Chemistry BC2001x: General Chemistry I

Lecture 20: Tuesday November 14, 2009

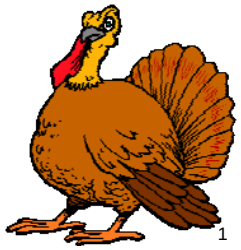
Topic: The Nernst Equation and Batteries

Pick up: Graded problem set 9

Study guide for Exam 3 is on WWW

No office hours tomorrow, Wednesday Nov 15.

Happy thanksgiving.



The Nernst Equation

Cell diagram example: $Ag(s) \mid Ag^+ \parallel Cu^{2+} \mid Cu(s)$

reduction (cathode) on right: $\Delta\mathscr{E}^o_{cell} = \mathscr{E}^o_R - \mathscr{E}^o_L$

Subtract reduction potentials (do not change signs)

Nernst equation: (subscript cell often omitted)

$\Delta\mathscr{E}_{cell} = \Delta\mathscr{E}^o_{cell} - (RT/n\mathscr{F})\ \ln Q$

$\Delta\mathscr{E}_{cell} = \Delta\mathscr{E}^o_{cell} - (2.303\ RT/n\mathscr{F})\ \log Q$

$\Delta\mathscr{E}_{cell} = \Delta\mathscr{E}^o_{cell} - (k_N/n)\ \log Q$

$k_N = (2.303\ RT/\mathscr{F}) = 0.05916\ V$ at 25°C

n = # of electrons transferred in balanced equation

$\mathscr{F}$ = Faraday constant (Coulombs per mole of charge)

Using the Nernst Equation

Nernst equation:

$\Delta\mathscr{E}_{cell} = \Delta\mathscr{E}^o_{cell} - (RT/n\mathscr{F})\ \ln Q$

$\Delta\mathscr{E}_{cell} = \Delta\mathscr{E}^o_{cell} - (k_N/n)\ \log Q$

$k_N \equiv (2.303\ RT/\mathscr{F}) = 0.05916\ V$ at 25°C

n = # of electrons transferred in balanced equation

$\mathscr{F}$ = Faraday constant (Coulombs per mole of charge)

At equilibrium $\Delta\mathscr{E}_{cell} = 0$ and $Q = K$

Therefore $\Delta\mathscr{E}^o_{cell} = (k_N/n)\ \log K$

and $\log K = n\Delta\mathscr{E}^o_{cell}/k_N$

Batteries

car battery: lead-acid