

Gibbs energy

$$\Delta G = -n F E_{\text{cell}}$$

ΔG → Gibbs energy change
 n → no. of electrons exchange in chemical equation
 F → Faraday
 E_{cell} → cell potential

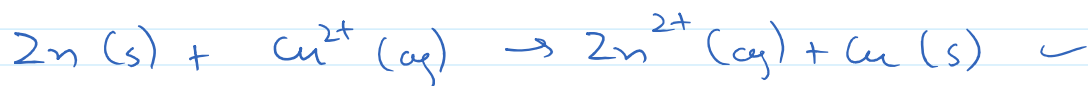
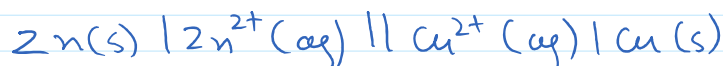
$$E/w = q \times V$$

$$\begin{aligned} 1 \text{ mole } e^- &= 1 F \\ n \text{ mole } e^- &= n F \end{aligned}$$

$$W = (n F) E$$

$$\Delta G^\circ = -n F E_{\text{cell}}^\circ$$

Equilibrium constant from Nernst Equation



By Nernst equation (298K)

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.059}{n} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

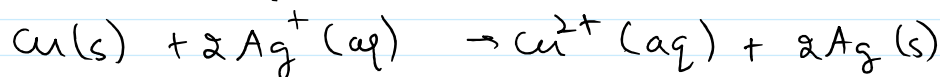
At equilibrium $E_{\text{cell}} = 0$, $\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = K_c$

$$0 = E_{\text{cell}}^\circ - \frac{0.059}{n} \log K_c$$

$$\log K_c = \frac{n E_{\text{cell}}^\circ}{0.059}$$

Question

Calculate the equilibrium constant of the reaction



$$E_{\text{cell}}^\circ = 0.46 \text{ V}$$

Answer

$$\log K_c = \frac{n E_{\text{cell}}^\circ}{0.059} = \frac{2 \times 0.46}{0.059} = 15.6$$

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$$\log K_c = \frac{n E^\circ}{0.059} = \frac{2 \times 0.46}{0.059} = 15.6$$

$$\begin{aligned}\log K_c &= 15 + 0.6 \\ &= \log 10^{15} + \log 4 \\ &= \log 4 \times 10^{15} \\ K_c &= 4 \times 10^{15}\end{aligned}$$

Question

The standard electrode potential for Daniell cell is 1.1V. Calculate the standard Gibbs energy for the reaction:
 $\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$

Answer

$$\begin{aligned}\Delta G^\circ &= -n F E^\circ_{\text{cell}} \\ &= -2 \times 96500 \times 1.1 = -212300 \text{ J mol}^{-1} \\ &= -212.3 \text{ kJ mol}^{-1}\end{aligned}$$

Question

The cell in which the following reaction occurs

$2\text{Fe}^{3+}(\text{aq}) + 2\text{I}^{-}(\text{aq}) \rightarrow 2\text{Fe}^{2+}(\text{aq}) + \text{I}_2(\text{s})$ has $E^\circ_{\text{cell}} = 0.236\text{V}$ at 298K. Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction.

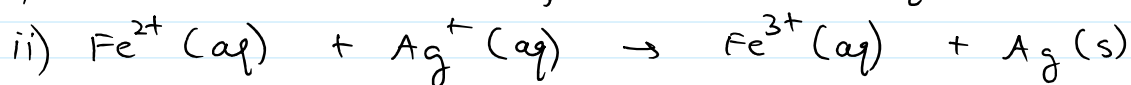
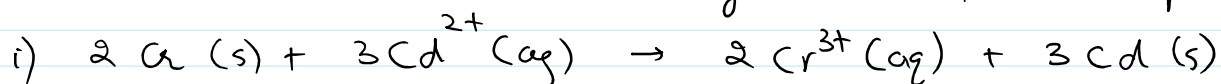
Answer

$$\begin{aligned}\text{i) } \Delta G^\circ &= -n F E^\circ_{\text{cell}} = -2 \times 96500 \times 0.236 = -45548 \text{ J mol}^{-1} \\ &= -45.55 \frac{\text{kJ}}{\text{mol}}\end{aligned}$$

$$\begin{aligned}\text{ii) } \log K_c &= \frac{n E^\circ_{\text{cell}}}{0.059} \\ &= \frac{2 \times 0.236}{0.059} \\ &= 8 \\ K_c &= 10^8\end{aligned}$$

Question

Calculate the standard cell potentials of galvanic cell in which the following reaction take place



Calculate the ΔG° and equilibrium constant of the reaction

$$E^\circ_{\text{Cr}^{3+}/\text{Cr}} = -0.74 \text{V}, \quad E^\circ_{\text{Cd}^{2+}/\text{Cd}} = -0.40 \text{V}$$

$$E^\circ_{\text{Fe}^{3+}/\text{Fe}^{2+}} = \underline{0.77 \text{V}}, \quad E^\circ_{\text{Ag}^+/\text{Ag}} = \underline{0.80 \text{V}}$$

Answer

$$i) \quad E^\circ_{\text{cell}} = E^\circ_{\text{Cd}^{2+}/\text{Cd}} - E^\circ_{\text{Cr}^{3+}/\text{Cr}} = -0.40 - (-0.74) \\ = 0.34 \text{V}$$

$$\Delta G^\circ = -nFE^\circ_{\text{cell}} = -6 \times 96500 \times 0.34 = -196860 \text{ J mol}^{-1} \\ = -196.86 \frac{\text{kJ}}{\text{mol}}$$

$$\log K_c = \frac{n E^\circ_{\text{cell}}}{0.059} = \frac{6 \times 0.34}{0.059}$$

$$\log K_c = 34.6$$

$$= 34 + 0.6$$

$$= \log 10^{34} + \log 4$$

$$= \log 4 \times 10^{34}$$

$$K_c = 4 \times 10^{34}$$

$$ii) \quad E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$= E^\circ_{\text{Ag}^+/\text{Ag}} - E^\circ_{\text{Fe}^{3+}/\text{Fe}^{2+}}$$

$$= 0.80 - 0.77$$

$$= 0.03 \text{V}$$

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ = -1 \times 96500 \times 0.03 = -2895 \text{ J mol}^{-1}$$

$$= -2.895 \text{ kJ mol}^{-1}$$

$$\log k_c = \frac{n E_{\text{cell}}^\circ}{0.059} = \frac{1 \times 0.03}{0.059} = 0.5$$

$$k_c = 10^{0.5}$$

$$= 3.2$$

$$\log 2 = 0.3$$

$$\log 3 = 0.48$$

$$\log 5 = 0.7$$

$$\log 7 = 0.85$$

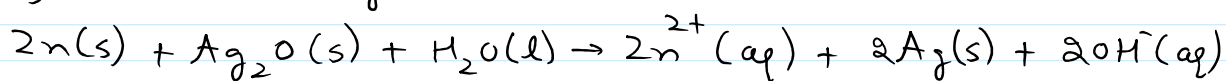
$$\log a b = \log a + \log b, \quad \log \frac{a}{b} = \log a - \log b$$

$$\log a^b = b \log a$$

$$\sqrt{10} = 3.16$$

Question

In the button cell widely used in watches and other devices, the following reaction takes place:



Determine ΔG° and E° for the reaction

$$E_{\text{Zn}^{2+}/\text{Zn}}^\circ = -0.76 \text{ V}$$

$$E_{\text{Ag}_2\text{O}/\text{Ag}}^\circ = 0.344 \text{ V}$$

Answer

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$$

$$= E_{\text{Ag}_2\text{O}/\text{Ag}}^\circ - E_{\text{Zn}^{2+}/\text{Zn}}^\circ = 0.344 - (-0.76)$$

$$= 1.104 \text{ V}$$

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ = -2 \times 96500 \times 1.104 = -213072 \text{ J mol}^{-1}$$

$$= -213.072 \text{ kJ mol}^{-1}$$