Live Revision

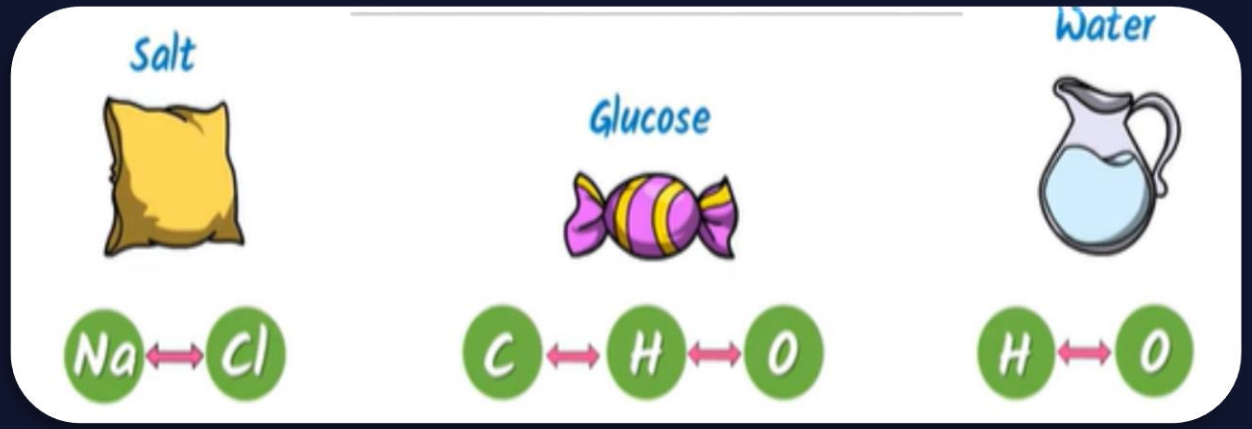
Grade 11 chemistry

Unit 2

Chemical Bonding

Definition of Chemical Bond

- A **chemical bond** is the attractive force that holds two or more atoms together in a molecule or compound.
- It forms due to the interaction of electrons in the outermost energy levels (valence electrons) of atoms.



Why atoms form chemical bonds?

- Atoms form chemical bonds to achieve stability by attaining a full outer shell of electrons (like the noble gases). This process lowers their energy state and makes them more stable.

Octet rule

- **Octet rule** the rule states that Atoms tend to gain or lose electrons or share electrons until they have achieved an outer shell that contains an octet of electrons (eight electrons).

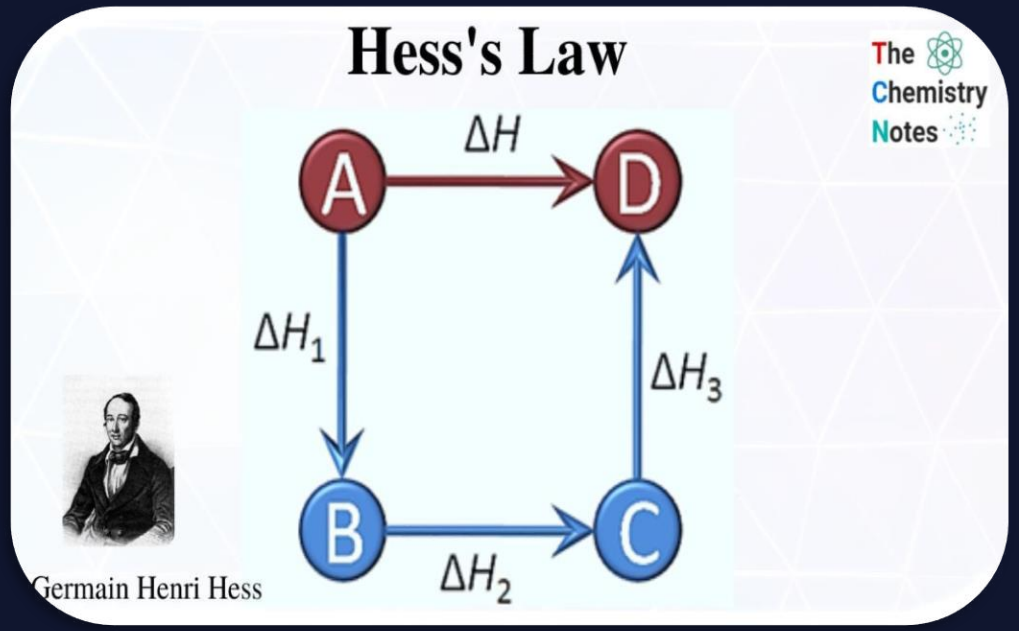
- Type of chemical bonding
  - ionic bond(Metal with non-metal)
  - covalent bond (Non-metal with non-metal)
  - metallic bond (Metal with metal).

Ionic bond

- ***Definition***
  - **An ionic bond is a type of chemical bond formed by the transfer of electrons from one atom to another, resulting in the formation of positively charged (cation) and negatively charged (anion) ions.**
  - **It also called electrovalent bond**

Formation of Ionic Compounds

- The formation process can be understood using **Hess's Law**, which states that the total enthalpy change of a reaction is the sum of the enthalpy changes of individual steps, regardless of the reaction pathway.



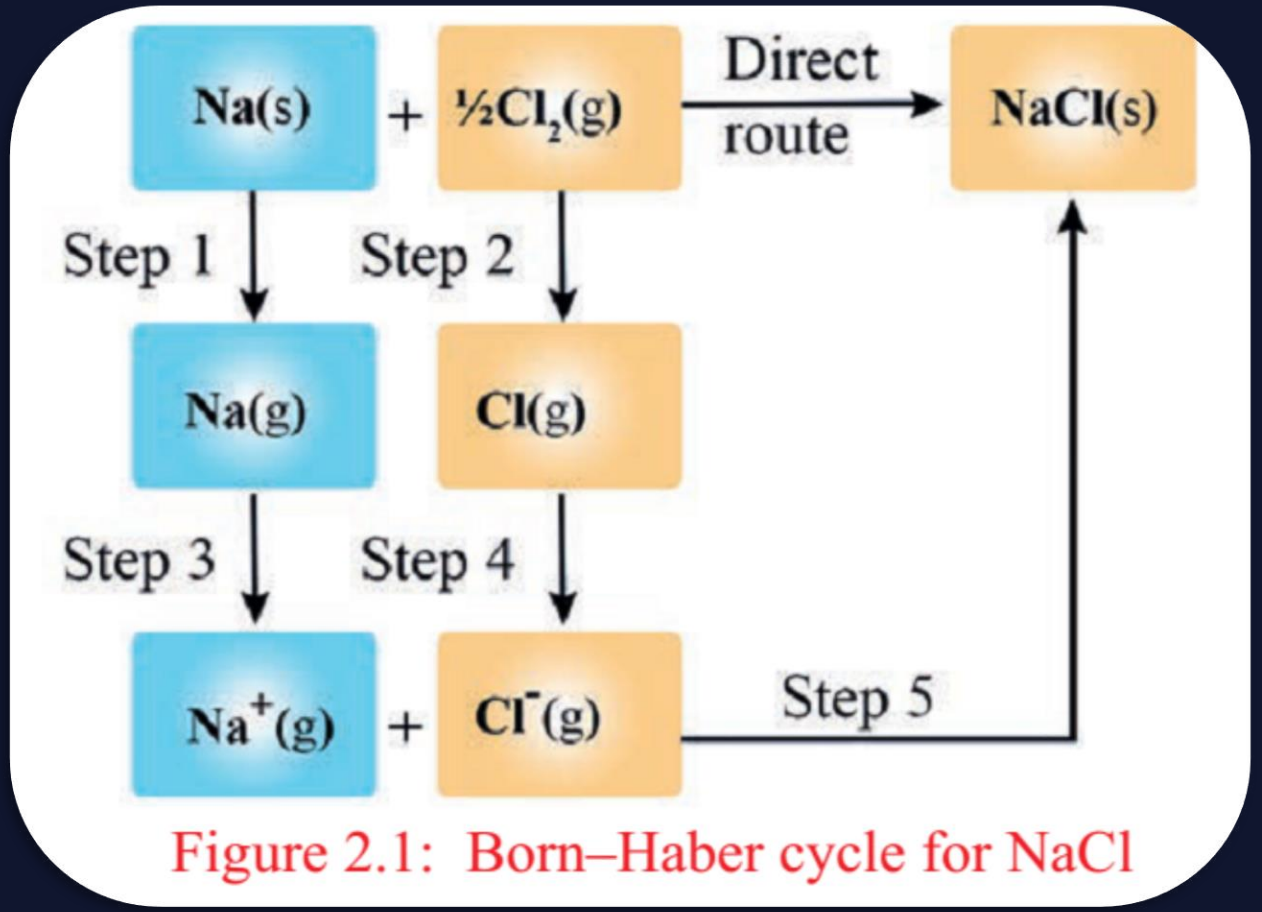


Figure 2.1: Born–Haber cycle for NaCl

Lattice Energy

- Lattice energy is the amount of energy released when one mole of an ionic compound forms from its gaseous ions.
- It measures the strength of the ionic bond.
- Higher lattice energy = Stronger ionic bond = More stable compound
- It depends on:
  - **Charge of the ions:** Higher charges = Higher lattice energy
  - **Size of the ions:** Smaller ions = Higher lattice energy

- **Factors Affecting the Formation of Ionic compound**
  - Ionization energy (IE)
  - Electron affinity (EA)
  - Lattice energy
- In conclusion, low IE, high EA, and high lattice energy all favor the formation of ionic bonds, making ionic compounds more stable.

- **Ionic compounds are characterized by their**
  - typically solid at room temperature.
  - High Melting and Boiling
  - Hard and Brittle
  - solubility in water and
  - generally non-volatile
  - electrical conductivity in molten or aqueous form.

Covalent Bond

Definition

- A **covalent bond** is a type of chemical bond formed when two atoms share one or more pairs of electrons in order to achieve a more stable electron configuration.
- Covalent bonds typically form between non-metal atoms.

How Covalent Bond Forms

- **Electron Sharing:** Atoms share electrons to fill their outermost electron shell (achieving a stable configuration).
- **Bond Formation:** The shared electrons are attracted by the nuclei of both atoms, creating a bond.
  - **Single Bond:** One pair of electrons is shared (e.g., H_2, Cl_2).
  - **Double Bond:** Two pairs of electrons are shared (e.g., O_2).
  - **Triple Bond:** Three pairs of electrons are shared (e.g., N_2).
- The shared pairs of electrons in a molecule are called **bonding pairs**.
- The other electron pairs that stay with on atom and are not shared are called **non-bonding pairs** or **lone pairs.**

Steps to Writing a Lewis structure

Step 1. Determine the total number of valence electrons.

Step 2. Determine the Central Atom

Step 3. Write the skeletal structure.

Step 4. Place electron pairs around terminal atoms so that each (except hydrogen) has an octet.

Step 5: If at this stage, a central atom has fewer than eight electrons, a multiple bond(s) is likely.

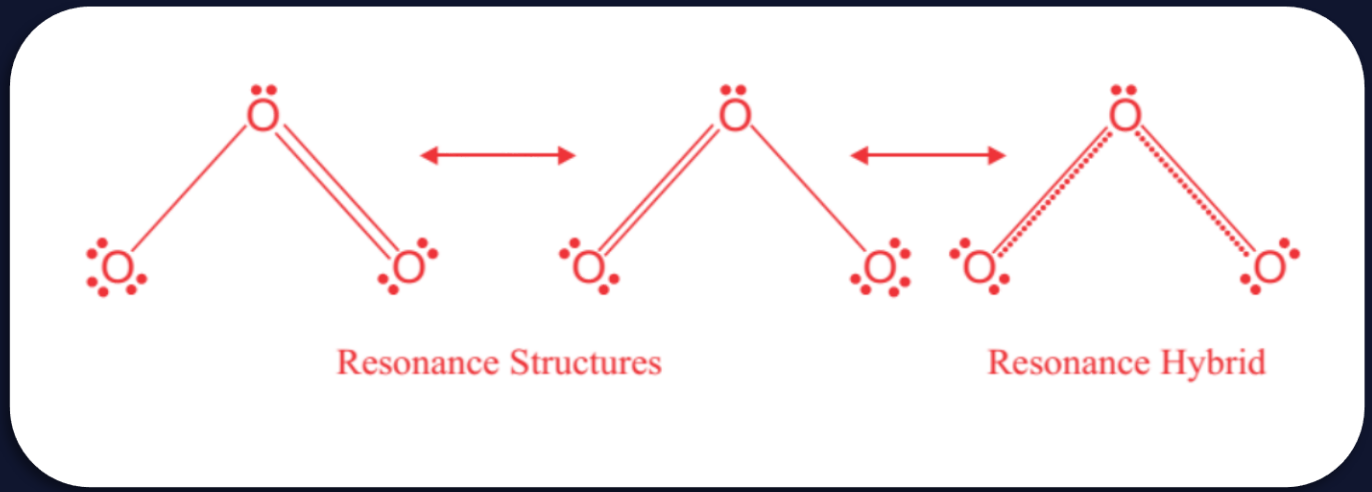
Coordinate-Covalent Bonding

Definition

- A **coordinate-covalent bond** (also called a **dative bond**) is a type of covalent bond where both electrons in the bond come from the same atom, rather than each atom contributing one electron.

Resonance Structures

- **Resonance structures** are different ways of drawing the same molecule, where the electron distribution is represented in more than one valid structure.
- These structures are used when a single Lewis structure cannot accurately represent the bonding in a molecule due to the delocalization of electrons.
- Instead, a hybrid or average of the resonance structures is considered the actual structure.



Exceptions to the Octet rule in Covalent Bonding

Although many Lewis structures follow the octet rule, there are exceptions. There can be categorized into three groups:

i. Less than octet (central atom is deficient of electrons)

ii. More than octet (central atom has excess of electrons)

iii. Molecules containing an odd number of electrons

- Based on the electronegativity difference between the bonded atoms and the distribution of shared electrons in the bond a covalent bond is classified as:
  - Non polar covalent bond
  - Polar covalent bond

Non-Polar Covalent Bond

- **A covalent bond is classified as non-polar if:**
  - The two atoms have equal or very similar electronegativity.
  - Electrons are shared equally between the atoms.
  - There is no partial positive (δ^+) or partial negative (δ^-) charge on the atoms.
- Examples of Non-Polar Bonds:
  - H_2 (Hydrogen-Hydrogen bond)

Polar Covalent Bond

- **A covalent bond is classified as polar if:**
  - There is a significant difference in electronegativity between the two atoms.
  - Electrons are shared unequally, causing the more electronegative atom to pull the shared electrons closer.
  - This creates a partial negative charge (δ^-) on the more electronegative atom and a partial positive charge (δ^+) on the less electronegative atom.
- Examples of Polar Bonds:
  - HCl (Hydrogen-Chlorine bond)

Molecule will be non-polar

- If the molecular shape around the central atom has no lone pair or if it either square planar or linear
- All the terminal atom around the central atom is the same

Classify the following molecules as polar or non-polar

$CHCl_3$

CCl_4

NF_3

WHAT IS DIPOLE MOMENT?

It measures the net polarity of a molecule.

DENOTED

μ

FORMULA

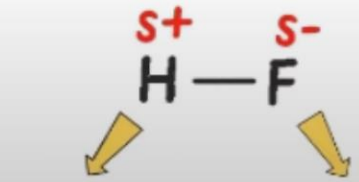
$\mu = q \times d$

Charge Distance

VECTOR

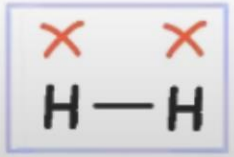


POLAR MOLECULE



Less E.N More E.N

NON-POLAR MOLECULE



- **Properties of Covalent Compounds**
  - Many of the covalent compounds are found in gaseous state at room temperature.
  - Covalent compounds are volatile.
  - Generally, they have low melting points and boiling points.
  - Covalent compounds are generally insoluble in water.
  - Most covalent compounds are soluble in non-polar solvents
  - Non-polar covalent compounds are non-electrolytes because they do not conduct electricity.

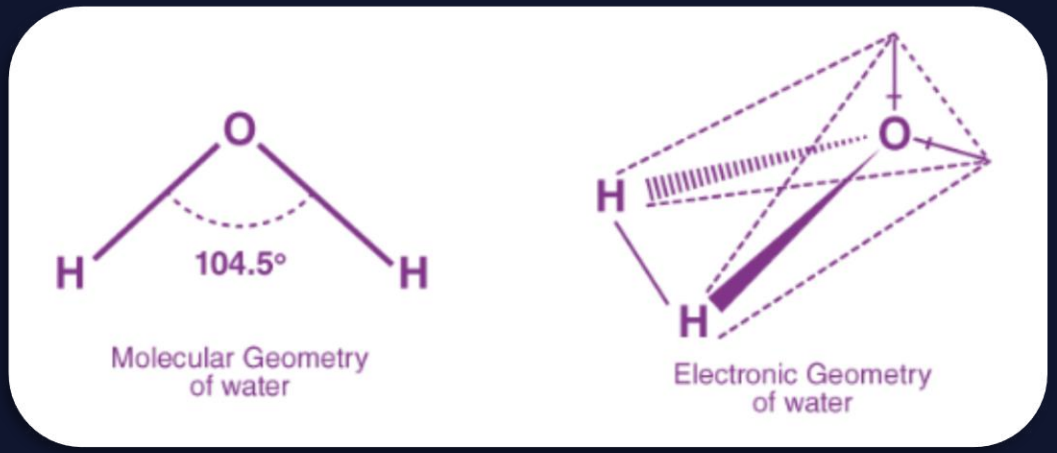
Molecular geometry

- **Molecular geometry** is the three-dimensional arrangement of atoms in a molecule.
- There is a simple procedure that enables us to predict the geometry of a molecule or ion if we know the number of electrons surrounding a central atom in its Lewis structure.
- The basis of this approach is the assumption that electron pairs in the valence shell of an atom repel one another.

Valence Shell Electron Pair Repulsion (VSEPR) Theory

- **VSEPR Theory**
  - VSEPR theory states that electron pairs (bonding and lone pairs) around a central atom will arrange themselves as far apart as possible to minimize electron pair repulsion.
  - Lone pair-lone pair > lone pair-bond pair > bond pair-bond pair
  - The geometry is determined by the total number of electron pairs around the central atom.

- **Electron Pair Geometry vs Molecular Geometry**
  - **Electron Pair Geometry:** Considers all electron pairs (bonding and lone pairs) around the central atom.
  - **Molecular Geometry:** Considers only the bonding pairs of electrons



- For the prediction of geometrical shapes of molecules with the help of VSEPR theory, it is convenient to divide molecules into two categories as

    1. Molecules in which the central atom has no lone pair and
    2. Molecules in which the central atom has one or more lone pairs.

Easy trick to learn shapes of molecules

	E. Pair	Bond Pair	Lone Pair	Shapes
2^1	2	2	0	linear
3^2	3	3	0	trigonal
	3	2	1	bent
4^3	4	4	0	tetrahedral
	4	3	1	trigonal pyramidal
	4	2	2	v shape
5^4	5	5	0	trigonal bipyramidal
	5	4	1	sea saw
	5	3	2	t shape
	5	2	3	linear





Below

top

valley,

SS

teacher

live.

- **Bond Angle**
  - **Bond Angle is the angle formed between two adjacent bonds around a central atom in a molecule.**
  - **It is measured in degrees (°) and plays a key role in determining the shape and geometry of the molecule.**

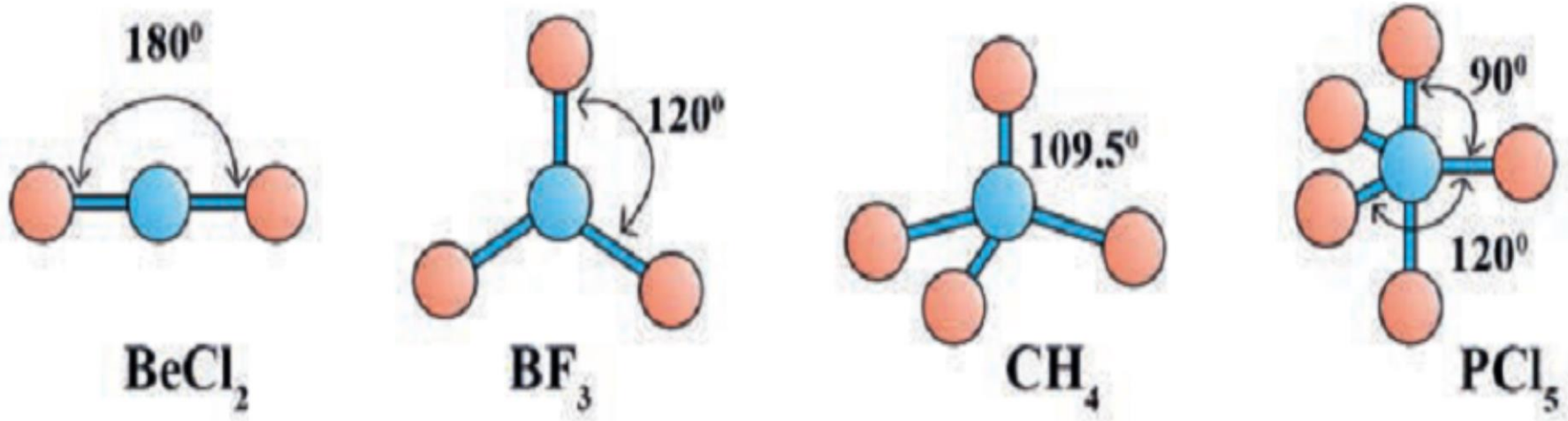


Figure 2.7: The shapes of molecules in which the central atom has no lone pair

Intermolecular force in the covalent bond

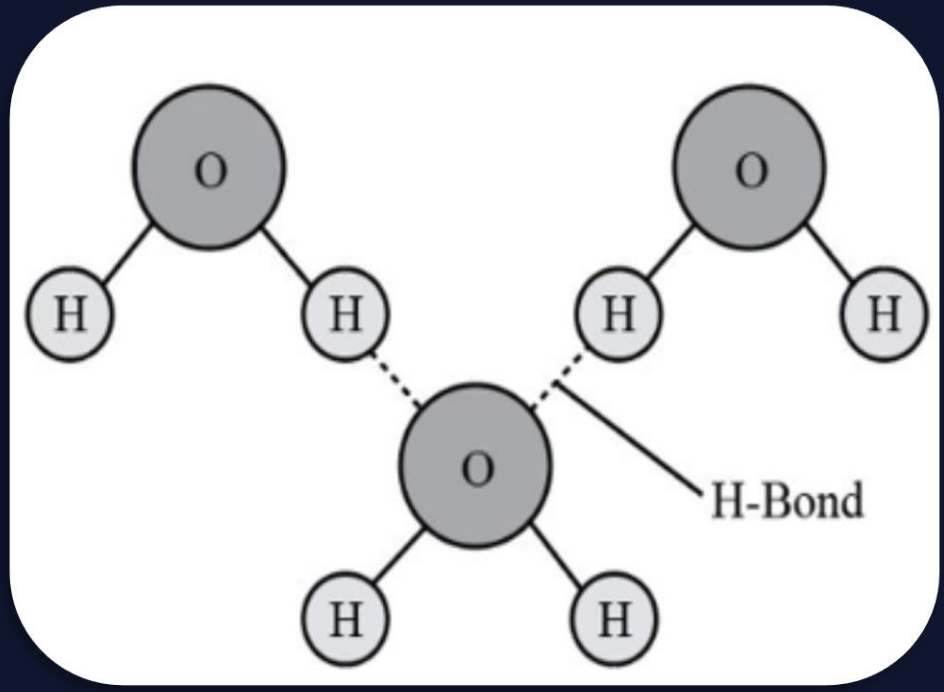
- There are two types of force that hold matter together. These are
  - Intramolecular forces
  - Intermolecular forces.
- **Intramolecular force** are the forces that hold atoms together within a molecule or compound and affects the chemical property of the species.

Intermolecular forces

- Intermolecular forces are the forces that act between molecules and hold them together in a substance.
- These forces are weaker than intramolecular forces but are crucial in determining the physical properties (boiling point, melting point, etc.) of substances.
- There are three primary types of intermolecular forces:
    - Dipole–dipole forces
    - London (or dispersion) forces,
    - Hydrogen bonding forces.

- **Dipole-Dipole Interactions:**
  - Occur between polar molecules where the positive end of one molecule attracts the negative end of another molecule.
  - Example: Hydrogen chloride (HCl) molecules interact through dipole-dipole forces.

- **Hydrogen Bonding:**
  - A special type of dipole-dipole interaction that occurs when hydrogen is bonded to a highly electronegative atom (such as oxygen, nitrogen, or fluorine), resulting in a strong dipole.

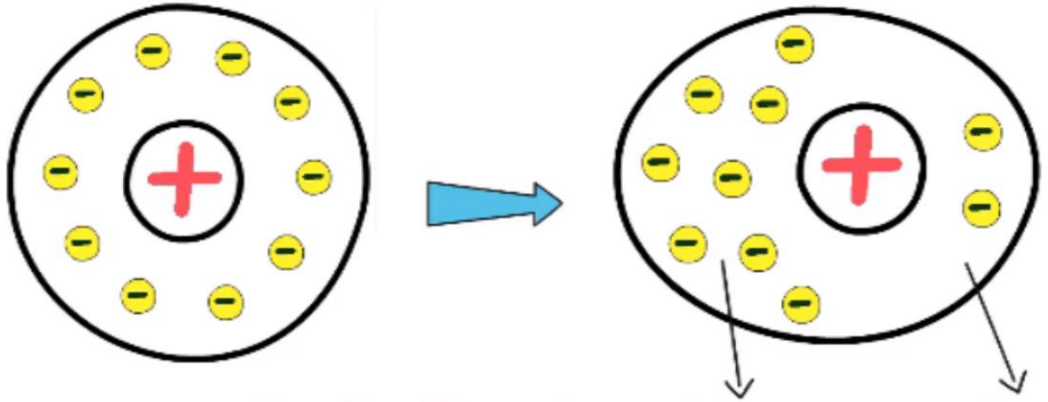

O
O
H
H
H
H
O
H-Bond
H
H

- **Dispersion or London Forces**
  - **The weakest intermolecular force, arising from temporary dipoles induced in atoms or molecules.**

TEMPERORY DIPOLE?

NEUTRAL ATOM

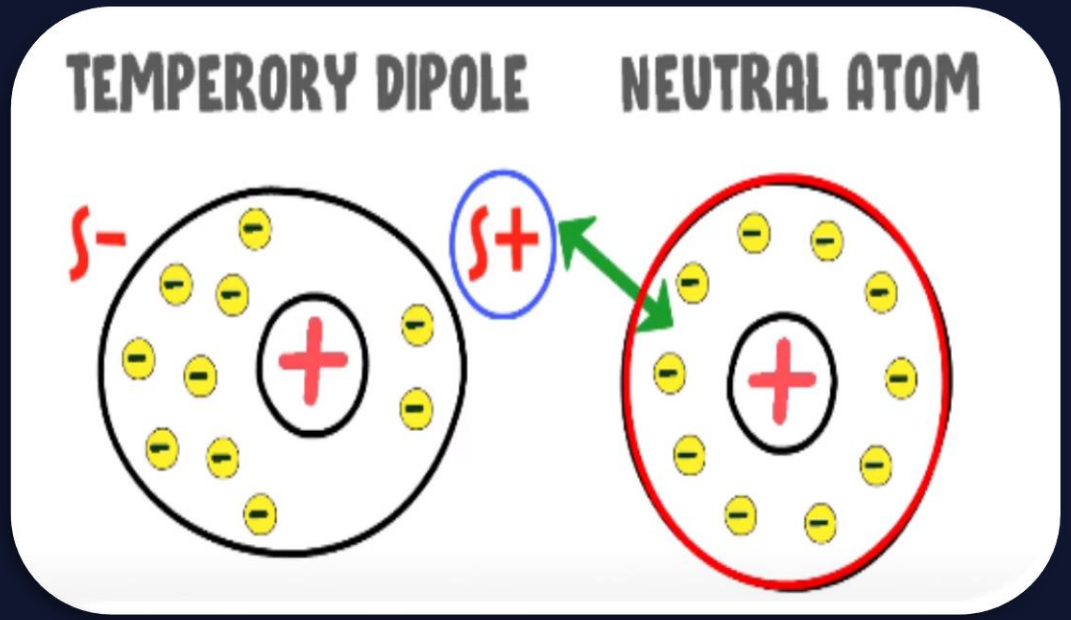
THERE WILL BE
A time when...

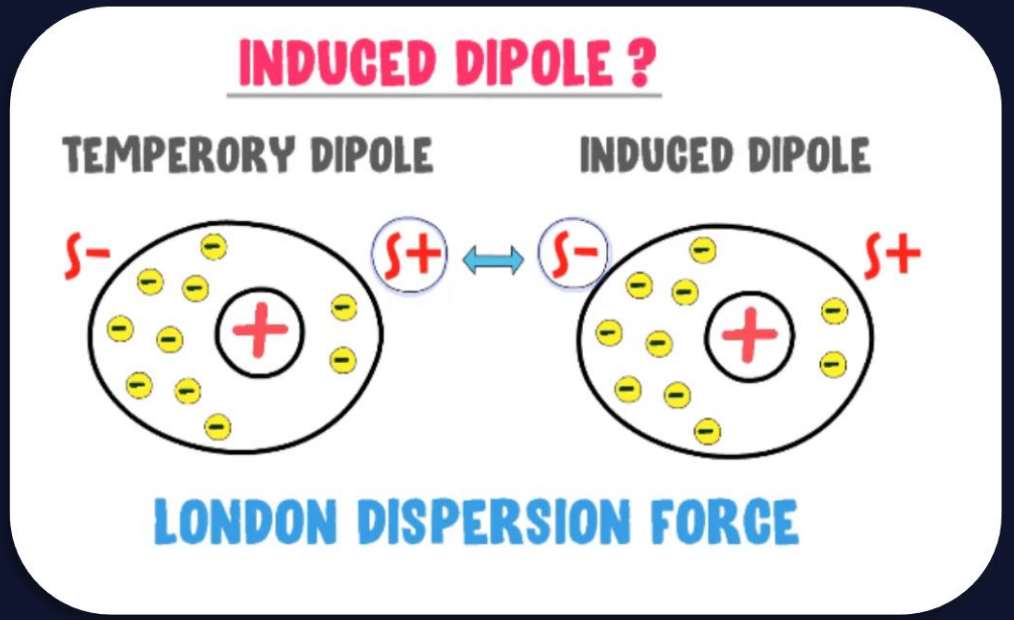


Electrons are equally distributed
Electrons constantly revolve...

More Electrons

Less Electrons


TEMPERORY DIPOLE
NEUTRAL ATOM
δ-
δ+


INDUCED DIPOLE ?
TEMPERORY DIPOLE
INDUCED DIPOLE
δ-
δ+
δ-
δ+
LONDON DISPERSION FORCE

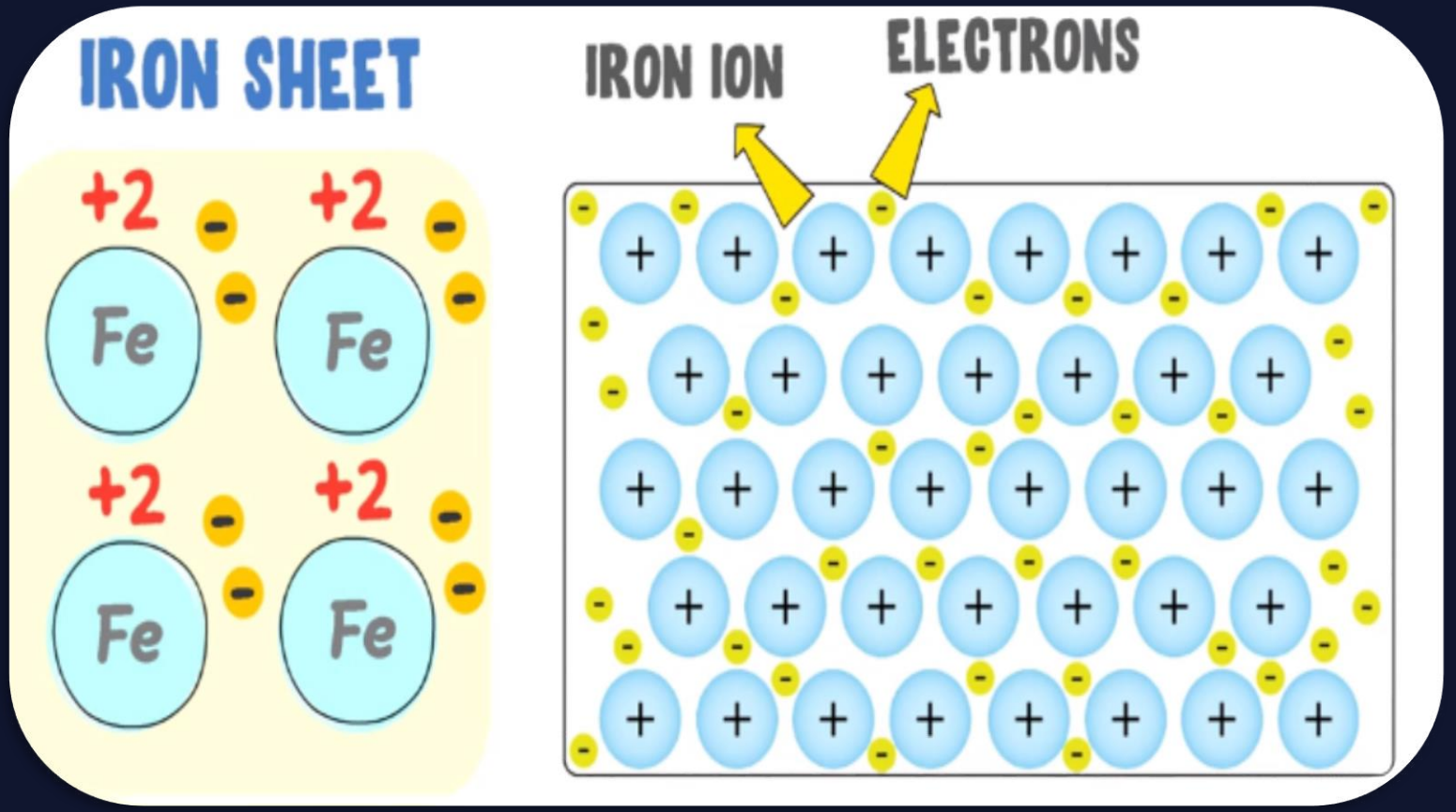
Easy trick to identify intermolecular forces

Hydrogen bonding	dipole-dipole bond	London dispersion
H - F,O,N	Polar molecule	non-polar molecule Polar molecule

Metallic bonding

Formation of metallic bond

- Metallic bonds are the chemical bonds that hold atoms together in solid metals such as copper, iron, and aluminum.
- In these metals, each metal atom is bonded to several neighboring atoms.
- Metallic bonding is the sharing of free electrons (delocalized electrons) among a lattice of positively charged metal ions.


IRON SHEET
IRON ION
ELECTRONS
+2
+2
+2
+2
Fe
Fe
Fe
Fe

The strength of the metallic bond depends on:

1.The number of electrons in the delocalized “sea” of electrons.

- More delocalized electrons result in a stronger bond and a higher melting point.

2. The packing arrangement of the metal atoms.

- The more closely packed the atoms are the stronger the bond is and the higher the melting point.

Electron-sea model

- The electron-sea model is a very simple model, which pictures the metal as an array of metal cations in a “sea” of electrons.
- The metal ions (nuclei plus core electrons) are submerged within this electron sea in an orderly array.

Properties of Metals Due to Metallic Bonding

- ***Electrical Conductivity:***

  The delocalized electrons can move freely, allowing metals to conduct electricity.

- ***Malleability and Ductility:***

  Malleable (can be hammered into thin sheets) and ductile (can be drawn into wires).

- ***High Melting and Boiling Points:***

  The strong electrostatic attraction between metal ions and the electron sea results in high energy required to break the bonds, leading to high melting and boiling points.

- ***Luster:***

  The free-moving electrons can absorb and re-emit light, giving metals their shiny appearance or luster.

Chemical Bonding Theories

- Covalent bonds are more accurately explained using modern bonding theories
  - The valence bond theory (VBT)
  - The molecular orbital theory(MOT)

- **The valence bond (VB) theory** assumes that the electrons in a molecule occupy atomic orbitals of the individual atoms.
- **The molecular orbital (MO) theory**, assumes the formation of molecular orbitals from the atomic orbitals.

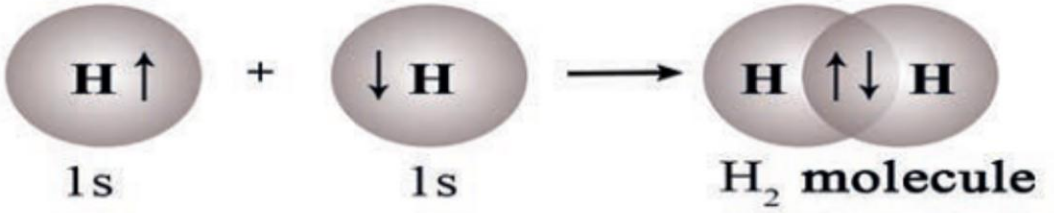


Figure 2.9: Atomic orbital overlap and bonding in H_2

Valence Bond (VB) Theory

- **The basic principle of valence bond theory** is that a covalent bond forms when orbitals of two atoms overlap and the overlap region, which is between the nuclei, is occupied by a pair of electrons.
- By overlap, we mean that the two orbitals share a common region in space.

For example, consider the formation of the H_2 molecule from two hydrogen atoms.

- Each atom has the electron configuration $1s^1$.
- As the H atoms approach each other, their 1s orbitals begin to overlap and a covalent bond forms.
- Electron density is higher in the overlap region than anywhere else,

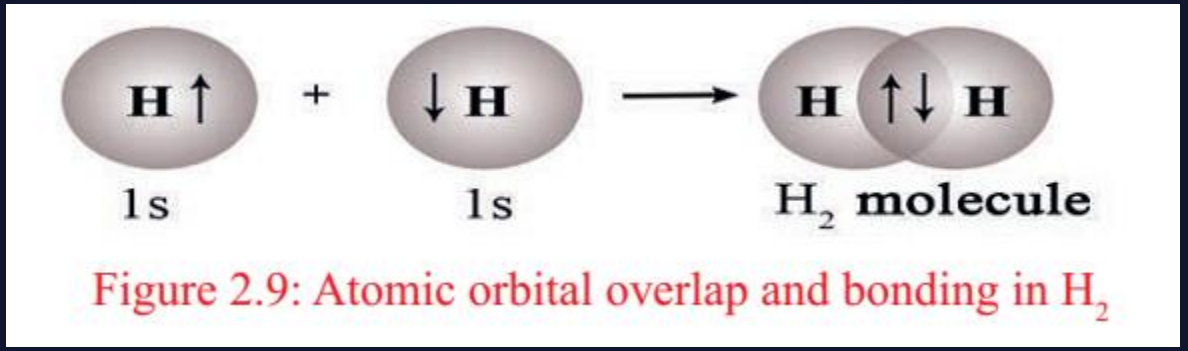


Figure 2.9: Atomic orbital overlap and bonding in H_2

- When two atoms are brought closely together, the repulsion of the atomic nuclei becomes more important than the electron-nucleus attraction and the bond becomes unstable.
- There is a condition of optimal orbital overlap that leads to maximum bond strength (bond energy)at a particular internuclear distance (bond length)

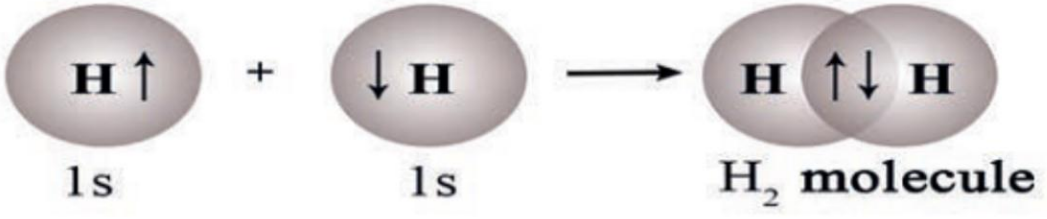


Figure 2.9: Atomic orbital overlap and bonding in H_2

- The bond strength depends on the attraction of the nuclei for the shared electrons, so the greater the orbital overlap, the stronger (more stable) the bond.
- The extent of overlap depends on the shapes and directions of the orbitals.
- Thus, whenever possible, a bond involving p or d orbitals will be oriented in the direction that maximizes overlap.

- In the HF bond, for example, the 1s orbital of H overlaps the half-filled 2p orbital of F along the long axis of that orbital.
- Any other direction would result in less overlap and, thus, a weaker bond.

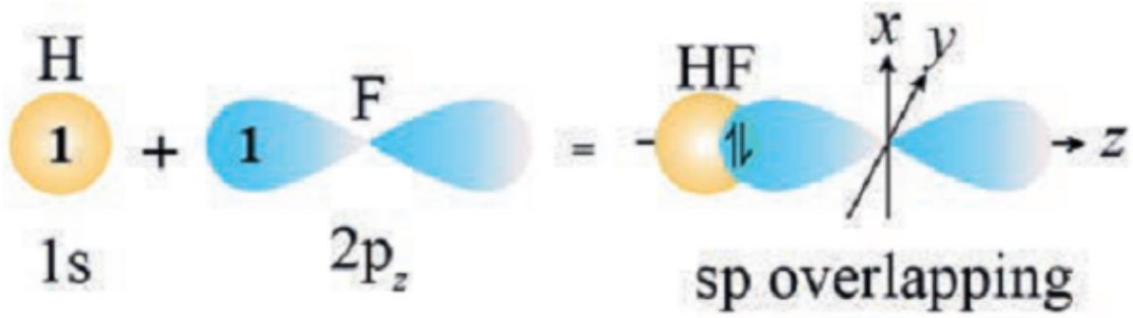


Figure 2.10: Orbital and spin pairing in the formation of a HF molecule

- Similarly, in the F–F bond of F_2, the two 2p orbitals interact end-to-end, that is along the orbital axes, to attain maximum overlap

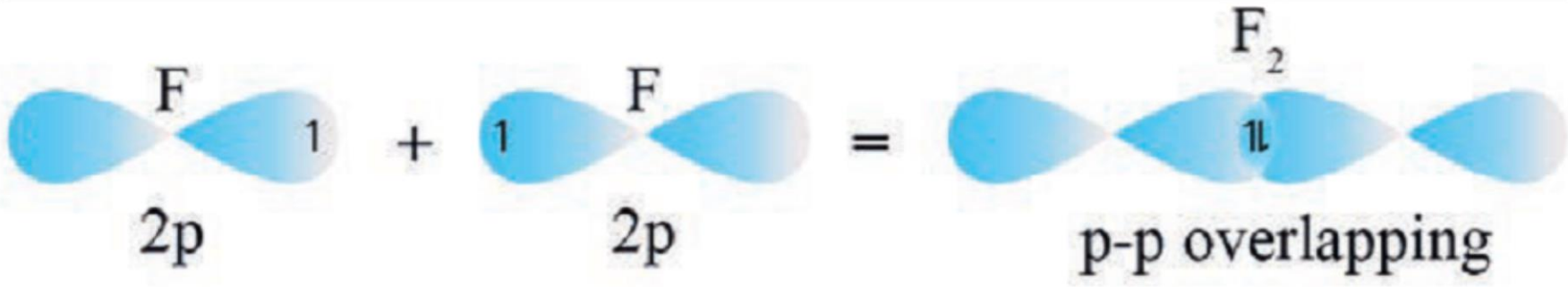


Figure 2.11: Orbital and spin pairing in the formation of a F_2 molecule

- When two atomic orbitals overlap to build up electron density along the two axis between the two nuclei, the resulting localized bond is called a sigma bond (σ).
- Such an overlap is sometimes referred to as a head on, end-to-end, end on or linear overlap.
- In Figures 2.9, 2.10 and 2.11, head-to-head overlap resulted in sigma bonds.

- The shape of p atomic orbitals allows them to overlap not only in a head-to-head fashion.
- If they are oriented in a parallel position, they can overlap sideways or laterally as .A bond formed in this way is called pi-bond (π).

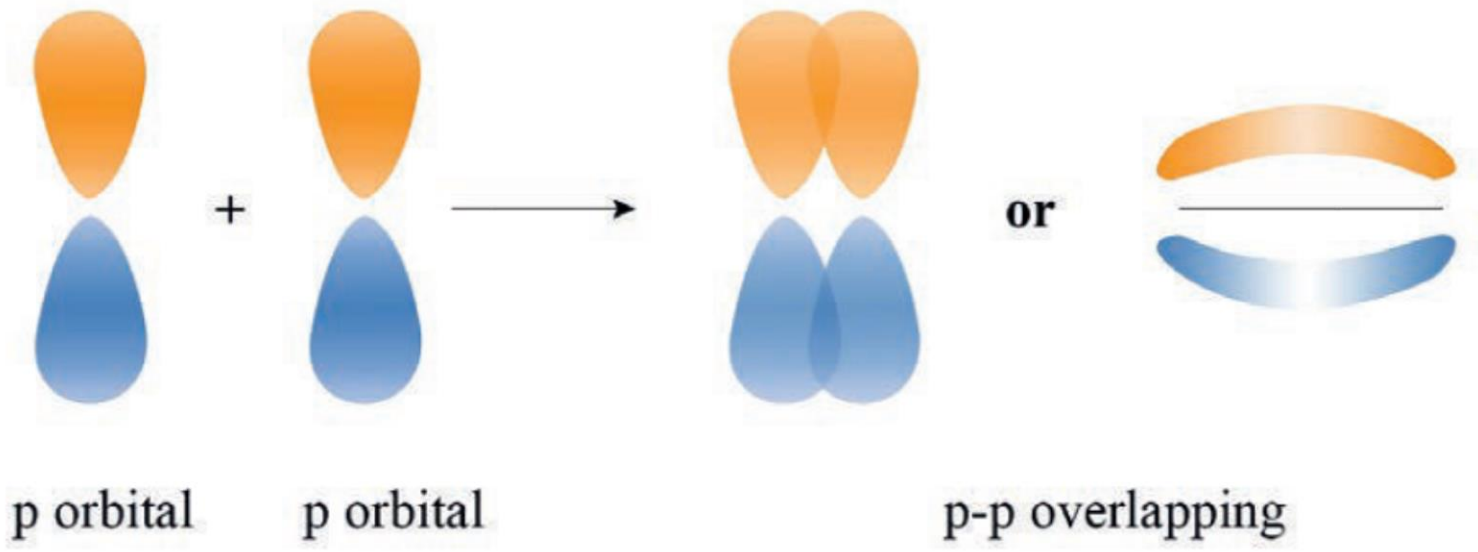
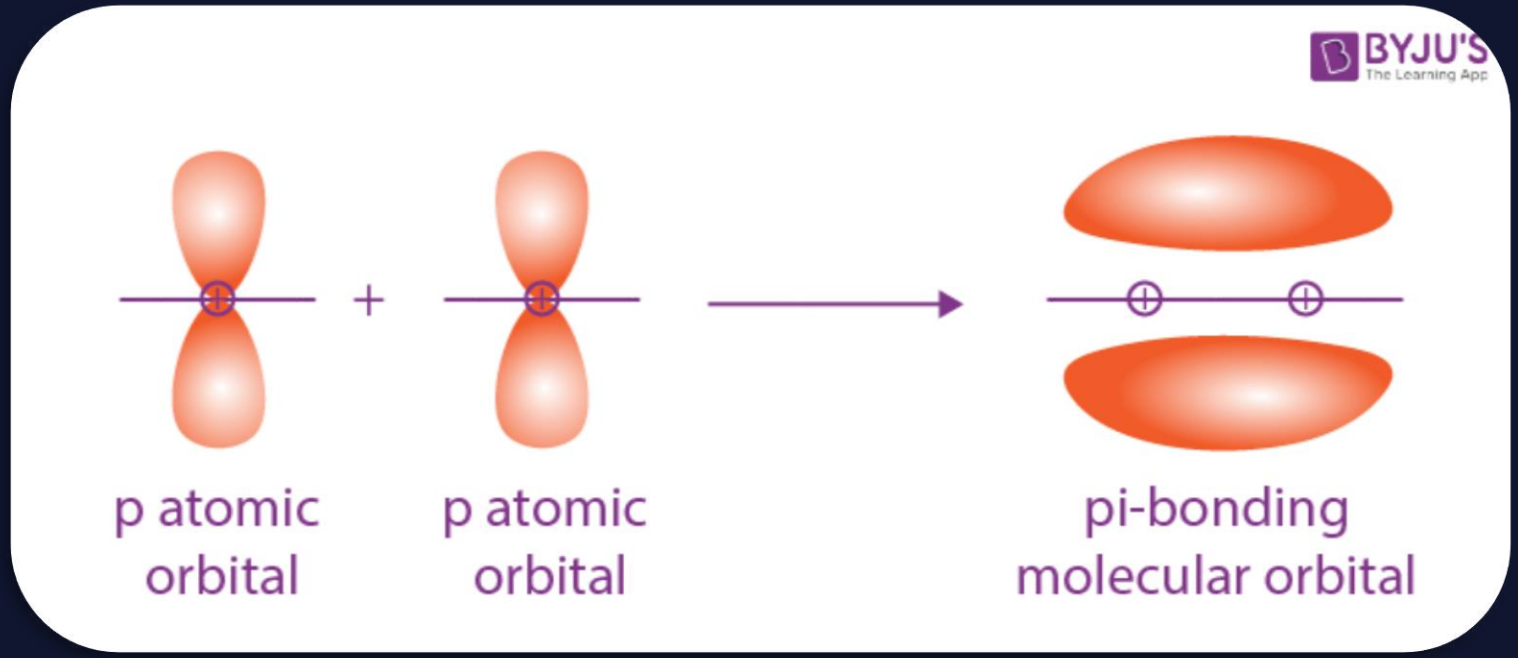


Figure 2.12: Formation of a pi bond (π bond) by the overlap of two half-filled p orbitals that are perpendicular to the internuclear axis


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p atomic orbital
p atomic orbital
pi-bonding molecular orbital

- Note that the directional nature of p orbital allows to overlap in two ways: end-to end or sideways.
- These two modes give rise to the two types of covalent bond: sigma (σ) and pi (π) bond, respectively.
- However, s-s and s-p will always overlap along the nuclear axis, resulting only in sigma bonds.
- Pi bonds are present in molecules containing multiple bonds.

- **Note:**
  - A **single bond** is always a sigma bond.
  - In a **double bond**, one bond is a sigma bond, and the other is a pi-bond.
  - In a **triple bond**, there is one sigma bond and two pi-bonds.

Hybridization of Orbitals

What is hybridization?

- It is an imaginary mixing process in which the orbitals of an atom rearrange themselves to form new orbitals called **hybrid orbitals.**

Why we study hybridization?

- Predicting Molecular Geometry
- Explaining Bond Angles
- understanding Molecular Stability

- **Key Features of Hybridization**
  - The orbitals involved in hybridization must only have a slight variation in energy.
  - The number of hybrid orbitals formed is equal to the number of atomic orbitals mixed.
  - Half filled, completely filled and vacant orbitals can take part in hybridization.
  - Hybridization happens only during the bond formation and not in an isolated gaseous atom.
  - The shape of the molecule can be predicted if the hybridization of the molecule is known.

- Based on the types of orbitals involved in mixing, the hybridization can be classified as
  - sp
  - sp^2
  - sp^3
  - sp^3d
  - sp^3d^2
  - sp^3d^3

- **Sp hybrid orbital**
  - **Orbitals Involved**: 1 s-orbital + 1 p-orbital
  - **Number of Hybrid Orbitals**: 2
  - **Geometry**: Linear
  - **Bond Angle**: 180°
  - **Percent Character:** 50% s character and 50% p character.
  - **Example:** $BeCl_2$, CO_2

- **sp^2 Hybridization**
  - **Orbitals Involved**: 1 s-orbital + 2 p-orbitals
  - **Number of Hybrid Orbitals**: 3
  - **Geometry**: Trigonal Planar
  - **Bond Angle**: 120°
  - **Percent Character**: 33.33% s character and 66.67% p character
  - **Example**: BF_3, SO_3

- **sp^3 Hybridization**
  - **Orbitals Involved:** 1 s-orbital + 3 p-orbitals
  - **Number of Hybrid Orbitals:** 4
  - **Geometry:** Tetrahedral
  - **Bond Angle:** 109.5°
  - **Percent Character**: 25% s character and 75% p character
  - **Example:** CH_4, NH_3, H_2O

- **sp^3d Hybridization**
  - **Orbitals Involved:** 1 s-orbital + 3 p-orbitals + 1 d-orbital
  - **Number of Hybrid Orbitals:** 5
  - **Geometry:** Trigonal Bipyramidal
  - **Bond Angles:** 90°, 120°
  - **Percent Character**: 20% s character, 60% p & 20% d character
  - **Example:** PCl_5, SF_4

- **sp^3d^2 Hybridization**
  - **Orbitals Involved:** 1 s-orbital + 3 p-orbitals + 2 d-orbitals
  - **Number of Hybrid Orbitals:** 6
  - **Geometry:** Octahedral
  - **Bond Angle:** 90°,180°
  - **Percent Character**: 16.67% s, 50% p and 33% d character.
  - **Example:** SF_6, XeF_4

Hybrid Orbitals

2	→	sp	→	linear
3	→	sp^2	→	trigonal
4	→	sp^3	→	tetrahedral
5	→	sp^3d	→	trigonal bipyramidal
6	→	sp^3d^2	→	octahedral

Identify the hybridization of the central atom

- CO_2
- $S0_2$
- BF_4^-
- SF_4
- SF_6
- CO_3^{2-}
- NH_4^+

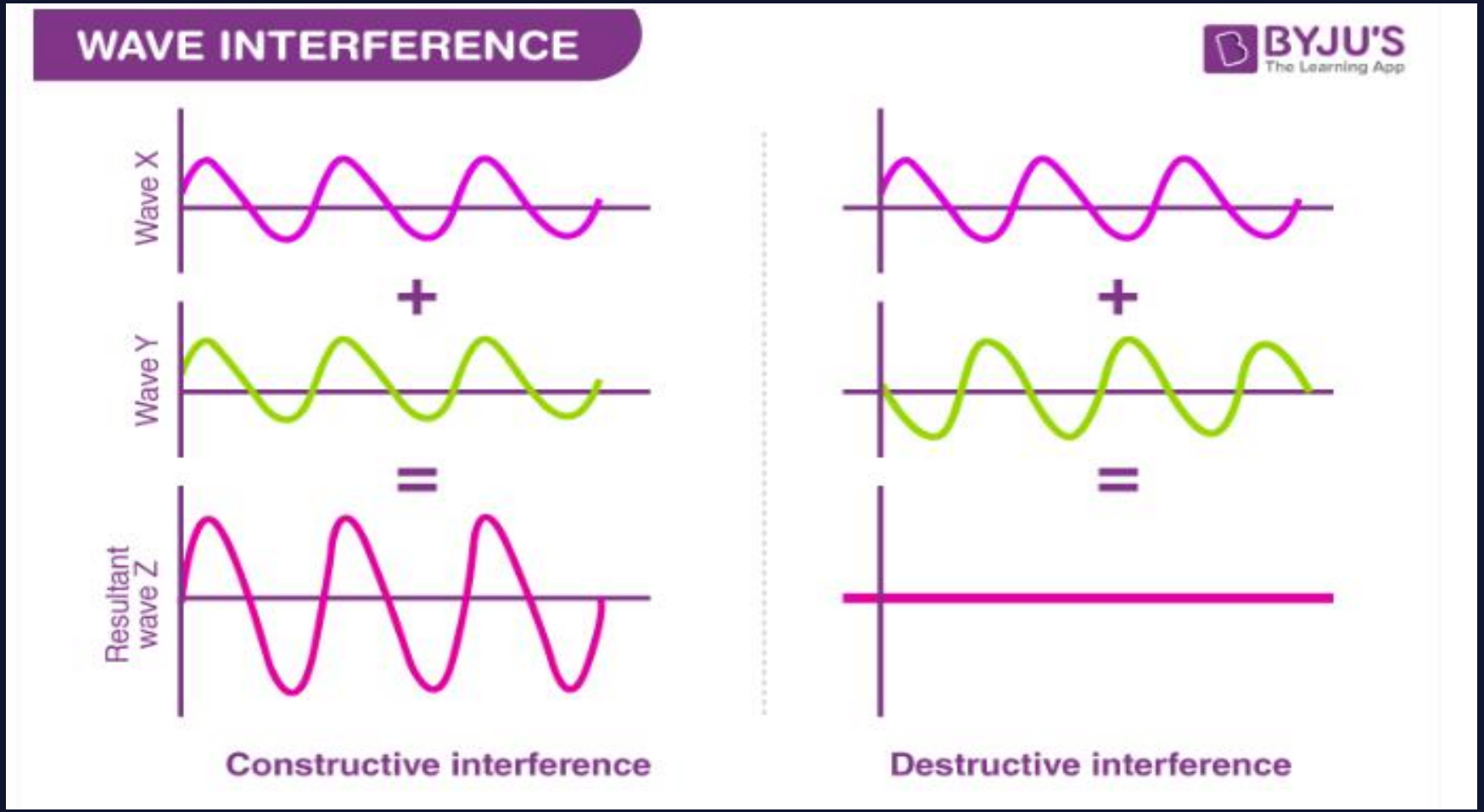
Molecular Orbitals Theory(MOT)

- The molecular orbital theory assumes that when atoms come together, their orbitals not only overlap, but are also simultaneously transformed into new orbitals.
- These new orbitals, called **molecular orbitals (MOs)** These play the same role for molecules that atomic orbitals play for atoms.
- In contrast to VB theory, in which one bonding orbital is formed as a result of the interaction of two atomic orbitals, in MO theory two molecular orbitals result from the combination of two atomic orbitals these are

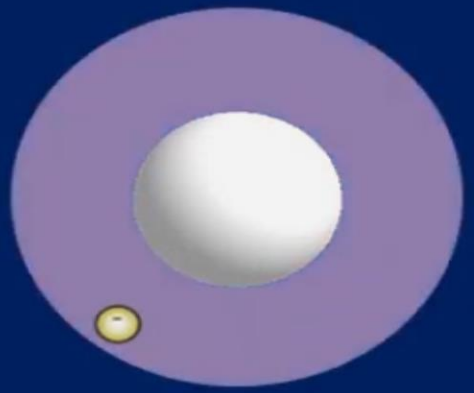
1. Bonding Molecular Orbitals
2. Anti-Bonding Molecular Orbitals

Bonding Molecular Orbitals

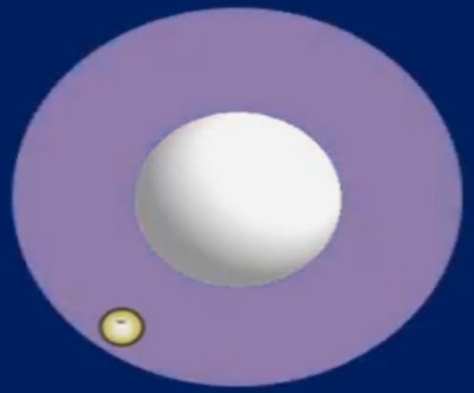
- The molecular orbitals that results from the addition of the atomic orbitals that overlap is called a bonding orbital.
- In a bonding molecular orbital, because the electron density is concentrated between the two nuclei, the energy of the system is lowered compared with that of isolated atomic orbitals.


WAVE INTERFERENCE
BYJU'S
The Learning App
Wave X
+
Wave Y
=
Resultant wave Z
Constructive interference
+
+
=
Destructive interference

Molecular Orbitals Explained



1s

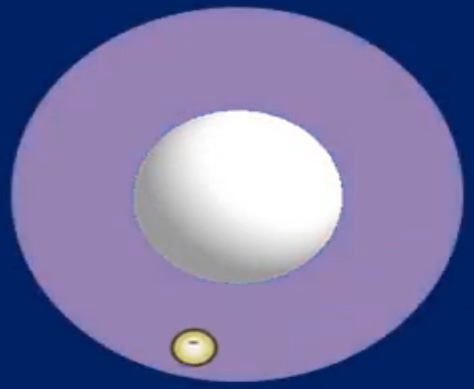


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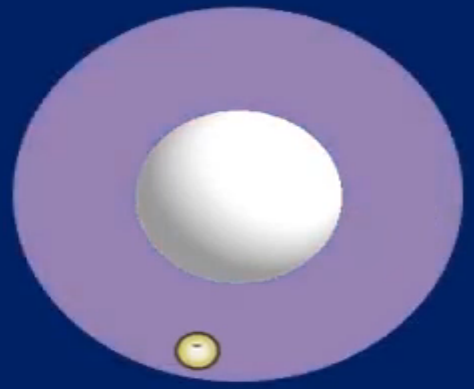
Anti-Bonding Molecular Orbitals

- The molecular orbital that results from the subtraction of atomic orbitals that overlap is called an antibonding molecular orbital.
- In antibonding molecular orbitals, the electron density is concentrated away from the region between the two nuclei.
- That is, anti-bonding molecular orbitals have a region of zero electron density (a node) between the nuclei.
- The net effect of having a low electron density between the nuclei is that the two nuclei repel each other.
- Therefore, energy of an antibonding orbital is increased compared with that of the two isolated atomic orbitals.

Molecular Orbitals Explained



1s



1s

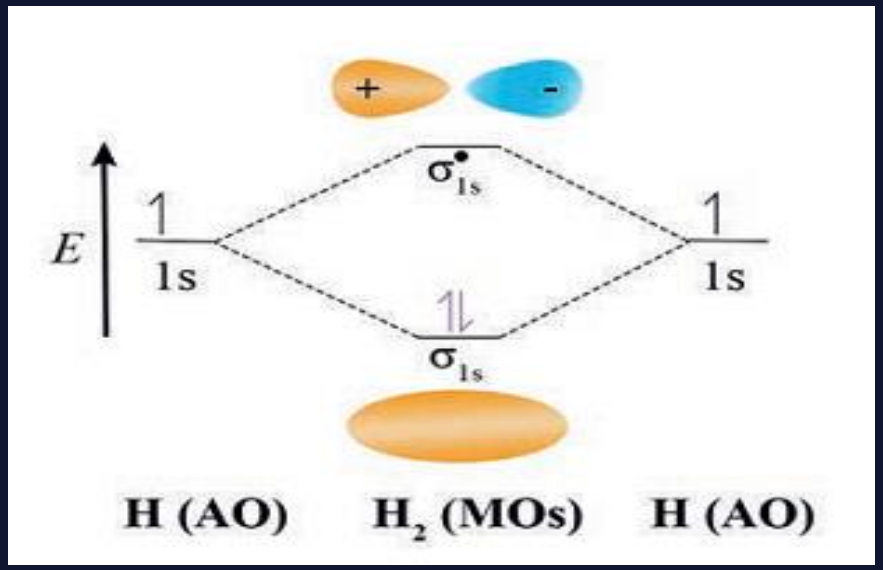
Electron Configuration of Diatomic Molecules

- Electrons occupy molecular orbitals following the same rules that apply to the filling of atomic orbitals: the **aufbau principle** and the **Pauli Exclusion Principle** and **Hund's rule** are obeyed.

a. In simple homonuclear diatomic molecules (total electrons = 14 or less) the order is: $\sigma_{1s}\sigma^*_{1s}\sigma_{2s}\sigma^*_{2s}(\pi_{2p_y}=\pi_{2p_z})\sigma_{2p_x}(\pi^*_{2p_y}=\pi^*_{2p_z})\sigma^*_{2p_x}$

b. For simple homonuclear diatomic molecules (total electrons greater than 14) the $(\pi_{2p_y}=\pi_{2p_z})$ comes after σ_{2p_x} and the order is:

$\sigma_{1s}\sigma^*_{1s}\sigma_{2s}\sigma^*_{2s}\sigma_{2p_x}(\pi_{2p_y}=\pi_{2p_z})(\pi^*_{2p_y}=\pi^*_{2p_z})\sigma^*_{2p_x}$


+
-
E
1
1s
σ1s
1
1s
σ1s
H (AO)
H2 (MOs)
H (AO)

Bond border

- MO theory uses the term **"bond order"** to indicate whether a covalent bond is single (bond order = 1), double (bond order = 2) or triple (bond order = 3).
- **Bond order** is defined as the number of electrons in bonding MOs minus the number in antibonding MOs, divided by two:

$$\text{Bond order} = \frac{1}{2}\left[\begin{pmatrix}\text{Number of } e^- \text{ in}\\ \text{bonding MOs}\end{pmatrix} - \begin{pmatrix}\text{Number of } e^- \text{ in}\\ \text{anti-bonding MOs}\end{pmatrix}\right]$$

- A bond order greater than zero indicates that the molecule is stable relative to the separate atoms, whereas a bond order of zero implies no net stability.
- In general, the higher the bond order, the stronger the bond.

Magnetic Properties

- Molecules with unpaired electrons exhibit paramagnetic properties. The species is attracted by an external magnetic field.
- When all the electrons are paired, there is diamagnetism. Such species are not attracted (and, in fact, are slightly repelled) by a magnetic field.

Types of Solid

- Depending on the arrangement of atoms, there are two types of solid

  Crystalline

  Amorphous

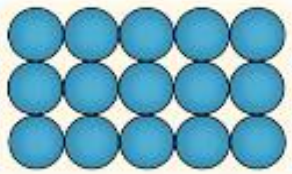
- all crystals are solids, but not all solids are crystals

TYPES OF SOLIDS

CRYSTALLINE SOLIDS

1) Regular arrangement of particles.

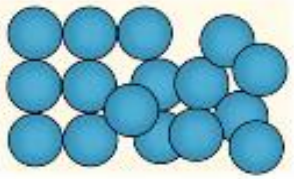
REGULARLY ARRANGED



AMORPHOUS SOLIDS

1) Irregular arrangement of particles.

IRREGULARLY ARRANGED



Types of Crystal

- The crystalline solids can be classified as
  - Ionic crystals
  - Molecular crystals
  - Covalent network crystals
  - Metallic crystals

Ionic Crystals

- **Bonding:** Ionic bonds (electrostatic attraction between oppositely charged ions).
- **Contains:** Ions (e.g., cations and anions).
- **Properties:** Hard, brittle, high melting points, good electrical conductors in molten or dissolved form.
- **Example:** Sodium chloride (NaCl).

Molecular Crystals

- **Bonding:** Weak intermolecular forces like van der Waals forces, hydrogen bonds, or dipole-dipole interactions.
- **Contains:** Molecules (e.g., H_2O, I_2).
- **Properties:** Soft, low melting points, poor electrical conductivity, often volatile usually transparent, brittle, and break easily when stressed.
- **Example:** Ice (H_2O), solid iodine (I_2), dry ice (CO_2).

- **Covalent Network Crystals:**
  - **Bonding:** Strong covalent bonds forming a continuous network across the entire crystal.
  - **Contains:** Atoms (e.g., in a lattice structure).
  - **Properties:** Hard, high melting points, poor electrical conductors (except graphite), insoluble in most solvents.
  - **Example:** graphite ,diamond , silicon dioxide (SiO_2), quartz.

Covalent Network Crystals

- In this type of crystalline solids, separate particles are not present.
- Instead, strong covalent bonds link the atoms together throughout the network of covalent solid.
- As a consequence of the strong bonding, all these substances have extremely high melting and boiling points, but their conductivity and hardness depend on the nature of their bonding.
- graphite and diamond are examples of network covalent solids

- **Metallic Crystals:**
  - **Bonding:** Metallic bonds (delocalized electrons freely moving around a lattice of metal ions).
  - **Contains:** Metal atoms surrounded by a "sea" of delocalized electrons.
  - **Properties:** Good electrical and thermal conductivity, malleable, ductile, high melting points.
  - **Example:** Copper (Cu), iron (Fe), gold (Au).

THE END