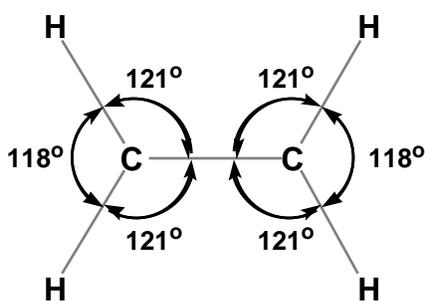


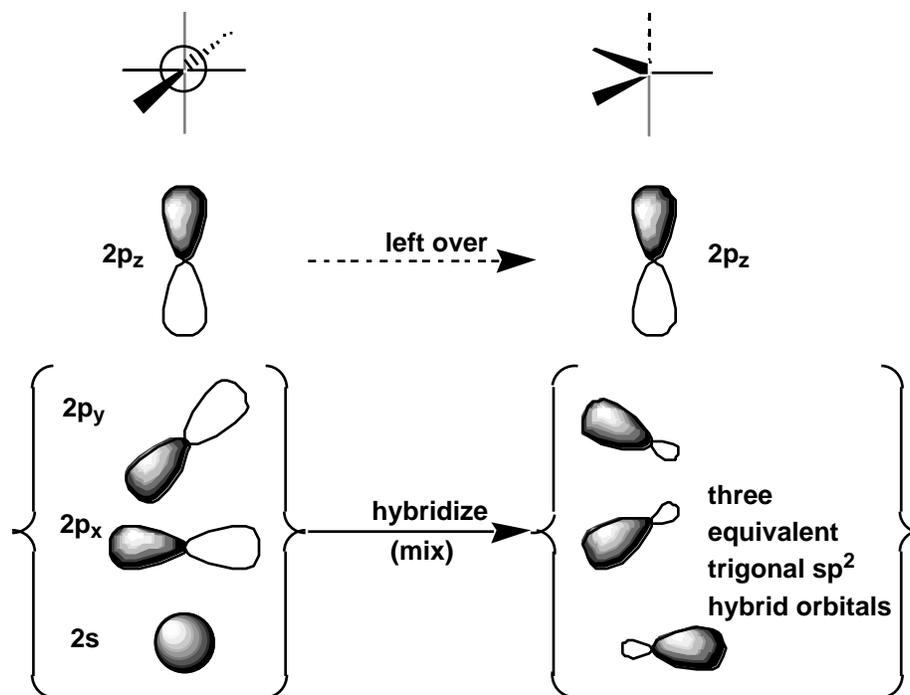
**Text Related to Segment 7.03 ©2002 Claude E. Wintner**

Carbon by no means is limited to a tetrahedral geometry, or to being bonded to four other groups. For example, the molecule ethylene has the formula  $C_2H_4$ . Its atoms all lie in a plane and are arranged with each carbon atom attached to three other atoms: one carbon atom and two hydrogen atoms. By electron diffraction the bond angles in ethylene are found to be almost, although not quite perfectly, trigonal (H—C—H angle of approximately  $118^\circ$ , rather than  $120^\circ$ ):



**arrangement of atoms in ethylene**

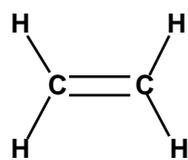
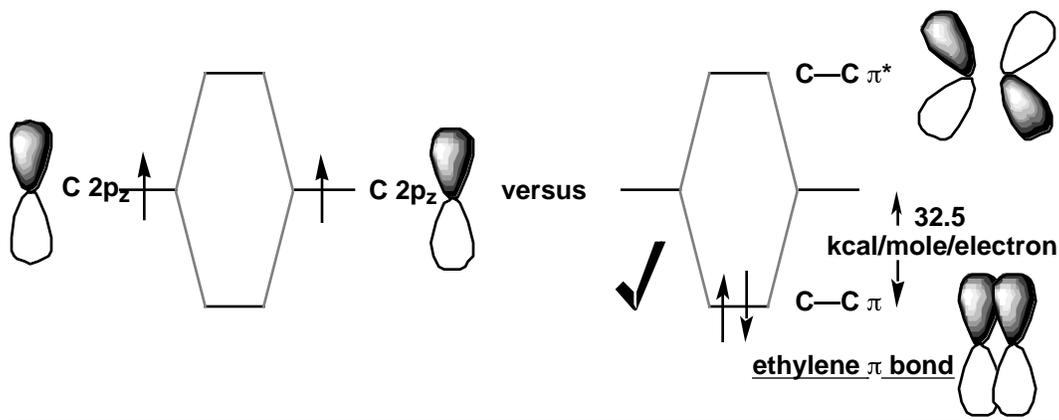
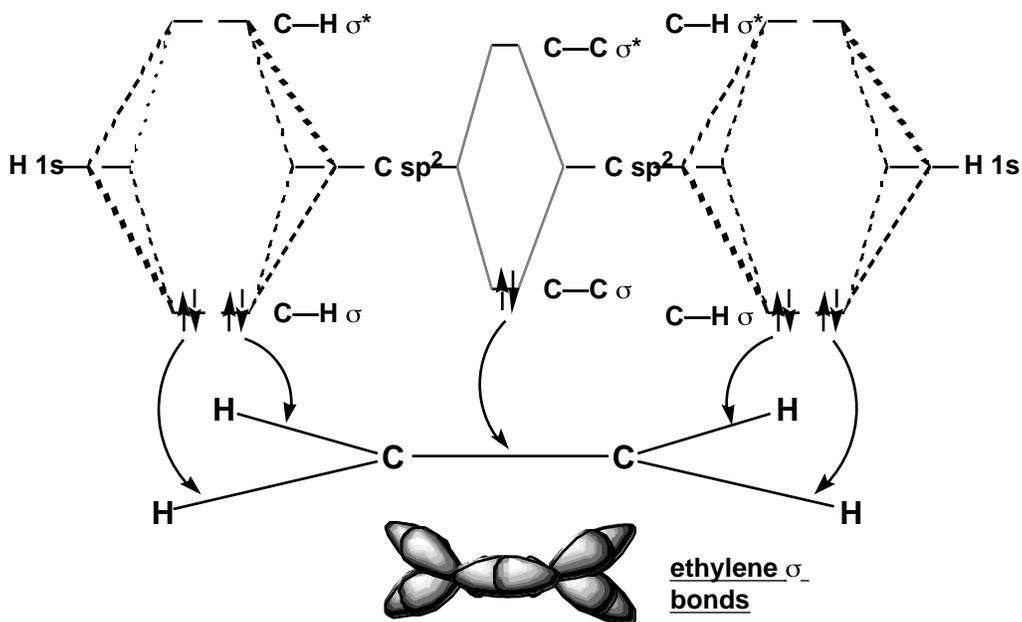
What is the bonding picture here? We have used carbon's valence orbitals to form tetrahedral  $sp^3$  orbitals; now we deploy them differently, by mixing the  $2s$ ,  $2p_x$ , and  $2p_y$  orbitals to form three equivalent trigonal  $sp^2$  hybrid orbitals lying in a plane. Left over is the  $2p_z$  orbital, perpendicular to the plane. Note again how the directionality of the array as a whole is conserved: starting with a  $2s$  orbital (non-directional) and two  $2p$  orbitals that define direction in two dimensions, we end up with the two-dimensional trigonal array predicted for three equivalent vectors pointing outward in two dimensions:



construction of  $sp^2$  hybrid atomic orbitals by mixing 2s and two 2p orbitals

A bond framework for ethylene now can be constructed, just as for ethane, but with the central C—C bond resulting from overlap of two  $sp^2$  orbitals, one from each carbon atom, and the use of two electrons, again one from each carbon, to form a  $Csp^2—Csp^2$  bond, as in the figure below. The remaining two  $sp^2$  orbitals of each carbon atom, and two electrons from each, are employed for bonds with hydrogen. Left with the two unused  $2p_z$  orbitals, and with an unused electron from each carbon, we might simply consider each  $2p_z$  orbital to hold a single localized electron; however, in light of what we have discovered already about the consequences of neglecting delocalization, we should be concerned that such a model will predict an energy that is too high. Instead, to reap the benefits of the delocalization that can take place, the two  $2p_z$  orbitals should be mixed to form two new molecular orbitals, (bonding) and \* (antibonding), having approximate shapes as shown in the figure, and the orbital filled with the two available electrons.

In this model (the  $\sigma/\pi$  model), then, the "double bond" in ethylene is considered to have a  $\sigma$  component and a  $\pi$  component:



$\sigma/\pi$  bonding scheme for ethylene

The shorter C—C bond length in ethylene (1.34Å) as opposed to ethane (1.54Å) is interpreted as a reflection of the greater "tightness" of the double bond as

opposed to the single bond. (In correct systematic nomenclature ethylene is named *ethene*. Also, we should mention that hydrocarbons possessing carbon-carbon double bonds often are called *unsaturated* hydrocarbons, or *olefins*.)

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