Gibbs energy

$\Delta G = -nFE_{cell}$

Faraday

ΔG — Gibbs energy change

n — no. of electrones exchange in chemical equation

E_{cell} — cell potential

$E/w = q \times V$

1mole e^- = 1 F
n mole e^- = n F

E

$w = (nF)E$

$\Delta G^\circ = -nFE^\circ_{cell}$

Equilibrium constant from Nernst Equation

$Zn(s) | Zn^{2+}(aq) || Cu^{2+}(aq) | Cu(s)$

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

By Nernst equation (298K)

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

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

$$0 = E^\circ_{cell} - \frac{0.059}{n} \log K_c$$

$$\boxed{\log K_c = \frac{nE^\circ_{cell}}{0.059}}$$

Question

Calculate the equilibrium constant of the reaction

$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

$E^\circ_{cell} = 0.46V$

Answer

$\log K_c = \frac{nE^\circ}{} = \frac{2 \times 0.46}{} = 15.6$

Answer

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

$$\log K_c = 15 + 0.6$$
$$= \log 10^{15} + \log 4$$
$$= \log 4 \times 10^{15}$$
$$K_c = 4 \times 10^{15}$$

Question

The standard electrode potential for Daniell cell is 1.1V. Calculate the standard Gibbs energy for the reaction:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

Answer:

$$\Delta G^\circ = -nFE^\circ_{cell}$$
$$= -2 \times 96500 \times 1.1 = -212300\ J\ mol^{-1}$$
$$= -212.3\ kJ\ mol^{-1}$$

Question

The cell in which the following reaction occurs

$2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(s)$ has $E^\circ_{cell} = 0.236V$ at 298K. Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction.

Answer

i) $\Delta G^\circ = -nFE^\circ_{cell} = -2 \times 96500 \times 0.236 = -45548\ J mol^{-1}$

$$= -45.55\ \frac{kJ}{mol}$$

ii) $$\log K_c = \frac{nE^\circ_{cell}}{0.059}$$
$$= \frac{2 \times 0.236}{0.059}$$
$$= 8$$
$$K_c = 10^8$$

Question

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

i) $2\,Cr(s) + 3\,Cd^{2+}(aq) \rightarrow 2\,Cr^{3+}(aq) + 3\,Cd(s)$

ii) $Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$

Calculate the ΔG° and equilibrium constant of the reaction

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

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

Answer

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

$= 0.34V$

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

$= -196.86\ \frac{kJ}{mol}$

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

$\log k_c = 34.6$

$= 34 + 0.6$

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

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

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

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

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

$= 0.80 - 0.77$

$= 0.03V$

$$\Delta G^\circ = -nFE^\circ_{cell} = -1 \times 96500 \times 0.03 = -2895 \ J\,mol^{-1}$$
$$= -2.895 \ kJ\,mol^{-1}$$

$$\log K_c = \frac{nE^\circ_{cell}}{0.059} = \frac{1 \times 0.03}{0.059} = 0.5$$

$$K_c = 10^{0.5}$$
$$= 3.2$$

$\log 2 = 0.3$

$\log 3 = 0.48$

$\log 5 = 0.7$

$\log 7 = 0.85$

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

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

$\sqrt{10} = 3.16$

Question

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

$$Zn(s) + Ag_2O(s) + H_2O(l) \rightarrow Zn^{2+}(aq) + 2Ag(s) + 2OH^-(aq)$$

Determine ΔG° and E° for the reaction

$E^\circ_{Zn^{2+}/Zn} = -0.76 \ V$ $\qquad$ $E^\circ_{Ag_2O/Ag} = 0.344V$

Answer

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$
$$= E^\circ_{Ag_2O/Ag} - E^\circ_{Zn^{2+}/Zn} = 0.344 - (-0.76)$$
$$= 1.104V$$

$$\Delta G^\circ = -nFE^\circ_{cell} = -2 \times 96500 \times 1.104 = -213072 \ J\,mol^{-1}$$
$$= -213.072 \ kJ\,mol^{-1}$$