Basic inorganic principles

Electronic structures of atoms

Atoms consist of three fundamental types of particles: protons, electrons and neutrons. Neutrons and protons have approximately the same mass and, in contrast to this, the mass of an electron is negligible. A proton carries a positive charge, a neutron has no charge and an electron is negatively charged. An atom contains equal numbers of protons and electrons and therefore, overall, an atom has no charge. The nucleus of an atom contains protons and neutrons only, and therefore is positively charged. The electrons occupy the region of space around the nucleus. Therefore, most of the mass is concentrated within the nucleus.

Within an element, the atomic number (Z), that is, the number of protons and electrons is always the same, but the number of neutrons and therefore the mass number (A) can vary. These possible versions of an element are called isotopes



# 1- Principle quqntum nunber (n):

. It can take values of 1, 2,3, …… it represent the energy level at which the electron rotates.



# . The quantum number( l)

Represents the shape of the atomic orbital AO. It is called the orbital quantum number as it represents the orbital angular momentum of the electron. It can have values of l =0, 1, 2, …, (n−1), which correspond to the orbital labels s, p, d and f. (shape)



# The magnetic quantum number ml

Provides information about the orientation (directionality) of the AO and can take values between +l and −l. This means that there is only one direction for an s- orbital, as l =0, and therefore ml also is equal to 0. For a p-orbital, l=1 and therefore ml is −1, 0 or 1, which means it can occupy three orientations. In this case, they are classified as the px, py and pz orbitals. (orientation)



# The spin quantum number

Each individual electron can be described by an additional fourth quantum number, the so-called spin quantum number s (value of either +1/2 or −1/2). Each orbital can be filled with one or two electrons. Once an orbital is filled with two electrons, they will occupy opposite spin directions in order to fulfill the Pauli Exclusion Principle

The Pauli Exclusion Principle states that no two electrons in the same atom can have the same values for their four quantum numbers



**Valence Electrons**

The valence electrons are the electrons in the last shell or energy level of an atom.



Lewis structures and Lewis Theory

The octet rule:

All elements except hydrogen ( hydrogen have a duet of electrons) have octet of electrons once they from ions and covalent compounds.

The Lewis dot symbols for atoms and ions shows how many electrons are need for an atom to fill the octet. Normally there are octets of electrons on most monatomic ions.

A Lewis symbol is a symbol in which the electrons in the valence shell of an atom or simple ion are represented by dots placed around the letter symbol of the element. Each dot represents one electron.



For any compounds, other than Nobel gases, it is necessary to maintain the octet rule. Octet rule is obtained by bond formation.

A covalent bond

Is a chemical bond formed by the sharing of a pair of electrons between two atoms.

Bond pair: electron pair shared between two atoms.

Lone pair: electron pair found on a single atom.

Ionic bond:

It involves formation of ions, complete the octet rule, and the attraction between different ions

**Orbital hybridization**

In [chemistry](https://en.wikipedia.org/wiki/Chemistry), orbital hybridisation (or hybridization) is the concept of mixing [atomic](https://en.wikipedia.org/wiki/Atomic_orbital) [orbitals](https://en.wikipedia.org/wiki/Atomic_orbital) to form new hybrid orbitals (with different energies, shapes, etc., than the

component atomic orbitals) suitable for the pairing of electrons to form [chemical](https://en.wikipedia.org/wiki/Chemical_bond) [bonds](https://en.wikipedia.org/wiki/Chemical_bond) in [valence bond theory](https://en.wikipedia.org/wiki/Valence_bond_theory).

# Types of hybridisation

SP3:

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

Carbon's [ground state](https://en.wikipedia.org/wiki/Ground_state) configuration is 1s2 2s2 2p2 or more easily read:

The carbon atom can use its two singly occupied p-type orbitals to form two [covalent](https://en.wikipedia.org/wiki/Covalent_bond) [bonds](https://en.wikipedia.org/wiki/Covalent_bond) with two hydrogen atoms, yielding the singlet [methylene](https://en.wikipedia.org/wiki/Methylene_%28compound%29) CH2, the simplest [carbene](https://en.wikipedia.org/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,[[9]](https://en.wikipedia.org/wiki/Orbital_hybridisation#cite_note-9) which are the four sp3 hybrids.



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



SP2:

Other carbon compounds and other molecules may be explained in a similar way. For example, [ethene](https://en.wikipedia.org/wiki/Ethene) (C2H4) has a double bond between the carbons.

For this molecule, carbon sp2 hybridises, because one [π (pi) bond](https://en.wikipedia.org/wiki/Pi_bond) is required for the [double bond](https://en.wikipedia.org/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](https://en.wikipedia.org/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.



SP:

The chemical bonding in compounds such as [alkynes](https://en.wikipedia.org/wiki/Alkyne) with [triple bonds](https://en.wikipedia.org/wiki/Triple_bond) is explained by sp hybridization. In this model, the 2s orbital is mixed with only one of the three p orbitals, resulting in two sp orbitals and two remaining p orbitals. The chemical bonding in [acetylene](https://en.wikipedia.org/wiki/Acetylene) (ethyne) (C2H2) consists of sp–sp overlap between the two carbon atoms forming a σ bond and two additional [π bonds](https://en.wikipedia.org/wiki/Pi_bonds) formed by p–p overlap. Each carbon also bonds to hydrogen in a σ s–sp overlap at 180° angles.