Chapter 4

Structure of Molecules

Major Concepts

4.1 Why do atoms react?
4.2 Chemical bonds
4.3 Types of bonds
4.4 Intermolecular forces
4.5 Nature of bonding and properties

Time allocation	
Teaching periods	16
Assessment periods	04
Weightage	8%

Students Learning Outcomes

Students will be able to:

- Find the number of valence electrons in an atom using the Periodic Table.
- Describe the importance of noble gas electronic configurations.
- State the octet and duplet rule.
- Explain how elements attain stability.
- Describe the ways in which bonds may be formed.
- State the importance of electronic configurations in formation of ion.
- Describe formation of cations from an atom of a metallic element.
- Describe formation of anion from a non-metallic element.
- Describe characteristic of ionic bond.
- Recognize a compound as having ionic bonds.
- Identify characteristics of ionic compounds.
- Describe formation of covalent bond between two non-metallic elements.
- Describe with examples single, double and triple covalent bonds.
- Draw electron cross and dot structure of simple covalent molecules containing single, double and triple covalent bonds.

Introduction

IThe things around us are composed of matter. All matter is made up of the building units 'atoms'. These atoms combine to form molecules, which appear in different states of matter around us. The forces responsible for binding the atoms together in a molecule are called chemical forces or chemical bonds. These bonding forces which keep the atom together will be discussed in this chapter.

4.1 WHY DO ATOMS FORM CHEMICAL BONDS?

It is a universal rule that everything in this world tends to become more stable. Atoms achieve stability by attaining electronic configuration of noble gases (He, Ne or Ar, etc) i.e. ns^2 np^6. Having 2 or 8 electrons in the valence shell is sign of stability. *Attaining two electrons in the valence shell is called* ***duplet*** *rule while attaining eight electrons in the valence shell is called* ***octet*** *rule.*

The noble gases do have 2 or 8 electrons in their valence shells. It means all the noble gases have their valence shells completely filled. Their atoms do not have vacant space in their valence shell to accommodate extra electrons. Therefore, noble gases do not gain, lose or share electrons. That is why they are non-reactive.

The importance of the noble gas electronic configuration lies in the fact that all other atoms try their best to have the noble gas electronic configuration. For this purpose, atoms combine with one another, which is called chemical bonding. In other words, atoms form chemical bonds to achieve stability by acquiring inert gas electron configuration.

An atom can accommodate 8 electrons in its valence shell in three ways:

i. By giving valence shell electrons (if they are less than three) to other atoms.

ii. By gaining electrons from other atoms (if the valence shell has five or more electrons in it).

iii. By sharing valence electrons with other atoms.

It means every atom has a natural tendency to have 2 or 8 electrons in its valence shell. The atoms having less than 2 or 8 electrons in their valence shells are unstable.

Now the question arises that how can we identify the way an atom reacts? The position of an atom in the periodic table indicates its group number. As we have studied in chapter 3, the group number is assigned on the basis of valence shell electrons. For example, group 1 has only 1 electron in its valence shell and group 17 has 7 electrons in its valence shell. Mode of reaction of an atom depends upon its number of valence shell electrons. It is discussed in the next sections.

4.2 CHEMICAL BOND

A **chemical bond** *is defined as a force of attraction between atoms that holds them together in a substance.* In other words, during bond formation there is some force which holds the atoms together.

This attaining of 8 electrons configuration in the outermost shell either by sharing, by losing or by gaining electrons, is called **octet rule**. This octet rule only symbolizes that noble gas electronic configuration should be attained by atoms when they combine or react. For elements like hydrogen or helium; which have only s-subshell, this becomes *'duplet rule'*. It plays a significant role in understanding the formation of chemical bond between atoms.

If the bond formation is between ions, it is due to an electrostatic force of attraction between them. But if bond formation is between similar atoms or between the atoms that have comparable electronegativities, then the chemical bond formation is by '*sharing*' of electrons. This sharing of electrons may be mutual or one sided.

When two approaching atoms come closer, the attractive as well as repulsive forces become operative. The formation of a chemical bond is a result of net attractive forces which dominate. The energy of that system is lowered and molecule is formed. Otherwise if repulsive forces become dominant no chemical bond will be formed. In that case there will be increase in the energy of the system due to creation of repulsive forces.

4.3 TYPES OF CHEMICAL BOND

The valence electrons, which are involved in chemical bonding, are termed as **bonding electrons**. They usually reside in the incomplete or partially filled outermost shell of an atom. Depending upon the way how these valence electrons are involved in bonding, they result in following four types of chemical bonds:

- Ionic Bond
- Covalent Bond
- Dative Covalent or Coordinate Covalent Bond
- Metallic Bond

4.3.1 Ionic Bond

The elements of Group-1 and Group-2 being metals have the tendency to lose their valence electrons forming positively charged ions. Whereas non-metals of Group-15 to Group-17 have the tendency to gain or accept electrons. They are electronegative elements with high electron affinities. If atoms belonging to these two different groups, metals and non-metals, are allowed to react, chemical bond is formed. *This type of chemical bond, which is formed due to complete transfer of electron from one atom to another atom, is called* **ionic bond**.

The formation of NaCl is a good example of this type of bond.

$$2Na_{(s)} \quad + \quad Cl_{2(g)} \longrightarrow \quad 2NaCl_{(s)}$$

Sodium chloride is a simple compound formed by sodium (**Z** =11) and chlorine (**Z**=17) atoms. The ground state electronic configuration of these elements is shown below:

$$_{11}Na \;=\; 1s^2, 2s^2\,2p^6, \boxed{3s^1} \quad \text{or} \quad Na^{\bullet}$$

$$_{17}Cl \;=\; 1s^2, 2s^2\,2p^6, \boxed{3s^2\,3p^5} \quad \text{or} \quad \overset{\times\times}{\underset{\times\times}{\times\times Cl\times}}$$

The frames indicate electrons in the valence shells of these elements; sodium has only one electron and chlorine has seven electrons. Sodium being electropositive element has the tendency to lose electron and chlorine being an electronegative element has the tendency to gain electron. Therefore, they form positive and negative ions by losing and gaining electrons, respectively. They attain electronic configuration to the nearest noble gases.

$$Na^{\bullet} \longrightarrow Na^{+} + e^{-}$$

$1s^2, 2s^2, 2p^6$ (Ne)

$$\overset{\times\times}{\underset{\times\times}{\times\times Cl \times}} + e^{-} \longrightarrow \overset{\times\times}{\underset{\times\times}{\times\times Cl \overset{\bullet}{\times}}}^{-} \quad or \quad Cl^{-}$$

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ (Ar)

By losing one electron from the outermost shell, sodium becomes Na^+ ion and it is left with 8 electrons in the second shell which will now become the valence shell. By gaining one electron, chlorine atom now also has eight electrons in its outermost shell and becomes Cl^- ion. Both of these atoms are now changed into oppositely charged ions. They stabilize themselves by combining with each other due to electrostatic force of attraction between them such as:

$$Na^{+} + Cl^{-} \longrightarrow NaCl$$

It is to be noted that only valence shell electrons take part in this type of bonding, while other electrons are not involved. In such type of reaction heat is usually given out. The compounds formed due to this type of bonding are called **ionic compounds**.

Test yourself 4.1

i. Why does sodium form a chemical bond with chlorine?
ii. Why does sodium lose an electron and attains +1 charge?
iii. How do atoms follow octet rule?
iv. Which electrons are involved in chemical bonding?
v. Why does group 1 elements prefer to combine with group 17 elements.
vi. Why chlorine can accept only 1 electron?
vi. Why and how elements are arranged in a period?

4.3.2 Covalent Bond

The elements of Group-13 to Group-17 when allowed to react with each other, they form a chemical bond by mutual sharing of their valence shell electrons. This type of *bond, which is formed due to mutual sharing of electrons,* is called a **covalent bond**.

The energy changes during the covalent bond formation are of considerable value. When two atoms approach each other attractive forces develop between electrons of one atom and nucleus of the other atom. Simultaneously, repulsive forces between electrons of the two atoms as well as between their nuclei are also created. When the attractive forces dominate due to decrease in distance between those two atoms, a chemical bond is formed between them. The formation of hydrogen, chlorine, nitrogen and oxygen gases are few examples of this type of bonding.

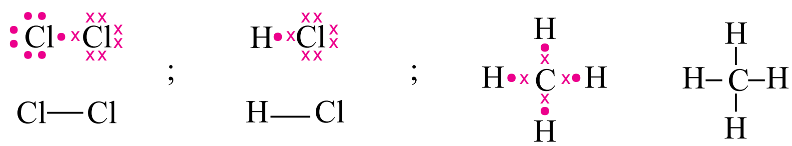
Types of covalent bonds

As described above, the covalent bond is formed by mutual sharing of electrons between two atoms. *The electrons that pair up to form a chemical bond are called* '**bond pair**' electrons. Depending upon the number of bond pairs, covalent bond is classified into following three types:

Single Covalent bond

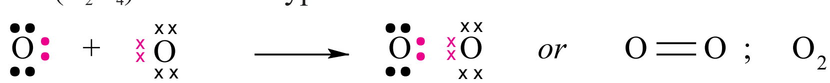
When one electron is contributed by each bonded atom, one bond pair is formed and it forms a **single covalent bond**. While drawing the structure of such molecules the single bond pair is indicated by a line between those two atoms. A few examples of molecules with single covalent bonds are hydrogen (H_2), chlorine (CI_2), hydrochloric acid (HQ) and methane (CH_4).

single *covalent bond*



Double Covalent bond

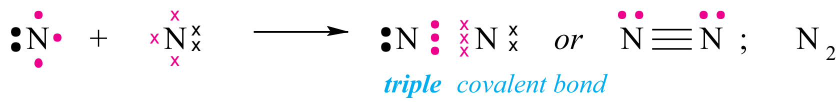
When each bonded atom contributes two electrons, two bond pairs are shared and a **double covalent bond** is formed. These bond pairs are indicated as double line between those atoms in the structure of such molecules. The molecules like oxygen (O_2) gas and ethene (C_2H_4) show such type of double covalent bonds.



double *covalent bond*

Triple Covalent Bond

When each bonded atom contributes three electrons, three bond pairs are involved in bond formation. This type is called **triple covalent bond**. Three small lines are used to indicate these three pairs of electrons between those atoms in the molecules of such compounds. The examples of molecules having triple covalent bonds are nitrogen (N_2) and ethyne (C_2H_2).



By this mutual sharing of valence shell electrons, each of the contributing atom attains the 'Octet' or nearest noble gas electronic configuration.

Do you know?

*The electronic configuration of the valence shells of atoms is shown in small 'dots' or 'crosses' around the symbol of the element. Each dot or cross represents an electron. This is a standard method of Lewis to describe the electronic configuration of valence shell of an atom. It is called **Lewis Structure Diagram**.*

4.3.3 Dative Covalent or Coordinate Covalent Bond

Coordinate covalent or dative covalent bonding is a type of *covalent bonding in which the bond pair of electrons is donated by one bonded atom only.* The atom which donates the electron pair is called **donor** and the atom which accepts the electron pair is called **acceptor**. A small arrow is usually used to indicate the atom and pair of electron being donated. The head of arrow is towards the acceptor atom.

The non-bonded electron pair available on an atom, like the one available on nitrogen in ammonia, (NH_3) is called a **lone pair**. When a proton (H^+) approaches a molecule with a lone pair of electrons, that *lone pair* is donated to H^+ and a **coordinate covalent bond** is formed, e.g. formation of ammonium radical (NH_4^+).

In the formation of BF_3 (boron trifluoride) molecule, three valence electrons of boron atom (Z= 5) pair up with three electrons, one from each three fluorine atoms. The boron atom even after this sharing of electrons (covalent bond formation), remains short or deficit of two electrons in its outermost shell. Now if a molecule with a lone pair approaches this molecule, it accepts lone pair from that donor and forms a coordinate covalent bond. The lone pair on nitrogen of ammonia molecule makes it a good donor molecule to form a coordinate covalent bond as shown in figure 4.1.

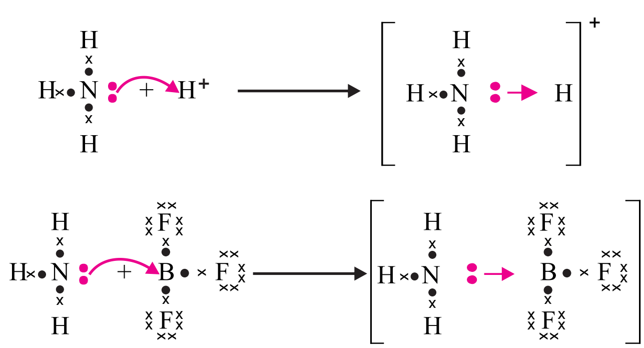


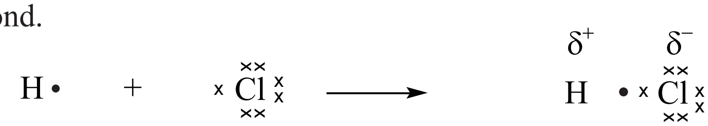
Fig. 4.1 Dative covalent bond (red arrow)

4.3.4 Polar and Non-polar Covalent Bond

If a covalent bond is formed between two similar atoms (homo-atoms), the shared pair of electrons is attracted by both the atoms equally. Such type of bond is called **non-polar** covalent bond. These bonds are formed by equal sharing of electron pair between the two bonding atoms. This type of bond is called a pure covalent bond. For example, bond formation in H_2 and CI_2.

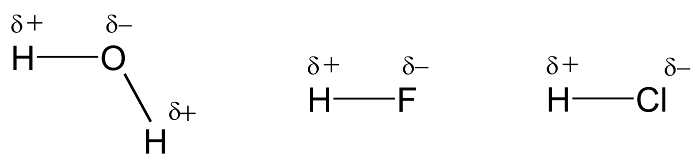
If the covalent bond is formed between two different types of atoms (hetro-atoms) then the bond pair of electrons will not be attracted equally by the bonded atoms. One of the atoms will attract the bond pair of electrons more strongly than the other one. This atom(element) will be called as more electronegative.

When there is difference of electronegativity between two covalently bonded atoms, there will be unequal attraction for the bond pair of electrons between such atoms. It will result in the formation of polar covalent bond. The difference between electronegativities of hydrogen and chlorine is 1.0. As the electronegativity of chlorine is more, it attracts the shared pair of electron towards itself with a greater force. A partial negative charge is therefore created on chlorine and in turn a partial positive charge on hydrogen due to electronegativity difference. It creates polarity in the bond and is called polar covalent bond.



E.N. = 2.2 ***E.N.*** = 3.2

The delta (δ) sign indicates partial positive or partial negative charge that is developed due to unequal sharing of shared pair or bonded pair of electrons. The compounds resulting from polar covalent bonds are called **polar compounds**. For example: water, hydrogen fluoride and hydrogen chloride.



By using electronegativity values, it is possible to predict whether a chemical bond will be ionic or covalent in nature. A bond formed between elements of high electronegativity (halogen group) and elements of low electronegativity (alkali metals) are ionic in nature. There is complete transfer of electrons between them. The bond between elements of comparable electronegativities will be covalent in nature as the bond between carbon and hydrogen in methane, or nitrogen and hydrogen in ammonia. *If the difference of electronegativities between two elements is more than 1.7 the bond between them will be predominantly ionic bond and if it is less than 1.7, the bond between two atoms will be predominantly covalent.*

Test yourself 4.2

i. *Give the electronic configuration of carbon atom.*
ii. *What type of elements have tendency of sharing of electrons?*
iii. *If repulsive forces dominate to attractive forces will a covalent bond form?*
iv. *Considering the electronic configuration of nitrogen atom, how many electrons are involved in bond formation and what type of covalent bond is formed.*
v. *Point out the type of covalent bonds in the following molecules*
CH_4, C_2H_4, H_2, N_2, and O_2
vi. *What is a lone pair? How many lone pairs of electrons are present on nitrogen in ammonia?*
vii. *Why is the BF_3 electron deficient?*
viii. *What types of electron pairs make a molecule good donor?*
ix. *What is difference between bonded and lone pair of electron and how many bonded pair of electrons are present in NH_3 molecule?*
x. *What do you mean by delta sign and why it develops?*
xi. *Why does oxygen molecule not form a polar covalent bond?*
xii. *Why has water polar covalent bonds?*

4.3.5 Metallic Bond

The **metallic bond** is *defined as a bond formed between metal atoms (positively charged ions) due to mobile or free electrons*. The different properties shown by metals such as high melting and boiling points, good conductions of heat and electricity, hard and heavy nature, suggest existence of different type of chemical bond between atoms of metals.

In case of metals, the hold of nucleus over the outermost electrons is weak because of large sized atoms and greater number of shells in between nucleus and valence electrons. Furthermore, because of low ionization potentials, metals have the tendency to lose their outermost electrons easily. Resultantly, these loose or free electrons of all metal atoms move freely in the spaces between atoms of a metal. None of these electrons is attached to any particular atom. Either they belong to a common pool, or belong to all the atoms of that metal. Nuclei of metal atoms appear submerged in sea of these free mobile electrons. These mobile electrons are responsible for holding the atoms of metals together forming a metallic bond. A simple metallic bond is shown in figure 4.2.

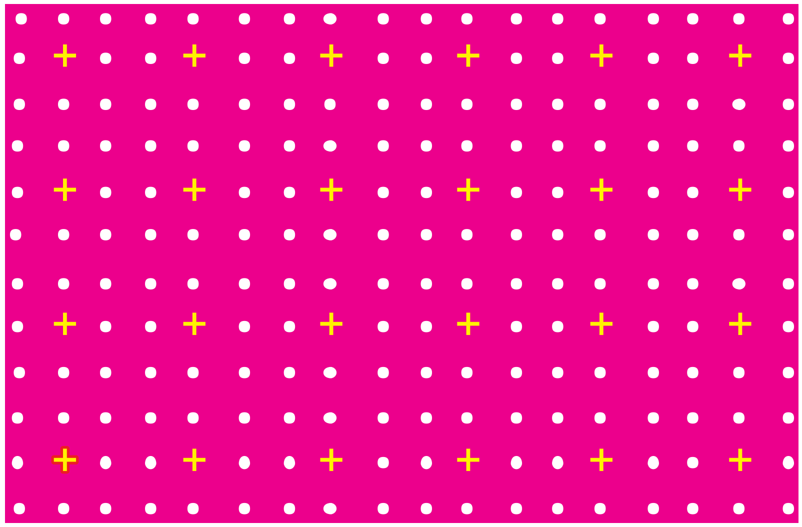
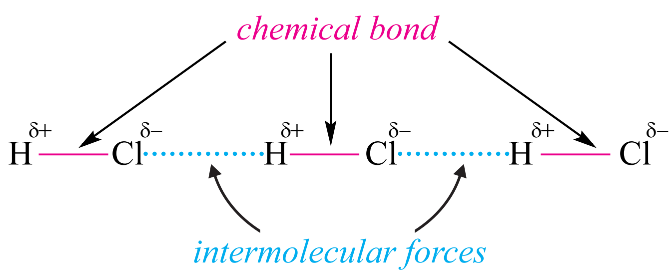


Fig. 4.2 A schematic diagram of Copper wire showing its positive nuclei (+) embedded in sea of free electrons (o) making 'Metallic Bonding'

4.4 INTERMOLECULAR FORCES

As discussed earlier, the forces that hold atoms in a compound are chemical bonds. In addition to these strong bonding forces, relatively weak forces also exist in between the molecules, which are called **intermolecular forces**. The bonding and intermolecular forces of hydrochloric acid are shown below:



It requires about 17 *kJ* energy to break these *intermolecular forces* between one mole of liquid hydrogen chloride molecules to convert it into gas. Whereas, about 430 *kJ* energy's required to break the chemical bond between hydrogen and chlorine atoms in 1 mole of hydrogen chloride.

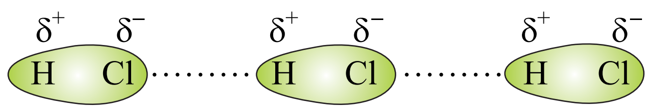
4.4.1 Dipole - Dipole Interaction

All intermolecular forces, which are collectively called **van der Waals** forces, are electrical in nature. They result from the attractions of opposite charges which may be temporary or permanent. The unequal sharing of electrons between two different types of atoms make one end of molecule slightly positive and other end slightly negatively charged. As shared pair of electron is drawn towards more electronegative atom, it is partially negatively charged, as chlorine in hydrogen chloride. The other end automatically becomes partially positively charged.

$$\overset{\delta^+}{H} — \overset{\delta^-}{Cl}$$

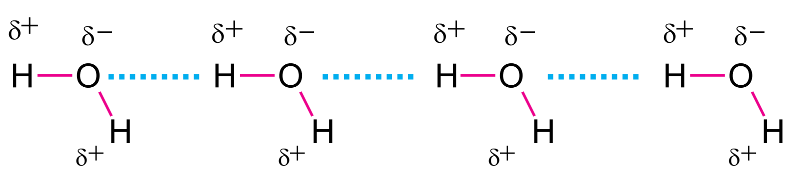
When partial positive and partial negative charges exist at different positions in a molecule, the adjacent molecules will arrange themselves in such a way that negative end of that molecule comes near to positive end of other molecule. It results in a net forces of attraction between oppositely charged ends of two adjacent molecules. These attractive forces are called dipole – dipole interactions as represented in HC1:

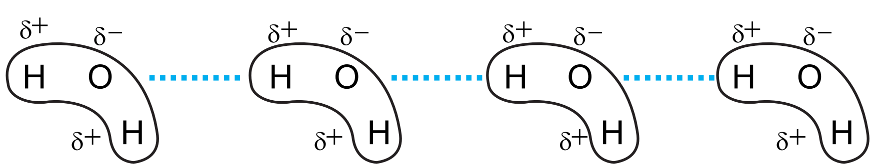
$$\overset{\delta^+}{H} — \overset{\delta^-}{Cl} \cdots\cdots\cdots \overset{\delta^+}{H} — \overset{\delta^-}{Cl} \cdots\cdots\cdots \overset{\delta^+}{H} — \overset{\delta^-}{Cl}$$



4.4.2 Hydrogen Bonding

Hydrogen bonding is a special type of intermolecular forces present in the permanently polar molecules. This bonding can be considered unique dipole-dipole attraction. This force of attraction develops between molecules that have a hydrogen atom bonded to a small, highly electronegative atom with lone pairs of electrons such as nitrogen, oxygen and fluorine. The covalent bond between hydrogen atom and other atom becomes polar enough to create a partial positive charge on hydrogen atom and a partial negative charge on the other atom. The small size and high partial positive charge on the hydrogen atom enables it to attract highly electronegative (N,O or F) atom of the other molecule. *So, partially positively charged hydrogen atom of one molecule attracts and forms a bond with the partially negatively charged atom of the other molecule, the bonding is called* **hydrogen bonding**. This force of attraction is represented by a dotted line between the molecules as shown below:





Hydrogen bonding affects the physical properties of the molecules. Due to this boiling points of the compounds are affected greatly. For example, boiling point of water

(100 °C) is higher than that of alcohol (78 °C) because of more and stronger hydrogen bonding in water.

The important phenomenon of floating of ice over water is because of hydrogen bonding. The density of ice at 0 °C (0.917 gem^{-3}) is less than that of liquid water at 0°C (1.00 $gem^{-3)}$. In the liquid state water molecules move randomly. However, when water freezes, the molecules arrange themselves in an ordered form, that gives them open structure. This process expands the molecules, that results in ice being less dense as compared to water.

Test yourself 4.3

i. What type of elements form metallic bonds?
ii. Why is the hold of nucleus over the outermost electrons in metals weak?
iii. Why the electrons move freely in metals?
iv. Which types of electrons are responsible for holdings the atoms together in metals.
v. Why a dipole develops in a molecule ?
vi. What do you mean by induced dipole ?
vii. Why are dipole forces of attraction not found in halogen molecules?
viii. What types of attractive forces exist between HCl molecules?
ix. Define intermolecular forces; show these forces among HCl molecule.

4.5. NATURE OF BONDING AND PROPERTIES

Properties of the compounds depend upon the nature of bonding present in them. Let us discuss the effects of nature of bonding on the properties of compounds.

4.5.1 Ionic Compounds

Ionic compounds are made up of positively and negatively charged ions. Thus they consist of ions and not the molecules. These positively and negatively charged ions are held together in a solid or crystal form with strong electrostatic attractive forces. The orderly arrangement of Na+ and CI'' ions in a solid crystal of sodium chloride is shown in figure 4.3.

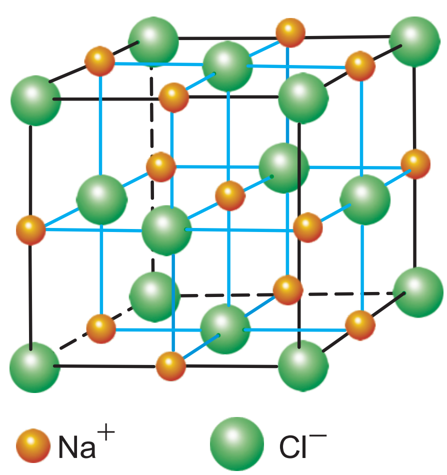


Figure 4.3 Regular arrangement of Na^+ and Cl^- ions in solid crystal of NaCl

The ionic compounds have following properties:

i. Ionic compounds are mostly crystalline solids.

ii. Ionic compounds in solid state have negligible electrical conductance but they are good conductors in solution and in the molten form. It is due to presence of free ions in them.

iii. Ionic compounds have high melting and boiling points. For example, sodium chloride has melting point 800 °C and a boiling point 1413 °C. As ionic compounds are made up of positive and negative ions, there exist strong electrostatic forces of attraction between oppositely charged ions. So, a great amount of energy is required to break these forces.

iv. They dissolve easily in polar solvents like water. Water has high dielectric constant that weakens the attraction between ions.

4.5.2 Covalent Compounds

The covalent compounds are made up of molecules that are formed by mutual sharing of electrons between their atoms i.e. covalent bonds. A covalent bond is generally regarded as weaker than an ionic bond. Covalent compounds are made up of two or more non-metals, e.g. H_2, CH_4, CO_2, H_2SO_4, $C_6H_{12}O_6$. Lower molecular mass covalent compounds are gases or low boiling liquids. Contrary to it, higher molecular mass covalent compounds are solids. General properties shown by covalent compound are as follows:

i. They have usually low melting and boiling points.

ii. They are usually bad conductors of electricity. The compounds having polar character in their bonding are conductor of electricity when they dissolve in polar solvents.

iii. They are usually insoluble in water but are soluble in non-aqueous solvents like benzene, ether, alcohol and acetone.

iv. Large molecules with three dimensional bonding form covalent crystals which are very stable and hard. They have very high melting and boiling points.

Polar and Non-Polar Compounds

As discussed earlier the polarity in a chemical bond is due to difference in electronegativities of the bonding atoms. On the **Pauling Scale**, fluorine has been given an electronegativity value of 4.0. The values for other elements are calculated relative to it.

Properties of non-polar and polar covalent compounds differ to some extent. Non-polar covalent compounds usually do not dissolve in water while polar covalent compounds usually dissolve in water. Similarly non-polar compounds do not conduct electricity but an aqueous solution of a polar compound usually conduct electricity due to the formation of ions as a result of its reaction with water.

4.5.3 Coordinate Covalent Compounds

Their properties are mostly similar to those of covalent compounds. As the nuclei in these compounds are held by shared pair of electrons, therefore, they do not form ions in water. Due to their covalent nature they form solutions in organic solvents and are very less soluble in water. Usually they are rigid compounds with a dipole.

4.5.4 Metals

Metals have common property of conducting heat and electricity. It gives them prime role in many industires. Major properties shown by the metals are as follows:

i. They show metallic luster.

ii. They are usually malleable and ductile. Malleability is the property by virtue of which a metal can be rolled into sheets, while ductility is the property by virtue of which a metal can be drawn into wires.

iii. They have usually high melting and boiling points.

iv. Being greater in size they have low ionization energies and form cations (M^+) very easily.

v. They are good conductors of heat and electricity in solid and liquid state due to mobile electrons.

Test yourself 4.4

i. Why the ionic compounds have high melting and boiling points?
ii. What do you mean by malleability?
iii. Why are ionic compounds easily soluble in water?
iv. What type of bond exists in sodium chloride ?
v. Why the covalent compounds of bigger size molecules have high melting points?
vi. (a): What is the electronegativity difference between the following pair of elements (atoms). Predict the nature of the bond between them?
(a) H and CI (b)H and Na (c) Na and I (d) K and CI
(b): Comparing the electronegativity differences, arrange these compounds in increasing ionic strength.

Synthetic Adhesives

Although natural adhesives are less expensive to produce, but most important adhesives used now a days are synthetic. Adhesives based on synthetic resins and rubbers excel in versatility and performance. Synthetic adhesives can be produced in a sufficient supply with uniform properties and they can be modified in many ways. The polymers or resins used in synthetic adhesives fall into two general categories—thermoplastics and thermoseting. One form of polymer used industrially is epoxy adhesive.

AIR CRAFTS, CARS, TRUCKS AND BOATS ARE PARTIALLY HELD TOGETHER WITH EPOXY ADHESIVES

Epoxy is polymer that is formed from two different chemicals. These are referred to as resin and the hardener. Epoxy adhesives are called structural adhesives. These high-performance adhesives are used in the construction of aircraft, automobiles, bicycles, boats, golf clubs, where high strength bonds are required. Epoxy adhesives can be developed to suit almost any application. They can be made flexible or rigid, transparent or opaque even colored as well as fast or slow setting. Epoxy adhesives are good heat and chemical resistant. Because of these properties, they are given the name of engineering adhesives.

Key Points

- Atoms of different elements react to attain noble gas configuration, which is stable one.
- Chemical bonds may be formed by complete transfer of electrons (ionic); mutual sharing (covalent) or by donation from an atom(coordinate or dative covalent).
- Metals have the tendency to lose electrons easily forming cations.
- Non-metals have tendency to gain electrons and form anions.
- In ionic bonding strong electrostatic force hold ions together.
- Ionic compounds are solids with high melting and boiling points.
- Covalent bonds among non-metals are weaker than ionic bonds.
- Ionic bonds are non-directional, but covalent bonds are formed in a particular direction.
- Covalent bonds formed between similar atoms are non-polar while between different atoms are polar.
- In covalent bonding single, double or triple covalent bond is formed by sharing of one, two or three electron pairs by the bonded atoms.
- Coordinate covalent bond is formed between electron pair donors and electron pair acceptors.
- Metallic bond is formed between metal atoms due to free electrons.
- In addition to chemical bonds, intermolecular forces of attraction exist between polar molecules.
- Hydrogen bonding exists between the hydrogen atom of one molecule and highly electronegative atom of other molecule.
- Hydrogen bonds affect the physical properties of the compounds.

- Properties of the compounds depend upon the nature of bonding present in the compound.
- Ionic compounds are crystalline solid with high melting and boiling points.
- Covalent compounds exist in molecular form in three physical states.
- Polar and non- polar covalent compounds have different properties.
- Metals have shining surface. They are good conductor of electricity and are malleable and ductile

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer

1. Atoms react with each other because:
(a) they are attracted to each other. (b) they are short of electrons
(c) they want to attain stability (d) they want to disperse

2. An atom having six electrons in its valence shell will achieve noble gas electronic configuration by:
(a) gaining one electron (b) losing all electrons
(c) gaining two electrons (d) losing two electrons

3. Considering the electronic configuration of atoms which atom with the given atomic number will be the most stable one?
(a) 6 (b) 8 (c) 10 (d) 12

4. Octet rule is:
(a) description of eight electrons
(b) picture of electronic configuration
(c) pattern of electronic configuration
(d) attaining of eight electrons

5. Transfer of electrons between atoms results in:
(a) metallic bonding (b) ionic bonding
(c) covalent bonding (d) coordinate covalent bonding

6. When an electronegative element combines with an electropositive element the type of bonding is:
(a) covalent (b) ionic
(c) polar covalent (d) coordinate covalent

7. A bond formed between two non-metals is expected to be:
(a) covalent (b) ionic
(c) coordinate covalent (d) metallic

8. A bond pair in covalent molecules usually has:
(a) one electron (b) two electrons
(c) three electrons (d) four electrons

9. Which of the following compounds is not directional in its bonding?
(a) CH_4(b) Kbr (c) CO_2 (d). H_2O

10. Ice floats on water because:
(a) ice is denser than water (b) ice is crystalline in nature
(c) water is denser than ice (d) water molecules move randomly

11. Covalent bond involves the
(a) donation of electrons (b) acceptance of electrons
(c) sharing of electrons (d) repulsion of electrons

12. How many covalent bonds does C_2H_2 molecule have?
(a) two (b) three (c) four (d) five

13. Triple covalent bond involves how many electrons?
(a) eight (b) six (c) four (d) only three

14. Which pair of the molecules has same type of covalent bonds?
(a) O_2 and HC1 (b) O_2 and N_2
(c) O_2 and C_2H_4 (d) O_2 and C_2H_2

15. Identify the compound which is not soluble in water.
(a) C_6H_6 (b)NaCl (c) KBr (d) MgC_{12}

16. Which one of the following is an electron deficient molecule?
(a) NH_3 (b)BF_3 (c) N_2 (d) O_2

17. Identify which pair has polar covalent bonds.
(a) O_2 and Cl_2 (b) H_2O and N_2
(c) H_2O and $_2H_2$ (d) H_2O and Hc1

18. Which one of the following is the weakest force among the atoms?
(a) ionic force (b) metallic force
(c) intermolecular force (d) covalent force

Short answer questions.

1. Why do atoms react?
2. Why is the bond between an electropositive and an electronegative atom ionic in nature?
3. Ionic compounds are solids. Justify.
4. More electronegative elements can form bonds between themselves. Justify.
5. Metals are good conductor of electricity. Why?
6. Ionic compounds conduct electricity in solution or molten form. Why?
7. What type of covalent bond is formed in nitrogen molecule.
8. Differentiate between lone pair and bond pair of electrons.

9. Describe at least two necessary conditions for the formation of a covalent bond.
10. Why HC1 has dipole-dipole forces of attraction?
11. What is a triple covalent bond, explain with an example?
12. What is difference between polar and non-polar covalent bonds, explain with one example of each?
13. Why a covalent bond becomes polar?
14. What is relationship between electronegativity and polarity?
15. Why does ice float on water?
16. Give the characteristic properties of ionic compounds.
17. What characteristic properties do the covalent compound have?

Short answer questions.

1. What is an ionic bond? Discuss the formation of ionic bond between sodium and chlorine atoms?
2. How can you justify that bond strength in polar covalent compounds is comparable to that of ionic compound?
3. What type of covalent bonds are formed between hydrogen, oxygen and nitrogen? Explain their bonding with dot and cross model.
4. How a covalent bond develops ionic character in it? Explain.
5. Explain the types of covalent bonds with at least one example of each type.
6. How a coordinate covalent bond is formed? Explain with examples?
7. What is metallic bond? Explain the metallic bonding with the help of a diagram.
8. Define hydrogen bonding. Explain that how these forces affect the physical properties of compounds.
9. What are intermolecular forces? Compare these forces with chemical bond forces with reference to HC1 molecule?
10. What is a chemical bond and why do atoms form a chemical bond?
11. What is octet rule? Why do atoms always struggle to attain the nearest noble gas electronic configuration?