Study Guide to Electrochemistry

1. Term to know:

- Faraday Constant: The charge (in Coulombs) of one mole of electrons:

$$F = 96485 \text{ C / mol e-}$$

2. The relationship between ΔG and $E_{cell:}$

$$\Delta G = -nFE_{cell}$$

- Where:
  - ΔG: Change in Gibbs Free Energy
  - n: number of moles of electrons exchanged in the redox reaction
  - F: Faraday's constant
  - E_{cel}l: The general cell potential (not necessarily under standard conditions)

3. How to tell if a reaction is spontaneous from the value of E_{cell}:

- If E_{cell} is positive ($E_{cell} > 0$), then the reaction is spontaneous
- If E_{cell} is negative ($E_{cell} < 0$), then the reaction is not spontaneous

4. How to combine known E° values to find an unknown E° value

Consider the following reaction:

$$Fe^{3+}_{(aq)} + 3e^- \rightarrow Fe_{(s)}$$

This reaction (and its E° value) are not found in the standard reduction potentials table. Instead, we find the following:

1. $Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$ (E° = -0.440 V)
2. $Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$ (E° = 0.771 V)
3. $Fe^{3+}_{(aq)} + 3e^- \rightarrow Fe_{(s)}$ (E° = ?)

We cannot just add the E° values because the number of electrons transferred is different for both equations. Instead, we calculate the ΔG values, which we can then add:

$$\Delta G = -nFE°_{cell}$$

1. $\Delta G = -(2 \text{ mol } e^-)(96485 \text{ C/mol } e^-)(-0.440 \text{ V}) = 0.880F \text{ V}$
2. $\Delta G = -(1 \text{ mol } e^-)(96485 \text{ C/mol } e^-)(0.771 \text{ V}) = -0.771F \text{ V}$

$$\Delta G_{total} = 0.880F \text{ V} – 0.771F \text{ V} = 0.109F \text{ V}$$

We can then use ΔG_{total} to calculate E°:

$$\Delta G_{total} = 0.109F \text{ V}$$

$$-nFE° = 0.109F\ V$$

$$-3FE° = 0.109F\ V$$

$$E° = -0.0363\ V$$