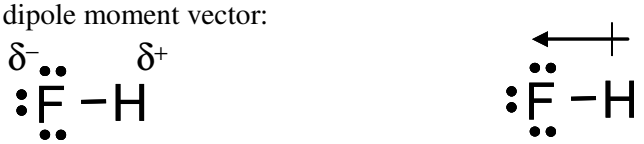
One Page Lesson: Assessing the Polarity of Molecules

When the two atoms sharing electrons in a covalent bond have *very different electronegativity values* (a difference greater than 0.4, but less than 2.1), the electrons in the bond are *not shared* ***equally*** between the two atoms. The element that has the *higher* electronegativity has greater "*electron pulling power*;" that is, the atom is able to attract the electrons in the bond toward itself. The result is that the covalent bond is "*polarized*." Or to put it more simply, we say the bond is "*polar*." Since the bonding electrons spend more of their time in the region around the element with the higher electronegativity, that atom takes on a *partial negative charge* (represented by the symbol δ^-). As the bonding electrons spend very little time in the vicinity of the less electronegative element, that atom takes on a *partial positive charge* (represented by the symbol δ^+).

So similar to a battery, a polar covalent bond has ***two*** "*poles*:" one partial positive (δ^+), and the other partial negative (δ^-). Thus we can say that it's a "*dipole*," and we can represent the polarized bonding electrons in an alternative way: by using a *dipole moment vector*. That is, we can draw a vector (simply an arrow) pointing in the direction that the electrons are being pulled: toward the atom with the higher electronegativity. To emphasize the fact that the atom at the other end of the bond has a partial positive charge, we make the "tail" of the arrow into a "plus" sign. Thus the very *polar* hydrogen fluoride bond (electronegativities: H = 2.1, F = 4.0) could be represented either by labeling the two atoms with the partial charge symbols or by drawing a dipole moment vector:



The geometry of a molecule is tremendously important when evaluating the polarity of each covalent bond around a central atom (in order to determine the type of intermolecular attractive forces that a molecule is capable of using). Quite often the three-dimensional geometry of the central atom's bonding and nonbonding electron groups can *enhance* the overall polarity of a cluster of atoms. For example, consider the water molecule, H_2O.

The hydrogen-oxygen single bonds in the angular H_2O molecule are *polar* (electronegativities: H = 2.1, O = 3.5). Thus the electrons in each bond are drawn toward the oxygen atom. This can be illustrated using either the partial charge symbols or dipole moment vectors.

But now consider the *tetrahedral geometry* of the *four* groups of electrons around the oxygen atom (2 bonds, 2 long pairs). This geometry greatly enhances the overall polarity of the angular (bent) water molecule. The dipole moment vectors are extremely useful in illustrating this additive effect of the two polar bonds. The partial negative charge of *both* dipoles is focused at the oxygen atom. Additionally, the oxygen's two lone pairs of electrons contribute significantly to its overall partial negative charge. The result is that the water molecule is extremely polar, with the partial negative charge centered at the oxygen atom, and a partial positive charge at each of the two hydrogen atoms.

Occasionally, for very small molecules, the symmetry of the bonding electrons around the central atom can cause two or more dipole moment vectors to completely *cancel* one another, so that the molecule as a whole has *no* ***net*** *dipole*. For example, consider the carbon dioxide (CO_2) molecule.

The carbon-oxygen double bonds in the *linear* CO_2 molecule are *polar* (electronegativities: C = 2.5, O = 3.5). The electrons in each of the double bonds are drawn toward the oxygens, so both oxygen atoms have a partial negative charge. This can be illustrated using either the partial charge symbols or dipole moment vectors. The vectors are especially helpful in illustrating that because of the linear geometry of the molecule, the bond dipoles are completely symmetric to one another. They are equal and opposite, so they cancel (the centers of partial-negative and partial-positive charge in the molecule are at exactly the same place). Thus the carbon dioxide molecule as a whole is *nonpolar*, even though it contains polar bonds.