**Hybridization** is the concept of mixing [atomic orbitals](file:///\\wiki\Atomic_orbital) into new *hybrid orbitals* (with different energies, shapes, etc., than the component atomic orbitals) suitable for the pairing of electrons to form [chemical bonds](file:///\\wiki\Chemical_bond) in [valence bond theory](file:///\\wiki\Valence_bond_theory).

**For example**, in a carbon atom which forms four single bonds the valence-shell s orbital combines with three valence-shell p orbitals to form four equivalent sp3 mixtures which are arranged in a [tetrahedral](file:///\\wiki\Tetrahedral_molecular_geometry) arrangement around the carbon to bond to four different atoms. Hybrid orbitals are useful in the explanation of [molecular geometry](file:///\\wiki\Molecular_geometry) and atomic bonding properties and are symmetrically disposed in space. Usually hybrid orbitals are formed by mixing atomic orbitals of comparable energies.

HISTORY

(Chemical Bonding)

[Chemist](file:///C:\wiki\Chemist) [**Linus Pauling**](file:///C:\wiki\Linus_Pauling) first developed the hybridisation theory in **1931** to explain the structure of simple molecules such as **methane (CH4)** using **atomic orbitals.**

Pauling pointed out that a **carbon atom forms four bonds by using one s and three p** orbitals, so that "it might be inferred" that a **carbon atom would form three bonds at right angles (using p orbitals) and a fourth weaker bond using the s orbital in some arbitrary direction**. In reality, methane has four C-H bonds of equivalent strength. The angle between any two bonds is the tetrahedral bond angle of **109°28'**(approx. 109.5°).

This concept was developed for such simple chemical systems, but the approach was later applied more widely, and today it is considered an effective [heuristic](file:///C:\wiki\Heuristic) for rationalizing the structures of [organic compounds](file:///C:\wiki\Organic_compounds). It gives a simple orbital picture equivalent to [Lewis structures](file:///C:\wiki\Lewis_structures).

Why don’t u thanks me for such a theory. Also thanks Arnab for putting me before you!



Types of hybridization:-

**sp3**

Hybridisation describes the bonding of atoms from an atom's point of view. For a **tetrahedrally coordinated carbon** (e.g., [methane](file:///C:\wiki\Methane) CH4), the carbon should have **4 orbitals** with the correct symmetry to bond to the 4 hydrogen atoms.

Carbon's [ground state](file:///C:\wiki\Ground_state) configuration is 1s2 2s22p2 or more easily read:

The carbon atom can use its two singly occupied p-type orbitals, to form two [covalent bonds](file:///C:\wiki\Covalent_bond) with two hydrogen atoms, yielding the singlet [methylene](file:///C:\wiki\Methylene_(compound)) CH2, the simplest [carbene](file:///C:\wiki\Carbene). The carbon atom can also bond to four hydrogen atoms by an excitation (or promotion) of an electron from the doubly occupied 2s orbital to the empty 2p orbital, producing four singly occupied orbitals.

The energy released by the formation of two additional bonds more than compensates for the excitation energy required, energetically favoring the formation of four C-H bonds.

Quantum mechanically, the lowest energy is obtained if the four bonds are equivalent, which requires that they are formed from equivalent orbitals on the carbon. A set of four equivalent orbitals can be obtained that are linear combinations of the valence-shell (core orbitals are almost never involved in bonding) s and p wave functions,which are the four sp3 hybrids.

In CH4, four sp3 hybrid orbitals are overlapped by [hydrogen](file:///C:\wiki\Hydrogen) 1s orbitals, yielding four [σ (sigma) bonds](file:///C:\wiki\Sigma_bond) (that is, four single covalent bonds) of equal length and strength.

sp2

For this molecule, carbon sp2 hybridises, because one [π (pi) bond](file:///C:\wiki\Pi_bond) is required for the [double bond](file:///C:\wiki\Double_bond) between the carbons and only three σ bonds are formed per carbon atom. In sp2 hybridisation the 2s orbital is mixed with only two of the three available 2p orbitals, usually denoted 2px and 2py. The third 2p orbital (2pz) remains unhybridised.

Forming a total of three sp2 orbitals with one remaining p orbital. In ethylene ([ethene](file:///C:\wiki\Ethene)) the two carbon atoms form a σ bond by overlapping one sp2 orbital from each carbon atom. The π bond between the carbon atoms perpendicular to the molecular plane is formed by 2p–2p overlap. Each carbon atom forms covalent C–H bonds with two hydrogens by s–sp2 overlap, all with 120° bond angles. The hydrogen–carbon bonds are all of equal strength and length, in agreement with experimental data.