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Faraday's 1st law:→

Amount of substance produced at anode and cathode is directly proportional to the charge passed.

$W \propto Q$

Amount produced ↗ W; ↑ charge (Q)

$$\boxed{W = Z Q}$$

Z ↑ electrochemical equivalent

no. of equivalent = no. of moles × n-factor

n-factor of e^- = 1

charge on 1 mole e^- $= N_A \times 1.6 \times 10^{-19}$ Coulombs

$= 96500$ Coulombs

$= 1F$

Charge on 1 mole e^- = charge on 1 equivalent e^- $= 1F = 96500\ C$

= 1 mole charge = 1 equivalent charge

$$\text{equivalent of charge} = \frac{Q}{96500} = \frac{I \times t}{96500}$$

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$Na^+ + e^- \longrightarrow Na(s)$

1 mol   1 mol

1 eq   1 eq

$Ca^{2+} + \boxed{2e^-} \longrightarrow Ca(s)$

1 mol   2 mol

2 eq   2 eq

$Al^{3+} + 3e^- \longrightarrow Al(s)$

1 mol   3 mol

3 eq   3 eq

equivalents of charge passed = equivalents of substance produced at anode = equivalents of substance produced at cathode

Q. find the mass of Na(s) & $Cl_2(g)$ produced by 5 Faraday charge.

$$\text{equivalents of charge} = \frac{5 \times 96500}{96500} = 5$$

→ $Na^+ + e^- \longrightarrow Na$

n-factor = no. of e^- involved per molecule

no. of eq = no of moles = 5

of Na

→ $2Cl^- \longrightarrow Cl_2 + 2e^-$

$n_f = 1$ $n_f = 2$

5 eq

2.5 moles.

2.5 × 71 gm

Cl_2

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Q. find charge requires to produce 54 gm Al from Al^{3+} solution.

$$Al^{3+} + 3e^- \longrightarrow Al$$

$n_f = 3$

$= 2\,mol$

$= 2 \times 3\,eq = 6F = 6 \times 96500$ coulomb $= I \times t$

eq of charge $= \dfrac{Q}{96500}$

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$$W = Z\,Q$$

$$\rightarrow \quad E = Z \times 96500$$

$$\left(Z = \frac{E}{96500}\right)$$

$$\Rightarrow \boxed{W = \frac{E}{96500} \times Q}$$

mass of 1 mol = M

mass of 1 eq = E = Eq mass

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

$$E = \frac{M}{n}$$

Faraday's 2nd law :→ when same amount is passed through the different cells, the amount of substance produced at each electrode is directly proportional to their equivalent mass

$$W \propto E$$

$$\frac{W_1}{W_2} = \frac{E_1}{E_2}$$

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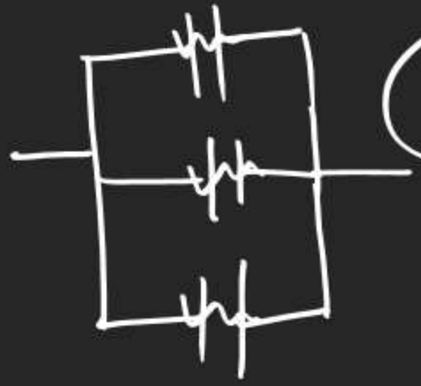
Q. Three electrolytic cells A, B & C containing $CuSO_4$, $AgNO_3$ & $AuCl_3$ respectively are electrolysed by 0.1 F charge. find the ratio of mass of metal deposited at cathode if these cells are connected in

Cu : 64
Ag : 108
Au : 198

(A) series :- $E_1 : E_2 : E_3$

$$\frac{64}{2} : \frac{108}{1} : \frac{198}{3}$$

(B)  Parallel and ratio of charge passed is 1 : 2 : 3

$$E_1Q_1 : E_2Q_2 : E_3Q_3$$

$$\frac{64}{2}\times 1 : \frac{108}{1}\times 2 : \frac{198}{3}\times 3$$

(C) Parallel and ratio of resistance is 3 : 2 : 1

$$\frac{1}{3} : \frac{1}{2} : \frac{1}{1}$$

ratio of charge $2 : 3 : 6$

$$\frac{64}{2}\times 2 : \frac{108}{1}\times 3 : \frac{198}{3}\times 6$$

$$64 : 324 : 396$$

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Case-II If solution contains more than one cation and anion

$NaCl(aq)$

The substance with higher electrode potential will be deposited first.

$\boxed{2A^+} + 2e^- \longrightarrow 2A \qquad E = 1\ \text{volt}$

$\boxed{B^{2+}} + 2e^- \longrightarrow B \qquad E = 0.75\ \text{volt}$

$\Delta G = -1 \times F \times 1$

$\Delta G = -2F \times 0.75$
$= -1.5F$

$\Delta G = -nFE$

the Rxn with more -ive value of $\left(\frac{\Delta G}{n}\right)$ will be deposited first

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Reduction potential of H_2O

$A^+(aq) + e^- \longrightarrow A \quad E^\circ = -0.3$ volt

$H^+(aq) + e^- \longrightarrow \frac{1}{2}H_2 \quad E^\circ_{H^+/H_2} = 0$

$2H_2O \longrightarrow \cancel{2H^+} + 2OH^- \quad K_w^2 = 10^{-28}$

$\cancel{2H^+} + 2e^- \longrightarrow H_2 \quad E^\circ_{H^+/H_2} = 0$

$2H_2O + 2e^- \longrightarrow H_2 + 2OH^- \quad E^\circ_{H_2O/H_2} = -0.84$

10^{-7} M

for 1 mol H^+

$10^{-7} \times V = 1$

$V = 10^7$ lit

$$\left(E_{H_2O/H_2}\right)_{pH=7} = -0.84 - \frac{0.06}{2}\log[OH^-]^2$$

$$= -0.84 - 0.06 \log 10^{-7}$$

$$= -0.84 + 0.42 = -0.42 = \left(E_{H_2O/H_2}\right)_{pH=7}$$

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J-Mains

3, 5, 6, 9, 10, 11, 12, 13

14, 15, 21, 22, 23, 24

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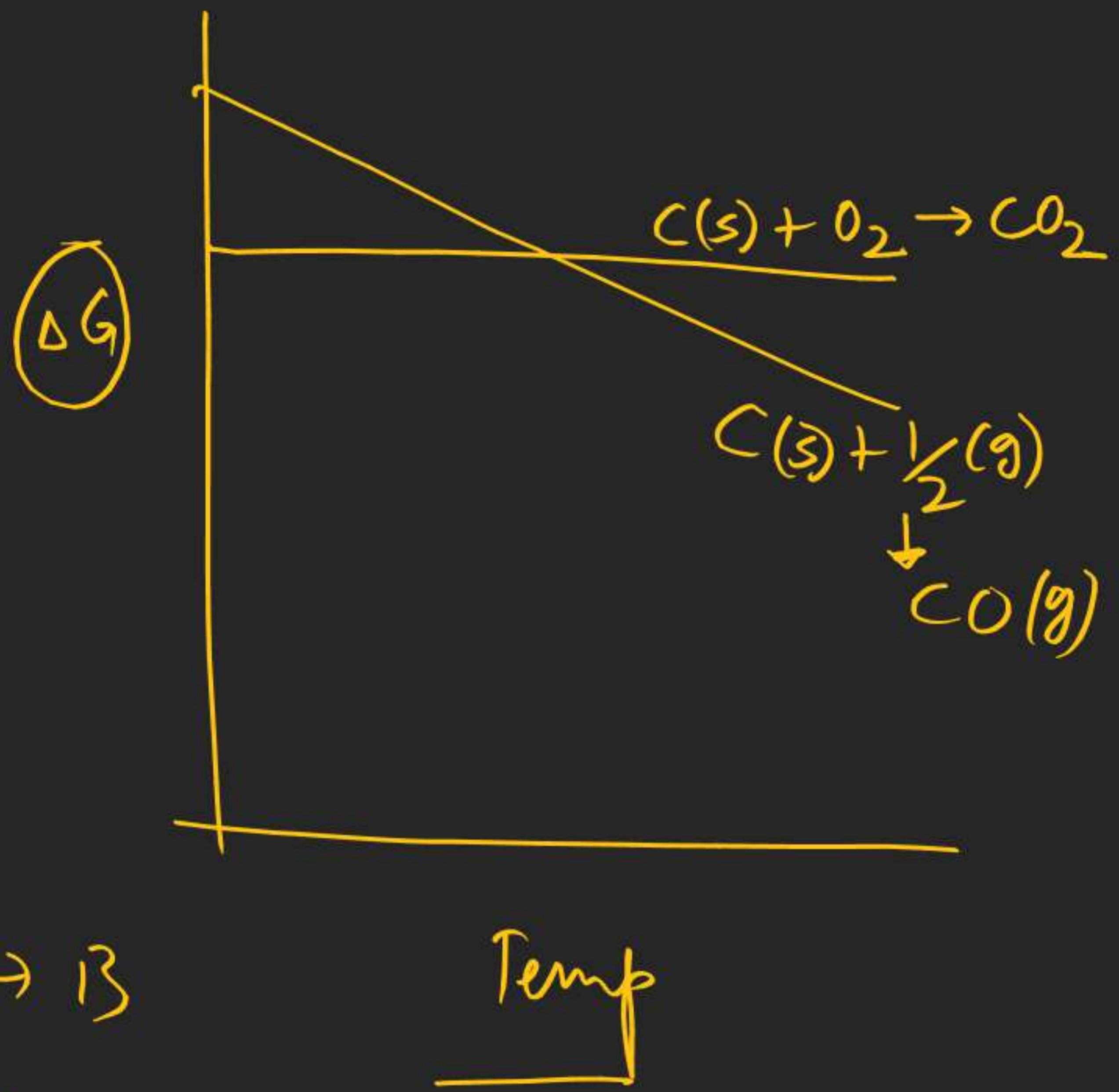
$\Delta G = \Delta H - T\Delta S$

$\Delta G_r = -600 \times \frac{2}{3} = -400 \qquad \frac{4}{3}Al + O_2 \rightarrow \frac{2}{3}Al_2O_3$

$\Delta G_r = -600 \qquad 2Al + \frac{3}{2}O_2 \rightarrow Al_2O_3$

$\Delta G_r = -500 \qquad 2Mg + O_2 \rightarrow 2MgO$

$\cancel{\Delta G_r = -1000 \qquad 4Mg + 2O_2 \rightarrow 4MgO}$



$A \rightarrow B$

$C \rightarrow D$

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$$A + O_2 \longrightarrow AO_2 \quad \Delta G = -3w$$

$$B + \cancel{O_2} \longrightarrow BO_2 \quad \Delta G = -5w$$

$$AO_2 \longrightarrow A + \cancel{O_2} \quad \Delta G = +3w$$

$$\underline{B} + \underline{\underline{AO_2}} \longrightarrow A + BO_2 \quad \Delta G = -2w$$

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S - I 38 - 42

O - I 46 - 51

J - AdV 2, 7, 8, 10, 11, 12

$$\frac{193000}{nFE}$$

$$2.67 = 2.7 - \frac{RT}{nF} \log \frac{Mg^{2+}}{Cu^{2+}}$$

$$= 2.7 - \frac{300}{2 \times 11500} \log x$$