Redox Reactions

Terminology and Units

- *Oxidation* – To lose electrons
- *Reduction* - To gain electrons
- *Oxidizing agent* - causes the other reactant to oxidize; it is itself reduced.
- *Reducing agent* - causes the other reactant to reduce; it is itself oxidized.
- *Oxidation state (or number)* - A "virtual charge" on an atom within a molecule; used to track changes in electrons.

Units of electron transfer

- 96,500 *Coulombs* (C) = 1 mole of electrons
  - ▹ 96,500 is *Faraday's Constant* (F)
- 1 *Ampere* (Amps, A) = 1 C/s

Oxidation State ("OS") Rules

▸ Uncombined element: 0	*e.g.*, O_2, Fe
▸ Sum of OS in a neutral species is 0 and in an ion is equal to the charge.	*e.g.*, H_2SO_4, CO_3^{2-}, Na^+
▸ Group 1 metals, +1; Group 2 metals, +2	*e.g.*, NaCl, $BaCl_2$
▸ Fluorine in compounds: –1	*e.g.*, BaF_2
▸ Oxygen: –2 in most covalent compounds	*e.g.*, Na_2O
▹ Exception: peroxides, in which oxygen's OS is –1	*e.g.*, H_2O_2
▹ Exception: Superoxides, in which oxygen's OS is –½	*e.g.*, KO_2
▸ H in compounds: +1 in covalent compounds with nonmetals	*e.g.*, H_2S
▸ Binary metallic compounds, Group 15: –3; Group 16: –2; Group 17, –1	*e.g.*, Na_3P, H_2S, SrF_2

Balancing Redox Equations

Acid & Basic solutions

Note that the balancing process is nearly identical for balancing redox reactions in acidic and basic solutions; the basic solutions add two extra steps.

1. Write the equations for the oxidation and reduction half reactions.
2. For each half reaction:
   - Balance all the elements except H and O.
   - Balance oxygen using H_2O
   - Balance hydrogen using H^+
   - Basic solutions: Neutralize the H^+ by adding OH^- to both sides of the reaction
   - Basic solutions: Combine H^+ and OH^- to make H_2O
   - Balance the charge using electrons
3. Balance the electrons lost and gained by multiplying the half-reactions by integers as necessary
4. Add the half-reactions, cancelling items that appear on both sides.