

# The Four “Charges” in Chemistry

As you study chemistry, you will encounter four “charges”, or what look like charges. Some of these represent actual electric charges you may have studied before, but others represent calculated values that, although they look like electric charges, are actually something quite different. It the purpose of this study sheet to help you understand this confusing situation. The four “charges” are ionic charge, oxidation numbers, formal charge, and partial charges. Each is explained below.

## Ionic Charge

This is the actual electric charge that is caused by an imbalance between the number of protons and the number of electrons that are found on an atom or polyatomic ion. If an atom or polyatomic ion has fewer protons than it has electrons, it will carry a negative electric charge and be termed an anion. If there are more protons than electrons there will be a positive charge and the ion will be called a cation. Examples include a sodium ion,  $\text{Na}^{+1}$ ; and a sulfate ion,  $\text{SO}_4^{-2}$ . Ionic charges commonly range from -3 to +6, although other values are possible.

## Oxidation Number

The oxidation number is a *calculated number* that may equal the ionic charge for monoatomic ions, but may also be calculated for atoms in compounds that are actually sharing electrons and have no electric charge at all. The oxidation number applies only to individual atoms, although they can be calculated for those atoms within a compound or polyatomic ion. Because oxidation numbers are often equal to ionic charges, these two are easily confused. The calculation of oxidation numbers is accomplished using a few simple rules. It is often expressed by writing the value in roman numerals enclosed within parenthesis immediately following the symbol for the element. An example is Cr(VI), for chromium in the chromate ion ( $\text{CrO}_4^{-2}$ ). Oxidation numbers are never used in the formula of a compound, but are always used in the name of a compound containing a variable valence element like iron. For example: iron (III) oxide ( $\text{Fe}_2\text{O}_3$ ). This is an example of the oxidation number (III) equaling the charge of the iron ion ( $\text{Fe}^{+3}$ ). Although equal, they are technically slightly different things.

## Partial Charge

Certain covalent compounds have an uneven distribution of electronegative atoms and/or unshared pairs of electrons in three-dimensional space. This causes the molecule to both have portions that concentrate electrons in certain area and areas that are deficient in electrons. This causes part of the molecule to have a partial negative electric charge, and another part to have a partial positive charge. Such molecules have a dipole moment and are said to be polar, even though they may have no net electric charge at all. The concentration of partial charge is indicated by writing the Greek letter delta ( $\delta$ ) followed by a plus or minus symbol to indicate the nature of the charge. An example is the gas hydrogen fluoride.  $\delta^+ \text{H} \text{--} \text{F} \delta^-$

## Formal Charge

The formal charge is also a *calculated number* that is normally applied to atoms of covalent compounds. It is determined by subtracting from the number of valence electrons an atom has, the total number of unshared electrons and one-half of the total number of shared electrons the atom has in a compound or ion.  $\text{Formal Charge} = \# \text{ of valence electrons} - (\text{all unshared electrons} + \frac{1}{2} \text{ of the shared electrons})$ . The number of shared and unshared electrons is determined from the Lewis structure of the compound or ion. An example is oxygen in hydroxide ( $\text{OH}^{-1}$ ).  $\text{F.C.} = 6 - (6 + \frac{1}{2} \times 2) = -1$ . The formal charge on oxygen is -1. This value is never represented like an ionic charge ( $\text{O}^{-1}$  is incorrect) or an oxidation number ( $\text{O}(-1)$  is incorrect). It is simply indicated by stating the formal charge as -1 for oxygen. The formal charge is often zero and normally is not equal to the oxidation number or the ionic charge.