Chapter 19 Electrochemistry Math Summary

Relating Standard Cell Potential to Standard Half Cell Potentials

$E^o_{cell}=E^o_{oxidation} + E^o_{reduction}$ (standard conditions assume 1.0 M concentrations)

Relating Half Cell Potentials when Written in Opposite Directions

$E^o_{ox} = -E^o_{red}$ for half reactions written in opposite directions

Relating Standard Cell Potentials to ΔG

$\Delta G^o = -nFE^\circ_{cell}$ (to give answer in kJ, use F = 96.485)

F = 96,500 C/mol
n=number of electrons transferred

Relating Actual Cell Potential to Standard Cell Potential when Concentrations aren't 1.0-M

$E_{cell} = E^o_{cell} - [0.0592/n] \log Q$ (Q = ratio of actual concentrations)

Relating Standard Cell Potential to Equilibrium Constant

$\log K = nE^o/0.0592$

Relating Actual Cell Potential to Actual Concentrations in Concentration Cells

$E_{cell} = -[0.0592/n] \log Q$ for concentration cells, where anode and cathode differ only in concentration, but otherwise have same ions

Relating # of Moles of Electrons Transferred as a Function of Time and Current in Electrolysis

1 mol e^- = 96,500 C

moles of electrons = [current (A)•time (sec)]/96,500 for electrolysis, moles, current, and time are related.

rearranged: time (sec)=(moles of electrons)(96500)/current (in A)
Note: 3600 sec/hour
so time (hours)=(moles of electrons)(26.8)/current (in A)

Electrochemistry-Related Units

C = Coulomb = quantity of electrical charge = $6.24 \cdot 10^{18}$ electrons

- 1 mole of electrons = 96,500 C

A = amp = rate of charge flow per time = C/sec
V = volt = electrical power/force/strength = J/C

$$F = \text{Faraday} = \frac{96{,}500\text{C}}{\text{mole } e^-} = \frac{96.5 \text{ kJ}}{\text{mole } e^- \bullet \text{V}}$$