DPP
DAILY PRACTICE PROBLEMS

Class : XIIth | **Solutions** | Subject : CHEMISTRY
Date : | | DPP No. : 4

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.

$\Delta G = -nFE^\circ_{cell}$

or $= -nFE$

2 **(b)**

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

Given,

$E^\circ_{RP} = -2.36$ V, $[M^+] = 0.1\ M$

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

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

$$= -2.36 + \frac{0.0591}{1}\log 0.1$$

$$= -2.36 + 0.0591 \times (-1)$$

$$= -2.36 - 0.0591$$

$$= -2.419\ V$$

3 **(d)**

1 faraday deposits 1 g equivalent of any substance.

64 **(b)**

$$\frac{1}{a} = k \times R = 0.002765 \times 400$$

$$= 1.106\ \text{cm}^{-1}.$$

5 **(d)**

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

$$0.2905 = E^\circ_{cell} + \frac{0.059}{2}\log\frac{0.01}{0.10}$$

$\therefore\quad E^\circ_{cell} = 0.32$

No, $E^\circ_{cell} = \frac{0.059}{2}\log_{10} K$

$\therefore\quad 0.32 = \frac{0.059}{2}\ \log_{10} K$

$$K = 10^{0.32/0.0295}$$

7 **(a)**

$Ni^{-} + 2e^{-} \rightarrow Ni$ (at cathode)

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

$= \frac{58.7}{2}$

$= 29.35$

$i = 12\text{A}, t = 1\text{h} = 60 \times 60s.,$

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

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

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

$= 7.883$ g

68 **(a)**

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

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

$\therefore Q = i \times t = 96500 \times 10^{-2} = 965$ 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, KNO_3, 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

($\kappa = C \times$ cell constant).

11 **(a)**

Given, $i = 2.5A$

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

Number of coulomb passed $= i \times t$

$= 2.5 \times 386$

$= 965$ C

$Cu^{2+} + 2e^- \longrightarrow Cu$

$\therefore$ 2×96500 C charge deposits Cu $= 63.5$ g

$\therefore$ 965 C charge deposits

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

$= 0.3175$ g

12 **(c)**

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

73 **(c)**

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

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

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

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

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

$Co + 2Ag^{+} \rightarrow Co^{2+} + 2Ag$

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

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

17 **(c)**

Electrolysis of water takes place as follows

$H_2O \rightleftharpoons \underset{\text{Cathode}}{H^{+}} + \underset{\text{anode}}{OH^{-}}$

At anod

$OH^{-} \xrightarrow{\text{oxidation}} OH + e^{-}$

$4OH \longrightarrow 2H_2O + O_2$

At cathode

$2H^{+} + 2e^{-} \xrightarrow{\text{Reduction}} H_2$

Given, time, $t = 1930s$

Number of moles of hydrogen collected

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

$= 0.05$ moles

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

$\because$ 0.05 moles of hydrogen will be deposited by

$= 2 \times 0.05$

$= 0.10$ mole of electrons

Charge, $Q = nF$

$= 0.1 \times 96500$

Charge, $Q = it$

$0.1 \times 96500 = i \times 1930$

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

$= 5.0 \text{ A}$

18 **(d)**

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

2. $E^{\circ}_{cell} = \dfrac{2.303\ RT}{nF} \log K_c$

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

19 **(b)**

$$Cr_2O_7^{2-} + 6e^- \longrightarrow 2Cr^{3+}$$

Reduction of 1 mol of $Cr_2O_7^{2-}$ to Cr^{3+} required 6 moles of electrons. Hence, charge required $= 2 \times 96500$ C

20 **(b)**

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

$\therefore$ unit is cm^{-1}.

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**