Chapter 3

Chemical Bonding

Student Learning Outcomes

After studying this chapter, students will be able to:

- Describe that noble gas electronic configuration, octet and duplet rules help predict chemical properties of main group elements
- Compare between the formation of cations and anions
- Account for the electropositive and electronegative nature of metals and non-metals.
- Define ionic, covalent, coordinate covalent and metallic bonds
- Differentiate between ionic compounds and covalent compounds. (The following points need to be included in the respective definitions:
  a. Ionic Bond as strong electrostatic attraction between oppositely charged ions
  b. Covalent bond as strong electrostatic attraction between shared electrons and two nuclei
  c. Metallic bond as strong electrostatic attraction between cloud/sea of delocalized electrons and positively charged cations)
- Explain the properties of compounds in terms of bonding and structure
- Compare uses and properties of materials such as strength and conductivity as determined by the type of chemical bond present between their atoms.
- Interpret the strength of forces of attraction and their impact on melting and boiling points of ionic and covalent compounds.
- Justify the availability of free charged particles (electrons or ions) for conduction of electricity in ionic compounds (solid and molten) covalent compounds and metallic bonds.
- Recognize that some substances can ionize when dissolved in water. (e.g. acids dissolves in water and conduct electricity)
- Justify the suitability of usage of graphite, diamond and metals for industrial purposes (Some examples may include: **a.** graphite as lubricant or an electrode **b.** diamond in cutting tools **c.** metals for wires, and sheets)
- Draw the structure of ionic and covalent compounds along with their format ion. (some examples can include: **a.** ionic bonds in binary compounds such as NaBr, NaF, $CaCl_2$ using dot-and- cross diagrams and Lewisdot structures simple molecules including H_2, Cl_2, O_2, N_2, H_20, CH_4, NH_3, HCl, CH_3O H, C_2H_4, CO_2, HCN, and similar molecules using dot and-cross diagrams and Lewis- dot structures)

3.1 Why do atoms form chemical bonds?

Atoms have a tendency to decrease their energy. They can do this by combining with other atoms. It is a natural phenomenon because it increases the stability of atoms.

How do atoms succeed in lowering their energy? The early chemists had started thinking about this a long time ago. They finally succeeded to get an answer only when the noble gases He, Ne, Ar, Kr, Xe were discovered. Helium has two electrons in its outer shell while all other noble gases have eight electrons in their outermost shells. We also know about these gases that neither their atoms combine with themselves nor with other atoms. The probable reason for this lack of reactivity was their stability. It was suggested that these gases were stable due to the presence of two electrons in helium and eight electrons in the outermost shells of the rest of gases. This gave rise to a principle that having two electrons (for hydrogen and helium which have only the first shell) or eight electrons in the outermost shell meant stability and hence unreactivity as well. This principle was named as Duplet or Octet Rule.

The discovery of duplet or octet rule was followed by another similar suggestion that atoms form bonds because they would like to lower their energy by completing their duplet or octet. For example, for sodium atom it is easy to lose one electron and stabilize itself than to gain seven electrons while completing its octet. Sodium atom, therefore, adopts the energetically easier path and loses its electron to form a bond. In the same way, it is energetically favourable for hydrogen atom to lose one electron to become proton (H^+) or gain one electron to become hydride ion (H^-). In the latter case, it completes its duplet.

Alkali and alkaline earth metals are therefore expected to be electropositive metals which will form bonds with electronegative elements of 6th and 7th groups. Although, in the beginning, octet rule played a significant role in understanding the nature of a chemical bond, yet further investigations found it to be less important.

3.2 Chemical Bond

A chemical bond is a force of attraction between atoms which holds them together in the form of a molecule or a compound.

When atoms of different substances approach each other, there are two possibilities. They may attract or repel each other. If the forces of attraction between them dominate the forces of repulsion, the energy of the system gets lowered and as a result the two atoms will react to form a new molecule. Conversely, the two atoms simply move away from each other.

Important Information!

The arrangement of electrons around the nucleus of an atom in shells and sub-shells is called electronic configuration.

Types of Bonds

We shall consider here three types of bonds.

(1) Ionic bond
(2) Covalent bond
(3) Coordinate covalent bond

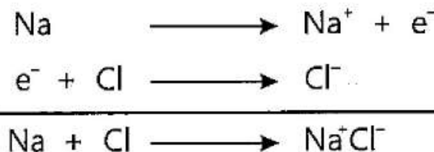
3.2.1 Ionic Bond

A chemical bond is formed as a result of the tendency of atoms to lose or gain electron or electrons to acquire the electronic configuration of the nearest noble gas because this is a more stable electronic structure. Let us take the example of the formation of a simple and important compound, sodium chloride. This compound is formed when the elements sodium and chlorine react chemically. The electronic configurations of these elements are shown in Fig (3.1).

	1st shell	2nd shell	3rd shell
$_{11}Na$	2	8	1
$_{17}Cl$	2	8	7

Fig (3.1): Electronic Configurations of Sodium and Chlorine

An electron from the outermost shell of sodium atom is transferred to the outermost shell of chlorine atom and in doing so, both these atoms acquire the electronic configurations of their nearest noble gases. (Fig 3.2)



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

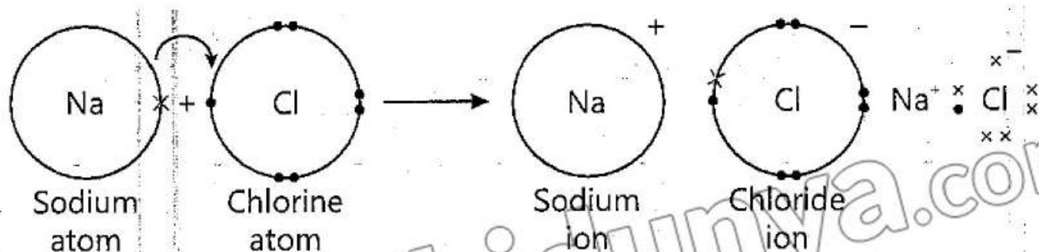
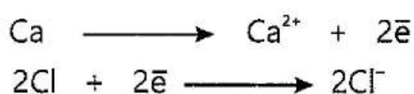


Fig (3.2): Transference of Electron from Sodium to Chlorine Atom

Similarly, sodium also reacts with fluorine and bromine to give sodium fluoride and sodium bromide respectively.

It should be noted here that an electron or electrons, which take part in a chemical reaction, come only from the outermost shells of the atoms. Sodium chloride, formed as a result of the chemical reaction mentioned on the previous page contains the positively charged sodium ions (Na^+) and the negatively charged chloride ions (Cl^-). These oppositely charged ions are then held together by the electrostatic force of attraction. The chemical bond, thus formed, is called an Ionic or an Electrovalent Bond and the compounds having such a bond are called ionic compounds.

Calcium, an alkaline earth metal, loses two electrons to form calcium chloride ($CaCl_2$). Fig (3.3)



$$Ca \longrightarrow Ca^{2+} + 2\bar{e}$$

$$2Cl + 2\bar{e} \longrightarrow 2Cl^-$$

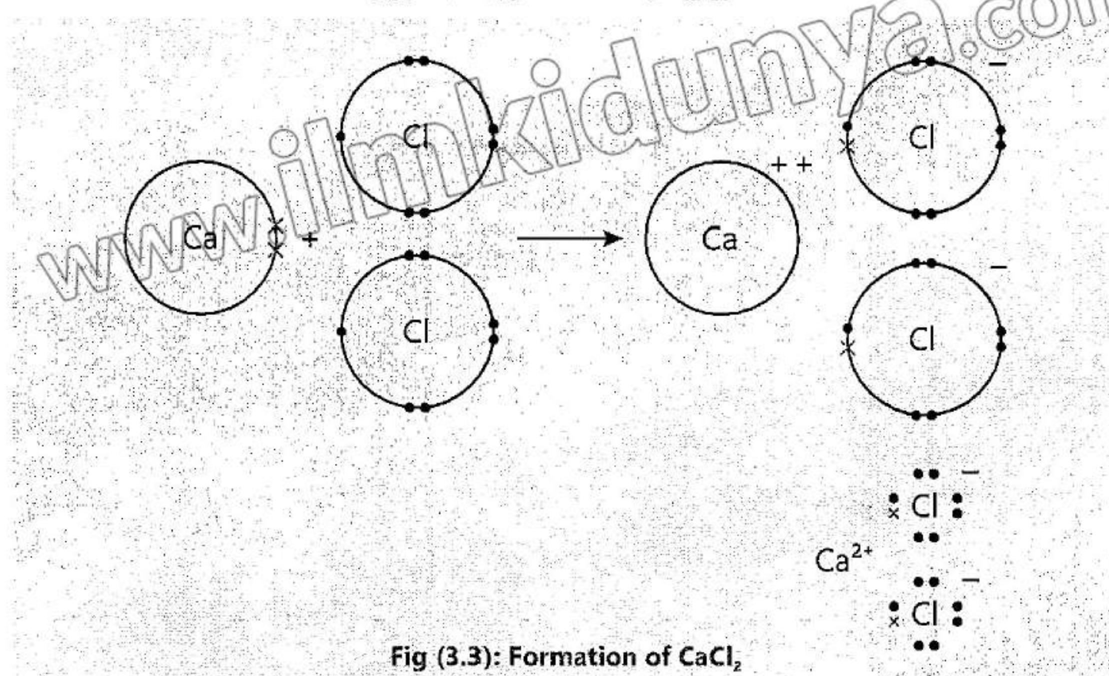


Fig (3.3): Formation of $CaCl_2$

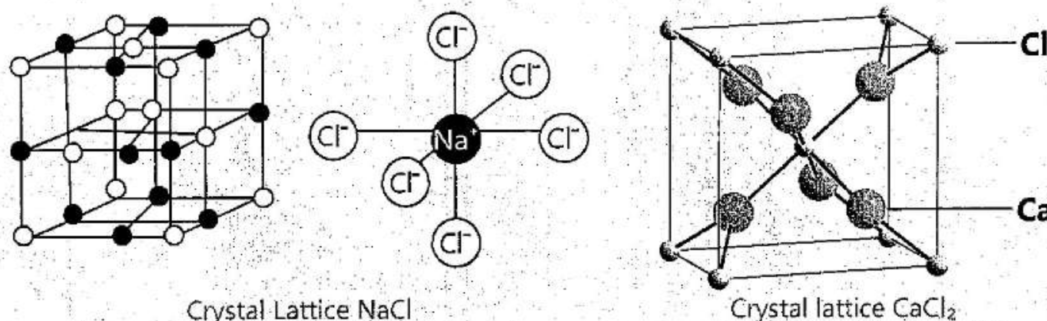
There ions then surround each other three dimensionally to form a crystal lattice.

Examples of ionic compounds are KCl, Mg F, NaF, Kbr, CaF_2

Exercise

1.	What types of elements form ionic bonds?
2.	What are the conditions for an ionic bond to form?

Figures of crystal lattices of NaCl, NaBr, NaF and $CaCl_2$. (fig 3.4)



Crystal Lattice NaCl Crystal lattice $CaCl_2$

Fig (3.4) Crystal Lattices of NaCl and $CaCl_2$

An ionic bond is therefore a bond which is formed by the complete transference of electron or electrons from one atom to another atom.

3.2.2 Covalent Bond

During the formation of an ionic bond, the atoms lower their energy by the transference of an electron and thus acquire the electronic configuration of the nearest noble gas. However, it is not the only way by which atoms can lower their energy. Some atoms decrease their energy by mutually sharing their electrons. This can be explained as follows.

When two atoms approach each other in order to form a bond, they undergo important changes in their energy. The electrons belonging to one atom will come under the attractive influence of the nucleus of the other atom. This is the new force of attraction and will be responsible for lowering the energy. The electrons and the nucleus of one atom will also repel the electrons and the nucleus of the other atom. This is the force of repulsion and will obviously increase the energy. The two atoms will bring themselves at such a distance so that the attractive forces dominate the repulsive forces. The total energy at this distance will be minimum and thus a stable molecule is formed. **A covalent bond is therefore a bond formed by the mutual sharing of an electron pair provided by the bonded atoms. This is called a single covalent bond.**

In some compounds, the atoms share two electrons each to form a double covalent bond. In the same way atoms can share three electrons each to form a triple covalent bond. Double and triple covalent bonds have two and three electron pairs respectively which are mutually shared between the two atoms. A single covalent bond is represented by a single line(–), a double covalent bond is represented by two lines (=) while a triple covalent bond is represented by three lines(≡). The mutually shared electrons may be shown by a dot or a cross. The formation of single, double and triple covalent bonds in different molecules is explained in the examples shown in Fig (3.5).

Exercise

What type of elements form covalent bond?

How covalent bond is different from an ionic bond?

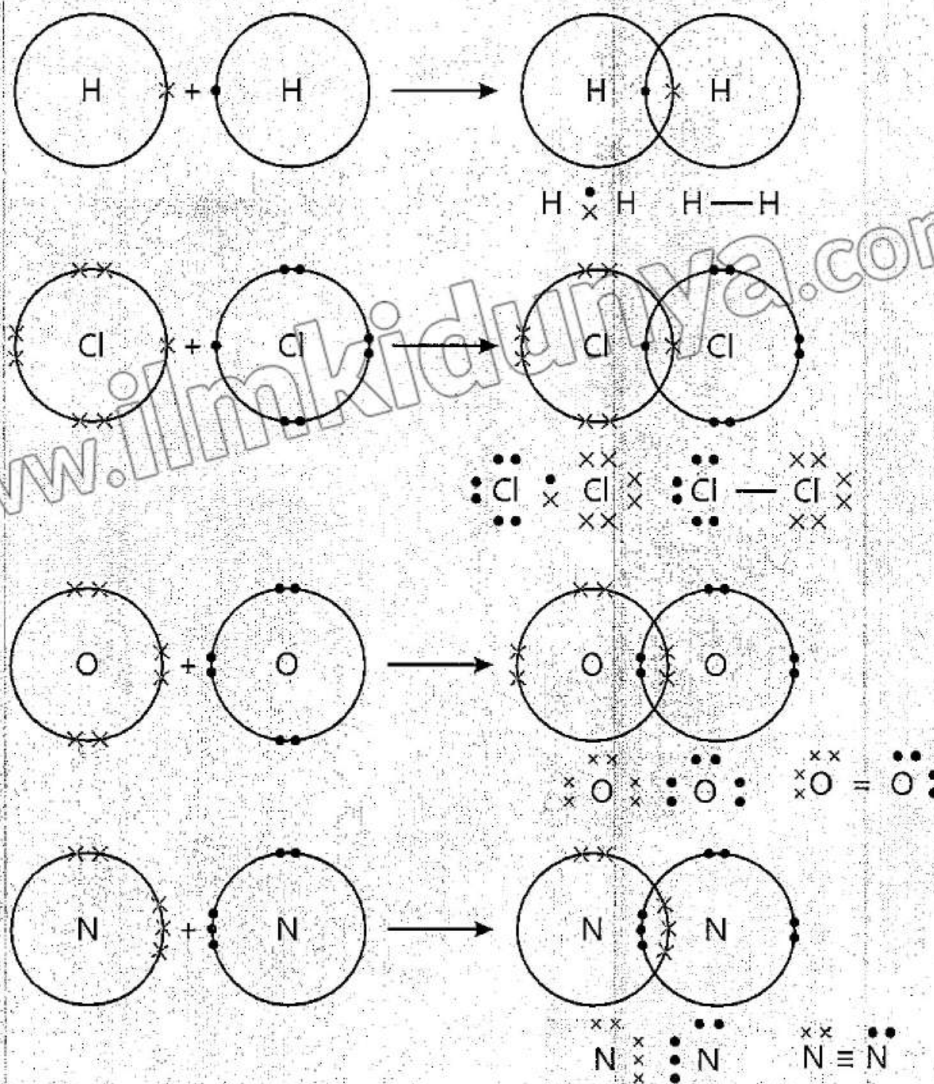
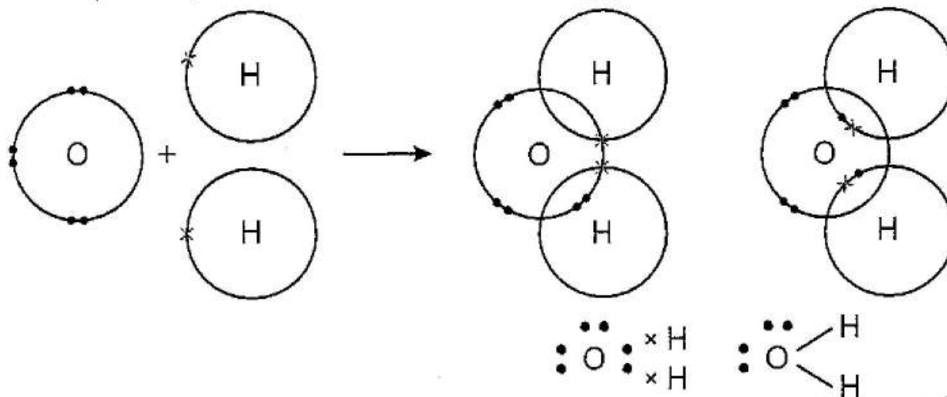


Fig (3.5): Formation of Single, Double and Triple Covalent Bonds

Formation of Covalent Compounds

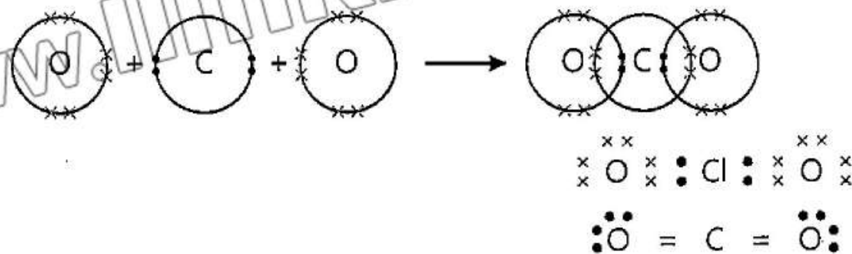
Water

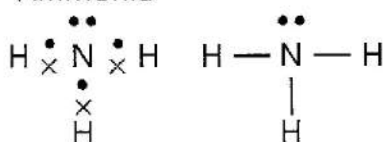
A water molecule is formed when two hydrogen atoms share their electrons separately with the electrons of one oxygen atom.

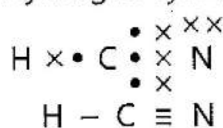


Carbon dioxide

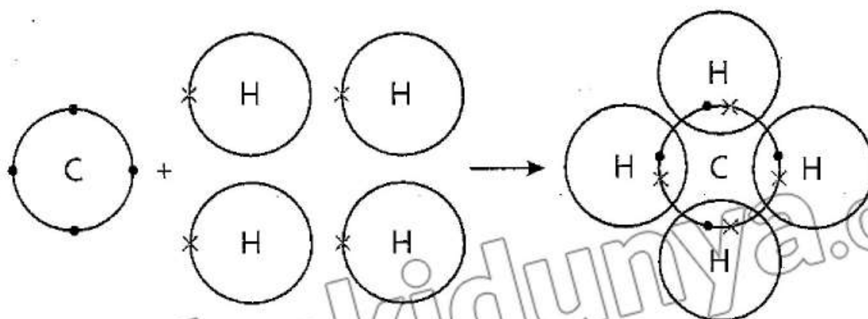
A carbon dioxide molecule is formed when an atom of carbon shares its four electrons with two oxygen atoms. Each oxygen atom also shares two electrons.



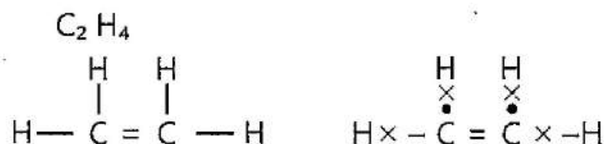




Methane



Ethene



Methanol

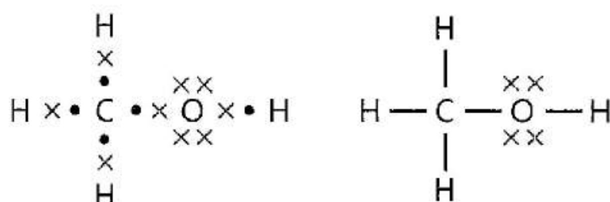
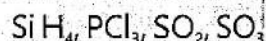


Fig (3.6): Formation of Covalent Compounds



Draw electron dot and cross structure of the following compounds.
SiH_4, PCl_3, SO_2, SO_3

It is quite clear from the examples shown above that after mutually sharing their electrons, the bonded atoms acquire the electronic configuration of the nearest noble gas.

3.2.3 Coordinate Covalent Bond

Coordinate covalent bond is a type of covalent bond in which the shared electron pair is donated by one atom only. This bond is formed when a molecule has an electron pair to donate to another molecule. The molecule which donates the electron pair, is called a donor while that which accepts it is called an acceptor. An arrow head (→) pointing towards the acceptor represents this type of bond. Following examples will help to explain this bond.

Hydronium Ion (H_3O^+)

Acids provide protons (H^+) when dissolved in water. This proton has an empty outer shell and can accept a pair of electrons present on the oxygen atom in water molecule. As a result of this, a hydronium ion (H_3O^+) is formed. (Fig 3.7)

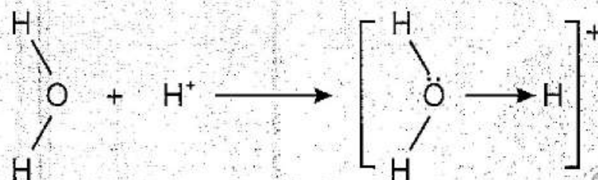


Fig (3.7): Formation of a Coordinate Covalent Bond Between H_2O and H^+

The positive charge covers whole of the hydronium ion. After the formation of hydronium ion, there does not remain any difference between a coordinate covalent bond and a covalent bond. All the three bonds of oxygen behave exactly alike.

Reaction Between $\ddot{N}H_3$ and BF_3

A reaction between ammonia (NH_3) and boron trifluoride (BF_3) is another example of the formation of a coordinate covalent bond. During the reaction, an electron pair from nitrogen of ammonia fills the partially empty outer shell of boron present in boron trifluorinde Fig (3.8).

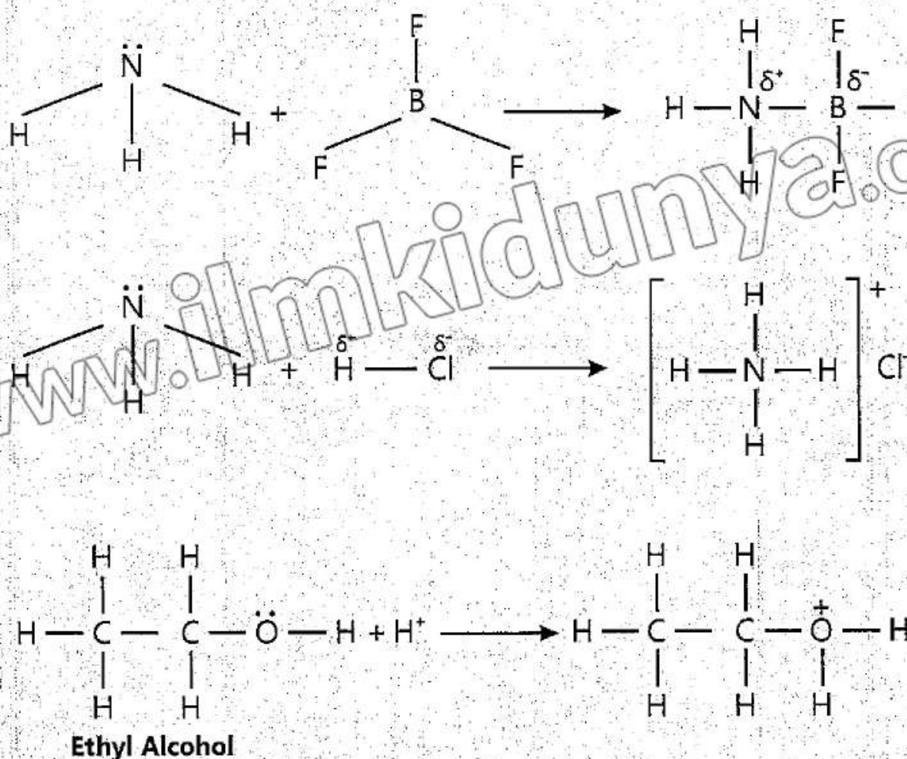


Fig (3.8): Formation of Boron trifluorinde ammonia, ammonium chloride and protonated ethyl alcohol

In the above example, a coordinate covalent bond in ammonium chloride links nitrogen of ammonia and the proton. The positive charge is spread all over ammonium ion. All the four bonds between nitrogen and hydrogen in ammonium ion behave exactly alike. This proves the point that the difference between a covalent bond and a coordinate covalent bond lies in the way they are formed. Once such bonds are formed, there does not remain any difference.

Exercise

Draw the pictures of coordinate covalent bond formed between:
(a) BF_3 and $AlCl_3$
(b) CH_3OCH_3 and H^+

Exercise

Which compound is not able to form a coordinate covalent bond?

3.3 Metallic Bond

The characteristics shown by metals are very different from those of ionic and covalent compounds. This suggests the presence of different types of binding forces among the metallic atoms.

Properties of Metals

1. Metals usually show metallic luster.
2. Metals usually have high melting and boiling points.
3. Metals are good conductors of heat and electricity.
4. Metals are usually hard and heavy.
5. Metals can be made into different shapes by applying pressure.

These characteristics of metals can be explained if we know the nature of binding forces present between their atoms.

Usually metals have low values of ionization energy. Their atoms can therefore, lose their outer electron or electrons easily. In other words, the nuclei of metallic atoms cannot hold their outer electrons firmly. For example, in sodium metal, each sodium atom is surrounded by eight other sodium atoms. The outer electrons of these atoms move freely between the vacant spaces present between atoms because of the loose linkage they have with their nuclei. No electron remains attached with any particular nucleus. Instead, all the electrons, at the same time, get attached with all the nuclei. When all the atoms attract all the electrons collectively, obviously they will be bound together. A metal will appear to have a sea of electrons in which all the nuclei of atoms are submerged. A metallic bond, is therefore a type of chemical bond which has positively charged ions bound together by the mobile electrons. Fig (3.9)

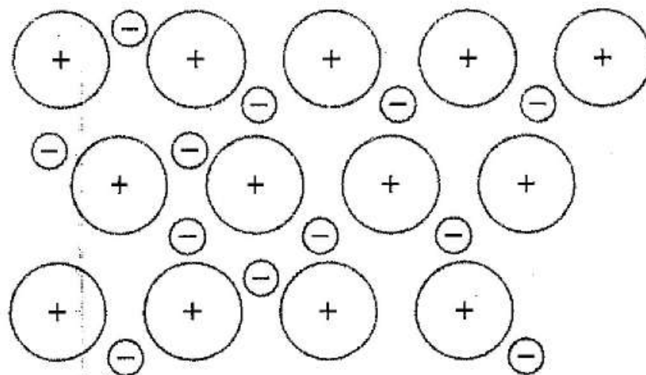


Fig (3.9): Metallic Bond in Sodium Metal

The strength of a metallic bond depends upon two factors: the number of positive charges present on the positive ions and the number of mobile electrons set free by each atom. In sodium metal, for example, each sodium atom sets free only one electron. The metallic bond in sodium metal is, therefore, not very strong. In magnesium metal, each magnesium atom releases two electrons to acquire two positive charges. The metallic bond in magnesium metal will evidently be stronger than that in sodium metal. This explains why the magnesium metal melts at a higher temperature than sodium metal. Fig (3.10)

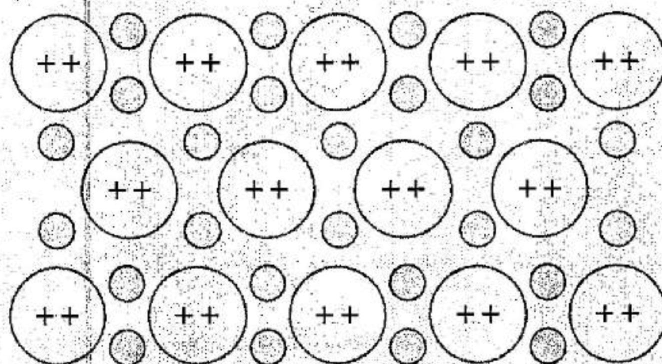


Fig (3.10): Metallic Bond in Magnesim Metal

The presence of freely moving electrons in metals makes them good conductor of heat and electricity. Moreover, in metals, the atoms are strongly held and arranged in the form of rows one above the other. This arrangement makes them hard and heavy. When pressure is applied on the metals, the upper rows of atoms slip pass the lower rows. As a result, their shapes are changed. Metals can, therefore, be easily drawn into wires and sheets. Fig (3.11)

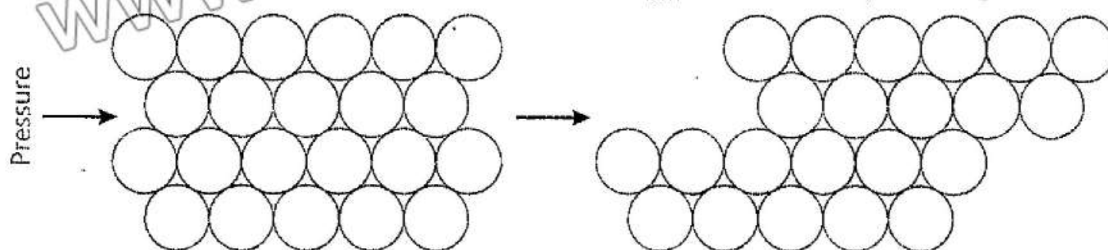


Fig (3.11)

Exercise

What type of atoms form metallic bond?
Give a comparison of metallic bond with an ionic bond.

Interesting Information!

Metals are extensively used in many industries. They are used in machinery, automobiles, railways, air crafts, rockets, in construction industry, in electronics industry, in jewellery, in electric wires and many more.

3.4 Electropositive Character of Metals

Metals generally have a tendency to lose electrons to form positive ions called cations. This property is called the electropositive character of metals. This property is also related to the reactivity of the metals. Metals which lose electron or electrons easily are considered more reactive. For example, alkali metals (Na, K) are highly electropositive elements and thus they undergo reactions very easily. Sodium and potassium react vigorously with water and halogens to give their respective hydroxides and halides. They also react with acids to give salts and water.

Alkaline earth metals (Mg, Ca), on the other hand, lose their outer electrons less easily and thus they are less electropositive than alkali metals. Their reactions towards water and halogens are also less vigorous.

Aluminum is also highly electropositive metal. It reacts readily with mineral acids to form salts and water.

3.5 Electronegative character of Non-metals

Non-metals have an affinity towards electrons. They tend to gain electrons and become negatively charged ions called anions. They are therefore, named as electronegative elements. Fluorine is the most electronegative element in the periodic table followed by oxygen, nitrogen and chlorine. Non-metals readily react with metals forming ionic bonds. Non-metals also combine with other non-metals to form a wide variety of molecular substances.

3.6 Compare the properties of ionic and covalent compounds.

1. In ionic compounds oppositely charged ions are properly arranged to give a crystalline structure. As a whole the compound is neutral. There exists a strong electrostatic force between their ions.
2. Ionic compounds are usually solids having high melting and boiling points. The melting point of sodium chloride is 800°C because it is difficult to break the strong electrostatic forces of attraction between the oppositely charged ions.
3. Ionic compounds are generally soluble in polar solvent like water.
4. They are usually good conductor of electricity in molten state or in solution form. Their conductance is due to the presence of free ions.

1. Covalent compounds mostly exist as discrete neutral molecules. There exists a strong electrostatic attraction between the two nuclei and the shared electrons.
2. Covalent compounds are made of two or more non-metals. Lower molecular mass covalent compounds are gases or low boiling liquids. High molecular mass covalent compounds exist as solids. Generally, they have lower melting and boiling points.
3. They are usually insoluble in water but soluble in non-polar solvents like ether, benzene and acetone.
4. They are usually bad conductor of electricity.

3.7 Intermolecular Forces of Attraction

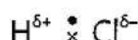
The forces of attraction which are present between the molecules of elements and compounds are named as intermolecular forces of attraction. There attractive forces are generally very weak as compared to the bonding forces present between the atoms of substances. Among the three states of matter, these forces are the weakest among the molecules of the gases and the strongest among the molecules of solids.

The intermolecular forces of attraction are of many type: some are weak and other are relatively strong. They affect the physical properties of the substances. The melting and boiling points of substances depend on the strength of these forces. The stronger the forces among the molecules of a liquid the higher is its boiling point and vice versa. Similarly stronger the intermolecular forces the higher will be the melting point of a solid.

We shall explain here two type of such forces.

1. Dipole – Dipole Forces of Attraction

These attractive forces are present between the molecules of a polar compound like HCl. Hydrogen and chlorine attract the shared pair of electron between them with different force. This force of attraction of an atom is called its electronegativity. Since the electronegativity of chlorine is greater than that of hydrogen it attracts the shared pair of electron with greater force. As a result the bond between hydrogen and chlorine becomes polar as shown in the following



$H^{\delta+} \overset{\bullet}{\times} Cl^{\delta-}$

Due to these partial charges the molecules of HCl start attracting each other. These forces of attraction are called dipole-dipole forces. (Fig 3.12)

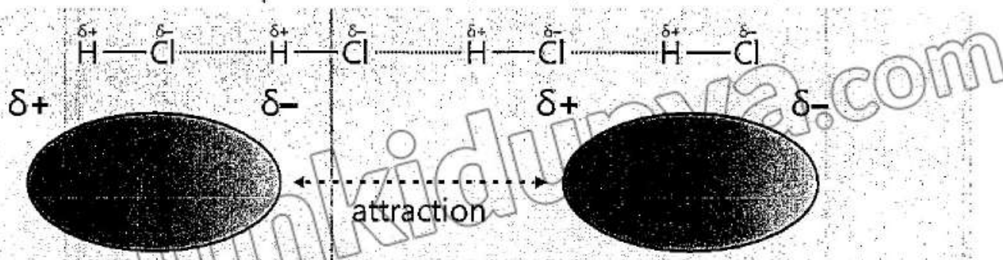


Fig (3.12): Dipole-Dipple Attraction

The compounds which have this type of attractive forces will show relatively higher melting and boiling points.

2. Hydrogen Bonding

Hydrogen bonding is a special case of dipole-dipole attractive forces. When hydrogen is covalent bonded to highly electronegative elements like F, O or N then the large difference of electronegativity values will make the covalent bond highly polar. As a result strong dipole-dipole attractions are observed among the molecules. For example, in H_2O, The O—H bonds are highly polar. Due to this strong attractive forces are developed between water molecules as shown in the Fig (3.13).

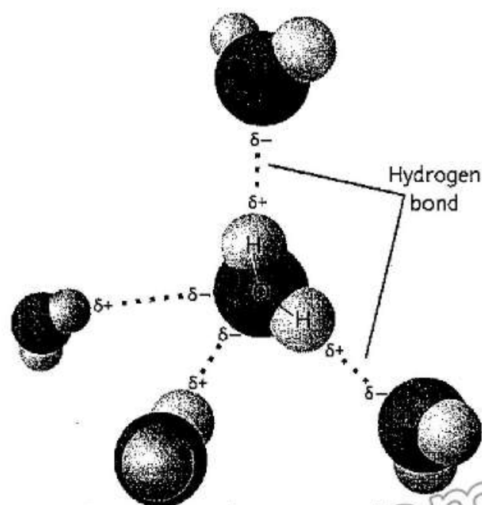


Fig (3.13): Hydrogen bond

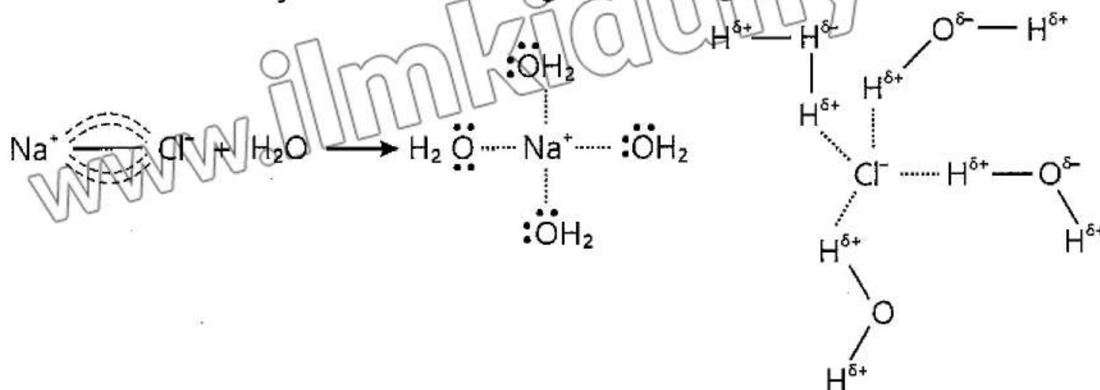
This attractive force present between the molecules of water is called **Hydrogen Bonding**.

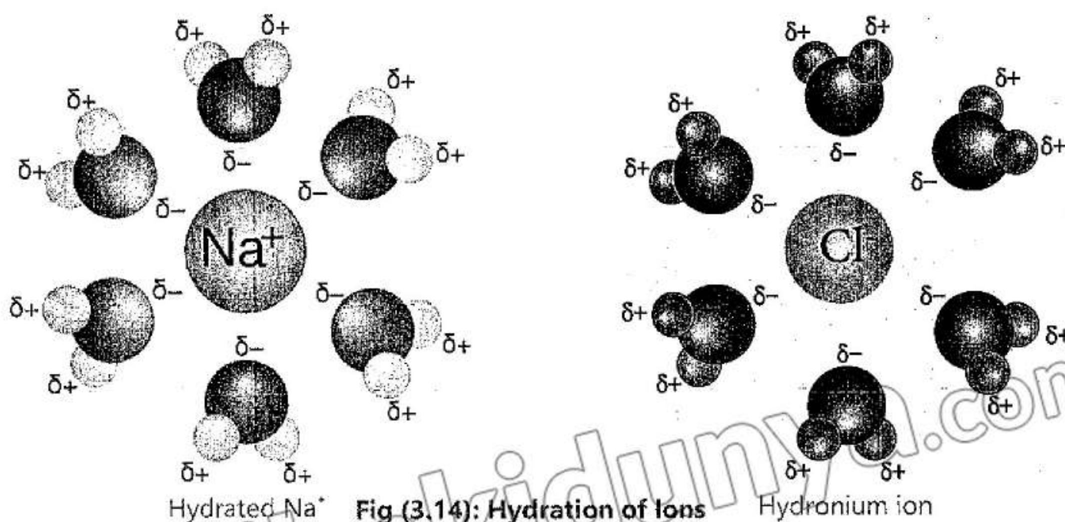
The strength of the hydrogen bonds causes water to have relatively higher melting and boiling points as compared to compounds like H_2S and NH_3.

3.8 Nature of Bonding and Properties

In ionic compounds, the oppositely charged ions are held together by the strong electrostatic force of attraction in the form of a crystal lattice. Since the ions are rigid in ionic compounds, such compounds therefore exist in the form of very stable solids with significantly high melting points. Since ions are spherical and oppositely charged they can surround each other from all the sides, ionic bonds are non-directional. This arrangement of ions is called crystal lattice.

If an external force is applied on the crystal lattice, it breaks easily. It shows that ionic solids are highly brittle. In the solid form, ionic compounds do not conduct electricity because ions are tightly held and cannot move. However, in the molten state, the ions get free and start conducting electricity. Ionic solids are also generally soluble in water. Water not only breaks the electrostatic force of attraction but also hydrates the resulting free ions. Fig (3.14)





Hydrated Na⁺ **Fig (3.14): Hydration of Ions** Hydronium ion

Ionic compounds in an aqueous solution also conduct electricity because the free ions can now move towards their respective electrodes. Ionic compounds generally react in an aqueous solution. When we mix two solutions of ionic compounds, the positive ions of one compound may react with the negative ions of the other to form a new compound.

$$NaCl_{(3)} + AgNO_{3\,(s)} \xrightarrow{H_2O} Na^{+}_{(aq)} + Cl^{-}_{(aq)} + Ag^{+}_{(aq)} + NO^{-}_{3(aq)}$$

$$\downarrow$$

$$AgCl_{(s)} + Na^{+}_{(aq)} + NO^{-}_{3(aq)}$$

White precipitate of silver chloride comes out of the aqueous solution.

Interesting Information!

Conduction of ionic compounds in molten state and in form of an aqueous solution has been utilized to prepare many important elements and compounds. For example, electrolysis of molten sodium chloride gives us sodium metal and chlorine gas. Similarly electrolysis of aqueous sodium chloride gives sodium hydroxide and chlorine gas.

Diamonds, due to their exceptional hardness, are highly valued in industries. Diamond tipped glass cutters are used to make clean cuts in glass. Diamond-tipped drill bits are used to drill through hard rocks in mining operation.

Graphite is used in pencils, in polishes and to make crucibles. Graphite electrodes are used in battery cells and in electric arc furnaces to produce steel.

ALKALI METALS · METALLOIDS · NOBLE GASES
ALKALINE EARTH METALS · OTHER NONMETALS · LANTHANIDES
TRANSITION METALS · HALOGENS · ACTINIDES
POST-TRANSITION METALS · UNKOWN PROPERIES

10	11	12	13	14	15	16	17	18
								He
			B	C	N	O	F	Ne
			Al	Si	P	S	Cl	Ar
Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Ds	Rg	Cn	Uut	Fl	Uup	Lv	Uus	Uuo
Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	
Cm	Bk	Cf	Es	Fm	Md	No	Lr	

Covalent elements and compounds behave very differently from ionic compounds. Elements present at the right side of the periodic table exist as covalently bonded diatomic molecules, for example nitrogen(N_2), oxygen(O_2), fluorine(F_2) and chlorine(Cl_2). Due to very weak forces of attraction between their molecules, their densities and boiling points are very low. Bromine (Br_2) exists as volatile fuming liquid while elements like carbon, phosphorous and sulphur exist as covalent solids. All these solid elements exist both in amorphous and crystalline forms.

Coal is the amorphous form of carbon whereas diamond and graphite are its crystalline forms. Coal is used as a fuel in electricity generating plants. In diamond, each carbon atom is surrounded by four other carbon atoms linked together by strong covalent bonds. Due to this rigid structure, diamond is the hardest thing on this planet. It is used as a cutting, polishing and drilling tool.

Graphite consists of a layered structure, made of hexagonal rings of carbon. Since layers are not bonded strongly, they can slip past each other. Graphite is thus used as a lubricant in industry. Further, these layers in graphite have mobile electrons in between them. Graphite is a good conducter of electricity and it is also used as an electrode.

Binary covalent compounds generally exist as low temperature boiling gases except water. Methane (CH_4), ammonia (NH_3), hydrogen sulphide (H_2S), hydrogen chloride (HCl), nitrogen dioxide (NO_2), carbon dioxide (CO_2) and sulphur dioxide are all covalent compounds which are gases at room temperature.

Water and hydrogen fluoride, on the other hand, are liquids at room temperature. Liquid water has a high boiling point because strong intermolecular forces are present between its molecules. Covalent molecules like hydrogen chloride, sulphuric acid and nitric acid ionize completely in water behaving as very strong acids.

Key Points

1. Atoms form bonds with other atoms to stabilize themselves by obeying duplet and octet rules.
2. The force of attraction which keeps the atoms together is called a chemical bond.
3. Bond which is formed by the transference of one or more electrons is called ionic bond.
4. A covalent bond is formed by the mutual sharing of electrons between atoms. A covalent bond may be single, double or triple.
5. When an electron pair is shared by one atom only, it is called a cordinate covalent bond.
6. Ionic solids are crystalline compounds with high melting and boiling points. They are generally soluble in an aqueous solution.
7. Lower molecular mass covalent compounds are gases or low boiling liquids. Higher molecular mass covalent compounds exist as solids. They are bad conductors of electricity and are soluble in organic solvents.
8. Properties of ionic and covalent compounds are adequately explained on the basis of the type of attractive forces present between them.

Exercise

1. Tick (✓) the correct answer.

i. When molten copper and molten zinc are mixed together, they give rise to a new substance called brass. Predict what type of bond is formed between copper and zinc.

(a) Coordinate covalent bond
(b) Ionic bond
(c) Metallic bond
(d) Covalent bond

(ii) Which element is capable of forming all the three types of bonds; covalent, coordinate covalent or ionic?

(a) Carbon
(b) Oxygen
(c) Magnesium
(d) Silicon

(iii) Why is H_2O a liquid while H_2S is a gas?

(a) Because in water, the atomic size of oxygen is smaller than that of sulphur
(b) Because water is a polar compound and there exists strong forces of attraction between its molecules
(c) Because H_2O molecule is lighter than H_2S
(d) Because water can easily freeze into ice

(iv) Which of the following bonds is expected to be the weakest?

(a) C−C (b) Cl−Cl

(c) O−O (d) F−F

(v) Which form of carbon is used as a lubricant?

(a) Coal (b) Diamond

(c) Graphite (d) Charcoal

(vi) Keeping in view the intermolecular forces of attraction, indicate which compound has the highest boiling point.

(a) H_2O (b) H_2S

(c) HF (d) NH_3

(vii) Which metal has the lowest melting point?

(a) Li (b) Na

(c) K (d) Rb

(viii) Which ionic compound has the highest melting point?

(a) NaCl (b) KCl

(c) LiCl (d) RbCl

(ix) Which compound contains both covalent and ionic bonds?

(a) $MgCl_2$ (b) NH_4Cl

(c) CaO (d) PCl_3

(x) Which among of the following has a double covalent bond?

(a) Ethane (b) Methane

(c) Ethylene (d) Acetylene

2. Questions for Short Answers

i. What type of elements lose their outer electron easily and what type of elements gain electron easily?

ii. Why does lower molecular mass covalent compound exist as gases or low boiling liquids.

iii. Give one example of an element which exists as a crystalline solid and it has covalent bonds in its atoms.

iv. Which property of metals makes them malleable and ductile?

v. Is cordinate covalent bond a strong bond?

vi. Write down dot and cross formula of HNO_3.

3. Constructed Response Questions

i. Why HF is a liquid while HCl is a gas?

ii. Why covalent compounds are generally not soluble in water?

iii. How do metals conduct heat?

iv. How many oxides does nitrogen form. Write down the formulae of oxides?

v. What will happen if NaBr is treated with $AgNO_3$ in water?

vi. Why does iodine exist as a solid while Cl_2 exist as a gas?

4. Descriptive Questions

i. Explain the formation of an ionic bond and a covalent bond.

ii. How do ions arange themselve to form NaCl crystal.

iii. Explain the properties of metals keeping in view the nature of metellic bond.

iv. Compare the properties of ionic and covalent compounds.

v. How will you explain the electrical conductivity of graphite crystals?

vi. Why are metals usually hard and heavy?

5. Investigative Questions

i. The formula of $AlCl_3$ in vapour phase is Al_2Cl_6 which means it exists as a dimer. Explain the bonding between its two molecules?

ii. Explain the structure of sand (SiO_2).