

Topic :- Electro Chemistry

1 (c)

The Gibb's free energy change ΔG and emf (E°) of a reversible electrochemical cell are related by the following expression.

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

$$\text{or } = -nFE$$

2 (b)

$$E = E_{\text{RP}}^\circ + \frac{0.0591}{n} \log[M^+]$$

Given,

$$E_{\text{RP}}^\circ = -2.36 \text{ V}, [M^+] = 0.1 \text{ M}$$

$$n = 1 \text{ (for } M^+ \rightarrow M)$$

$$\begin{aligned} E &= E_{\text{RP}}^\circ + \frac{0.0591}{n} \log[M^+] \\ &= -2.36 + \frac{0.0591}{1} \log 0.1 \\ &= -2.36 + 0.0591 \times (-1) \\ &= -2.36 - 0.0591 \\ &= -2.419 \text{ V} \end{aligned}$$

3 (d)

1 faraday deposits 1 g equivalent of any substance.

64 (b)

$$\begin{aligned} \frac{1}{a} &= k \times R = 0.002765 \times 400 \\ &= 1.106 \text{ cm}^{-1}. \end{aligned}$$

5 (d)

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

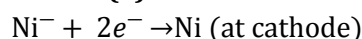
$$0.2905 = E_{\text{cell}}^\circ + \frac{0.059}{2} \log \frac{0.01}{0.10}$$

$$\therefore E_{\text{cell}}^\circ = 0.32$$

$$\text{No, } E_{\text{cell}}^\circ = \frac{0.059}{2} \log_{10} K$$

$$\begin{aligned} \therefore 0.32 &= \frac{0.059}{2} \log_{10} K \\ K &= 10^{0.32/0.0295} \end{aligned}$$

7 (a)



$$\text{Equivalent weight of Ni} = \frac{\text{mol.wt.}}{\text{gain electron}}$$

$$= \frac{58.7}{2}$$

$$= 29.35$$

$$i = 12A, t = 1h = 60 \times 60s,$$

$$Z = \frac{\text{eq.wt.}}{96500}$$

$$\text{Weight of deposit Ni} = \frac{Zit \times \text{efficiency}}{100}$$

$$= \frac{29.35 \times 12 \times 60 \times 60 \times 60}{96500 \times 100}$$

$$= 7.883 \text{ g}$$

68 (a)

$$\frac{W}{E} = \frac{i \times t}{96500}$$

$$\therefore \frac{W}{E} = 10^{-2} (\text{Ag is monovalent})$$

$$\therefore Q = i \times t = 96500 \times 10^{-2} = 965 \text{ C}$$

69 (a)

The tendency to gain electron is in the order $z > y > x$

Thus, $y + e^- \rightarrow y^-$

$x \rightarrow x^- + e^-$

10 (d)

NaCl, KNO₃, HCl are strong electrolytes but the size of H⁺ is smallest. Smaller the size of the ions, greater is the conductance and hence greater is the conductivity ($\kappa = C \times \text{cell constant}$).

11 (a)

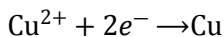
Given, $i = 2.5A$

$$t = 6 \text{ min } 26 \text{ s} = 6 \times 60 + 26 = 386s$$

$$\text{Number of coulomb passed} = i \times t$$

$$= 2.5 \times 386$$

$$= 965 \text{ C}$$



$$\therefore 2 \times 96500 \text{ C charge deposits Cu} = 63.5 \text{ g}$$

$$\therefore 965 \text{ C charge deposits}$$

$$\text{Cu} = \frac{63.5}{2 \times 96500} \times 965$$

$$= 0.3175 \text{ g}$$

12 (c)

Metal placed above in electrochemical series replaces the other from its salt solutions.

73 (c)

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

$$\therefore 1.1 = 0.78 + E_{RP_{Cu}}^{\circ} + \frac{0.059}{2} \cdot 1$$

$$\therefore E_{RP_{Cu^{2+}/Cu}}^{\circ} = 0.32$$

$$\therefore E_{RP_{Cu^{2+}/Cu}}^{\circ} = -0.32V$$

14 (d)

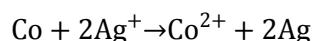
More the reduction potential, more is the power to get itself reduced or lesser is reducing power or greater is oxidizing power

15 (a)

Quantity of current is charge, *i.e.*, coulomb or ampere sec.

16 (d)

Cobalt is anode, *ie*, oxidation takes place on cobalt electrode *ie*, cell reaction is

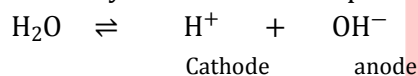


$$E_{cell} = E_{cell}^{\circ} - \frac{RT}{nF} \ln \frac{[Co^{2+}]}{[Ag^{+}]^2}$$

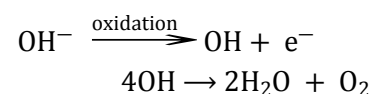
Thus, less is the factor $\frac{[Co^{2+}]}{[Ag^{+}]}$, greater is the E_{cell}°

17 (c)

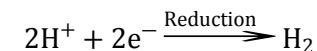
Electrolysis of water takes place as follows



At anode



At cathode



Given, time, $t = 1930s$

Number of moles of hydrogen collected

$$= \frac{1120 \times 10^{-3}}{22.4} \text{ moles}$$

$$= 0.05 \text{ moles}$$

\therefore 1 mole of hydrogen is deposited by = 2 moles of electrons

\therefore 0.05 moles of hydrogen will be deposited by

$$= 2 \times 0.05$$

$$= 0.10 \text{ mole of electrons}$$

Charge, $Q = nF$

$$= 0.1 \times 96500$$

Charge, $Q = it$

$$0.1 \times 96500 = i \times 1930$$

$$i = \frac{0.1 \times 96500}{1930}$$

PE

$$= 5.0 \text{ A}$$

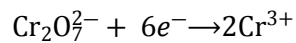
18 **(d)**

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

2. $E^\circ_{\text{cell}} = \frac{2.303 RT}{nF} \log K_c$

3. $k = Ae^{-E_a/RT}$

19 **(b)**



Reduction of 1 mol of $\text{Cr}_2\text{O}_7^{2-}$ to Cr^{3+} required 6 moles of electrons. Hence, charge required = $2 \times 96500 \text{ C}$

20 **(b)**

$$\text{Cell constant} = \frac{1}{a} = \frac{\text{length}}{\text{area}}$$

\therefore unit is cm^{-1} .

PE

ANSWER-KEY										
Q.	1	2	3	4	5	6	7	8	9	10
A.	C	B	D	B	D	D	A	A	A	D
Q.	11	12	13	14	15	16	17	18	19	20
A.	A	C	C	D	A	D	C	D	B	B

PE