Redox Reactions

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 with Redox Half-Reactions

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^-
   - 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 an integer as necessary
4. Add the half-reactions, cancelling items that appear on both sides.