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***Lecture 2 Notes, Immaculata Week, July-August 2013, Charles H. Mahler, Lycoming College***

My email address is [mahler@lycoming.edu](https://mail.lycoming.edu/owa/redir.aspx?C=ab05c5cbaa4a4ba88781c5c6d6c17bd7&URL=mailto%3amahler%40lycoming.edu) – please contact me if you have questions.

Our topic was oxidation-reduction and its relation to electrochemistry and we started with a review of redox chemistry and assigning oxidation states (oxidation numbers) in balanced equations. We first assigned an oxidation state or number to each atom in the reaction:

2 MnO2 (s) + H2 (g) 🡪 Mn2O3 (s) + H2O (l)

In MnO2 each O atom is -2, so the Mn is +4, in H2 each H is 0, in Mn2O3 each O is -2, so each Mn is +3, and in H2O each H is +1 and the O is -2.

Next we looked at what changes oxidation states in this reaction. Each Manganese (Mn) atom goes from +4 to +3, so each Mn gains 1 electron (and together the 2 Mn gain 2 electrons). Gaining electron is reduction, so the Mn are reduced. The oxidation is seen in Hydrogen, where each H atom goes from 0 to +1. Each H loses 1 electron and is oxidized, and the 2 H together lose 2 electrons. Stressed that oxidation and reduction always take place together and what one species loses, the other gains. Also you can’t have more than one species being oxidized and one being reduced.

We also looked at the balanced reaction of zinc in acid:

Zn (s) + 2 HCl (aq) 🡪 ZnCl2 (aq) + H2 (g)

Here each Zinc (Zn) goes from 0 (in Zn) to +2 in ZnCl2 and is oxidized by 2 electrons. Now the H is reduced, going from +1 in HCl to 0 in the hydrogen gas. Chlorine is -1 in each species (some things do not change oxidation states).

The rest of what I covered is pretty much in the Power Point presentation.

Looked at the Daniell Cell, how to have an electrochemical cell you need to separate the oxidation and reduction reactions and force the electrons to flow through a wire.

We looked at an analogy with tanks of water and ranking their energies from highest to lowest. Then we looked at Standard Reduction Potentials which are a series of reactions all written as a reduction and listed from highest potential (voltage) to lowest.

To make a cell, you need both an oxidation and a reduction. The standard potential (voltage) for a galvanic cell is just:  
  
Eocell = Eoreduction- Eooxidation

For the Daniell Cell, Zn(s) + Cu2+(aq) → Zn2+(aq) + Cu(s)  
  
Cu2+ (aq)  +  2e-  🡪  Cu (s) 0.34 V  
Zn2+ (aq)  +  2e-  🡪   Zn (s) –0.76 V  
  
Eocell = 0.34 V – (–0.76 V) = 1.10 V

We did at least one example where I just picked half reactions at random from the table of reduction potentials. We also looked at the silver tree reaction (Lecture 1 demo) as a cell:

Cu (s) + 2 Ag+ (aq) 🡪 Cu2+ (aq) + 2 Ag (s)

Cu2+ (aq)  +  2e-  🡪  Cu (s) 0.34 V

Ag+ (aq) + e-  🡪 Ag (s) 0.80 V

Eocell = 0.80 V – (0.34 V) = 0.46 V

Pointed out that we always get a positive voltage for a spontaneous (Galvanic) cell reaction, and if you get a negative number, you just subtracted the smaller number from the bigger (instead of the other way around).

Galvanic cell – has a positive potential (voltage) and is a spontaneous reaction, used to do electrical work / is a source of energy (an example is a battery)  
  
Anode – electrode where oxidation takes place in a galvanic cell  
   
Cathode - electrode where reduction takes place in a galvanic cell

Electrolytic cell - has a negative potential (voltage) and is **NOT** a spontaneous reaction, consumes electrical work / needs an outside source of energy to run (an example is a battery being charged, or splitting water into hydrogen and oxygen)  
  
Galvanized – a metal (usually iron or steel) that has a layer of zinc (Zn) applied to prevent rusting has been galvanized. The zinc will oxidize before the iron will, protecting the iron.

Demos: (Broken) Potato Clock / Battery, Elephant’s Toothpaste (redox as each O atom goes from -1 peroxide to 0 in oxygen gas)

Activity: Glue plus Borax to make “Glue-p” (silly putty like substance)