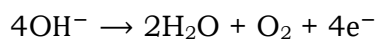


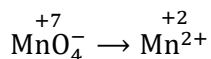
## Solutions

**S1.** Ans. (d)

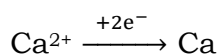


for 2 mole of  $\text{H}_2\text{O} = 4F$  charge is required

for 1 mole of  $\text{H}_2\text{O} = \frac{4F}{2} = 2F$  required

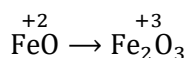


for 1 mole  $\text{MnO}_4^-$  5F charge is required



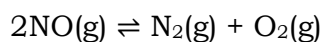
for 1 mole  $\text{Ca}^{2+}$  ion required = 2F

1.5 mole  $\text{Ca}^{2+}$  ion required =  $\frac{2}{1} \times 1.5 = 3F$



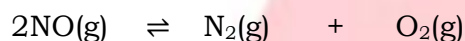
for 1 mole  $\text{FeO}$ , 1F charge is required.

**S2.** Ans. (c)



$$K_c = \frac{[\text{N}_2][\text{O}_2]}{[\text{NO}]^2}$$

$$= \frac{3 \times 10^{-3} \times 4.2 \times 10^{-3}}{2.8 \times 10^{-3} \times 2.8 \times 10^{-3}} = 1.607$$



	t = 0	0.1	0	+	0		0
		0.1 - 0.1α	0.05α		0.05α		0.05α

$$K_c = \frac{0.05\alpha \times 0.05\alpha}{(0.1 - 0.1\alpha)^2}$$

$$K_c = \frac{0.5\alpha \times 0.05\alpha}{0.01(1 - \alpha)^2}$$

$$1.607 = \frac{(0.05)^2 \alpha^2}{0.01(1 - \alpha)^2}$$

$$\frac{\alpha^2}{(1 - \alpha)^2} = \frac{1.607 \times (0.1)^2}{(0.05)^2}$$

$$\frac{\alpha}{1 - \alpha} = \frac{1.27 \times 0.1}{0.05}$$

$$\frac{\alpha}{1 - \alpha} = 2.54$$

$$3.54 \alpha = 2.54$$

$$\alpha = \frac{2.54}{3.54} = 0.717$$

**S3.** Ans. (b)

$\text{Ti}^{3+}$  act as an oxidising agent not reducing agent.

**S4.** Ans. (d)

B and D statements are correct.

**S5.** Ans. (a)

$$\begin{aligned} E_{\text{cell}}^{\circ} &= E_{\text{C}}^{\circ} - E_{\text{A}}^{\circ} \\ &= (1.33) - (-0.44) \\ &= +1.77 \text{ V} \end{aligned}$$

**S6.** Ans. (b)

Centimolar solution = 1/100 M = 0.01 M

Conductivity (k) = 0.0210 ohm<sup>-1</sup> cm<sup>-1</sup>

Resistance (R) = 60 ohm

$$k = 1/R \left( \frac{l}{A} \right)$$

$$\Rightarrow 0.0210 = 1/60 \left( \frac{l}{A} \right) \Rightarrow \frac{l}{A} = 1.26 \text{ cm}^{-1}$$

**S7.** Ans. (d)

$$\Delta_{\text{rG}} = -nFE_{\text{cell}}$$

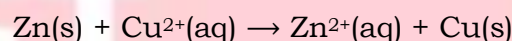
$E_{\text{cell}}$  is an intensive property and  $\Delta_{\text{rG}}$  is an extensive property as it depends on number of  $\text{e}^-$  transferred in cell reaction.

**S8.** Ans. (d)

Since,  $E_{\text{OP}}^{\circ}$  of Al is more than  $\text{Co}^{2+}$ , so at anode Al will oxidise and at cathode  $\text{Co}^{3+}$  will reduce.

$$\begin{aligned} E_{\text{cell}}^{\circ} &= (E_{\text{Cathode}}^{\circ})_{\text{RP}} - (E_{\text{Anode}}^{\circ})_{\text{RP}} \\ &= E_{\text{Co}^{3+}/\text{Co}^{2+}}^{\circ} - E_{\text{Al}^{3+}/\text{Al}}^{\circ} \\ &= (1.81) - (-1.66) \\ &= +3.47 \text{ V} \end{aligned}$$

**S9.** Ans. (c)



$$E_{\text{cell}}^{\circ} = 1.1 \text{ V}$$

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

$$\therefore n = 2$$

$$\Delta G^{\circ} = -2 \times 96487 \times 1.1$$

$$\Delta G^{\circ} = -212271.4 \text{ J mol}^{-1}$$

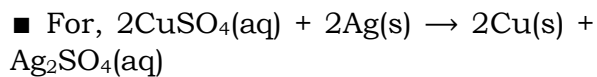
$$\Delta G^{\circ} = -212.27 \text{ kJ mol}^{-1}$$

**S10.** Ans. (d)

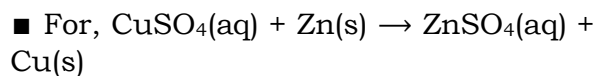
For a reaction to be spontaneous,  $E_{\text{cell}}^{\circ}$  must be positive.

■ For,  $\text{FeSO}_4(\text{aq}) + \text{Zn}(\text{s}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{Fe}(\text{s})$

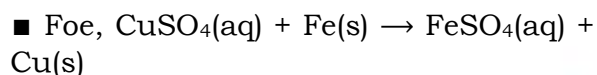
$$\begin{aligned} E_{\text{cell}}^{\circ} &= E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \\ &= -0.44 \text{ V} - (-0.76 \text{ V}) = 0.32 \text{ V} \end{aligned}$$



$$E_{\text{cell}}^{\circ} = 0.34 \text{ V} - 0.80 \text{ V} \\ = -0.46 \text{ V}$$

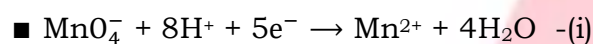


$$E_{\text{cell}}^{\circ} = 0.34 \text{ V} - (-0.76 \text{ V}) \\ = 1.1 \text{ V}$$

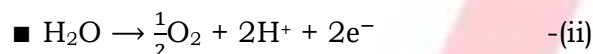


$$E_{\text{cell}}^{\circ} = 0.80 \text{ V} - (-0.44 \text{ V}) \\ = 1.24 \text{ V}$$

**S11.** Ans. (a)

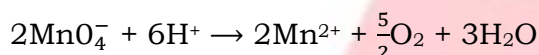


$$E_{\text{MnO}_4^-/\text{Mn}^{2+}}^{\circ} = -E_{\text{Mn}^{2+}/\text{MnO}_4^-}^{\circ} = 1.51 \text{ V}$$



$$E_{\text{O}_2/\text{H}_2\text{O}}^{\circ} = 1.223 \text{ V}$$

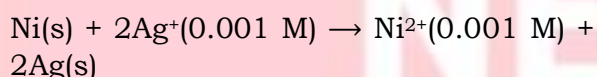
Using  $2 \times$  (i) +  $5 \times$  (ii), net cell reactions is



$$E_{\text{cell}}^{\circ} = E_{\text{C}}^{\circ} - E_{\text{A}}^{\circ} = E_{\text{MnO}_4^-/\text{Mn}^{2+}}^{\circ} - E_{\text{O}_2/\text{H}_2\text{O}}^{\circ} \\ = 1.51 - 1.223 = 0.287 \text{ V}$$

Since  $E_{\text{cell}}^{\circ} > 0$ , therefore net cell reaction is spontaneous and so  $\text{MnO}_4^-$  liberate  $\text{O}_2$  from  $\text{H}_2\text{O}$  in presence of an acid.

**S12.** Ans. (c)



$$E_{\text{cell}}^{\circ} = 1.05 \text{ V}$$

$$E_{\text{cell}}^{\circ} = E_{\text{cell}}^{\circ} - \frac{0.059}{n} \log \frac{[\text{Ni}^{2+}]}{[\text{Ag}^+]^2}$$

$$\Rightarrow 1.05 - \frac{0.059}{2} \log \frac{(10^{-3})}{(10^{-3})^2}$$

$$\Rightarrow 1.05 - \frac{0.059}{2} \log(10)^3$$

$$\Rightarrow 1.05 - 0.0295 \times 3$$

$$\Rightarrow 1.05 - 0.0885$$

$$\Rightarrow 0.9615 \text{ V}$$

**S13.** Ans.(a)

$$\Lambda_{\text{NaCl}} = \Lambda_{\text{Na}^+} + \Lambda_{\text{Cl}^-}$$

$$\Lambda_{\text{HCl}} = \Lambda_{\text{H}^+} + \Lambda_{\text{Cl}^-}$$

$$\Lambda_{\text{CH}_3\text{COONa}} = \Lambda_{\text{Na}^+} + \Lambda_{\text{CH}_3\text{COO}^-}$$

$$\text{Let, } \Lambda_{\text{Na}^+} = x, \Lambda_{\text{Cl}^-} = y, \Lambda_{\text{H}^+} = \Lambda_{\text{CH}_3\text{COO}^-} = w$$

Given,

$$x + y = 126.45 \quad \dots(\text{i})$$

$$y + z = 426.16 \quad \dots(\text{ii})$$

$$x + w = 91 \quad \dots(\text{iii})$$

From the above 3 equations, value of

$$z + w = 390.71 \text{ d}$$

**S14.** Ans.(b)

$$\Lambda_{\text{M}(\text{CH}_3\text{COOH})}^{\circ} = \Lambda_{\text{M}(\text{H}^+)}^{\circ} + \Lambda_{\text{M}(\text{CH}_3\text{COO}^-)}^{\circ} \\ = 350 + 50 = 400 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\alpha = \frac{\Lambda_{\text{M}}^{\text{c}}}{\Lambda_{\text{M}}^{\circ}}$$

$$\alpha = \frac{20}{400} = 5 \times 10^{-2}$$

$$K_{\text{a}(\text{CH}_3\text{COOH})} = C\alpha^2$$

$$= 0.007 \times (5 \times 10^{-2})^2 = 1.75 \times 10^{-5} \text{ mol L}^{-1}$$

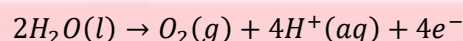
**S15.** Ans.(a)

The following reactions take place on electrolysis of dilute sulphuric acid on using pt electrodes are:

At cathode:



At anode:



Thus, the product obtained at anode is  $\text{O}_2$ .

**S16.** Ans.(d)

1 equivalent of any substance is deposited by 1 F of charge.

We have, 20 g calcium

The balance reaction will

The charge on Ca in  $\text{CaCl}_2$

Cl has -1 charge so that

$$\text{Ca} + 2(-1) = 0$$

$$\text{Ca} = 2$$

We have to get Ca from  $\text{Ca}^{2+}$

Number of required moles = mass/molar mass

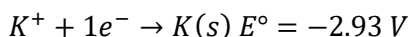
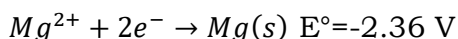
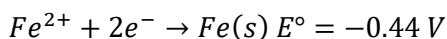
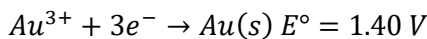
Molar mass of Ca is 40 g/mol and required of Ca is 20 g

Hence number of moles =  $20/40$   
= 0.5 mol

Electricity required to produce 1 mol of calcium =  $2F$

The electricity required to produce 0.5 mol of calcium =  $0.5 \times 2F$   
=  $1F$

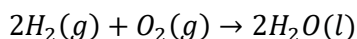
**S17.** Ans.(b)



$Au^{3+}$  occupies the top position in the electrochemical series.

**S18.** Ans.(a)

Cell reaction involved in hydrogen-oxygen fuel cell is



Thus,  $R = H_2(g), O_2(g); P = H_2O(l)$

**S19.** Ans.(c)

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{n} \log Q \quad \dots(i)$$

At equilibrium,  $Q = K_{eq}$  and  $E_{cell} = 0$

$$0 = E^\circ_{cell} - \frac{0.059}{1} \log K_{eq} \quad (\text{from equation (i)})$$

$$\log K_{eq} = \frac{E^\circ_{cell}}{0.059} = \frac{0.59}{0.059} = 10$$

$$K_{eq} = 10^{10} = 1 \times 10^{10}$$

**S20.** Ans.(a)

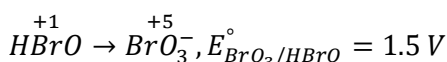
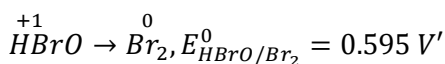
$$\Delta G^\circ = -nF E^\circ_{cell}$$

$$= -2 \times 96500 \times 0.24$$

$$= -46320 \text{ J mol}^{-1} = \frac{-46320}{1000}$$

$$= -46.32 \text{ kJ/mol}$$

**S21.** Ans.(c)



$E^\circ_{cell}$  for the disproportionation of  $HBrO$ ,

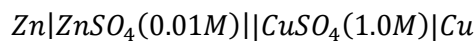
$$E^\circ_{cell} = E^\circ_{HBrO/Br_2} - E^\circ_{BrO_3^-/HBrO}$$

$$= 1.595 - 1.5$$

$$= 0.095 V = +ve$$

Hence, option (c) is correct answer.

**S22.** Ans.(d)



$$\therefore E_1 = E^\circ_{cell} - \frac{2.303RT}{2 \times F} \times \log \frac{(0.01)}{1}$$

When concentrations are changed

$$\therefore E_2 = E^\circ_{cell} - \frac{2.303RT}{2F} \times \log \frac{1}{0.01}$$

i.e.,  $E_1 > E_2$

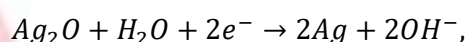
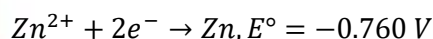
**S23.** Ans.(c)

$$\Lambda_m^\circ(AgCl) = \Lambda_m^\circ(AgNO_3) + \Lambda_m^\circ(KCl) - \Lambda_m^\circ(KNO_3)$$

$$= (133.4 + 149 - 144.9) S \text{ cm}^2 \text{ mol}^{-1}$$

$$= 138.4 S \text{ cm}^2 \text{ mol}^{-1}$$

**S24.** Ans.(c)

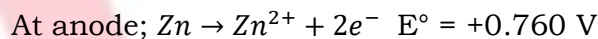


$$E^\circ = 0.344 V$$

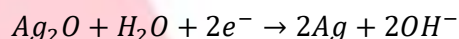
Both are reducing potential :

$$\text{As, } E^\circ_{Ag_2O/2Ag} > E^\circ_{Zn^{2+}/Zn}$$

$\therefore$  Cell reaction will be

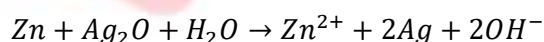


At cathode;



$$E^\circ = +0.344 V$$

Cell reaction:



$$n = 2 \quad E^\circ_{cell} = 1.104 V$$

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

$$\Delta G^\circ = -2 \times 96500 \times 1.10$$

$$= -213072 \text{ J mol}^{-1}$$

$$= -213.072 \text{ kJ mol}^{-1}$$

$$|\Delta G^\circ| = 213.072 \text{ kJ mol}^{-1}$$

**S25.** Ans.(d)

According to Faraday's first law:

$$w = z.i.t$$

$$z = \frac{E}{96500} (\text{molar mass})$$

$$0.1 \times 71 = \frac{35.5}{96500} \times 3 \times t \quad \text{as n-factor is 2}$$

$$t = 6433 \text{ sec} = 107.2 \text{ min}$$

$$\approx 110 \text{ min}$$

**S26.** Ans.(b)

Zn have a higher (-ve) electrode potential that is more reactive than Fe. It is coated on iron substances to provide resistance against rusting such a process is called galvanization. But in reverse, that is Fe cannot be coated on Zn, as corrosion will occur. In above, Zn displaces Fe from its salt solution.

**S27.** Ans.(a)

According to Faraday's law

$$Q = ne$$

$$Q = it$$

$$ne = it$$

$$n = \frac{1 \times 60}{1.6 \times 10^{-19}} = 3.75 \times 10^{20} \text{ electrons}$$

**S28.** Ans.(c)

$$\Delta G^\circ = -nFE^\circ \text{ cell} \quad E^\circ \text{ cell} = (-ve)$$

$$\text{So, } \Delta G^\circ = (+ve) \quad \Delta G > 0$$

$$\text{Also, } \Delta G^\circ = -2.303 RT \log K_{eq}$$

$$\therefore K_{eq} < 1$$

**S29.** Ans.(d)

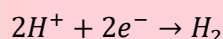
Concentration =  $0.5 \text{ mol dm}^{-3}$ ,

$$K = 5.76 \times 10^{-3} \text{ S cm}^{-1}$$

$$T = 298 \text{ K}$$

$$\lambda_m = \frac{K \times 1000}{M}$$
$$= \frac{5.76 \times 10^{-3}}{0.5} = 11.52 \text{ S cm}^2 \text{ mol}^{-1}$$

**S30.** Ans.(b)



$$E = E^\circ - \frac{0.059}{2} \log \frac{P_{H_2}}{(H^+)^2}$$

$$0 = 0 - \frac{0.059}{2} \log \frac{P_{H_2}}{(10^{-7})^2}$$

$$\log 1 = 0$$

$$P_{H_2} = (10^{-7})^2 = 10^{-14} \text{ atm} .$$

**S31.** Ans.(d)

**S32.** Ans.(b)

$\text{Sn}^{2+}$  cannot reduce  $\text{Fe}^{2+}$ , so  $\text{FeCl}_2$  and  $\text{SnCl}_2$  can exist together.

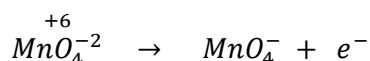
**S33.** Ans.(a)

$$\Delta G^\circ = -2.303 RT \log K_{sp}$$

$$63300 = -2.303 \times 8.314 \times 298 \log K_{sp}$$

$$K_{sp} \sim 8 \times 10^{-12}$$

**S34.** Ans.(c)



$$0.1 \text{ mole} \quad 0.1 \text{ mole}$$

$$Q = nF = 0.1 \times 96500 = 9650 \text{ C}$$

**S35.** Ans.(c)

$$W_{O_2} = \frac{5600}{22400} \times 32 = 8 \text{ g} = 1 \text{ equivalent}$$

$$= 1 \text{ equivalent of Ag}$$

$$= 108 \text{ g}$$

**S36.** Ans.(c)

According to Kohlrausch's law of limiting molar conductivity:

$$\alpha = \frac{\Lambda_m}{\Lambda_m^\infty} = \frac{9.54}{238} = 0.04008$$

$$\% \text{ Dissociation } (\alpha) = 4.008\%$$

**S37.** Ans.(b)

$$\text{pH} = 10$$

Oxidation potential of standard hydrogen electrode is given by:

$$E_{op} = 0.0592 (\text{pH})_{\text{Anode}}$$

$$= 0.0592 \times 10$$

$$E_{op} = 0.59 \text{ V}$$

**S38.** Ans.(a)

$$E^\circ \text{ cell} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$= 0.76 - (-0.34)$$

$$= 1.1 \text{ V}$$