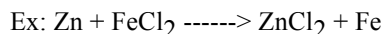# NOTES #59 The *Electrifying* WORLD OF ELECTROCHEMISTRY AP Chemistry

What is electrochemistry? It is the conversion of electrical energy to chemical energy (or vise versa) using REDOX Rxns.

I. **Review of Redox Rxns and 1/2 rxns....**   Ex: $Zn + FeCl_2$ ------> $ZnCl_2 + Fe$ **OX**:

Ox Agent: _____   Red Agent: _____   **RED**:

II. Reduction Potentials

1. Every half reaction has an electric potential or voltage associated with it. Why? Electrons, like heat, travel spontaneously from <u>High Potential</u> (less stable) to <u>Low Potential</u> (more stable). Voltage is a measure of the driving force for electron transfer from high to low potential. The process is analogous to water flow; water flows downhill (high to low potential E), energy release can be harnessed to to produced electricity. Water flowing from a waterfall flows spontaneously downhill with gravity. To get the water in the pool beneath the waterfall to flow back up the waterfall is not spontaneous (but it could be with a pump and an input of energy).

2. STANDARD REDUCTION POTENTIALS TABLE - pg 766 in book.

a. $E^{\bullet}$ at standard conditions - _____ b. All 1/2 reactions are written as _____ rxns. To get $E^{\bullet}$ for the oxidation reaction, just look at reactions backwards and *change the sign* for $E^{\bullet}$. EX: RED: $Li^+ + e^-$ -----> $Li(s)$ $E^{\bullet}$=-3.05V   *or*   OX: $Li(s)$ -----> $Li^+ + e^-$ $E^{\bullet}$=_____

c. Changing stoichiometric coefficients DOES NOT affect the value of $E^{\bullet}$. Why? Electric potentials are _____ properties, independent of # of the times the reaction must occur. It will produce this same voltage for a *longer duration,* however.
EX: $2Li^+ + 2e^-$ ------> $2Li(s)$ $E^{\bullet}$ = _____
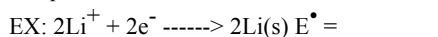
d. The more positive $E^{\bullet}$, the MORE favorable! Q: Which substance on the table is most easily reduced (the best _____ agent)?

Q: Notice which sorts of substances are most easily oxidized? _____ Does this make sense?

e. You can use this magical table as an ACTIVITY SERIES!   EX: Using this table, will Mg (s) displace Fe^{2+} in this single replacement reaction:
$Mg (s) + Fe^{2+}$ ----> $Mg^{2+} + Fe (s)$?

f. How did someone come up with this table? The $E^{\bullet}$ value for the reduction rxn of H^+ to H_2 (g), $2 H^+ + 2e^-$ ----> H_2 (g), has been arbitrarily set to ZERO. Every other $E^{\bullet}$ value was determined *relative* to this.

II. Calculating $E^{\bullet}$ for a Redox Reaction. This involves a THREE step process:

(1) Break rxn into half reactions (2) Balance 1/2 rxns (3) Add the $E^{\bullet}$ values from the two half reactions . $E_T = E_{ox} + E_{red}$

EX: Calculate $E^{\bullet}$ for: $Zn (s) + 2 Ag^+$ ------> $Zn^{2+} + 2 Ag (s)$ $E^{\bullet}$ = ?

OX:

RED:

III. ELECTROCHEMICAL CELLS   EX: **$Zn(s) + Cu^{2+}$ -----> $Zn^{2+} + Cu(s)$ $E^{\bullet}$ = 1.10V**
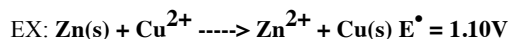
1. Galvanic or Voltaic Cells   OX:

a. Involves a SPONTANEOUS redox rxn. E° value has to be _____.   RED:
b. Batteries are common examples of galvanic cells!!
c. General set-up.

2. Electrolytic Cells

a. Involves a NON-SPONTANEOUS redox rxn. $E^{\bullet}$ value has to be _____.
An outside voltage source is needed to FORCE this non-spontaneous rxn to take place.

b. Electrolysis reactions (Hofmann Apparatus) and electroplating processes (like the Cu plating lab we did....) are examples.

c. General Set-up   EX: **Cu (s) + Cu²⁺ (aq) -------> Cu²⁺ (aq) (g) + Cu**   E°cell = _____   OX:

RED:

d. Common Electroplating Question.
Ex: What mass of Cu could be plated out by electrolyzing $CuCl_2$ (aq) for 16.0 hrs at a constant current of 3.00 amperes.

1. Calculate CHARGE. Remember   Amperes x Seconds = AmpSeconds   What is another name for ampsec? _____!

2. Convert Coulombs to moles of e⁻. Use the Faraday constant as a conversion factor: 1 Faraday ($\mathcal{F}$) or 1 mol of e⁻ = 96,500 Coulombs

3. Using a balanced equation, convert from moles of e⁻ to moles of metal. Convert moles of metal to grams.....Can use determine Molar Mass of a Metal too.....

IV. **E°, ΔG° and Equilibrium....**1. KEY EQUATIONS: (for derivations of equations, read in book - pg 769 - 770)

a. **ΔG° = - nFE°**

$G°$ = Gibb's Free Energy (kJ/mol)   F = Faraday's constant: 1 mol of e⁻ has a total charge of 96,500 Coulombs
n = # of e⁻ (in moles) transferred in redox rxn $E°$ = Standard rxn potential (V)

b.   **E° = $\dfrac{RT \ln K}{nF}$**   R=8.31J/mol·K   T = Temp in K
                                                         K = Equilibrium constant   n, F, and E° are same as above!

OR

K = nE°/0.0592

2. Effect of changing concentrations on E?
a. Up until now, all cell potentials have been at standard conditions of 1M solutions. What if we are NOT using a 1M soln?

EX: Consider the following: Cu (s) + 2 Ce⁴⁺ → Cu²⁺ + 2 Ce³⁺ E° = 1.36 V
If $[Ce^{4+}]$ is INCREASED, what will happen to E? *THINK about Le Chatelier's Principle!*

If $[Cu^{2+}]$ and $[Ce^{3+}]$ were INCREASED, E would _____.

b. You can calculate E at non-standard conditions by using the NERST EQUATION.   **E = E° $\dfrac{-0.0592 \ln Q}{n}$**

EX: Consider the following: 2 Al (s) + 3 Mn²⁺ -----> 2 Al³⁺ + 3 Mn (s)   E° = 0.48 V

If $[Mn^{2+}]$ = 0.5 M and $[Al^{3+}]$ = 1.5 M, what would E be? _____ ☐

         *** Notice what happened to E? It _____. Is this consistent with what you would expect from LeChatelier?

                    This Is The End! YEAH!! Whoopee! Yeeeee-Haw!