Activity Series

An **activity series** is a list of metals arranged in order of decreasing ease of oxidation. The metals at the top of the table, especially the alkali and alkali earth metals, are the most easily oxidized. They are called active metals. The metals at the bottom of the activity series, especially the transition metals, are very stable and are called noble metals because of their low reactivity.

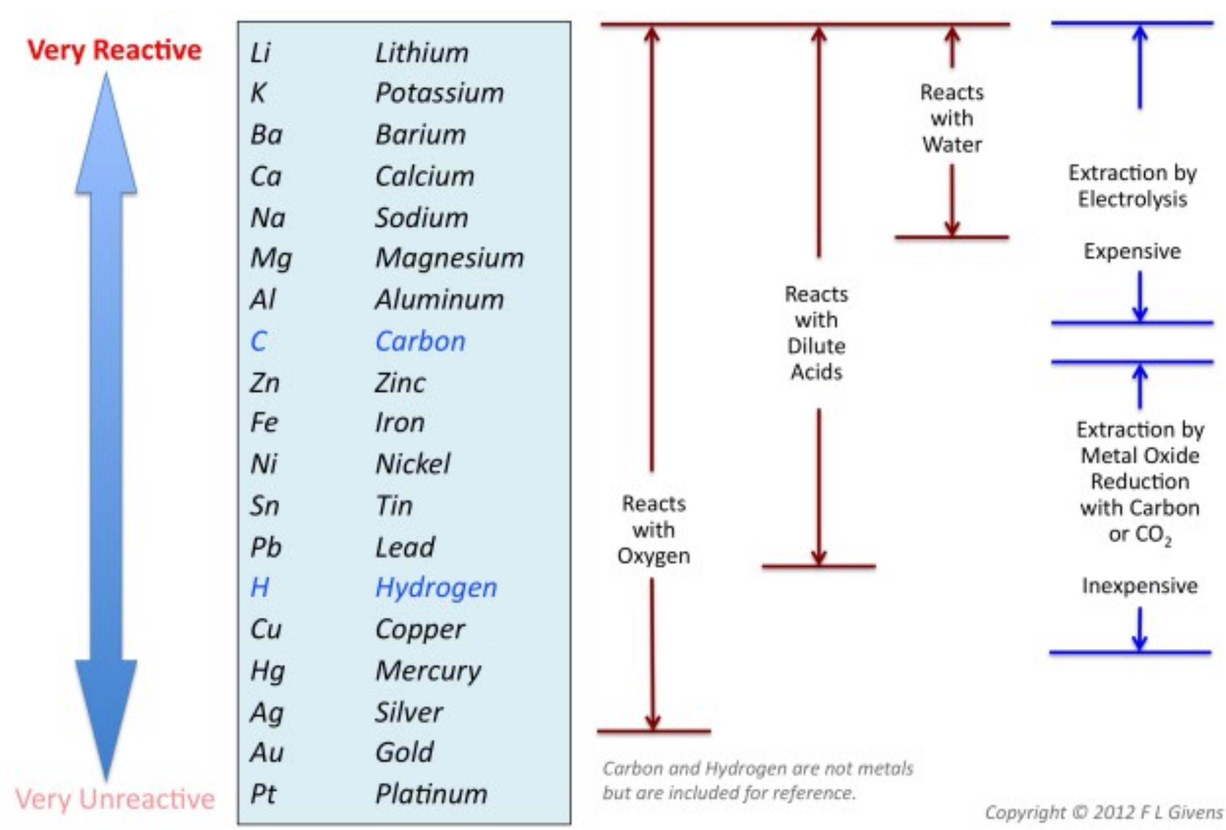
The activity series can be used to predict the outcome of certain reactions. Any metal on the list can be oxidized by the ions of the element below it. For example: copper is above silver in the activity series, and thus will be by oxidized by the silver ions.

$Cu (s) + 2Ag^{+} (aq) \rightarrow Cu^{2+} (aq) + 2Ag(aq)$

In this reaction, copper is oxidized and silver is reduced.
Definition according to above reaction

Oxidation – losing electrons
Reduction – gaining electrons

In another way of defining oxidation one compound will lose oxygen (reduced) and the other compound will gain oxygen(oxidized)



Oxidation Reduction Reactions

redox reaction between copper and silver

Oxidation-Reduction Reactions, sometimes known as redox reactions, are chemical reactions which involve a change in the oxidation number of two or more atoms. These reactions are often found as reactions which are used as sources of heat or work, such as the burning of natural gas, the breakdown of sugar in our bodies, and the corrosion of metal. Redox reactions involve a transfer of one or more electrons and thus the change in oxidation numbers.

Oxidation occurs when the oxidation number increases, when it loses electrons. It becomes more positively charged. **Reduction** occurs when the oxidation number decreases, when it gains electrons. It becomes more negatively charged.
TIP: a good way to remember which process loses electrons and which process gains electrons is this mnemonic device: LEO the lion goes GER.
LEO: Lose Electrons Oxidation
GER: Gain Electrons Reduction

ALSO: When stumbling across the words "cathode" and "anode", remember this. The cathode is the solution that reduces. The anode is the the solution that oxidizes. To remember this, Anode and Oxidize start with a vowel.

Oxidation Numbers

An **oxidation number** of an atom is the charge that would be present on an atom if the compound was made of ions, regardless of whether the compound actually contained ions. Oxidation numbers were developed as a way to keep track of the electrons gained or lost by substances in a redox reaction. These are a set of rules which help assign oxidation numbers:

- An atom in its elemental form, a neutral substance containing atoms of only one element, has an oxidation number of zero.
  - example: each atom of H_2, P_4, or Al has an oxidation number of 0
- A monatomic ion has the oxidation number that is equal to its charge.
  - example: K+ has an oxidation number of +1, S2- has an oxidation number of -2.
  - alkali metals always have a +1 charge, and always a +1 oxidation number
  - alkali earth metals always have a +2 charge, and always have a +2 oxidation number
- Non-metals usually have a negative oxidation number. However, sometimes they can be positive
  - Oxygen usually has an oxidation number of -2, with the exception of peroxides where oxygen has an oxidation number of -1.
  - Hydrogen has an oxidation number of +1 when it is bonded to nonmetals and -1 when bonded to metals.
  - Fluorine always has an oxidation number of -1. All other halogens also have an oxidation number of -1, except when combined with oxygen, they have a positive oxidation number.

- The sum of the oxidation numbers of all atoms in neutral compounds is equal to 0. The sum of oxidation numbers in an ion is equal to the charge of the ion.

Examples: Determine the oxidation number of sulfur in each of the following.

H_2S	When bonded to a non-metal, hydrogen has an oxidation number of +1. Because H_2S compound is neutral, the sum of the oxidation numbers must be 0. Since there are two hydrogen atoms, that amounts to a +2 charge. To balance the charge out, Sulfur must have an oxidation number of -2.
S_8	This is the an elemental form of sulfur, and therefore its oxidation number must be 0.
SCl_2	The sum on the oxidation numbers must equal 0 because it is a neutral compound. Chlorine is a halogen, so its oxidation number is -1. Since there are two Chlorine atoms, their charge adds up to -2. To balance the oxidation numbers, the oxidation number of sulfur must be +2.
Na_2SO_3	The sum of the oxidation numbers must equal 0 because it is a neutral compound. Sodium is an alkali metal so its oxidation number is +1. There are two sodium atoms, so the oxidation number of sodium adds up to +2. Oxygen has an oxidation number of -2, and because there are three oxygen atoms, the oxidation number of oxygen adds up to be -6. To balance out the +2 from the sodium atoms and the -6 from oxygen atoms, the oxidation number of sulfur must equal +4.
SO_4-2	Because this is an ion, the sum of the oxidation numbers must equal -2. Oxygen has an oxidation number of -2 and because there are four oxygen atoms, the total oxidation number of oxygen adds up to be -8. In order for the total oxidation number for the entire compound to equal -2, the oxidation number of sulfur must equal +6.

Kinds of Redox reactions

Combination- reactions that involve combining two elements for form a chemical compound
Formation of water from hydrogen and oxygen gas:

$2H_2(g) + O_2(g) \text{ --> } 2H_2O(l)$

- On the reactant side, both hydrogen and oxygen have oxidation numbers of 0 because they are elements in their elemental form.
- On the product side, each hydrogen atom has an oxidation number of +1, and each oxygen has an oxidation number of -2. Hydrogen is oxidized and oxygen is reduced.

Decomposition- A compound that is broken down into the components from which it was formed.

Decomposition of $KClO_3$ to KCl and oxygen gas:

$2KClO_{3(aq)}$ --> 2KCl + $3O_2$

On both the reactant and product side, potassium has an oxidation number of +1. On the reactant side, chlorine has an oxidation number of +5 and oxygen has an oxidation number of -2 for each atom. On the product side, chlorine has an oxidation number of -1 and oxygen has an oxidation number of 0 because it is in its elemental form. Thus, chlorine is reduced and oxygen is oxidized.
Single Displacement- an element replaces another in a compound. The replacing ion is always oxidized.

Displacement of hydrogen gas by iron:

2Fe + 6HCl --> $2FeCl_3$ + $3H_2$

On the reactant side, iron has an oxidation number of 0, hydrogen has an oxidation number of +1, and chlorine has an oxidation number of -1.

On the product side, Fe has an oxidation number of +3, chlorine's oxidation number is still -1 for each atom, and the oxidation number of hydrogen is 0. Thus, iron is oxidized and hydrogen is reduced.

Balancing Oxidation Reduction Reactions

Just as when balancing any other chemical equation, the law of conservation of mass must be observed: the amount of each element must be the same on both sides of the equation. If a substances loses a certain number of electrons, another substance must gain that same number of electrons.

It is often helpful to consider the oxidation and reduction processes separately, although they take place at the same time. These two separate reactions are called **half reactions**. The total number of electrons lost in one half reaction must equal those gained in the other half reaction.
For example:

Sn2+ (aq) + 2Fe3+ (aq) --> Sn4+ (aq) + 2Fe2+ (aq) can can be broken down into two half reactions

Sn2+ --> Sn4+ + 2e- (oxidation)
2Fe3+ + 2e- --> 2Fe2+ (reduction)

The number of electrons lost in the oxidation reaction equals that gained in the reduction process.

Steps for balancing Redox equations using half reactions

(using $Cr_2O_7^{2-}$ + Cl- --> Cr3+ + Cl_2 in acidic solution)

1. Assign oxidation numbers to see which atoms are being oxidized (losing electrons) and which are being reduced (gaining electrons)

$Cr_2O_7^{2-}$:

- The total oxidation number must add up to be -2 because -2 is the overall charge.

- Oxygen has an oxidation number of -2, and there are 7 atoms so the total oxidation number for all the oxygen atoms is -14 (-2 * 7 = -14).

- To make the total oxidation number equal to -2, the total oxidation number for all chromium atoms must equal +12 (-14 + 12 = -2). Since there are two atoms of chromium in the ion, each atom has an oxidation number of +6.

- Cl^- : Because it is a monatomic ion, the oxidation number is equal to its charge, so the oxidation number of Cl^- is -1.
  Cr^{3+} : This is also a monatomic ion, so its oxidation number is +3
  Cl_2 : Because it is in its elemental form, the oxidation number of Cl_2 is 0

- Thus, each chromium atom is reduced by 3 electrons (+6 + 3e- = +3), and each chlorine atom is oxidized by 1 electron (-1 = 0 + 1e-)

2. Divide the equation into two half reactions

$Cr_2O_7^{2-} \rightarrow Cr^{3+}$ (reduction)

$Cl^- \rightarrow Cl_2$ (oxidation)

3. Balance each half-reaction
    - balance elements rather than H and O
        - $Cr_2O_7^{2-} \rightarrow 2Cr3$
        - $2Cl^- \rightarrow Cl_2$
    - balance the O atoms by adding H_2O
        - $Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$
    - balance the H atoms by adding H^+
        - $14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$
    - balance the charge by adding e-
        - $14H^+ + Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
        - $2Cl^- \rightarrow Cl_2 + 2e^-$

4. Multiply the half reactions by integers necessary so the number of electrons lost by the oxidation process equals those gained by the reduction process.

$14H^+ + Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

$3(2Cl^- \rightarrow Cl_2 + 2e^-) = 6Cl^- \rightarrow 3Cl_2 + 6e^-$

5. Add the two half reactions together and cancel anything appearing on both sides of the combined equation

$14H^+ + Cr_2O_7^{2-} + 6Cl^- \rightarrow 2Cr^{3+} + 7H_2O + 3Cl_2$

6. Check to make sure the atoms of each species is equal on both sides; make sure the charge is equal on both sides.
The above reaction took place in acidic solution so H_2O and H^+ were added in order to balance the O's and H's.

- In a basic solution, the reaction requires H_2O and OH^- instead of H_2O and H^+.

- To balance this, add H+ as if it was an acidic solution, then add the same number of OH- atoms to the solution on both sides as there are H+ atoms. The H+ and OH- atoms will react to form H_2O