4.8 Oxidation- Reduction Reactions

Reactions in which one or more electrons are transferred.

Ex. 1: Ionic $2Na_{(s)} + Cl_{2(g)} \rightarrow 2\ NaCl_{(s)}$

Ex. 2: Non-ionic, Combustion (or rapid oxidation) $C_2H_{6(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$ + heat

Rules for Assigning Oxidation States Numbers: Table 4.2

Note: In a covalent bond, the electrons are shared "equally", however, we will assign the more electronegative atom the negative charge.

Charge is assigned to individual atom (ion), not the group. Ex. H_2O, each H = +1

**Convention; actual charges are written as n^+ or n^-. While oxidation states (not actual charges) are written ^+n or ^-n, the number after the + or – sign.*

Rules: Steven Zumdahl, 4th Edition Houghton Mifflin

Rules for Assigning Oxidation States

1. The oxidation state of an atom in an element is 0. For example, the oxidation state of each atom in the substances Na(*s*), $O_2(g)$, $O_3(g)$, and Hg(*l*) is 0.
2. The oxidation state of a monatomic ion is the same as its charge. For example, the oxidation state of the Na^+ ion is +1.
3. Oxygen is assigned an oxidation state of −2 in its covalent compounds, such as CO, CO_2, SO_2, and SO_3. The exception to this rule occurs in peroxides (compounds containing the O_2^{2-} group), where each oxygen is assigned an oxidation state of −1. The best-known example of a peroxide is hydrogen peroxide (H_2O_2).
4. In its covalent compounds with nonmetals, hydrogen is assigned an oxidation state of +1. For example, in the compounds HCl, NH_3, H_2O, and CH_4, hydrogen is assigned an oxidation state of +1. w/ metals = H^{-1}
5. In binary compounds the element with the greater attraction for the electrons in the bond is assigned a negative oxidation state equal to its charge in its ionic compounds. For example, fluorine is always assigned an oxidation state of −1. That is, for purposes of counting electrons, fluorine is assumed to be F^-. Nitrogen is usually assigned −3. For example, in NH_3, nitrogen is assigned an oxidation state of −3; in H_2S, sulfur is assigned an oxidation state of −2; in HI, iodine is assigned an oxidation state of −1; and so on.
6. The sum of the oxidation states must be zero for an electrically neutral compound and must be equal to the overall charge for an ionic species. For example, the sum of the oxidation states for the hydrogen and oxygen atoms in water is 0; the sum of the oxidation states for the carbon and oxygen atoms in CO_3^{2-} is −2; and the sum of oxidation states for the nitrogen and hydrogen atoms in NH_4^+ is +1.

Special Cases: oxygen usually O^{2-}, except in peroxides, O_2^{2-} is O^{-1}, Ex. Na_2O_2

Hydrogen H^{+1}, except with metal hydrides, H^{-1}. Ex. LiH

Fractional Charges are possible. Ex. Fe_3O_4, iron pyrite Fe= $^+8/3$, 2 Fe = $^+3$, 1 Fe= $^+2$

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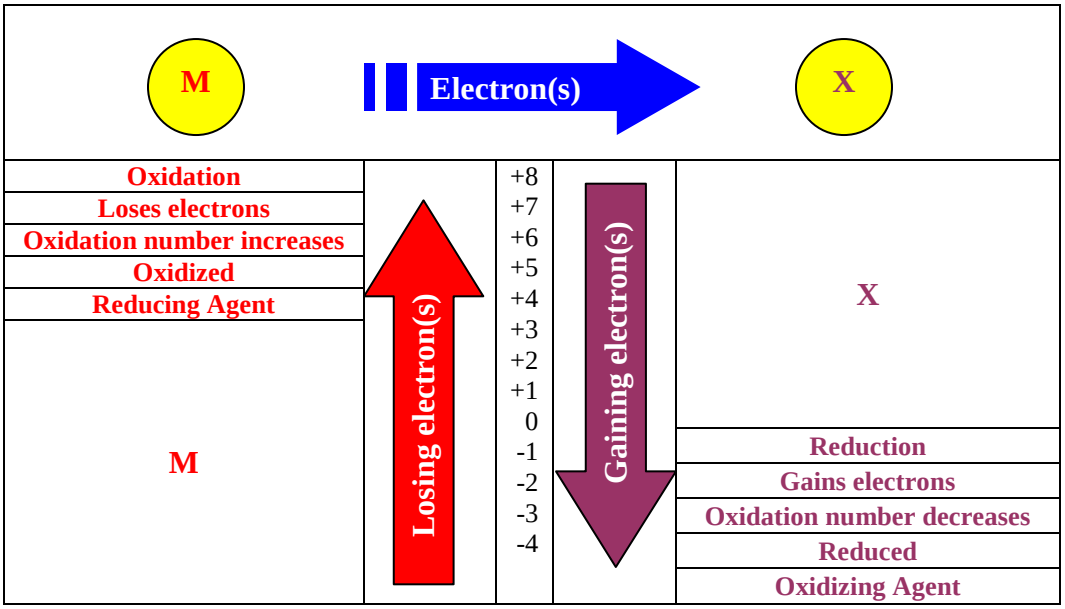
Characteristics of Redox (oxidation reduction reactions)

OIL RIG: **O**xidation **I**nvolves **L**oss (of electrons), **R**eduction **I**nvolves **G**ain (of electrons).



Ex. Thermite (Glencoe Video)

$Fe_2O_{3(s)} + 2\ Al_{(s)} \rightarrow Al_2O_{3(s)} + 2\ Fe_{(l)}$

Each $Fe^{+3} \rightarrow Fe^0$, gains 3 electrons, reduced, oxidizing agent

Each $Al^0 \rightarrow Al^{3+}$, loses 3 electrons, oxidized, reducing agent

Quick Glance: $NO_3 \rightarrow NO_2$, note fewer oxygen atoms with nitrogen, N is being reduced.

In Class Practice: P. 174 # 67, 70, 71, 72