Valences & Orbitals

On the periodic table, the elements are arranged according to atomic number. As the atomic number increases, the number of electrons found in a field outside of the nucleus of the atom increases according to certain rules. The fields outside of an atom where electrons are found are called orbitals. There are four different kinds of orbitals, s, p, d, and f orbitals. The number of electrons that can be put in an orbital increases from s to p, from p to d, and from d to f. 2, 6, 10, and 14 are the number of electrons that can be found in each orbital respectively. Each atom has a specific electron configuration. The number of electrons in each orbital for an atom is listed in exponential notation, to the upper right of the orbital letter.

The elements for the main group elements follow a pattern of filling similar orbitals. Group One elements begining with hydrogen fill Xs^1 For example, Hydrogen is $1s^1$. An "X" is used in brackets to represemt the rare gas configuration. Because a rare gas has a "full" configuration, it is simpler to use the symbol of the rare gas to represent all of the full orbitals. In the base state all elements have electrons in their orbitals according to electron configurations. To become most stable an element will gain or lose electrons. Group One have a valence number of (+1) because the elements will lose one electron and have a net postive charge because of the loss of one electron.

Potassium's electron configuration is $2s^1$. Moving Helium leftward to sit beside Hydrogen, all of the elements in Group Two follow the electron configuration pattern of $[X]rs^2$, where r is the row number. For example, Magnesium has a configuration of $[Ne]3s^2$. Group Two elements have a valence number of (+2) because the two outer electrons in it's s orbital are lost to become more stable.

The central elements in the middle of the table fill either an inner d orbital or an inner f orbital with electrons. These elements are called the transition elements. Elements attached to the main structure of the periodic table fill an inner d orbital while elements found in an extended part of the table fill inner f orbitals. For example, Gold, a central table element has a partly filled d orbital, with configuration $[Xe]4f^{14}5d^{10}6s^1$. Uranium has a partly filled f orbital, and has a configuration of $[Rn]5f^3 6d^1 7s^2$.

For groups Three through Eight, as the row number increases, an atom will have more than a rare gas configuration in addition to the s and p orbitals as mentioned on this page. Row 4 configuration is $(r-1)d^{10}$, Row 5 configuration is $(r-1)d^{10}$, and for Row 6 the configuration is $(r-1)f^{14}(r-2)\ d^{10}$.

The next group is located in the first vertical group to the right of the gap at the top of the table. It is called Group Three. All of the elements in this vertical group are said to have a (+3) valence because its electron configuration is Xs^2p^1. For example, Boron has an electron configuration of $1s^2 2s^2 p^1$.

The elements in Group Four all have four electrons in their outermost orbitals and have valences of (+4) or (-4) with a configuration of Xs^2p^2. Notice that the element Silicon (Si) is found on the periodic table one block below Carbon (C). Silicon has a configuration of $[Ne]3s^2p^2$.

The Nitrogen group is called Group Five, and has a valence charge of (-3). A valence of (-3) suggests that the element will take three electrons from its environment to fill its shells and become most stable. Group Five elements have configurations of Xs^2p^3. Phosphorus has a configuration of $[Ne]3s^2p^3$.

Groups Six, Seven and Eight continue filling the outer p orbital and have electron configurations of $[X]rs^2p^4$, $[X]rs^2p^5$, and $[X]rs^2p^6$ respectively. The valence number for Group Six is (-2), and for Group Seven is (-1). The last column on the periodic table has a valence number of (0) because the outer p orbital is filled and electrons do not need to be gained or lost.

Another method to represent the configurations of electrons in atoms is through Lewis structures. Proceed to the Lewis structures page for information on this concept.