Oxidation-Reduction Reactions

Oxidation-reduction (redox) reactions are reactions in which oxidation numbers change. *Oxidation numbers* are either real charges or formal charges which help chemists keep track of electron transfer. In practice, oxidation numbers are best viewed as a bookkeeping device.

- Oxidation cannot occur without reduction.
- In a redox reaction, the *substance oxidized* contains atoms which *increase* in oxidation number. Oxidation is associated with *electron loss* (helpful mnemonic: ***LEO*** = *Loss of Electrons, Oxidation*).
- The *substance reduced* contains atoms which *decrease* in oxidation number during the reaction. Reduction is associated with *electron gain* (helpful mnemonic: ***GER*** = *Gain of Electrons, Reduction*).
- An *oxidizing agent* is a substance which *oxidizes something else*: it itself is reduced! Also, a *reducing agent* is a substance that *reduces another reactant*: it itself is oxidized.
- A *disproportionation reaction* is a reaction in which the *same element* is *both* oxidized and reduced.

How to Assign Oxidation Numbers: The Fundamental Rules

- The oxidation number of any *pure element* is *zero*. Thus the oxidation number of H in H_2 is zero.
- The oxidation number of a *monatomic ion* is equal to its *charge*. Thus the oxidation number of Cl in the Cl^- ion is -1, that for Mg in the Mg^{+2} ion is +2, and that for oxygen in O^{2-} ion is -2.
- The *sum* of the oxidation numbers in a compound is *zero if neutral*, or *equal to the charge if an ion*.
- The oxidation number of alkali metals in compounds is +1, and that of alkaline earths in compounds is +2. The oxidation number of F is -1 in all its compounds.
- The oxidation number of *H is +1 in most compounds*. Exceptions are H_2 (where H = 0) and the ionic hydrides, such as NaH (where H = -1).
- The oxidation number of *oxygen (O) is -2 in most compounds*. Exceptions are O_2 (where O = 0) and peroxides, such as H_2O_2 or Na_2O_2, where O = -1.
- For other elements, you can usually use the *sum rule* above to solve for the unknown oxidation number.

Examples:

NO(g) has O = -2, so N = +2.
NO_2(g) has two oxygen atoms and each has O = -2. Thus N + 2(-2) = 0, so N = +4.
SO_4^{2-} has O = -2. Thus S + 4(-2) = -2. Solving the equation gives S = -2 + 8 = +6.
$K_2Cr_2O_7$ has K = +1 and O = -2. Thus 2(+1) + 2 Cr + 7(-2) = 0; 2 Cr = 12; Cr = +6.

Recognizing Oxidation-Reduction Reactions

Oxidation-reduction reactions are reactions in which one type of atom *increases* in oxidation number (*is oxidized*) and another type of atom *decreases* in oxidation number (*is reduced*). Thus to show that a reaction is a redox reaction, you need to calculate oxidation numbers for the atoms in the reactants and products, and document that changes are taking place. There are, however, a few useful generalizations.

- A large number (but not all!) of oxidation-reduction reactions contain *one or more reactants or products which are pure elements.* Why is this true? Also, *all* electrochemical reactions are redox reactions.
- Most acid-base reactions and most precipitation reactions are not redox reactions. Why? Give some examples!