Oxidation and Reduction (a.k.a. Redox)- Chapter 20

20.1 Oxidation-Reduction Reactions

Oxidation- process by which substances lose electrons (or old and more specific, gain oxygen)

Reduction- process by which substances gain electrons (or old and more specific, lose oxygen); the oxide compound loses oxygen, so there is a *reduction* in mass

Cool mnemonics-

1) "**LEO** the lion says **GER**!" **L**ose **E**lectrons **O**xidation, **G**ain **E**lectrons **R**eduction
2) **OIL RIG**- **O**xidation **I**s **L**oss of electrons; **R**eduction **I**s **G**ain of electrons

Redox reactions- reactions in which electrons are transferred between the reactants; reduction and oxidation reactions occur simultaneously

Oxidation numbers- it is the charge of an atom if the electrons within a bond belonged to the most electronegative atom; it helps us balance redox reactions and identify species being oxidized and reduced; *these are not actual charges!*

Rules for Assigning Oxidation Numbers:

(*Notice: oxidation numbers are written with the charge before the number, as opposed to ions where the charge follows the number*)

1. The oxidation number of an uncombined element is zero.
2. The oxidation number of any monatomic ion equals its ionic charge.

3. In compounds, the oxidation number of many elements corresponds to the element's position in the periodic table:
    a. Elements in Group 1A are always +1.
    b. Elements in Group 2A are always +2.
    c. Aluminum is always +3.
    d. Fluorine is always -1.
    e. Hydrogen has an oxidation number of +1 when combined with nonmetals.
    f. Oxygen has an oxidation number of -2 in most compounds and ions.
4. The oxidation numbers of elements in compounds are written per atom.
5. The algebraic sum of the individual oxidation numbers of all the atoms in the formula for a compound is zero.
6. The algebraic sum of the individual oxidation numbers of all the atoms in the formula for a polyatomic ion is equal to the charge of the ion.

**Sample and practice problems on page 661*

- ***Oxidation*** has occurred when the oxidation number increases (becomes more positive)
- ***Reduction*** has occurred when the oxidation number decreases (becomes more negative)

Oxidizing agent- oxidizes another substance by accepting electrons by that substance (high electronegativity or high positive oxidation numbers); it itself is being reduced
Reducing agent- reduces another substance by donating electrons to that substance (low electronegativity- highly active metals); it itself is being oxidized

**Sample problems, practice problems, and section review on pp. 664-665*

20.2 Types of Redox Reactions

- **Combination and decomposition reactions** are types of Redox reactions
- **Single-replacement reactions** is a type of Redox reaction (either a free element becomes an ion *or* an ion become a free element)

Ex. Copper metal is placed into a solution of silver nitrate

$$\mathbf{Cu_{(s)} + 2AgNO_{3\,(aq)} \rightarrow Cu(NO_3)_{2\,(aq)} + 2Ag_{(s)}}$$

↓oxidized

$$\mathbf{Cu_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}}$$ (net ionic equation)

↑reduced

Activity series of metals (p. 668)- a rank order list of metals according to their relative reactivity (tendency to give up electrons or *oxidize*); metals most easily oxidized replace those less willing to release electrons [some metals are so reactive they replace hydrogen from water (e.g. Na, K, Ca, Li)]; hydrogen in the series serves as a reference by which metals are compared

**Section review on page 669*

20.3 Applications of Redox

Batteries
Corrosion
Bleaching
Fuels and Explosives
Photography

20.4 Balancing Redox Reactions

Redox reactions in aqueous solutions can be balanced by using half-reactions (a.k.a. half-reaction method); the number of electrons lost in the oxidation process must equal the number of electrons gained in the reduction process

Steps for balancing redox reactions in *acidic solutions*:
(*Notice: Steps 2b and 2c are skipped if the redox reaction doesn't occur in an acidic solution*)

1) Write both the reduction and oxidation half-reactions separately
2) Then:
    a) Balance elements (except H & O)
    b) Balance oxygen using H_2O
    c) Balance hydrogen using H^+
    d) Balance the overall charge on both sides of the half-reactions by using electrons
3) Make the number of electrons equal in both half-reactions [if needed, multiply a number through the half-reaction(s)]
4) Add half-reactions together and cancel out identical species (elections should cancel out, due to step 3)
5) Verify your answer by checking if it's in fact balanced

**Sample problems, practice problems, and section review on pp. 676-677*