CBSE Class 12 physics

Important Questions

Chapter 3

Electrochemistry

3 Marks Questions

1.What is the cell potential for the cell at 25^0C $Cr/Cr^{3+}10.1\ m]//Fe^{2+}(0.01m)/Fe$

$E^0{}_{cr+/cr}=-0.74V$; $E^0Fe^{2+}/Fe=-0.44V$.

Ans. The cell reaction is

$2Cr+3Fe^{2+}6e^{-}\rightarrow 2Cr^{3+}+3Fe$

Nernst Equation –

$$E_{cell}=\left(E^0{}_{Fe}{}^{2+}/{}_{Fe}-E^0{}_{cr}{}^{3+}/{}_{cr}\right)-\frac{0.059}{6}\log\frac{\left[Cr^{3+}\right]^2}{\left[Fe^{2+}\right]^3}$$

=(-0.44v – (-0.74v) - $\frac{0.059}{6}\log\frac{(0.10)^2}{(0.01)^3}$

= $0.3V-\frac{0.059}{6}\log 10^4$

= 0.3V – 0.0394V

= +0.2606 V

2.Calculate ΔG^0 **for the reaction at** 25^0C
$Zn(s)1Zn^2+[0.0004m]11cd^2+(0.2m)1cd(s)$ $E^0{}_{Zn}{}^{2+}{}_{/Zn}=-0.763V$,
$E^0{}_{cd}{}^{2+}/cd=-0.403v,$ $F=96500\ C\,Mol^{-1}$**, R = 8.314 J/K.**

Ans. The half cell reactions are

Anode: $Zn(s) \rightarrow Zn^{2} + (aq) + 2e^{-}$

Cathode: $Cd^{2} + (aq) + 2e^{-} \rightarrow Cd(s)$

Nernst Equation

$$E_{cell} = \left(E^{0}{}_{Cathode} - E^{0}{}_{anode}\right) - \frac{0.059}{n} \log \frac{\left[Zn^{2+}\right]}{\left[Cd^{2+}\right]}$$

= (- 0.403 – (-0.763) - $\frac{0.059}{2} \log \frac{0.0004}{0.2}$

= 0.36V - 0.0798V = 0.4398V

$\Delta G^{0} = -\text{n F E}^{0}{}_{cell}$

$= \frac{-2mol \times 96500\ C}{mol \times 0.4398V}$

= -8488 J mol^{-1}

3.Calculate Equilibrium constant K for the reaction at
$298K\ Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+} / aq) + Cu\ E^{0}{}_{Zn}{}^{2+} 1Zn = -0.076v; E^{0}Cu^{2} + / Cu$ =
+0.34v.

Ans.From the reaction, n =2

$E^{0}{}_{cell} = E^{0}cu^{2} + / cu - E^{0}Zn^{2} + / Zn$

= + 0.34v - (-0.76v) = 1.10V

$E^{0}{}_{cell} = \frac{2.303RT}{nF} \log k_{c}$

At 298k , $E^{0}{}_{cell} \times \frac{n}{0.059} \log k_{c}$

Log k_c = $E^0_{cell} \times \frac{n}{0.059}$

= $1.10 \times \frac{2}{0.059} = 37.29$

$K_c = Antilog\ 37.29$

$= 1.95 \times 10^{37}$

4. For what concentration of $Ag+(aq)$ will the emf of the given cell be zero at 25^0C if the concentration of $CU^{2+}(aq)$ is 0.1 M ?

$Cu(s)/Cu^{2+}(0.1M)//Ag^+(aq)/Ag(s)\ E^0Ag^+/Ag = +0.80V;$
$E^0{}_{Cu}{}^{2+/Cu} = 0.34\ V$

Ans. $\left[Ag^+\right] = 5.3 \times 10^{-9} M$

5.Calculate the standard free energy change for the cell- reaction.

$Fe^{2+}(aq) + Ag^+(s)\dot{a} \rightarrow Fe^3 + (aq) + Ag(s)$ **How is it related to the equilibrium constant of the reaction?** $E^0{}_{Fe}{}^{3+}/Fe^{2+} = +0.77V$, $E^0{}_{Ag}{}^{+1/Ag} = +0.08V\ F$ $= 96500\ C/mol.$

Ans. $E^0{}_{cell} =\ 0.03V$

$\Delta G^0 = \text{-}2895\text{J}$

6.How much charge is required for the following reductions:

(i) 1 mol of Al^{3+} **to Al.**

(ii) 1 mol of Cu^{2+} **to Cu.**

(iii) 1 mol of MnO_4^- **to** Mn^{2+} **.**

Ans.(i) $Al^{3+} + 3e^- \rightarrow Al$

Therefore, Required charge = 3 F

$= 3 \times 96487\ C$

= 289461 C

(ii) $Cu^{2+} + 2e^{-} \rightarrow Cu$

Therefore, Required charge = 2 F

$= 2 \times 96487\ C$

= 192974 C

(iii) $MnO_4^{-} \rightarrow Mn^{2+}$

i.e., $Mn^{7+} + 5e^{-} \rightarrow Mn^{2+}$

Therefore, Required charge = 5 F

$= 5 \times 96487\ C$

= 482435 C

7. How much electricity in terms of Faraday is required to produce

(i) 20.0 g of Ca from molten $CaCl_2$.

(ii) 40.0 g of Al from molten Al_2O_3.

Ans.(i) According to the question,

$Ca^{2+} + 2e^{-1} \rightarrow \underset{40g}{Ca}$

Electricity required to produce 40 g of calcium = 2 F

Therefore, electricity required to produce 20 g of calcium = $\frac{2 \times 20}{40} F$

= 1 F

(ii) According to the question,

$$Al^{3+} + 3e^{-1} \rightarrow \underset{27g}{Al}$$

Electricity required to produce 27 g of Al = 3 F

Therefore, electricity required to produce 40 g of Al = $\frac{3 \times 40}{27}F$

= 4.44 F

8. How much electricity is required in coulomb for the oxidation of

(i) 1 mol of H_2O to O_2.

(ii) 1 mol of FeO to Fe_2O_3.

Ans.(i) According to the question,

$$H_2O \rightarrow H_2 + \frac{1}{2}O_2$$

Now, we can write:

$$O^{2} \rightarrow \frac{1}{2}O_2 + 2e^{-}$$

Electricity required for the oxidation of 1 mol of H_2O to O_2 = 2 F

= 2×96487 C

= 192974 C

(ii) According to the question,

$$Fe^{2+} \rightarrow Fe^{3+} + e^{-1}$$

Electricity required for the oxidation of 1 mol of FeO to Fe_2O_3 = 1 F

= 96487 C

9. A solution of $Ni(NO_3)_2$ is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

Ans. Given,

Current = 5A

Time = 20×60 = 1200 s

Therefore, $\text{Charge} = \text{current} \times \text{time}$

$= 5 \times 1200$

= 6000 C

According to the reaction,

$$Ni^{2+}_{(aq)} + 2e^- \rightarrow \underset{58.7g}{Ni_{(s)}}$$

Nickel deposited by $2 \times 96487\ C$ = 58.71 g

Therefore, nickel deposited by 6000 C $= \dfrac{58.71 \times 6000}{2 \times 96487} g$

= 1.825 g

Hence, 1.825 g of nickel will be deposited at the cathode.

10. Depict the galvanic cell in which the reaction takes place. Further show:

(i) Which of the electrode is negatively charged?

(ii) The carriers of the current in the cell.

(iii) Individual reaction at each electrode.

Ans. The galvanic cell in which the given reaction takes place is depicted as:

$$Zn_{(s)} \mid Zn^{2+}_{(aq)} \parallel Ag^{+}_{(aq)} \mid Ag_{(s)}$$

(i) Zn electrode (anode) is negatively charged.

(ii) Ions are carriers of current in the cell and in the external circuit, current will flow from silver to zinc.

(iii) The reaction taking place at the anode is given by,

$$Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2e^-$$

The reaction taking place at the cathode is given by,

$$Ag^+{}_{(aq)} + e^- \rightarrow Ag_{(s)}$$

11. Write the chemistry of recharging the lead storage battery, highlighting all the materials that are involved during recharging.

Ans.A lead storage battery consists of a lead anode, a grid of lead packed with lead oxide (PbO_2) as the cathode, and a 38% solution of sulphuric acid (H_2SO_4) as an electrolyte.

When the battery is in use, the following cell reactions take place:

At anode: $Pb_{(s)} + SO_4^{2-}{}_{(aq)} \rightarrow PbSO_{4(s)} + 2e^-$

At cathode: $PbSO_{4(s)} + SO_4^{2-}{}_{(aq)} + 4H^-{}_{(aq)} + 2e^- \rightarrow PbSO_{4(s)} + 2H_2O_{(l)}$

The overall cell reaction is given by,

$$Pb_{(s)} + PbO_{2(s)} + 2H_2SO_{4(aq)} \rightarrow 2PbSO_{4(s)} + 2H_2O_{(l)}$$

When a battery is charged, the reverse of all these reactions takes place.

Hence, on charging, $PbSO_{4(s)}$ present at the anode and cathode is converted into $Pb_{(s)}$ and $PbO_{2(s)}$ respectively.

12. If a current of 0.5 ampere flows through a metallic wire for 2 hours, then how many electrons would flow through the wire?

Ans. I = 0.5 A

t = 2 hours = $2 \times 60 \times 60\ \text{s} = 7200\ \text{s}$

Thus, $Q = It$

= $0.5\ \text{A} \times 7200\ \text{s}$ = 3600 C

We know that $96487\ \text{C} = 6.023 \times 10^{23}$ number of electrons.

Then, $3600\ \text{C} = \dfrac{6.023 \times 10^{23} \times 3600}{96487}$ number of electrons

$= 2.25 \times 10^{22}$ number of electrons

Hence, 2.25×10^{22} number of electrons will flow through the wire.

13. Calculate the potential of hydrogen electrode in contact with a solution whose pH is 10.

Ans. For hydrogen electrode, $H^+ + e^- \rightarrow \frac{1}{2}H_2$, it is given that pH = 10

Therefore, $[H^+]$ = 10 - 10M

Now, using Nernst equation:

$$H_{\left(H^+/\frac{1}{2}H^2\right)} = E^{\circ}_{\left(H^+/\frac{1}{2}H^2\right)} - \frac{RT}{nF}\ln\frac{1}{\left[H^+\right]}$$

$$= E^{\circ}_{\left(H^+/\frac{1}{2}H^2\right)} - \frac{0.0591}{1}\log\frac{1}{\left[H^+\right]}$$

$$0 - \frac{0.0591}{1}\log\frac{1}{\left[10^{-10}\right]}$$

= - 0.0591 log 10^{10} = - 0.591 V

14. Calculate the emf of the cell in which the following reaction takes place:

$$Ni_{(s)} + 2Ag^{+}(0.002M) \rightarrow Ni^{2+}(0.160M) + 2Ag_{(s)}$$

Given that $E^{\ominus}_{(cell)} = 1.05V$

Ans. Applying Nernst equation we have:

$$E_{(cell)} = E^{\ominus}_{(cell)} - \frac{0.0591}{n}\log\frac{[Ni^{2+}]}{[Ag^{2}]^{+}}$$

$$= 1.05 - \frac{0.0591}{2}\log\frac{(0.160)}{(0.002)^{2}}$$

$$= 1.05 - 0.02955\log\frac{0.16}{0.000004}$$

= 1.05 - 0.02955 log 4 × 104

= 1.05 - 0.02955 (log 10000 + log 4)

= 1.05 - 0.02955 (4 + 0.6021)

= 0.914 V

15. The cell in which the following reactions occurs:

$2Fe^{3+}_{(aq)} + 2I^{-}_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + I_{2(s)}$ **has** $E^{\ominus}_{cell}$ **= 0.236 V at 298 K. Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction.**

Ans. Here, *n* = 2, $E^{\ominus}_{cell} = 0.236V$ T = 298 K

We know that:

$$\Delta_r G^{\ominus} = nFE^{\ominus}_{cell}$$

= - 2 × 96487 × 0.236

= - 45541.864 J mol - 1

= - 45.54 kJ mol - 1

Again, $\Delta_r G^{\circ} = -2.303RT\log Kc$

$$\log K_c = -\frac{\Delta_r G^{\circ}}{2.303RT}$$

$$= -\frac{-45.54 \times 10^3}{2.303 \times 8.314 \times 298}$$

= 7.981

∴*K*c= Antilog (7.981) = 9.57 × 107

16. How would you determine the standard electrode potential of the systemMg2+| Mg?

Ans. The standard electrode potential of Mg^{2+} | Mg can be measured with respect to the standard hydrogen electrode, represented by Pt(*s*), $H_{2(g)}$ (1 atm) | H^+(*aq*) (1M).

A cell, consisting of Mg | $MgSO_4$(*aq* 1 M) as the anode and the standard hydrogen electrode as the cathode, is set up.

$$Mg \mid Mg^{2+}(aq, 1M) \parallel H^+(aq, 1M) \mid H_2(g, 1bar), Pt_{(s)}$$

Then, the emf of the cell is measured and this measured emf is the standard electrode potential of the magnesium electrode.

$$E^{\circ} = E_R^{\circ} - E_L^{\circ}$$

Here, E_R° for the standard hydrogen electrode is zero.

Therefore, $E^{\circ} = 0 - E_L^{\circ}$

$= -E_L^{\circ}$