Consider the following half-reactions and their standard reduction potentials.

\[2\text{ClO}_3^- + 12\text{H}^+ + 10e^- \rightarrow \text{Cl}_2 + 6\text{H}_2\text{O} \quad E^\circ = 1.47 \text{ V}\]
\[\text{S}_2\text{O}_8^{2-} + 2e^- \rightarrow 2\text{SO}_4^{2-} \quad E^\circ = 2.01 \text{ V}\]

Give the overall cell reaction.

The half cell with the least positive reduction potential is reversed:

\[\text{Cl}_2 + \text{H}_2\text{O} + 5\text{S}_2\text{O}_8^{2-} \rightarrow 2\text{ClO}_3^- + 12\text{H}^+ + 10\text{SO}_4^{2-}\]

Calculate \(\Delta G^\circ\) and hence the value of \(K_c\) for the cell reaction at 298 K.

As the \(\text{ClO}_3^- / \text{Cl}_2\) half cell is reversed, the standard cell potential is:

\[E^\circ = ((2.01) + (-1.47)) \text{ V} = 0.54 \text{ V}\]

Using \(\Delta G^\circ = -nFE^\circ\), the free energy change for this 10e\(^-\) process is:

\[\Delta G^\circ = -(10) \times (96485 \text{ C mol}^{-1}) \times (0.54 \text{ V}) = -521000 \text{ J mol}^{-1} = -521 \text{ kJ mol}^{-1}\]

Using \(\Delta G^\circ = -RT\ln K_c\), the equilibrium constant is given by:

\[-521000 \text{ kJ mol}^{-1} = (8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}) \times \ln K_c \quad \text{so} \quad K_c = 2.14 \times 10^{90}\]

Alternatively, \(K_c\) can be obtained directly from \(E^\circ\) using the Nernst equation:

\[E^\circ = \frac{RT}{nF} \ln K_c = 0.54 \text{ V}\]

\[0.54 \text{ V} = \frac{(8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}) \ln K_c}{(10 \times 96485 \text{ C mol}^{-1})} \quad \text{so} \quad K_c = 2.14 \times 10^{90}\]

The reaction essentially proceeds to completion.

\[\Delta G^\circ = -521 \text{ kJ mol}^{-1} \quad K_c = 2.14 \times 10^{90}\]
Coordination complexes can display a number of types of isomerism. Draw a simple diagram showing a pair of geometric isomers. Label your diagram with the systematic name of each isomer.

The compound \([\text{CoCl}_2(\text{NH}_3)_4]\) can exist as in two isomeric forms which differ in the arrangement of the ligands in space.

\[
\begin{align*}
\text{cis-tetraamminedichloridocobalt(II)} &\quad \text{trans-tetraamminedichloridocobalt(II)} \\
\end{align*}
\]

Pt(\text{NH}_3)_2\text{Cl}_2 can also exist as \textit{cis} and \textit{trans} isomers:

\[
\begin{align*}
\text{cis-diamminedichloridoplatinum(II)} &\quad \text{trans-diamminedichloridoplatinum(II)} \\
\end{align*}
\]
- Draw all of the geometric isomers for the complex ion \([\text{CoI}_2(\text{NH}_3)_4]^+\). Label each isomer with its systematic name.

\[
\begin{align*}
\text{cis-tetraamminediiodocobalt(II)} & \quad \text{trans-tetraamminediiodocobalt(II)} \\
\text{cis-tetraamminediiodocobalt(II)} & \quad \text{trans-tetraamminediiodocobalt(II)}
\end{align*}
\]
Calculate the standard free-energy change for the following reaction at 298 K.

\[ 2 \text{Au}(s) + 3 \text{Ca}^{2+} (1.0 \text{ M}) \rightarrow 2 \text{Au}^{3+}(1.0 \text{ M}) + 3 \text{Ca}(s) \]

The reduction half cell reactions and \( E^0 \) values are:

\[
\begin{align*}
\text{Au}^{3+}(\text{aq}) + 3e^- &\rightarrow \text{Au}(s) & E^0 &= + 1.50 \text{ V} \\
\text{Ca}^{2+}(\text{aq}) + 2e^- &\rightarrow \text{Ca}(s) & E^0 &= -2.87 \text{ V}
\end{align*}
\]

In the reaction, Au is being oxidized and so the overall cell potential is:

\[ E^0 = ((-2.87) - (+1.50)) \text{ V} = -4.37 \text{ V} \]

The reaction involves 6 electrons so, using \( \Delta G^0 = -nFE^0 \):

\[
\Delta G^0 = - (6) \times (96485 \text{ C mol}^{-1}) \times (-4.37 \text{ V}) = +2530000 \text{ J mol}^{-1} \\
= +2.53 \times 10^3 \text{ kJ mol}^{-1}
\]

Answer: \(+2.53 \times 10^3 \text{ kJ mol}^{-1}\)

Complete and balance the following equation for the reaction between iron(II) ions and permanganate ions in an acidic solution.

\[ \text{Fe}^{2+} + \text{MnO}_4^- \rightarrow \text{Fe}^{3+} + \text{Mn}^{2+} \]

In acid, the relevant half cells are:

\[
\begin{align*}
\text{Fe}^{2+}(\text{aq}) &\rightarrow \text{Fe}^{3+}(\text{aq}) + e^- \\
\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5e^- &\rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(l)
\end{align*}
\]

Giving an overall reaction:

\[ 5\text{Fe}^{2+}(\text{aq}) + \text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) \rightarrow 5\text{Fe}^{3+}(\text{aq}) + \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(l) \]

Answer continues on the next page
What is the value of the equilibrium constant for the following reaction at 298 K?

\[ 2Fe^{3+}(aq) + Sn(s) \rightleftharpoons Sn^{2+}(aq) + 2Fe^{2+}(aq) \]

The reduction half cell reactions and \( E^0 \) values are:

- \( Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq) \) \( E^0 = +0.77 \text{ V} \)
- \( Sn^{2+}(aq) + 2e^- \rightarrow Sn(s) \) \( E^0 = -0.14 \text{ V} \)

In the reaction, Sn is being oxidized and so the overall cell potential is:

\[ E^0 = ((+0.77) - (-0.14)) \text{ V} = +0.91 \text{ V} \]

The reaction involves 2 electrons so, using \( E^0 = \frac{RT}{nF} \ln K \):

\[ \ln K = E^0 \times \frac{nF}{RT} = (+0.91 \text{ V}) \times \left( \frac{2 \times 96485 \text{ C mol}^{-1}}{8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}} \right) = 70.9 \]

\[ K = e^{-70.9} = 6.05 \times 10^{30} \]

Answer: \( 6.05 \times 10^{30} \)