Hybrid Orbitals - sp3

Ck12 Science

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

Hybrid Orbitals - sp3

[Figure1]

Do you recognize this plant?

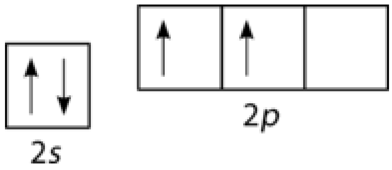
If we were walking on the beach, the plants shown above would look very different. They would be short and sticking out of the sand. When we see them this way, we do not immediately recognize them as beach plants. Often we need to look at the world around us in different ways to understand things better.

Hybrid Orbitals – sp^3

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.

sp^3 Hybridization

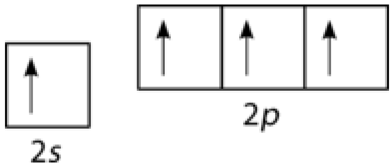
Unfortunately, overlap of existing atomic orbitals (*s*, *p*, etc.) is not sufficient to explain some of the bonding and molecular geometries that are observed. Consider the element carbon and the methane (CH_4) molecule. A carbon atom has the electron configuration of $1s^2\ 2s^2\ 2p^2$, meaning that 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).



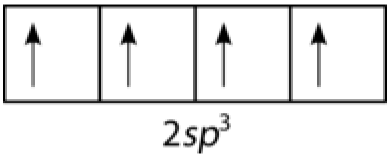
[Figure3]

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 CH_4 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^{o} relative to one another. However, as was seen from VSEPR theory, the observed H−C−H bond angle in the tetrahedral CH_4 molecule is actually 109.5^{o}. 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.

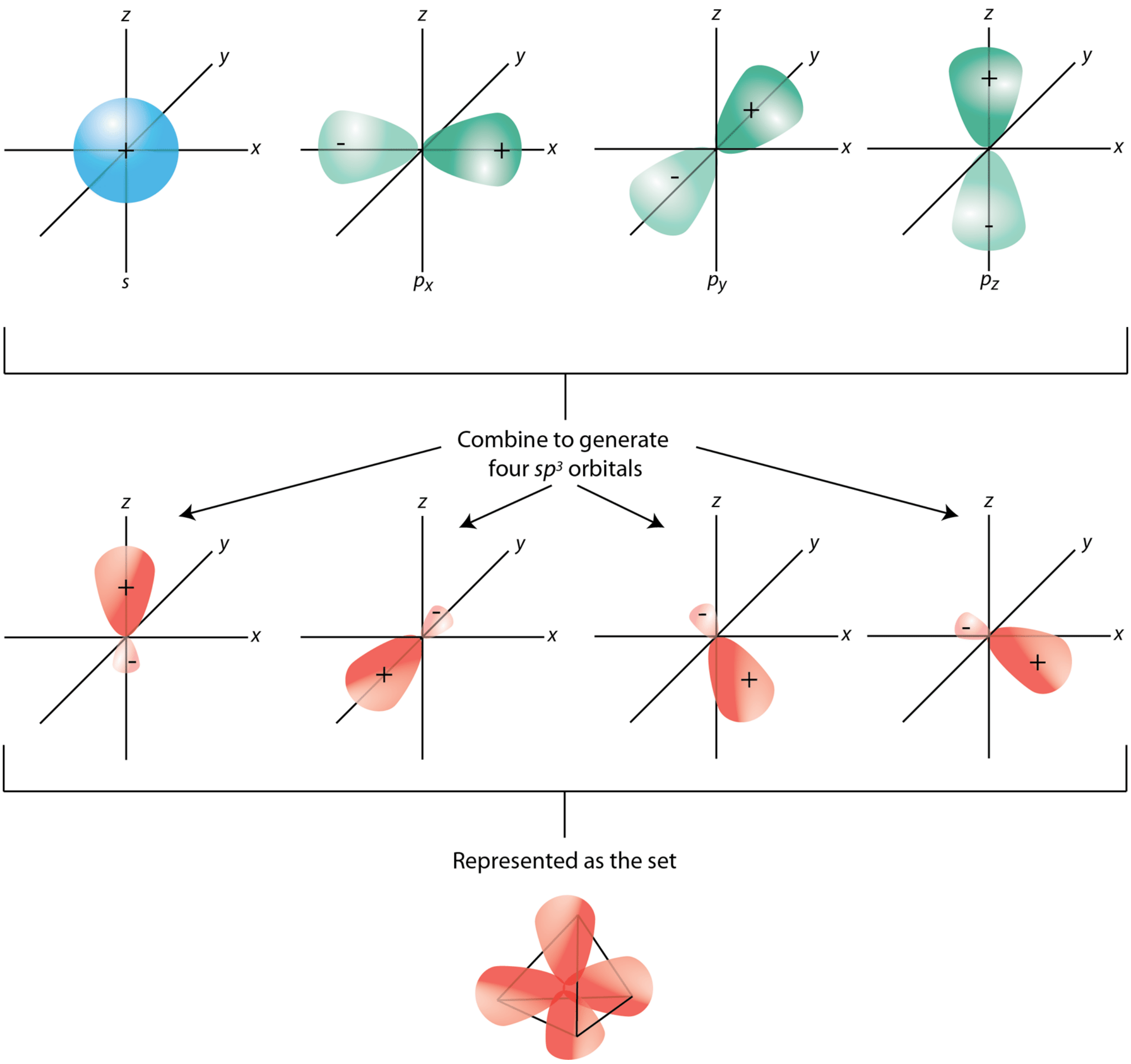
To explain the bonding in methane, it is necessary to introduce the concept of hybridization and hybrid atomic orbitals. **Hybridization** is the mixing of the atomic orbitals in an atom to produce a set of hybrid orbitals. When hybridization occurs, it must do so as a result of the mixing of nonequivalent 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 form 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 sp^3 hybrids (see **Figure** below).



[Figure4]

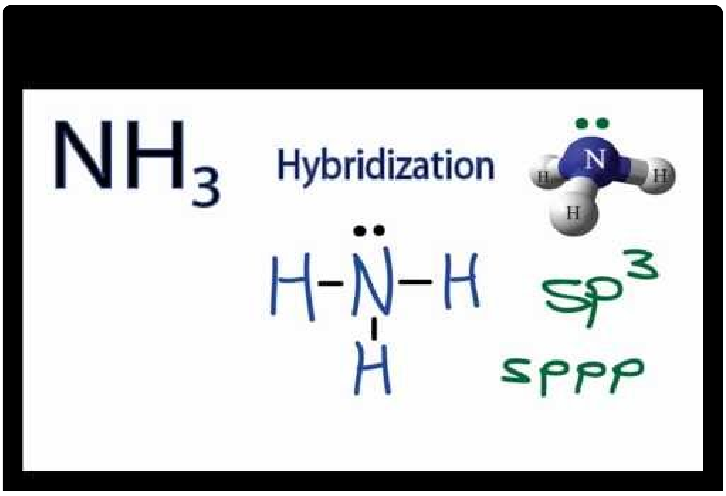
Carbon sp^3 hybrid orbitals.

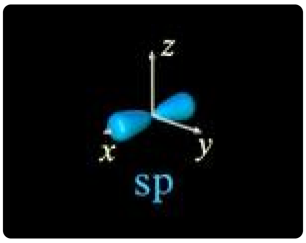
The sp^3 hybrids are all equivalent to one another. Spatially, the hybrid orbitals point towards the four corners of a tetrahedron (see **Figure** below).



[Figure5]

The process of sp^3 hybridization is the mixing of an s orbital with a set of three p orbitals to form a set of four sp^3 hybrid orbitals. Each large lobe of the hybrid orbitals points to one corner of a tetrahedron. The four lobes of each of the sp^3 hybrid orbitals then overlap with the normal unhybridized 1s orbitals of each hydrogen atoms to form the tetrahedral methane molecule.





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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?

1.0 REFERENCES

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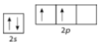
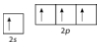
	**Credit:** CK-12 Foundation - Joy Sheng **License:** CC BY-NC-SA 3.0
	**Credit:** CK-12 Foundation - Joy Sheng **License:** CC BY-NC-SA 3.0
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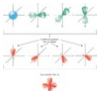	**Credit:** CK-12 Foundation - Jodi So **Source:** CK-12 Foundation **License:** CC BY-NC-SA 3.0