Oxidation Numbers

The oxidation number of an element is the charge the element would have if it were an ion. It helps us keep track of electrons in an oxidation-reduction reaction; it may be real or make-believe.

Oxidation: The process whereby the oxidation number of an element increases
- Becomes more positive
- Involves the loss of electrons
- Electrons are a product
  - $M \rightarrow M^{n+} + ne^-$
  - $X^- \rightarrow X + e^-$
  - $M^{2+} \rightarrow M^{3+} + e^-$

Reduction: The process whereby the oxidation number of an element decreases
- Becomes more negative
- Involves the gain of electrons
- Electrons are a reactant
  - $M^{n+} + ne^- \rightarrow M$
  - $X_2 + 2e^- \rightarrow 2X^-$
  - $M^{4+} + 2e^- \rightarrow M^{2+}$

Oxidation – reduction reactions
- Called “redox” reactions for short
- Always occur as a pair
- One element “loses” electrons
  - Oxidation
- One element “gains” electrons
  - Reduction

Redox reaction examples:
- $Zn^+ + Cu^{2+} \rightarrow Zn^{2+} + Cu^+$
- $2 \, Cl^- + F_2^+ \rightarrow 2 \, F^- + Cl_2^+$

Determining oxidation numbers
1. The oxidation number of a free element is zero, regardless of if it is “monatomic” or if it has a subscript.
   - Examples: Mg, O$_2$, P$_4$, Zn
2. The oxidation number of a “monatomic” ion is the same as the charge of the ion
   - Na$^+$ has an ox# of +1
   - S$^2$ has an ox# of -2
   - Fe$^{3+}$ has an ox# of +3
   - I$^-$ has an ox# of -1
3. The sum of all the oxidation numbers of all the elements in a substance is the same as the charge of the substance.
   - The ox#’s in a neutral compound must all add up to zero
   - The ox#’s in a polyatomic ion must all add up to the charge of the polyatomic ion

In a compound...
4. The ox# of fluorine is -1
5. The ox# of hydrogen is +1
   - except in a hydride, where it is -1
   - ex: LiH, CaH$_2$
6. The ox# of oxygen is -2
   - except in peroxides and superoxides
   - ex: H$_2$O$_2$, KO$_2$, OF$_2$