Chapter Summary: Electrochemistry

Oxidation-Reduction: Oxidation is loss of electrons and reduction is gain of electrons.
Half Reactions: show oxidation and reduction in individual reactions.
Redox reactions in acid and basic solutions and balancing redox reactions.

Galvanic/Voltaic Cells: A cell that has a spontaneous redox reaction. The terminology of a cell are:

- Electrode – a metal piece at which the electrochemical reaction takes place.
- Anode: where oxidation (loss of e^-) occurs.
- Cathode: where reduction (gain of e^-) occurs.
- Half-cell: the reduction or the oxidation part of the cell.
- Coulomb (C): the unit of electric charge.
- Volt (V): one joule per coulomb.
- Voltmeter: measures volts.
- Cell potential (E_{cell}): electro-potential difference that moves the electrons from the anode to cathode.

Cell Diagram: anode on the left side and cathode on the right. Single line (|) to separate electrode and solution and double line (||) to separate half cells.
$Al(s) \mid Al^{3+}(aq) \parallel H^+(g) \mid Pt(s)$

Standard Hydrogen Electrode Potential (SHE): the potential when a 1M solution of H^+ is reduced to H_2 gas at one atmosphere on Platinum electrode. At this point the electro-potential $E^o = 0$ V. Equilibrium is established between H^+ and H_2 gas.
Standard Electrode Potential – the tendency for reduction to occur at an electrode under the conditions of SHE.
Standard Cell Potential: $E^o_{cell} = E^o \text{ (cathode)} - E^o \text{ (anode)}$
$E^o_{cell} = E^o \text{ (right)} - E^o \text{ (left)}$

Activity Series and E^o_{cell} values (table 19.1): Note that the top element is the most powerful oxidizing agent (i.e. causes other elements to lose electrons), while the lowest one is the most powerful reducing agent (i.e. causes other elements to gain electrons).
So the element above e.g. Cu, will get displaced by an element below it e.g. Zn.
OR you can say Cu will get reduced by Zn. So Cu will gain electrons in presence of Zn but will cause Fe to gain electrons

Effect of concentration on cell voltage: Nerst equation: used for non-standard electrode potential calculations. It also explains why voltage of cell changes over time or with change in concentration.
$E_{cell} = E^o_{cell} - \frac{0.0592V}{n} \log Q$ where n = moles of electrons, Q = reaction quotient = $\frac{[\text{products}]}{[\text{reactants}]}$

Applications: (read on your own if not covered in class)

1. Batteries: the dry cell, lithium battery, lead-acid battery, fuel cells.
2. Corrosion chemistry: e.g. of iron.
3. Electrolysis: non-spontaneous voltaic cell.
4. Electroplating: application of electrolysis, helps to protect certain metals against corrosion.