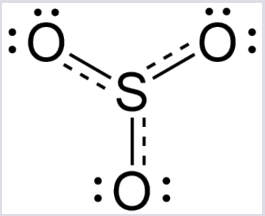
FIGURE 37



Bonding in the molecule SO_3.

The original idea was that somehow the molecule oscillated between the two, hence the terminology of "resonance." However, now we know that the bonding is truly equivalent and can best be thought of as 1.5 bonds on each side.[19] This concept is represented in Figure 36, which has partial bonds and charges added.

The molecule SO_3 can be treated similarly since all three bonds in this molecule are experimentally identical. The result can be represented as is shown in Figure 37. The dashed bonds represent ⅓ of a bond, so each resulting bond total for the link between the two atoms is 1⅓ bonds.

With detailed molecular orbital treatments that allow for sigma and pi bond formation and include "delocalized" orbitals that can bond more than two atoms together with a pair of electrons, a very modern theory—**molecular orbital theory**, or MO theory—has come to replace the resonance model.

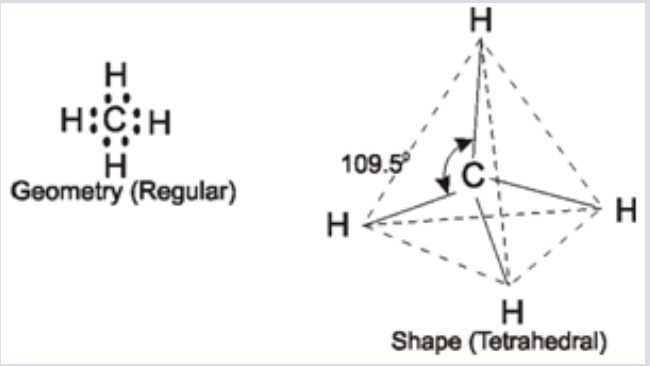
The Geometry of Molecules and Ions

VSEPR makes predictions about the geometry of simple molecules by focusing on the central atom; for example, if there are four electron pairs forming bonds around a central atom, these will point toward the corners of a tetrahedron, as is shown Figure 38, which shows a model for a methane molecule.

Dipole Moments of Molecules

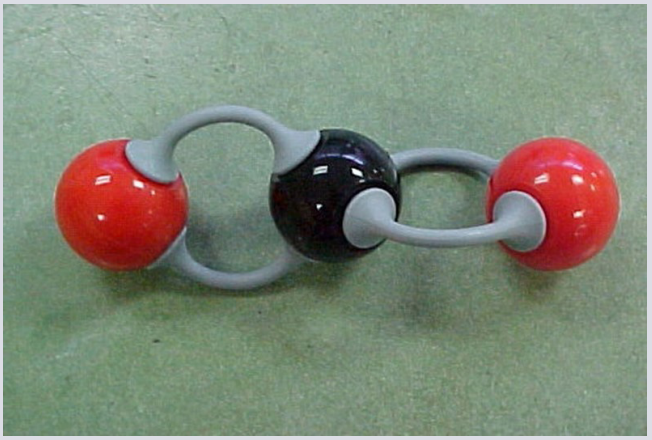
Earlier, we described **dipoles** and polar bonds. One of the clues to the geometry of a molecule that has polar bonds is whether the entire molecule has a dipole moment. If it does not, then symmetry must have allowed for the sum of dipoles pointing one way to cancel the sum of dipoles pointing the other way. (If you have learned about vectors in math, this means that the addition of vectors equals zero.) If the molecule does have a significant dipole moment, then there must be a geometry in which the bond dipoles do not cancel.

FIGURE 38



A VSEPR model (right) for a methane molecule.

FIGURE 39



Depiction of a linear molecular structure.

For example, in CO_2 (carbon dioxide) the C=O bonds are polar (i.e., they have dipoles, or charge separation), but the molecule has zero dipole moment. So, we conclude that one C=O bond must exactly oppose the other. Thus, the structure of the molecule is linear as is shown in the model pictured in Figure 39.

The Relationship between Molecular Properties and Molecular Structure

The polarity and shapes of molecules affect the way in which intermolecular forces are distributed. Highly polar molecules will exhibit strong van der Waals forces, making the molecules more difficult to separate. Such molecules therefore have higher melting points and boiling points than molecules with lower polarity. Molecules that are nonpolar have the lowest melting and boiling points.