OXIDATION NUMBERS

Oxidation numbers are made-up or hypothetical numbers assigned to each atom in a reaction, individual or within a molecule. They represent, loosely, the number of electrons available for shuffling around during the course of a reaction. Reactions can proceed with or without the exchange of electrons. If electrons are exchanged, that will be reflected in the difference in oxidation number of atoms on the right and left side of the chemical equation, and the reaction is called a **redox** (reduction-oxidation) reaction.

There are agreed-upon rules for assigning oxidation numbers. I'll go through them here, but they're recapped in the table below. You can click on the table to download a .pdf copy of it.

The rules

1. Atoms in elemental form have an oxidation number of zero. For example, **Mg**, $\mathbf{H_2}$, **Ar** and $\mathbf{Fe_{(s)}}$ are all examples of atoms in their elemental states. In the case of Mg, if no charge and no state are shown, we have to assume it's metallic **Mg**. Hydrogen exists as a diatomic gas in its elemental form.

2. Group 1A and 2A elements have the same oxidation number that the ion of the metal would, unless it's in its elemental form.

3. Hydrogen almost always has an oxidation number of **+1**. A rarely-encountered exception is when **H** is bound to a metal in a **hydride** compound.

4. Oxygen almost always has an oxidation number of **-2**. In rare exceptions, when oxygen is in a peroxide (**O_2^-**, like **H_2O_2**), its oxidation number is **-1**.

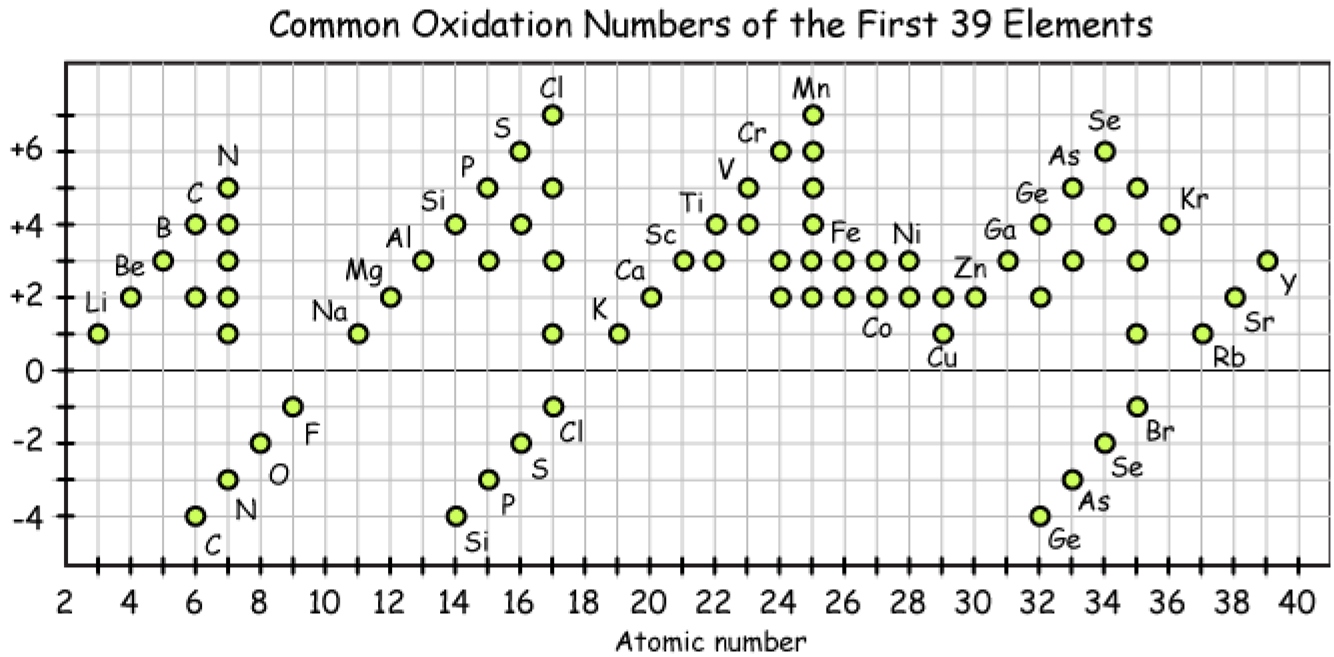
5. Fluorine always has a **-1** oxidation number, and **Cl**, **Br** and **I** almost always do.

6. This may be the most important rule: The oxidation numbers of a molecule have to add up to the total charges on the molecule. If the molecule is neutral, that's zero. For example, the sums of the oxidation numbers of **CO_2** and **CO_3^{2-}** are**0** and **-2**, respectively.

Elemental form	**zero (0)**. Only one kind of atom present, no charge
Atomic ions	**= the charge on the atom** (monatomic ion)
Group 1A Li, Na, K, Rb, Cs	**+1** unless in elemental form
Group 2A Be, Mg, Ca, Sr, Ba	**+2** unless in elemental form
Hydrogen (H)	**+1** when bonded to a nonmetal, **-1** when bonded to a metal
Oxygen (O)	**-1** in peroxides O_2^-, **-2** in all other compounds (most common)
Fluorine (F)	**-1**, always
Neutral compounds	The sum of all oxidation numbers of atoms or ions in a neutral compound is **zero**.
Ionic compounds	The sum of all oxidation numbers of atoms in an ionic compound is the **charge** on the polyatomic ion.

You've noticed that some oxidation numbers are fixed, and others can vary (otherwise we wouldn't *have* redox reactions). It turns out that the oxidation numbers of some atoms can vary quite a lot.

The chart below should help you to visualize the possible oxidation numbers that can occur for the first 39 atoms. If you're working out the oxidation states of the atoms in a reaction and you get one that's *not* on this chart, it's probably worth checking your work. You can download the chart and the table above by clicking on either.



Below you'll find a few examples of how we use oxidation numbers to make some judgments about chemical reactions.

The general approach is to assign oxidation numbers to each atom (remember, its each *atom*, not molecule), then compare the oxidation number of any given atom on both the left and right sides of the reaction. If the oxidation state (number) of the atom increases, that atom is **oxidized** (loses electrons). If the oxidation state decreases from left to right, that atom is **reduced**.

If the oxidation number of an atom increases from left to right in a reaction, the atom is oxidized in the process. If it decreases, the atom is reduced.

Source: http://www.drcruzan.com/OxidationNumbers.html