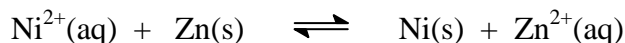


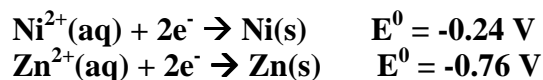
- Consider the following reaction at 298 K.



Calculate  $\Delta G^{\circ}$  for the cell. (Relevant electrode potentials can be found on the data page.)

**Marks**  
**5**

The half-cell reduction reactions and potentials are:



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

$$E_{\text{cell}}^{\circ} = (-0.24) - (-0.76) = +0.52 \text{ V}$$

Using  $\Delta G^{\circ} = -nFE^{\circ}$  for this two electron reaction:

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

Answer:  $-100 \text{ kJ mol}^{-1}$

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

$$\text{Using } E^{\circ} = \frac{RT}{nF} \ln K,$$

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

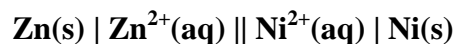
Alternatively, using  $\Delta G^{\circ} = -RT \ln K$ ,

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

Answer:  $3.89 \times 10^{17}$

Express the overall reaction in voltaic cell notation.

In the reaction, Zn is being oxidized and hence is the anode.  $\text{Ni}^{2+}$  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:



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  $1.14 \times 10^{-2}$  M  $\text{Ag}^+(\text{aq})$  solution?

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

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

**The reduction of  $\text{Ag}^+(\text{aq})$  is a one electron process,  $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$ , 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,**

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

$$t = 137 \text{ s} = 2.29 \text{ minutes}$$