Marks
5

- Consider the following reaction at 298 K.

$$Ni^{2+}(aq) + Zn(s) \rightleftharpoons Ni(s) + Zn^{2+}(aq)$$

Calculate ΔG^{o} for the cell. (Relevant electrode potentials can be found on the data page.)

The half-cell reduction reactions and potentials are:

$$\mathbf{Ni^{2+}(aq) + 2e^{-} \rightarrow Ni(s) \qquad E^{0} = -0.24\ V}$$
$$\mathbf{Zn^{2+}(aq) + 2e^{-} \rightarrow Zn(s) \qquad E^{0} = -0.76\ V}$$

In the reaction above, the Zn is undergoing oxidation so its potential is reversed and the overall cell potential is:

$$\mathbf{E^{0}_{cell} = (-0.24) - (-0.76) = +0.52\ V}$$

Using ΔG^{0} = -nFE^{0} for this two electron reaction:

$$\mathbf{\Delta G^{0} = -(2) \times (96485) \times (+0.52) = -100000\ J\ mol^{1} = -100\ kJ\ mol^{1}}$$

Answer: **-100 kJ mol^{-1}**

What is the value of the equilibrium constant for the reaction at 298 K?

Using $\mathbf{E^{0} = \dfrac{RT}{nF}\ln K}$,

$$\mathbf{+0.52 = \frac{(8.314)\times(298)}{(2)\times(96485)}\ln K \qquad so\ K = 3.89 \times 10^{17}}$$

Alternatively, using ΔG^{0} = -RTlnK,

$$\mathbf{-100 \times 10^{3} = -(8.314) \times (298) \times \ln K \qquad so\ K = 3.89 \times 10^{17}}$$

Answer: **3.89×10^{17}**

Express the overall reaction in voltaic cell notation.

In the reaction, Zn is being oxidized and hence is the anode. Ni^{2+} is being reduced and so Ni is the cathode. In the standard cell notation, the anode is written on the left and the cathode on the right:

$$\mathbf{Zn(s)\ |\ Zn^{2+}(aq)\ ||\ Ni^{2+}(aq)\ |\ Ni(s)}$$

ANSWER CONTINUES ON THE NEXT PAGE

- Using a current of 2.00 A, how long (in minutes) will it take to plate out all of the silver from 0.250 L of a 1.14×10^{-2} M $Ag^{+}(aq)$ solution? **2**

The number of moles of $Ag^{+}(aq)$ in a 0.250 L of a 1.14×10^{-2} M solution is,

number of moles = volume×concentration = $0.250 \times 1.14 \times 10^{-2} = 2.85 \times 10^{-3}$ mol

The reduction of $Ag^{+}(aq)$ is a one electron process, $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$, so this number of moles of electrons are required.

As the number of moles of electrons delivered by a current I in a time t is,

$$\textbf{number of moles of electrons} = \frac{It}{F} = \frac{2.00 \times t}{96485} = 2.85 \times 10^{-3}$$

t = 137 s = 2.29 minutes