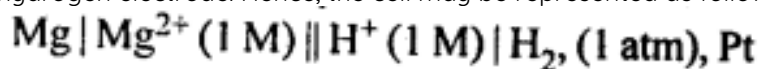




3.1. How would you determine the standard electrode potential of the system $\text{Mg}^{2+}|\text{Mg}$?

Ans: A cell will be set up consisting of Mg/MgSO_4 (1 M) as one electrode and standard hydrogen electrode Pt, H_2 , (1 atm) $|\text{H}^+$ (1 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^- 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 :



$$E^\circ_{\text{cell}} = E^\circ_{\text{H}^+ / \frac{1}{2}\text{H}_2} - E^\circ_{\text{Mg}^{2+} / \text{Mg}}$$

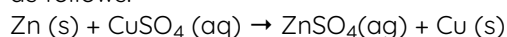
$$\text{Put } E^\circ_{\text{H}^+ / \frac{1}{2}\text{H}_2} = 0$$

$$\therefore E^\circ_{\text{Mg}^{2+} / \text{Mg}} = -E^\circ_{\text{cell}}$$

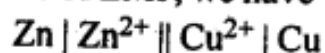
3.2. Can you store copper sulphate solutions in a zinc pot?

Ans:

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



In terms of EMF, we have



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

As E°_{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. Consult the 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., $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^-$; $E^\circ_{\text{ox}} = -$

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 $\text{Fe}^{3+}/\text{Fe}^{2+}$ 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,

$$\begin{aligned} E_{H^+, \frac{1}{2}H_2} &= E^\circ_{H^+, \frac{1}{2}H_2} - \frac{0.0591}{n} \log \frac{1}{[H^+]} \\ &= 0 - \frac{0.0591}{1} \log \frac{1}{10^{-10}} \\ &\quad \left\{ \begin{array}{l} \text{pH} = 10 \\ \Rightarrow [H^+] = 10^{-10} \text{ M} \end{array} \right\} \\ &= -0.0591 \times 10 \\ &= -0.591 \text{ V} \end{aligned}$$

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

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

Ans:

Applying Nernst equation,

$$\begin{aligned} E_{cell} &= E^\circ_{cell} - \frac{0.0591}{n} \log \frac{[Ni^{2+}]}{[Ag^+]^2} \\ &= 1.05 \text{ V} - \frac{0.0591}{2} \log \frac{0.160}{(0.002)^2} \\ &= 1.05 - \frac{0.0591}{2} \log(4 \times 10^4) \\ &= 1.05 - \frac{0.0591}{2} (4.6021) \\ &= 1.05 - 0.14 \text{ V} \\ &= 0.91 \text{ V} \end{aligned}$$

3.6. The cell in which the following reaction occurs: $2Fe^{3+} (aq) + 2I^- (aq) \rightarrow 2Fe^{2+} (aq) + I_2 (s)$ has $E^\circ_{cell} = 0.236 \text{ V}$ at 298 K. Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction.

Ans:

Applying Nernst equation,

$$\begin{aligned}E_{\text{cell}} &= E^{\circ}_{\text{cell}} - \frac{0.0591}{n} \log \frac{[\text{Ni}^{2+}]}{[\text{Ag}^{+}]^2} \\&= 1.05\text{V} - \frac{0.0591}{2} \log \frac{0.160}{(0.002)^2} \\&= 1.05 - \frac{0.0591}{2} \log(4 \times 10^4) \\&= 1.05 - \frac{0.0591}{2} (4.6021) \\&= 1.05 - 0.14\text{ V} \\&= 0.91\text{ V}\end{aligned}$$

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.

***** END *****