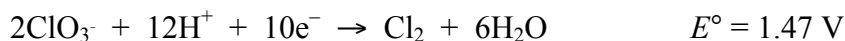
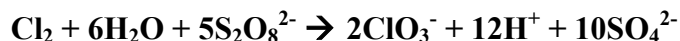


- Consider the following half-reactions and their standard reduction potentials.



Give the overall cell reaction.

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



Calculate ΔG° 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:

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

Alternatively, K_c can be obtained directly from E° using the Nernst equation:

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

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

The reaction essentially proceeds to completion.

$$\Delta G^\circ = -521 \text{ kJ mol}^{-1}$$

$$K_c = 2.14 \times 10^{90}$$