Electrochemistry Study Guide

1. Terms to know:

- Electrode Potential/Voltage: The amount of energy per unit of charge that a half-reaction either requires or releases (typically in units of Volts)
- Standard Hydrogen Electrode (SHE): A platinum electrode that provides a surface for the following reaction:

$$2H^+_{(aq)}\ (1M) + 2e^- \rightarrow H_{2(g)}\ (1\ bar)$$

This electrode is arbitrarily assigned a potential of 0.

- Standard Electrode Potential/Voltage: The potential of an electrode in comparison with the SHE.
- Standard Cell potential/Voltage: The difference in standard potential between the cathode and the anode of an electrochemical cell:

$$E°_{cell} = E°_{anode} – E°_{cathode}$$

2. How to calculate the Standard Cell Potential for a redox reaction:

Take the following reaction:

$$Zn_{(s)} + Cl_{2(g)} \rightarrow ZnCl_{2(aq)}$$

a. Assign oxidation numbers:
   - $Zn_{(s)} = 0$
   - $Cl_{2(g)} = 0$
   - $ZnCl_2$:
     - Cl = -1
     - Zn = +2

b. Write oxidation and reduction half-reactions:

$$\text{Red: } Cl2_{(g)} + 2e^- \rightarrow 2Cl^-_{(aq)}$$

$$\text{Ox: } Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$$

c. Look up standard reduction potentials in a reference table:

$$Cl_{2(g)} + 2e^- \rightarrow 2Cl^-_{(aq)}\ E° = +1.358V$$

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-\ E° = -0.763V$$

d. Calculate Cell potential using the cell potential equation:

$$E°_{cell} = E°_{anode} – E°_{cathode}$$

$$E^\circ_{cell} = 1.358 - (-0.763)$$

$$E^\circ_{cell} = 2.121V$$

Notice that the standard cell potentials in the reference table are all reduction potentials, even though you have one reduction half-reaction and one oxidation half-reaction. This is just convention. Generally, the more positive a reduction potential is, the more spontaneous the reduction reaction, so it makes sense that we get negative reduction potentials for your oxidation half-reactions.