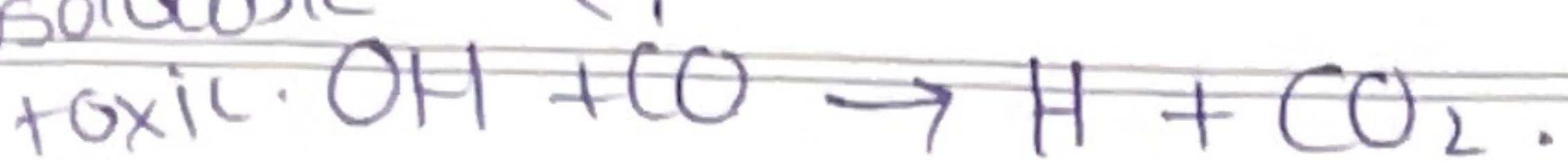


The most imp react  
in the atmosphere

insoluble  $\leftarrow$



$\text{OH}^-$  (Hydroxide ion).

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## CHEMICAL SPECIE

$\hookrightarrow$  soluble used.  
in intoxication in photo-synthesis

FREE Radicals:- The type of chemical species that possesses at least one (pair) of unpaired electrons (in the valence shell).

Example:-  $\text{H}_\cdot, \text{O}_\cdot, \text{OH}_\cdot, \text{OOH}, \text{NO}_\cdot, \text{NO}_2, \text{NO}_3$ .  
(hydroxyl free Radical)

\* Atomic mass & molecular masses are expressed in kg and amu.

Mole:- A standard scientific unit for measuring very large quantities of very small entities like atoms, molecules etc.

" SI unit for expressing the amount of any substance".

$$1 \text{ mole} = 6.022 \times 10^{23} \text{ units}$$

(Avogadro's no.).

## CONCENTRATION UNITS:-

$$1 \text{ L} = 1 \text{ dm}^3$$

i) Molarity :-  $\frac{\text{no. of moles of solute}}{1 \text{ dm}^3 / 1 \text{ L of solution}}$ .

$$1 = 1000 \text{ mL}$$

ii) Molality :-  $\frac{\text{no. of moles of solute}}{1 \text{ kg of solution}}$ .

Example ; Volume percent ; Mass percent.  
If a 100L sol. contains 30%.

$\text{H}_2\text{SO}_4$  i.e:-

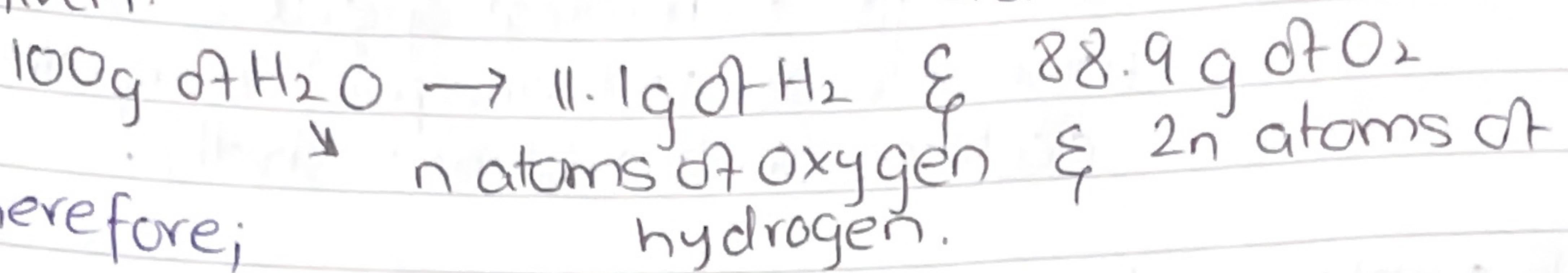
$$30\% \text{ H}_2\text{SO}_4 + 70\% \text{ H}_2\text{O}$$

If 100 kg of sol have  
30% of  $\text{C}_2\text{H}_2\text{O}_1$ .

50kg of sugar + 50kg of  $\text{H}_2\text{O}$ .

The Atomic Mass Scale (first generation)  
 that used O<sub>2</sub> as a standard.

Given:-



Therefore;

$$\text{mass of one hydrogen} = \frac{11.1}{2n}$$

$$\text{mass of one oxygen} = \frac{88.9}{n}$$

$$\text{Mass ratio of O & H} \left( \frac{88.9}{n} \right) : \left( \frac{11.1}{2n} \right) = 16.018$$

Oxygen is 16.018 times heavier than H-atom.

The Atomic mass scale (Mass spectrometer).

- (i) Separation of ions based upon the mass difference
- + Vacuum pump
- + Magnet
- Electric field.

"Roadway method of Calculations"

GRAMS → MOLES → MOLECULES → AMU

How many amu(s) in 1g of C<sup>12</sup> atom?

$$\begin{aligned} \text{amu?} &= 1 \text{ g of Carbon atom} \times \frac{1 \text{ mol of C}^{12} \text{ atom}}{12 \text{ g of C}^{12} \text{ atom}} \\ &\quad \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol of C}^{12} \text{ atom}} \times \frac{12 \text{ amu}}{1 \text{ C}^{12} \text{ atom}} \end{aligned}$$

$$\Rightarrow 6.022 \times 10^{23} \text{ amu.}$$



2 moles 1 mole. 2 moles.

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Q How many moles of atoms does 136.9 g of iron metal contain?

### MASS → MOLES

$$\begin{aligned} ? \text{ mols of Fe atom} &= 136.9 \text{ g of Fe} \times \frac{1 \text{ mol of Fe}}{55.85 \text{ g of Fe}} \\ &= 2.451 \text{ mol Fe atoms.} \end{aligned}$$

Q. How many atoms are contained in 2.451 mols of iron.

### MOLES → PARTICLES

$$\begin{aligned} ? \text{ no. of atoms} &= 2.451 \text{ moles of Fe} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol of Fe}} \\ &= 1.476 \times 10^{24} \text{ Fe atoms.} \end{aligned}$$

Q. Suppose 0.05 g of water evaporate in one hour and then how many molecules evaporate each second?

### GRAMS → MOLES → MOLECULES

$$\begin{aligned} (\text{evapour}) \text{ H}_2\text{O molecules} &\Rightarrow 0.05 \text{ g H}_2\text{O} \times \frac{1 \text{ mol of H}_2\text{O}}{18 \text{ g of H}_2\text{O}} \times \\ &\quad \text{s} \end{aligned}$$

$$\begin{aligned} &6.022 \times 10^{23} \text{ molecules} \times \frac{1 \text{ hr}}{1 \text{ mol of H}_2\text{O}} \times \frac{1}{60 \text{ min}} \\ &\quad \text{60 min} \end{aligned}$$

$$\begin{aligned} &\Rightarrow 5 \cdot 10^{17} \text{ H}_2\text{O} \\ &\quad 60 \text{ s.} \end{aligned}$$

molecules/s.

Q. Calculate the no. of hydrogen atoms in 39.6 g of ammonium sulphate.  $(\text{NH}_4)_2\text{SO}_4$ .

### GRAM → MOLES → PARTICLES → No. OF H-atoms.

$$\begin{aligned} \text{no. of H atoms} &\Rightarrow 39.6 \text{ g of } (\text{NH}_4)_2\text{SO}_4 \times \frac{1 \text{ mol of }}{} \\ &\quad \times \frac{6.022 \times 10^{23} \text{ formula units}}{1 \text{ mole}} \times \frac{8 \text{ H-atoms}}{1 \text{ formula unit}} \end{aligned}$$

1 mol of any element = 24.414 L

1 m<sup>3</sup> = 1000L

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1 atm, 0°C, 273K.

Q. Assuming the dimension of class are 20m x 20m x 6m.  
Calculate the no. of molecules in CO<sub>2</sub> in this room if the concentration of CO<sub>2</sub> is 0.04%.

$$\text{no. of molecules} \rightarrow \frac{?}{\text{of CO}_2} \left( \frac{2400 \text{ L}}{\text{m}^3 \text{ of air}} \right) \left( \frac{0.04 \text{ L of CO}_2}{100 \text{ L of air}} \right)$$
$$\left( \frac{1 \text{ mole of CO}_2}{24.414 \text{ L}} \right) \left( \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mole}} \right)$$

## "COMBUSTION ANALYSIS"

Question: A sample space of 0.14 g of purified glucose is burned in C-H combustion to produce 0.1486 g of CO<sub>2</sub> and 0.0609 g of H<sub>2</sub>O.

$$? \text{ g of C} = 0.1486 \text{ g of CO}_2 \times \frac{12 \text{ g of C}}{44.01 \text{ g of CO}_2} = 0.04 \text{ g}$$

$$\text{given mass} \times \frac{\text{mass of C-12 atom}}{\text{molar mass of CO}_2}$$

$$? \text{ g of H} = 0.0609 \text{ g of H}_2\text{O}$$

$$\frac{2 \text{ g of H}}{18 \text{ g of H}_2\text{O}} = 0.0666 \text{ g}$$

$$\text{given mass} \times \frac{\text{mass of H-atom}}{\text{molar mass of H}_2\text{O}}$$

0.1486-

$$? \text{ g } \delta \text{f O} \Rightarrow \uparrow(0.04 + 0.006) \Rightarrow 0.0540 \text{ g } \delta \text{f O.}$$

to find the relative no. of atoms.

$$C = \frac{0.04}{12} = 0.0033$$

Divide

$$H = \frac{0.006}{1} = 0.006 \text{ to lowest } 2 \text{ of H}$$

$$O = \frac{0.005}{16} = 0.0033 \text{ of O}$$

Therefore the empirical formula is  $\text{CH}_2\text{O}$ .

$n$  (Empirical formula) = Molecular formula



$n = \text{molecular mass/empirical formula mass}$

$$n = 180/30 = 6$$

$$6(\text{CH}_2\text{O}) \Rightarrow (\text{C}_6\text{H}_{12}\text{O}_6 \text{ Ans!} \& \text{ molecular formula.})$$

Excess of phosphorus in water bodies (eutrophication).  
(i) Algae blooms      (iii) water q. degradation.  
(ii) oxygen depletion      (iv) loss of aquatic life.

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Determining the atomic mass of an unknown element from the empirical formula.

Empirical formula of a given metal oxide is  $\text{MO}$ . Mass of metal taken  $0.490\text{g}$ . Burn the unknown metal in oxygen. Metal oxide formed is  $= 0.813\text{g}$ .

as the mass of the metal & metal oxide are given ; the mass of Oxygen can be determined.  
mass of oxygen.

$$\uparrow \text{MO} = 0.813\text{g} - 0.490\text{g} = 0.323\text{g of O.}$$

$$\text{Now moles of oxygen} = \frac{0.323\text{g}}{(18)\text{g}} \times \frac{1 \text{ mole}}{1 \text{ mole}}$$

$$\text{Oxygen} = 0.0202 \text{ moles.}$$

Now.

$$\text{no. of moles of oxygen} = 0.0202 \text{ moles} = \text{no. of moles of metal.}$$

$$\text{hence } \frac{\text{mass in gram}}{\text{molar mass}} = \text{no. of moles.}$$

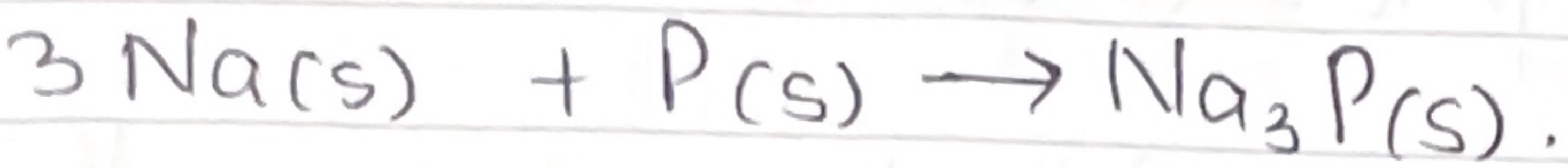
$$\Rightarrow \text{mass in g} = " \times \text{molar mass}$$

$$\Rightarrow \text{molar mass} = \frac{\text{mass in g}}{\text{no. of moles}} = \frac{0.490}{0.0202} = 24.3\text{g.}$$

C, O, H, N, S, P are responsible for the growth of plant in water bodies and Phosphorus is the limiting reagent & group <sup>Date: 20</sup> ~~WTFOP~~.

## CALCULATIONS BASED ON CHEMICAL EQUATIONS:-

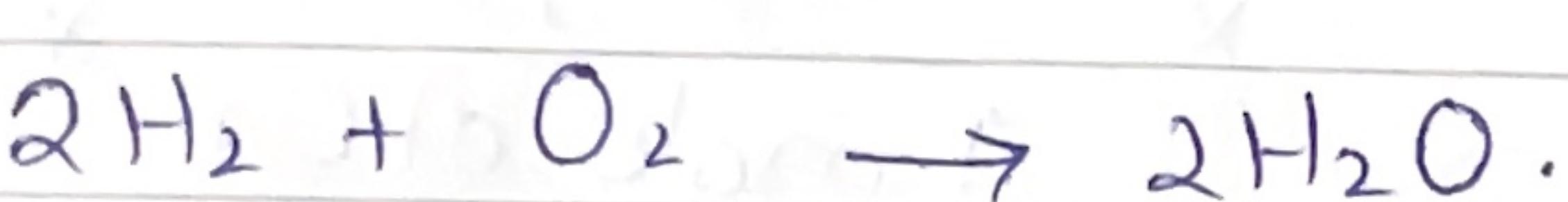
Phosphorus reacts with sodium to produce Na-  
Phosphorus / sodium phosphide.



How many grams of  $\text{Na}_3\text{P(s)}$  can be made from 10g of Sodium.

$$\begin{aligned} \text{mass of } \text{Na}_3\text{P} &= 10 \text{g of Na} \times \frac{1 \text{mole}}{22.9 \text{g of Na}} \times \frac{1 \text{mole of } \text{Na}_3\text{P}}{3 \text{mol of Na}} \\ &\quad \times \left( \frac{99.94 \text{g of } \text{Na}_3\text{P}}{1 \text{mole of } \text{Na}_3\text{P}} \right) \\ &= 14.5 \text{g } \text{Na}_3\text{P}. \end{aligned}$$

## CALCULATING LIMITING REAGENTS:-



→ always produces lesser product

7 moles of  $\text{O}_2$  } Given.  
10 moles of  $\text{H}_2$  }  
↓

$$\begin{aligned} \text{Moles of } \text{H}_2\text{O} &\Rightarrow (7 \text{ moles of } \text{O}_2) \left( \frac{2 \text{ moles of } \text{H}_2\text{O}}{1 \text{ mole of } \text{O}_2} \right) \\ &\Rightarrow 14 \end{aligned}$$

$$\begin{aligned} \text{Moles of } \text{H}_2\text{O} &\Rightarrow (10 \text{ moles of } \text{H}_2) \left( \frac{\text{H}_2\text{O}}{\frac{2 \text{ moles of } \text{H}_2}{2 \text{ moles of H}}} \right) \\ &\Rightarrow 10 \end{aligned}$$

## GASOLINE NUMERICALS:-

A vehicle consumes 14kg of gasoline C<sub>8</sub>H<sub>18</sub> in a round trip to Islamabad.

Assumption:-  $2\text{C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O}$ .

- (i) Calculate; - the kg of O<sub>2</sub> required
- (ii) kg of CO<sub>2</sub> produced (and H<sub>2</sub>O).
- (iii) kg of N<sub>2</sub> involved in this round trip.

$$\text{moles of C}_8\text{H}_{18} = 14 \text{ kg of C}_8\text{H}_{18} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mole}}{114 \text{ g of C}_8\text{H}_{18}}$$

$$= 123 \text{ moles of C}_8\text{H}_{18}$$

$$\text{O}_2 \text{ produced} \Rightarrow 123 \text{ moles of C}_8\text{H}_{18} \times \frac{25 \text{ mols of O}_2}{2 \text{ mol of C}_8\text{H}_{18}}$$

$$\times \frac{32 \text{ g O}_2}{1 \text{ mole of O}_2} \times \frac{1 \text{ kg}}{1000 \text{ g}}$$

$$= 98.4 \text{ kg, ?}$$

$$\text{CO}_2 \text{ produced} \Rightarrow 123 \text{ moles of C}_8\text{H}_{18} \times \frac{16 \text{ mols of CO}_2}{1 \text{ mol of C}_8\text{H}_{18}}$$

$$\times \frac{44 \text{ g of CO}_2}{1 \text{ mol of CO}_2} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 86.5 \text{ kg.}$$

$$\text{H}_2\text{O} \Rightarrow 123 \text{ moles of C}_8\text{H}_{18} \times \frac{18 \text{ mols}}{1 \text{ mol of C}_8\text{H}_{18}} \times \frac{18 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ kg}}{1000 \text{ g}}$$

$$\Rightarrow 39.8 \text{ kg.}$$

## Quiz Questions:-

(i) Find volume of air if 100kg of CO<sub>2</sub> and 0.04% v of CO<sub>2</sub> is present.

$$\text{Volume of air} \Rightarrow 100\text{kg of CO}_2 \times \frac{1 \text{ mole}}{44\text{g of CO}_2} \times \frac{1000\text{g}}{1\text{kg}} \times \frac{24.4 \text{ L}}{1\text{mol}}$$

$$\times \frac{100\text{L of air}}{0.04\text{L of CO}_2} \times \frac{1000\text{ml}}{1\text{L}} \Rightarrow \text{Ans!}$$

(ii) M<sub>2</sub>O<sub>3</sub> is 1.6 grams. and mass of M is 1.4 g (metal). Find the metal?

$$\text{grams of oxygen} \Rightarrow 1.6 - 1.4 = 0.2 \text{ grams!}$$

$$\text{Moles of M} \Rightarrow (\frac{0.2}{\text{g of O}}) \left( \frac{1 \text{ mol of O}}{16 \text{ g of O}} \right) \left( \frac{2 \text{ moles of M}}{3 \text{ moles of O}} \right)$$

\_\_\_\_ Ans!

How much KMnO<sub>4</sub> from 250ml of 0.01M solution

grams in KMnO<sub>4</sub>?  $\uparrow$

KMnO<sub>4</sub>

$$\text{g of KMnO}_4 = 250\text{mL} \times \frac{0.01 \text{ moles of KMnO}_4}{1\text{L SO}} \times \frac{158\text{g}}{1\text{mol}}$$

KMnO<sub>4</sub>

$$\times \frac{1000\text{mL}}{1\text{L}} \Rightarrow 0.395 \text{ g are present in this solution!}$$

37

Q. A bottle contains 46 mass% HCl solution in water and has a density of 120 g/ml. Calculate volume in ml from this bottle for preparing 500ml sol. of 0.2M.

$$\text{Vol of HCl} = 500 \text{ml} \times \frac{0.2 \text{ moles of HCl}}{1 \text{ L}} \times \frac{36 \text{ g of HCl}}{1 \text{ mole of HCl}} \\ \times \frac{100 \text{ g}}{37} \times \left( \frac{1 \text{ ml of sol}}{1.2 \text{ g of sol}} \right) \Rightarrow 8.2 \text{ ml (Ans)}$$

Q. Calculate conc. as molarity & ppm of O<sub>2</sub> in the solution if O<sub>2</sub> dissolves in sea water to the extent of 8.3mg/L at 25°C.

(8.3 mg)  
1L

density of water  $\Rightarrow$   
1.03 g/ml

molarity  $\Rightarrow$   $\frac{\text{no. of moles of solute}}{1 \text{ L}}$

$$\text{ppm} = \frac{8.3}{1000 \text{ mL}} \times 10^6$$

Na<sub>2</sub>CO<sub>3</sub>

# STANDARDIZATION OF SOLUTION

Standard solution are used because

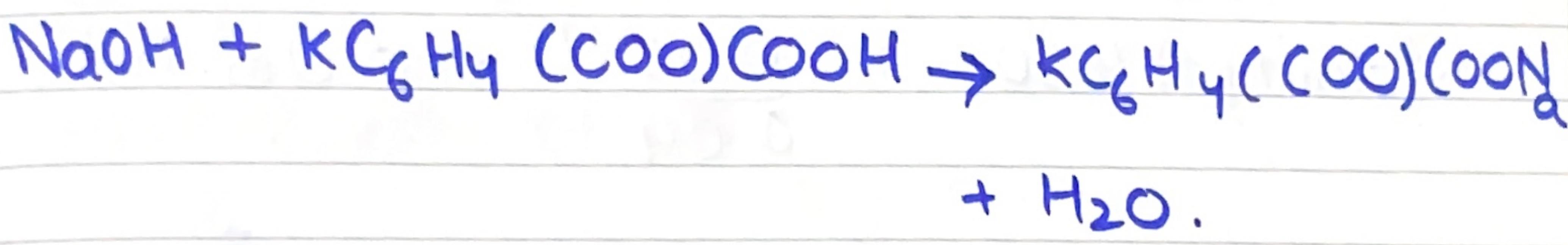
- (i) Solution must not react with the components of atmosphere.  
( $O_2$ ,  $CO_2$ , & water vapours)
  - (ii) Must react accordingly to one well defined reaction.
  - (iii) Must be solvable in the desired solvent.
  - (iv) Preferably available in solid state.
  - (v) Should be environmental friendly.

(ALKALINE SOLUTION BY)

# STANDARDIZATION OF 1 KHP.

(Potassium hydrogen phosphate).

calculate the molality of NaOH solution if 20.00 ml of the solution reacts with 0.3644 g of dry accurately weighed KHP.



$$\text{moles of NaOH} \Rightarrow 0.3644 \times \frac{1\text{mole}}{204.2\text{g}} \Rightarrow 0.001783 \text{ moles}$$

as 1:1 of both reactants

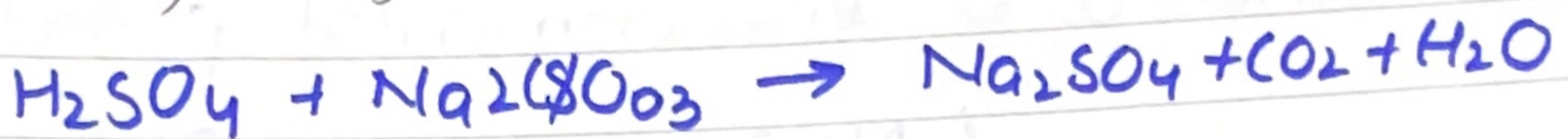
i - 0.001783 mole  $\Rightarrow$  0.001783 moles  
of NaOH A kHP.

$$\text{Molarity} \Rightarrow 0.001783 \text{ mole} \times \frac{1}{0.0200} \\ \Rightarrow 0.08915 \text{ M. of NaOH.}$$

## STANDARDIZATION OF ACID SOLUTIONS

By Na<sub>2</sub>CO<sub>3</sub>.

Calculate the molarity of H<sub>2</sub>SO<sub>4</sub> if 40.0 ml of this solution neutralizes 0.564 g of dry & accurately weighed Na<sub>2</sub>CO<sub>3</sub>.



$$\text{moles of } (\text{Na}_2\text{CO}_3) \Rightarrow \frac{0.364}{106 \text{ g}} = 0.0034 \text{ mol.}$$

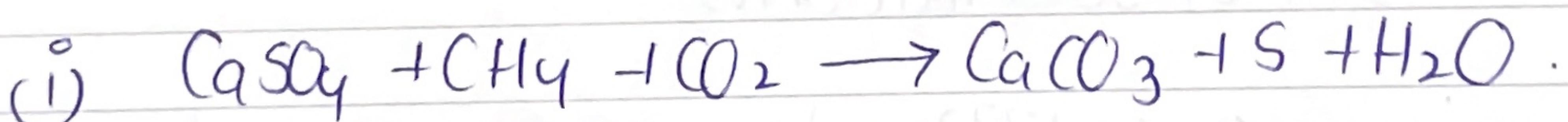
as 1:1 ratio in equation.

$$\text{moles of H}_2\text{SO}_4 = \text{moles of Na}_2\text{CO}_3.$$

$$\text{Molarity of H}_2\text{SO}_4 \Rightarrow \frac{0.0034}{0.04} = 0.085 \text{ M.}$$

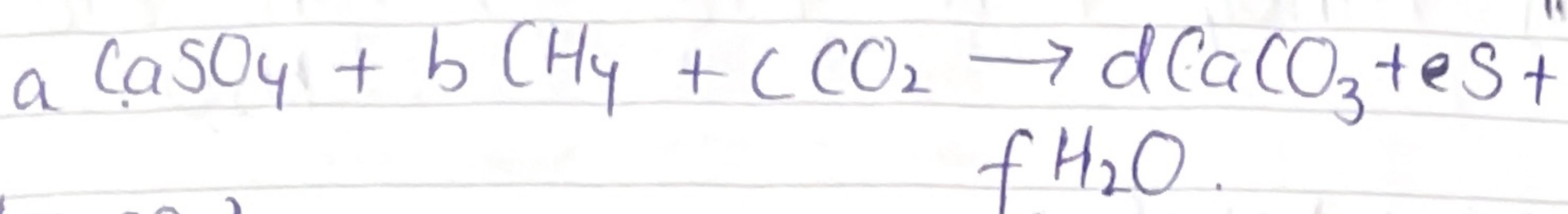
## "BALANCING CHEMICAL EQUATIONS"

Example:-



Step 1.

Let alphabets with every component.



Step no. 2

Now balance each component:-

$$\text{Ca; } a = d \quad \text{(i)}$$

$$\text{C; } b + c = d \quad \text{(iv)}$$

$$\text{S; } a = e \quad \text{(ii)}$$

$$\text{H; } 4b = 2f \quad \text{(v)}$$

$$\text{O; } 4a + 2c = 3d + f \quad \text{(iii)}$$

Step no. 3 :-

As 6 unknown and only 5 equations generated;  
let the 6th equation;

$$a = 1.$$

$$a = d = e = 1.$$

Step no. 4

Now find the values of all alphabets by  
simultaneously solving.

Step 5:-

Now place in eqn.

$$a = 1$$

$$b = 8 \div 2 = 3 \frac{1}{2},$$

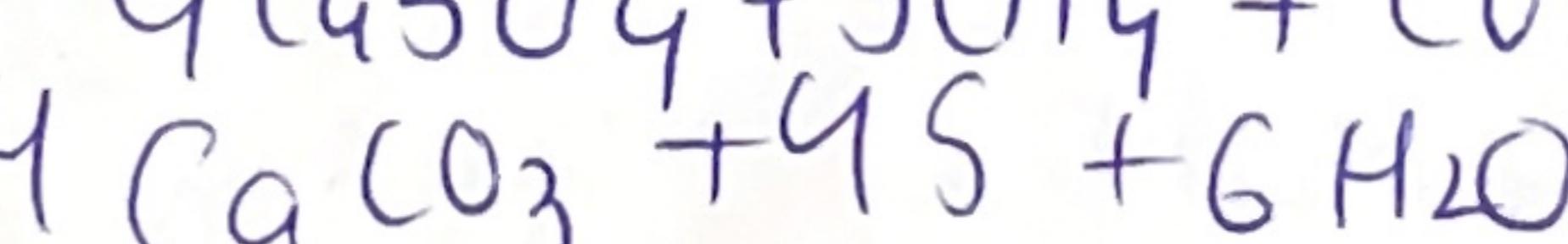
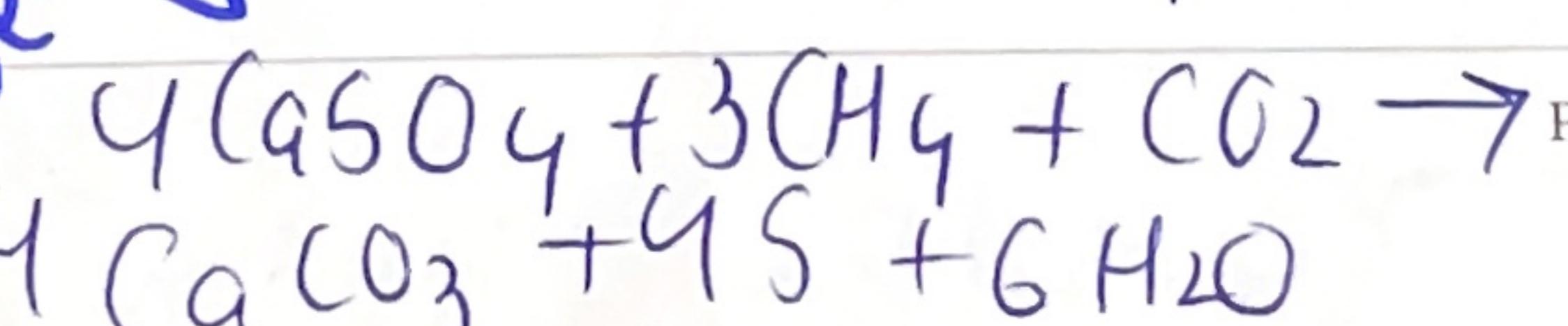
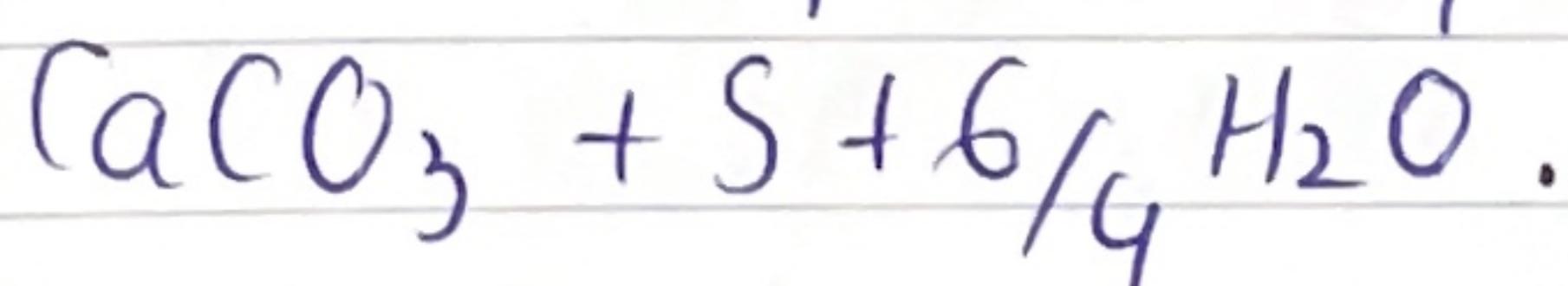
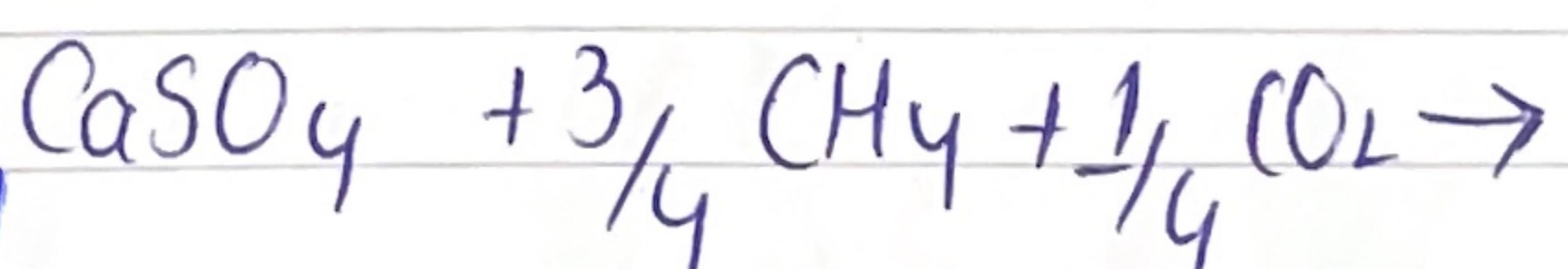
$$c = 15 - 1 = 14,$$

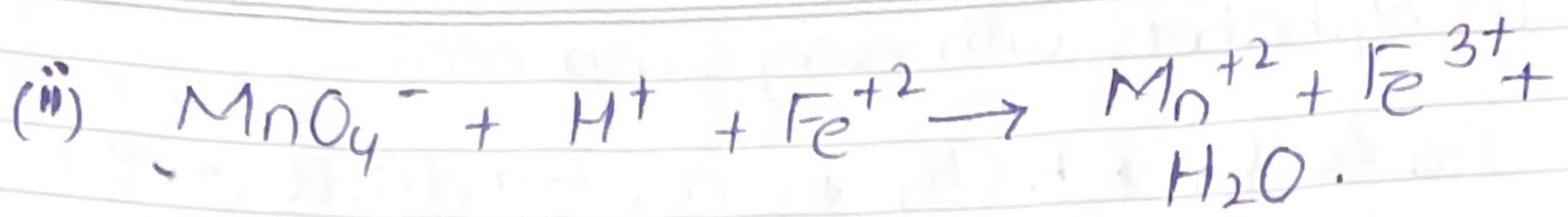
$$d = 1$$

$$e = 1$$

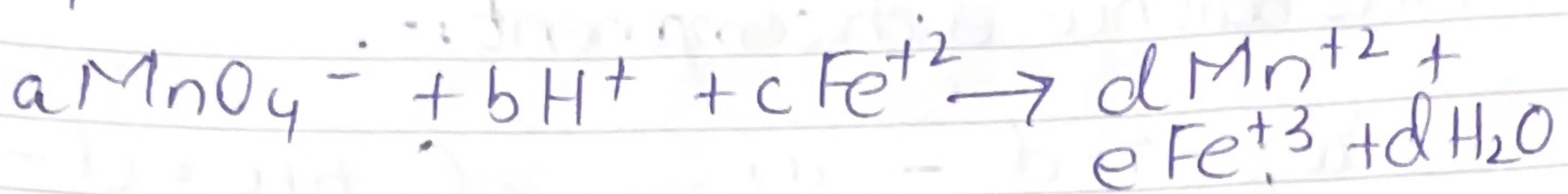
$$f = 6 \frac{1}{4}$$

By dividing always remove ratione answer!





Let alphabets :-



Mass balance :-

$$a = d$$

$$c = e$$

$$4a = f$$

$$b = 2f$$

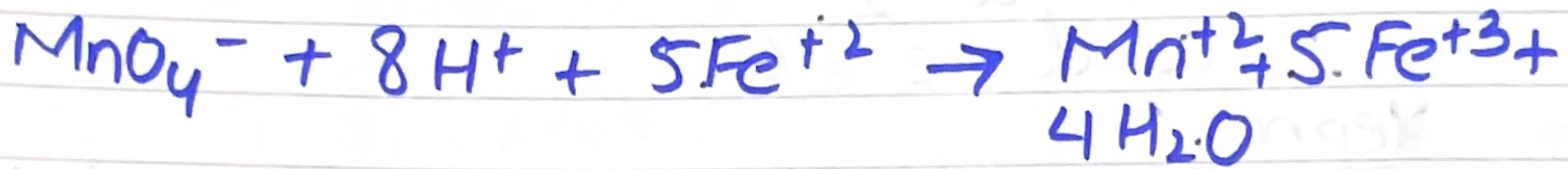
Charge balance :-

$$-a + b + 2c = 2 + 3e$$

Assuming

$$a = 1$$

$$\therefore e = 5, c = 5, a = 1, d = 1, \\ f = 4, b = 8.$$



Ans!

# CHAPTER NO.2

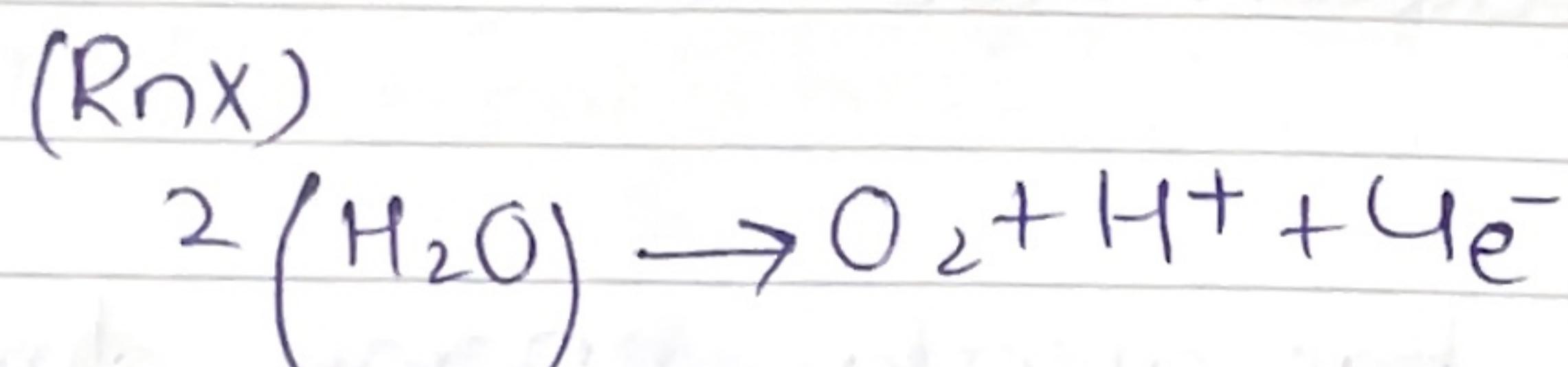
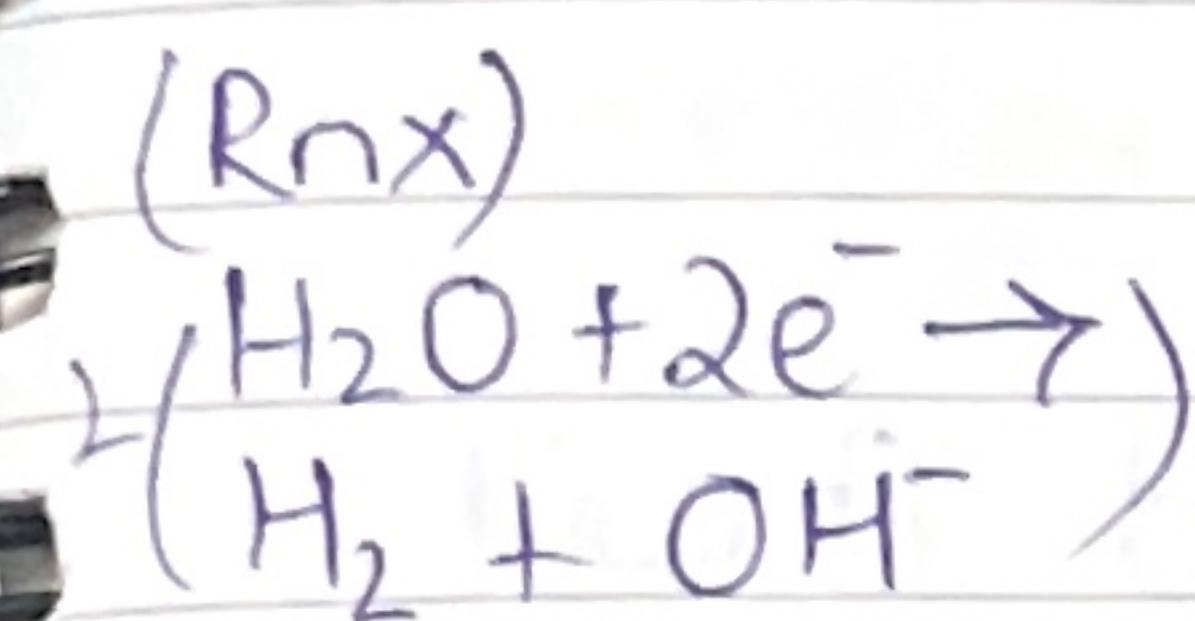
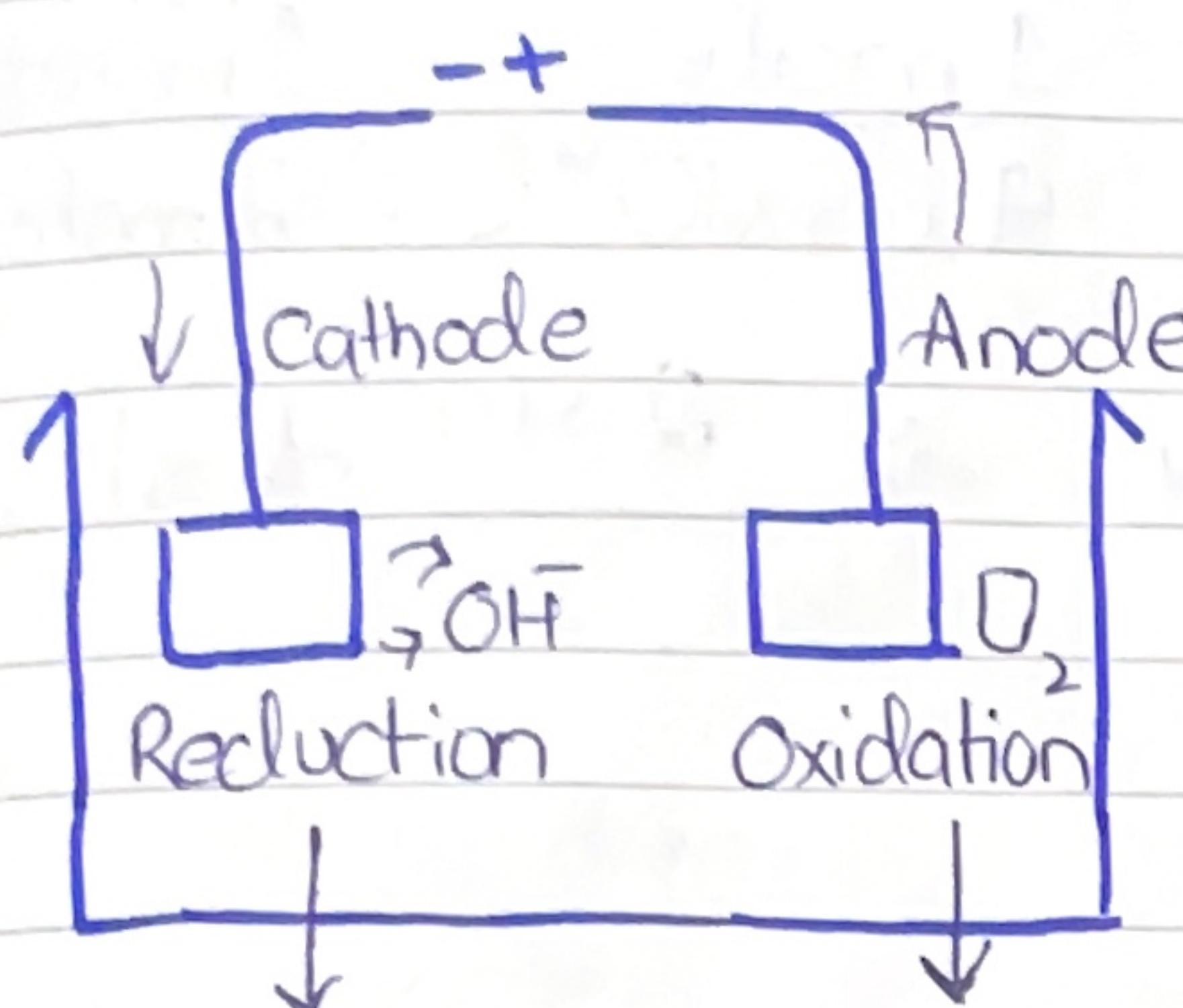
# ELECTROCHEMISTRY

Complete combustion products  $\Rightarrow \text{CO}_2, \text{H}_2\text{O}$ .

incomplete combustion' products  $\Rightarrow$  CO, unburnt hydrocarbons.

Galvanic  
Chemical + electrical.

Electrolytic.  
electrical to chemical.



## Conventions:-

- (i) Charge ( $z$ )
  - (ii) Voltage ( $v$ )
  - (iii) Current ( $I$ )

column(c)

$$V = \underline{\cup} \Rightarrow J \cdot C^{-1} \Rightarrow \underline{1}V$$

$$I = \frac{Z}{\Delta A} \Rightarrow \Delta A = \frac{1C \cdot s^{-1}}{Z}$$

$$1\bar{e} = 1.602 \times 10^{-19} C$$

$$\frac{1e^-}{1F} = \text{Charge on } 1\text{mol } A^- = 9.65 \times 10^4 \text{ C}$$

Q. calculate mass of Cu in g produced during the passage of 2.50 amperes of current for 50 minutes through a solution of Cu(II) sulphate.

Given

$$t = 50 \text{ minutes}$$

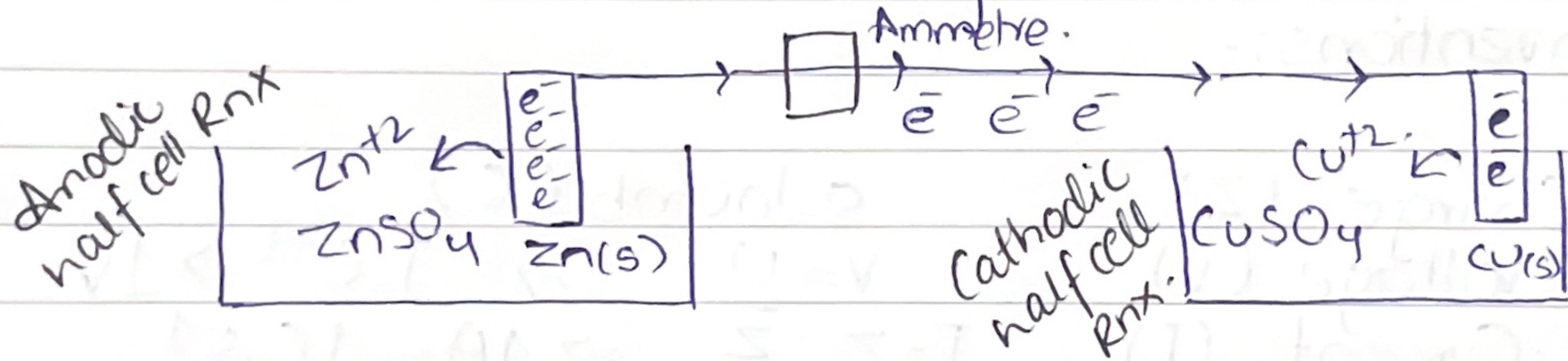
$$I = 2.5 \text{ A.}$$



$$\begin{aligned} \text{Mass of Cu} &= 2.5 \text{ A} \times 50 \text{ min} \\ &\times \frac{1 \text{ C} \cdot \text{s}^{-1}}{1 \text{ A}} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{1 \text{ mole}}{9.65 \times 10^4 \text{ C}} \times \frac{1 \text{ mol of Cu}}{2 \text{ mole of } 2\text{e}^-} \\ &\Rightarrow \times \frac{63.5 \text{ g of Cu}}{1 \text{ mol of Cu}} \Rightarrow \underline{\underline{2.47 \text{ gns}}} \end{aligned}$$

### Oxidation - Reduction Reaction:-

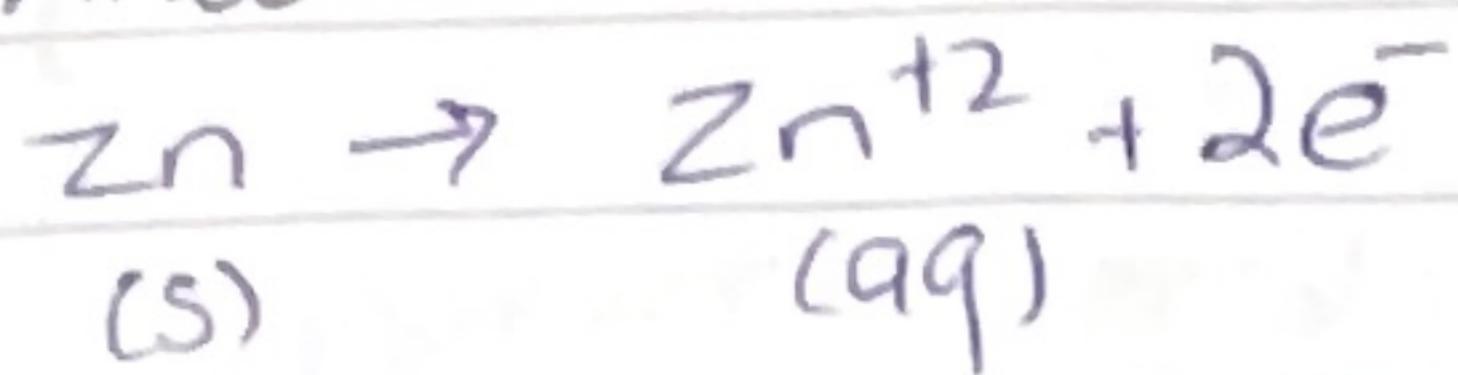
As the Zinc has greater energy stored , it will release more electrons during oxidation in a cell.



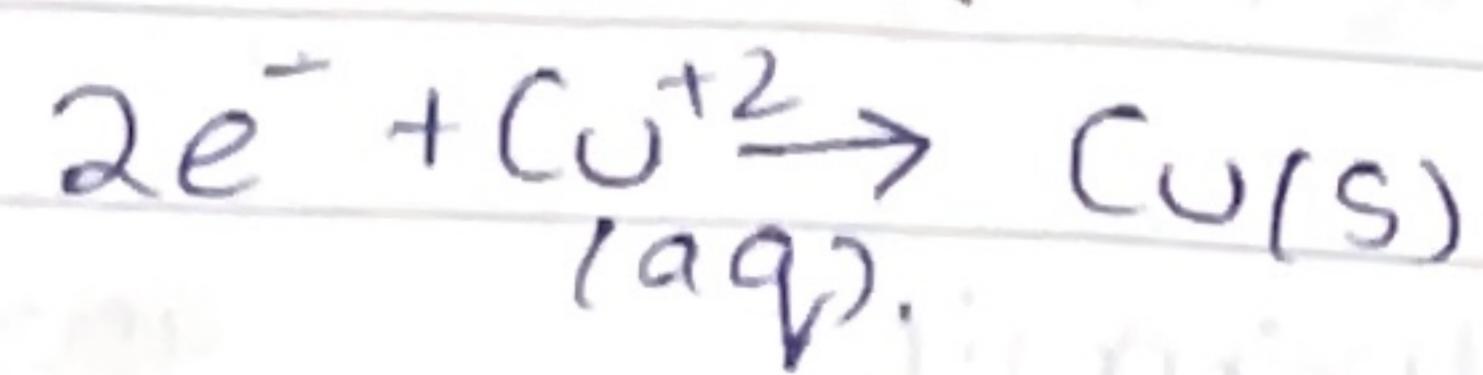
Now through the conductor electrons move from higher conc. to lower conc. i.e from zinc rod to Cu

Both are called Redox reaction.

Anodic Rnx



Cathodic Rnx

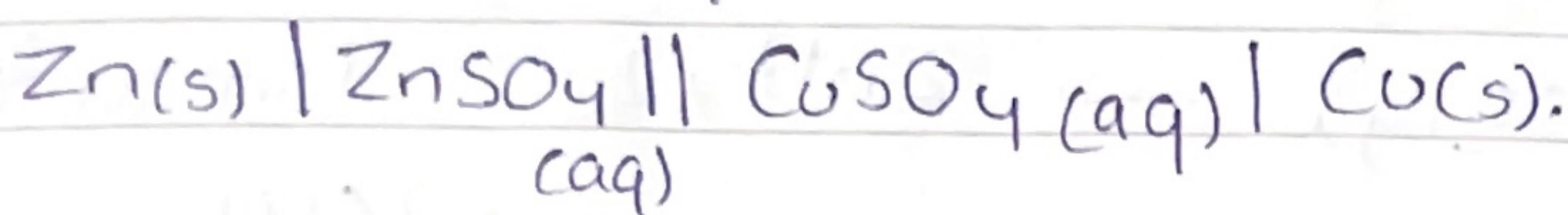


It has oxidation potential.

It has reduction potential.

- \* This reaction occurs in discharge however reverse of it would occur when all zinc is consumed & recharging occurs. (becomes a <sup>no</sup> galvanic cell), i.e electrolytic.

Cell diagram:- 'salt bridge'



                Anodic halfcell

                Cathodic halfcell.

Salt Bridge:-

- Maintains the neutral solution by discharging necessary ions in each half cell.
  - To act as a ionic charge carrier.
- however  $E_{Ox}$  &  $E_{Red}$  is measured at a specific point.

Cell overall voltage:-

$$E_{cell} = E_{Ox} + E_{Red.} \quad \begin{matrix} \xrightarrow{\text{no. of es}} \\ \xrightarrow{\text{F. C.}} \end{matrix} \begin{matrix} \text{Faraday const.} \\ \text{cell potential.} \end{matrix}$$

Gibbs Free Energy ( $\Delta G = -n \cdot F \cdot E$ ) tell the spontaneous or not of a rxn.

$\Delta G$  is an energy. (Gibbs Free Energy) tell the spontaneous or not of a rxn.

If  $G > 0$  non-spontaneous.  
 $G < 0$  spontaneous.

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

Date: \_\_\_\_\_ 20 \_\_\_\_\_  
 M T W T F S S

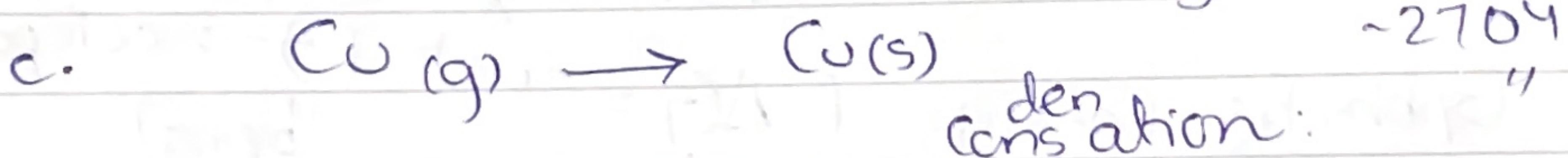
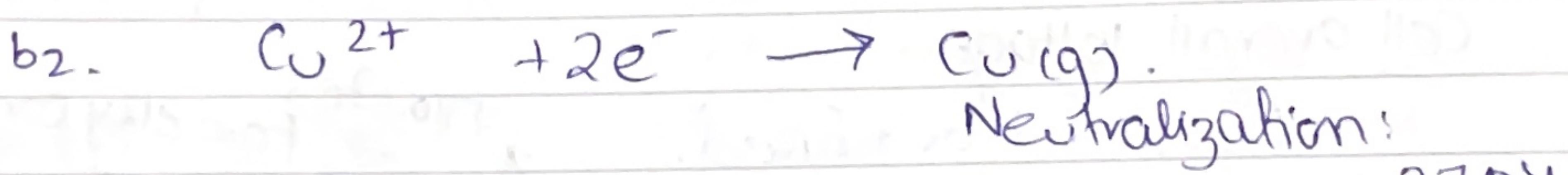
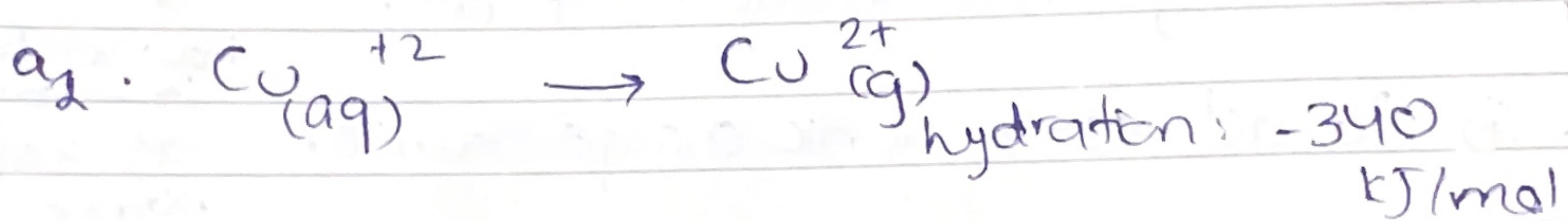
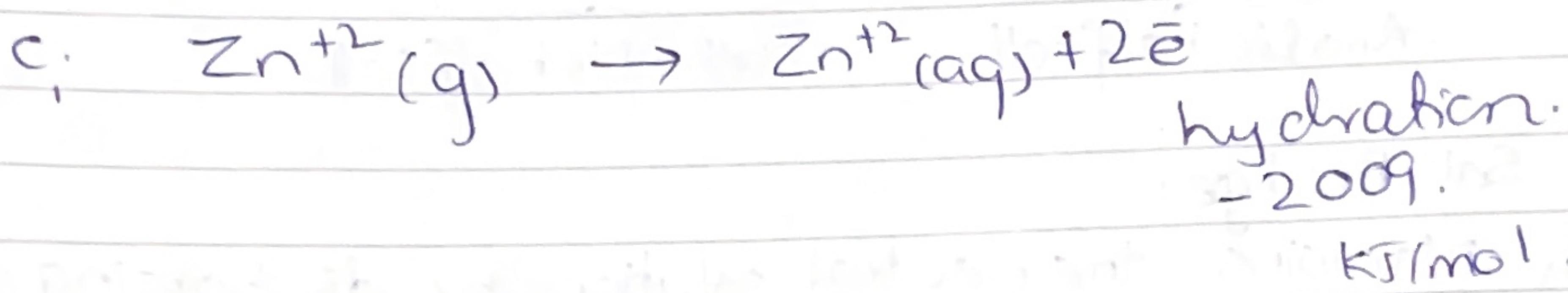
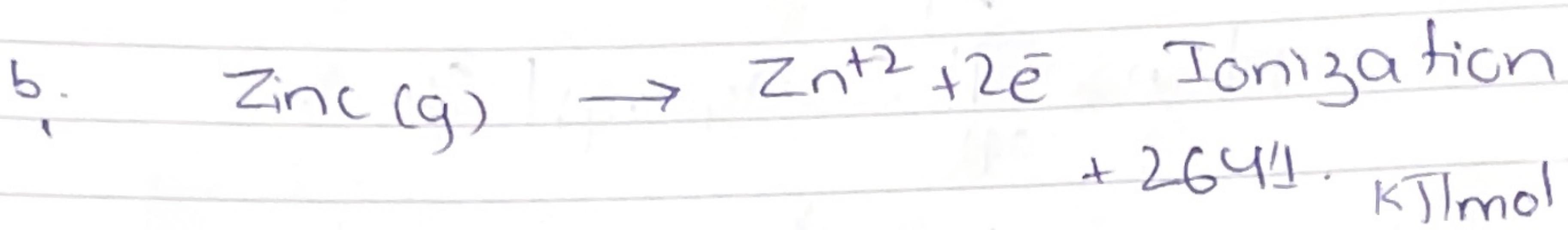
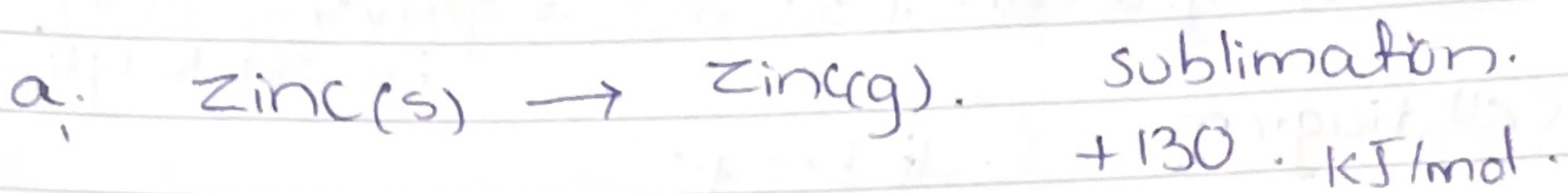
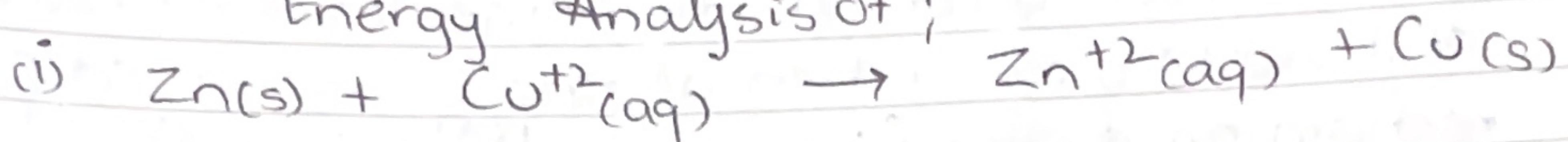
negative sign indicates release of energy

Endothermic non-spontaneous, needs energy.

Exothermic spontaneous, releases energy.

\* The energy delivered by a galvanic cell is spontaneous and exothermic.

### Energy Analysis of:



Sum of all value =  $-218 \text{ kJ/mol}$ .

How  $H_2 + 2e^-$  in SHE

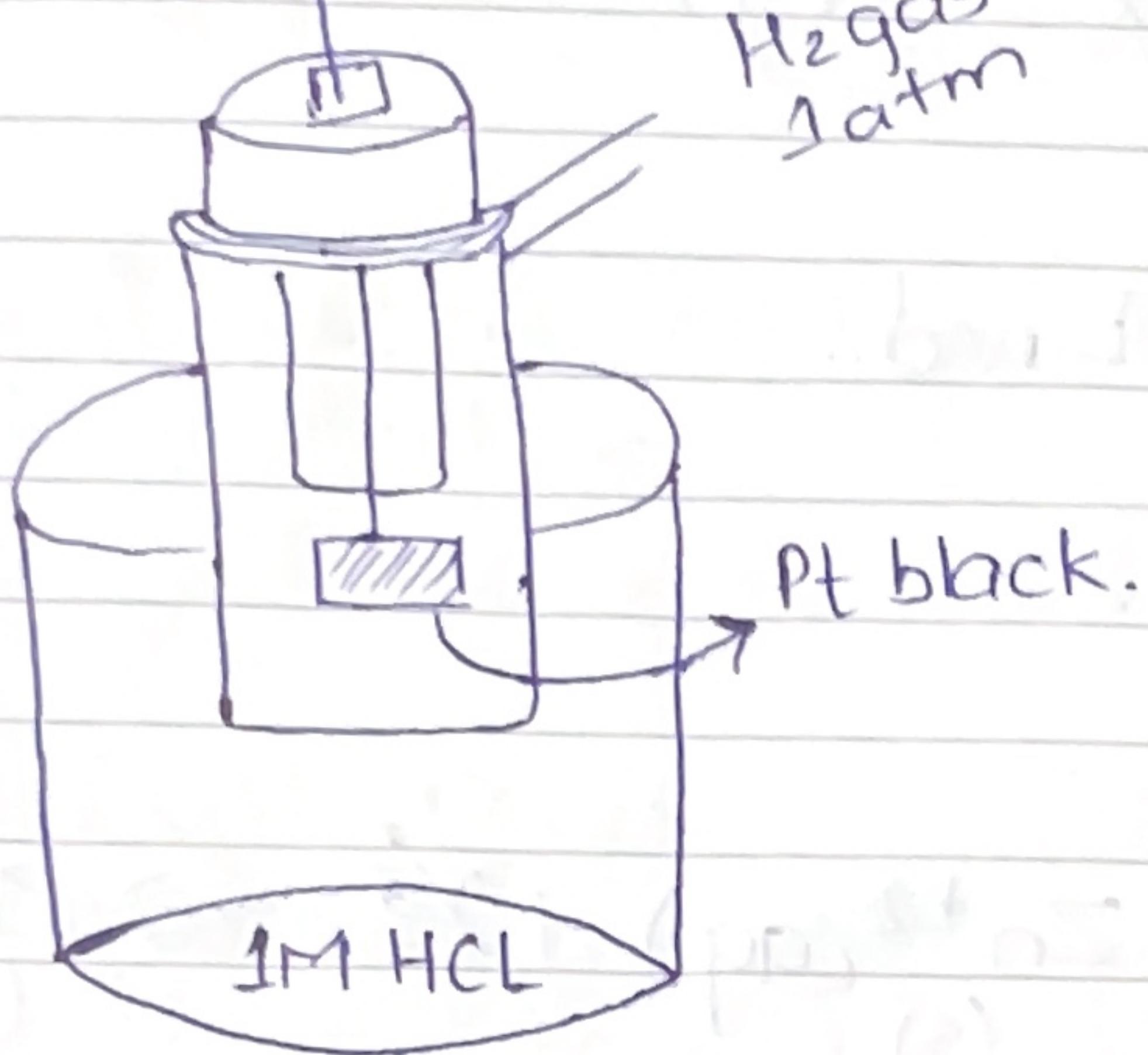
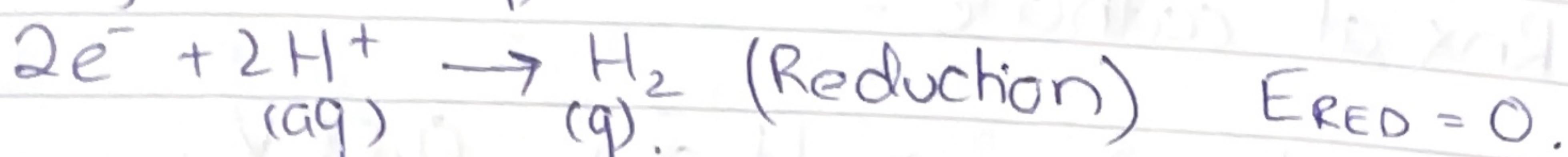
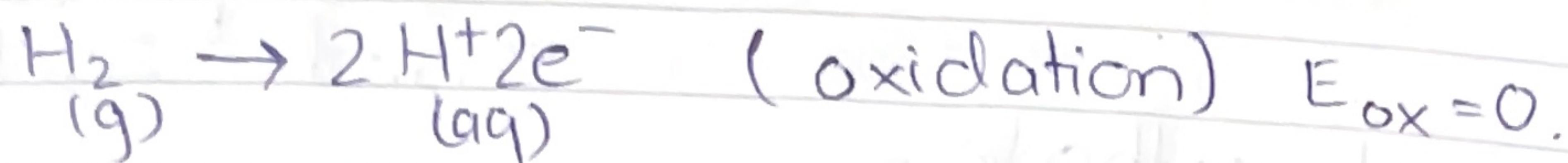
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Calculating electrode Potentials.

SHE (Standard Hydrogen electrode)

Pressure  $\rightarrow 1 \text{ atm}$ ; 1 M HCl, due to breaking bond by catalyst  $\uparrow$  by i.p.t

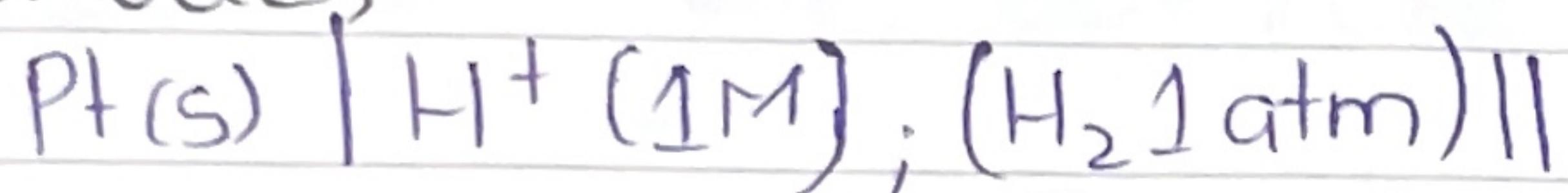
Hydrogen gas filled in which split into  $H^+$  ions and  $2e^-$  where  $H^+$  ions move to the sol. &  $2e^-$  remain at the platinum rod.



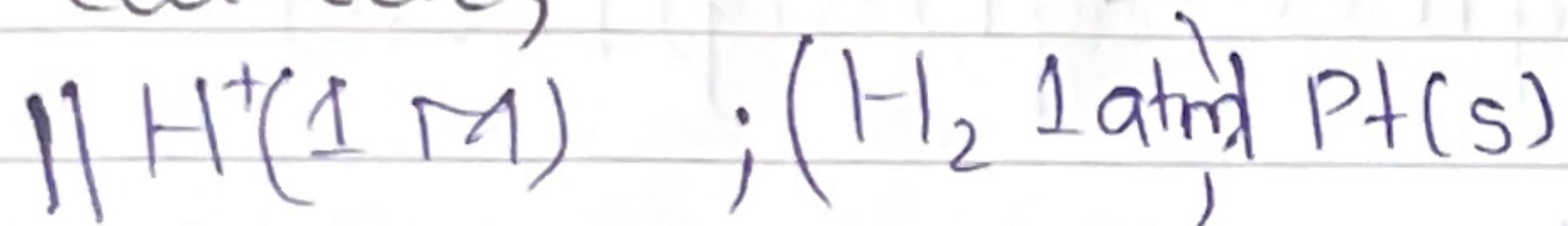
\* "It can act as both anode & cathode."

cell diagram;

As anode;



As cathode;

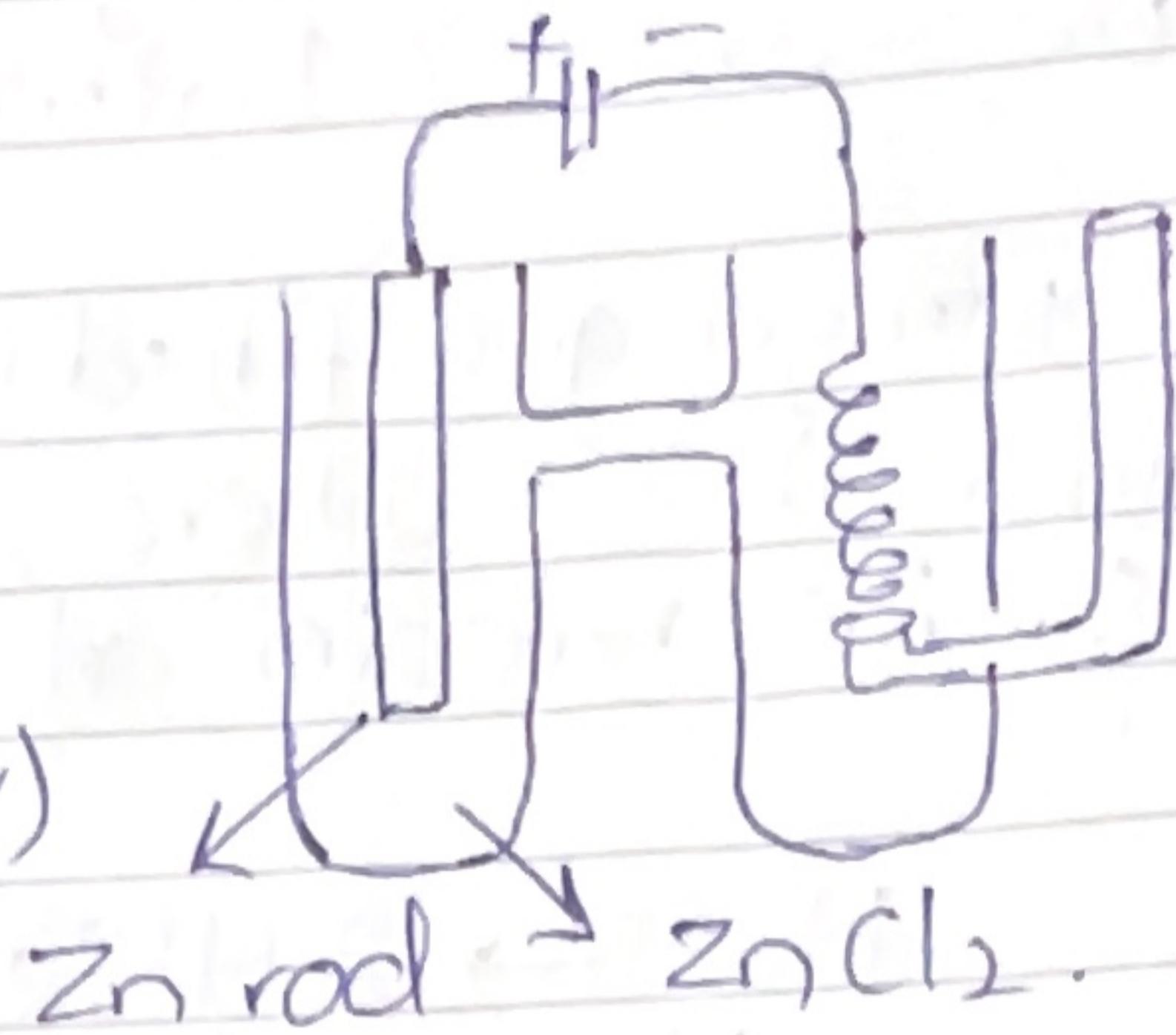
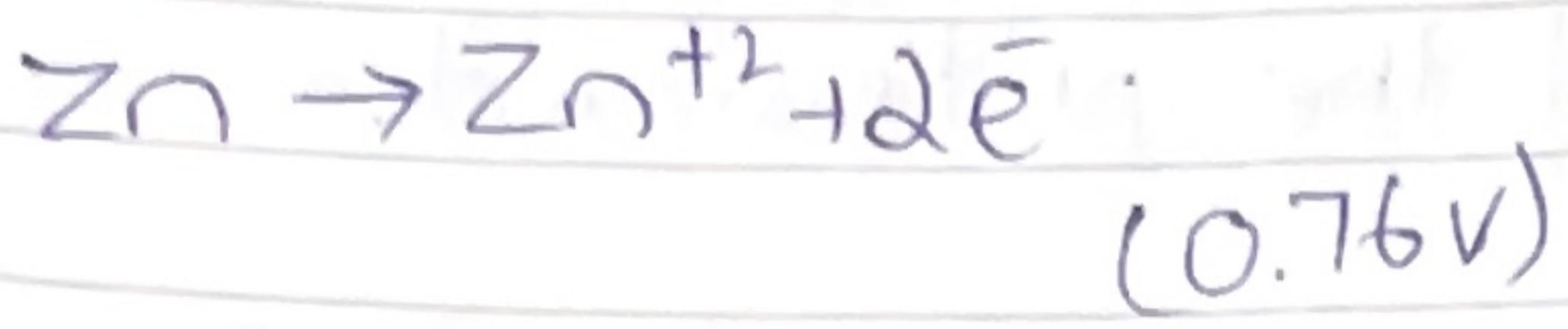


$\Delta G$  when negative } spontaneous  
E when positive rx.

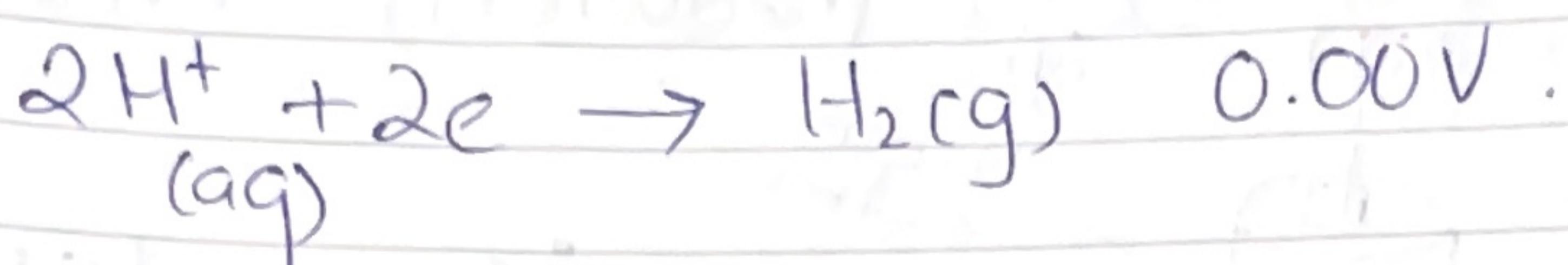
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### Zn - SHE cell.

Rnx at anode:-



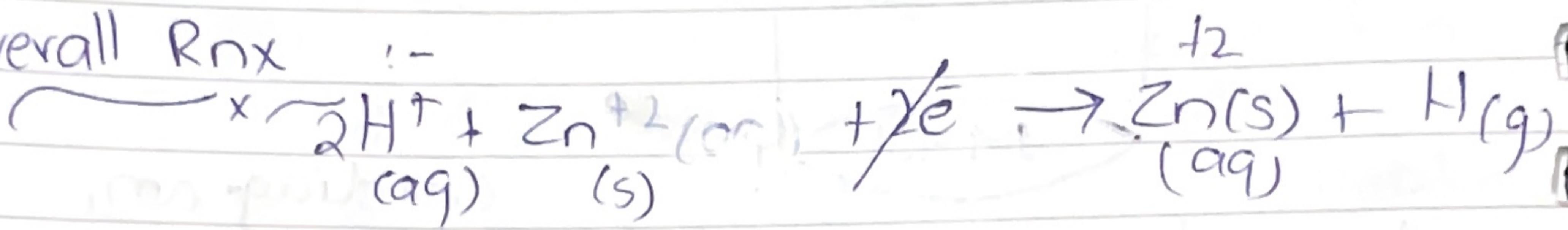
Rnx at cathode:-



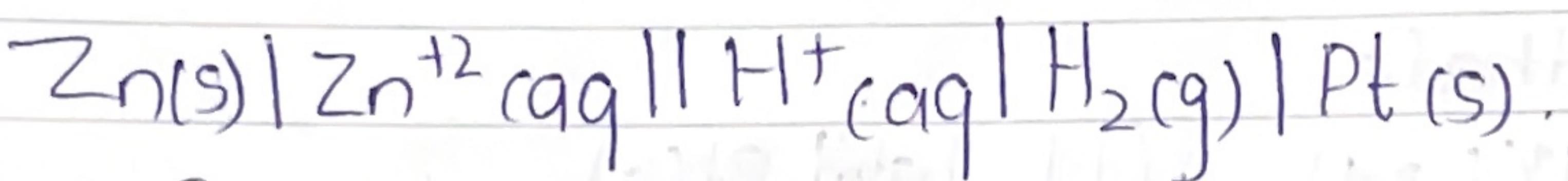
$$\therefore E_{cell} = E_{oxd} - E_{red}.$$

$$E_{cell} = 0.76V.$$

Overall Rnx :-

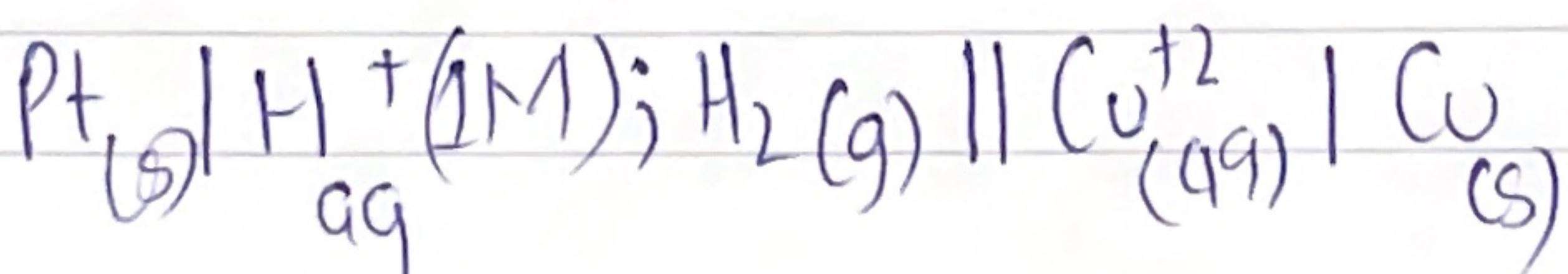


Cell diagram.



Cu - SHE cell :-

$\downarrow$        $\downarrow$   
Reduction    Oxidation.



Ohyah



(this tells about the change in voltage  
due to change in conditions; i.e. non-standard electrode)

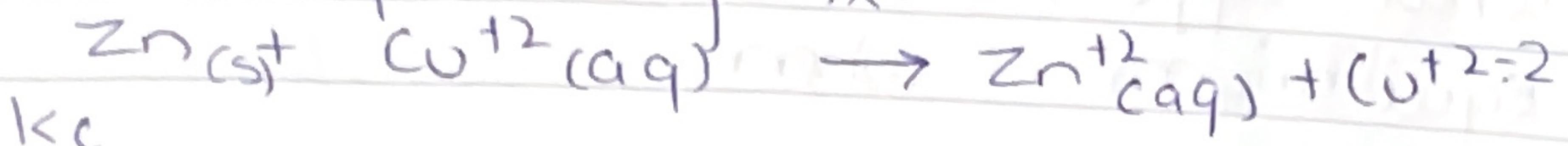
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## "NERNST EQUATION":-

i) In a galvanic cell rx when the cell is dead; (equilibrium is maintained).

Consider the following rx:-



$k_c$

$\Rightarrow \frac{\text{conc. of reactants}}{\text{conc. of reatant}} = \frac{\text{products.}}{}$

$$\Rightarrow \frac{[\text{Zn}][\text{Cu}^{+2}]}{[\text{Cu}][\text{Zn}^{+2}]}$$

ii) In a galvanic cell rx when the cell is working / alive; (equilibrium will not be maintained).

Rnx Quotient /  $\varphi \Rightarrow \frac{[\text{Zn}][\text{Cu}^{+2}]}{[\text{Cu}][\text{Zn}^{+2}]}$

(Remember that solid & liquids are pure and have 1 conc.)

## Non standard Electrode / Cell Potential :-

$$i) E = E^\circ - \frac{2303 RT}{nF} \log \varphi \rightarrow \text{rx quotient.}$$

(no. of es)  $\hookrightarrow$  Faraday const.  
at equilibrium,  $E=0$ . (no voltage discharge)

\* Conc. is time dependant (the time at which the rx proceeds until it stops.)

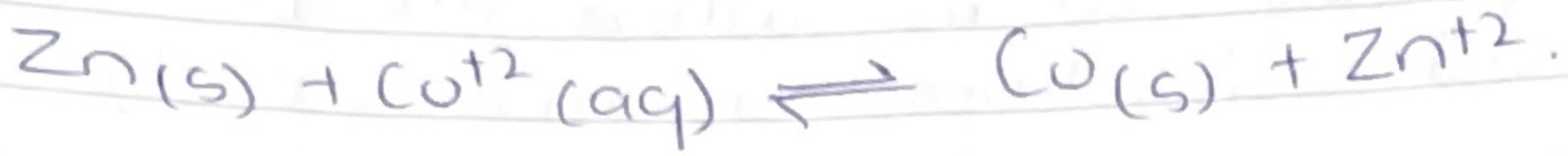
$$\frac{P}{RT} = \text{Conc of} = \frac{n}{V} \left( \frac{\text{moles}}{\text{volume}} \right)$$

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at (E)

## AFFECT OF CHANGE IN CONC:-



What happens when the cell voltage at equilibrium when

(i) conc. of  $Cu^{+2}$  is increased:- (Suppose  $\xrightarrow{1 \rightarrow 1.1M}$   
(thus the equilibrium is disturbed).

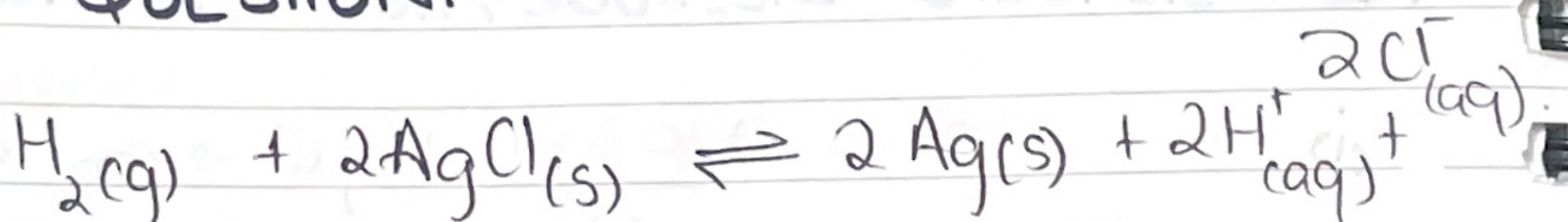
$$E = E_0 - \frac{0.03}{nF} RT \log \left[ \frac{Zn^{+2}(aq)}{Cu^{+2}(aq)} \right]$$

$$\log \left( \frac{1}{1.1} \right) \Rightarrow -$$

non-standard  
eventually it would increase the cell potential.  
 $(E)$

vice versa, i.e. the E would decrease if the conc. of  $Zn^{+2}$  is increased.

## PRACTICE QUESTION:-



takes place. In this rxn,  $H_2(g)$  reduces to  $AgCl(s)$  to silver metal. Predict the effect of the following changes observed in the cell voltage with

(i) increase in  $H_2(g)$  pressure.

\* equilibrium  $\Phi_c$  is converted into  $K_c$  which is a const value as conc. of reactants and products do not change. — (i)

$$E = E^\circ - \frac{2.303 RT}{nF} \log \frac{[H^+(aq)]^2 [Cl^-(aq)]^2}{[H_2(g)]}$$

(i) increasing the conc./pressure of  $H_2$  would result in more negative log factor and thus  $E^\circ$  would increase.

(ii) similarly  $E$  would decrease.

(iii) has no affect.

### EQUATION OF $K_c$

consider — (i)

$$E = 0$$

$$\Phi = K_c$$

$$2.303 \log = \ln$$

Now evolving Nernst equ.

$$\left[ E^\circ = \frac{RT \ln K_c}{nF} \right] — (ii).$$

Relating Nernst equ & (ii)

$$\text{Nernst equ} \Rightarrow E^\circ_{\text{cell}} = E^\circ - \frac{RT \ln \Phi}{nF}$$

$$(V) E_{\text{cell}} = \frac{RT \ln K_c}{nF} - \frac{RT \ln \Phi}{nF} \rightarrow (iii)$$

↓ voltage.

0.0570

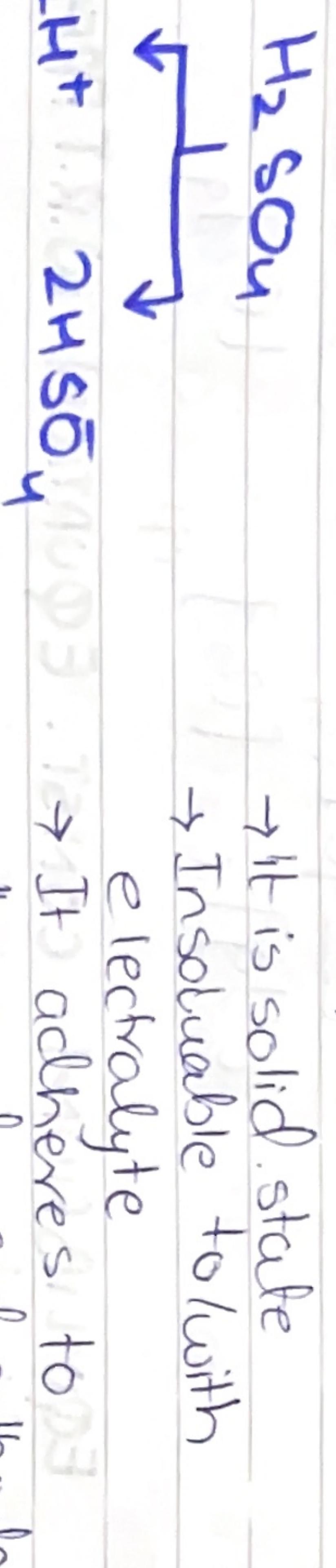
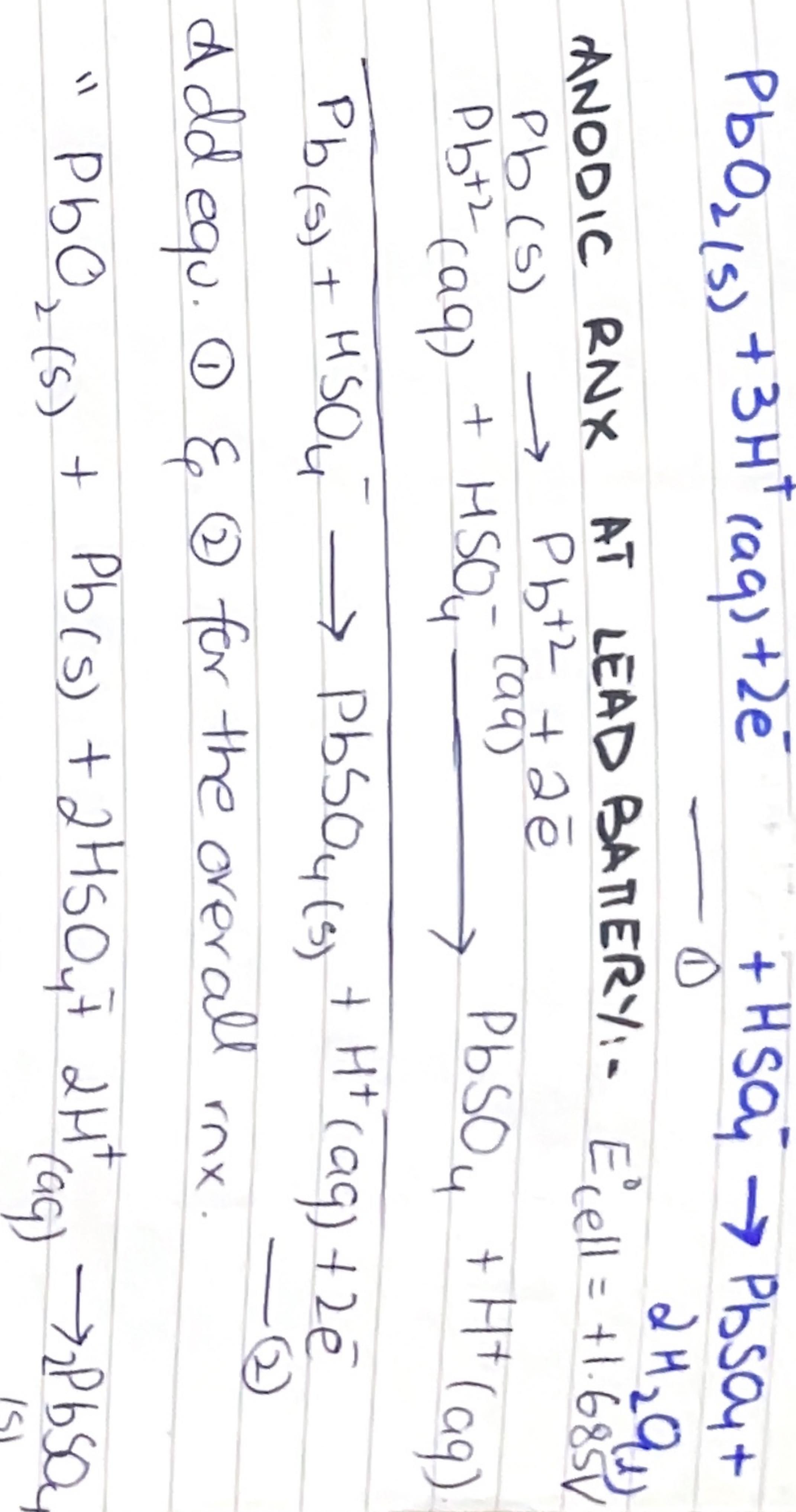
e<sup>8.6.</sup>

## LEAD STORAGE BATTERY:-

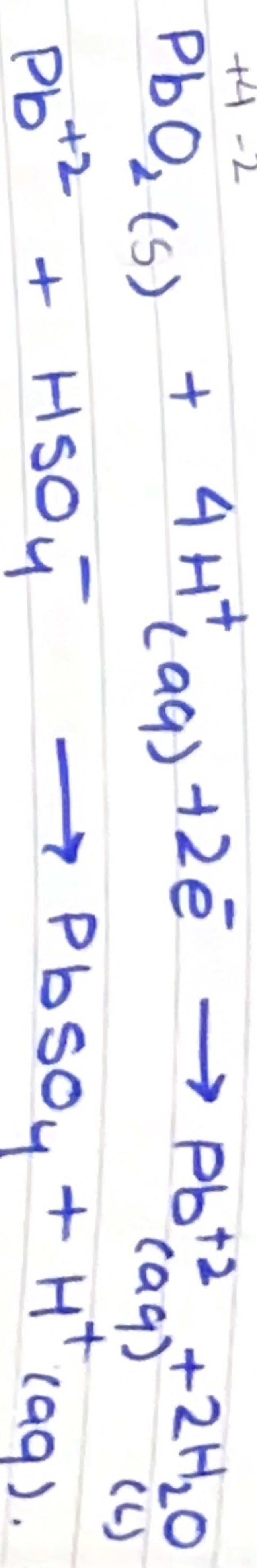
Invented in 1859.  
Temp range :  $-35^{\circ}\text{C}$  to  $+45^{\circ}\text{C}$   
due to the electrolyte.

40Wt% aq  $\text{H}_2\text{SO}_4$

- 2V cells in series to give a battery of 12 V.
- It can deliver 1A  $\text{I}_{\text{max}}$  for 45 hrs.
- It is rechargeable useful for cranking, lightening and ignition in id.
- The anode and cathode are first grinded and compressed to form a wall such that wall is then perforated & surface area increases.
- ∴ increasing rate of rxn.
- ∴ Decreases Resistance.



## CATHODIC RXN AT LEAD BATTERY:-



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