

Hybrid Orbitals - sp^3

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Printed: November 3, 2021 (PST)



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Hybrid Orbitals - sp³

Combination of two or more non equivalent orbitals from the same atom.

Hybrid Orbitals - sp³



[Figure1]

Do you recognize this plant?

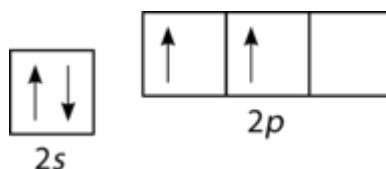
If we were walking on the beach, the plants shown above would look very different. They would be short and sticking out of the sand. When we see them this way, we do not immediately recognize them as beach plants. Often we need to look at the world around us in different ways to understand things better.

Hybrid Orbitals – sp³

The bonding scheme described by [valence bond theory](#) must account for molecular geometries as predicted by [VSEPR theory](#). To do that, we must introduce a concept called hybrid orbitals.

sp³ Hybridization

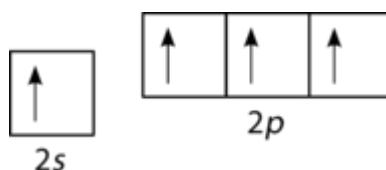
Unfortunately, overlap of existing atomic orbitals (*s*, *p*, etc.) is not sufficient to explain some of the bonding and molecular geometries that are observed. Consider the [element](#) carbon and the methane (CH₄) molecule. A carbon [atom](#) has the [electron](#) configuration of $1s^2 2s^2 2p^2$, meaning that it has two unpaired electrons in its *2p* orbitals, as shown in **Figure** [below](#).



[Figure2]

Orbital configuration for carbon [atom](#).

According to the description of [valence bond theory](#) so far, carbon would be expected to form only two bonds, corresponding to its two unpaired electrons. However, methane is a common and stable molecule, with four equivalent C–H bonds. To account for this, one of the 2s electrons is promoted to the empty 2p [orbital](#) (see **Figure below**).



[Figure3]

Promotion of carbon s [electron](#) to empty p [orbital](#).

Now, four bonds are possible. The promotion of the electron “costs” a small amount of [energy](#), but recall that the process of bond formation is accompanied by a decrease in energy. The two extra bonds that can now be formed results in a lower overall energy and thus greater stability to the CH₄ molecule. Carbon normally forms four bonds in most of its compounds.

The number of bonds is now correct, but the geometry is wrong. The three *p* orbitals (*p_x*, *p_y*, *p_z*) are oriented at 90° relative to one another. However, as was seen from [VSEPR theory](#), the observed H–C–H bond angle in the tetrahedral CH₄ molecule is actually 109.5°. Therefore, the methane molecule cannot be adequately represented by simple overlap of the 2s and 2p orbitals of carbon with the 1s orbitals of each hydrogen atom.

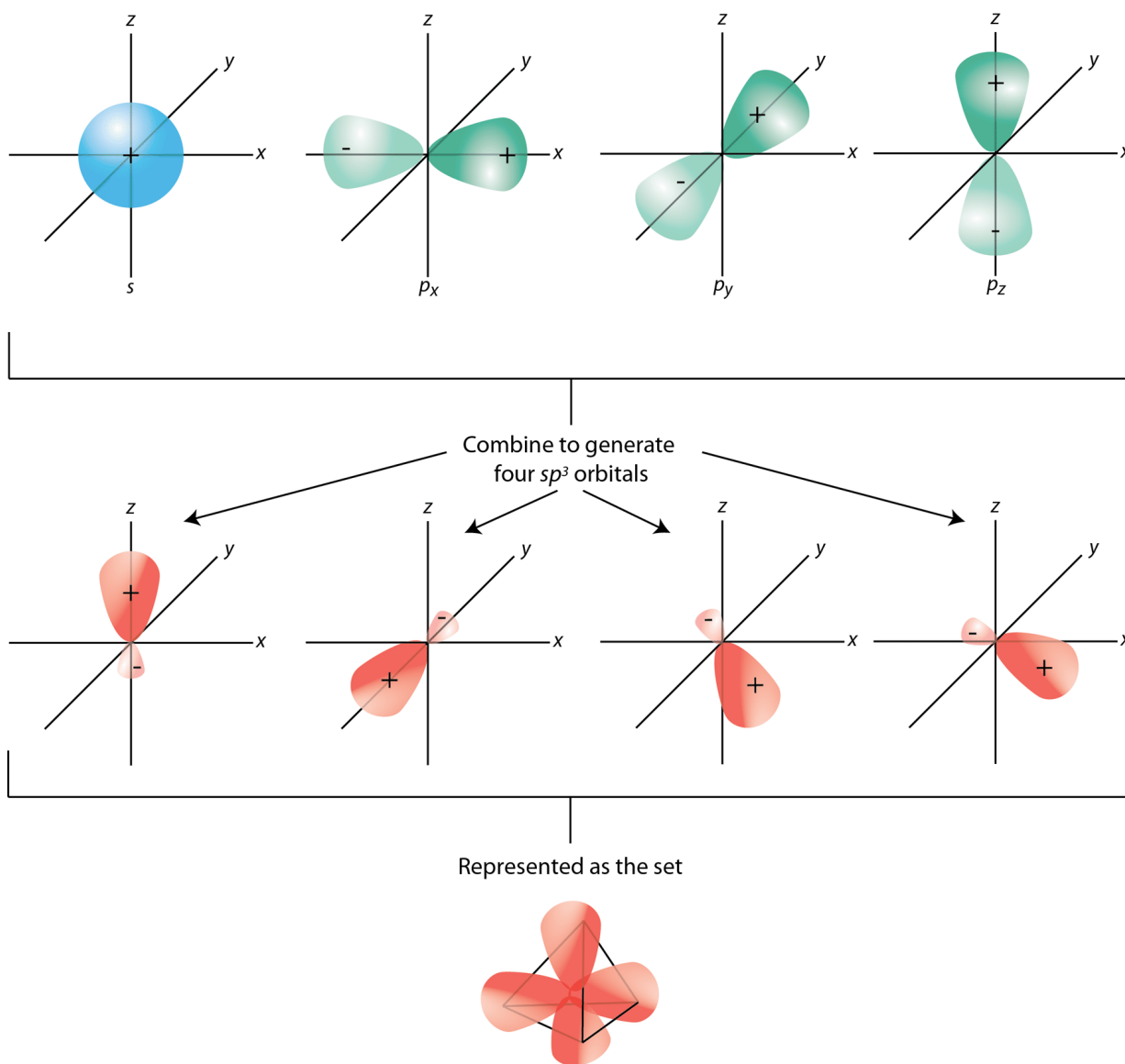
To explain the bonding in methane, it is necessary to introduce the concept of hybridization and hybrid atomic orbitals. **Hybridization** is the mixing of the atomic orbitals in an atom to produce a set of hybrid orbitals. When hybridization occurs, it must do so as a result of the mixing of nonequivalent orbitals. In other words, s and *p* orbitals can hybridize but *p* orbitals cannot hybridize with other *p* orbitals. **Hybrid orbitals** are the atomic orbitals obtained when two or more nonequivalent orbitals from the same atom combine in preparation for bond formation. In the [current](#) case of carbon, the single 2s orbital hybridizes with the three 2p orbitals to form a set of four hybrid orbitals, called sp³ hybrids (see **Figure below**).

 $2sp^3$

[Figure4]

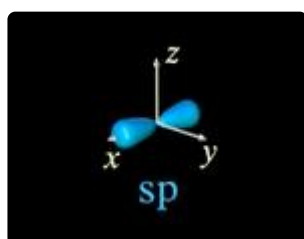
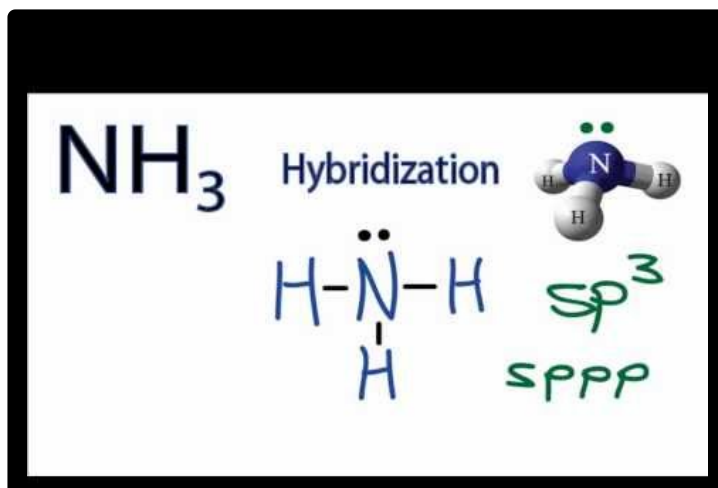
Carbon sp³ hybrid orbitals.

The sp³ hybrids are all equivalent to one another. Spatially, the hybrid orbitals point towards the four corners of a tetrahedron (see **Figure below**).



[Figure5]

The process of sp³ hybridization is the mixing of an s orbital with a set of three p orbitals to form a set of four sp³ hybrid orbitals. Each large lobe of the hybrid orbitals points to one corner of a tetrahedron. The four lobes of each of the sp³ hybrid orbitals then overlap with the normal unhybridized 1s orbitals of each hydrogen atoms to form the tetrahedral methane molecule.



<https://www.ck12.org/flx/render/embeddedobject/184802>

Summary

- Electrons hybridize in order to form covalent bonds.
- Nonequivalent orbitals mix to form hybrid orbitals.

Review

1. Why is carbon expected to form only two covalent bonds?
 2. How many covalent bonds does carbon actually form?
 3. What needs to happen to allow carbon to form four bonds?
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1.0 REFERENCES

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