**Oxidation – Reduction Reactions**

**aka *Redox* Reactions**

Chemistry – AP

*Basic Principles:*

* *oxidation – lose electrons*
* *reduction – gain electrons*
* ***LEO*** *the lion goes* ***GER*** *(****L****=lose* ***E****=electrons* ***O****=oxidized;* ***G****=gain* ***E****=electrons* ***R****=reduced)*
* *oxidizing agent becomes reduced*
* *reducing agent becomes oxidized*

**The Half – Reaction Method for Balancing Equations for Oxidation – Reduction Reactions Occurring in Acidic Solution**

1. Write separate equations for the oxidation and reduction half – reactions.
2. For each half – reaction,
	1. Balance all the elements except hydrogen and oxygen.
	2. Balance oxygen using H2O.
	3. Balance hydrogen using H+.
	4. Balance the charge using electrons.
3. If necessary, multiply one or both balanced half reactions by an integer to equalize the number of electrons transferred in the two half – reactions.
4. Add the half reactions, and cancel identical species.
5. Check that the elements and charges are balanced.

**Examples in Acidic Solution**

MnO4- (aq) + Fe2+ (aq) → Fe3+ (aq)+ Mn2+ (aq)

H+ (aq) + Cr2O72- (aq) + C2H5OH (l) → Cr3+ (aq) + CO2 (g) + H2O (l)

**The Half – Reaction Method for Balancing Equations for Oxidation – Reduction Reactions Occurring in Basic Solution**

1. Use the half – reaction method as specified for acidic solutions to obtain the final balanced equation as *if* H+ ions were present.
2. To both sides of the equation obtained above, add a number of OH- ions that is *equal* to the number of H+ ions*. (We want to eliminate H+ by forming H2O.)*
3. Form H2O on the side containing both H+ and OH- ions, and eliminate the number of H2O molecules that appear on both sides of the equation.
4. Check that the elements and charges are balanced.

**Examples in Basic Solution**

Ag (s) + CN- (aq) + O2 (g) → Ag(CN)2- (aq)

Al (s) + MnO4- (aq) → MnO2 (s) + Al(OH)4- (aq)