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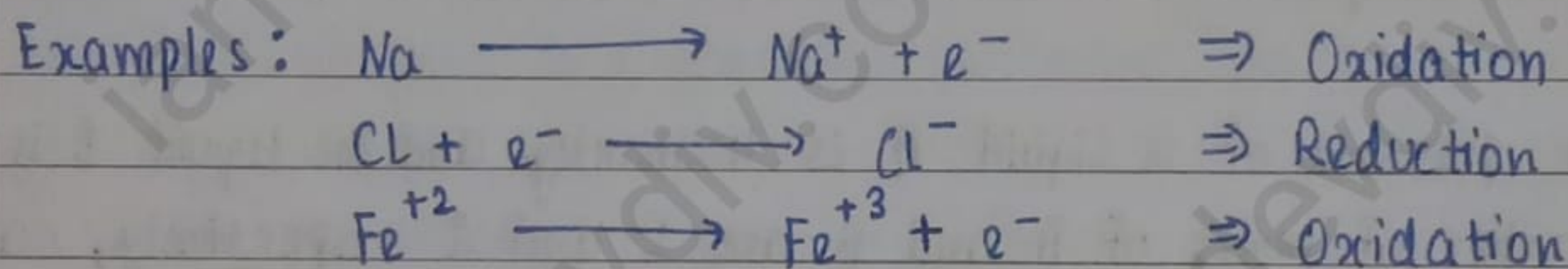
CHAPTER - 2

ELECTROCHEMISTRY

Electrochemistry is the branch of chemistry which deals with the study of chemical change on passing electric current or generation of electricity by carrying chemical reaction.

★ ELECTRONIC CONCEPT OF REDUCTION AND OXIDATION

- (i) Loss of electrons is called oxidation.
- (ii) Gain of electrons is called reduction.



★ ~~RESISTANCE~~

★ CONDUCTORS

METALLIC

- (i) Carrier of electric current is electron.
- (ii) No chemical change.
- (iii) It can conduct high amount of electric.

ELECTROLYTIC

- (i) Carrier of electric current is free ions.
- (ii) Chemical changes take place.
- (iii) Its conductivity is not so high.

• RESISTANCE

$$R = \frac{V}{I}$$

Unit of Resistance = Ω

• CONDUCTANCE

It is the reciprocal of resistance.

$$G = \frac{1}{R}$$

$$\text{Unit of } G = \frac{1}{\text{ohm}} = \text{ohm}^{-1} = \text{mho} = \text{siemen}$$

• SPECIFIC RESISTANCE / RESISTIVITY

$$R = \frac{\rho l}{A} \rightarrow \begin{array}{l} l \rightarrow 1 \text{ cm} \\ A \rightarrow 1 \text{ cm}^2 \end{array}$$

$$R = \rho$$

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

$$\text{Unit of } \rho = \frac{\text{ohm-cm}^2}{\text{cm}} = \text{ohm-cm} \rightarrow \text{C.G.S. Unit}$$

• SPECIFIC CONDUCTANCE / CONDUCTIVITY

$$\frac{1}{R} = K \quad (\text{Kappa})$$

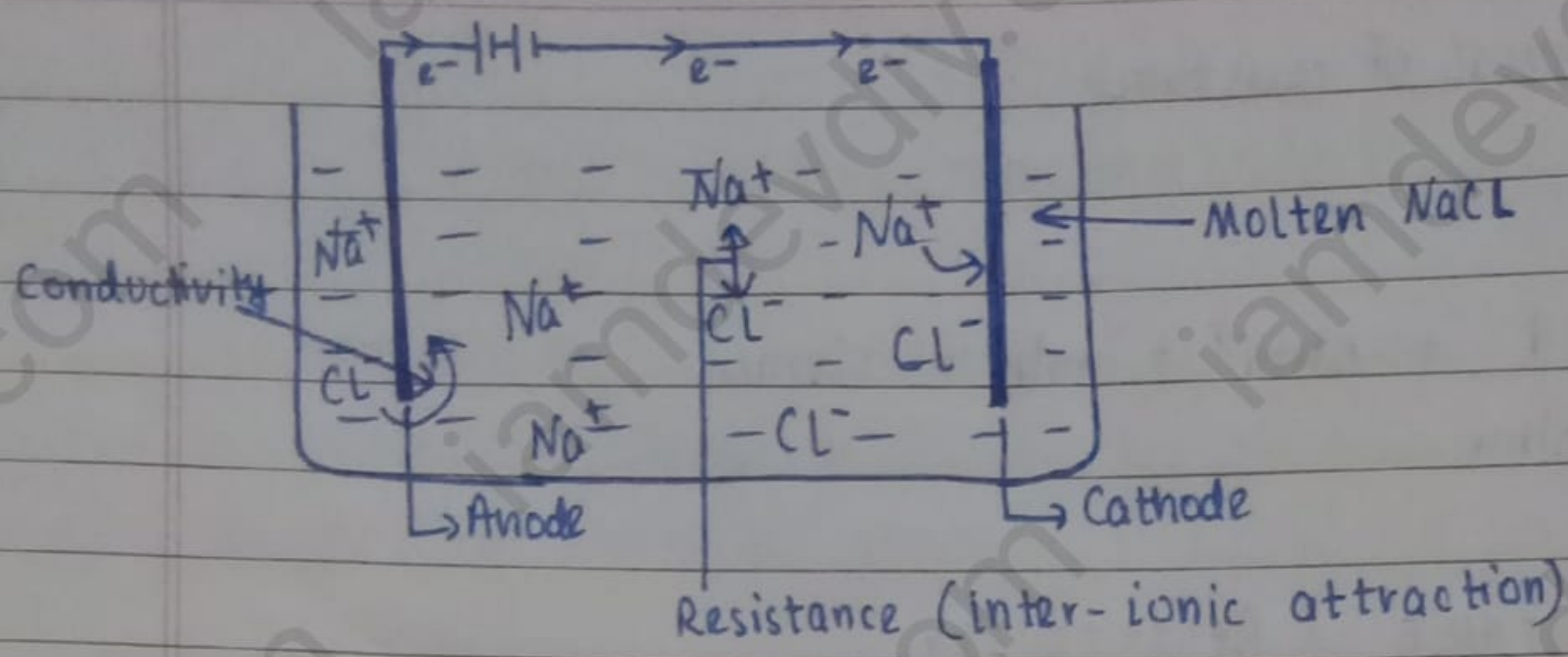
$$R = \frac{\rho l}{A}$$

$$\Rightarrow \frac{1}{R} = \frac{1}{\rho l} \cdot A$$

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

$$K = \frac{G l}{A} \rightarrow \begin{array}{l} l \rightarrow 1 \text{ cm} \\ A \rightarrow 1 \text{ cm}^2 \end{array}$$

$$\boxed{K = G}$$



• MOLAR CONDUCTIVITY [Λ_m] [DEFINITION]

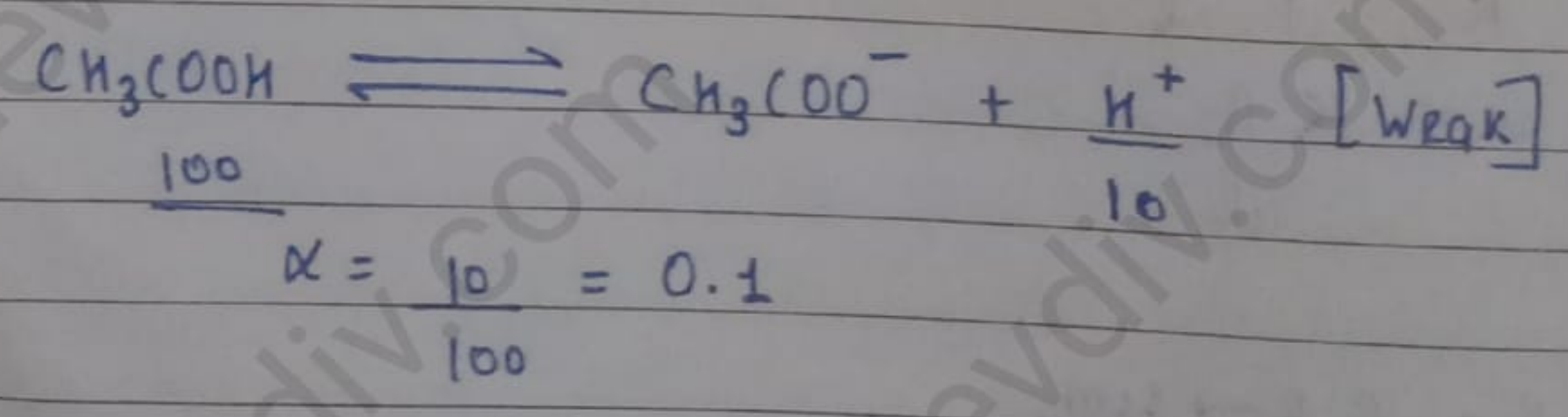
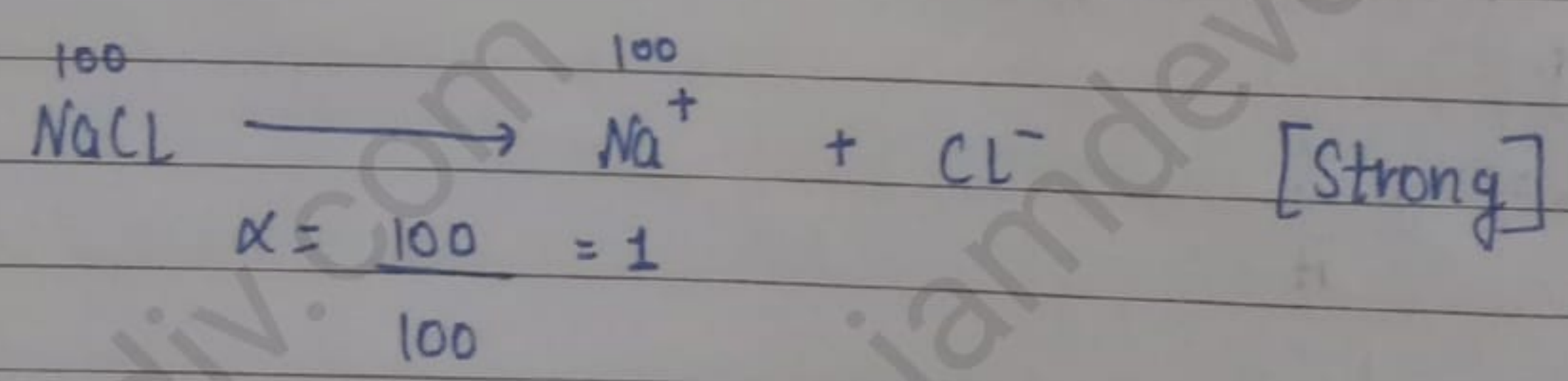
It is the conductivity of all the ions produced by one mole of electrolyte in the solution, when the dilution is $V \text{ cm}^3$.

• EQUIVALENT CONDUCTANCE [DEFINITION]

The conductance of all the ions produced by 1 gram equivalent of electrolyte in the solution and it is denoted by Λ_{eq} .

• STRONG AND WEAK ELECTROLYTE

$\hookrightarrow \alpha = 1$ $\hookrightarrow \alpha < 1$



• ~~DILUTION~~ MOLAR CONDUCTANCE [FORMULA]

Conductance of 1 cm^3 solution = K
 Conductance of $V \text{ cm}^3$ solution = $K \times V$

$$\Delta_m = 1 \text{ mol}$$

$$\text{molarity} = \frac{\text{no. of moles} \times 1000}{\text{Vol (ml)}}$$

$$\Rightarrow M = \frac{1000}{V}$$

$$\Rightarrow V = \frac{1000}{M}$$

$$\therefore \Delta_m = \frac{K \times 1000}{M}$$

\uparrow Molar conductance \uparrow Specific $M \leftarrow$ Molarity

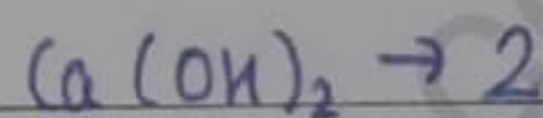
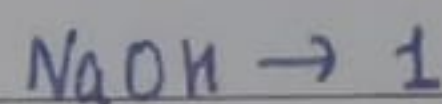
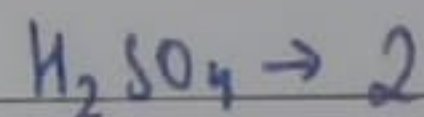
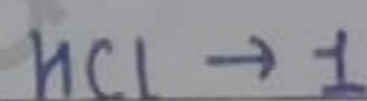
• EQUIVALENCE CONDUCTANCE [FORMULA]

$$\text{Normality} = \frac{\text{No. of gm equivalent}}{\text{Volume of solution}}$$

$$\text{No. of gram equivalent} = \frac{\text{Weight}}{\text{Eq. weight}}$$

$$\text{Eq. weight} = \frac{\text{Mol. weight}}{n\text{-factor}}$$

n-factors



≠ VARIATION OF CONDUCTIVITY WITH DILUTION

$$\text{Normality} = \frac{\text{No. of gm equivalent} \times 1000}{V(\text{ml})}$$

$$\Rightarrow V = \frac{1000}{N} \quad \text{①}$$

$$\Delta_m = \Delta_{eq} \times n\text{-factor}$$

$$\Delta_m = \frac{1000 \times K}{M}$$

• VARIATION OF CONDUCTIVITY WITH DILUTION

Specific conductance decreases on dilution because the number of ions per unit volume decreases and we know that the ions are the carrier of current and hence conductivity decreases.

The case is similar for both strong and weak electrolyte.

• MOLAR CONDUCTIVITY AND EQUIVALENT CONDUCTIVITY

(i) For strong electrolyte

On dilution, the number of ions do not increase because $\alpha = 1$ for strong electrolyte but the volume of solution increases which causes increment in inter-ionic distance, or resistance decreases and hence conductivity increases.

(ii) For weak electrolyte

a) The distance between ions increases on dilution and hence the conductivity increases.

$$\begin{array}{ccc} \alpha \downarrow \propto & \frac{1}{\sqrt{c}} \uparrow & \\ \text{Degree of dissociation} & \text{Concentration} & \end{array}$$

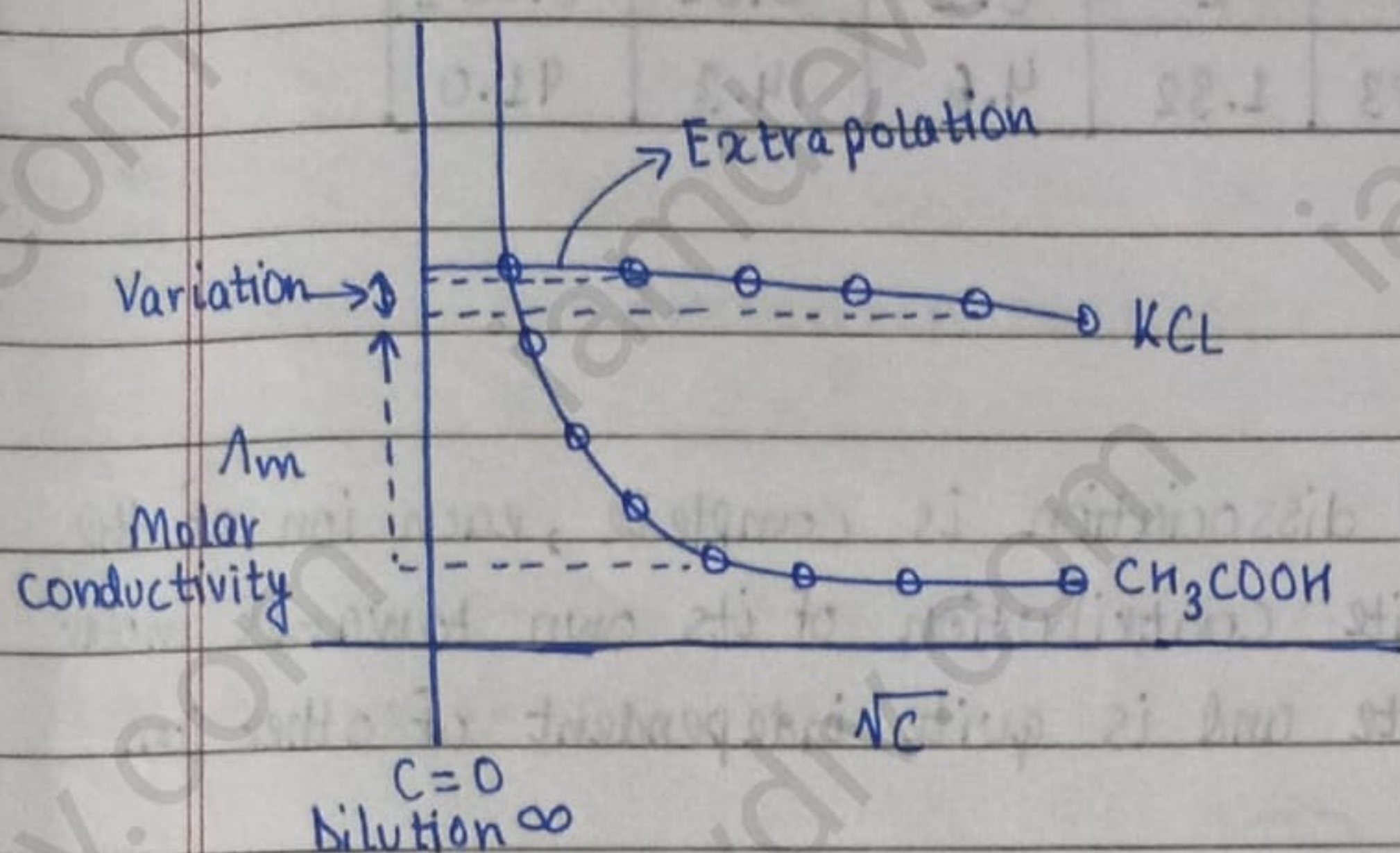
b) The conductivity increases on dilution. This can be related with Ostwald's dilution law.

According to this law, $\alpha \propto \frac{1}{\sqrt{c}}$

[For weak electrolyte]

As the concentration decreases (dilution increases) the no. of dissociation/dissociated ions increases according to above equation and we know that when the no. of ions increases conductance.

★ GRAPH FOR MOLAR CONDUCTIVITY OF WEAK AND STRONG ELECTROLYTES



Limiting $\Lambda_m \rightarrow \Lambda_m^\infty$

• STRONG ELECTROLYTES

This increase in molar conductivity will continue till a stage is reached when the inter-ionic forces become negligible and further increase is not possible. This is known as limiting molar conductivity $[\Lambda_m^\infty]$.

• DEBYE HUCKEL ONSAGER

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

where, $A =$ constant which depends upon the nature of solvent and temperature

For KCl

Concentration	3	2	1	0.1	0.01	0.001
Λ_m (KCl)	88.3	92.6	98.3	112.4	122	127.3

• WEAK ELECTROLYTES

- (i) This extrapolation is not possible for weak electrolytes.
- (ii) The increase in molar conductance is observed on dilution which is very high because it depends upon two factors.

For CH_3COOH

Concentration	3	2	1	0.1	0.01	0.001
$\Lambda_m(\text{CH}_3\text{COOH})$	0.54	0.83	1.32	4.6	14.3	41.0

★ KOHLRAUSCH LAW

- (i) At infinite dilution when dissociation is complete, each ion of the electrolyte makes a definite contribution of its own towards molar conductivity of electrolyte and is quite independent of other ion.
- (ii) This law states that limiting molar conductivity of an electrolyte can be represented as the ~~sum~~ ^{sum} of the individual contribution of anions and cations of the electrolyte.

$$\Lambda_m^\infty(\text{NaCl}) = \Lambda_m^\infty(\text{Na}^+) + \Lambda_m^\infty(\text{Cl}^-)$$

$$\Lambda_m^\infty(\text{NaBr}) = \Lambda_m^\infty(\text{Na}^+) + \Lambda_m^\infty(\text{Br}^-)$$

• APPLICATION OF KOHLRAUSCH LAW

- ① This law helps us to calculate limiting molar conductivity for any electrolyte from the conductivity of each ion.

$$\text{Formula} \Rightarrow \textcircled{1} \Lambda_m^\infty(\text{NaCl}) = \lambda_m^\infty(\text{Na}^+) + \lambda_m^\infty(\text{Cl}^-)$$

$$= 50.1 + 76.3$$

$$= 126.4 \text{ Scm}^2/\text{mol}$$

$$\textcircled{2} \Lambda_m^\infty(\text{CaCl}_2) = \lambda_m^\infty(\text{Ca}^{2+}) + 2 \times \lambda_m^\infty(\text{Cl}^-)$$

$$= 119.0 + 2 \times 76.3$$

$$= 271.6 \text{ Scm}^2/\text{mol}$$

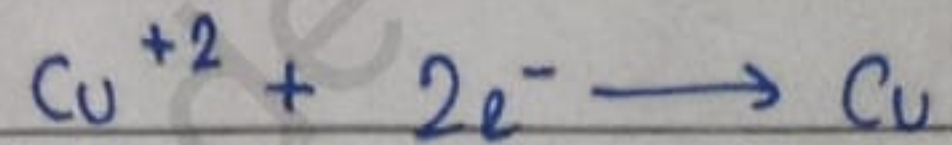
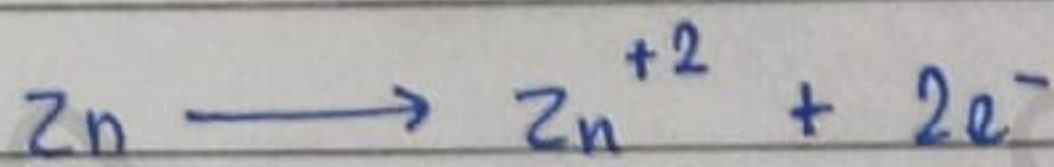
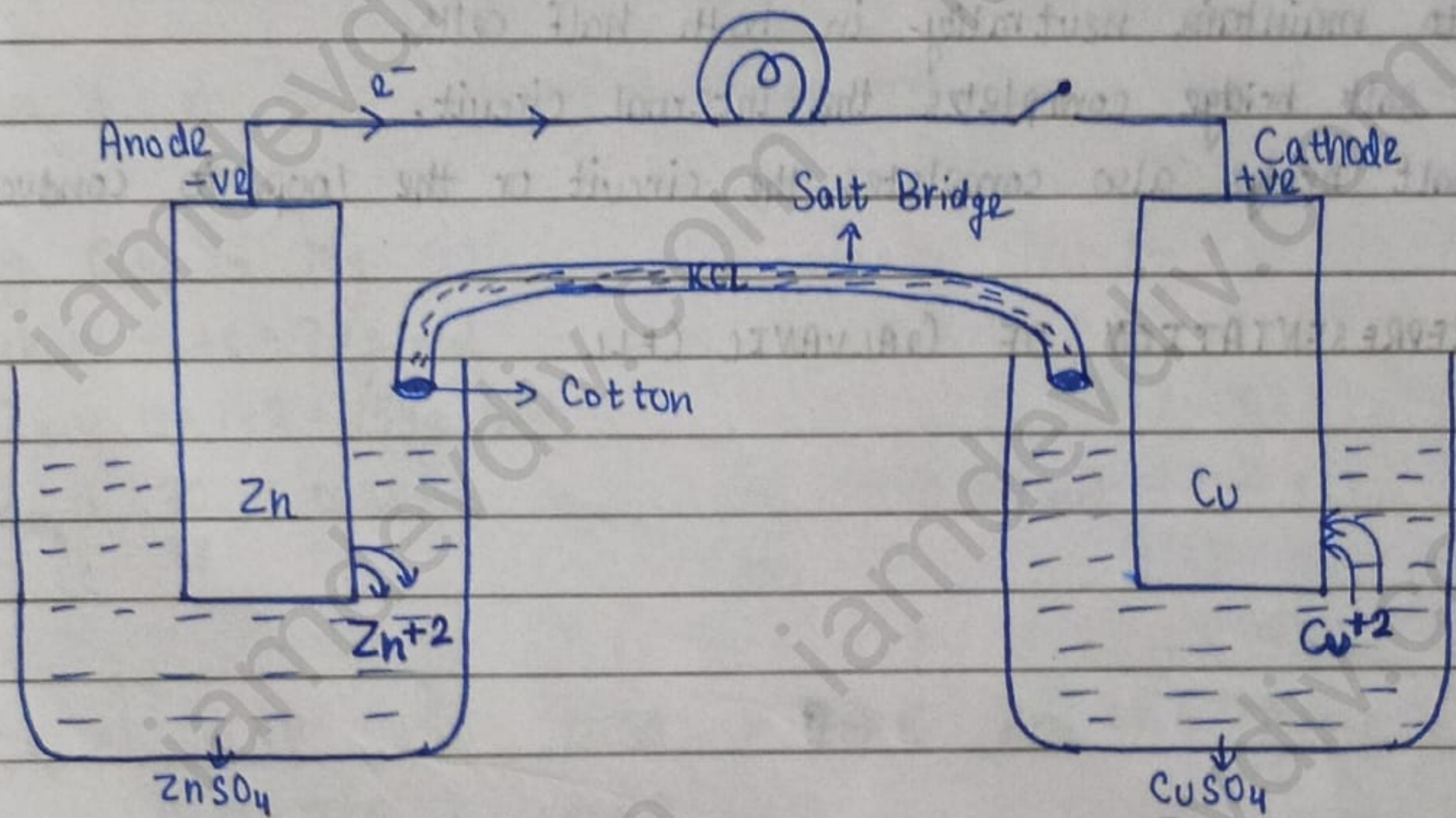
- ② Calculation of degree of dissociation (α) and dissociation constant
It can be calculated with the help of the following formula

degree of dissociation, $\alpha = \frac{\Lambda_m^c}{\Lambda_m^\infty}$ (molar conductance at any conc. c)
 Λ_m^∞ (molar conductance at infinite dilution)
 \rightarrow any value other than max
 \rightarrow maximum value $c \rightarrow 0$

Formula for dissociation constant,

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

★ GALVANIC CELL



L \rightarrow Left

O \rightarrow Oxidation

A \rightarrow Anode

N \rightarrow -ve

R \rightarrow Right

R \rightarrow Reduction

C \rightarrow Cathode

P \rightarrow +ve

A device in which the redox reaction is carried indirectly and decrease in free energy appears as the electrical energy.

• OBSERVATIONS

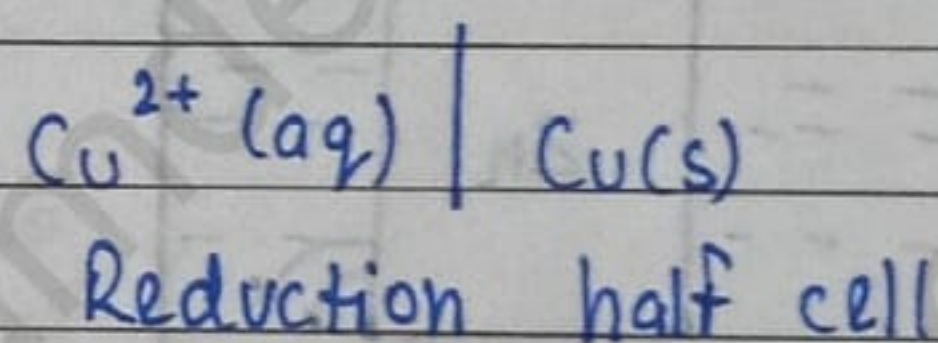
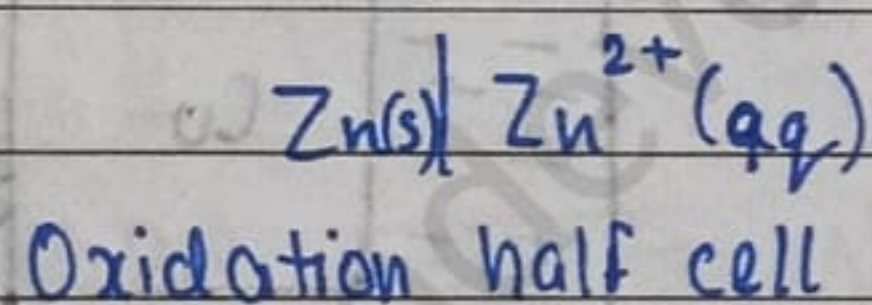
iv) Zinc rod will dissolve in the solution.

- (ii) Copper will deposit on the copper rod.
- (iii) External circuit \rightarrow wire \rightarrow electron \rightarrow anode \rightarrow cathode and electricity \rightarrow cathode \rightarrow anode
- (iv) Internal circuit \rightarrow salt bridge \rightarrow ions \rightarrow anode \rightarrow cathode
- (v) Solution in both beakers remain electrically neutral.

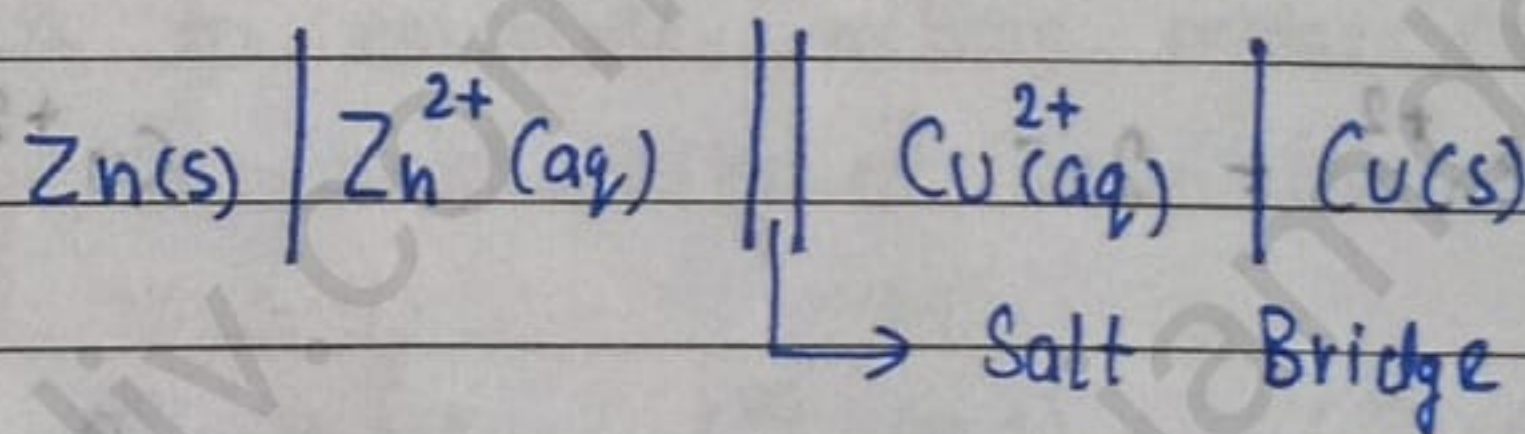
• SALT BRIDGE

- (i) Not concentration solution of any salt made in agar-agar jelly.
- (ii) It sends its ions to anode (anions) and cation to cathode in order to maintain neutrality in both half cells.
- (iii) A salt bridge completes the internal circuit.
- (iv) Salt bridge also completes the circuit or the loop to conduct electricity.

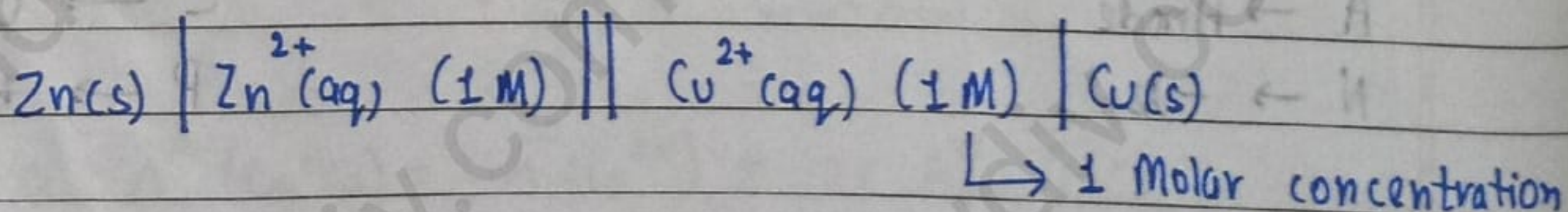
• REPRESENTATION OF GALVANIC CELL



OR



OR



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Electrochemistry continue after page no. 49

★ ELECTRODE POTENTIAL

The potential difference between the electrode and its ions in the solution is called electrode potential.

OR

It may also be regarded as the tendency of electrodes to lose or gain e^- when it is in contact with the solution of its own ions.

These are of two types:

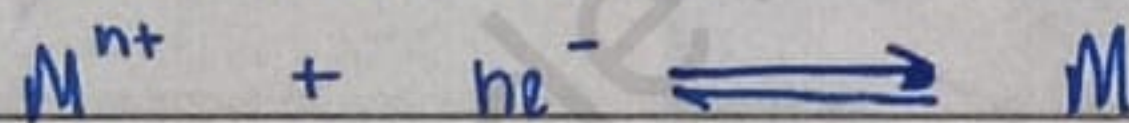
1) OXIDATION POTENTIAL



(at equilibrium)

This is called oxidation potential because in this reaction, e^- are lost.

2) REDUCTION POTENTIAL



(at equilibrium)

This is called reduction potential because in this reaction, e^- are gained.

• FACTORS EFFECTING ELECTRODE POTENTIAL

- 1) Nature of the metal
- 2) Concentration of the ions
- 3) Temperature

• STANDARD ELECTRODE POTENTIAL

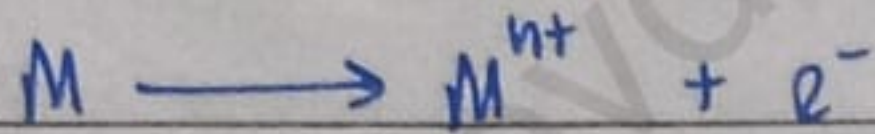
When temperature = 298K

pressure = 1 bar

concentration = 1 molar,

the potential of electrode and its ions is called standard electrode potential.

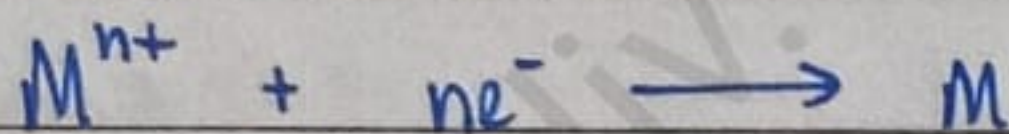
* OXIDATION POTENTIAL



$$E_{M/M^{n+}} \Rightarrow \text{Oxidation potential}$$

When this oxidation potential is measured at these conditions, it will be considered as $\Rightarrow E_{M/M^{n+}}^{\circ} \Rightarrow$ Standard oxidation potential

* REDUCTION POTENTIAL



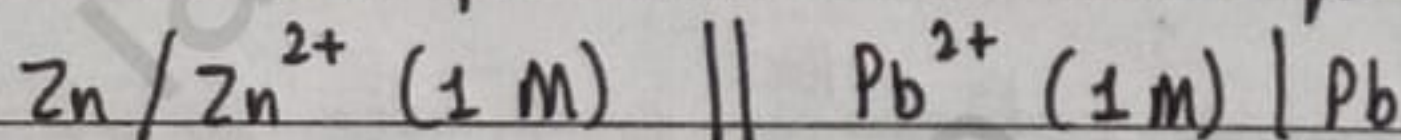
$$E_{M^{n+}/M} \Rightarrow \text{Reduction potential}$$

When this reduction potential is measured at these conditions, it will be considered as $\Rightarrow E_{M^{n+}/M}^{\circ} \Rightarrow$ Standard reduction potential.

NOTE \rightarrow According to IUPAC, standard reduction potential is always taken as standard electrode potential.

Find

Q. The standard cell potential for the equation



Given $E_{Zn/Zn^{2+}}^{\circ} = +0.76V$ and $E_{Pb^{2+}/Pb}^{\circ} = +0.13V$

SOL. $E_{Zn/Zn^{2+}}^{\circ} = +0.76V \Rightarrow E_{Zn^{2+}/Zn}^{\circ} = -0.76V$

$$E_{Pb^{2+}/Pb}^{\circ} = +0.13V$$

$$EMF = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$

$$= 0.13 - (-0.76)$$

$$= 0.89V \quad \text{Ans}$$

NOTE \rightarrow The EMF of a cell is always the difference between standard reduction potential of the two ~~cells~~ electrodes.

★ STANDARD HYDROGEN ELECTRODE

1. It is a gas ion electrode.
2. It is a reference electrode by IUPAC.
3. It can act as both anode half cell and cathode half cell.
4. Its standard reduction potential and standard oxidation potential is always zero.

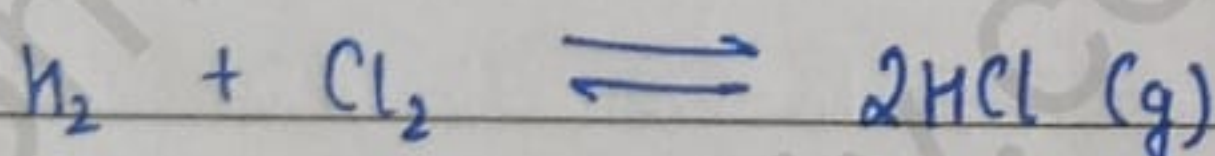
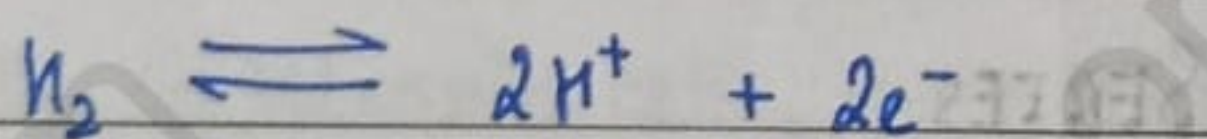
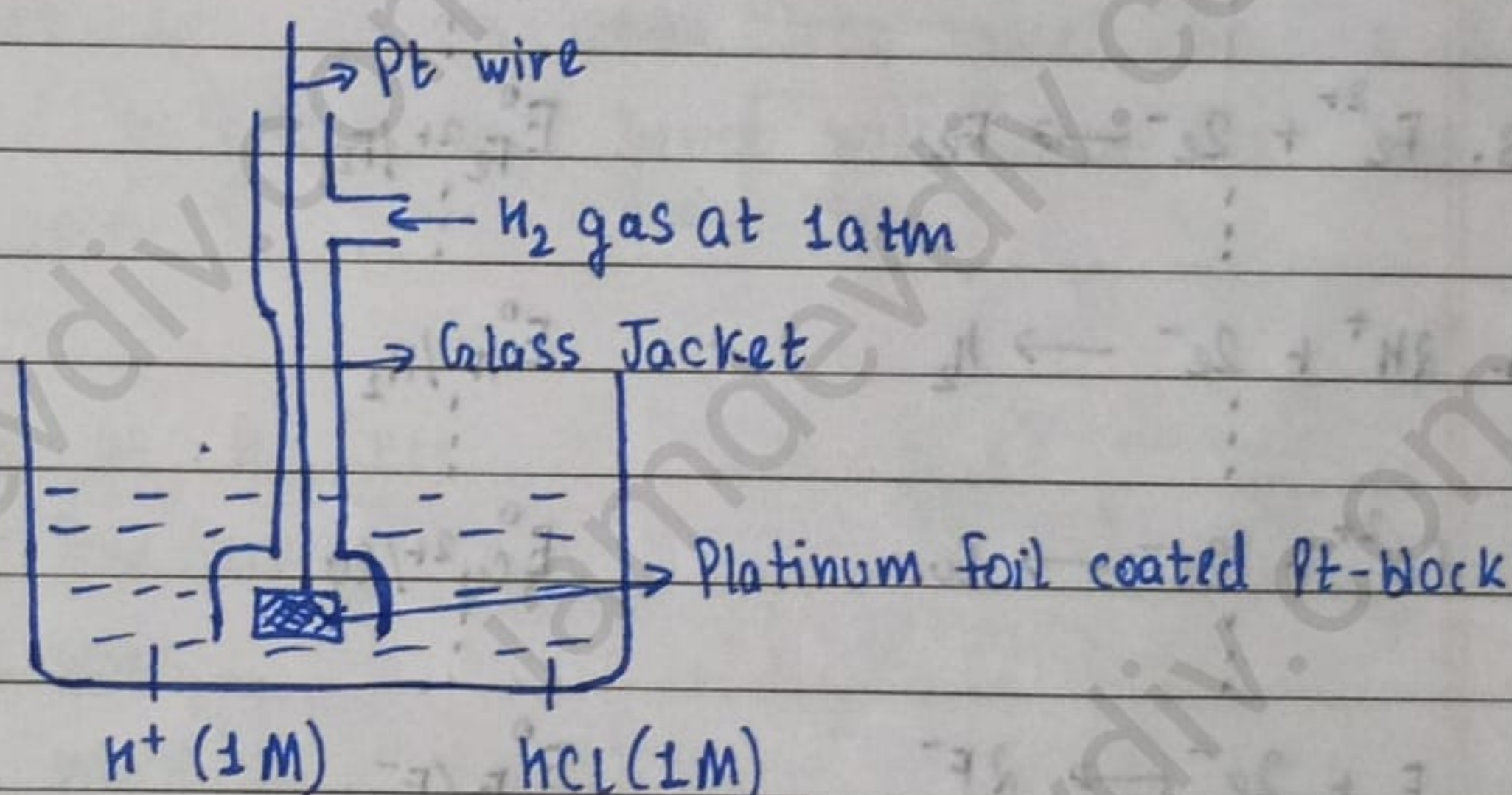
$$E^{\circ}_{\text{H}^+/\text{H}_2} = 0$$

$$E^{\circ}_{\text{H}_2/\text{H}^+} = 0$$

$$E^{\circ}_{\text{H}^+/\frac{1}{2}\text{H}_2} = 0$$

$$E^{\circ}_{\frac{1}{2}\text{H}_2/\text{H}^+} = 0$$

• CONSTRUCTION OF STANDARD HYDROGEN ELECTRODE



• LIMITATIONS

1. It is not possible to maintain 1 atm pressure for hydrogen gas for long time.
2. H^+ concentration cannot be maintained at 1M (molar) since HCl being a gas corporates from the solution.
3. Difficult to get pure and dry hydrogen gas.

★ ELECTROCHEMICAL SERIES

1. Elements are arranged in order of increasing standard reduction potential.
2. Standard reduction potential is taken with respect to standard hydrogen electrode.

REACTION	REPRESENTATION	STANDARD REDUCTION POTENTIAL	
1. $\text{Li}^+ + e^- \rightarrow \text{Li}$	$E^\circ_{\text{Li}^+/\text{Li}}$	-3.05 V	Tendency to get reduced ↓
2. $\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}$	$E^\circ_{\text{Zn}^{2+}/\text{Zn}}$	-0.76 V	
3. $\text{Fe}^{2+} + 2e^- \rightarrow \text{Fe}$	$E^\circ_{\text{Fe}^{2+}/\text{Fe}}$	-0.45 V	
4. $2\text{H}^+ + 2e^- \rightarrow \text{H}_2$	$E^\circ_{\text{H}^+/\text{H}_2}$	0 V	
5. $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$	$E^\circ_{\text{Cu}^{2+}/\text{Cu}}$	+0.34 V	
6. $\text{F}_2 + 2e^- \rightarrow 2\text{F}^-$	$E^\circ_{\text{F}_2/\text{F}^-}$	+2.87 V	

• APPLICATIONS OF ELECTROCHEMICAL SERIES

1. Comparison of relative reducing or oxidizing power of the element

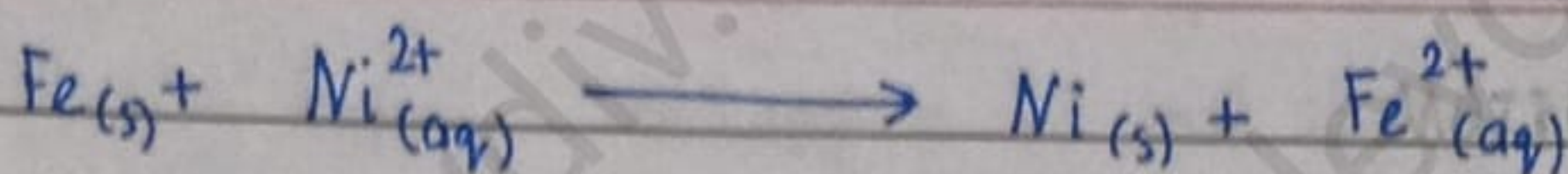
a) $E^\circ_{\text{Zn}^{2+}/\text{Zn}} = -0.76\text{V}$

b) $E^\circ_{\text{Ni}^{2+}/\text{Ni}} = -0.25\text{V}$

From the above values we can predict that Ni is a good oxidizing agent than Zn and vice-versa in case of reducing agent.

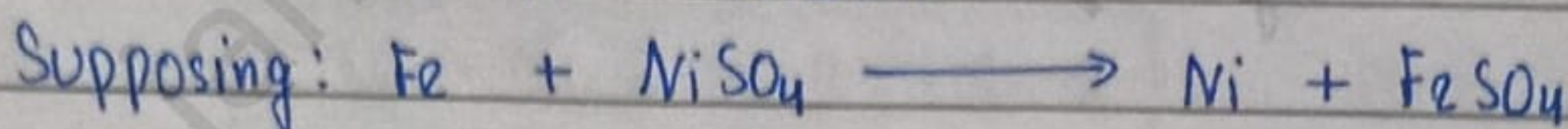
2. Predicting the Spontaneity / Feasibility of reaction

a) Redox reaction is feasible if e^- releasing species is present at the position lower in the series



Given: E° of Fe = -0.44V

E° of Ni = -0.25V



This reaction is feasible because Fe^{2+} is placed in the upper position as compared to Ni^{2+} .

Q Can we store CuSO_4 in the Fe vessel?

$$E^\circ_{\text{Fe}^{2+}/\text{Fe}} = -0.44\text{V}$$

$$E^\circ_{\text{Cu}^{2+}/\text{Cu}} = +0.34\text{V}$$

Sol. No, we cannot store Cu in an ~~iron~~ iron vessel because Cu having higher reduction potential and hence will accept e^- and will change into solid copper.

★ CALCULATION OF ΔG OF A CELL

$\Delta G =$ Gibb's free energy

$\Delta G = -$ useful work

$\Delta G = |\text{useful work}|$

In a galvanic cell, the useful work is equal to electrical work done by a cell

$$W_{\text{useful}} = Q \times V$$

$\begin{array}{l} \searrow \text{EMF} \\ \rightarrow \text{Charges transferred} \end{array}$

$$Q \text{ transferred} = -nF$$

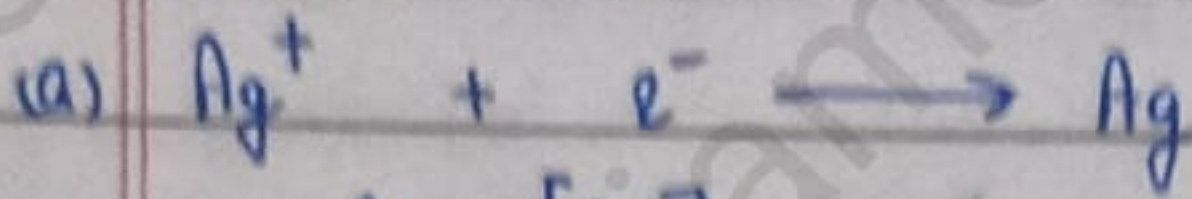
$$= -nF \times E^\circ_{\text{cell}}$$

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

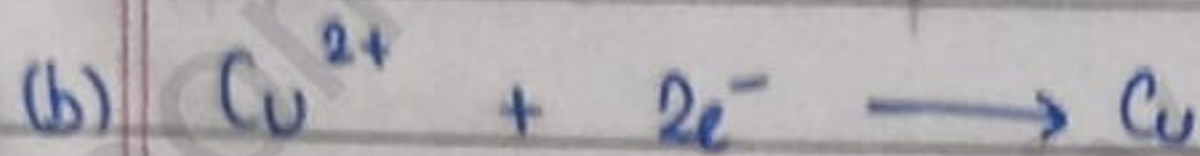
$$\Delta G = -nFE_{\text{cell}} \Rightarrow \text{Non-standard}$$

★ NERNST EQUATION

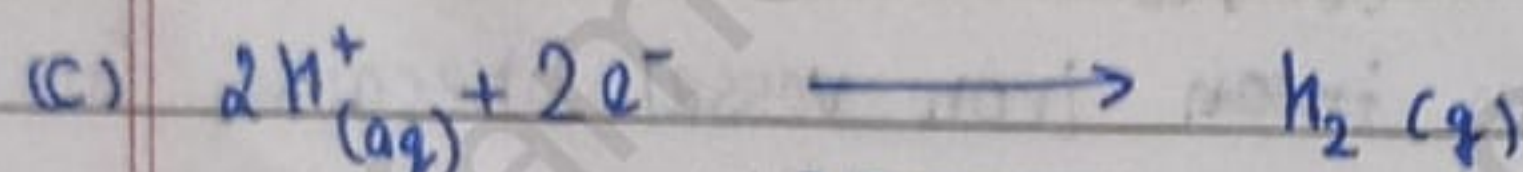
1. Determination of Reaction quotient



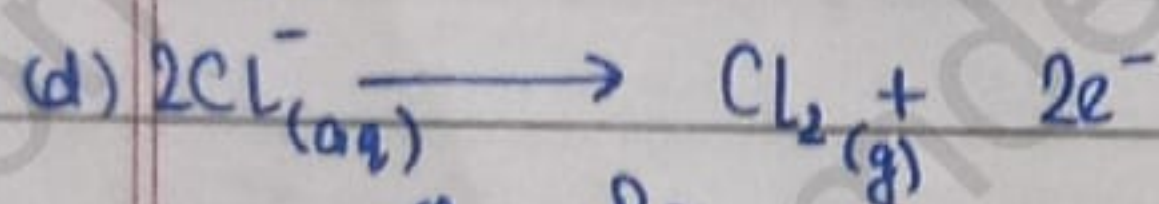
$$Q = \frac{[\text{Ag}]}{[\text{Ag}^+]}$$



$$Q = \frac{[\text{Cu}]}{[\text{Cu}^{2+}]}$$



$$Q = \frac{P_{\text{H}_2}}{[\text{H}^+]^2} \quad \left. \begin{array}{l} \text{---} \\ \text{---} \end{array} \right\} \text{Pressure for gases}$$



$$Q = \frac{P_{\text{Cl}_2}}{[\text{Cl}^-]^2}$$

2. Nernst equation tells us about effect of change in concentration on electrode potential.

Now the question arises why conc. changes and how to calculate the change

Because at anode, the concentration is going to increase due to oxidation

$[\text{M} \longrightarrow \text{M}^{n+} + ne^-]$ but in case of cathode the concentration is going to decrease due to reduction $[\text{M}^{n+} + ne^- \longrightarrow \text{M}]$

3. Nernst suggested a formula related to change in concentration which is called Nernst equation.

According to Gibbs Helmholtz equation

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

$$\Rightarrow -nF E_{\text{cell}} = -nF E_{\text{cell}}^{\circ} + RT \ln Q$$

Dividing by nF

$$\Rightarrow -E_{\text{cell}} = -E_{\text{cell}}^{\circ} + \frac{RT \ln Q}{nF}$$

$$\Rightarrow E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT \ln Q}{nF}$$

where, R = gas constant T = temperature Q = Reaction quotient

$$\text{As } \ln = 2.303 \log_{10}$$

$$\Rightarrow E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{2.303 RT \log_{10} Q}{nF}$$

$$R = 8.314 \text{ J/K mol}$$

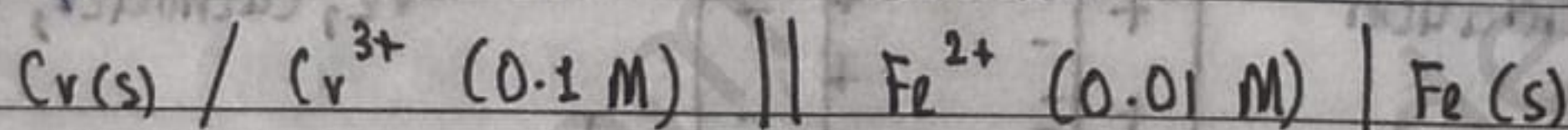
$$T = 25^{\circ}\text{C} = 298\text{K}$$

$$F = 96500 \text{ C/mol}$$

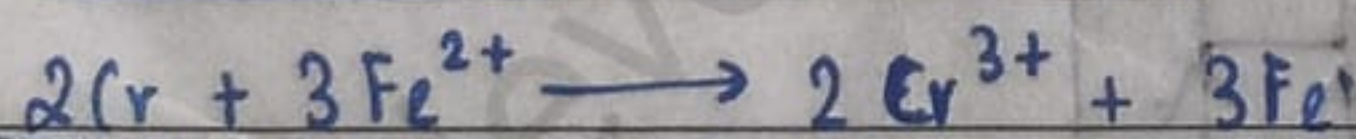
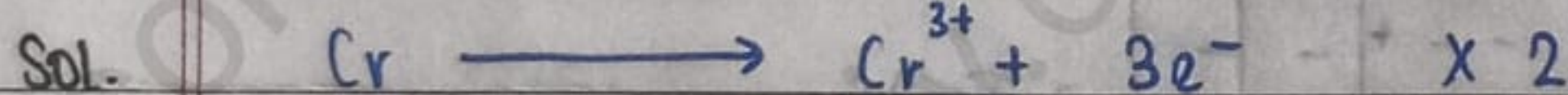
always

$$\Rightarrow E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591 \log_{10} Q}{n}$$

Q Calculate the EMF of the cell at 298K



$$\text{Given: } E_{\text{cell}}^{\circ} = +0.30\text{V}$$



$$n = 6$$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.06 \log_{10} \frac{[\text{Cr}^{3+}]^2 [\text{Fe}^{2+}]^3}{[\text{Fe}^{2+}]^3 [\text{Cr}^{3+}]^2}}{6}$$

$$= 0.30 - 0.01 \times \log_{10} \frac{10^{-2}}{10^{-6}}$$

$$= 0.30 - 0.01 \times 4$$

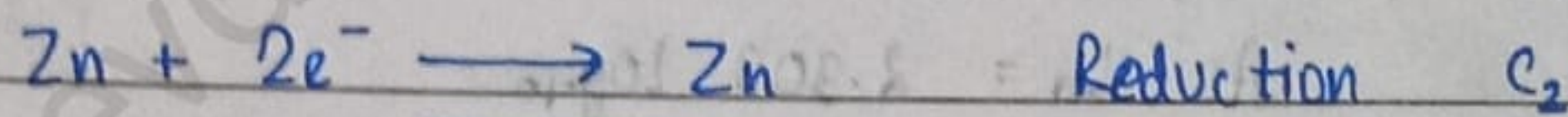
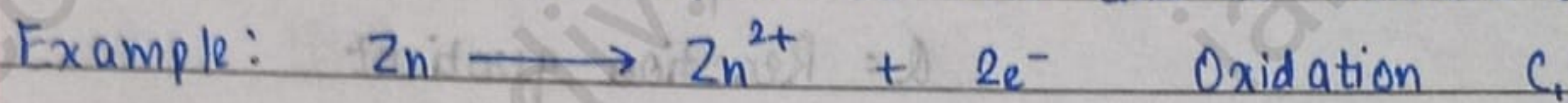
$$= 0.30 - 0.04$$

$$= 0.26\text{V} \quad \text{Ans}$$

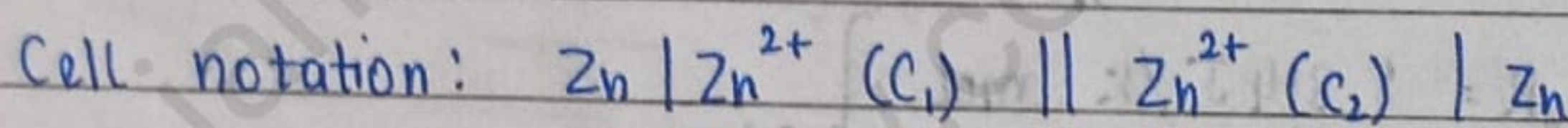
★ CONCENTRATION CELL

1) Cells can also be prepared in which the concentration of the same electrolyte solution or the material of the electrode may be different. However, the chemical reactions are same, such electrochemical cells are known as concentration cell.

2) The concentration at anode is less and at cathode is more.

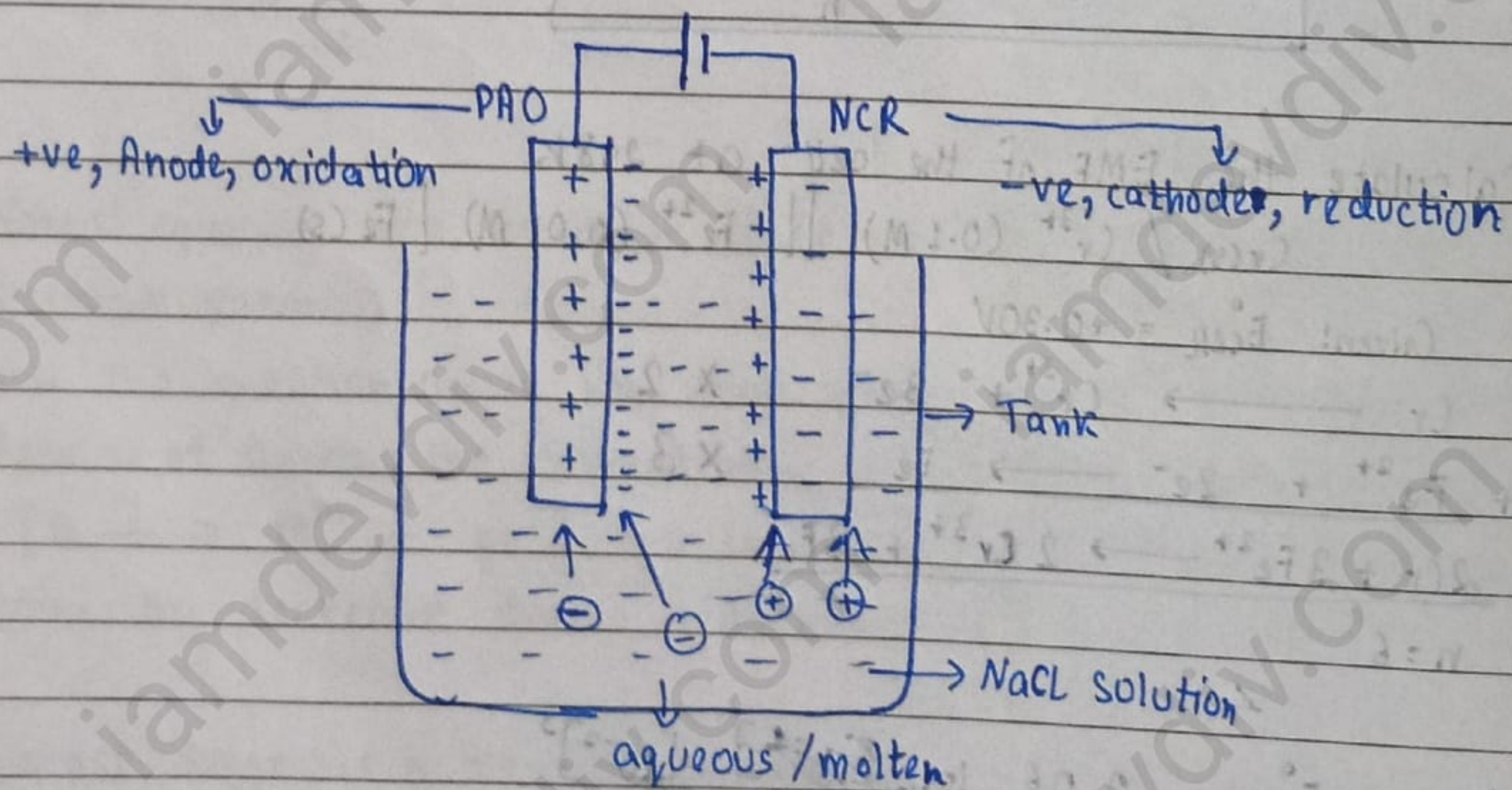


$$C_2 > C_1$$



$$C_2 > C_1$$

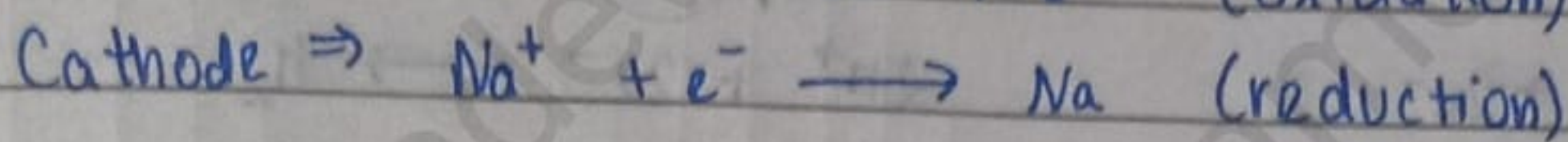
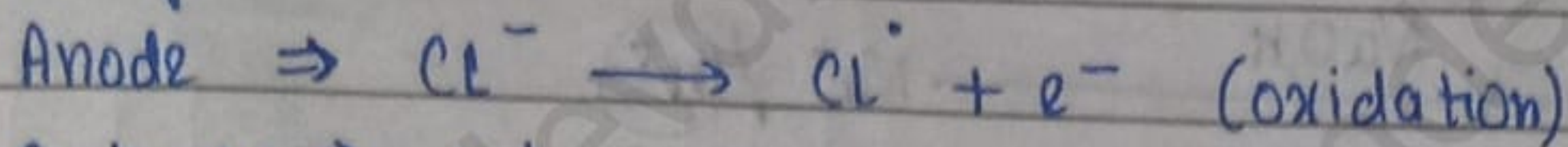
★ ELECTROLYSIS



1. The process of decomposition of an electrolyte when electric current is passed through either aqueous solution or molten state.

2. This is non-spontaneous process.

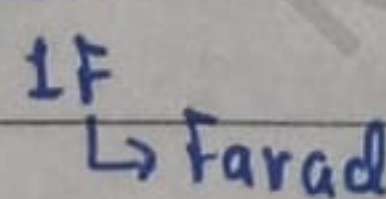
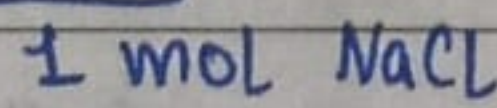
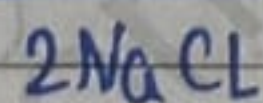
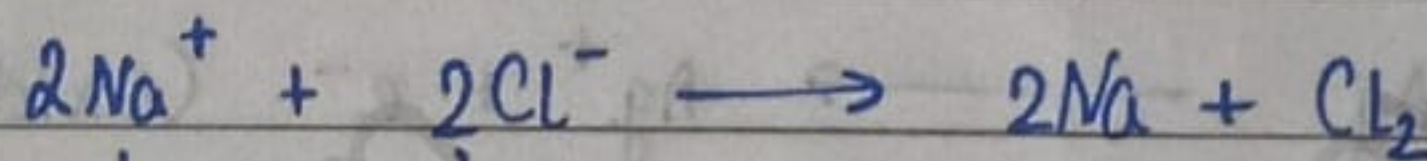
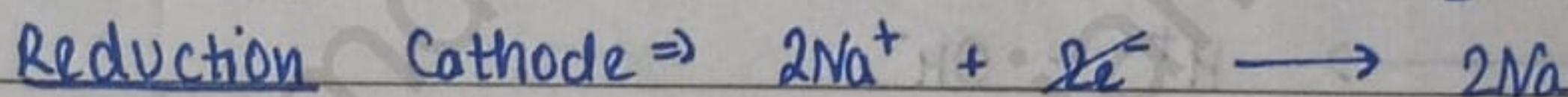
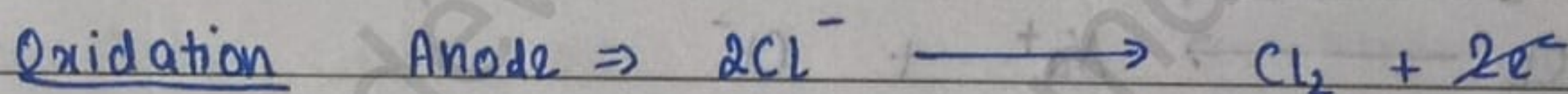
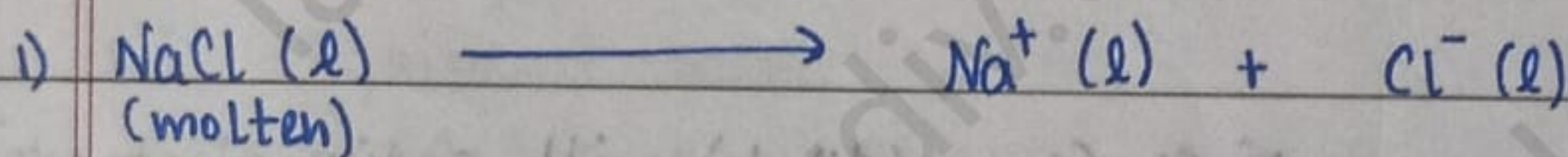
3. Discharge of ions:



• FACTORS ON WHICH PRODUCTS DEPEND

1. Nature of electrolyte
2. Medium of electrolyte (molten / aqueous)
3. Concentration of electrolyte (dilute / concentrated)
4. Nature of electrode (active / inert)

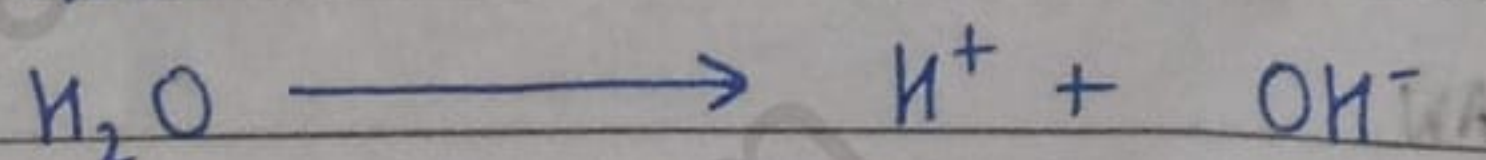
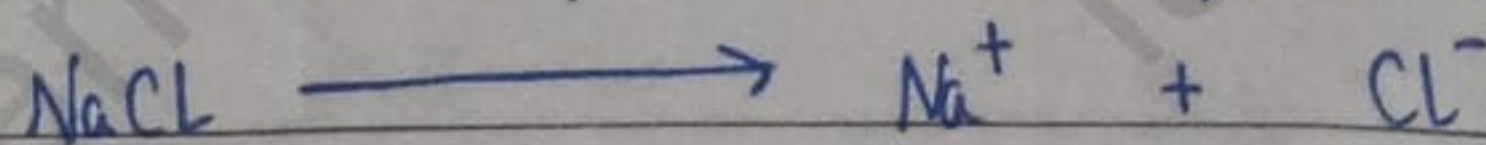
* CASE 1 [For molten electrolyte (NaCl) with platinum electrode]

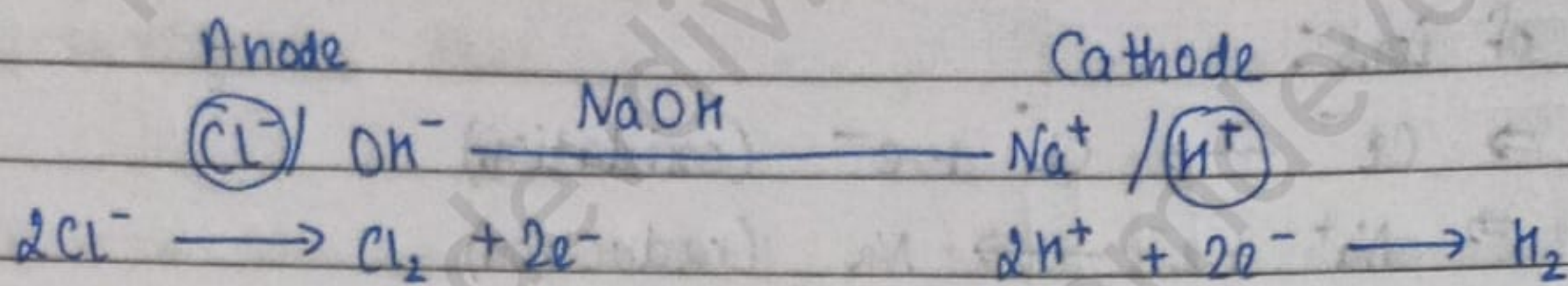


* CASE 2 [Electrolysis of aqueous solution]

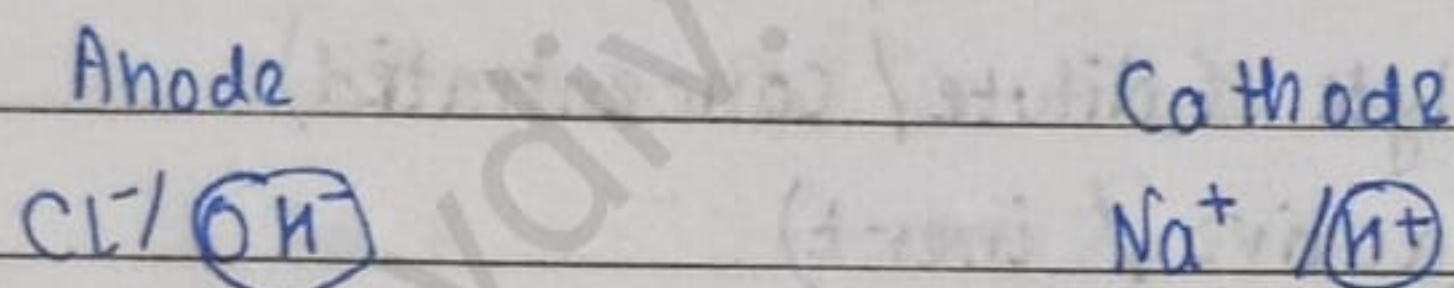
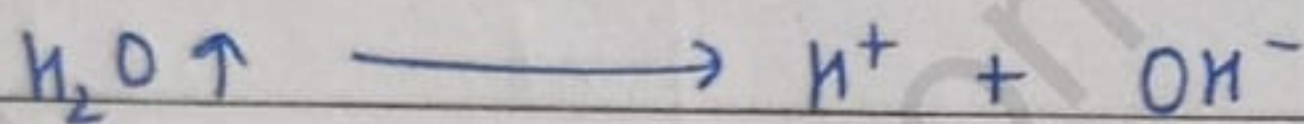
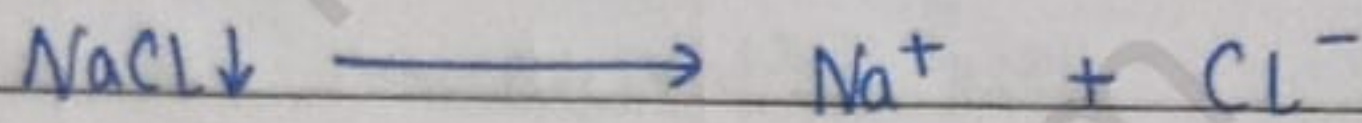
├ concentrated
└ dilute

1) Electrolysis of NaCl aq. solⁿ with platinum electrodes



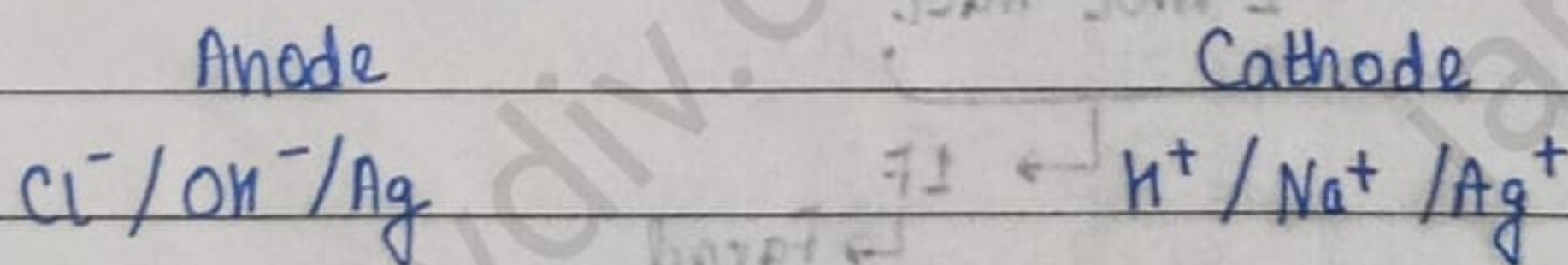
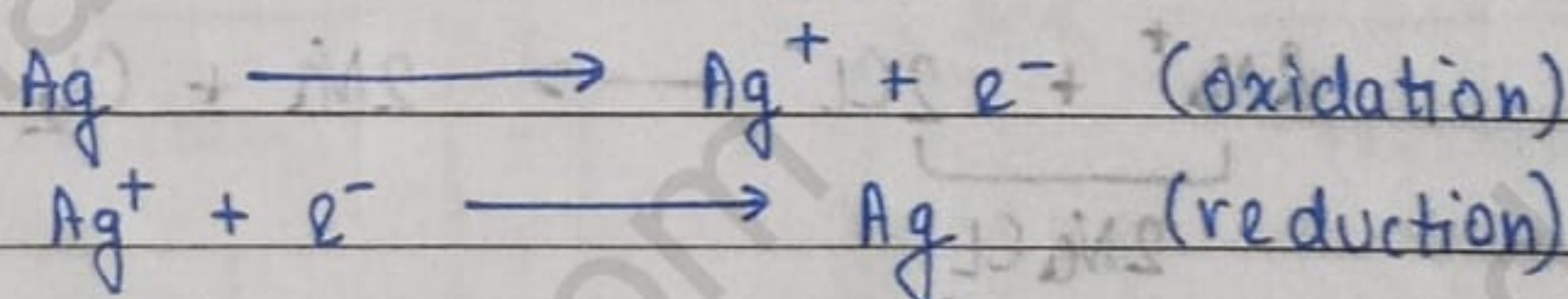
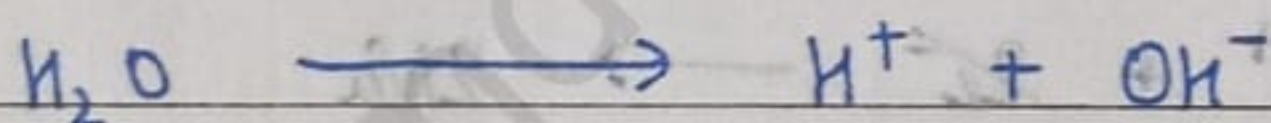
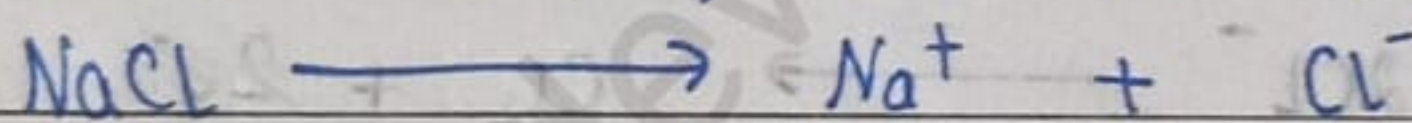


* CASE 3 [Aqueous NaCl (dilute) electrode]



Since number of H^+ and OH^- are far greater than that of NaCl, therefore H^+ and OH^- will dominate and it will take part in redox reaction.

* CASE 4 [Molten/Aqueous NaCl (conc./dilute) with active electrode]



In case of active electrode like ~~anode~~ ^{Ag, at} anode - only Ag will take part in reaction.

At cathode \longrightarrow reduction of Ag^+ will take place

Because of the nature of Ag (metal), it is acting as an active electrode.

★ FARADAY'S LAW OF ELECTROLYSIS

• FARADAY'S FIRST LAW

The mass of a substance deposited/gas evolved at particular electrode is directly proportional to the amount of charge passing through solution (electrolyte)

$$m \propto Q$$

← mass of the substance
→ Charge of battery

$$\Rightarrow m = ZQ$$

↳ Electrochemical equivalent (constant)

$$\therefore Q = it$$

$$\Rightarrow m = Zit$$

↓ Time
 Current

* CALCULATION OF Z

$$Z = \frac{\text{Equivalent weight}}{96500 \text{ C (1F)}}$$

$$\text{Eq. weight} = \frac{\text{molar mass}}{n\text{-factor}}$$

n-factor \Rightarrow no. of ions transferred

Q Find the amount of Cu deposited when 5 ampere current is used for 3600 sec (1 hour) to electrolyte solution which is CuSO_4 .

$$\text{Sol. } m = ZQ = Zit = \frac{\text{Eq. wt} \times i \times t}{F} = \frac{63.5 \times 5 \times 3600}{2 \times 96500 \times 10^2} = \frac{635 \times 3}{32}$$

$$\Rightarrow m = 5.9 \text{ g} \quad \text{Ans.}$$

• FARADAY'S SECOND LAW

If same amount of electric charge is passed through different electrolytes, then the amount of substance deposited/evolved is directly proportional to its equivalent weight.

$$M = ZQ$$

$$M_{\text{Cu}} = \frac{E_{\text{Cu}}}{F} \times Q \quad \text{--- (1)}$$

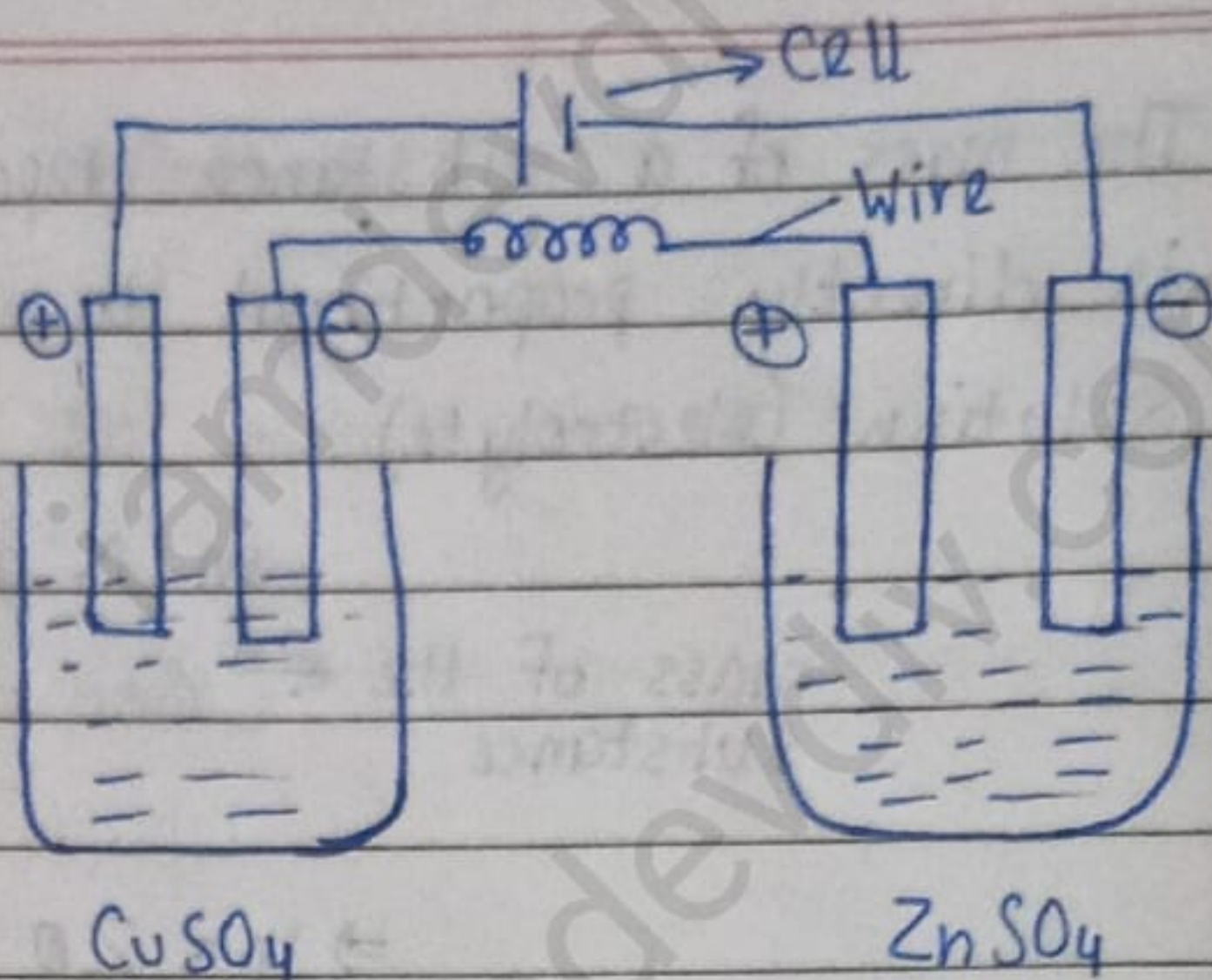
$$M_{\text{Zn}} = \frac{E_{\text{Zn}}}{F} \times Q \quad \text{--- (2)}$$

① ÷ ②

$$\frac{M_{Cu}}{M_{Zn}} = \frac{E_{Cu}/F \times Q}{E_{Zn}/F \times Q}$$

$$\Rightarrow \frac{M_{Cu}}{M_{Zn}} = \frac{E_{Cu}}{E_{Zn}}$$

$$\Rightarrow \boxed{\frac{M_1}{M_2} = \frac{E_1}{E_2}}$$



In series \Rightarrow since Q is constant

★ BATTERY [COMMERCIAL CELLS]

① The combination of several electrochemical cells connected in series or a single electrochemical cell is called a battery.

② Batteries have been classified into:

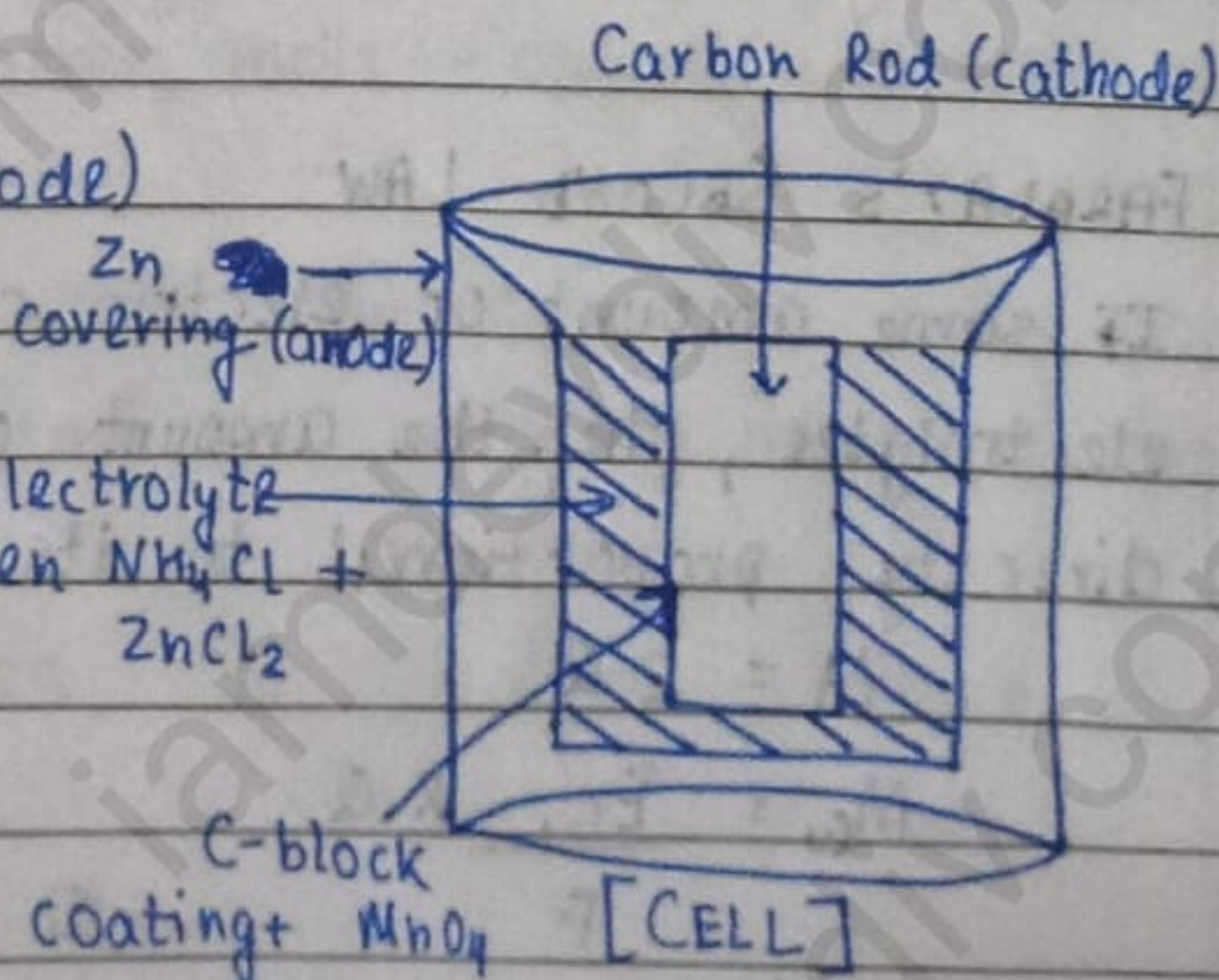
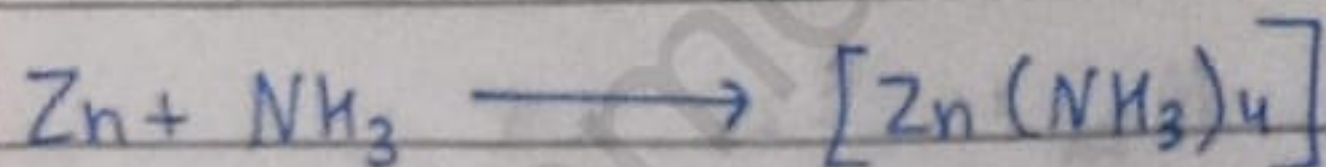
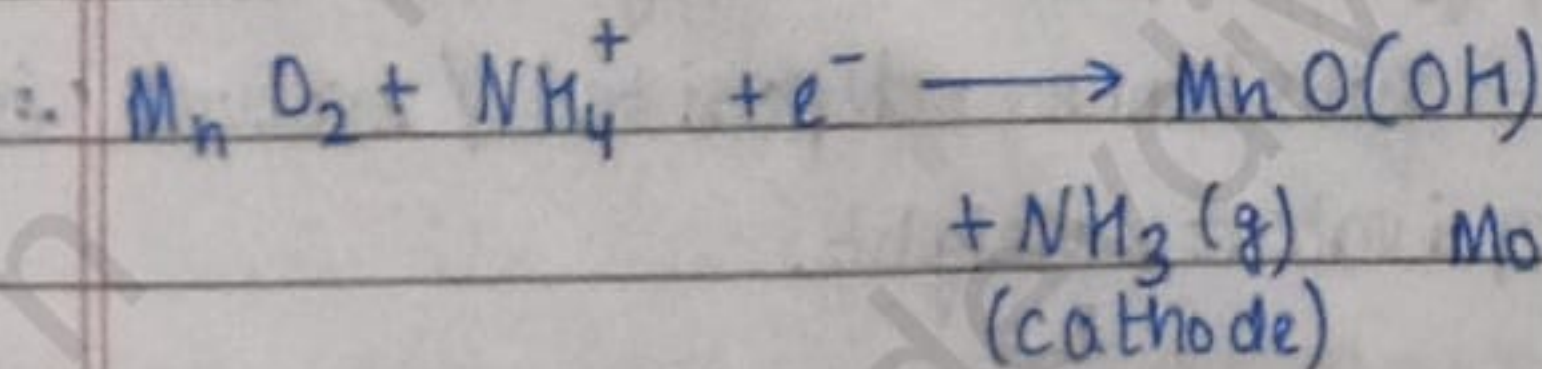
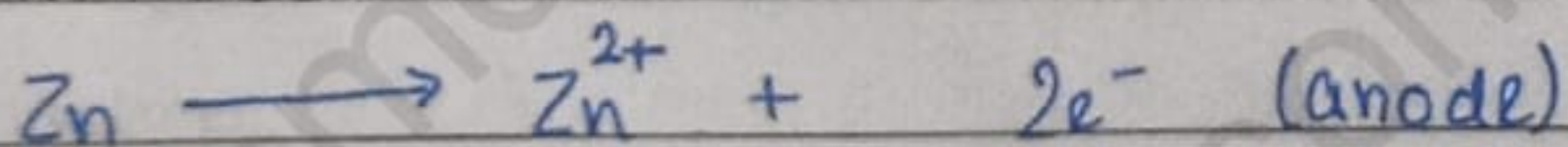
- a) Primary cells
- b) Secondary cells
- c) Fuel cells

• PRIMARY CELL

Cell which cannot be recharged and cannot be reused. There, the cell reaction is irreversible. Examples:

- a) Dry cell (Leclanche cell)
- b) Mercury cell (~~Leclanche cell~~)

* DRY CELL



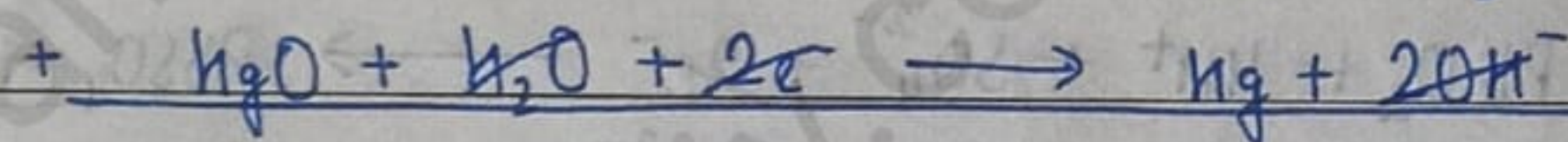
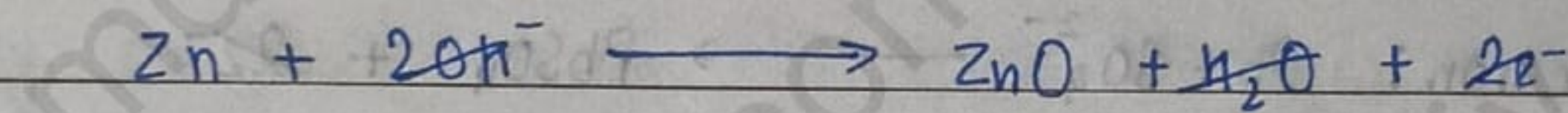
- ① This cell does not have ∞ life because NH_4Cl keeps on corroding Zn anode.
- ② It produced approximately 1.25V to 1.5V

* MERCURY CELL [BUTTON CELL]

Anode \rightarrow Zn - Amalgam [Zn - Hg]
 \rightarrow Mercury metal

Cathode \rightarrow Paste of HgO + Carbon

Electrolyte \rightarrow Paste of KOH + ZnO_2



(No ion is present in this reaction)

- ① Overall, cell reaction does not involve any ion whose concentration will change with time. So, it provides almost constant potential throughout its life.
- ② Potential of about ~~1.35~~ 1.35V is provided by this cell.

• SECONDARY CELL

These are the cells which can be recharged on passing electric current.

These are also known as storage cell because electrical energy can be stored in them (current). Examples: a) Lead storage cell

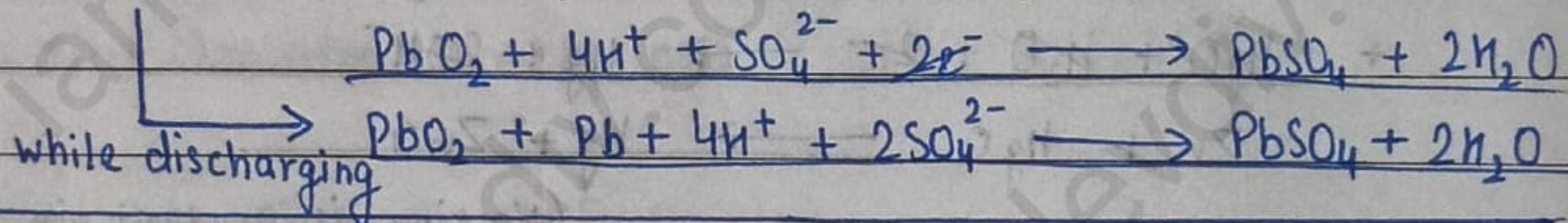
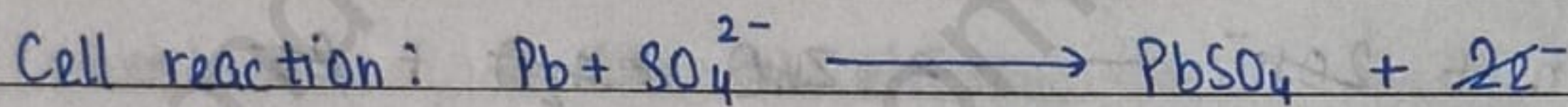
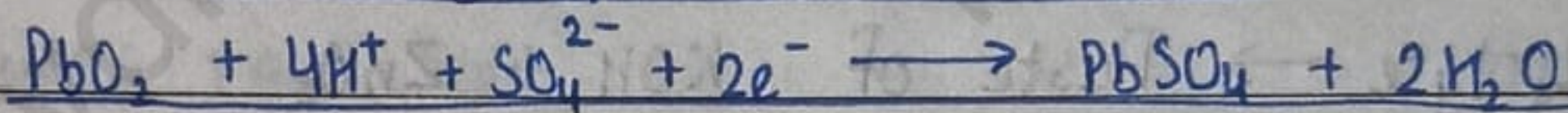
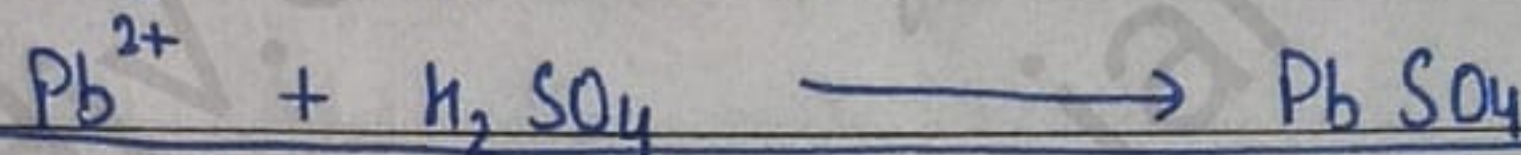
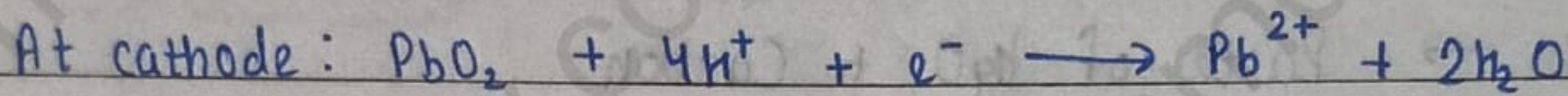
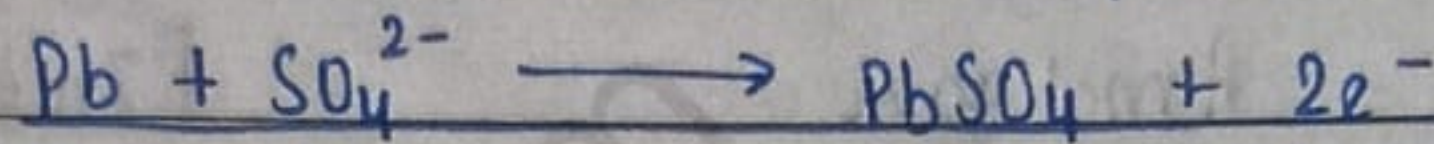
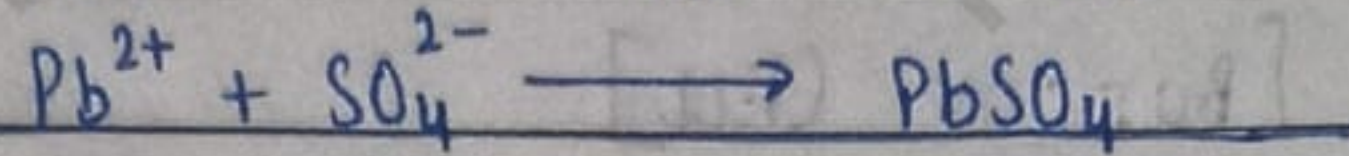
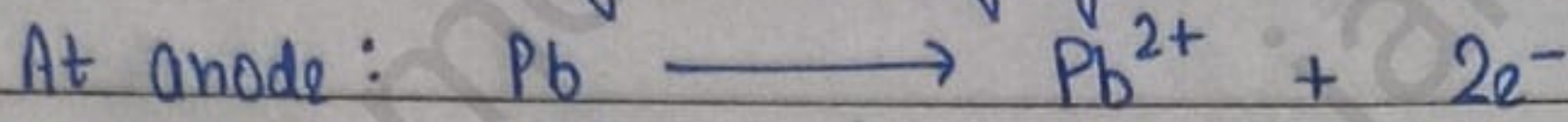
b) Nickel-cadmium storage cell

* LEAD STORAGE CELL

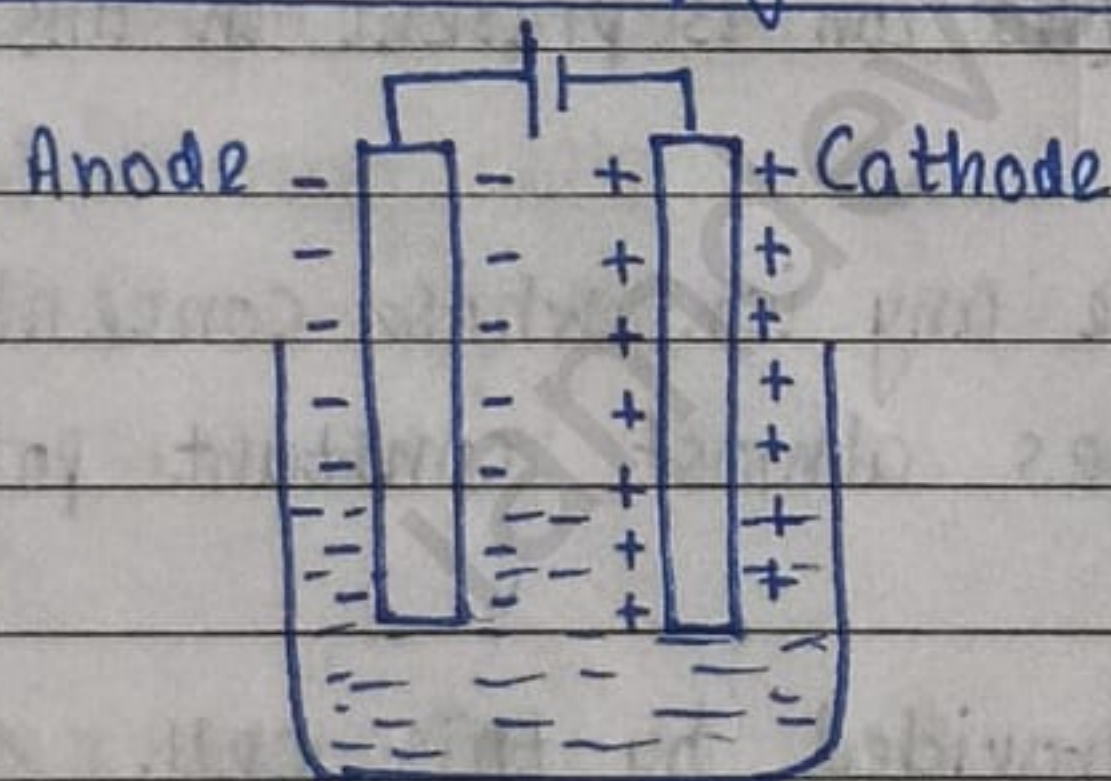
- ① It is a secondary cell.
- ② It is a combination of 6 cells of 2 volt each in series.
- ③ The total voltage produced by this cell is $6 \times 2 = 12$ volts.
- ④ In this cell, Anode \Rightarrow Pb (lead)

Cathode = PbO_2 (lead oxide)Electrolyte = H_2SO_4 (sulphuric acid)

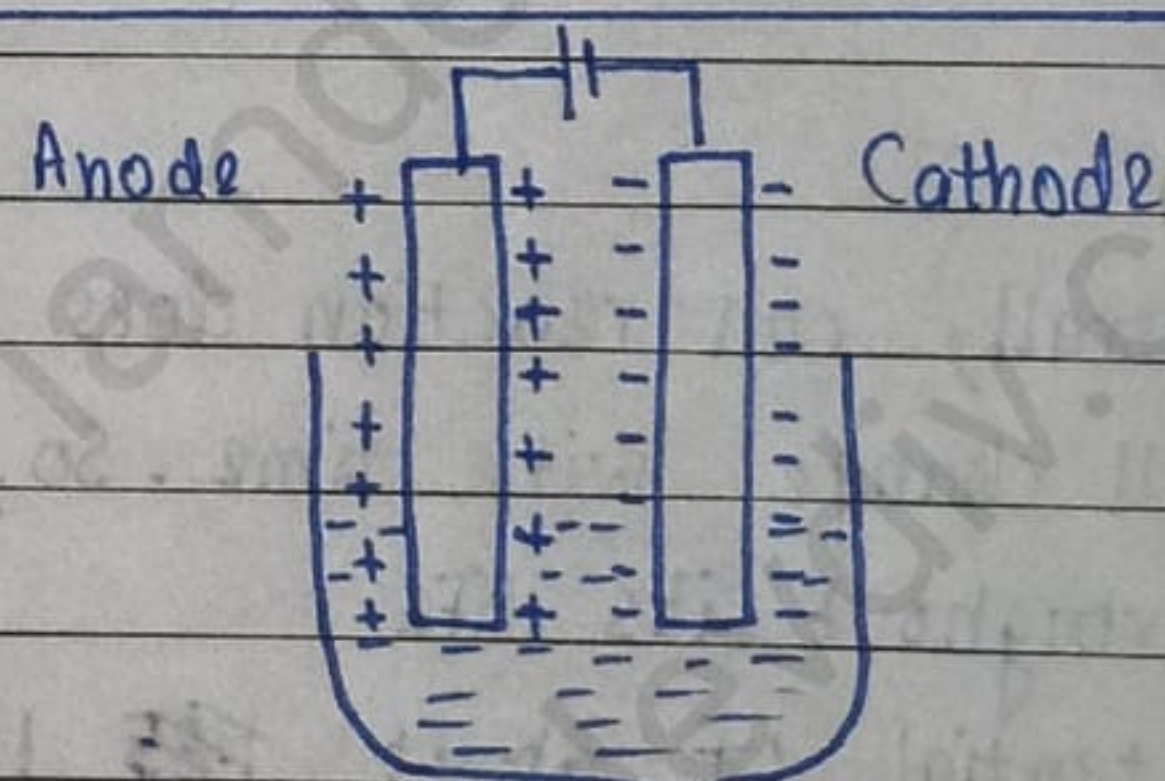
⑤ Cell reactions during discharging:



NOTE →



Galvanic Cell



Electrolytic Cell

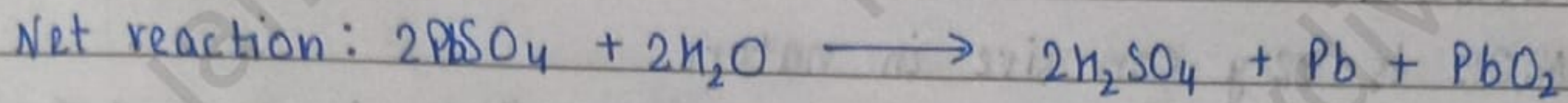
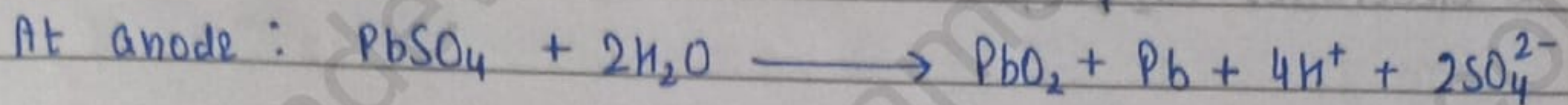
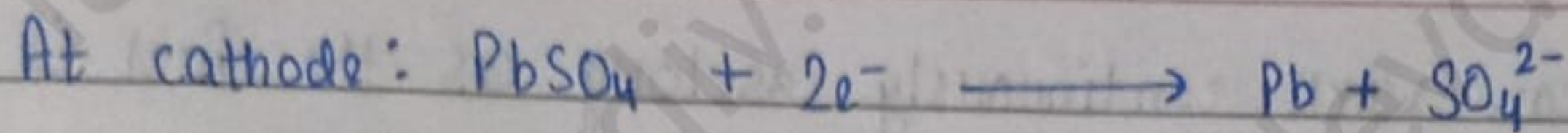
Cathode of Galvanic cell = Anode of electrolytic cell

Anode of Galvanic cell = Cathode of electrolytic cell

● During discharging,

a) H_2SO_4 gets consumedb) $PbSO_4$ gets accumulated / deposit on cathode and anode

⑥ Cell reactions during charging:

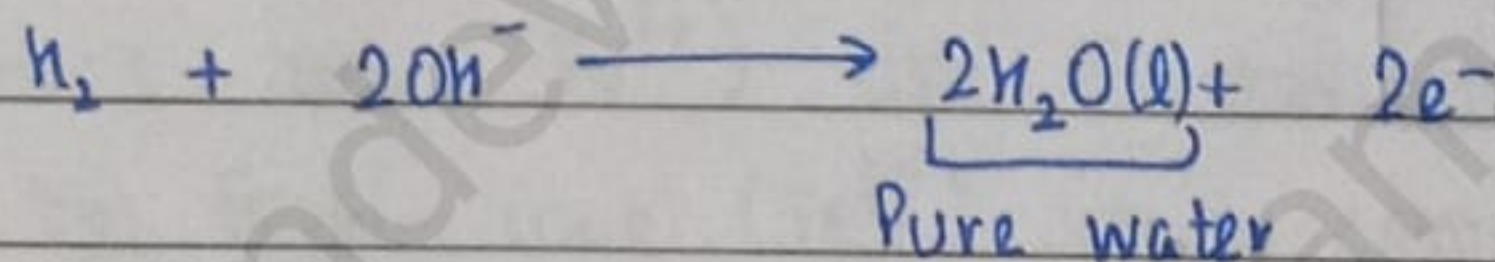


During charging, PbSO_4 dissolves to form H_2SO_4 again which is the electrolyte of this reaction.

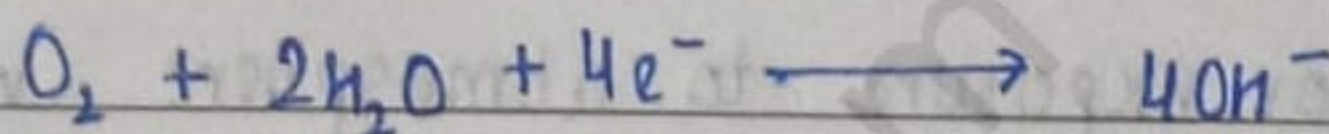
• FUEL CELL

Specially designed galvanic cell which directly converts enthalpy of combustion of fuels into electrical energy.

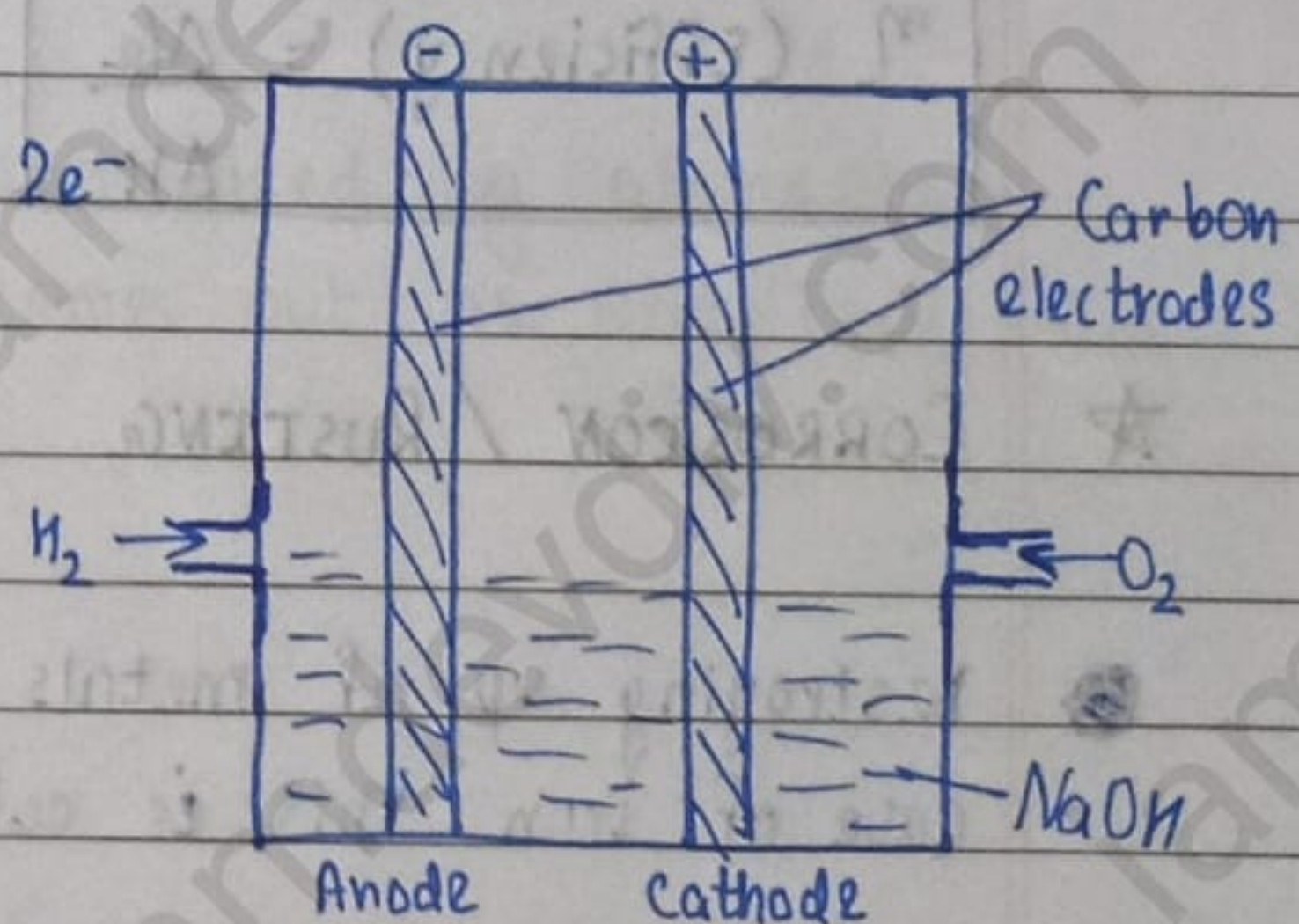
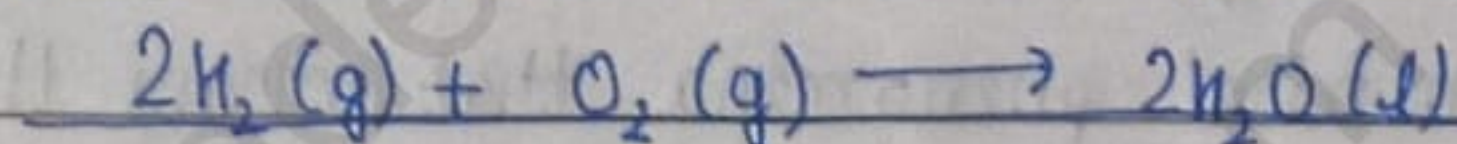
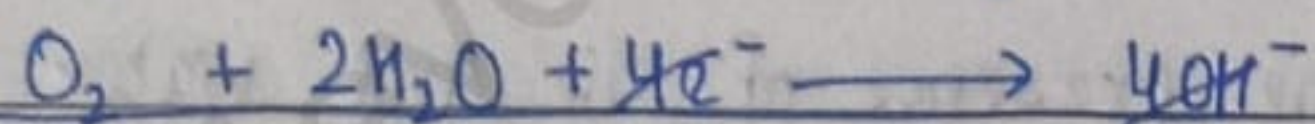
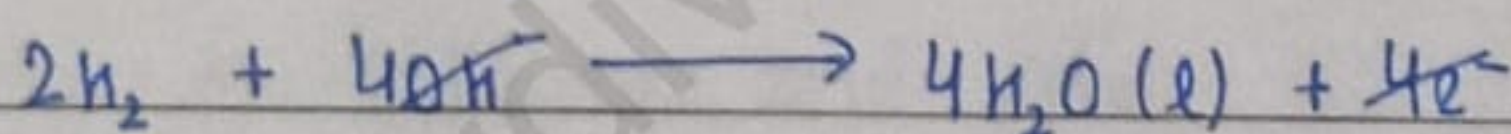
At ~~anode~~ anode:



At cathode:



Net reaction:



[FUEL CELL]

* H₂-O₂ FUEL CELL

- ① It was firstly used in APOLLO SPACE MISSION OF NASA.
- ② The water formed in this reaction is used for drinking purpose of Astronauts.
- ③ This cell is highly efficient approx 70%.
- ④ It is eco-friendly and pollution free.

- ⑤ A fuel cell can work continuously as long as the fuel is supplied to electrodes (H_2-O_2).

* LIMITATIONS

- ① NaOH is very corrosive in nature.
- ② It is very difficult to store gaseous fuels i.e. H_2 and O_2 . These are kept in liquid form in specially designed cylinders.

* EFFICIENCY OF FUEL CELL

- The theoretical efficiency of a fuel cell can be calculated in terms of ΔH and ΔG .

ΔH = heat of combustion

ΔG = Electrical cell energy.

$$\eta \text{ (Efficiency)} = \frac{\Delta G}{\Delta H}$$

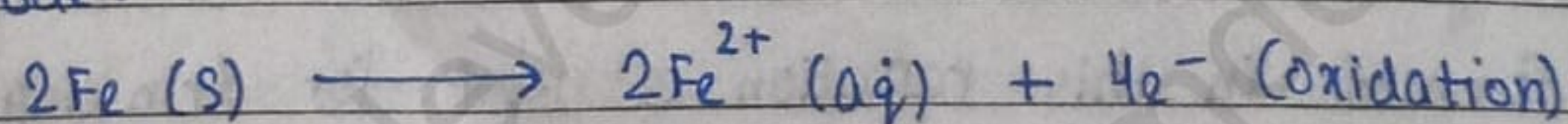
★ CORROSION / RUSTING

- Destroying up of metals or exposure to atmosphere (water and air) in case of iron. This is called rusting / corrosion.

• RUSTING OF IRON

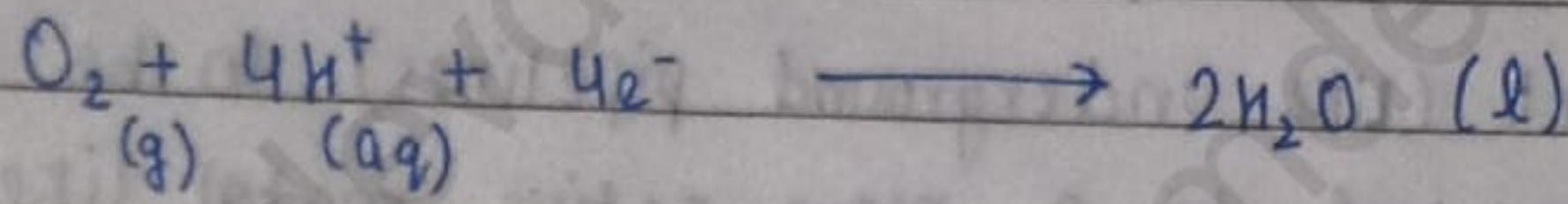
Rusting of iron occurs in the presence of air and water. It is considered as an electrochemical phenomena that means there is a chance of formation of a cell (anode, cathode, electrolyte)

At anode:

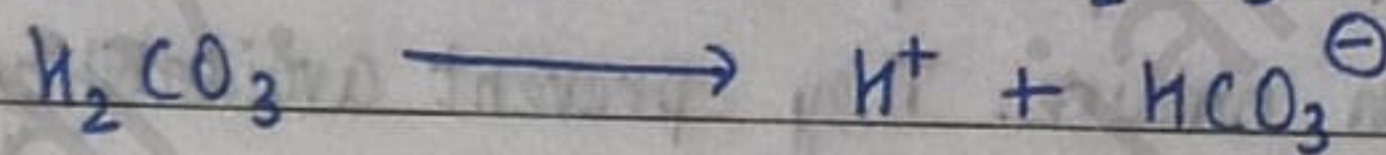
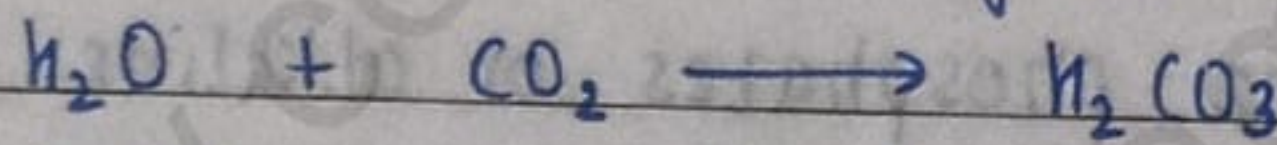


Oxidation takes place at a particular spot of iron which acts as anode.

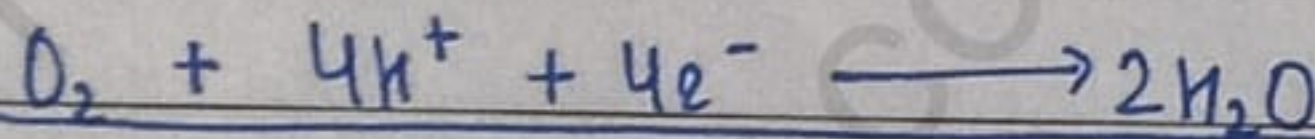
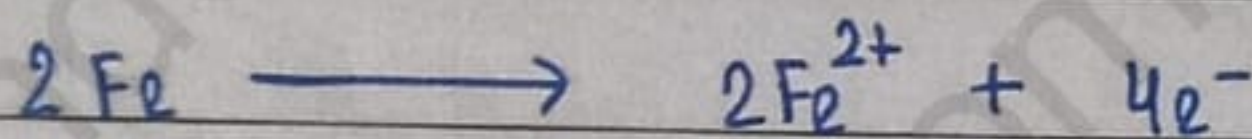
At cathode:



e^- released at the anodic spot move through the metal and go to another spot on a metal ^{and} reduced oxygen in the presence of H^+ . ~~which~~ ^{is} Acidic oxide is formed by the reaction of H_2O .



Net reaction:



The ferrous ions (Fe^{2+}) are further oxidised by atmospheric oxygen by ferric ions (Fe^{3+}) which comes out as rust. ($\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$)

• PREVENTION OF CORROSION / RUSTING

1. BARRIER PROTECTION

- In the Barrier protection the iron surface is not allowed to come in contact with moisture and air.
- This can be achieved either by coating iron surface or by paint or by applying any oil or by coating with non-corroding metals.
Example: Nickel (Ni) and Chromium (Cr)

2. SACRIFICIAL PROTECTION

- It means covering of iron surface with a layer of metal which is more reactive than iron and thus, prevents the oxidation of iron.
- Zinc is oftenly used for this purpose and the process is known as galvanisation.

3. ELECTRICAL PROTECTION / CATHODIC PROTECTION

The iron object (like underground pipelines) to be protected from rusting is connected to a more active metal like Mg. Thus, the iron object here will act as cathode and the protecting metal (Mg) will act as anode.

4. USING ANTI-RUST SOLUTION

These are alkaline phosphates or alkaline chromates solution being alkali in nature. They prevent availability of H^+ ions and stops the cathodic reaction. Thus, the rusting is prevented.