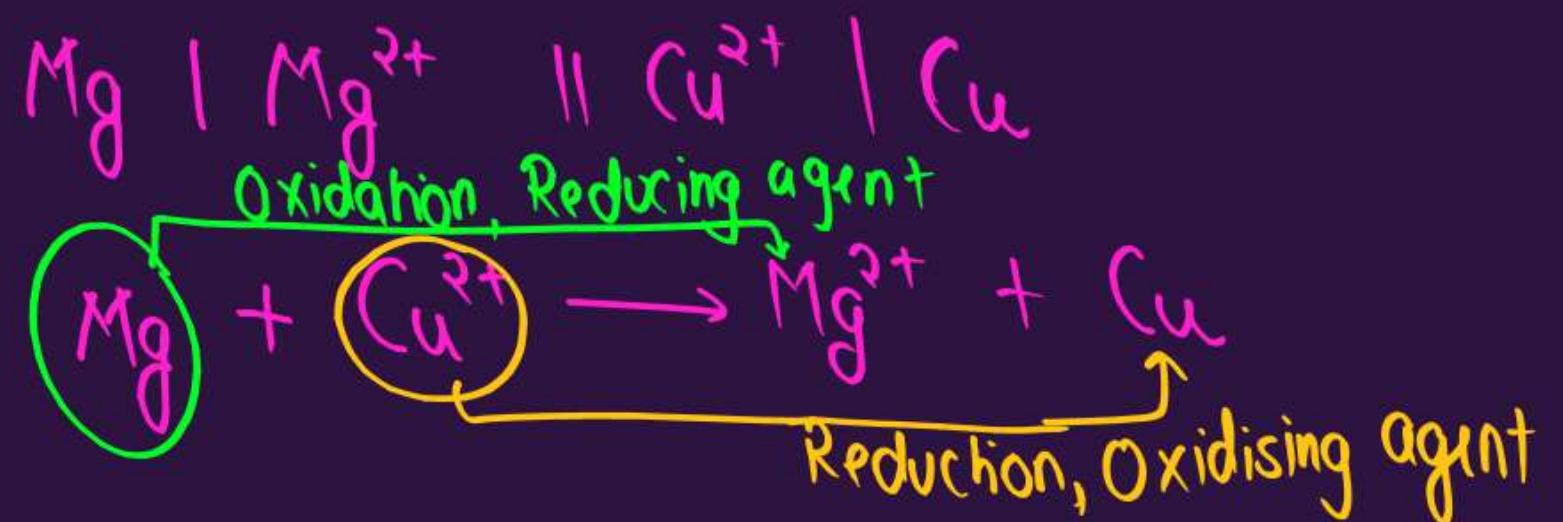
$Mg \rightarrow \text{anode} \rightarrow \text{oxidation}$

(ii) Cu is cathode

(iii) The cell reaction is $Mg + Cu^{2+} \longrightarrow Mg^{2+} + Cu$

(iv) Cu is the oxidising agent \times

Anode | Anode electrolyte \parallel Cathode electrolyte | Cathode



A galvanic cell can behave like an electrolytic cell when

- (A) $E_{\text{cell}} = E_{\text{ext}}$
- ~~(B) $E_{\text{ext}} > E_{\text{cell}}$~~
- (C) $E_{\text{cell}} > E_{\text{ext}}$
- (D) $E_{\text{cell}} = 0$

✓ Assertion (A) : For measuring resistance of an ionic solution an AC source is used. (✓)

Reason (R) : Concentration of ionic solution will change if DC source is used. (✓) ↳ electrolysis of electrolyte will start

- (A) Both Assertion (A) and Reason (R) are true and Reason (R) is the correct explanation of the Assertion (A).
- (B) Both Assertion (A) and Reason (R) are true, but Reason (R) is **not** the correct explanation of the Assertion (A).
- (C) Assertion (A) is true, but Reason (R) is false.
- (D) Assertion (A) is false, but Reason (R) is true.

Kohlrausch gave the following relation for strong electrolytes :

$$\Lambda_m = \Lambda_m^\circ - A\sqrt{C}$$

Which of the following equality holds true ?

(A) $\Lambda_m = \Lambda_m^\circ$ as $C \rightarrow 1$

(C) $\Lambda_m = \Lambda_m^\circ$ as $C \rightarrow \infty$

(B) $\Lambda_m = \Lambda_m^\circ$ as $C \rightarrow 0$

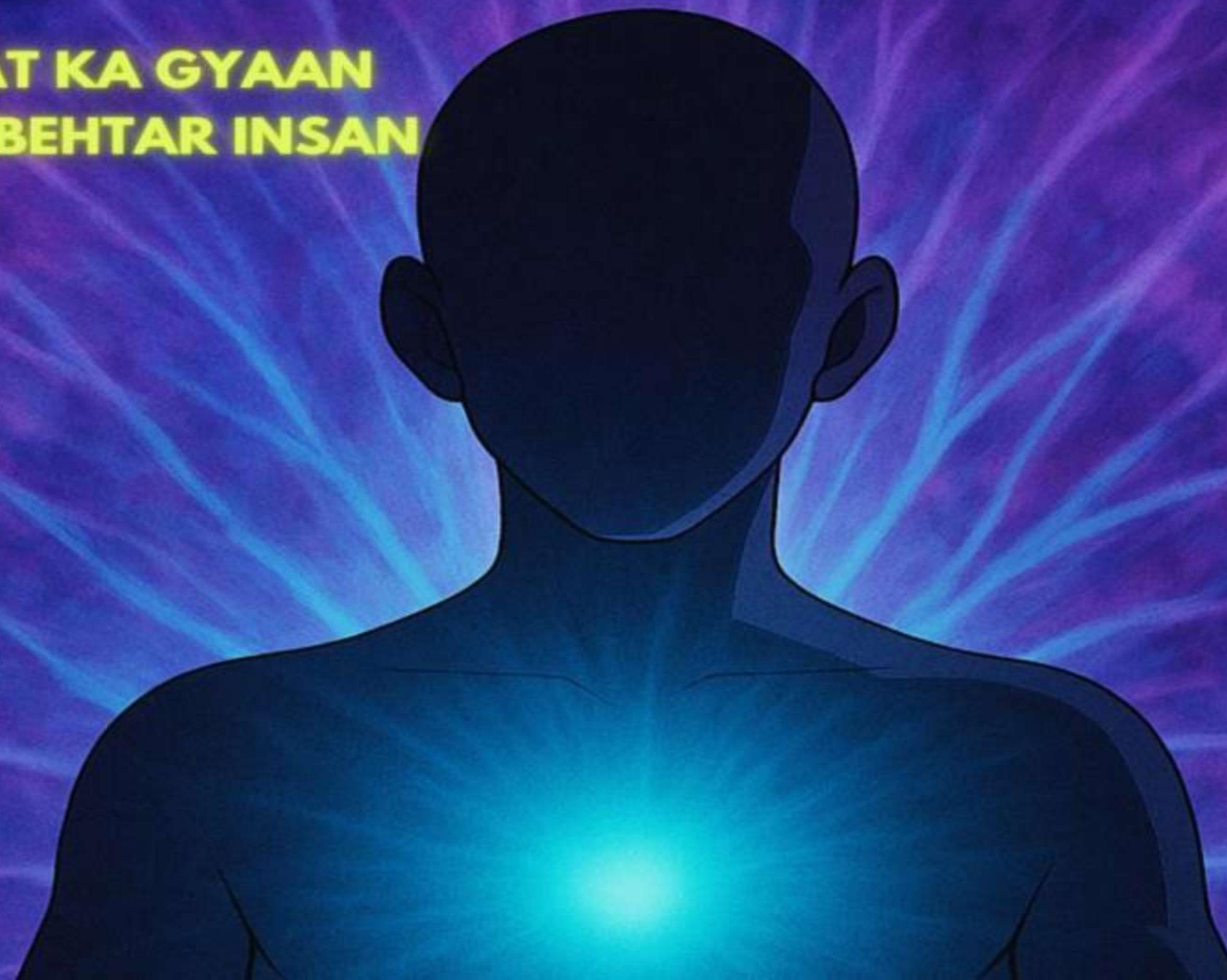
(D) $\Lambda_m = \Lambda_m^\circ$ as $C \rightarrow \sqrt{A}$

SAMAJ AAYA TOH
LIKH DO.

AYE BHAIYA



**INSANIYAT KA GYAAN
JO BANAE BEHTAR INSAN**



SUNIL BHAIYA

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CONCEPT POLISH – HOMEWORK



above question

Use the data given in Q.8 and find out which of the following is the strongest oxidising agent.

- (i) Cl^-
- (ii) Mn^{2+}
- (iii) MnO_4^-
- (iv) Cr^{3+}

H.W.

*(Interesting
question)*



SUNIL BHAIYA IS ALWAYS THERE FOR YOU.

#sbsathhai(✓)

#pwsathhai(✓)

Thank
You



PARISHRAM



2026

CHEMISTRY

Lecture 12

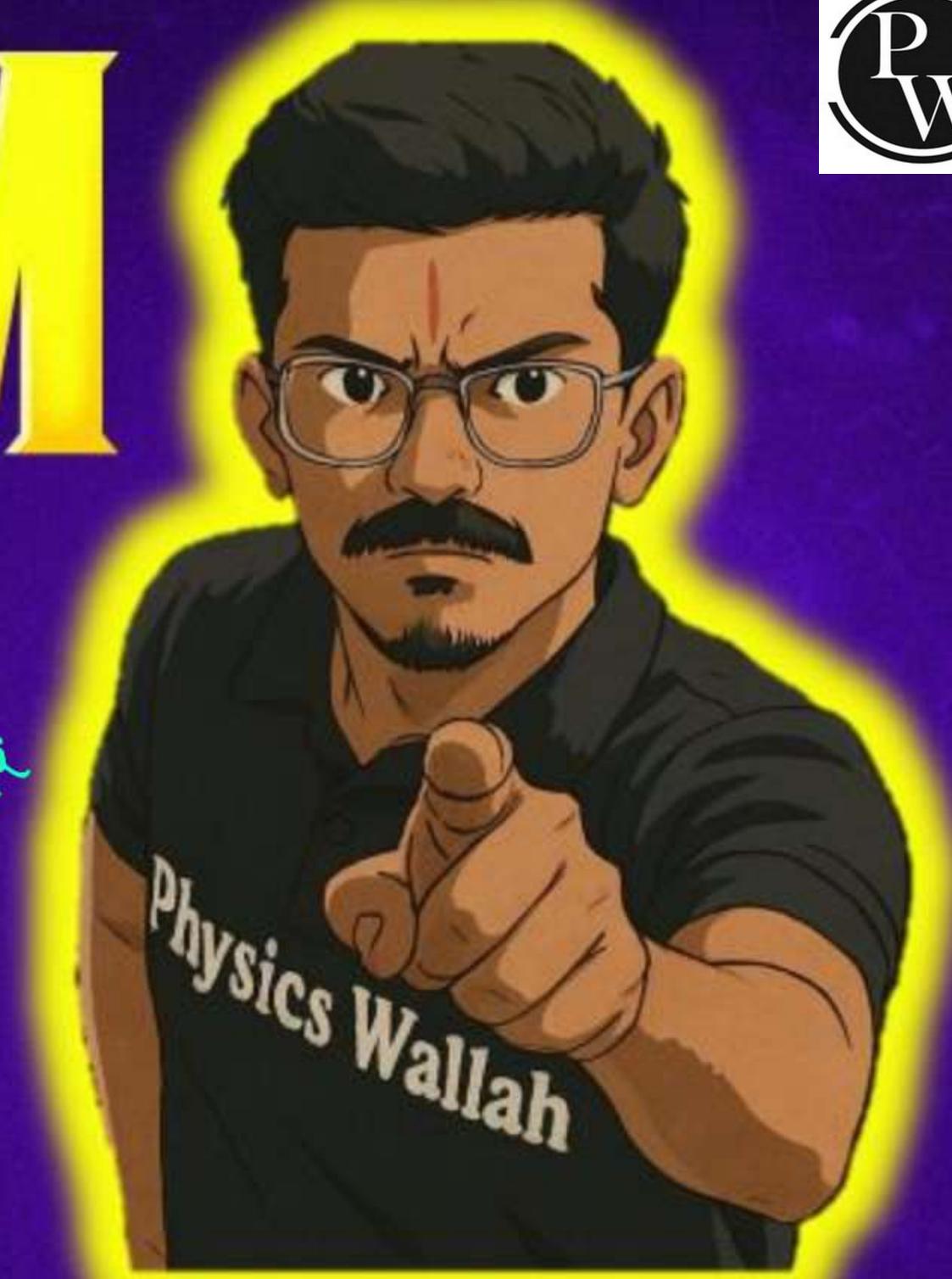
ELECTROCHEMISTRY

Corrosion and Batteries – Part I

Bharat Mata
Ki Jai

Physics Wallah

BY – PRIYA-PUTRA-SUNIL



TOPICS TO BE COVERED

(i) Type of Redox Reactions:
Corrosion 

(ii) [Batteries – Part I ]



SUNIL BHAIYA

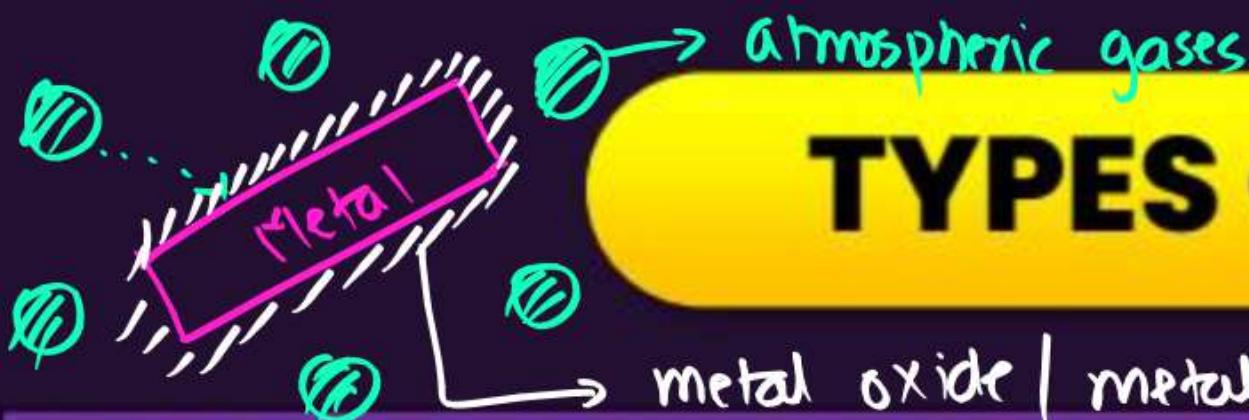
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TYPE OF REDOX REACTION: CORROSION

(\downarrow
संक्षिप्त)



TYPES OF REDOX REACTION



TYPES OF REDOX REACTIONS

Corrosion

It is a surface deterioration/degradation process of metals in which they oxidise** and convert into oxides, sulphides, carbonates and more, due to the attack of atmospheric gases.



OTHER DETAILS

Types of Corrosion - Rusting

Happens in iron and a reddish-brown layer is formed on iron which is called rust.

→ Something not in favour of

Cons of Corrosion - Rusting

- Rust is a flaky (non-sticky) layer that is very brittle and peels-off. It then exposes the fresh iron layer to moisture and oxygen. This continuous cycle makes iron objects weak and can collapse buildings and bridges, break oil pipelines and more.

TYPES OF REDOX REACTION

TYPES OF REDOX REACTIONS

Corrosion

It is a surface deterioration process of metals in which they convert to a more stable form, i.e. oxides, sulphides, carbonates and more, due to the attack of atmospheric gases.

OTHER DETAILS

Types of Corrosion - Tarnishing

- Tarnishing: Happens in copper, silver etc. and a green layer is formed on copper while a black layer is formed on silver. This protective layer is called patina.

→ something in favour of

Pros of Corrosion - Tarnishing

- Patina seems to be helpful for some of the metals like Copper (Cu), which on oxidation forms a protective layer that protects further corrosion (here tarnishing) of metal.

TYPES OF REDOX REACTION



RUSTING OF IRON

↓
will be discussing this!

① Oxygen

② Water / Water-vapour



TARNISHING OF SILVER



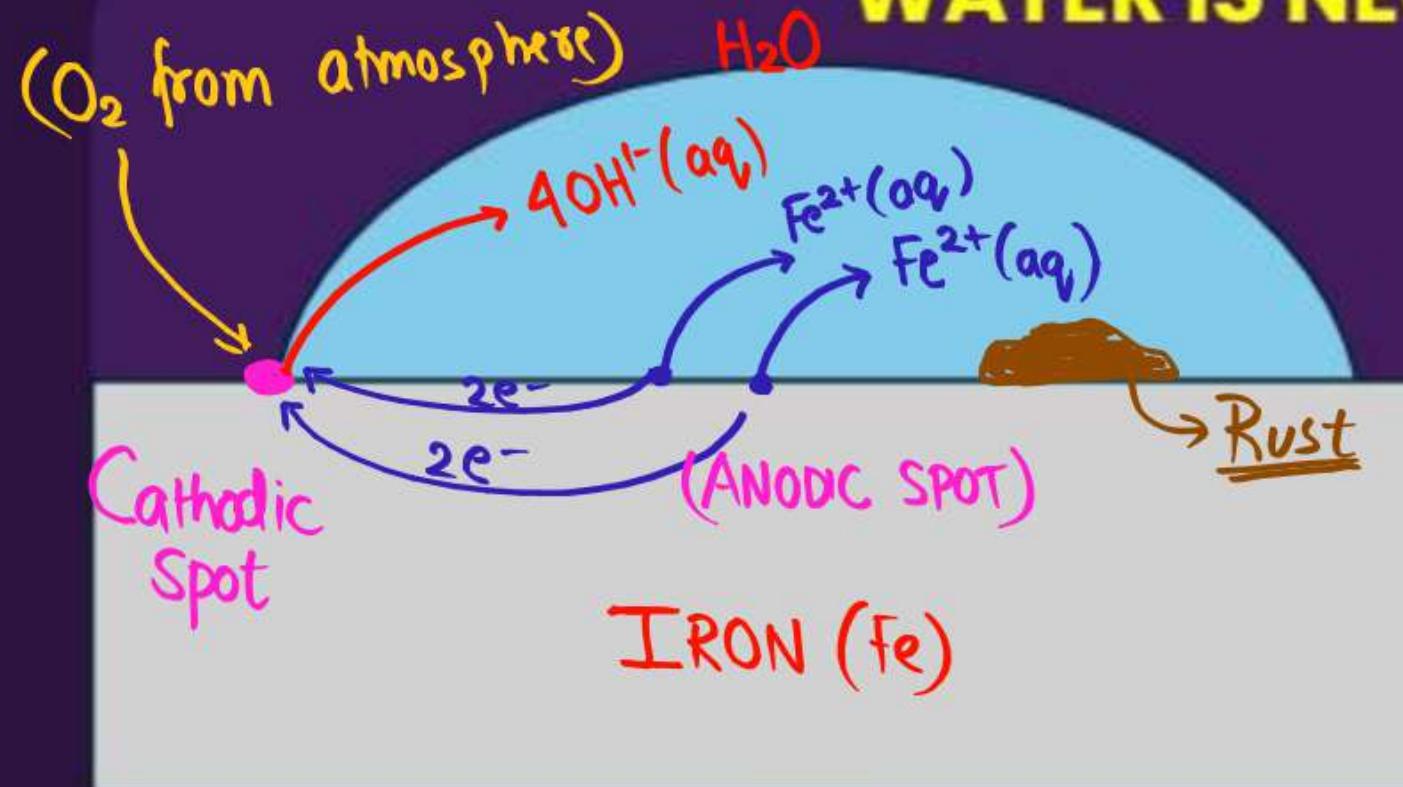
TARNISHING OF COPPER

→ Green | Blue-green layers

→ Basic copper carbonate

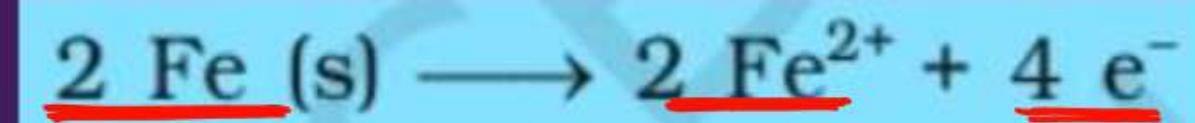
→ CuCO₃, Cu(OH)₂

RUSTING OF IRON – A REDOX REACTION

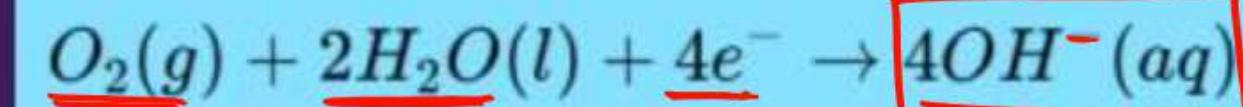


WATER IS NEUTRAL/SLIGHTLY BASIC

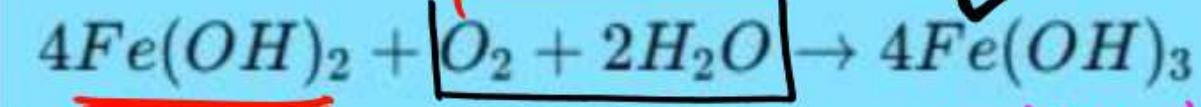
✓ (i) Reaction at **anodic spot:**



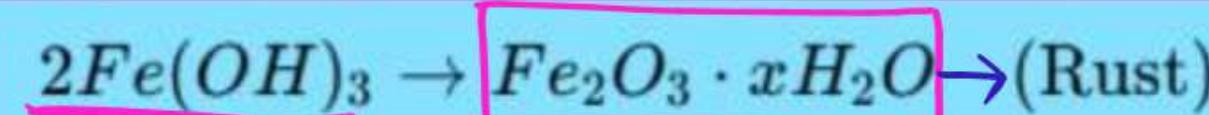
✓ (ii) Reaction at **cathodic spot:**



(iii) Oxidation



(iv) Dehydration



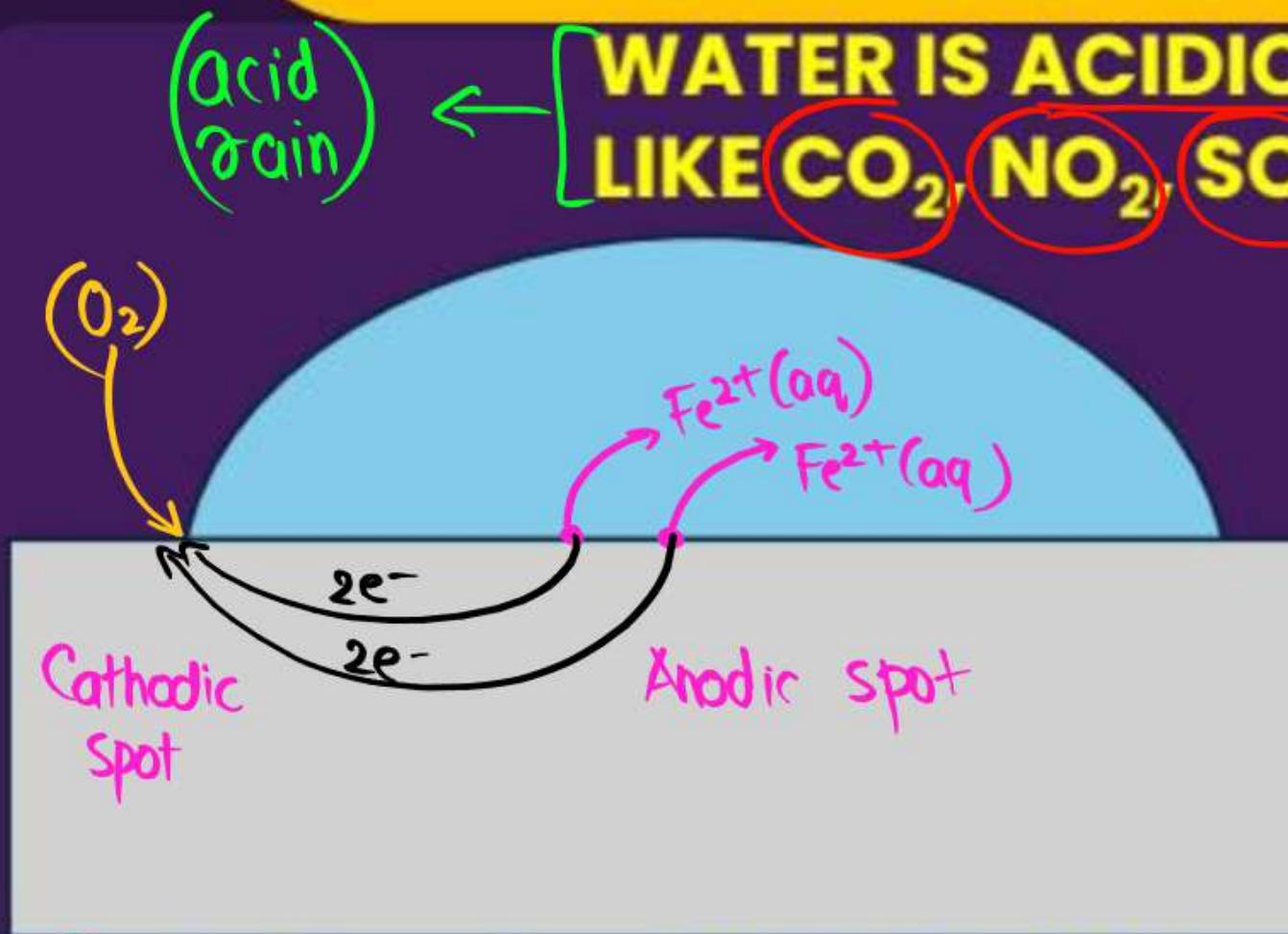
Hydrated ferric oxide

OXIDATION → Loss of electrons

REDUCTION → Gain of electrons

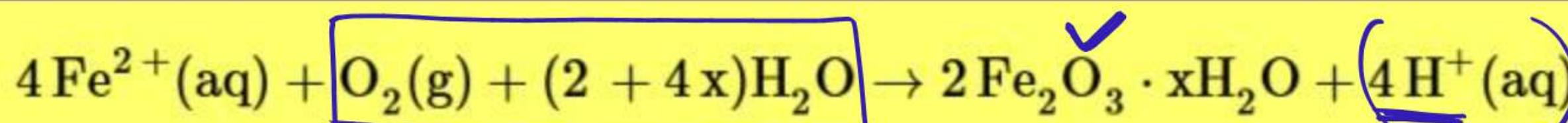
reddish-brown ppt.

RUSTING OF IRON – A REDOX REACTION



IV Fe²⁺ produced in the initial reaction are then oxidised by atmospheric oxygen to produce the insoluble hydrated oxide containing Fe³⁺, as:

RM
Not in
Syllabus



WAYS TO PREVENT RUSTING

Type of Protection

I Barrier Protection ✓



- Applying a layer of paint, oil or grease or compounds like bisphenol on iron

II Sacrificial Protection ✓

- In this a more reactive metal is placed on iron.
- For ex: Galvanisation (Applying a layer of zinc on iron)
- Layer of Zn acts as a barrier and zinc being more reactive than iron oxidises itself at place of iron.

III Tin or Chrome Plating ✓

- Tin or chromium which are less reactive metals than iron are placed on iron.
- Chromium makes a stable and inert oxide layer.
- Tin on the other hand is just a physical barrier on iron.

**SAMAJ AAYA TOH
LIKH DO.
AYE BHAIYA**



A cell or a group of cells connected in series (battery) that we use as a source of electrical energy is basically a galvanic cell where the chemical energy of the redox reaction is converted into electrical energy.



- Part I

BATTERIES AND FUEL CELLS

GIVE A THOUGHT



What are the requirements for a battery to be of a practical use?

GIVE A THOUGHT



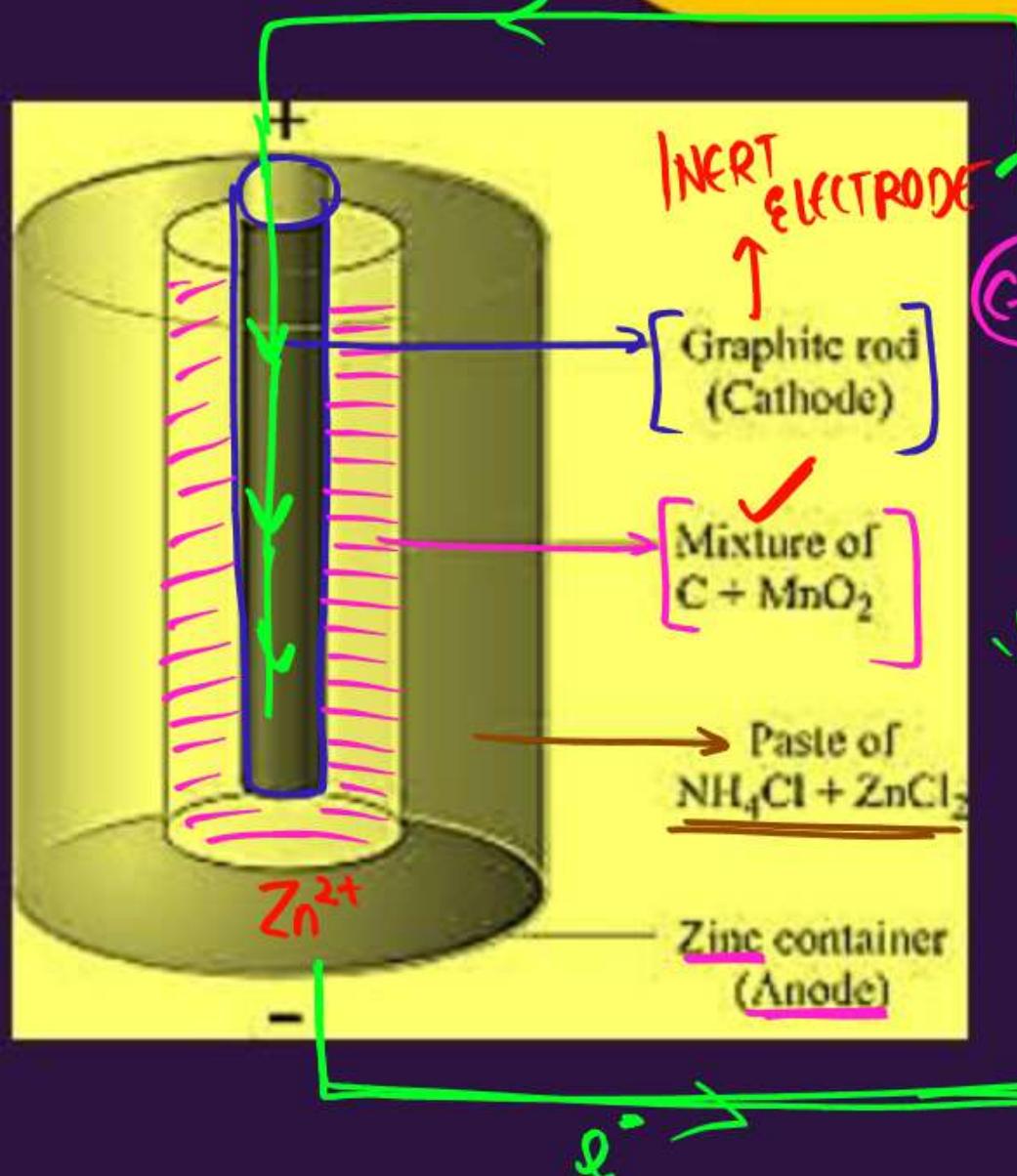
What are the requirements for a battery to be of a practical use?

For a battery to be of practical use it should be reasonably light, compact and its voltage should not vary appreciably during its use.

PRIMARY BATTERIES

- ✓ The reaction occurs (only once and after use over a period of time battery becomes dead and cannot be reused again.)
- ✓ Primary batteries are electrochemical cells in which the electrode reactions are 'irreversible' under normal conditions.
- Its 2 famous types are:
 - (i) Dry cell (known as Leclanche cell)
 - (ii) Mercury cell

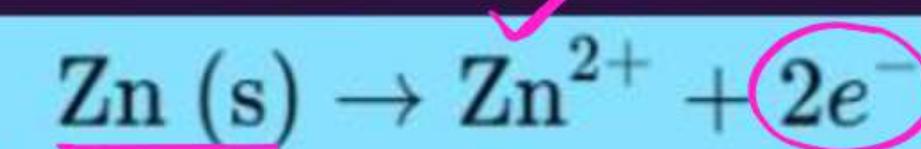
LECLANCHE CELL (DRY CELL)



STRUCTURE

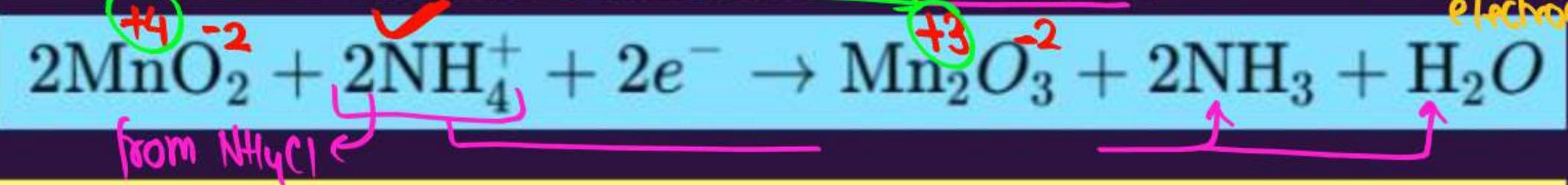
- Anode (-): Zinc container
- Cathode (+): Carbon rod (Graphite), surrounded by $MnO_2 +$ Carbon powder
- Electrolyte: Paste of $NH_4Cl + ZnCl_2$ (moist) $\rightarrow (H_2O)$

REACTION AT ANODE \rightarrow Oxidation



Loss of electron(s)

REACTION AT CATHODE \rightarrow Reduction



Gain of electron(s)

OVERALL REACTION $\rightarrow Zn + 2MnO_2 + 2NH_4^+ \rightarrow Zn^{2+} + Mn_2O_3 + 2NH_3 + H_2O$

GIVE A THOUGHT



Where did the NH_3 and H_2O are used up in the reaction?

GIVE A THOUGHT



Where did the NH_3 and H_2O are used up in the reaction?

Product	Final Fate	Role
NH_3	$\checkmark \text{Zn}^{2+}$ ke saath complex banata hai	<u>$[\text{Zn}(\text{NH}_3)_4]^{2+}$ banake reaction ko aage badhata hai</u>
H_2O	Electrolyte paste mein rehta hai	<u>Moisture maintain karta hai, ion flow mein help karta hai</u>

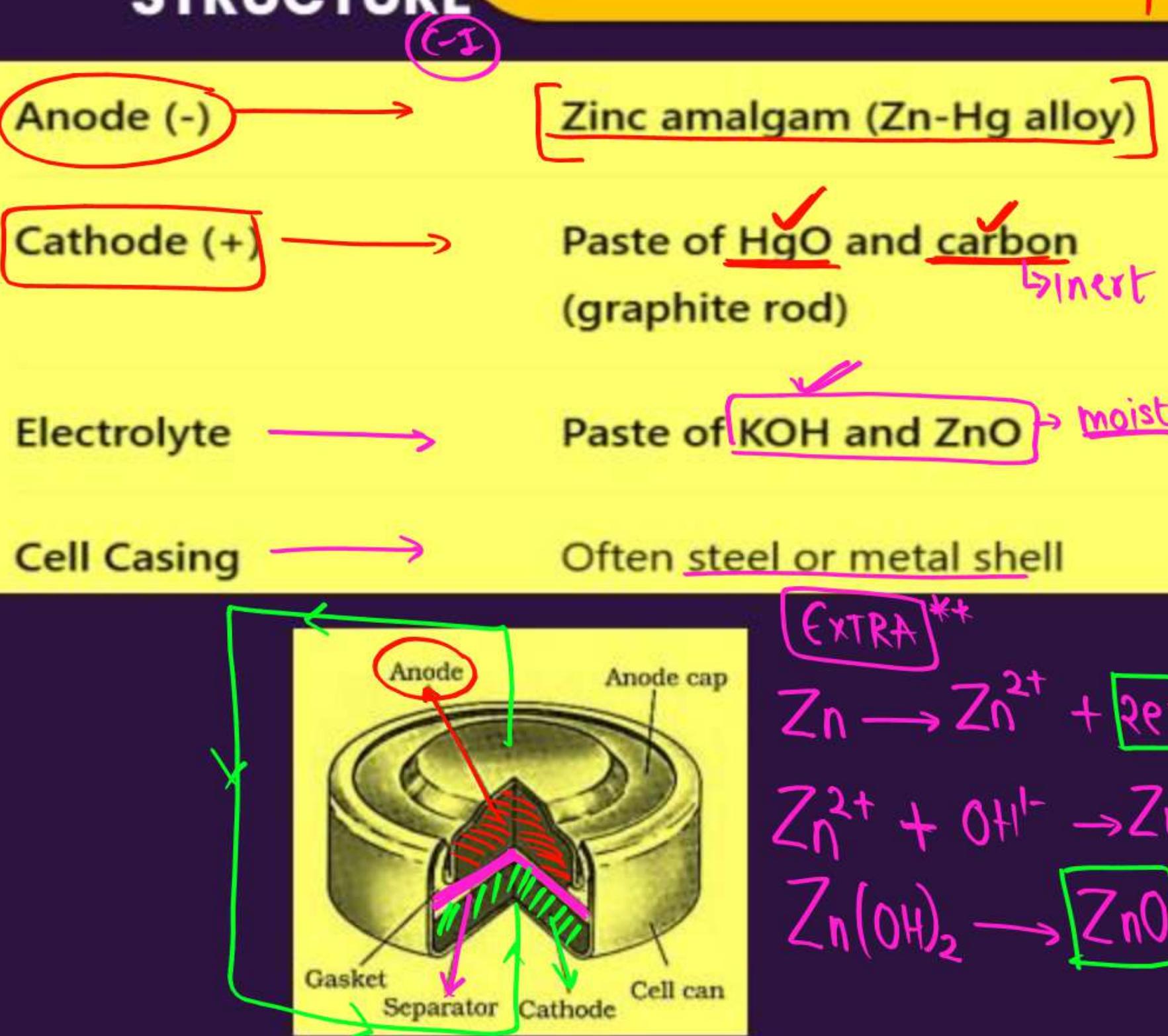
IMPORTANT TO REMEMBER

- The cell has a potential of nearly 1.5 V
- It is used commonly in our transistors and clocks.



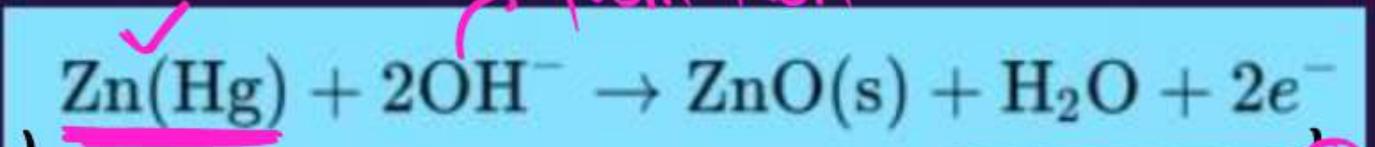
MERCURY CELL (BUTTON CELL)

STRUCTURE

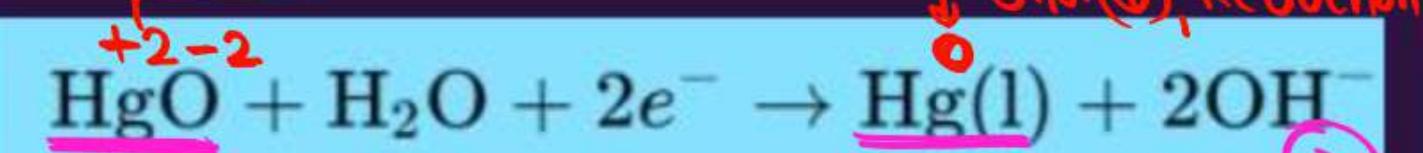


(C-II)

REACTION AT ANODE



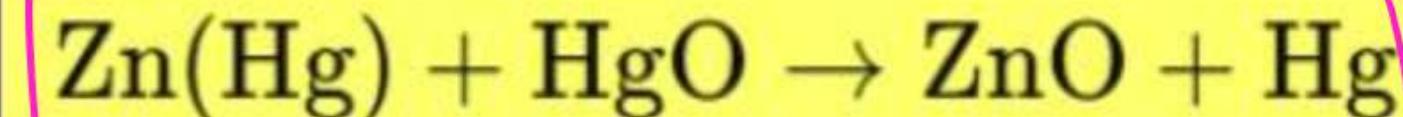
REACTION AT CATHODE



LOGIC

OVERALL REACTION

(1) + (2)



IMPORTANT TO REMEMBER

- The cell potential is approximately 1.35 V.
- It is used commonly in our hearing aids, watches, calculators etc.





PYQS' AND PRACTICE PROBLEMS



Unlike dry cell, the mercury cell has a constant cell potential throughout its useful life. Why?

In mercury cell, the cell potential remains constant during its life as [the overall reaction does not involve any ion in solution whose concentration can change during its life time.]

On the other hand, in a dry cell the (overall reaction contains ions which changes with concentration and gradually cell potential drops.)

Write the name of the cell which is generally used in hearing aids. Write the reactions taking place at the anode and the cathode of this cell.

Mercury cell

REACTION AT ANODE



REACTION AT CATHODE

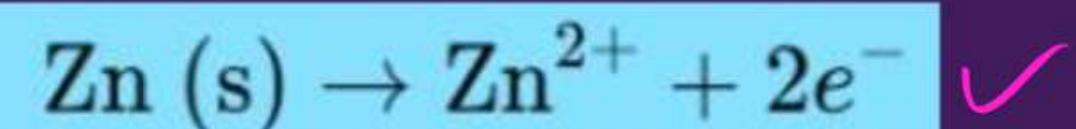
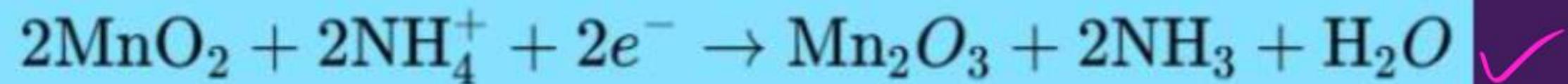
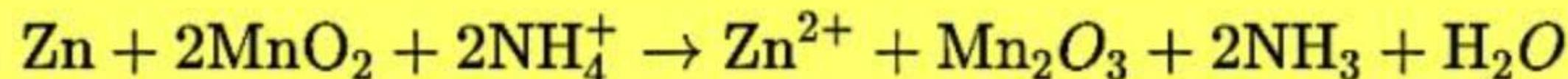


OVERALL REACTION



Write the name of the cell which is generally used in transistors. Write the reactions taking place at the anode and the cathode of this cell.

Leclanche Cell (DRY cell)

REACTION AT ANODE**REACTION AT CATHODE****OVERALL REACTION**

**SAMAJ AAYA TOH
LIKH DO.
AYE BHAIYA**



**INSANIYAT KA GYAAN
JO BANAE BEHTAR INSAN**

मन
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CONCEPT POLISH – HOMEWORK





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#sbsathhai ✓
#pwsathhai ✓

Thank
You

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2026

CHEMISTRY

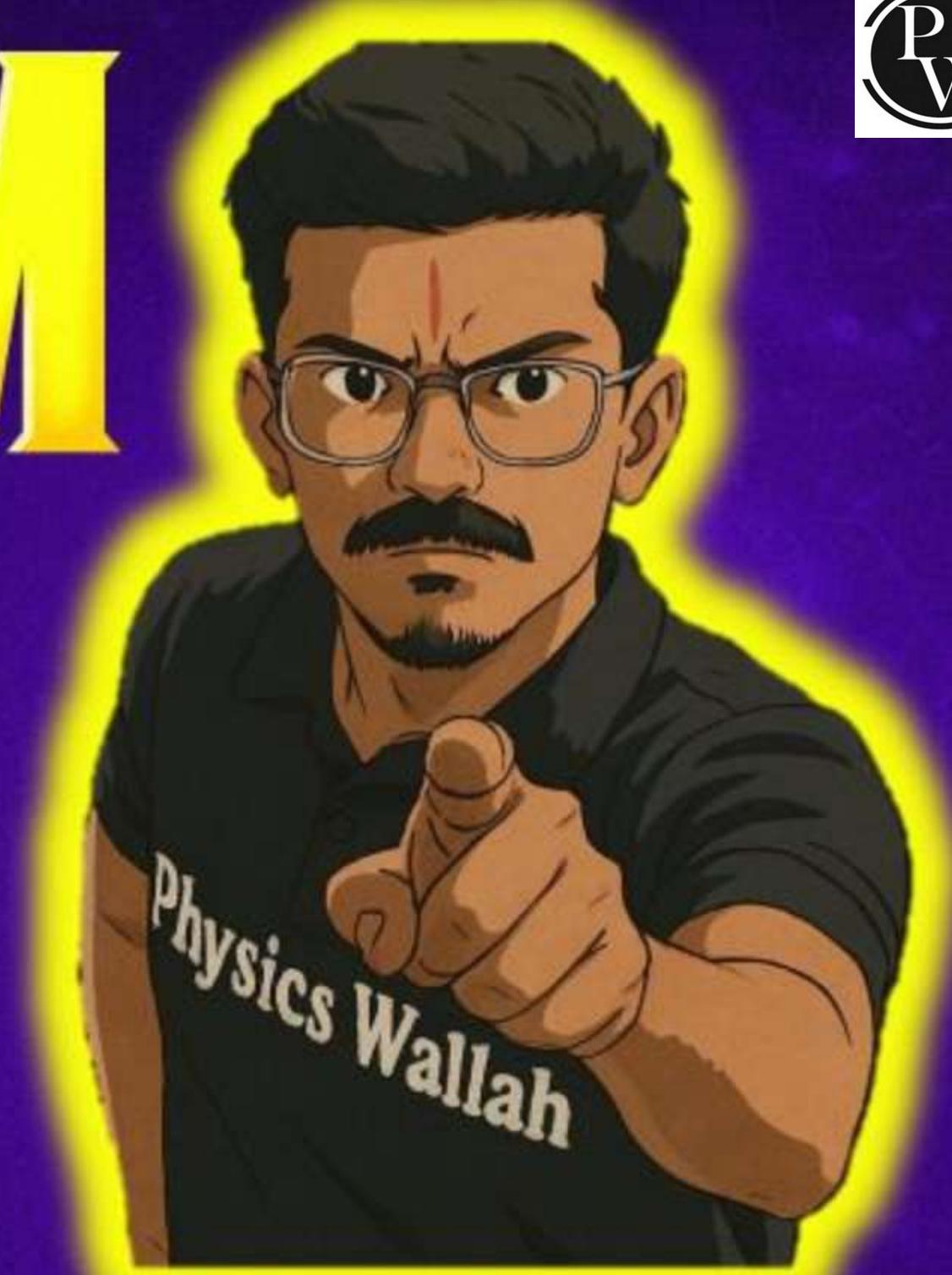
Lecture 13

ELECTROCHEMISTRY

Batteries – Part II, Fuel Cells and CBSE
PYQs'

Bharat Mata
Ki Jai ❤

BY – PRIYA-PUTRA-SUNIL



TOPICS TO BE COVERED

- (i) Batteries – Part II (✓)
- (ii) Fuel Cells (✓)
- (iii) CBSE Previous Year Questions (✓)



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A **cell** or a group of cells connected in series (**battery**) that we use as a source of electrical energy is basically a **galvanic cell** where the chemical energy of the redox reaction is converted into electrical energy.

'Recap'

BATTERIES – PART II

SECONDARY BATTERIES

(Reversible Redox Rxn)
^{II}

- It can be recharged by passing current through it in the opposite direction so that it can be used again.
- A good secondary cell can undergo a large number of discharging and charging cycles.

^I
- Recharged
- Reused

^{III}

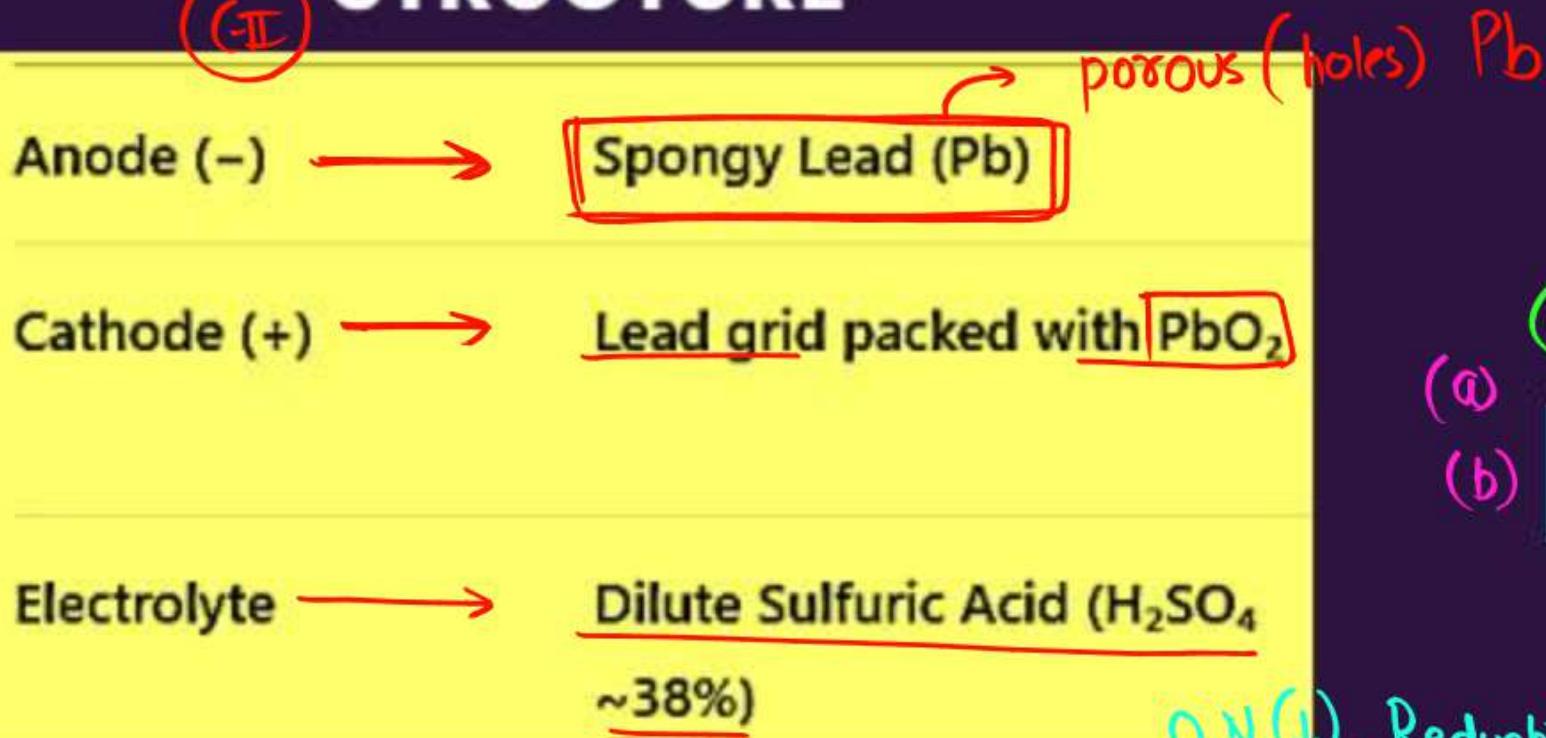
1 cycle : 1 complete discharge & charge

Ex: Life of a Ni-Cd battery → 1000 cycles

LEAD STORAGE BATTERY

Lead Acid
Battery

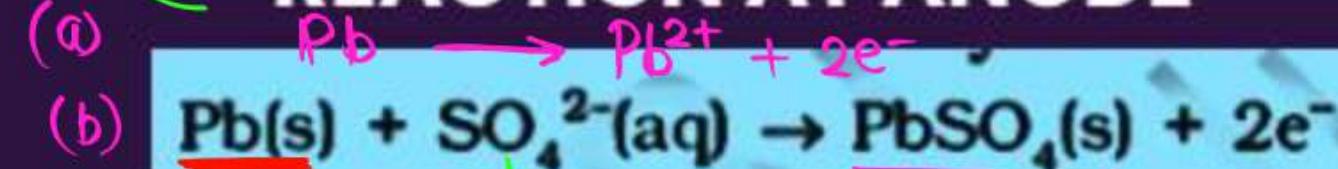
STRUCTURE



(C-I)

It is commonly used in
automobiles and invertors.

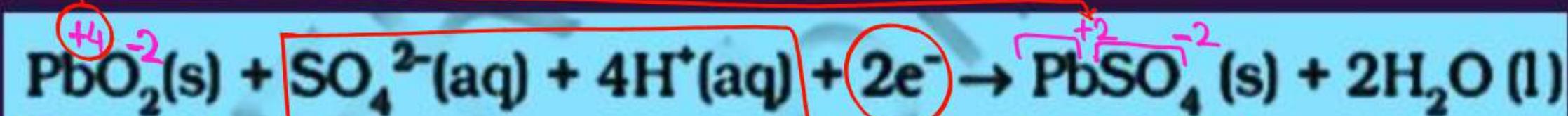
REACTION AT ANODE



from electrolyte

Oxidation → Loss of
electron(s)

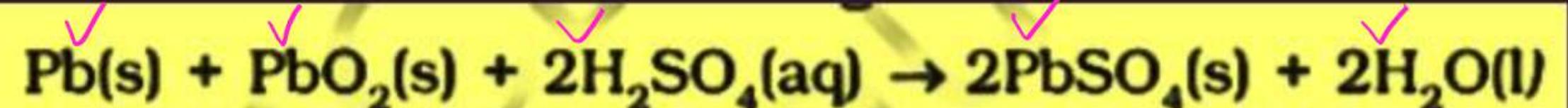
REACTION AT CATHODE



from electrolyte

Reduction → Gain of
electron(s)

OVERALL REACTION



GIVE A THOUGHT



On charging the battery the reaction is reversed and $\text{PbSO}_4(s)$ on anode and cathode is converted into Pb and PbO_2 , respectively.

- A. YES
- B. NO

*+

NICKEL-CADMIUM BATTERY

(C-I)

It is commonly used in emergency lights,
digital cameras etc.

NiCad
Battery

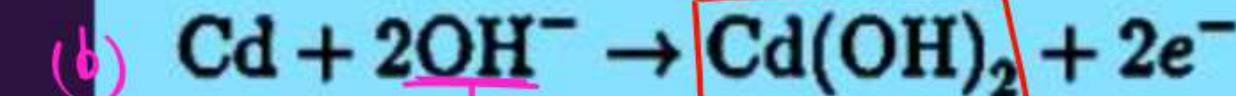
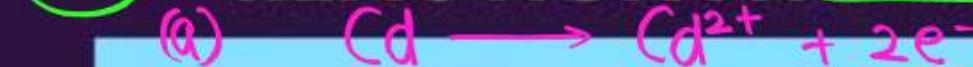
Ni-CD
Battery

(C-II) STRUCTURE

Part	Material Used
Anode (-) →	Cadmium (Cd)
Cathode (+) →	Nickel(III) oxide hydroxide (NiO(OH))
Electrolyte →	Aqueous Potassium Hydroxide (KOH) 'Water'

(C-III)

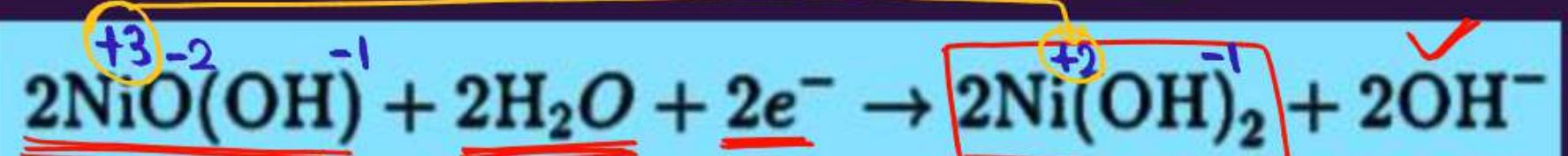
REACTION AT ANODE



from electrolyte

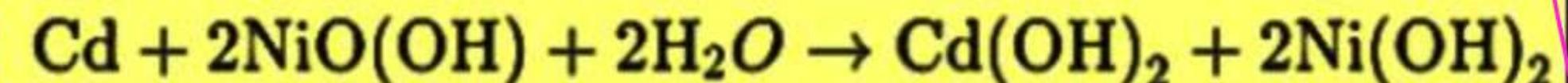
O.N. (↓), Reduction

REACTION AT CATHODE



Reduction

OVERALL REACTION



GIVE A THOUGHT

On charging the battery the reaction is reversed and
NiO(OH)(s) and Cd(s) are restored.

- A. YES
- B. NO

GIVE A THOUGHT

Ni-Cd battery has longer life than the lead storage battery but more expensive to manufacture.

- A. YES
- B. NO

IMPORTANT TO REMEMBER

Feature	◆ Primary Battery	◆ Secondary Battery
☛ Rechargeable?	<input checked="" type="checkbox"/> No (Single use only)	<input checked="" type="checkbox"/> Yes (Can be recharged)
☛ Usage	<u>Once discharged, cannot be reused</u>	<u>Discharged → can be recharged and reused</u>
☛ Reaction Type	Irreversible redox reactions	Reversible redox reactions
☛ Electrolyte	<u>Usually dry</u> (moist for flow of ions)	May be <u>liquid or gel</u>
☛ Example	<u>Dry cell (pencil battery), Mercury cell</u>	<u>Lead-acid battery, Ni–Cd battery, Li-ion battery</u>

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FUEL CELLS

The word "FUEL CELLS" is centered within a blue-outlined oval shape, which is itself set against a large yellow rounded rectangle.

FUEL CELLS

no release of heat & light energy

without combustion

- A fuel cell is an electrochemical cell that continuously converts the chemical energy of a fuel (like H_2) and oxidant (like O_2) into electrical energy – without combustion.

Fuel Cell → Galvanic Cell → Chemical directly into electrical energy

By combustion of coal (fuel)

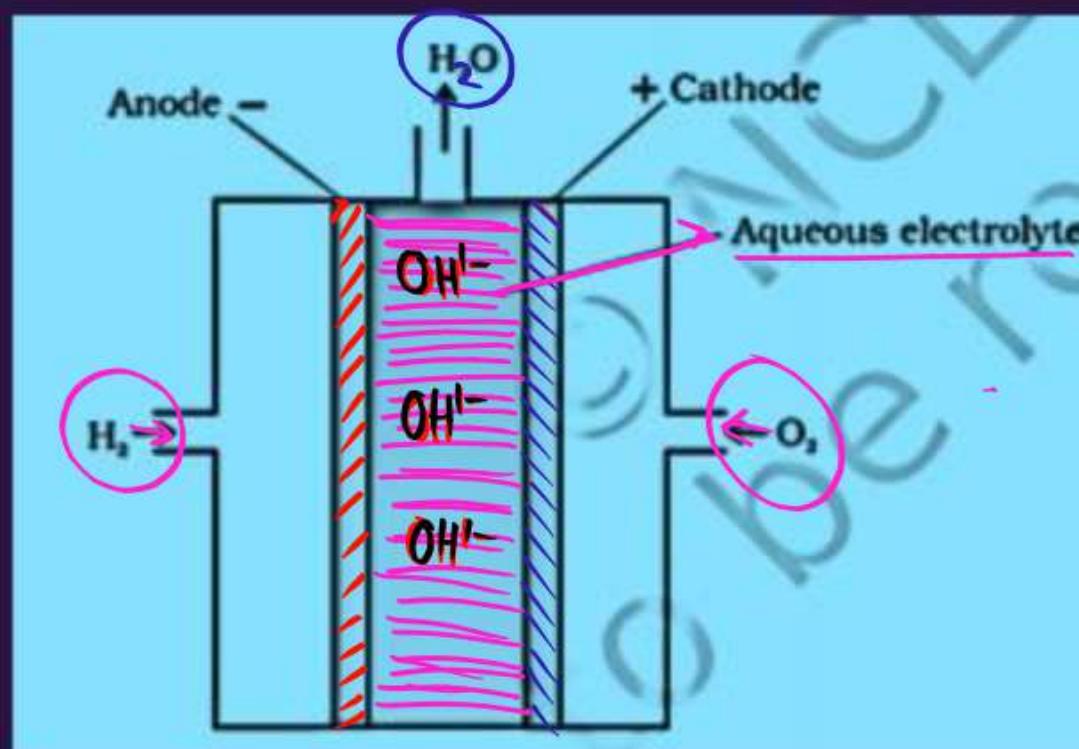
Need for Fuel Cells

- Thermal plants convert fuel → heat → steam → turbine → electricity. This is inefficient (efficiency is about 40%) and polluting.
- Fuel cells convert chemical energy directly to electricity. This is more efficient (efficiency is about 70%) and pollution free.

H₂-O₂ FUEL CELL

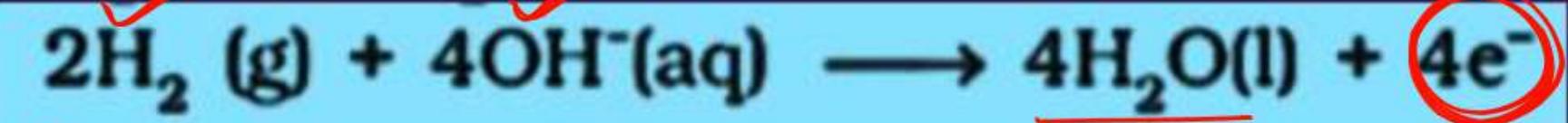
(C-II) STRUCTURE

Anode (-)	Hydrogen gas (H ₂) enters here
Cathode (+)	Oxygen gas (O ₂) enters here
Electrolyte	Aqueous KOH or OH ⁻ ions (alkaline medium)
Electrodes	Porous graphite with catalyst (Pt or Ni)



(C-I) Used for providing electrical power in the Apollo space programme.

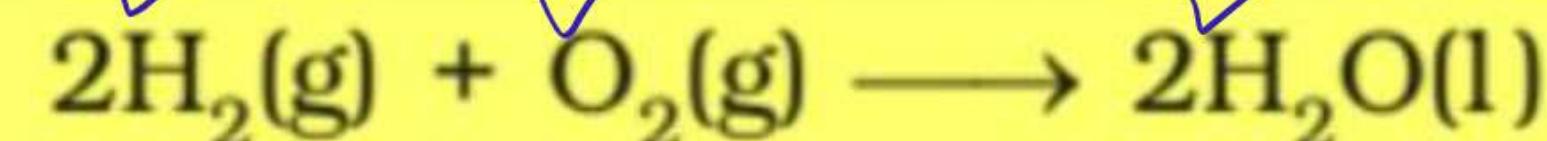
REACTION AT ANODE



REACTION AT CATHODE



OVERALL REACTION



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CBSE PREVIOUS YEAR QUESTIONS

From the given cells:

Lead storage cell, Mercury cell, Fuel cell and Dry cell

Answer the following:

- (i) Which cell is used in hearing aids?
- (ii) Which cell was used in Apollo Space Programme?
- (iii) Which cell is used in automobiles and inverters?
- (iv) Which cell does not have long life?

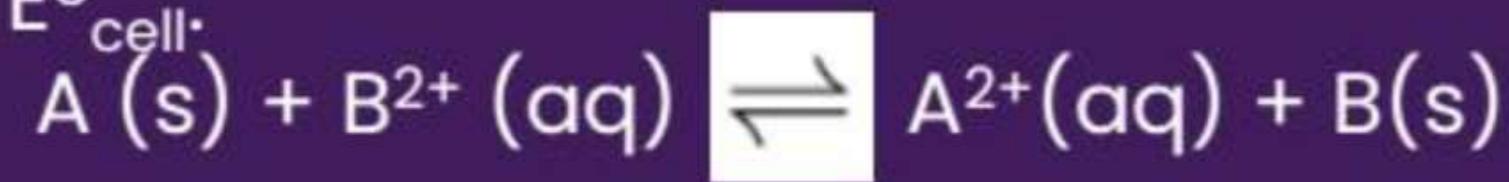
- (i) Mercury cell
- (ii) H_2-O_2 Fuel cell
- (iii) Lead storage cell
- (iv) Dry cell

What is the difference between primary and secondary battery? Give one example of each type.

Feature	◆ Primary Battery	◆ Secondary Battery
<input checked="" type="checkbox"/> Rechargeable?	<input checked="" type="checkbox"/> No (Single use only)	<input checked="" type="checkbox"/> Yes (Can be recharged)
<input checked="" type="checkbox"/> Usage	Once discharged, cannot be reused	Discharged → can be recharged and reused
<input checked="" type="checkbox"/> Reaction Type	Irreversible redox reactions	Reversible redox reactions
<input checked="" type="checkbox"/> Electrolyte	Usually dry	May be liquid or gel
<input checked="" type="checkbox"/> Example	Dry cell (pencil battery), Mercury cell	Lead-acid battery, Ni-Cd battery, Li-ion battery

- (i) Write two advantages of $\text{H}_2\text{-O}_2$ fuel cell over ordinary cell.
- (ii) Equilibrium constant (K_c) for the given cell reaction is 10.

Calculate E_{cell}° .



- (i) Advantages of $\text{H}_2\text{-O}_2$ fuel cell are:
- (a) Its efficiency is high.
- (b) It does not create pollution. (Pollution free)

$$\begin{aligned}
 (\text{ii}) \quad E_{\text{cell}}^{\circ} &= \frac{0.059}{n} \log K_c \\
 &= \frac{0.059}{2} \times \log 10 = \frac{0.059 \times 1}{2} = 0.0295
 \end{aligned}$$

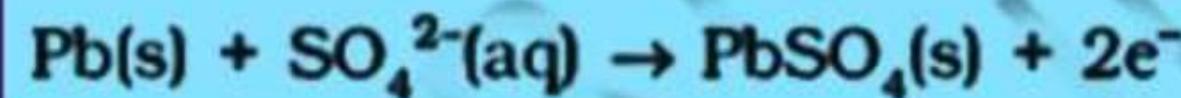
Write the name of cell used in inverters . Write the reactions taking place at cathode and anode of this cell?



- ① Lead storage battery

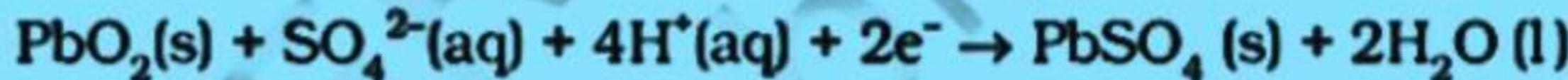
REACTION AT ANODE

(i)



REACTION AT CATHODE

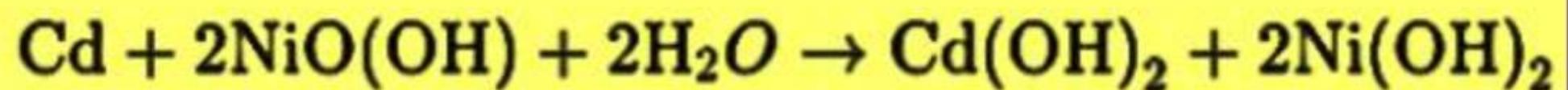
(ii)



Write the overall reaction that occurs during use (discharging) of nickel-cadmium cell. Is it a primary or a secondary cell? Mention its one merit over the lead storage cell.

① The overall reaction during discharge of battery:

⑪ Secondary Cell



 Ni-Cd battery has longer life than the lead storage battery.

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**INSANIYAT KA GYAAN
JO BANAE BEHTAR INSAN**

KARMA TRACKER

- Aaj maine kiski madad ki?
- Kya maine apne kaam mein imandari rakhi?
- Kya maine kisi ka dil dukhaya?

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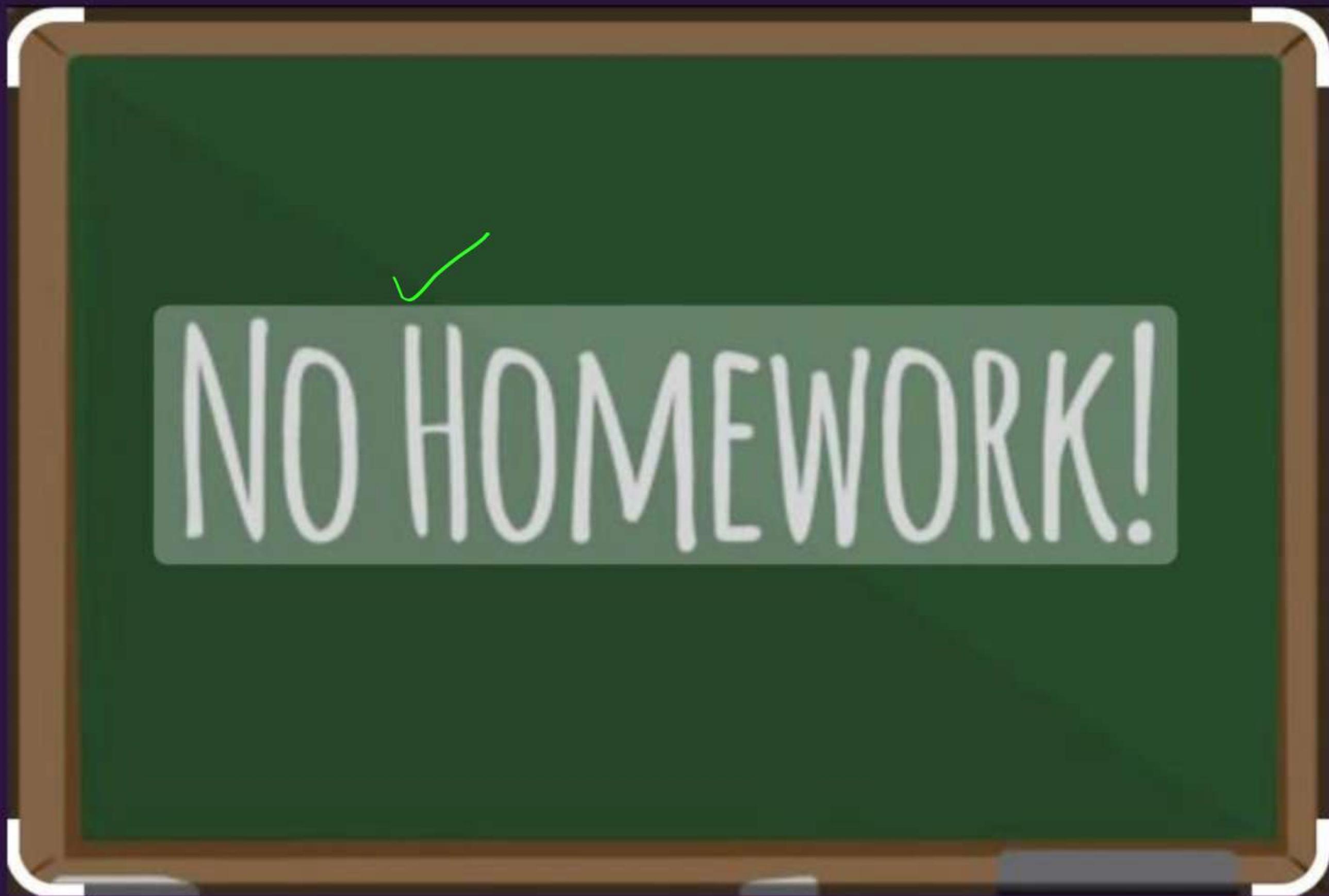


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CONCEPT POLISH – HOMEWORK





NO HOMEWORK!



SUNIL BHAIYA IS ALWAYS THERE FOR YOU.
#sbsathhai ✓
#pwsathhai ✓

Thank
You