Chem 116 Lecture 22 Notes 12/12/08 (MO)

SOHe method of balancing redox equations
*If you balance half reactions using this method, you will always have a balanced overall reaction

Practice using SOHe method
*Know difference between species and element

0 +5 +2 +4

1. $Cu(s) + NO_3^-(aq) \rightarrow Cu^{2+}(aq) + NO_2(g)$ in acid

Elements whose oxidation numbers are changing
Cu: $0 \rightarrow +2$ (oxidation)
N: $+5 \rightarrow +4$ (reduction)

Species on reactant side	*Species on product side*
Cu (s) (oxidized)	Cu^{2+} (aq)
NO_3^- (reduced)	NO_2 (g)

Oxidation half reaction
$Cu \rightarrow Cu^{2+} + 2e-$
*add 2e- to product side to make charge on reactant side = charge on product side

Reduction half reaction
$NO_3^- + 2H^+ + e- \rightarrow NO_2 + H_2O$

*Add oxygens needed by adding H_2O
*Add H^+ to side that needs hydrogens (do this if the redox reaction is in acid solution)
    if the redox reaction were happening in basic solution, you would add OH^- to one side and H_2O to the other side, to balance the hydrogens, because the side that gets H_2O is adding one more H
*Add e- to balance net charge

Overall reaction
Oxidation: $Cu \rightarrow Cu^{2+} + e-$
(Reduction: $NO_3^- + 2H^+ + e- \rightarrow NO_2 + H_2O$) $\times 2$
Overall rxn: $Cu + 2\,NO_3^- + 4H^+ \rightarrow 2NO_2 + 2\,H_2O + Cu^{2+}$

*Multiply reduction rxn by 2 to make e- cancel
note that the least common multiple of electrons (which canceled) was 2 electrons

Another Practice using SOHe method

0 +4 +2 +2

2. Cd (s) + NiO_2 (s) + H_2O (l) → $Cd(OH)_2$ (s) + $Ni(OH)_2$ (s)
   In acid

Elements whose oxidation numbers are changing
Cd: 0 → +2 (oxidized)
Ni: +4 → +2 (reduced)

Species on reactant side
Cd (s) is getting oxidized, so it is the reducing agent
NiO2 is getting reduced so it is the oxidizing agent

Oxidation half reaction
Cd (s) + $2H_2O$ → $Cd(OH)_2$ (s) + $2H^+$ + 2e-

net charge = 0 net charge = +2
2e- added to neutralize charge

*When adding e- check with oxidation numbers (oxidation side) to ensure correct number of e- were added.

Reduction half reaction
NiO_2 (s) + $2H^+$ + 2e- → $Ni(OH)_2$ (s)

*Check with oxidation numbers (reduction side) that correct number of e- were added

Overall reaction
Cd(s) + $2H_2O$ + NiO_2 (s) → $Cd(OH)_2$ + $Ni(OH)_2$

least common multiple of electrons that canceled when you were adding up the half reactions is n = 2 (need this for Nernst equation – save for later)

Electrochemical Cell Animation shown in class

$Cu \rightarrow Cu^{2+}$ or $Cu^{2+} \rightarrow Cu$ is happening in the Cu cell
$Zn \rightarrow Zn^{2+}$ or $Zn^{2+} \rightarrow Zn$ is happening in the Zn cell

*Use table of standard reduction potentials to figure out which one of cells is oxidation (anode) and which one is reduction (cathode)

Reduction potentials are:
$Cu^{2+} + 2e^- \rightarrow Cu$ $E^o = +.337$ V
$Zn^{2+} + 2e^- \rightarrow Zn$ $E^o = -.763$ V
note that both of these half-reactions are written as reductions
for a redox reaction, one of the half-reactions stays a reduction, and the other one turns around and becomes an oxidation
to get E^o for an oxidation, just flip the sign of E^o for the reduction when you turn the reaction backwards to get the oxidation half-reaction

Two possibilities

1. $Cu \rightarrow Cu^{2+} + 2e^-$ $E^o = -.337$ V (oxidized)
   $Zn^{2+} + 2e^- \rightarrow Zn$ $E^o = -.763$ V (reduced)
   E^o cell = (-.337)+(-.763)
   E^o cell = -1.10 V

2. $Zn \rightarrow Zn^{2+} + 2e^-$ $E^o = +.763$ V (oxidized)
   $Cu^{2+} + 2e^- \rightarrow Cu$ $E^o = +.337$ V (reduced)
   E^o cell = (+.763)+(+.337)
   E^o cell = +1.10
   ***This one is spontaneous because E^o cell is positive**

Parts of an electrochemical cell and cell notation
*Standard cell notation is not in book….only in class – will be on final exam