Our minds can handle two electrons interacting with one another in a sphere of space. But then we start putting in double bonds and triple bonds. The way we draw these bonds suggests we are squeezing more electrons into the same space, and that doesn't work. Electrons don't like to be pushed together (especially since they all have negative charges that repel one another). So we need a more complex picture that works for all these electrons.

**Sigma and Pi Bonds**

The hybridization model helps explain molecules with double or triple bonds (see figure below). Ethene \(\ce{C_2H_4}\) contains a double covalent bond between the two carbon atoms and single bonds between the carbon atoms and the hydrogen atoms. The entire molecule is planar.

![Geometry of ethene molecule](CC BY-NC; CK-12)

As can be seen in the figure below, the electron domain geometry around each carbon independently is trigonal planar. This corresponds to \(\text{sp}^2\) hybridization. Previously, we saw carbon undergo \(\text{sp}^3\) hybridization in a \(\ce{CH_4}\) molecule, so the electron promotion is the same for ethene, but the hybridization occurs only between the single \(\text{s}\) orbital and two of the three \(\text{p}\) orbitals. Thus generates a set of three \(\text{sp}^2\) hybrids along with an unhybridized \(2\text{p}_z\) orbital. Each contains one electron and so is capable of forming a covalent bond.

![Hybridization in ethene](CC BY-NC; CK-12)

The three \(\text{sp}^2\) hybrid orbitals lie in one plane, while the unhybridized \(2\text{p}_z\) orbital is oriented perpendicular to that plane. The bonding in \(\ce{C_2H_4}\) is explained as follows. One of the three \(\text{sp}^2\) hybrids forms a bond by overlapping with the identical hybrid orbital on the other carbon atom. The remaining two hybrid orbitals form bonds by overlapping with the \(1\text{s}\) orbital of a hydrogen atom. Finally, the \(2\text{p}_z\) orbitals on each carbon atom form another bond by overlapping with one another sideways.

It is necessary to distinguish between the two types of covalent bonds in a \(\ce{C_2H_4}\) molecule. A **sigma bond** (\(\text{\sigma bond}\)) is a bond formed by the overlap of orbitals in an end-to-end fashion, with the electron density concentrated between the nuclei of the bonding atoms. A **pi bond** (\(\text{\pi bond}\)) is a bond formed by the overlap of orbitals in a side-by-side fashion with the electron density concentrated above and below the plane of the nuclei of the
bonding atoms. The figure below shows the two types of bonding in \(\text{C}_2\text{H}_4\). The \(\text{sp}^2\) hybrid orbitals are purple and the \(p_z\) orbital is blue. Three sigma bonds are formed from each carbon atom for a total of six sigma bonds total in the molecule. The pi bond is the "second" bond of the double bonds between the carbon atoms and is shown as an elongated green lobe that extends both above and below the plane of the molecule. This plane contains the six atoms and all of the sigma bonds.

![Images of sigma and pi bonds](Figure 9.20.3: Sigma and pi bonds. (CC BY-NC; CK-12))

In a conventional Lewis electron-dot structure, a double bond is shown as a double dash between the atoms as in \(\text{C} = \text{C}\). It is important to realize, however, that the two bonds are different: one is a sigma bond, while the other is a pi bond.

Ethyne \(\text{C}_2\text{H}_2\) is a linear molecule with a triple bond between the two carbon atoms (see figure below). The hybridization is therefore \(\text{sp}\).

\[
\text{H} - \text{C} \equiv \text{C} - \text{H}
\]

![Ethyne structure](Figure 9.20.4: Ethyne structure. (CC BY-NC; CK-12))

The promotion of an electron in the carbon atom occurs in the same way. However, the hybridization now involves only the \(2s\) orbital and the \(2p_x\) orbital, leaving the \(2p_y\) and the \(2p_z\) orbitals unhybridized.

![Images of hybridization](Figure 9.20.5: Hybridization in ethyne. (CC BY-NC; CK-12))

The \(\text{sp}\) hybrid orbitals form a sigma bond between each other as well as sigma bonds to the hydrogen atoms. Both the \(p_y\) and the \(p_z\) orbitals on each carbon atom form pi bonds between each other. As with ethene, these side-to-side overlaps are above and below the plane of the molecule. The orientation of the two pi bonds is that they are
perpendicular to one another (see figure below). One pi bond is above and below the line of the molecule as shown, while the other is in front of and behind the page.

![Bond Diagram]

Figure 9.20.6: The \(\text{C}_2\text{H}_2\) molecule contains a triple bond between the two carbon atoms, one of which is a sigma bond, and two of which are pi bonds. (CC BY-NC; CK-12)

In general, single bonds between atoms are always sigma bonds. Double bonds are comprised of one sigma and one pi bond. Triple bonds are comprised of one sigma bond and two pi bonds.

**Summary**

- Sigma bonds form between two atoms.
- Pi bonds form from \((p)\) orbital overlap.

**Contributors**

- [CK-12 Foundation](https://www.ck12.org) by Sharon Bewick, Richard Parsons, Therese Forsythe, Shonna Robinson, and Jean Dupon.