concentration cell: same material, diff. conc. - flows from side with lower concentration to higher concentration

$\Delta G = -nFE^\circ_{cell} = -2\,mol\,(96{,}485\ J/V\text{-}mol)\,1.10$

$\Delta G = -nFE^\circ = -RT\ln K$

What is K for the copper-silver voltaic cell reaction?

$Cu_{(s)} + Ag^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + Ag_{(s)}$ $\quad (Ag^+_{(aq)} + 1e^- \rightarrow Ag_{(s)})2 \quad E^\circ_{red} = .8V$

separate into two half reactions: $\quad Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^- \quad E^\circ_{red} = .34V$

Ag being reduced - cathode

Cu being oxidized - anode

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

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 0.8\,V - 0.34V = 0.46V$

$E^\circ = \frac{RT}{nf}\ln K$ so $\ln K = \frac{nFE^\circ}{RT} = \frac{(2\,mol)(96{,}485\,J/V\text{-}mol)(0.46)}{RT}$

$K = 4.3 \times 10^{15}$ ← very favorable

Nernst Equations: accounting for non-standard conditions

- $\Delta G = \Delta G^\circ + RT\ln Q$
- substituting ΔG terms: $-nFE = -nFE + RT\ln Q$ - divide both sides by $-nF$

$$\boxed{E = E^\circ - \frac{0.0592\,V}{n}\log Q}$$

nonstandard conditions bc not 1M

$2Ag^+_{(aq,\,0.8M)} + Cu_{(s)} \rightarrow 2Ag^+_{(s)} + Cu^{2+}_{(aq,\,0.00M)}$

$E_{cell} = E^\circ_{cell} - \frac{0.0592V}{n}\log Q, \quad Q = \frac{[Cu^{2+}]}{[Ag^+]^2} = \frac{0.00}{.8^2}$

using standard temperature, the constants R and F, and converting to log gives an alternative form

$0.46V - \frac{0.0592\,V}{2\,mol}\log Q = 0.54V$ ← more favorable than under standard conditions bc positive which means that it is bigger

Ex. Voltaic cell based on Tin and Aluminum

$2\times[Al^{3+}_{(aq)} + 3e^- \rightarrow Al_{(s)}, \quad E^\circ = -1.66V$ ← probably going to give up electrons.

$3\times[Sn^{2+}_{(aq)} + 2e^- \rightarrow Sn_{(s)}, \quad E^\circ = 0.15V, \quad E^\circ_{cell} = 0.15V - (-1.66V)$

cathode anode

$2Al_{(s)} + 3Sn^{2+} \rightarrow 2Al^{3+}_{(aq)} + 3Sn_{(s)}$

$\Delta G^\circ_{rxn} = nFE^\circ = 6\,mol \cdot F \cdot 1.81V = -1050\,kJ$ ← a lot of energy. good battery!

product reactant

Case 1. $[Al^{3+}] = 3.0M, \quad [Sn^{2+}] = 0.75M$ ∴ products > reactants, log Q > 0, subtracting a positive #, which is less favorable

Case 2. $[Al^{3+}] = 0.75M, \quad [Sn^{2+}] = 3.0M$

∴ More reactant, shift right, more favorable

No difference in energy between half cells of equal M (see ppt slide 19)

Slide 15: flow from anode to cathode - by convention.

anode is on the left and cathode is on the righ