



# Chemical Bonding



## Learning Objectives

Upon completion of this topic, learners will be able to:

- Distinguish the different types of chemical bonding
- Discuss hybridization of atomic orbitals
- Demonstrate the types of bonding of molecules by using the Lewis Structure
- Discuss electronegativity relative to the concept of bond polarity
- Discuss the molecular geometry
- Distinguish between inter-atomic bonding and intermolecular bonding and
- Explain coordinate covalent (dative) bond.

## Introduction

Matter is made up of one or different type of elements. Under normal conditions no other element exists as an independent atom in nature, except noble gases. However, a group of atoms is found to exist together as one species having characteristic properties. Such a group of atoms is called a **molecule**. Obviously there must be some force which holds these constituent atoms together in the molecules. *The attractive force which holds various constituents (atoms, ions, etc.) together in different chemical species is called a **chemical bond**.*

### 5.1. BONDING TYPES

The strength and properties of the chemical bonds formed vary. There are four primary types of chemical bonding. These are:

1. Ionic Bonding
2. Covalent Bonding
3. Metallic Bonding
4. Coordinate Covalent Bonding

### 5.1.1. Ionic Bonding

The compounds which are made up of ions are known as ionic compounds. In an ionic compound, the positively charged ions (cations) and negatively charged ions (anions) are held together by the strong electrostatic forces of attraction. The forces which hold the ions together in an ionic compound are known as **ionic bonds** or **electrovalent bonds**. Since an ionic bond consists of an equal number of positive ions and negative ions, the overall charge on an ionic compound is zero. For example, sodium chloride (NaCl) is an ionic compound which is made up of equal number of positively charged sodium ions ( $\text{Na}^+$ ) and negatively charged chloride ions ( $\text{Cl}^-$ ). Some of the common ionic compounds, their formulae and the ions present in them are given in Table 5.1.

**Table 5.1.** Formulae and Nomenclature of Some Ionic Compounds

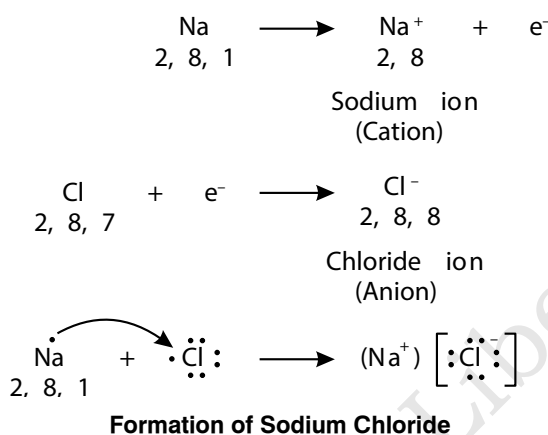
Nomenclature	Formula	Ions present
Aluminium oxide	$\text{Al}_2\text{O}_3$	$\text{Al}^{3+}$ and $\text{O}^{2-}$
Ammonium chloride	$\text{NH}_4\text{Cl}$	$\text{NH}_4^+$ and $\text{Cl}^-$
Calcium hydroxide	$\text{Ca}(\text{OH})_2$	$\text{Ca}^{2+}$ and $\text{OH}^-$
Calcium nitrate	$\text{Ca}(\text{NO}_3)_2$	$\text{Ca}^{2+}$ and $\text{NO}_3^-$
Calcium oxide	$\text{CaO}$	$\text{Ca}^{2+}$ and $\text{O}^{2-}$
Copper sulphate	$\text{CuSO}_4$	$\text{Cu}^{2+}$ and $\text{SO}_4^{2-}$
Magnesium chloride	$\text{MgCl}_2$	$\text{Mg}^{2+}$ and $\text{Cl}^-$
Potassium chloride	$\text{KCl}$	$\text{K}^+$ and $\text{Cl}^-$
Potassium hydroxide	$\text{KOH}$	$\text{K}^+$ and $\text{OH}^-$
Sodium carbonate	$\text{Na}_2\text{CO}_3$	$\text{Na}^+$ and $\text{CO}_3^{2-}$
Sodium hydroxide	$\text{NaOH}$	$\text{Na}^+$ and $\text{OH}^-$

Ionic compounds are made up of a metal and a non-metal (except ammonium chloride which is an ionic compound made up of only non-metals). So, whenever a bond involves a metal and a non-metal, we call it ionic bond.

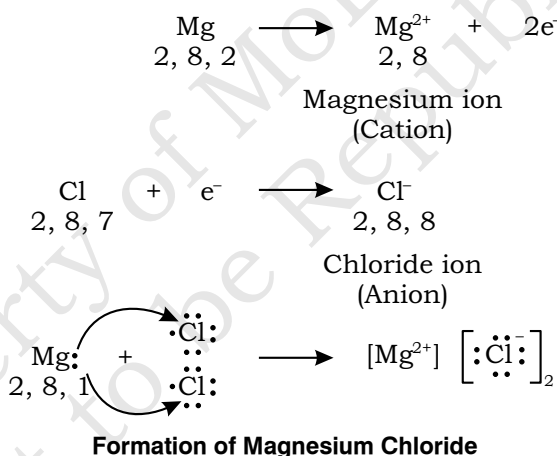
### 5.1.2. Formation of Ionic Bonding

An ionic bond changes the electronic configurations of the atoms. Metal atoms lose their outermost electron(s), forming cations. Non-metal atoms gain electron(s) to fill their outermost shell, forming anions. The electrostatic force of attraction between the oppositely charged ions holds the ions together. For example,

- (a) When a hot sodium atom is placed in chlorine gas, a reaction takes place resulting in formation of sodium chloride.



- (b) When a magnesium atom comes in contact with chlorine gas, it forms magnesium chloride.



### 5.1.2.1. Factors Influencing the Formation of Ionic Bonding

The following three factors influence the formation of ionic bonding:

1. **Ionization energy:** It is defined as the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom of an element. The lesser the ionization energy, the greater is the ease of the formation of a cation.
2. **Electron affinity:** It is defined as the amount of energy released when an electron is added to an isolated gaseous atom of an element. The higher the energy released during this process, the easier will be the formation of an anion.

3. **Lattice energy:** It is defined as the amount of energy released when cations and anions are brought from infinity to their respective equilibrium sites in the crystal lattice to form one mole of the ionic compound. The higher the lattice energy, the greater is the tendency of the formation of an ionic bond. The higher the charges on the ions and smaller the distance between them, the greater is the force of attraction between them.

### 5.1.3. Properties of Ionic Compounds



#### ACTIVITY 5.1

##### Illustrating Physical Properties of Ionic Compounds

1. Take a sample of sodium chloride or any other salt from the science laboratory.  
*What is the physical state of this salt?*
2. Take a small amount of a sample on a metal spatula and heat directly on the flame as shown in Fig. 5.1.

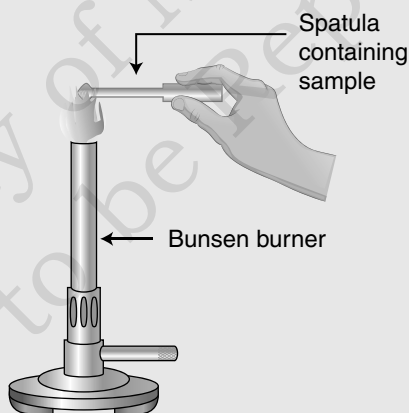
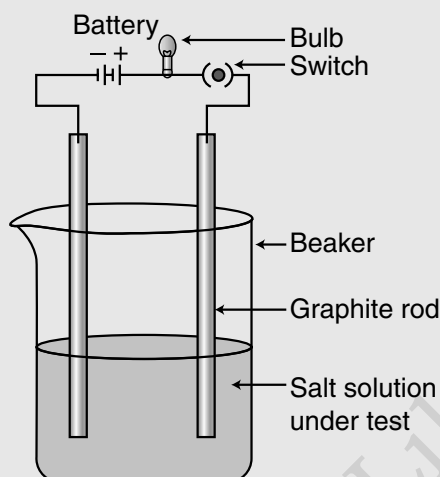


Fig. 5.1. Testing melting point of sodium chloride

*What did you observe? Did the sample impart any colour to the flame? Does this compound melt?*

3. Try to dissolve the sample in water, petrol and kerosene. Is it soluble?
4. Make a circuit as shown in Fig. 5.2 and insert the electrodes into a solution of salt. What did you observe?



**Fig. 5.2. Testing electrical conductivity of salt solution**

- What is your inference about the nature of this compound?

You may have observed the following general properties for ionic compounds:

- Ionic compounds are usually crystalline solids.
- Ionic compounds have high melting and high boiling points.
- Ionic compounds are usually soluble in water but insoluble in organic solvents like petrol and kerosene.
- Ionic compounds conduct electricity when dissolved in water or when melted. When we dissolve the ionic solid in water or melt it, the crystal structure is broken down to form ions. These ions help in conducting electricity.

## 5.2. HYBRIDIZATION OF ATOMIC ORBITALS

Hybridization is a theoretical concept which has been introduced to explain some structural properties such as shapes of molecules or equivalency of bonds, etc., which cannot be explained by simple theories of valency.

According to this concept, valence orbitals of the atom intermix to give rise to another set of equivalent orbitals before the formation of bonds. These orbitals are called **hybrid orbitals** or **hybridized orbitals** and the phenomenon is referred to as **hybridization**. Thus, **hybridization** may be defined as *the phenomenon of intermixing of atomic orbitals of slightly different energies of the atom (by redistributing their energies) to form new set of orbitals of equivalent energies and identical shape.*

There are many different types of hybridization depending upon the type of orbitals involved in mixing such as  $sp^3$ ,  $sp^2$ ,  $sp$ , etc. Let us discuss various types of hybridization alongwith some examples.

### $sp^3$ -Hybridization

The type of hybridization involves the mixing of one orbital of s-sub-level and three orbitals of p-sub-level of the valence shell to form four  $sp^3$  hybrid orbitals of equivalent energies and shape. Each  $sp^3$  hybrid orbital has 25% s-character and 75% p-character. These hybridized orbitals tend to lie as far apart in space as possible so that the repulsive interactions between them are minimum. The four  $sp^3$  hybrid orbitals are directed towards the four corners of a tetrahedron. The angle between the  $sp^3$  hybrid orbitals is  $109.5^\circ$  (Fig. 5.3).

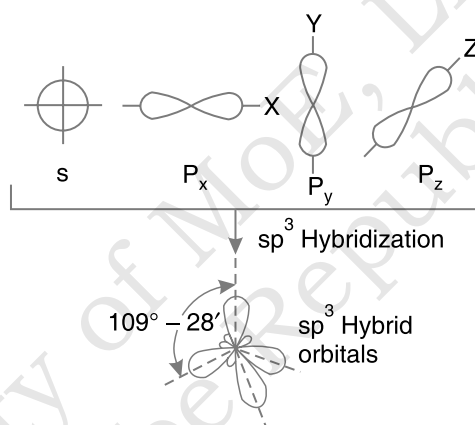


Fig. 5.3. Tetrahedral orientations of  $sp^3$  hybrid orbitals

$sp^3$  hybridization is also known as **tetrahedral hybridization**. The molecules in which central atom is  $sp^3$  hybridized and is linked to four other atoms directly, have **tetrahedral shape**.

Let us study an example of a molecule which involves  $sp^3$  hybridization.

1. **Formation of methane ( $CH_4$ ).** In methane molecule carbon atom acquires  $sp^3$  hybrid state.

Carbon atom has ground state configuration of  $2s^2 2p_x, 2p_y$  and in excited state its configuration is  $2s, 2p_x, 2p_y, sp_z$ . One orbital of 2s-sub-shell and three orbitals of 2p-sub-shell of excited carbon atom hybridise to form four  $sp^3$  hybrid orbitals.

The  $sp^3$  hybrid orbitals of carbon atom are directed towards the corners of regular tetrahedron. Each of the  $sp^3$  hybrid orbitals overlaps axially with half-filled 1s-orbital of hydrogen atom constituting a sigma bond (Fig. 5.4).

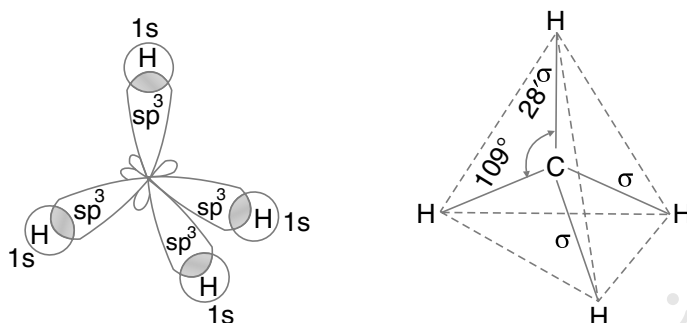


Fig. 5.4. Orbital picture of methane

Because of  $sp^3$  hybridization of carbon atom,  $\text{CH}_4$  molecule has **tetrahedral shape**.

### $sp^2$ -Hybridization

This type of hybridization involves the mixing of one orbital of  $s$ -sub-level and two orbitals of  $p$ -sub-level of the valence shell to form three  $sp^2$  hybrid orbitals. These  $sp^2$  hybrid orbitals lie in a plane and are directed towards the corners of equilateral triangle (Fig. 5.5).

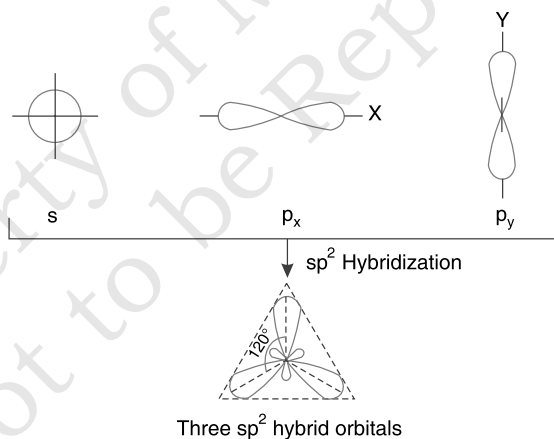


Fig. 5.5. Trigonal planar orientation of  $sp^2$  hybrid orbitals

Each  $sp^2$  hybrid orbital has one-third  $s$ -character and two-third  $p$ -character.  $sp^2$  hybridization is also called **trigonal hybridization**. The molecules in which central atom is  $sp^2$  hybridized and is linked to three other atoms directly have triangular planar shape.

Let us study an example of a molecule which involves  $sp^2$  hybridization.

1. **Formation of boron trichloride ( $\text{BCl}_3$ ).** Boron ( ${}_5\text{B}$ ) atom has ground state configuration as  $1s^2, 2s^2, 2p^1$ . But in the excited state its configuration is  $1s^2, 2s^1, 2p_x^1, 2p_y^1$ . One  $2s$ -orbital of

boron intermixes with two  $2p$ -orbitals of excited boron atom to form three  $sp^2$  hybrid orbitals.

The  $sp^2$  hybrid orbitals of boron are directed towards the corners of equilateral triangle and lie in a plane. Each of the  $sp^2$  hybrid orbitals of boron overlaps axially with  $3p$ -half-filled orbital of chlorine atom to form three B-Cl sigma bonds as shown in Fig. 5.6.

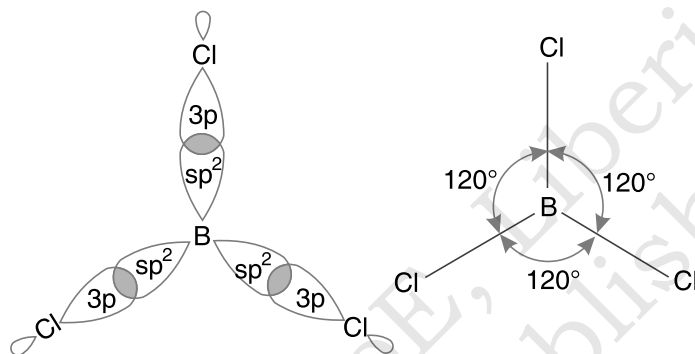


Fig. 5.6. Orbital diagram of  $\text{BCl}_3$

Because of  $sp^2$  hybridization of boron,  $\text{BCl}_3$  molecule has **triangular planar shape**.

### ***sp*-Hybridization**

This type of hybridization involves the mixing of one orbital of  $s$ -sub-level and one orbital of  $p$ -sub-level of the valence shell of the atom to form two  $sp$ -hybridized orbitals of equivalent shapes and energies. These  $sp$ -hybridized orbitals are oriented in space at an angle of  $180^\circ$  (Fig. 5.7). This hybridization is also called **diagonal hybridization**. Each  $sp$  hybrid orbital has equal  $s$  and  $p$  character, *i.e.*, 50%  $s$ -character and 50%  $p$ -character. The molecules in which the central atom is  $sp$ -hybridized and is linked to two other atoms directly have linear shape.

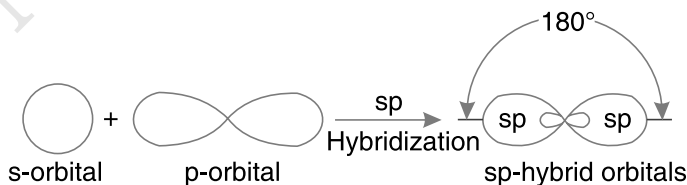


Fig. 5.7. Formation of  $sp$  hybrid orbitals

Let us study the **formation of beryllium fluoride ( $\text{BeCl}_2$ )**. Beryllium ( ${}_4\text{Be}$ ) atom has a ground state configuration as  $1s^2, 2s^2$ . In the excited state however, its configuration becomes  $v^2, 2^1, 2p^1x$ . One  $2s$ -orbital



and one  $2p$ -orbital of excited beryllium atom undergo  $2p$ -hybridization to form two  $sp$ -hybridized orbitals.

The two  $sp$ -hybrid orbitals are linear and oriented in opposite directions at an angle of  $180^\circ$ . Each of the  $sp$ -hybridized orbital overlap axially with  $3p$ -half-filled orbital of chlorine atom to form two Be—Cl sigma bonds (Fig. 5.8).

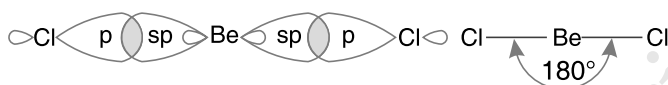


Fig. 5.8. Orbital diagram of BeCl<sub>2</sub>

Due to the  $sp$ -hybridised state of beryllium, **BeCl<sub>2</sub>** molecule has **linear shape**.

### 5.2.1. Covalent Bonding

The chemical bond formed by sharing of electrons between two atoms is known as a **covalent bond**. The compounds containing covalent bonds are known as covalent compounds. A covalent bond is formed when both the reacting atoms need electrons to achieve the inert gas electron arrangement. Now, the non-metals have usually 5, 6 or 7 electrons in the outermost shells of their atoms. So, all the non-metal atoms need electrons to achieve the inert gas structure. They get these electrons by mutual sharing. Thus, whenever a non-metal combines with another non-metal, covalent bond is formed.

**Table 5.2.** Formulae and Nomenclature of Some Covalent Compounds

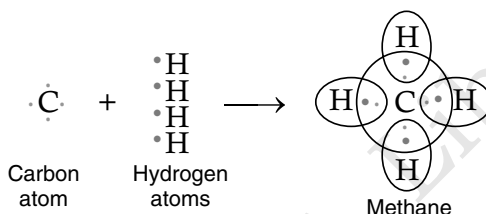
Nomenclature	Formula	Elements present
Methane	CH <sub>4</sub>	C and H
Ethane	C <sub>2</sub> H <sub>6</sub>	C and H
Ethene	C <sub>2</sub> H <sub>4</sub>	C and H
Ethyne	C <sub>2</sub> H <sub>2</sub>	C and H
Ammonia	NH <sub>3</sub>	N and H
Alcohol (Ethyl alcohol)	C <sub>2</sub> H <sub>5</sub> OH	C, H and O
Hydrogen sulphide gas	H <sub>2</sub> S	H and S
Carbon dioxide	CO <sub>2</sub>	C and O
Carbon disulphide	CS <sub>2</sub>	C and S
Carbon tetrachloride	CCl <sub>4</sub>	C and Cl

Glucose	$C_6H_{12}O_6$	C, H and O
Cane sugar	$C_{12}H_{22}O_{11}$	C, H and O
Urea	$CO(NH_2)_2$	C, O, N and H

### 5.2.1.1. Formation of Covalent Bonding

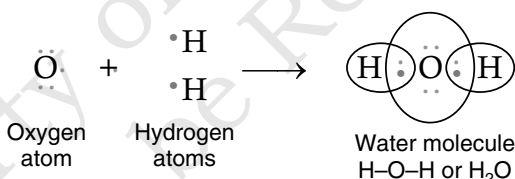
Covalent bonding between atoms of different elements.

- (i) Carbon atom shares four electrons to form methane.



Covalent bonding or sharing of electrons only takes place in outermost shells of atoms to attain inert gas electron arrangement.

- (ii) As in water molecule, 2 hydrogen atoms share electrons with oxygen atom.



### 5.2.1.2. Factors that Influence the Formation of Covalent Bonding

Since the covalent bond is formed by mutual sharing of electrons, therefore, the factors which favour the formation of covalent bonds are as follows:

1. **Electron Affinity:** A covalent bond is generally favoured between the two atoms if both the atoms have high electron affinity.
2. **Ionisation Energy:** The ionisation energy of both the atoms participating in bonding should be high.
3. **Atomic Size:** Atomic size of the atoms forming covalent bond should be smaller. Smaller the atomic radii of atoms, stronger the covalent bond will be. For example, H-H bond is stronger than Cl-Cl bond which in turn is stronger than Br-Br bond.

4. **Electronegativity:** The electronegativities of both the atoms should be high. The difference of electronegativities between the two atoms should be minimum.

### 5.2.1.3. Properties of Covalent Bonding



## ACTIVITY 5.2

### Illustrating Physical Properties of Covalent Compounds

Let us test some covalent compounds in different ways:

1. Take sample of cooking oil. Try to dissolve it in water and ethanol. Does it dissolve?
2. Have you ever observed a burning candle wax? If not, take a candle wax and observe it burning. How much time does it take to melt down?
3. Take a pan and add water to it. Let it boil. Do you know the boiling point of water?
4. Now add two electrodes to the water pan making a circuit. What did you observe? What would have happened if you would have added NaCl salt in the pan?

*What can you now say about these covalent compounds?*

You have observed the following properties of covalent compounds:

- Covalent compounds are usually liquids, gases or solids. For example, alcohol, benzene, water and cooking oil are liquids. Methane, ethane and chlorine are gases. Glucose, urea, and wax are solid covalent compounds.
- Covalent compounds have usually low melting points and low boiling points.
- Covalent compounds are usually insoluble in water, but they are soluble in organic solvents. Some of the covalent compounds like glucose, sugar and urea are soluble in water.
- Covalent compounds do not conduct electricity because they do not contain ions.



## ACTIVITY 5.3

### Detecting an Ionic Bond or Covalent Bond

1. Take the sample such as common salt (NaCl) provided.
2. Try to dissolve it in water.
3. If it dissolves, chances are it is likely to be an ionic compound. But, you already know some covalent compounds like glucose, urea and sugar are soluble in water.
4. Now, perform electrical conductivity test (Fig. 5.9).

- If the NaCl sample dissolves in water, arrange a circuit with two electrodes and a bulb.
- Figure out whether the bulb glows or not. According to your observation conclude the bond present in the sample.
- Make a report on the properties of ionic and covalent compounds.

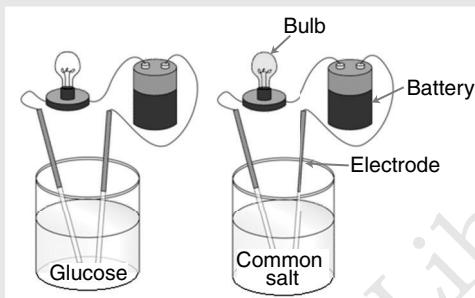


Fig. 5.9. Electrical conductivity test



## ACTIVITY 5.4

### Identifying Ionic and Covalent Compounds

Choose the ionic as well as covalent compounds from the bubbles and make a table in your exercise notebook.

Nitric acid		
Methane	Urea	Ionic compound
H <sub>2</sub> O	Acetic acid	
Ethyl alcohol	CuSO <sub>4</sub>	Covalent compound
Glucose	Formaldehyde	
H <sub>2</sub> SO <sub>4</sub>	CaO	
	NaOH	

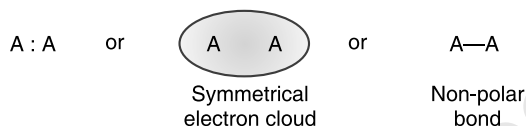
### 5.2.2. Bond Polarity

Covalent bonds between different atoms may acquire some ionic characters which can be explained on the basis of concept of electronegativity.

**Electronegativity** of an element is defined as *the power of its atoms to attract the bonding or shared pair of electrons towards itself.*

### 5.2.2.1. Polar Character of Covalent Bond

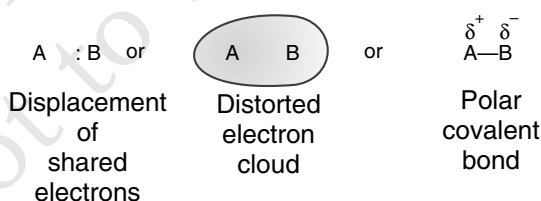
When covalent bond is formed between two similar atoms, the shared pair of electrons lies midway between the nuclei of the two atoms, because both the atoms have same attraction for the bonding electrons. The electron cloud constituting the covalent bond is symmetrically distributed around the atoms. Such a covalent bond is called a **non-polar covalent bond**.



For example, molecules like  $H_2$ ,  $Cl_2$ ,  $O_2$  and  $N_2$  contain non-polar bonds.

In case, the covalent bond is formed between two **dissimilar atoms**, one of which has a larger value of electronegativity, the bonding pair of electrons is displaced towards the more electronegative atom. In other words, electron cloud containing the bonding electrons gets distorted and the charge density concentrates around the more electronegative atom.

Due to the unequal distribution of electron charge density, the more electronegative atom acquires a partial negative charge (indicated as  $\delta^-$ ) whereas the less electronegative atom acquires a partial positive charge indicated as  $\delta^+$ . Thus, *a covalent bond develops a partial ionic character as a result of the difference of electronegativities of the atoms comprising the bond. Such a bond is called **polar covalent bond** (as shown in Fig. 5.10).*



**Fig. 5.10. Polarity in covalent bond**

For example, the bond between hydrogen and chlorine atoms in HCl molecule is polar because the shared electron pair is displaced towards chlorine atom which is more electronegative.



The extent of ionic character in a covalent bond depends upon the difference of electronegativities of the two atoms forming a bond. The difference of electronegativities, greater is the percentage of ionic

character in a bond. For example, hydrogen fluoride is more polar than hydrogen chloride because the difference of between H and F is more than that between H and Cl.

	$\delta^+$ $\delta^-$	$\delta^+$ $\delta^-$
	<b>H—F</b>	<b>H—Cl</b>
<i>Electronegativity</i>	2.1   4.0	2.1   3.0
<i>Difference</i>	<b>1.9</b>	<b>1.9</b>
	More ionic than H—Cl	More covalent than HF

It has been observed that the bond has 50% ionic character and 50% covalent character if the difference of electronegativities of the participating atoms is 1.7. On the other hand, the covalent character dominates if the difference of electronegativities is less than 1.7 while ionic character dominates if the difference of electronegativities is greater than 1.7.

#### 5.2.2.2. Dipole Moment ( $\mu$ )

Degree of polarity in a bond is measured in terms of **dipole moment** which is defined as *the product of the magnitude of charge ( $e$ ) on any one of the atoms and the distance ( $d$ ) between them*. Dipole moment is represented by a Greek letter ' $\mu$ '. It can be expressed mathematically, as

$$\mu = e \times d$$

Dipole moment is expressed in a unit called *Debye* and is represented by D.

$$1 \text{ D} = 1 \times 10^{-18} \text{ e.s.u.-cm.} = 3.335 \times 10^{-30} \text{ C m}$$

For example, dipole moment of HCl is  $1.03 \times 10^{-18} \text{ e.s.u.-cm}$  and is expressed as 1.03 D. Dipole moment can be determined experimentally and its-value can give an idea of the polar character of a molecule. It is a *vector quantity* as it has a direction as well as magnitude. The direction of dipole moment is usually represented by an arrow  $\rightarrow$  pointing from positive end towards the negative end.

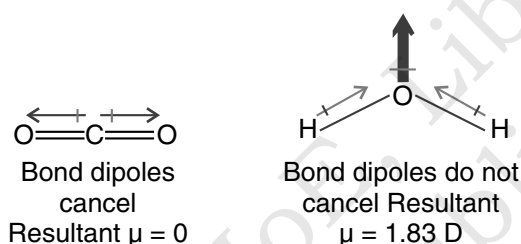
#### 5.2.2.3. Polar and Non-Polar Molecules

Polar character of bonds also affects the polarity in a molecule. In **diatomic** molecules, the two atoms are bonded by a covalent bond.

In such molecules, the dipole movement of bond also represents the dipole moment of the molecule, some examples are:

H—H	Cl—Cl	H→Cl	H→F
$\mu = 0$	$\mu = 0$	$\mu = 1.03\text{D}$	$\mu = 1.78\text{D}$

In polyatomic molecules having more than one bonds, the idea of a dipole can be applied to individual bonds within the molecule. The dipole moment of individual bond in a polyatomic molecule is referred to as **bond dipole**. The dipole moment of the molecule depends upon the orientations of various bond dipoles. For example, carbon dioxide ( $\text{CO}_2$ ) and water ( $\text{H}_2\text{O}$ ) are both triatomic molecules but dipole moment of carbon dioxide is zero whereas that of water is 1.83 D. This can be explained on the basis of their structures.



Carbon dioxide is a **linear molecule** in which the two C = O bonds are oriented in the opposite directions at an angle of  $180^\circ$ . The bond dipole of two C = O bonds cancel each other. Therefore, the resultant dipole moment of the molecule is zero. Hence,  $\text{CO}_2$  is a non-polar molecule. On the other hand, water molecule has a bent structure in which two O—H bonds are oriented at an angle  $104.5^\circ$ . Therefore, the bond dipoles of two O—H bonds do not cancel each other. Hence, the resulting dipole moment of  $\text{H}_2\text{O}$  molecule is not zero but has a value equal to 1.83 D.

It may be noted that the molecules having **zero dipole moment** ( $\mu = 0$ ) are called **nonpolar molecules**. On the other hand the molecules having dipole moment greater than zero ( $\mu > 0$ ) are **polar molecules**.

### 5.3. MOLECULAR GEOMETRY

Atoms bind together to form molecules, and atoms of each element are different. They are different sizes, they have different electron configurations, and their chemical behaviour means that they act differently inside a molecule. In this way, the variety of atoms and the behaviour of their electrons means that molecules form in certain ways, and their shapes can be predicted using molecular geometry. Thinking of atoms as balls and atomic bonds as sticks, one can construct “toothpick models” that are approximations of molecular shapes.

Molecular geometry is explained primarily through Valence Shell Electron Pair Repulsion (VSEPR) theory. While electrons can coexist within an orbital in pairs, they are still all negatively charged, so those pairs will repel each other. Most molecular geometry can be explained by counting the valence electrons in a molecule and accounting for how they want to be as far away from each other as possible. As molecules get larger, they develop stable configurations wherein electron pairs find a “comfortable” position apart from each other, and so familiar geometric shapes are formed.

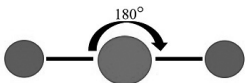
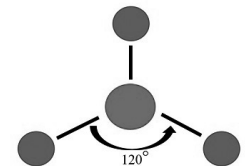
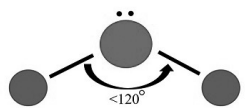
### Valence Shell Electron Pair Repulsion (VSEPR) theory

The valence shell electron pair repulsion (VSEPR) theory is a model used to predict 3-D molecular geometry based on the number of valence shell electron bond pairs among the atoms in a molecule or ion. This model assumes that electron pairs will arrange themselves to minimize repulsion effects from one another. In other words, the electron pairs are as far apart as possible.

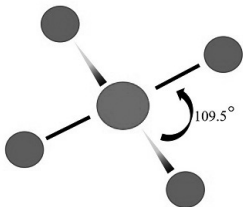
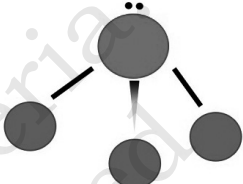
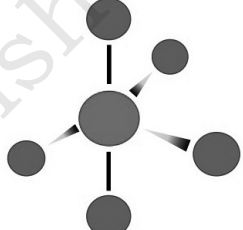
The VSEPR model is useful for predicting and visualizing molecular structures. The structures are: linear, trigonal planar, angled, tetrahedral, trigonal pyramidal, trigonal bipyramidal, disphenoidal (seesaw), t-shaped, octahedral, square pyramidal, square planar, and pentagonal bipyramidal.

#### 5.3.1. Shapes of Some Simple Molecules

Simple molecules have simple shapes, but things get increasingly more complicated when more atoms are involved. Some of the major molecular shapes are listed below.

S.No.	Shape	Structure
1.	<b>Linear:</b> In a linear molecule, the electron pairs take up opposite sides of the central atom.	
2.	<b>Trigonal Planar:</b> In a trigonal molecule, three pairs are evenly spaced around the central atom.	
3.	<b>Angular:</b> In an angular (also called bent) molecule, a lone pair of electrons takes what would be a flat molecule and pushes the arms further away from it and toward each other in the “down” direction.	



4.	<b>Tetrahedral:</b> In a tetrahedral molecule, the four electron pairs get further away from each other by leaving the flat plane and separating in 3 dimensions.	
5.	<b>Trigonal Pyramidal:</b> Much like an angular molecule, the multiple arms in a pyramidal molecule have been pushed out of the flat plane by the repulsion from a nonbonding pair.	
6.	<b>Trigonal Bipyramidal:</b> In a trigonal bipyramidal molecule, three of the pairs are on a flat plane, as in a trigonal planar molecule. Two more arms are aimed away from the flat plane, one “up” and one “down.”	

### 5.3.1.1. How to Calculate Molecular Geometry

In short, count the valence electrons to determine the shape of the molecule. Follow these steps to calculate molecular geometry:

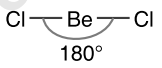

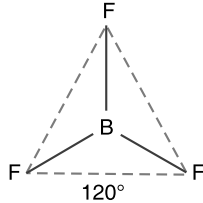

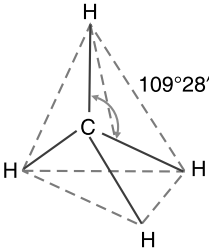

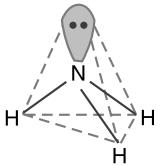
1. Determine the Lewis structure of the atom.
2. Count the valence electrons of the central atom in the molecule. The number of valence electrons, for the most part, can be determined by the element's group on the periodic table.
3. Determine the number of bonding electrons, and the number of electrons in lone pairs.
4. The number of bonds coming from the central atom and the number of lone pairs will determine the shape of the molecule. The bond angle is the degree measurement of the angle between any two bonds.

No. of Bonds on Central Atom	No. of Lone Pairs	Molecular Shape	Bond Angle	Examples
2	0	Linear	180	BeF <sub>2</sub> , BeCl <sub>2</sub> , CO <sub>2</sub>
3	0	Trigonal planar	120	BF <sub>3</sub> , AlCl <sub>3</sub>

2	1	Angular (bent)	<120	O <sub>3</sub> , SnCl <sub>2</sub>
4	0	Tetrahedral	109.5	CH <sub>4</sub> , SiF <sub>4</sub>
3	1	Trigonal pyramidal	<109.5	PCl <sub>3</sub> , NCl <sub>3</sub> , NH <sub>3</sub>
5	0	Trigonal bipyramidal	90, 120, 180	PF <sub>5</sub> , PCl <sub>5</sub> , SbCl <sub>5</sub>

**Example 5.1:** Describe the shapes of the following molecules on the basis of VSEPR Model: BeCl<sub>2</sub>, BF<sub>3</sub>, CH<sub>4</sub>, NH<sub>3</sub>, H<sub>3</sub>O<sup>+</sup>, CO<sub>2</sub>, SF<sub>6</sub>

**Solution:** The Lewis dot structures and shapes of various species are given below in tabular form:

Species	Lewis dot structure	Electron pairs around central atom		Arrangement of electron pairs	Shape
		bp	lp		
BeCl <sub>2</sub>	Cl •• Be •• Cl	2	0		Linear
BF <sub>3</sub>		3	0		Trigonal planar
CH <sub>4</sub>		4	0		Tetrahedral
NH <sub>3</sub>		3	1		Pyramidal
CO <sub>2</sub>	O •• C •• O	2	0	O = C = O	Linear

## 5.4. METALLIC BONDING

The force which binds various metal atoms together is called **metallic bond**. The metallic bond is neither a covalent bond nor an ionic bond because these bonds are not able to explain properties of metals.

For example, metals are very good conductors of electricity but in solid state. Both ionic and covalent compounds cannot do so with the exception of graphite.

### 5.4.1. Formation of Metallic Bonding

Lorentz proposed the theory of electron gas model or **electron sea model** for metallic bonding.

In this model, the metal is pictured as an array of metal cations in a “sea” of electrons. The atoms in a metallic solid contribute their valence electrons to form a “sea” of electrons that surrounds metallic cations. Delocalised electrons are not held by any specific atom and can move easily throughout the solid. A metallic bond is the attraction between these electrons and the metallic cation.

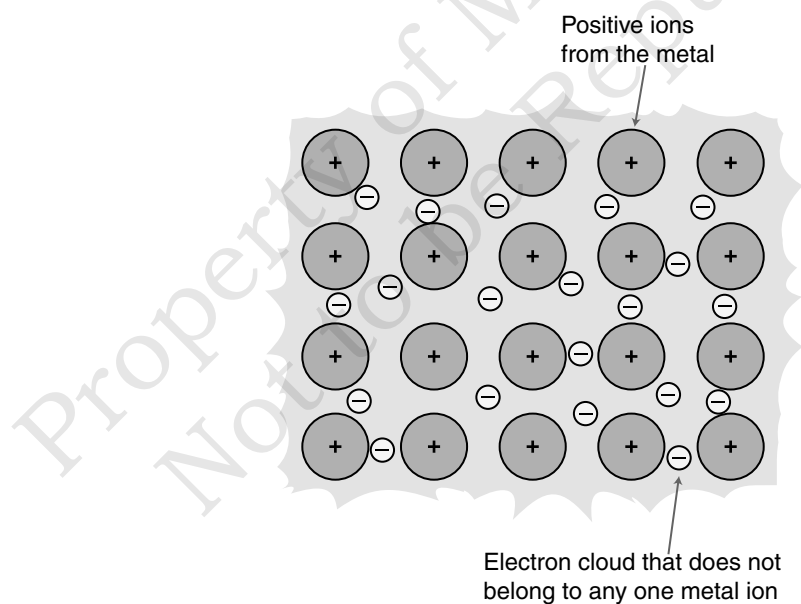


Fig. 5.11. Formation of metallic bond

#### 5.4.1.1. Factors that influence the formation of metallic bonding

The factors influencing the formation of metallic bonds are:

- The *ionisation energy* of the metal should be low so that the valence electrons are loosely held by the nucleus and thus become mobile.

- The metals should have high *electropositivity*, so they do not accept electrons and form anions.
- The *vacant orbital* should be more than the number of valence electrons so that the mobile electrons can easily move in the vacant orbitals.

### 5.4.2. Properties of Metals



#### ACTIVITY 5.5

##### Illustrating Conductivity of Heat and Electricity of Metals

1. Take an aluminium or copper wire. Clamp this wire on a stand, as shown in Fig. 5.12.
2. Fix a pin to the free end of the wire using wax.
3. Heat the wire with a spirit lamp, candle or a burner near the place where it is clamped.

*What do you observe after some time?*

4. Repeat the same with carbon or sulphur.
5. Note your observations.

*Does the element melt?*

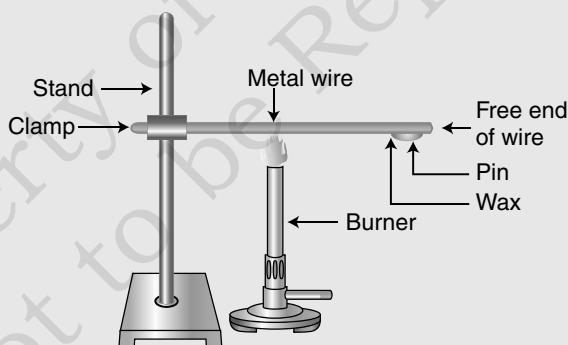
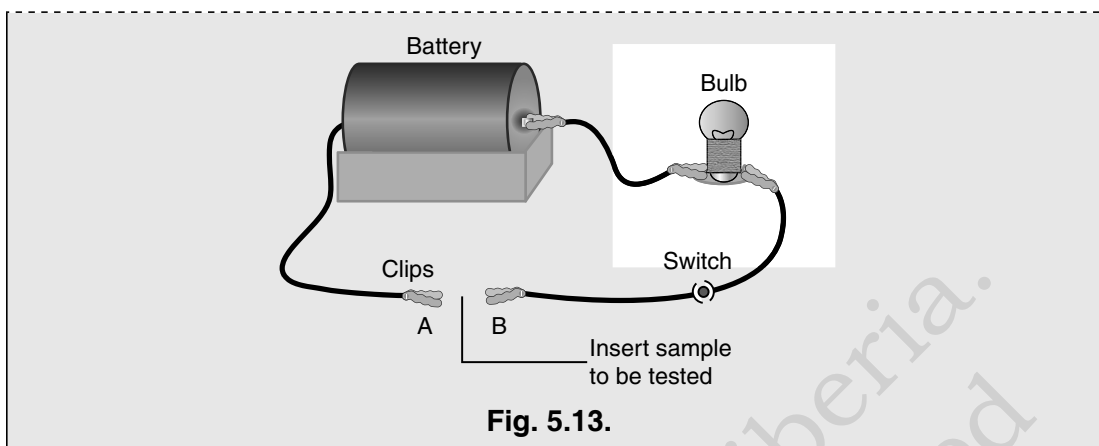


Fig. 5.12.

6. Consider elements aluminium, copper, sulphur and carbon.
7. Set up an electric circuit as shown in Fig. 5.13.
8. Place the element to be tested in the circuit between terminals A and B as shown.

*Does the bulb glow? What does this indicate?*

9. Compile your observations regarding properties of elements in your exercise book.



- Metals are good conductors of heat and electricity. This means that metals allow heat and electricity to pass through them easily. Silver metal is the best conductor of heat. Copper metal is a better conductor of heat than aluminium metal.
- Metals are lustrous (or shiny). This means that metals have a shiny appearance.
- Metals are usually **strong**. For example, iron metal (in the form of steel) is very strong when freshly cut and is used in the construction of bridges, buildings and vehicles. Some metals are not strong. For example, sodium and potassium.
- Metals are **ductile**. This means that metals can be drawn (or stretched) into thin wires.
- Gold and silver are among the best ductile metals.
- Metals are **malleable**. This means that metals can be hammered into thin sheets.

## 5.5. INTERMOLECULAR BONDING

We know that in covalent molecules, the constituent atoms are held by the covalent bonds. These molecules are further held by some attractive forces in different chemical entities. For example,  $\text{H}_2\text{O}$  molecule is a discrete unit which contains two O–H covalent bonds. But a sample of water contains large number  $\text{H}_2\text{O}$  units held by some attractive forces. *The attractive forces operating between the molecules of a given substances (solid, liquid or gas) are called **intermolecular forces**.*

The different types of intermolecular forces among the covalent molecular are

1. van der Waal forces or London dispersion forces.
2. Dipole-Dipole interactions
3. Hydrogen Bonds.

The magnitude of intermolecular forces control most of the physical properties of the substance such as *melting point*, *boiling point*, *viscosity*, *surface tension*, *molar heat of vaporisation*, *extent of evaporation*, etc. Higher the magnitude of the interparticle forces, higher will be boiling point. In many cases they also affect the chemical properties of the substances.

**Intermolecular forces** on the other hand, mostly control the chemical properties of the substance.

Let us now study the origin and nature of various intermolecular forces.

### 5.5.1. London Dispersion Forces or van der Waal Forces

These are the intermolecular forces among the **monoatomic** or **nonpolar molecules** such as  $N_2$ , He,  $H_2$ ,  $CO_2$ , etc. Non-polar molecules are electrically symmetrical because their electronic charge cloud is symmetrically distributed. However, an **instantaneous dipole** may develop in such molecules. Let us understand this by assuming two atoms of neon 'A' and 'B' in the close vicinity of each other (Fig. 5.14 (a)).

Rapid movement of electrons in A may cause its temporary polarization due to momentary accumulation of electron density on one side and making the charge distribution unsymmetrical (Figs. 5.14 (b) and (c)).

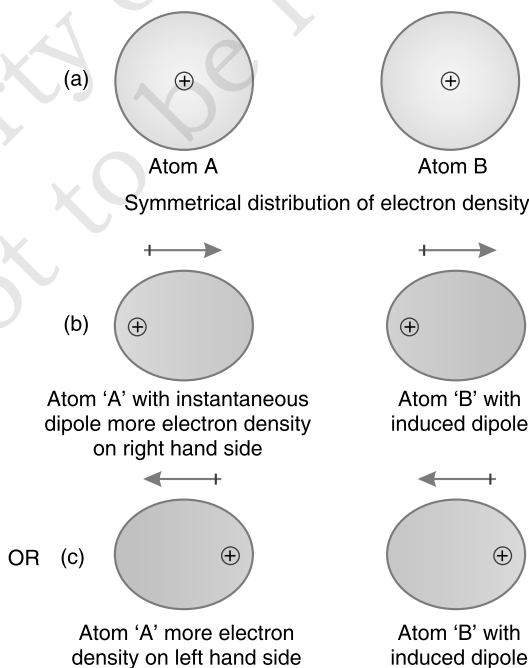


Fig. 5.14. Dispersion forces or London forces between atoms

This will result in the development of temporary instantaneous dipole on the atom 'A' for a very short time. This instantaneous or transient dipole distorts the electron density of the atom 'B', which is close to it. In other words, a dipole is induced in the atom 'B' also.

The **temporary dipoles** of atom 'A' and 'B' attract each other. Similar temporary dipoles are induced in non-polar molecules also. Magnitude of such a force of attraction was first calculated by the German physicist **Fritz London**. For this reason, force of attraction between two temporary dipoles is known as **London force**. Another name for this force is **dispersion force**. These forces are always attractive and interaction

energy is proportional to  $\left(\frac{1}{r^6}\right)$  where  $r$  is the distance between two interacting particles. These forces are important only at short distances ( $\sim 500$  pm) and their magnitude depends on the ability of the particles to undergo **polarization**.

### 5.5.2. Dipole-Dipole Forces

**Dipole-dipole forces** are attractive forces operating between the polar molecules possessing permanent dipole ( $\mu > 0$ ). Ends of the dipole possess "partial charges" and these charges are shown by Greek letter delta ( $\delta$ ). Partial charges are always less than the unit charge ( $1.6 \times 10^{-19}$  C) because of electron sharing effect. The polar molecules interact with neighbouring molecules through their oppositely charged ends. This interaction is weak as compared to ion-ion interactions. The interaction energy decreases with the increase of distance between the dipoles. Now, greater the dipole moment of polar molecule, greater is the magnitude of dipole-dipole forces. Figure 5.15 shows electron cloud distribution and attractive interactions between H—Cl dipoles.

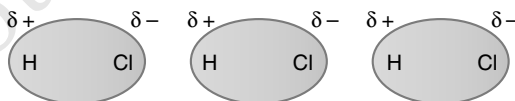
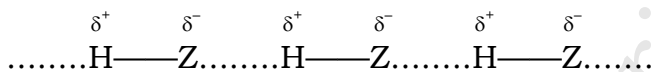


Fig. 5.15. Attractive interactions between H—Cl dipoles

In general, for the molecules of similar molecular masses, the magnitude of dipole-dipole forces is higher in case of more polar molecules. For example, molecular masses of  $\text{PH}_3$  and  $\text{H}_2\text{S}$  is same but  $\text{H}_2\text{S}$  is more polar ( $\mu = 1.10\text{D}$ ) than  $\text{PH}_3$  ( $\mu = 0.55\text{D}$ ). Hence boiling point of  $\text{H}_2\text{S}$  (187K) is higher than that of  $\text{PH}_3$  (137 K).

### 5.5.3. Hydrogen Bond

When a H-atom is bonded to a highly electronegative atom like F, O or N (say, Z) by a covalent bond, the bond pair of electrons is displaced towards the electronegative atom. When solitary electron of hydrogen atom lies away from it, its nucleus gets exposed and behaves as a bare proton. Such a bare hydrogen nucleus exerts a strong electrostatic attraction on the electronegative atom of the adjacent molecule. This



interaction between hydrogen atom of one molecule and the electronegative atom of the other molecule is referred to as **hydrogen bond**. Thus, **hydrogen bond** is defined as the electrostatic force of attraction which exists between the covalently bonded hydrogen atom of one molecule and the electronegative atom of the other molecule. The hydrogen bond is represented by dotted line (.....). For example, in case of hydrogen fluoride the hydrogen bond exists between H atom of one molecule and fluorine atom of another molecule as shown.



It may be noted that the hydrogen atom acts as a bridge between the electronegative atoms of two different molecules, to one atom it is linked through a covalent bond while to the other it is linked through a hydrogen bond. However, the hydrogen atom does not lie in the centres of the two electronegative atoms because hydrogen bond and covalent bond do not have same strength.

#### 5.5.3.1. Examples of Hydrogen Bonding

Let us study some examples of the compounds which exhibit hydrogen bonding.

- (i) **Hydrogen fluoride.** In hydrogen fluoride, hydrogen atom is bonded to highly electronegative atom fluorine (electronegativity = 4). It has been found that in solid state hydrogen fluoride consists of long zig-zag chains of H-F molecules associated by H-bonds as shown in Fig. 5.16 (a). On heating, progressively, the length of the chain shortens, and associated units become quite small.

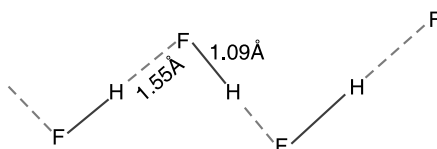


Fig. 5.16. (a) Zig-zag chains of H-F molecules



- (ii) **Water.** In water molecule, oxygen atom is bonded to two hydrogen atoms. Due to large electronegativity, oxygen atom forms the negative centre whereas each of the hydrogen atom acquires a partial positive charge. Each O atom can form two hydrogen bonds as shown in Fig. 5.16 (b).

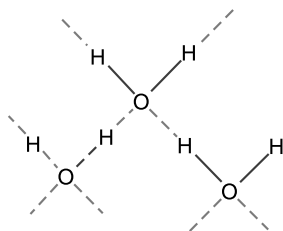


Fig. 5.16. (b) H-bonds in  $\text{H}_2\text{O}$  molecules

- (iii) **Ammonia.** In ammonia molecule, nitrogen, an electronegative atom is bonded to three hydrogen atoms. The nitrogen atom forms a negative site of the molecule whereas each of three H-atoms acquires a partial positive charge. The ammonia molecules are associated by H-bonds as shown in Fig. 5.16 (c).

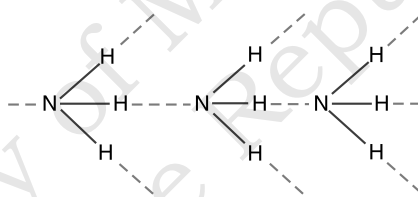
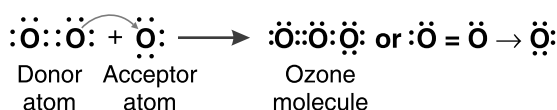


Fig. 5.16. (c) H-bonds in  $\text{NH}_3$  molecules

## 5.6. COORDINATE COVALENT (DATIVE) BONDING

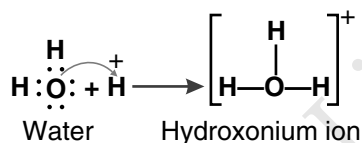
It is a special case of covalent bond the formation of which was postulated by **Perkins** (1921). It is formed by *mutual sharing of electrons between the two atoms but the shared pair of electrons is contributed only by one of the two atoms, the other atom simply participates in sharing*. The atom which donates an electron pair for sharing is called **donor** and it must have already completed its octet. On the other hand, the atom which accepts the electron pair in order to complete its octet is called **acceptor**. The bond is represented by an arrow ( $\rightarrow$ ) pointing from the donor towards the acceptor. Let us consider the formation of ozone molecule. A molecule of oxygen contains two oxygen atoms which share four electrons and complete their octets. In the formation of  $\text{O}_3$  molecule, one of the oxygen atom of  $\text{O}_2$  molecules acts as donor atom and shares two of its electrons with another oxygen atom. It can be represented as follows:



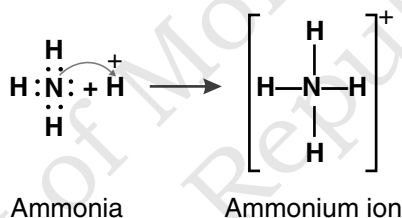
**It is important to note that co-ordinate bond once formed, cannot be distinguished from covalent bond.**

Some more examples of molecules/polyatomic ions having coordinate bond are as follows:

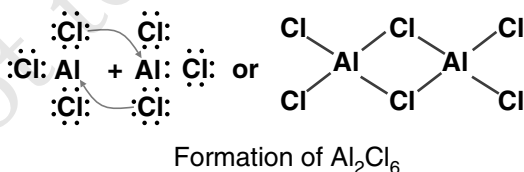
- (i) **Hydroxonium ion.** Here, oxygen atom of  $\text{H}_2\text{O}$  molecule is donor while  $\text{H}^+$  is acceptor.



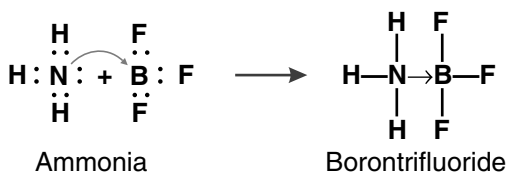
- (ii) **Ammonium ion.** Here, N atom of ammonia molecule is donor while  $\text{H}^+$  is acceptor.



- (iii) **Aluminium chloride.** Aluminium chloride exists in the form of dimer  $\text{Al}_2\text{Cl}_6$ . Here, chlorine atom of  $\text{AlCl}_3$  unit acts as donor while electron deficient Al atom of the other  $\text{AlCl}_3$  unit acts as acceptor as shown below.

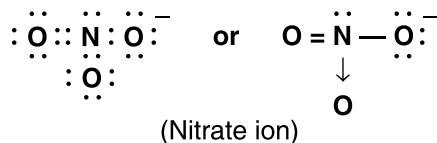


- (iv) **Combination of borontrifluoride and ammonia.** Boron atom of boron trifluoride being electron deficient acts as acceptor while nitrogen atom of ammonia molecule is donor. Their combination by donor acceptor mechanism is shown on follows.

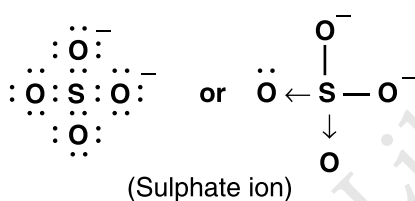


(v) **Lewis structures of nitrate ( $\text{NO}_3^-$ ) and sulphate ( $\text{SO}_4^{2-}$ ) ions**

**Nitrate ion ( $\text{NO}_3^-$ )**



**Sulphate ion ( $\text{SO}_4^{2-}$ )**



### ACTIVITY 5.6

In small groups, discuss the formation of covalent bond, dative bond and write the Lewis structures of simple covalent molecules/poly atomic ions. Watch video clips about formation of covalent and coordinate bond.

Make presentations by writing Lewis structures of the following molecules/ions.

- Water
- Sulphate ion
- Hydrogen peroxide
- Hydrogen chloride
- Carbon dioxide
- Carbon monoxide
- Methane
- Boron trifluoride
- Ammonia
- Nitrate ion
- Sulphur hexafluoride
- Ammonium ion

#### 5.6.1. Comparison of All Bond Types

A comparison chart of bond types is summarized in Table 5.3.

**Table 5.3.** Comparison of All Bond Types

	<b>Ionic</b>	<b>Covalent</b>	<b>Metallic</b>
<b>State at room temperature</b>	Crystalline	Liquid, gas or low-melting solid	Solid state (Exception in mercury)
<b>Occurs between</b>	A metal and a non-metal	Two non-metals	Two and more metals
<b>Character of bond</b>	Non-directional	Directional	Non-directional

<b>Conductivity</b>	Conducts electricity when molten and dissolves in water	Do not conduct electricity	Good conductor of electricity
<b>Solubility</b>	Soluble in water but insoluble in organic solvents like petrol and kerosene.	Insoluble in water, but soluble in organic solvents like petrol and kerosene.	Insoluble in water or organic solvents, unless they undergo a chemical reaction with them.
<b>Melting Point</b>	High	Low to moderate	Low
<b>Boiling Point</b>	High	Low to moderate	Low
<b>Ductility</b>	Non-ductile	Non-ductile	Ductile
<b>Malleability</b>	Non-malleable	Non-malleable	Malleable
<b>Examples</b>	NaCl, $\text{MgCl}_2$ , CaO, etc.	$\text{CH}_4$ , $\text{NH}_3$ , $\text{CO}_2$ , etc.	Na, Mg, Al, etc.

## GLOSSARY

- **Anion:** a negatively charged ion.
- **Boiling Point:** The temperature at which vapour pressure of a liquid becomes equal to the atmospheric pressure.
- **Cation:** a positively charged ion.
- **Covalent bonding:** A chemical bond formed when electrons are shared between two atoms.
- **Dative Bond** or **Co-ordinate Bond:** The bond is formed by sharing of electrons in which the shared pair of electrons is contributed by one of the atom called DONOR while the other atom is called ACCEPTOR.
- **Dipole Moment ( $\mu$ ):** A vector quantity defined by the product of charge developed on any of the atom and distance between the atoms; creating a dipole.
- **Ductile:** Able to be drawn out into a thin wire.
- **Electronegativity:** It is the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.
- **Hybridization:** The concept of mixing two atomic orbitals to give rise to a new type of hybridized orbitals.
- **Hydrogen Bond:** The electrostatic force of attraction between covalently bonded H-atom of one molecule and the electronegative atom (F or N or O) of the other molecule.

- **Intermolecular Forces:** The attractive forces which hold molecules of the substance together in different states of matter.
- **Ionic Bonding:** A chemical bond formed when one atom gives up one or more electrons to another atom.
- **Ionization Energy:** The energy required to remove the outermost electron from an isolated gaseous atom of the element.
- **Malleable:** Able to be hammered or pressed into shape without breaking or cracking.
- **Melting Point:** The temperature at which a given solid material changes from a solid state to a liquid, or melts.
- **Metallic Bonding:** A type of chemical bonding between two or more metal atoms.
- **Noble Gas:** The gaseous elements helium, neon, argon, krypton, xenon, and radon.
- **Van der Waals' Forces:** Dipole-dipole forces, dipole-induced dipole forces and dispersion forces are collectively called van der Waal forces.

## SUMMARY

- The attractive force which holds various constituents (atoms, ions, etc.) together in different chemical species is called a chemical bond.
- The compounds which are made up of ions are known as ionic compounds.
- The factors influencing the formation of ionic bonding are (i) ionization energy, (ii) electron affinity, and (iii) lattice energy.
- Ionic compounds are usually crystalline solids and have high melting and high boiling points.
- Hybridization is the phenomenon of intermixing of atomic orbitals of slightly different energies of the atom (by redistributing their energies) to form new set of orbitals of equivalent energies and identical shape.
- The chemical bond formed by sharing of electrons between two atoms is known as a covalent bond.
- The factors that influence the formation of covalent bonding are (i) electron affinity, (ii) ionization energy, (iii) atomic size, and (iii) electron egativity.
- Covalent compounds are usually liquids, gases or solids, and have low melting points and low boiling points.
- The force which binds various metal atoms together is called metallic bond.

- Degree of polarity in a bond is measured in terms of dipole moment.
- The VSEPR theory is a model used to predict 3-D molecular geometry based on the number of valence shell electron bond pairs among the atoms in a molecule or ion.
- Metals are good conductors of heat and electricity. They are strong, ductile and malleable.
- The attractive forces which hold together various molecules of the species in solid or liquid states are called intermolecular forces. The magnitude of these forces control the melting, boiling points of the substance.
- Different intermolecular forces are, van der Waal forces, hydrogen bonds, dipole-dipole forces and London dispersion forces.



## EVALUATION

### I. Multiple Choice Questions

1. Which of these is an ionic compound?  
(a) NaCl (b) H<sub>2</sub>O  
(c) NH<sub>3</sub> (d) CO<sub>2</sub>
2. Which of the following is not a property of ionic compound?  
(a) Ionic compounds are usually crystalline solids.  
(b) Ionic compounds conduct electricity in solid-state.  
(c) Ionic compounds are insoluble in organic solvents.  
(d) Ionic compounds have high melting and boiling points.
3. Which type of hybridization occurs in methane?  
(a) sp<sup>3</sup>-hybridization (b) sp<sup>2</sup>-hybridization  
(c) sp-hybridization (d) None of these
4. Which type of hybridization occurs in beryllium chloride?  
(a) sp<sup>3</sup>-hybridization (b) sp<sup>2</sup>-hybridization  
(c) sp-hybridization (d) None of these
5. Which of these is NOT a covalent compound?  
(a) Carbon dioxide (b) Methane  
(c) Ammonia (d) None of these
6. Covalent bond occurs between:  
(a) two metals (b) two non-metals  
(c) a metal and a non-metal (d) none of these

7. This is the force which binds various metal atoms together.
  - (a) ionic bond
  - (b) covalent bond
  - (c) metallic bond
  - (d) none of these
8. Molecular geometry is explained primarily through \_\_\_\_\_ theory.
  - (a) Valence Bond
  - (b) VSEPR
  - (c) Molecular Orbital
  - (d) None of these
9. Which of the following is a directional bond?
  - (a) Coordinate
  - (b) Covalent
  - (c) Metallic
  - (d) All these
10. This type of bonding is formed by mutual sharing of electrons between the two atoms but the shared pair of electrons is contributed only by one of the two atoms, the other atom simply participates in sharing.
  - (a) Coordinate
  - (b) Covalent
  - (c) Metallic
  - (d) Coordinate Covalent

## II. State True or False

1. Ionic compounds are insoluble in water.
2. Ionic bond involves a metal and a non-metal.
3. The ionic bond formed by sharing of electrons between two atoms.
4. Covalent compounds have usually low melting points and low boiling points.
5. Covalent compounds conduct electricity.
6. Molecular shape of  $\text{NH}_3$  is trigonal pyramidal.
7. Metals are ductile.
8. Metals are not lustrous.
9. The magnitude of intermolecular forces control the melting, boiling points of the substance.
10. Dipole moment of carbon dioxide is 1.83D.

## III. Answer the Following Questions

1. What do you mean by chemical bonding? What are its different type?
2. What is ionic bonding? Give at least two examples of ionic compounds.
3. List some factors that influence the formation of ionic bonding.
4. List some properties of ionic compounds.
5. Why do ionic compounds conduct electricity when dissolved in water?
6. What do you mean by hybridization?

7. What do you mean by covalent bonding? Give at least two examples of covalent compounds.
8. List the properties of covalent compounds.
9. Why do covalent compounds not conduct electricity?
10. List some factors that influence the formation of covalent bonding.
11. What do you mean by metallic bonding?
12. List some factors that influence the formation of metallic bonding.
13. Illustrate some physical properties of metals.
14. Give two examples each of polar and non-polar molecules.
15. What is a polar covalent bond? Give an example of a polar diatomic molecular. Draw its structure.
16. In a molecule  $AX_3$ , there are 3 bond pairs and one lone pair around A. What shape will it adopt? Draw it.
17. What is Dative bond? How it is different from covalent bond?
18. Show how the coordinate covalent bond is formed between ammonia and the hydrogen ion.
19. Differentiate between ionic, covalent and metallic bonding.

## PROJECT WORK

Make a chart to compare ionic covalent and metallic bond in terms of structure, heat and electrical conductivities, boiling & melting points, solubility in water and non-polar solvents and hardness. Submit it in the class as an assignment.