② A voltaic cell has an Ecell value of 1.536V. Calculate the concentration of Ag^+ in the cell.

$$Zn_{(s)} / Zn^{2+}(2.00\,M) // Ag^+(x\,M) / Ag_s$$

Solution

Oxidation: $Zn_{(s)} \longrightarrow Zn^{2+} + 2e \times 1$ $\quad E^\circ = -0.763V$

Reduction: $Ag^+ + \bar{e} \longrightarrow Ag \times 2$ $\quad E^\circ = 0.800V$

Net Rxn: $Zn_{(s)} + 2Ag^+_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + Ag$ $\quad E^\circ_{cell} = \underline{\underline{1.563V}}$

$$E^\circ_{cell} = E^\circ_c - E^\circ$$
$$0.800V - (-0.763 = \underline{\underline{1.563V}}$$

Nernst equation

oxidation

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{n} \log \frac{[Zn^{2+}]}{[Ag^+]^2}$$

reduction

$$1.536V = 1.563V - \frac{0.0591}{2e} \log \frac{(2.00)}{(X)^2}$$

$$\frac{-0.027V}{-0.02955} = \frac{-0.02955}{-0.02955} \log \frac{2}{X^2}$$

$$\log \frac{2}{X^2} = 0.9137$$

$$\frac{2}{X^2} = 10^{0.9137}$$

$$[Ag^+]\ X = \underline{\underline{0.495}}$$

~~antilog~~ log $\frac{2}{X^2}$ = antilog 0.9137

$$[Ag^+] = \underline{\underline{0.495}}$$

to obtain 10.0g of liquid mercury by passing a constant current of 0.17A through a solution containing $Hg_2(NO_3)_{2(aq)}$?

[4 marks]

Solution

(a) $Fe^{2+}_{(aq)} + 2e \rightarrow Fe \quad E^o_{cell} = -0.44V$

$Cd^{2+} + 2e \rightarrow Cd \quad E^o_{cell} = -0.40V$

$E^o_{cell} = E^o_C - E^o_A$

$-0.40V - (-0.44) = +0.04V$

(b) $E_{cell} = E^o_{cell} - \frac{0.0591}{2e} \log[Q]$

$E_{cell} = +0.04V - \frac{0.0591}{2} \log\frac{(0.800)}{(0.200)}$

$E_{cell} = +0.02V$

(c) $\Delta G^o = -nFE^o$

$-2 \times 96{,}500\ C/mol \times 0.04V$

$= -$

$C = nF$

It

$Q = I\,(C/s) \times s$

$mole^- = C \times \frac{1\,mole^-}{F}$

$\frac{Mass}{1} \times \frac{Q\,Mr}{nF} = \frac{It\,Mr}{nF}$

$Q\,Mr = \frac{Mass\,nF}{Mr} = \frac{10.0g \times 2 \times 96500}{525.19} = 3{,}674.9\ Cs$

$Q = It \qquad t = \frac{Q}{I} = \frac{3674.9}{0.17} = \frac{21616.83s}{60} = 360.3s$

since we have been given ...
we have;

Oxidation Anode

① $E_{cell} = E^{\circ}_{cell} - \frac{0.0591}{n} \log \frac{[Fe^{2+}]}{[Ni^{2+}]}$

cathode Reduction

$E_{cell} = +0.21 - \frac{0.0591}{2} \log \frac{(0.25)}{(0.5)}$

$E_{cell} = \underline{0.219V}$

free energy $\Delta G = -nFE$
$-2 \times 96500\ c/mol \times 0.219V$
$= \underline{-42267J}$

$\boxed{\Delta G^{\circ} \text{ Standard gibbs free energy} \neq \Delta G \text{ Gibbs free energy}}$

Practice Question

1. An electrochemical cell consists of 1.0L half-cells of Fe/Fe^{2+} and Cd/Cd^{2+} with the following initial concentrations; $[Fe^{2+}] = 0.800M$, $[Cd^{2+}] = 0.200M$
   - (a) write the overall cell reaction [3]
   - (b) What is the initial Ecell at 25°C [5]
   - (c) calculate the ΔG° of the reaction at 25°C [3]
   - (d) How many minute would be required

$\log K = 31.47$

antilog $\log K$ = antilog 31.47

$K = 2.95 \times 10^{31}$

4 [eqn? 2 marks]

(9) (a) Write the overall reaction for cell consist of Ni^{2+}/Ni and Fe^{2+}/Fe half cells.

(b) (i) Calculate cell potential and (ii) free energy change under standard conditions.

(c) Calculate ΔG for Ni^{2+}/Ni and Fe^{2+}/Fe half-cells if the concentrations of ~~Ni^{2+} = 0.5M~~ $Ni^{2+} = 0.50M$ and $Fe^{2+} = 0.25M$.

Solution:

(a) Half rxns; $Ni^{2+} + 2\bar{e} \rightleftharpoons Ni_{(s)}$ $E^0 = -0.23V$

$Fe^{2+} + 2e \rightleftharpoons Fe_{(s)}$ $E^0 = -0.44V$

(b) (i) cell potential

$E^0_{cell} = E^0_{cathode} - E^0_{anode}$

$-0.23V - (-0.44V)$

$= 0.21V$

(ii) standard

$\Delta G^0 = -nFE^0$

$= -2 \times 96485 C/s \times 0.21V$

$= -40,530 J$

(a) In shorthand notation, voltaic cell that would deliver the highest voltage.

FROM THESE

$$\boxed{\begin{array}{ll} Ag^+_{(aq)} + e^- \rightleftharpoons Ag_{(s)} & E^\circ = +0.80V \\ Cu^{2+}_{(aq)} + 2e^- \rightleftharpoons Cu_{(s)} & E^\circ = +0.34V \\ Pb^{2+}_{(aq)} + 2e^- \rightleftharpoons Pb_{(s)} & E^\circ = -0.13V \end{array}}$$

Anode: $Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$

Cathode: $Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$

Line notation (since done at 25°c we can include 1.0M)

$$Pb_{(s)} / Pb^{2+}_{(aq)}(1M) // Ag^+_{(aq)}(1M) / Ag_{(s)}$$

(b) Draw the voltaic cell in part (a)

Already done (check) *

(c) write the rexns for half cells (we have done it)

(d) calculate equilibrium constant for the reaction.

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{n} \log(Q)$$

at equilibrium ($Q = K$) where $E_{cell} = 0$

$$E^\circ_{cell} = \frac{0.0591}{n} \log(K)$$

$$+0.93V = \frac{0.0591}{2} \log K$$

Practice Question

Consider the following half reactions and their voltages:

$Ag^+_{(aq)} + e^- \rightleftharpoons Ag_{(s)}$; $E^o = +0.80V$

$Cu^{2+}_{(aq)} + 2e^- \rightleftharpoons Cu_{(s)}$; $E^o = +0.34V$

$Pb^{2+}_{(aq)} + 2e^- \rightleftharpoons Pb_{(s)}$; $E^o = -0.13V$

Check: We can construct three different voltaic cells (galvanic)

① Half cells

$Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)}$ $E^o = +0.80V$ (Cathode)

$Pb^{2+} + 2e^- \longrightarrow Pb_{(s)}$ $E^o = -0.13V$

Overal Reaction

$Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)} \times 2$

$Pb_{(s)} \longrightarrow Pb^{2+} + 2e \times 1$

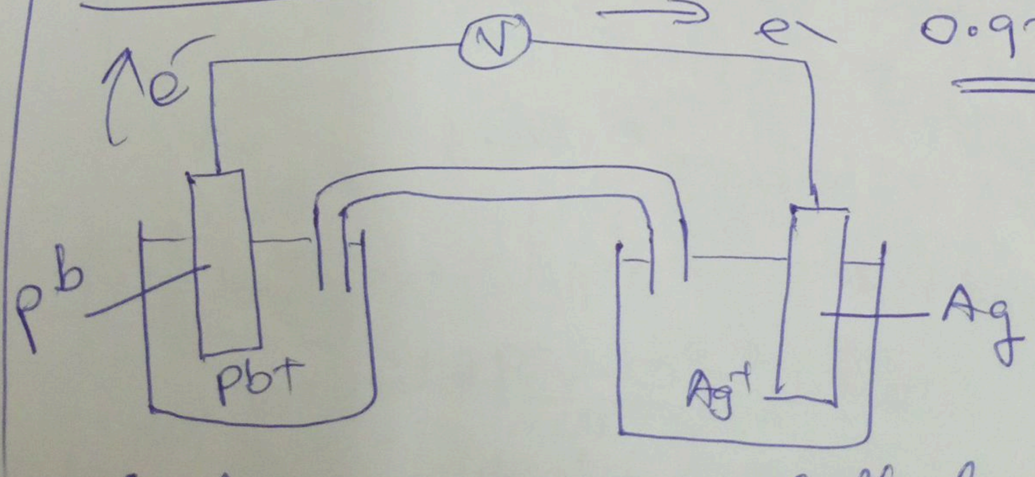
$2Ag^+_{(aq)} + Pb_{(s)} \longrightarrow 2Ag_{(s)} + Pb^{2+}_{(aq)}$ E_{cell}

$E^o_{cell} = E^o_c - E_{Anode}$

$+0.80V - (-0.13V)$

$\underline{\underline{0.93V}}$

Construction



Anode Cathode