

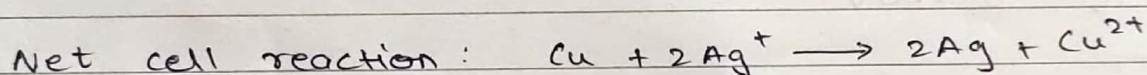
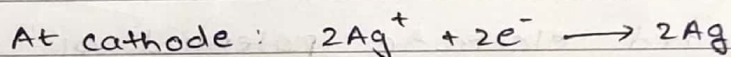
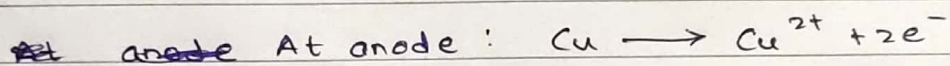
22/07/21

Engineering ChemistryDATE:

--	--	--

Corrosion - Tutorial 1

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

Solution

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

$$0.462 = E_{\text{Ag}}^{\circ} - E_{\text{Cu}}^{\circ}$$

$$\therefore E_{\text{Ag}}^{\circ} = 0.462 + 0.337$$

$$\therefore E_{\text{Ag}}^{\circ} = 0.799\text{V}$$

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

- 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 $[\text{Pb}^{2+}] = 0.0096\text{M}$

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

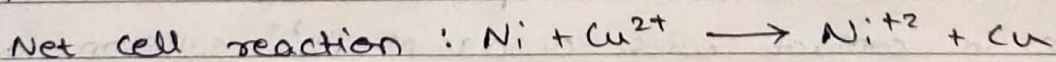
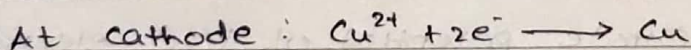
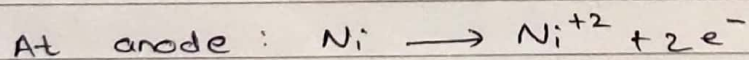
$$E_{\text{Pb}^{2+}/\text{Pb}} = E_{\text{Pb}^{2+}/\text{Pb}}^{\circ} - \frac{2.303RT}{nF} \log_{10} \frac{1}{[\text{Pb}^{2+}]}$$

$$\begin{aligned}
 \therefore E_{\text{Pb}^{2+}/\text{Pb}}^{\circ} &= E_{\text{Pb}^{2+}/\text{Pb}} - \frac{2.303RT}{nF} \log_{10} [\text{Pb}^{2+}] \\
 &= -0.18025 - \frac{2.303 \times 8.314 \times 301}{2 \times 96500} \times \log (0.0096) \\
 &= -0.18025 - 0.02986 \times (-2.0177) \\
 &= -0.1200 \text{ V}
 \end{aligned}$$

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

- 3) For the cell reaction $\text{Ni}/\text{Ni}^{2+}(0.01\text{M}) \parallel \text{Cu}^{2+}(0.5\text{M})/\text{Cu}$. 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



$$\begin{aligned}
 E_{\text{cell}}^{\circ} &= E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \\
 &= 0.34 - (-0.25)
 \end{aligned}$$

$$\therefore E_{\text{cell}}^{\circ} = 0.59 \text{ V}$$

NOW,

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

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

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

$$= 0.64029$$

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

\therefore The EMF of the cell is 0.6403 volts.