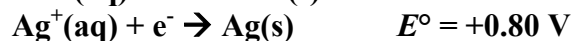
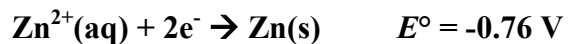
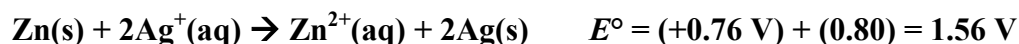


- A galvanic cell is made of a  $\text{Zn}^{2+}/\text{Zn}$  half cell with  $[\text{Zn}^{2+}] = 2.0 \text{ M}$  and an  $\text{Ag}^+/\text{Ag}$  half cell with  $[\text{Ag}^+] = 0.050 \text{ M}$ . Calculate the electromotive force of the cell at  $25^\circ\text{C}$ .

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



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



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

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

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

Answer: +1.47 V

Calculate the equilibrium constant of the reaction at  $25^\circ\text{C}$ .

The equilibrium constant is related to the standard cell potential:

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

Hence,

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

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

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

Calculate the standard Gibbs free energy of the reaction at  $25^\circ\text{C}$ .

Using  $\Delta G^\circ = nFE^\circ$ :

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

Answer:  $-301 \text{ kJ mol}^{-1}$

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

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

Express the overall reaction in the shorthand voltaic cell notation.

**$\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) (2.0 \text{ M}) \parallel \text{Ag}^+(\text{aq}) (0.050 \text{ M}) \mid \text{Ag(s)}$**