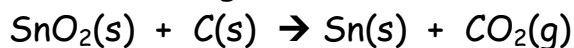


# Oxidation-Reduction Reactions

## Redox

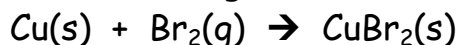
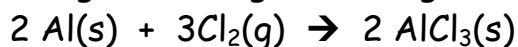
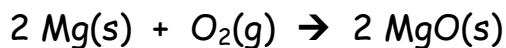
The technology of metallurgy has been has allowed humanity to progress from the Stone Age, through the Bronze age, and the Iron Age, to modern times. Very few metals (such as Au & Ag) exist as pure elements; most exist in a variety of compounds mixed with other substances in rocks called ores.

Although the technological processes of refining vary from one metal to another, the processes of refining involve a large volume of ore that is reduced to a smaller volume of metal. From metallurgy, the term reduction came to be associated with producing metals from their compounds. The production of iron, tin, and copper metals are typical examples:



Another substance, called a reducing agent, causes or promotes the reduction of a metal compound to an elemental metal.

Before metallurgy, humans discovered fire. The technology of fire has been crucial in the development of human cultures, but only relatively recently (18<sup>th</sup> century) have we come to realize the role of oxygen in burning. Understanding the connection of corrosion (rusting, tarnishing, etc.) and burning is an even more recent development. Reactions of substances with oxygen, whether they were the explosive combustion of gunpowder, the burning of wood, or the slow corrosion of iron came to be called oxidation. It soon became apparent that oxygen was not the only substance that could cause reactions with characteristics of oxidation. The rapid reaction process we call burning may even take place with gases other than oxygen. The term oxidation has been extended to include a wide range of combustion and corrosion reactions, such as:



A substance that causes or promotes oxidation is called an oxidizing agent.

## Theoretical Definitions of Reduction and Oxidation

Redox reactions involve a transfer of electrons. In a redox reaction, one substance loses electrons and is oxidized, while another substance gains electrons and is reduced. Oxidation and reduction are two halves of a reaction and one cannot happen without the other. The number of electrons lost by one substance must equal the number of electrons gained by the other. In order to track electron transfer, oxidation numbers are used.

- Oxidation**
- loss of electrons
  - increase in oxidation number
- Reduction**
- gain of electrons
  - decrease in oxidation number

### LEO GER

**Oxidizing Agent** - causes oxidation by removing (gaining) electrons and is reduced in the process

**Reducing Agent** - causes reduction by donating (losing) electrons and is oxidized in the process

## Rules for Assigning Oxidation Numbers

1. The oxidation number of any free element is zero.
2. The oxidation number of a monatomic ion ( $\text{Na}^+$ ,  $\text{Ca}^{2+}$ ,  $\text{Cl}^-$ , etc.) is equal to the charge on the ion.
3. The oxidation number of each hydrogen atom in a compound is 1+, except in metal hydrides ( $\text{NaH}$ ,  $\text{LiH}$ , etc.) where it is 1-.
4. The oxidation number of each oxygen atom in a compound is 2-, except in peroxides ( $\text{H}_2\text{O}_2$ ,  $\text{Na}_2\text{O}_2$ , etc.) where it is 1-.
5. For any neutral compound, the sum of the oxidation numbers of the atoms in the compound must equal zero.
6. For a polyatomic ion, the sum of the oxidation numbers must equal the ionic charge of the polyatomic ion.

## Balancing Redox Reactions Using the Half-Reaction Method

Step 1: Write the equation in ionic form

Step 2: Write the skeleton half-reactions for the oxidation and reduction processes

If necessary balance the atoms that are being oxidized/reduced and adjust the number of electrons

Step 3: Balance elements other than oxygen and hydrogen.

Step 4: Balance oxygen and hydrogen.

In aqueous solutions, either  $\text{H}^+$  and  $\text{H}_2\text{O}$  or  $\text{OH}^-$  and  $\text{H}_2\text{O}$  are available.