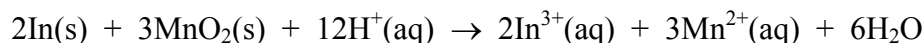


- Consider the following balanced redox reaction.



If  $E^\circ = 1.568\text{ V}$ , what would be the measured potential of this cell at 298 K at the following concentrations?

$$[\text{H}^+\text{(aq)}] = 0.25\text{ M}; \quad [\text{In}^{3+}\text{(aq)}] = 0.20\text{ M}; \quad [\text{Mn}^{2+}\text{(aq)}] = 0.42\text{ M}$$

**Marks**  
**2**

The reaction quotient  $Q$  is given by:

$$Q = \frac{[\text{In}^{3+}\text{(aq)}]^2 [\text{Mn}^{2+}]^3}{[\text{H}^+\text{(aq)}]^{12}}$$

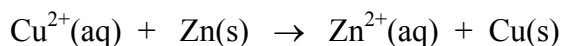
The cell reaction involves a  $6e^-$  process ( $2\text{In} \rightarrow 2\text{In(III)}$  and  $3\text{Mn(IV)} \rightarrow 3\text{Mn(II)}$ ). The cell potential can be calculated using the Nernst equation,

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

$$= (1.568\text{ V}) - \frac{(8.314\text{ J K}^{-1}\text{ mol}^{-1}) \times (298\text{ K})}{(6) \times (96485\text{ C mol}^{-1})} \ln \left[ \frac{(0.20)^2 (0.42)^3}{(0.25)^{12}} \right] = 1.52\text{ V}$$

Answer: **1.52 V**

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



Relevant electrode potentials can be found on the data page.

**2**

The standard reduction potentials for  $\text{Cu}^{2+}\text{(aq)} / \text{Cu(s)}$  and  $\text{Zn}^{2+}\text{(aq)} / \text{Zn(s)}$  are  $+0.34\text{ V}$  and  $-0.76\text{ V}$  respectively. The  $\text{Zn}^{2+}\text{(aq)} / \text{Zn}$  potential is the less positive and is reversed. The cell potential is therefore:

$$E_{\text{cell}}^\circ = ((0.34) - (-0.76))\text{ V} = 1.10\text{ V}$$

The equilibrium constant for this  $2e^-$  process can be calculated using  $E_{\text{cell}}^\circ$

$$= \frac{RT}{nF} \ln K :$$

$$\ln K = \frac{(2) \times (96485\text{ C mol}^{-1})}{(8.314\text{ J K}^{-1}\text{ mol}^{-1}) \times (298\text{ K})} \times (1.10) = 85.7$$

$$K = e^{85.7} = 1.62 \times 10^{37}$$

Answer:  **$1.62 \times 10^{37}$**