Crack the OAT

Crack Notes [General Chem 7]
Electrochemistry

Redox Reactions

LEO the lion says GER / OIL RIG
Lose electrons = *oxidized*, gain electrons = *reduced*
Something that is reduced is an *oxidizing agent*, if it's oxidized it's a *reducing agent*
Redox reactions involve something being reduced, something being oxidized

List of oxidation states:
Elements=0, F=-1, H=+1 (unless bonded with metal), O=-2 (unless peroxide)
Sum of oxidation states on molecule must equal to charge on molecule

Balancing: must balance both MASS and CHARGE.
1. Divide into half reactions (reduction, oxidation)
FOR EACH HALF REACTION...
2. Balance elements that aren't H, O
3. Balance O's by adding H_2O to one side
4. Balance H's by adding H^+ to one side
5. Balance charge by adding e^- to one side
NOW COMBINE THEM:
6. Multiply each half reaction by integer so e-'s used are equal
7. Add two half reactions together, e-'s should cancel out. Also cancel other things if you can.
8. If in basic solution, add OH- to both sides until all H's are used up

EXAMPLE $Cr(OH)_3\,(s) + ClO_3^-\,(aq) \rightarrow CrO_4^{2-}\,(aq) + Cl^-\,(aq)$ in basic conditions

Electric Potential

Reduction potential: how much something WANTS to be reduced (generally converting from cation to uncharged element)
Oxidation potential: how much something WANTS to be oxidized, equal to negative of reduction potential of oxidized state

All potentials given relative to oxidation potential of H_2 / reduction potential of $H^+ = 0$

Given something like this: $Au^{3+} + 3e^- \rightarrow Au\,(s)$, E^O=1.50 V:
Reduction potential for Au^{3+} is 1.50 V
Oxidation potential for Au (s) is -1.50 V
So Au^{3+} is a STRONG oxidizing agent (b/c high reduction potential) but a WEAK reducing agent (b/c low oxidation potential)

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Basically for two compounds reacting together, higher reduction potential one will be reduced, higher oxidation/lower reduction potential one will be oxidized.

Potential of an ionic reaction: add potentials for each *half reaction* (each individual ion)

Galvanic Cell
AN OX, RED CAT (anode=oxidation, cathode=reduction)
Has an *anode* and a *cathode* connected by a wire where electrons can flow. Oxidation happens at the anode (converting an element to a positive ion) and reduction happens at the cathode (converting an element to a negative ion)
Can figure out which is anode/cathode by looking at reduction potentials: higher reduction potential = will be reduced = cathode
Can have a *salt bridge* where ions in solution can flow to counterbalance the electrons flowing through wire.
Electromotive force/cell potential: reduction potential at cathode - oxidation potential at anode (or sum of reduction potentials for elements at cathode + anode)
Cell diagram: Anode|AnodeIon||CathodeIon|Cathode

Formulas
Free energy: $\Delta G = -nFE$ or $\Delta G^\circ = -nFE^\circ$, so positive cell potential = spontaneous reaction
F = Faraday's constant = 96500 C/mol
n = number of moles of electrons in balanced redox reaction
At non-standard state: $\Delta G = \Delta G^\circ + RT \ln Q$
Can also write $\Delta G^\circ = -RT \ln K$ where *K* is equilibrium constant for redox reaction
-> K=1 (equal amount reactants/products): $\Delta G^\circ = 0$
-> K>1 (more products): $\Delta G^\circ < 0$ or spontaneous reaction
-> K<1 (less products): $\Delta G^\circ > 0$ or non spontaneous reaction

Concentration Cell
Same chemical used for anode/cathode, just at different concentrations
More concentration of positive ion = reduction will occur, or cathode
Less concentration of positive ion = oxidation will occur, or anode
Nernst equation can be used to find cell potential based on concentration difference

Electrolytic Cell
Supplying power to cause opposite of galvanic cell (backwards from what we expect)