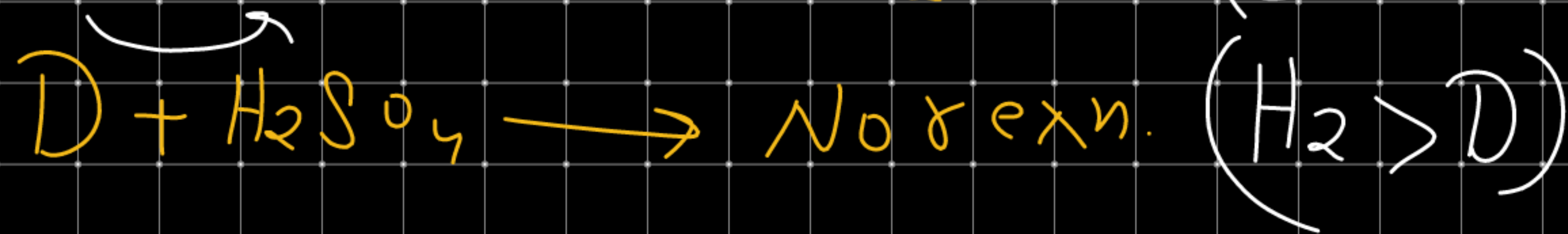
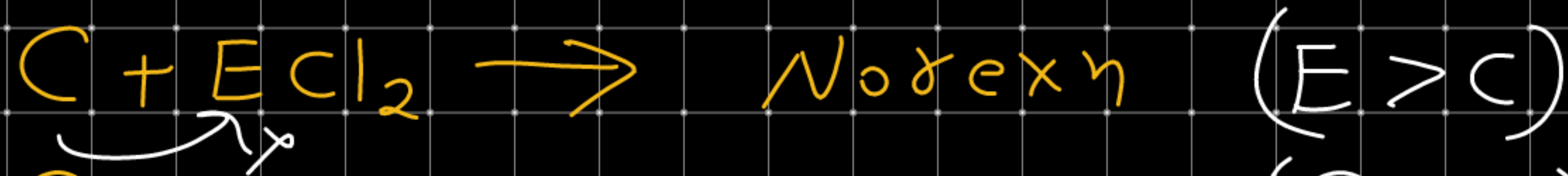
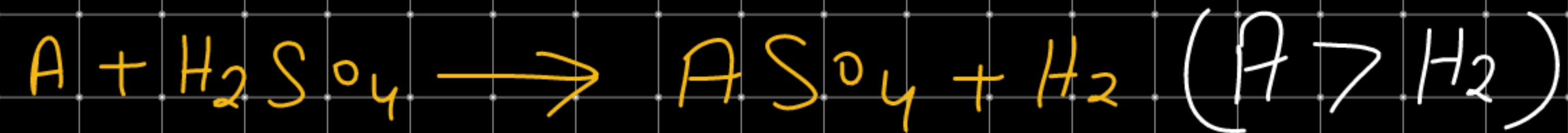


9.



{D, B, Si}

E
A
C
H₂
D
B

NEARNST EQUATION

This equation gives relation b/w ionic concentration, temperature and electrode potential.

let a rxn



Here :-

a, b, c, & d Stoichiometry
(coefficients).

Non P-V work
Charge potential difference = Useful ΔG

$$-\Delta G = n \cdot F \cdot E$$

potential
difference
/ E.M.F. of
Cell.

$$\Delta G = -nFE$$

From Thermodynamics.

$$\Delta G = \Delta G^\circ + RT \ln(Q_c)$$

Here ΔG = Gibbs energy
 ΔG° = Std Gibbs energy.
 Q_c = Reaction Quotient.

We know

$$1 e^- = 1.6 \times 10^{-19} C$$

Charge produce by
1 mole e^-

$$= N_A e^- \times 1.6 \times 10^{-19} C$$

$$= 6.023 \times 10^{23} \times 1.6 \times 10^{-19} C$$

$$\approx \underline{96500 C}$$

= 1 Faraday.

$$\mathcal{Q}_c = \frac{[P]^m}{[R]^n} \Rightarrow \text{Chemical Equilibrium.}$$

$$R = \text{Gas Constant} \left[\begin{array}{l} 8.314 \text{ J/k-mol} \\ \text{OR } 2 \text{ cal/k-mol} \end{array} \right]$$

$$T = \text{Temp in K}$$

$$\# \quad -nFE = -nFE^\circ + RT \ln(\mathcal{Q}_c)$$

$$E = E^\circ - \frac{RT}{nF} \ln(\mathcal{Q}_c)$$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{2.303 RT}{nF} \log \frac{[P]^m}{[R]^n} \Rightarrow \text{Nernst eq.}$$

$\frac{2.303 RT}{F} = 0.0591$

$R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$
 $T = 25^{\circ}\text{C} = 298 \text{ K}$
 $F = 96500 \text{ C}$

Cell potential

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591}{n} \log \frac{[P]^m}{[R]^n}$$

Std cell potential.

$n = \text{no. of transfer of } e^{-} \text{ in balance rxn.}$

Electrode potential depend on:

- (1) Nature of electrode
- (2) Concentration of electrolytic solution.
- (3) Temp.

Standard electrode potential depend on:

Only nature of electrode.

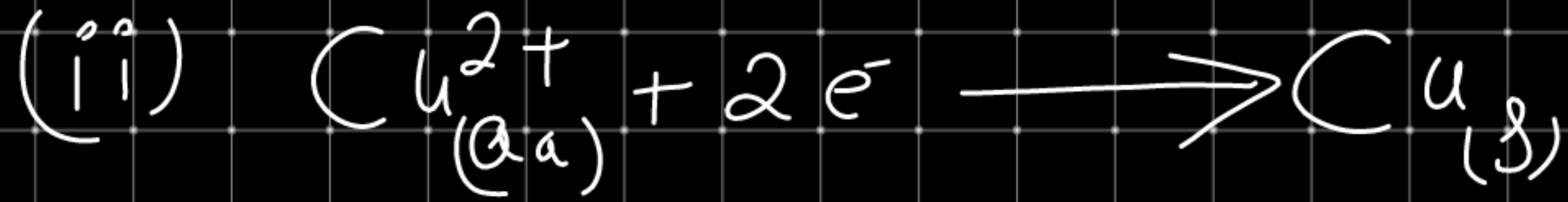
Application of Nernst Equation:

(1) To determine oxidation potential or reduction potential



$$E_{o.p.} = E^{\circ}_{o.p.} - \frac{0.059}{2} \log [Zn^{2+}]$$

IF $[Zn^{2+}] \uparrow$ $E_{o.p.} \downarrow$



$$E_{RP} = E^{\circ}_{RP} - \frac{0.0591}{2} \log \frac{1}{[\text{Cu}^{2+}]}$$

$$E_{RP} = E^{\circ}_{RP} + \frac{0.0591}{2} \log [\text{Cu}^{2+}]$$

If $[\text{Cu}^{2+}] \uparrow$ $E_{RP} \uparrow$

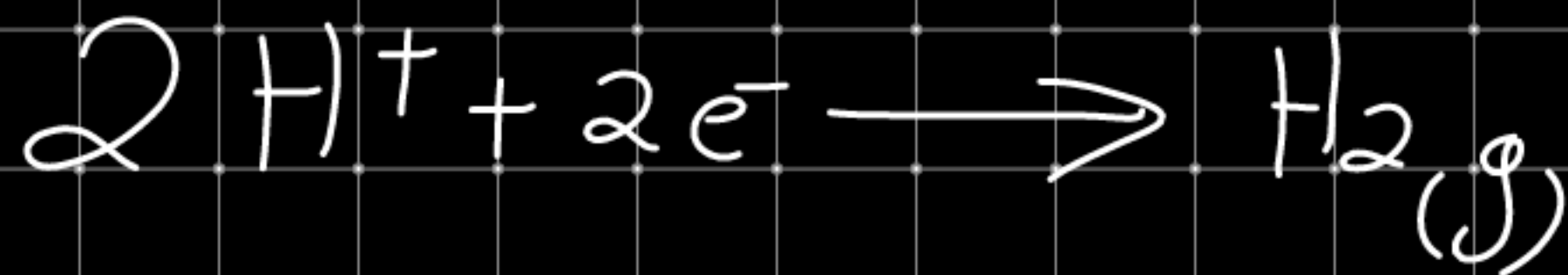
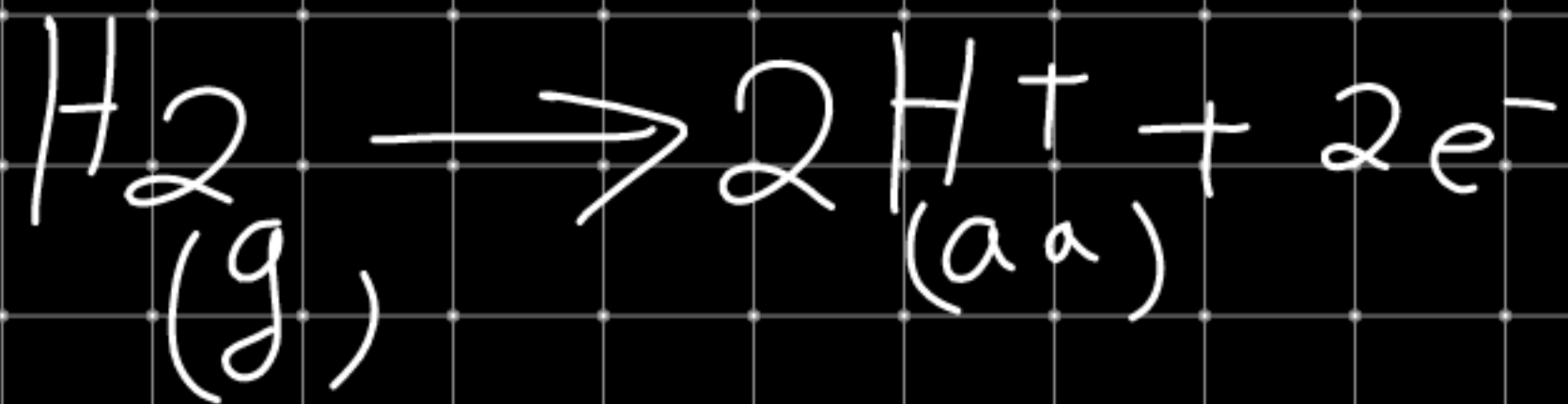
$$\left. \begin{aligned} \log\left(\frac{1}{x}\right) \\ = \log(x)^{-1} \\ = -\log(x) \end{aligned} \right\}$$

(2) To determine EMF of cell / E_{cell}°



$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

(3) To determine electrode potential H₂ electrode



$$E_{\text{O.P.}} = E_{\text{O.P.}}^{\circ} - \frac{0.059}{2} \log \frac{[\text{H}^+]^2}{(P_{\text{H}_2})}$$

if $P(\text{H}_2) = 1 \text{ atm}$

$$E_{\text{O.P.}} = 0 - \frac{0.059}{2} \log \frac{[\text{H}^+]^2}{1}$$

$$E_{o.p.} = - \frac{0.0591}{2} \times 2 \log [H^+]$$

$$E_{o.p.} = 0.0591 (-\log [H^+])$$

$$E_{o.p.} = 0.0591 \cdot pH$$

$$E_{p.p.} = -0.0591 pH$$

(4) To determine Equilibrium Constant (K_c) / (K_{eq})

At equilibrium, $E_{cell} = 0$ (No current flow in circuit)

$\Delta G = -nE_{cell} = -RT \ln K_c$

$$0 = E_{cell}^{\circ} - \frac{0.0591}{n} \log(K_c)$$

$$E_{cell}^{\circ} = \frac{0.0591}{n} \log(K_c)$$

$$\therefore \frac{2.303RT}{F} = 0.0591$$

$$\therefore E_{\text{cell}}^{\circ} = \frac{2.303RT}{nF} \log(K_c)$$

$$nF E_{\text{cell}}^{\circ} = 2.303RT \log(K_c)$$

$$-nF E_{\text{cell}}^{\circ} = -2.303RT \log(K_c)$$

$$\boxed{\Delta G^{\circ} = -2.303RT \log K_c}$$

#

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

(i) If $E^\circ_{\text{cell}} = +ve$
then $\Delta G^\circ = -ve$ $\{K_c > 1\}$
Cell is Feasible

(ii) If $E^\circ_{\text{cell}} = -ve$
then $\Delta G^\circ = +ve$ $\{K_c < 1\}$
Cell is non-feasible.

$\Delta G^\circ = -ve$
 \Rightarrow Spontaneous

$\Delta G^\circ = +ve$
 \Rightarrow Non Spontaneous