3.1. How would you determine the standard electrode potential of the system Mg^{2+1} Mg?

Ans: A cell will be set up consisting of Mg/$MgSO_4$ (1 M) as one electrode and standard hydrogen electrode Pt, H, (1 atm)H^+/(I M) as second electrode, measure the EMF of the cell and also note the direction of deflection in the voltmeter. The direction of deflection shows that e^{-1} s flow from mg electrode to hydrogen electrode, i.e., oxidation takes place on magnesium electrode and reduction on hydrogen electrode. Hence, the cell may be represented as follows :

$$Mg \mid Mg^{2+}(1\ M) \parallel H^+(1\ M) \mid H_2, (1\ atm), Pt$$

$$E^o_{cell} = E^o_{H^+/\frac{1}{2}H_2} - E^o_{Mg^{2+}/Mg}$$

Put $\quad E^o_{H^+/\frac{1}{2}H_2} = 0$

$$\therefore \quad E^o_{Mg^{2+}/Mg} = -E^o_{cell}$$

3.2.Can you store copper sulphate solutions in a zinc pot?

Ans:

Zn being more reactive than Cu, displaces Cu from $CuSO_4$ solution as follows:

Zn (s) + $CuSO_4$ (aq) → $ZnSO_4$(ag) + Cu (s)

In terms of EMF, we have

$$Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$$

$$\begin{aligned} E^o_{cell} &= E^o_{Cu^{2+}/Cu} - E^o_{Zn^{2+}/Zn} \\ &= 0.34\ V - (-0.76\ V) \\ &= 1.10\ V \end{aligned}$$

As E^o_{cell} is positive, reaction takes place, i.e., Zn reacts with copper and hence, we cannot store $CuSO_4$ solution in zinc pot.

3.3. Consultthe table of standard electrode potentials and suggest three substances that can oxidise ferrous ions under suitable conditions.

Ans:

Oxidation of Fe^{2+} converts it to Fe^{3+}, i.e., $Fe^{2+} \rightarrow Fe^{3+} + e^-$; $E°_{ox}$= -0.77 V Only those substances can oxidise Fe^{2+} to Fe^{3+} which are stronger oxidizing agents and have positive reduction potentials greater than 0.77 V, so that EMF of the cell reaction is positive. This is so for elements lying below Fe^{3+}/Fe^{2+} in the series ex: Br_2, Cl_2 and F_2.

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

Ans. For hydrogen electrode, $H^+ + e^- \rightarrow 1/2\ H_2$,

Applying Nernst equation,

$$E_{H^+, \frac{1}{2}H_2} = E^o_{H^+, \frac{1}{2}H_2} - \frac{0.0591}{n}\log\frac{1}{[H^+]}$$

$$= 0 - \frac{0.0591}{1}\log\frac{1}{10^{-10}}$$

$$\left\{\begin{array}{l} pH = 10 \\ \Rightarrow [H^+] = 10^{-10} M \end{array}\right\}$$

$$= -0.0591 \times 10$$

$$= -0.591\ V$$

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

$Ni(s) + 2Ag^+\ (0.002\ M) \rightarrow Ni^{2+}\ (0.160\ M) + 2Ag(s)$ Given that $E^{(-)}_{(cell)} = 1.05\ V$.

Ans:

Applying Nernst equation,

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

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

$$= 1.05 - \frac{0.0591}{2}\log(4 \times 10^4)$$

$$= 1.05 - \frac{0.0591}{2}(4.6021)$$

$$= 1.05 - 0.14\ V$$

$$= 0.91\ V$$

3.6. The cell in which the following reaction occurs: $2Fe^{3+}\ (aq) + 2I^-\ (aq) \rightarrow 2Fe^{2+}\ (aq) + I^2\ (s)$ has $E^\circ_{cell} = 0.236\ V$ at 298 K. Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction.

Ans:

Applying Nernst equation,

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

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

$$= 1.05 - \frac{0.0591}{2}\log(4\times10^4)$$

$$= 1.05 - \frac{0.0591}{2}(4.6021)$$

$$= 1.05 - 0.14\,V$$

$$= 0.91\,V$$

3.7. Why does the conductivity of a solution decrease with dilution?
Ans: Conductivity of a solution is the conductance of ions present in a unit volume of the solutions. On dilution, no. of ions per unit volume decreases. Hence, the conductivity decreases.

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