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SENIOR FIVE TERM 1

TOPIC 3/3: Bonding and Structure

Competency: The learner analyses the types of chemical bonds and molecular structures, and relates them to the properties and uses of substances in real-life contexts.

Chemical bonding

Chemical bond: A stable attraction between atoms due to the sharing, transfer, or pooling of electrons. The purpose of to atoms bond is to lower their energy and achieve stable electron configurations (often resembling noble gases).

There are three types of chemical bonding namely: -

1. Ionic (electrovalent) bonding.
2. Covalent bonding.
   (a) Normal covalent bonding.
   (b) Dative (coordinate) bonding.
3. Metallic bonding.
4. Van der Waal forces

Ionic bonding

This involves transfer of one or more electrons from one atom to another, e.g.

$$Na \rightarrow Na^+ + e^-$$

$(1s^22s^22p^63s^1)$ $(1s^22s^22p^6)$

$$Cl + e^- \rightarrow Cl^-$$

($1s^22s^22p^63s^2\,3p^5$)   ($1s^22s^22p^6\,3s^2\,3p^6$)

By losing one electron a sodium atom achieves the stable electronic structure (similar to that of neon), and addition of an electron to a chlorine atom, causes it to achieve a stable electronic configuration (similar to that argon).

It is the electrostatic attraction resulting between the opposite charges that constitute the ionic/electrovalent bond between Na^+ and Cl^- ions in sodium chloride.

Factors that favor ionic bonding

(i) One of the two atoms involved should have **lower ionization energy** and the other must have a **higher electron** affinity. For instance, elements such as aluminium metal form compound most of which are not ionic due to the high ionization energy of aluminium (i.e., a large amount of energy is needed to remove all the 3 valence electrons from an aluminium atom).

(ii) The **electronegativity difference** of the two atoms involved must be high. For example, a compound formed between an element of group 1A and an element in group 7B is more ionic than one formed with a group 2A element and group 7B element. Usually ionic compounds are formed between elements of the left-hand side of the Periodic Table and elements of the right-hand side of the Periodic Table.

Characteristic properties of ionic compounds

(i) They consist of ions and not molecules.

(ii) They are electrolytes, i.e., when dissolved in water or when fused, conduct an electric current with decomposition.

(iii) They are solids of high melting points. The melting point increases with the strength of ionic character for instance the melting point of NaCl > NaBr > NaI because the strength ionic bond decrease in the same order due to decreasing electronegativity of the anions

(iv) They are insoluble in organic solvents but many of them are soluble in water.

Factors affecting the strength of ionic bond

(i) size of the cation and anion; ionic bonds are stronger when the size of the cation is big and that of the anion is small due to low polarization, e.g. NaCl is more ionic than LiCl

(ii) Electronegativity/electropotivity differences; the bigger the difference of these factor between the reacting elements, the stronger the ionic bond e.g. NaCl is more ionic than NaI

Polarization of ions

Polarization occurs when a positively charged ion (cation) attracts and distorts the electron cloud of a negatively charged ion (anion), Fig. 1

Result: The anion's electron cloud is pulled toward the cation, reducing the purely ionic nature of the bond and introducing covalent character

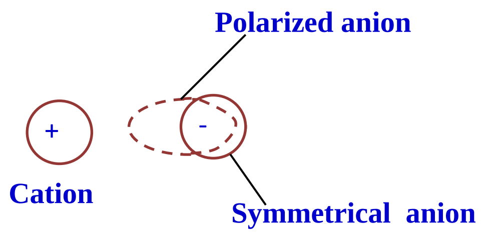


Fig. 3. 1 Polarization of ions

Consequences of Ion Polarization

(i) Covalent character in ionic compounds**:** Polarization explains why some ionic compounds (like $AlCl_3$) show covalent properties.
(ii) Variation in melting points**:** Stronger polarization lowers lattice energy, reducing melting points compared to purely ionic compounds.
(iii) Solubility differences**:** Polarization affects how compounds dissolve in polar vs nonpolar solvents.
(iv) Color and conductivity**:** Distorted electron clouds can influence optical and electrical properties.

Factors affecting the magnitude of polarizing power of a cation

(a) **Magnitude of positive charge on the cation:** the higher the positive charge on the cation, the more strongly it attracts the outermost shell electron cloud of the anions towards itself and hence polarizes a given anion more strongly. E.g. polarizing power of isolectric cation is in order $Na^+ < Mg^{2+} < Al^{3+}$. This explains why salts of aluminium are covalent and thus soluble in organic solvents have and lower boiling points than those of sodium salts.

Trial 3.1

The table below shows the ionic radii of sodium, magnesium and aluminium and the melting points of the chloride formed by the three cations

	Na^+	Mg^{2+}	Al^{3+}
Ionic radium (nm)	*0.095*	*0.065*	*0.050*
Mpt./^{0}C	*801*	*712*	*180(sublimes)*

(i) *Calculate the charge radius ratio for each cation* *(1½)*
(ii) *State and explain the trend in the melting points of the chlorides of the three metals (5marks)*

(b) **Size of cation:** small cations have high polarizing power than big cations of the same charge i.e. $Li^+ > Na^+ > K^+$. Small positive ions with high charge will have the highest electrical fields and therefore the highest polarizing power. That is why compounds of lithium, beryllium and aluminium are much more covalent than corresponding compound of other elements in the periodic table group.

(c) Electron configuration:

If we compare the polarizing power of two cations which have the same positive charge on them and same or almost same size, then cations having $ns^2p^6d^{1-10}$ as its valence-shell configuration will polarize the given anion more strongly than the cation having ns^2np^6 as its valence configuration because d-electron shield the nucleus poorly compared to s- and p-electron shells. For example, although Cu^+ and Na^+ ions have almost the same size (Cu^+ =96pm, Na^+ =96) and the same charge (=+1), Cu^+ ion polarize Cl^- ion in CuCl more strongly than Na^+ in NaCl because of the fact that Cu^+ ion has $3s^23p^6d^{10}$ configuration while sodium ion has $2s^22p^6$ configuration (8electron). The $3d^{10}$ electrons in Cu^+ shield nucleus less effectively than the $2s^2$ and $2p^6$ electrons in Na^+ ion do.

Polarisability

Polarisability is a measure of how easily the electron cloud of an atom, ion, or molecule can be distorted by an external electric field or by the presence of a nearby charged particle. The larger an ion or an atom, the more diffuse electron clouds and the more its electron is easily distorted → higher polarisability.

Factors affecting the polarisability

1. **Magnitude of the negative charge on the anion**. The higher the negative charge on the anion, the more easily and strongly its outer-most electron cloud attracted by the cation.
2. **Size of the anion:** the larger is the size of the anion the higher its polarisability will be; that is the reason why fluorides are more ionic than other halides and oxides are more ionic than sulphides.

Pajan's rules convalency

Fajan's rules predict whether a bond will be more **ionic** or **covalent** based on the ability of a cation to **polarize** an anion's electron cloud.

Key Points

- **Small cations with high charge** → strong polarizing power.
- **Large anions with high charge** → easily polarizable.
- **Result:** Distortion of the anion's electron cloud leads to **sharing of electrons** → covalent character.

Factors Increasing Covalency (According to Fajan's Rules)

(i) **Size of cation:** Smaller cations (e.g., Al^{3+}, Be^{2+}) strongly attract and distort anions.
(ii) **Charge of cation:** Higher charge increases polarizing power (e.g., $Al^{3+} > Mg^{2+} > Na^{+}$).
(iii) **Size of anion:** Larger anions (e.g., I^{-}, S^{2-}) are more easily distorted.
(iv) **Charge of anion:** Highly charged anions (e.g., O^{2-}, S^{2-}) are more polarizable.
(v) **Electron configuration:** the cation with $ns^2p^6d^{10}$ configuration has high polarizing power

Examples

- **NaCl:** Large cation (Na^{+}) with low charge → mostly ionic.
- **$AlCl_3$:** Small, highly charged cation (Al^{3+}) → strong polarization → significant covalent character.
- **BeF_2 vs BeI_2:** Be^{2+} polarizes I^{-} more than F^{-}, so BeI_2 is more covalent than BeF_2.

Trial 3.2

(a) Explain what is meant by the term melting point (1mark)
(b) State the factors which affect the melting points of
  (i) Metals
  (ii) Non metals
(c) Explain the trend in melting points of the elements in group II and group (VII) of the periodic table (8marks)

(d) Explain why the transition metals of period 4 tend to have higher melting points than nom transition metals of the same period (2marks)

(e) *The table below shows the melting points of some compounds*

Compound	*melting point/K*
Aluminium oxide	*2290*
Aluminium chloride	*451*
Calcium oxide	*2850*
Calcium chloride	*1051*

Explain why:

(i) *The melting point of aluminium chloride is abnormally low compared to calcium chloride (2marks)*

(ii) The melting point of calcium oxide is much higher than that of calcium chloride (2marks)

Covalent bonding

This is divided into two types, namely, the normal covalent bond and the dative or a coordinate bond.

(a) The normal covalent bonding

This involves sharing one or more pair of electrons to attain a stable electron configuration similar to that of noble gases. Considering the chlorine atom, which has seven electrons in its outer shell, if one electron is provided by each atom and shared equally, then each chlorine atom can acquire a share in eight electrons - a complete octet configuration as it is called (figure 3.1). The shared pair of electrons constitutes a **covalent bond**.

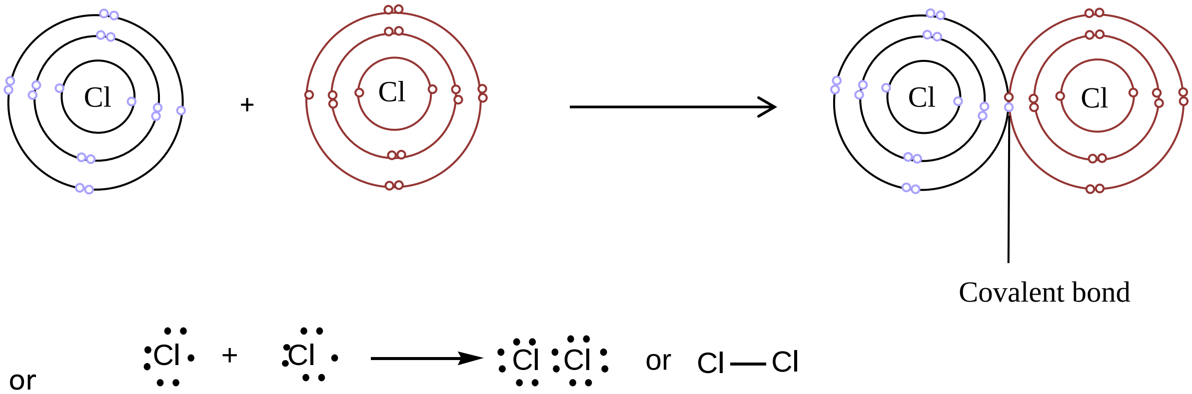


Fig. 3.2. Formation of a covalent bond by chlorine atoms

Factors that favor covalent bonding

(i) The elements involved must have high ionization energy or low electron affinity.

E.g. the ionization energy of hydrogen is quite high, the electron affinity is low, and therefore formation of either H^+ or H^- is not favored.

(ii) The electronegativity differences of the elements involved must be very small.

Characteristic of covalent compounds

a) They consist of discrete molecules.

b) They are non-electrolytes.

c) They have low boiling point and melting point.

d) They are insoluble in water and other polar solvents but soluble in organic or non polar solvents.

(b) Dative or coordinate bonding

The dative bond is like a covalent bond once formed except that both electrons in the shared pair are provided by one atom. The atom providing the two electrons is called the **donor** and the atom which accepts the two electrons is the **acceptor**. The donor atom must have an unshared paired of electrons available and such a pair of electrons is called a **lone pair**. An example is reaction of ammonia with H^+ or $AlCl_3$. By reacting with ammonia, Al^{3+} attains an octet configuration whereas H^+ attains doublet structure of helium.

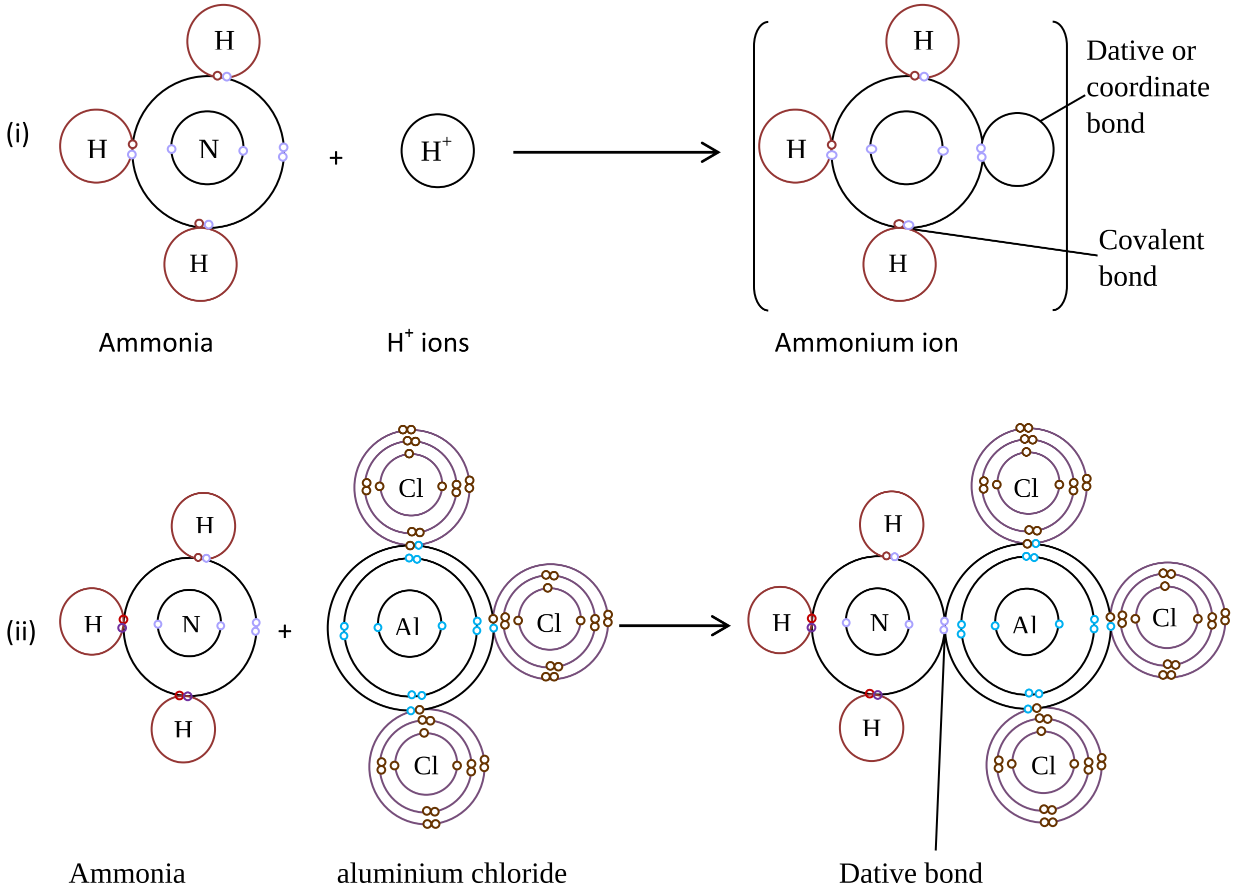


Fig. 3.3. Formation of dative bonds between ammonia and (i) hydrogen ion and (ii) aluminium chloride

Bond polarity

Bond polarity refers to the unequal sharing of electrons between two atoms in a covalent bond due to differences in their **electronegativity**.

If two atoms have different electronegativities, the shared electrons are pulled closer to the more electronegative atom.

For instance, the molecule HF, as fluorine is more electronegative than hydrogen, a fluorine atom pulls the shared electron pair toward itself, i.e., $H^{\delta+} \rightarrow F^{\delta-}$.

This creates a **partial negative charge (δ^-)** on the more electronegative atom and a **partial positive charge (δ^+)** on the less electronegative atom. The bond becomes of intermediate type i.e., partly covalent and partly ionic.

The larger the electronegativity difference between the atoms involved in the bond, the greater the polarity will be. The attraction or repulsion of bonding electrons by an atom or a group of atoms is called **inductive effect.** A dipole moment is the measure of degree of polarity of the bond:

When two and opposite charges **+e** and **-e** are at a distance, **l**, the system is called a dipole and is characterized by dipole moment (**μ**), which is the product of the charge **e** and the distance **l** between the charges, i.e., **μ** = e **l**

The dipole is indicated by using an arrow whose head points to the negative end and the tail at the positive end of the molecules, i.e., H→F. If two dipoles come together head to tail H→F; there is a net attraction between the hydrogen and fluoride atoms that are closest together. i.e

H→F ………H→F

However, a molecule made up of more than two atoms and having some bond polarity may or may not have a dipole moment. This depends on whether the molecule is symmetrical or not. For the symmetrical molecules the individual dipole moments of the bonds in the entire molecule cancel, e.g., the unsymmetrical trichloromethane ($CHCl_3$) has a dipole moment well as tetra chloromethane (CCl_4) a symmetrical molecule has none.

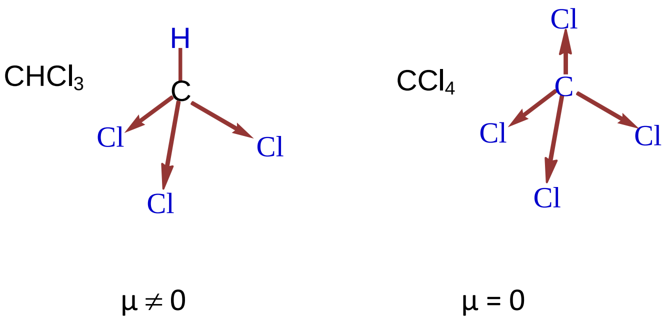


Fig. 3.4. shows the bonding in $CHCl_3$ and CCl_4

Therefore, the lack of a dipole moment where it might be expected is an indication of symmetry within a molecule.

Application of bond polarity

(i) **Predicting molecular polarity:** Bond polarity helps determine whether a molecule is overall polar or nonpolar. This affects solubility, boiling/melting points, and intermolecular forces. *Example:* Water (H_2O) is polar, while methane (CH_4) is nonpolar. **Hydrogen floride is highly polar;** it is soluble in water and has high melting and boiling points than expected.

(ii) **Explaining solubility:** Polar molecules dissolve in polar solvents (like water), while nonpolar molecules dissolve in nonpolar solvents (like hexane). *Rule:* "Like dissolves like."

(iii) **Intermolecular forces:** Polar bonds lead to dipole–dipole interactions and hydrogen bonding, which strongly influence physical properties. *Example:* Hydrogen bonding in DNA base pairs.

(iv) **Chemical reactivity:** Bond polarity creates partial charges (δ^+ and δ^-) that guide chemical reactions, especially nucleophilic and electrophilic attacks. *Example:* In H–Cl, Cl^{δ^-} attracts nucleophiles, while H^{δ^+} is attacked by bases.

(v) **Material properties:** Polarity affects conductivity, dielectric constants, and refractive indices, making it crucial in designing polymers, semiconductors, and liquid crystals.

(vi) **Biological systems:** Polar bonds in water and biomolecules explain protein folding, enzyme activity, and cell membrane behavior.

Trial 3.3

Explain the following observations

(a) *The dipole moment of $BeCl_2$, CO_2, BF_3 is zero while that of H_2S, SO_2 is not.*

(b) *CCl_4 is non-polar while $CHCl_3$ is polar.*

(c) *carbon tetrachloride molecules is not polar yet the bonds in carbon tetrachloride are polar (03marks each)*

Hydrogen bonding

A hydrogen bond is a dipole - dipole attraction that occurs between a hydrogen atom attached to a strong electronegative atom and a second strong electronegative atom with a lone pair of electrons.

The hydrogen bonds in HF are so strong that, they persist even in the vapor state and gaseous hydrogen fluoride consist mainly a mixture of H_2F_2 and H_3F_3 molecules. The hydrogen bonding of aqueous HF is shown in figure 3.5.

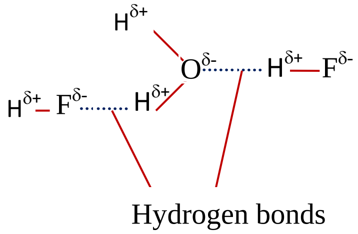


Fig. 3.5. Shows hydrogen bonds in a solution of HF in water.

Effects of hydrogen bonding

Physical Effects

(i) **Higher boiling and melting points:** Compounds with hydrogen bonding (like H_2O, HF, NH_3) have unusually high boiling/melting points compared to similar molecules without hydrogen bonding.
(ii) **Increased viscosity and surface tension**: Water's ability to form multiple hydrogen bonds makes it "stickier," giving it high surface tension and viscosity.
(iii) **Density anomalies**: Ice is less dense than liquid water because hydrogen bonds create an open lattice structure.

Chemical Effects

(i) **Solubility**: Polar molecules capable of hydrogen bonding dissolve more easily in water (e.g., alcohols, sugars).
(ii) **Stability of compounds**: Hydrogen bonding stabilizes structures like hydrogen fluoride chains or carboxylic acid dimers.
(iii) **Reaction pathways**: Hydrogen bonds influence acidity/basicity and reaction mechanisms by stabilizing intermediates.

Biological Effects

(i) **DNA structure**: Hydrogen bonds between base pairs (A–T and G–C) hold the double helix together.
(ii) **Protein folding**: Secondary structures (α-helices, β-sheets) are stabilized by hydrogen bonds.
(iii) **Enzyme activity**: Hydrogen bonding helps substrates bind to active sites, guiding biochemical reactions.

(iv) **Water's role in life**: Hydrogen bonding explains water's solvent power, thermal stability, and importance in biological systems.

Trial 3.3

1. *Explain the following observations: The boiling point of water is 100^0C while that of H_2S is -60.75^0C.*
2. *The table below shows the boiling points of the hydrides of some elements a group of the Periodic Table*

Hydrides	Boiling point in ^0C
NH_3	- 33
PH_3	- 90
ASH_3	- 55
SbH_3	- 17

Explain

(i) *the abnormally high boiling point of NH_3.*
It is due to stronger hydrogen bonds compared to other hydrides in the group

(ii) *The general trend in the boiling point of PH_3, ASH_3, and SbH_3*
Boiling points increase down the group due to increase in molecular mass and strength of molecular bonds

Metallic bonds

Metallic bonding is the type of chemical bond that holds metal atoms together in a solid. It arises from the attraction between positively charged metal ions and a "sea" of delocalized electrons that move freely throughout the structure.

In the metallic bond; each metal atom releases its valence electrons in a crystal pool. It is these `free' electrons `cement' the positive ions together.

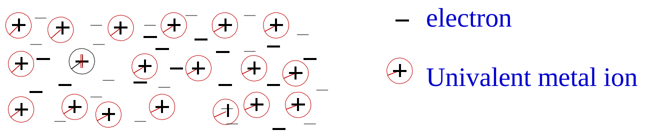


Fig. 3.6. Cross section metallic crystal.

Factors that favor metallic bond

(i) **Low ionization energy:** Metals with low ionization energy can easily release valence electrons, allowing them to become delocalized. *Example:* Alkali metals (Na, K) readily lose their outer electron.
(ii) **High number of valence electrons:** More delocalized electrons strengthen the metallic bond. *Example:* Magnesium (Mg, 2 valence electrons) forms stronger metallic bonds than sodium (Na, 1 valence electron).
(iii) **Small atomic size:** Smaller metal ions pack closely, increasing attraction between ions and the electron cloud. *Example:* Lithium (Li) has stronger metallic bonding than cesium (Cs).
(iv) High positive charge on ions: Metals with higher charges (like Al^{3+}) exert stronger **attraction on delocalized electrons, favoring metallic bond formation.**
(v) **Crystal lattice structure:** Closely packed arrangements (FCC, HCP) maximize overlap between ions and electrons, stabilizing metallic bonds.

Properties Explained by Metallic Bonding

(i) **Electrical conductivity:** Free electrons carry charge easily.
(ii) **Thermal conductivity:** Electrons transfer kinetic energy rapidly.
(iii) **Malleability and ductility:** Non-directional bonds allow atoms to slide past each other without breaking the bond.
(iv) **Luster:** Free electrons reflect light, giving metals their shiny appearance.
(v) **High melting and boiling points:** Strong attraction between ions and electron cloud requires significant energy to break.

Examples

- **Copper (Cu):** Excellent conductor due to strong metallic bonding.
- **Iron (Fe):** Metallic bonds explain its strength and ability to form alloys.
- **Aluminum (Al):** Lightweight but strong, thanks to metallic bonding.

Factors that affect the strength of metallic bonds

(i) **Number of delocalized electrons:** More valence electrons available for bonding → stronger metallic bonds. *Example:* Magnesium (Mg, 2 valence electrons) has stronger metallic bonds than sodium (Na, 1 valence electron).
(ii) **Charge of the metal ions:** Higher ionic charge increases attraction between ions and electrons. *Example:* Al^{3+} forms stronger metallic bonds than Na^{+}.

(iii) **Size of the metal ions:** Smaller ions hold electrons more tightly, strengthening the bond. *Example:* Lithium (Li) has stronger metallic bonds than potassium (K).

(iv) **Packing of metal ions:** Closely packed structures (like FCC or HCP lattices) enhance bonding strength.

(v) **Electronegativity of the metal:** Metals with higher electronegativity attract delocalized electrons more strongly, increasing bond strength. *Example:* Lithium (Li) has stronger metallic bonds than potassium (K).

Molecular or van der Waals bond

A **van der Waals bond** (sometimes called a **molecular bond**) refers to weak intermolecular forces that arise from temporary or permanent dipoles in molecules

Van der Waals forces are weak forces the reason why substances bounds by these forces like iodine have low melting and boiling point. In fact iodine sublimes at room temperature.

Molecular structures (shapes)

Molecular structures are important in the following ways:

Chemical Importance

- **Bond polarity and molecular polarity:** Shape decides whether dipoles cancel out or reinforce each other, influencing solubility and reactivity.
- **Reaction mechanisms**: The orientation of atoms affects how molecules collide and react.
- **Intermolecular forces**: Shapes determine hydrogen bonding, van der Waals forces, and dipole–dipole interactions.

Biological Importance

- **Enzyme activity**: Enzymes have active sites shaped to fit specific substrates (“lock and key” model).
- **DNA structure**: The double helix shape allows base pairing and genetic information storage.
- **Protein folding**: Secondary and tertiary structures depend on molecular geometry, which dictates biological function.

Material Science Importance

- **Crystal structures**: Shapes of molecules influence packing, hardness, and conductivity in solids.
- **Polymers**: The geometry of repeating units affects flexibility, strength, and thermal properties.
- **Nanotechnology**: Molecular shapes are engineered to create specific functions in sensors, catalysts, and drug delivery systems.

Prediction of molecular structures

The **shape of a molecule** is determined by the arrangement of atoms and electron pairs around the central atom. This is explained by the **Valence Shell Electron Pair Repulsion (VSEPR) theory**, which states that electron pairs repel each other and arrange themselves to minimize repulsion.

Diatomic molecules, like H_2, Cl_2, O_2, HCl, HF are linear in shape. However, for poly atomic molecules, i.e., H_2O, CH_4, CO_2, NH_3, etc., the shape adopted by their molecules depends on: -

1. The total number of electron pair around the central atom.
2. The number of bonding electron pair and lone pairs.

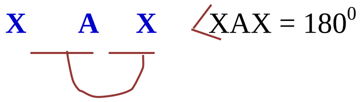
Case 1

Total number of electron pairs around the central atom = 2.

Number of bonding pairs = 2.

Number of lone pairs = 0.

The molecule will be of the form AX_2 and is linear in shape.



$\angle XAX = 180^0$

Case 2

Total number of electron pair around central atom = 3

Two possibilities:

a) Total number of electron pairs around the central atom = 3.

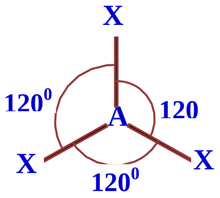
Number of bonding pairs = 3.

Number of lone pairs = 0.

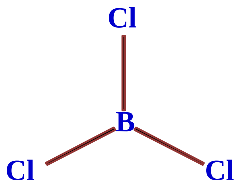
The molecule will be of the form AX_3.

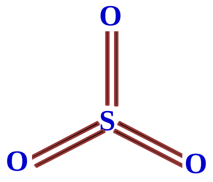
Shape: triangular or trigonal planar, bond angle 120^0.

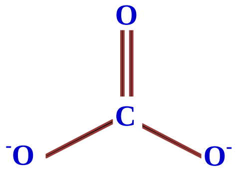
i.e.



Examples: BCl_3, SO_3, CO_3^{2-}







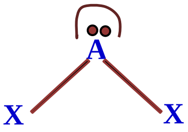
b) Total number of electron pairs around the central atom = 3.

Number of bonding pairs = 2.

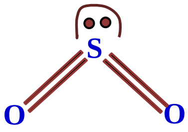
Number of lone pairs = 1.

The molecule will be of the form AX_2 but with a lone pair

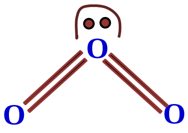
Shape: Angular /V-shaped.



e.g. SO_2



O_3



Case 3

Total number of electron pairs around central atom = 4

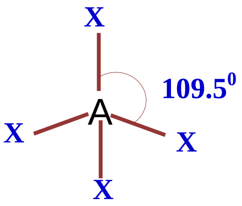
hree possibilities:

a) All the four pairs are bonding pairs.

No. of lone pairs = 0

Form of the molecule = AX_4

Shape Tetrahedral, bonding angle = 109.30' or 109.5^0



Examples, CH_4, CCl_4, SO_4^{2-}, PO_4^{2-}, MnO_4-, NH_4^+

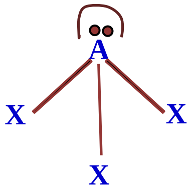
b) Total number of electron pairs around the central atom = 4.

Number of bonding pairs = 3.

Number of lone pairs = 1.

The molecule will be of the form AX_3 but with a lone pair of electron

Shape: Trigonal Pyramidal.



e.g. NH_3, PH_3

The bond angle XÂX, depends on how close to or far from the central atom, the shared pairs of electrons are. The closer they are to the central atom the stronger the repulsion between them and hence the larger the bonds angle, e.g., NH_3 ($\angle HNH = 107^0$), PH_3($\angle HPH = 93^0$ 20'), AsH_3 ($\angle HAsH = 91^0$ 50'), and SbH_3 ($\angle HSbH = 91^0$ 50')

The decrease in the bond angle in passing from NH_3 to SbH_3 is due to the decrease in electronegativity of the central atom. The electronegativity of the central atoms decreases in the order: N 〉 P 〉 As 〉 Sb.

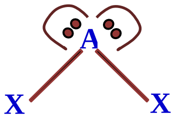
c) Total number of electron pairs around the central atom = 4.

Number of bonding pairs = 4.

Number of lone pairs = 2.

The molecule will be of the form AX_2 but with two loe pair

Shape:- angular



E.g. H_2O, H_2S, H_2Se, H_2Te

Case 4

Total number of electron pairs around central atom = 5

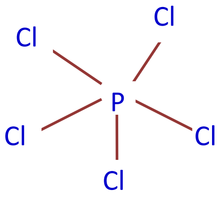
(a) Total number of electron pair = 5

Number of lone pairs = 0

Number of bond pairs = 5

Formula AX_5 e.g. PCl_5

Shape triangular bipyramidal



Case 5

Total number of electron pairs around central atom = 6

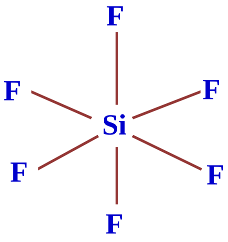
(a) Total number of electron pair = 6

Number of lone pairs = 0

Number of bond pairs = 6

Formula AX_6 e.g. SiF_6, $[Al(H_2O)_6]^{3+}$

Shape octahedral



NB: (I) Structures of molecules that do not involve lone pairs on the central atom are highly symmetrical, i.e., linear, triangular, and tetrahedral. This is because such structures involve only one kind of repulsion (bond pair - bond pair) and in such structure there is a constant bond angle, i.e., linear - 180^0, triangular = 120^0 and tetrahedral 109.5^0, triangular bipyramidal, octahedral, etc.

Trial 3.4.

1. Draw and name the structures of the following compounds

	Structure	Name
(i)	NCl_3	
(ii)	PCl_5	

2. (a) Sketch and name the shapes of the following molecules (3marks)

   (i) NH_3 (ii) BF_3 (iii) H_2S

   (b) Briefly explain why each molecule adapts the shape in (a) above (6marks)

   (c) Sketch the following molecules, SO_2Cl_2, PO_4^{3-}, ClO_4^-, MnO_4^-, ClF_3, ClO_3^-, ClO_2^-.

Thank you
Dr. Bbosa Science