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 $\text{C}_2\text{H}_4$. 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 $120^\circ$):

![Diagram of ethylene molecule]

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<sup>2</sup> hybrid atomic orbitals by mixing 2s and two 2p orbitals

A σ 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<sup>2</sup> orbitals, one from each carbon atom, and the use of two electrons, again one from each carbon, to form a Csp<sup>2</sup>—Csp<sup>2</sup> σ bond, as in the figure below. The remaining two sp<sup>2</sup> orbitals of each carbon atom, and two electrons from each, are employed for bonds with hydrogen. Left with the two unused 2p<sub>z</sub> orbitals, and with an unused electron from each carbon, we might simply consider each 2p<sub>z</sub> 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<sub>z</sub> orbitals should be mixed to form two new molecular orbitals, π (bonding) and π* (antibonding), having approximate shapes as shown in the figure, and the π 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:

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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