Reduction/Oxidation Reactions

Terminology

- *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 in a molecule used to track changes in electrons.

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
▸ 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

Activity Series

- An *activity series* is a list of metals sorted in order from most to least reactive.

  e.g., K, Na, Sr, Mg, Al, Mn, Fe, Ni, Pb, Hg, Au
- The most reactive metals will react with cold water to produce H_2 and a hydroxide.
- Mid-range metals don't react with water, but do react with acids.
- Least reactive metals don't react with very many things at all (and are often found in pure state in nature.)
- More reactive metals
  - Will oxidize more readily than less reactive metals (and are therefore stronger reducing agents)
  - In single-replacement reactions will displace metals of lower reactivity

Half Reactions

A half reaction represents the reduction or oxidation that takes place in a reaction. It includes electrons as part of the equation.

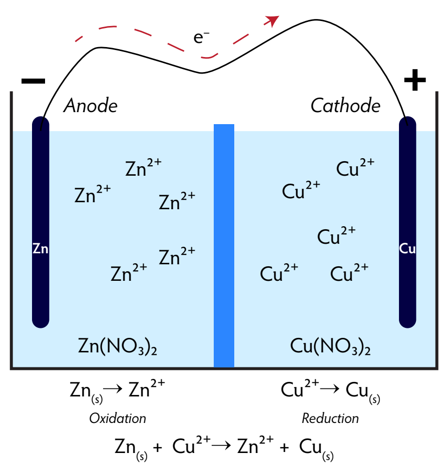
$$Sn^{2+} \rightarrow Sn^{4+} + 2e^- \quad \text{Oxidation half-reaction}$$

Balancing chemical reactions with half reactions

This method applies only to reactions happening in an acidic solution.

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^+
   - 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.

Electrochemistry and Batteries



- You create a battery by placing a reducing and oxidizing agent in separate containers connected by a wire.

  Properly, the whole battery is called an *electrochemical cell*.

  Each of the two containers is called a *half-cell*.
- Each half-cell has a solid strip of metal (an *electrode*) immersed in an aqueous solution of a compound of the same metal.
  - ▹ The half-cells are separated by a membrane or "bridge" that will allow ions to pass from one side to the other if they are sufficiently motivated.
- Electrons flow through the wire from the oxidizing side (that gives up electrons) to the reducing agent (that grabs them).
  - ▹ The electrode on the oxidizing side is called the *anode*.

    It is conventionally labelled with a minus sign because it supplies electrons to the current.
  - ▹ The electrode on the reducing side is called the *cathode*.

    It is conventionally labelled with a plus sign because it receives electrons from the current.
  - ▹ Thus, electrons flow from the anode to the cathode.