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Engineering Chemistry

Corrosion - Tutorial 1

1> The standard emf of the cell is 0.462 V. $Cu/Cu^{+2}(1M) \| Ag^{+}(1M)/Ag$
If the standard potential of Cu is 0.337 V, what is the standard potential of Ag electrode.

Solution

~~At~~ ~~anode~~ At anode : $Cu \longrightarrow Cu^{2+} + 2e^-$

At cathode : $2Ag^+ + 2e^- \longrightarrow 2Ag$

Net cell reaction : $Cu + 2Ag^+ \longrightarrow 2Ag + Cu^{2+}$

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

$0.462 = E^\circ_{Ag} - E^\circ_{Cu}$

$\therefore E^\circ_{Ag} = 0.462 + 0.337$

$\therefore E^\circ_{Ag} = 0.799\,V$

$\therefore$ The standard potential of Ag electrode is 0.799 V.

2> Calculate standard electrode potential of lead electrode when it is in contact with 0.0096 M Pb^{2+} solution at 301 K. The E value of lead is −0.18025 V.

Solution $[Pb^{2+}] = 0.0096M$

$E_{Pb^{2+}/Pb} = -0.18025\,V$

$$E_{Pb^{2+}/Pb} = E^\circ_{Pb^{2+}/Pb} - \frac{2.303RT}{nF} \log_{10} \frac{1}{[Pb^{2+}]}$$

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$$\therefore E^{\circ}_{Pb^{2+}/Pb} = E_{Pb^{2+}/Pb} - \frac{2.303RT}{nF} \log_{10}[Pb^{2+}]$$

$$= -0.18025 - \frac{2.303 \times 8.314 \times 301}{2 \times 96500} \times \log(0.0096)$$

$$= -0.18025 - 0.02986 \times (-2.0177)$$

$$= -0.1200\,V$$

$\therefore$ Standard electrode potential of lead electrode is $-0.1200\,V$.

3) For the cell reaction $Ni/Ni^{+2}(0.01M) || Cu^{2+}(0.5M)/Cu$. The Standard reduction potentials of Ni and Cu are -0.25 and $0.34\,V$ respectively. Write the electrode reactions and calculate the EMF of the cell at 298 K.

Solution

At anode : $Ni \longrightarrow Ni^{+2} + 2e^-$

At cathode : $Cu^{2+} + 2e^- \longrightarrow Cu$

Net cell reaction : $Ni + Cu^{2+} \longrightarrow Ni^{+2} + Cu$

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

$$= 0.34 - (-0.25)$$

$$\therefore E^{\circ}_{cell} = 0.59\,V$$

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Now,

$$E_{cell} = E^\circ_{cell} - \frac{2.303RT}{nF} \log \frac{[Ni^{2+}]}{[Cu^{2+}]}$$

$$= 0.59 - \frac{2.303 \times 8.314 \times 298}{2 \times 96500} \times \log \frac{0.01}{0.5}$$

$$= 0.59 - \frac{0.0592}{2} \times (-1.6989)$$

$$= 0.64029$$

$\therefore E_{cell} = \underline{\underline{0.6403V}}$

$\therefore$ The EMF of the cell is 0.6403 volts.