Marks 5

- A galvanic cell is made of a Zn^{2+}/Zn half cell with $[Zn^{2+}]$ = 2.0 M and an Ag^{+}/Ag half cell with $[Ag^{+}]$ = 0.050 M. Calculate the electromotive force of the cell at 25 °C.

The standard reduction reactions and potentials for the two half cells are:

$Zn^{2+}(aq) + 2e^{-} \rightarrow Zn(s)$ $\quad E° = -0.76$ V

$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$ $\quad E° = +0.80$ V

The least positive (Zn^{2+}/Zn) couple is reversed giving the overall reaction:

$Zn(s) + 2Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$ $\quad E° = (+0.76\text{ V}) + (0.80) = 1.56$ V

As non-standard concentrations are used, the cell potential is calculated using the Nernst equation. The reaction involves the transfer of $2e^{-}$ so with n = 2 this becomes:

$$E = E° - \frac{RT}{nF}\ln Q = E° - \frac{RT}{nF}\ln\left(\frac{[Zn^{2+}(aq)]}{[Ag^{+}(aq)]^2}\right)$$

$$= (+1.56\text{ V}) - -\frac{(8.314\text{ J K}^{-1}\text{ mol}^{-1})(298\text{ K})}{(2 \times 96485\text{ C mol}^{-1})}\ln\left(\frac{2.0}{0.050^2}\right) = +1.47\text{ V}$$

Answer: **+1.47 V**

Calculate the equilibrium constant of the reaction at 25 °C.

The equilibrium constant is related to the standard cell potential:

$$E° = \frac{RT}{nF}\ln K$$

Hence,

$$\ln K = E° \times \frac{nF}{RT} = (1.56\text{ V}) \times \frac{(2 \times 96485\text{ C mol}^{-1})}{(8.314\text{ J K}^{-1}\text{ mol}^{-1})(298\text{ K})} = 121.5$$

$K = 5.9 \times 10^{52}$

Answer: $K = 5.9 \times 10^{52}$

Calculate the standard Gibbs free energy of the reaction at 25 °C.

Using $\Delta G° = nFE°$:

$\Delta G° = -(2 \times 96485\text{ C mol}^{-1}) \times (+1.56\text{ V}) = –301\text{ kJ mol}^{-1}$

Answer: $–301\text{ kJ mol}^{-1}$

ANSWER CONTINUES ON THE NEXT PAGE

Indicate whether the reaction is spontaneous or not. Give a reason for your answer.

As $E > 0$, $\Delta G° < 0$ and K is very large: the reaction is spontaneous.

Express the overall reaction in the shorthand voltaic cell notation.

$Zn(s) | Zn^{2+}(aq)$ (2.0 M) || $Ag^{+}(aq)$ (0.050 M) | $Ag(s)$