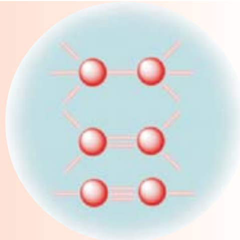
UNIT 3



Chemical Bonding and Intermolecular Forces

Unit Outcomes

After completing this unit, you will be able to:

- *discuss the formation of ionic, covalent, and metallic bonds;*
- *know the general properties of substances containing ionic, covalent, and metallic bonds;*
- *develop the skills of drawing the electron-dot (Lewis structures) for simple ionic and covalent compounds;*
- *understand the origin of polarity within molecules;*
- *appreciate the importance of intermolecular forces in plant and animal life;*
- *understand the formation and nature of intermolecular forces; and*
- *demonstrate scientific inquiry skills: observing, predicting, making model, communicating, asking questions, measuring, applying concepts, comparing and contrasting, relating cause and effects.*

MAIN CONTENTS

Start-up Activity

Objective of the Activity

Scientists have identified different types of attractive forces between atoms in forming bonds. The strength of the forces relies on the types of bonds. For instance, in covalent bonds, the strength of the bonds depends on whether the bonds are single, double or triple bonds.

In this activity, you will use bundles of sticks to develop your ideas about strength of bonds.

1. Collect the following materials and bring to school:
   - Six sticks of wood of the same length and of the same thickness.
2. Place your sticks on the table in the classroom,

   Your teacher will place your sticks in three groups.

   - A single stick
   - Pairs of sticks
   - Sets of three sticks

   Your teacher will assign three students and will give for each student a single stick, a pair of sticks, and a set of three sticks.

Figure 3.1 A bundle of sticks.

Figure 3.2 Students trying to break the sticks.

Analysis

1. Which of your bundles of sticks was the strongest? What is the reason for the different strengths of the bundle of sticks?
2. Draw your conclusions and present to the rest of the class.

3.1 CHEMICAL BONDING

Competencies

By the end of this section, you will be able to:

- define chemical bonding;
- explain why atoms form chemical bonds.

Activity 3.1

Most of the elements are not found free in nature. Why do the elements exists in combined form ? Discuss in group and present your findings to the class.

A chemical bond is the attractive force that binds atoms together in a molecule, or a crystal lattice.

After the periodic table and the concept of electron configuration were developed, scientists began to develop ideas about molecules and compounds. In 1916, G.N Lewis concluded that atoms combine in order to achieve a more stable electron configuration resulting in molecules or compounds.

Historical Note

G.N. Lewis

Gilbert Newton Lewis (1875 - 1946), introduced the theory of the shared-electron-pair chemical-bond formation in a paper published in the journal of the American chemical society. In honor of this contribution, we sometimes refer to "electron-dot" structures as "Lewis structures".

As independent particles, atoms are at relatively high potential energy. Nature, favors arrangements in which potential energy is minimized. Most individual atoms exist in a less stable state than in their combined form.

When atoms form bonds with each other, they attain lower potential energy states. This decrease in atomic energy generally results in a more stable arrangement of matter. When atoms interact to form a chemical bond, only their outer regions are in contact. In the process of the interaction, atoms achieve stable outermost shell configuration. For this reason, when we study chemical bonding, we are concerned primarily with the valence electrons. As described in Unit 2, valence electrons are electrons that exist in the outermost shell of an atom.

Activity 3.2

Form a group and discuss the following:

The following table shows elements of group VIIIA and their atomic numbers. In the space provided fill the valence shell electron configuration and number of valence electrons for the elements. In your discussion include why helium is placed in group VIIIA even though it has only 2 valence electrons.

Elements	Atomic Number	Valence-Shell Electron Configuration	Number of Valence Electrons
Helium, He	2		
Neon, Ne	10		
Argon, Ar	18		
Krypton, Kr	36		
Xenon, Xe	54		
Radon, Rn	86		

Share your discussion with the rest of the class.

Noble gas atoms with eight electrons (except for, He) in the outermost shell are stable. Thus, the ns^2np^6 electronic valence structure has maximum stability. Atoms containing less than eight electrons in their outermost shell are unstable. To attain stability, these atoms tend to have eight electrons in their valence shells. This leads to the explanation of the octet rule. The rule states that atoms tend to gain, lose or share electrons until there are eight electrons in their valence shell.

The type or characteristic of the resulting arrangement depends largely on the type of chemical bonding that exists between the atoms. These are ionic, covalent, and metallic bonds.

Exercise 3.1

1 Why do atoms combine to form compounds?

2 How is a chemical bond formed to make a compound or molecule?

3 Which electron (s) of an atom take (s) part in bond formation?

4 How does the chemical reactivity of halogens compare with that of the noble-gas family?

3.2 IONIC BONDING

Competencies

By the end of this section, you will be able to:

- explain the term ion;
- illustrate the formation of ions by giving examples;
- define ionic bonding;
- describe the formation of an ionic bond;
- give examples of simple ionic compounds;
- draw Lewis structures or electron-dot formulas of simple ionic compounds;
- explain the general properties of ionic compounds; and
- investigate the properties of given samples of ionic compounds.

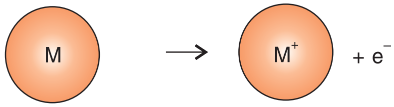
Activity 3.3

Form a group and discuss each of the following concepts. Share your ideas with the class.

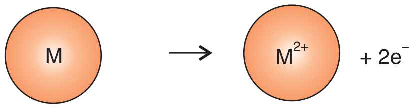
1. Why do some atoms easily lose electrons and others do not?
2. Sodium chloride, NaCl is a good conductor in the form of liquid state, but a non conductor in the form of solid state.

When an atom either loses or gains electrons, it becomes an ion. An ion is an electrically charged particles. Two different types of ions exist. These are the positive ions called cations and the negative ions called anions.

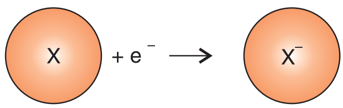
The chemical properties of metals differ from those of non-metals. A metal has 1, 2, or 3 electrons in its outermost shell. Metals tend to lose these electrons and become positively charged ions. For example, if a metal (M) loses one electron, it becomes an ion with a charge of +1.



However, if it loses two electrons it becomes an ion with a charge of +2:



A non-metal may have 4, 5, 6, or 7 electrons in its outermost shell. Non-metals tend to gain electrons to form negatively charged ions. For example, if a non-metal (X) gains one electron, it becomes an ion with a charge of –1.



Note that hydrogen can form both a cation, H^+ (*hydrogen ion*) as in HCl, or an anion H^- (*hydride ion*) as in NaH.

Metals in Group IA, the alkali metals, tend to lose one electron when they combine with other elements, producing cations of +1 charge. For example, Na and K each lose one electron to form ions of +1 charge.

Na	$\rightarrow$	Na^+	+	e^-
sodium atom		sodium ion		
K	$\rightarrow$	K^+	+	e^-
potassium atom		potassium ion		

On the other hand, Group VIIA elements, the halogens, usually gain one electron and produce an ion with –1 charge. For example, each Cl and Br atom accepts one electron to produce an ion with –1 charge.

Cl	+	e^-	$\rightarrow$	Cl^-
chlorine atom				chloride ion
Br	+	e^-	$\rightarrow$	Br^-
bromine atom				bromide ion

Activity 3.4

Form a group and perform the following tasks:

Consider the elements:

1. Calcium (atomic number = 20)
2. Barium (atomic number = 56)
3. Oxygen (atomic number = 8)
4. Sulphur (atomic number = 16)

   a determine whether each of the elements gain or lose electrons in chemical bond formation.

   b write the type of ions they form; and

   c indicate the charges on the ions formed.

Present your findings to the class.

The following table relates the position of some elements in the periodic table to the ions they normally produce.Note that the charge is the same for each ion in a given group or column.

Table 3.1 Selected ions in the periodic table.

Period	IA	IIA	IIIA	IVA	VA	VIA	VIIA
1	H^+						
2	Li^+	Be^{2+}			N^{3-}	O^{2-}	F^-
3	Na^+	Mg^{2+}	Al^{3+}		P^{3-}	S^{2-}	Cl^-
4	K^+	Ca^{2+}					Br^-
5	Rb^+	Sr^{2+}					I^-
6	Cs^+	Ba^{2+}					
7							

Ionic Bond formation

When two atoms combine, one of the atoms gains electrons and becomes an anion and the other loses electrons to form a cation. When a cation and an anion are brought close to one another, an electrostatic force of attraction is set up between them. This force of attraction between oppositely charged ions is called an ionic bond. It is also called an electrovalent bond. The bond is produced when electrons are transferred from the outermost shell of a metal atom to the outermost shell of a nonmetal atom.

To illustrate ionic bonding, let us consider the formation of the bond between sodium and chlorine. A sodium atom has 1 valence electron. In order to attain the electron configuration of the nearest noble gas (Ne), it has to lose its valence electron and form a sodium ion (Na^+). Chlorine has 7 valence electrons. By gaining 1 electron, chlorine attains the electron configuration of argon (Ar) and form a chloride ion (Cl^-).

In general, an ionic bond is formed by the transfer of electron from a metal to a nonmetal- for example, sodium and chlorine. Atoms that are bound together by an ionic bond form ionic compounds. For example, Na^+ and Cl^- ions form sodium chloride, NaCl. The transfer of an electron from sodium to chlorine and the formation of the ionic bond in sodium chloride is shown with the following shell diagrams.

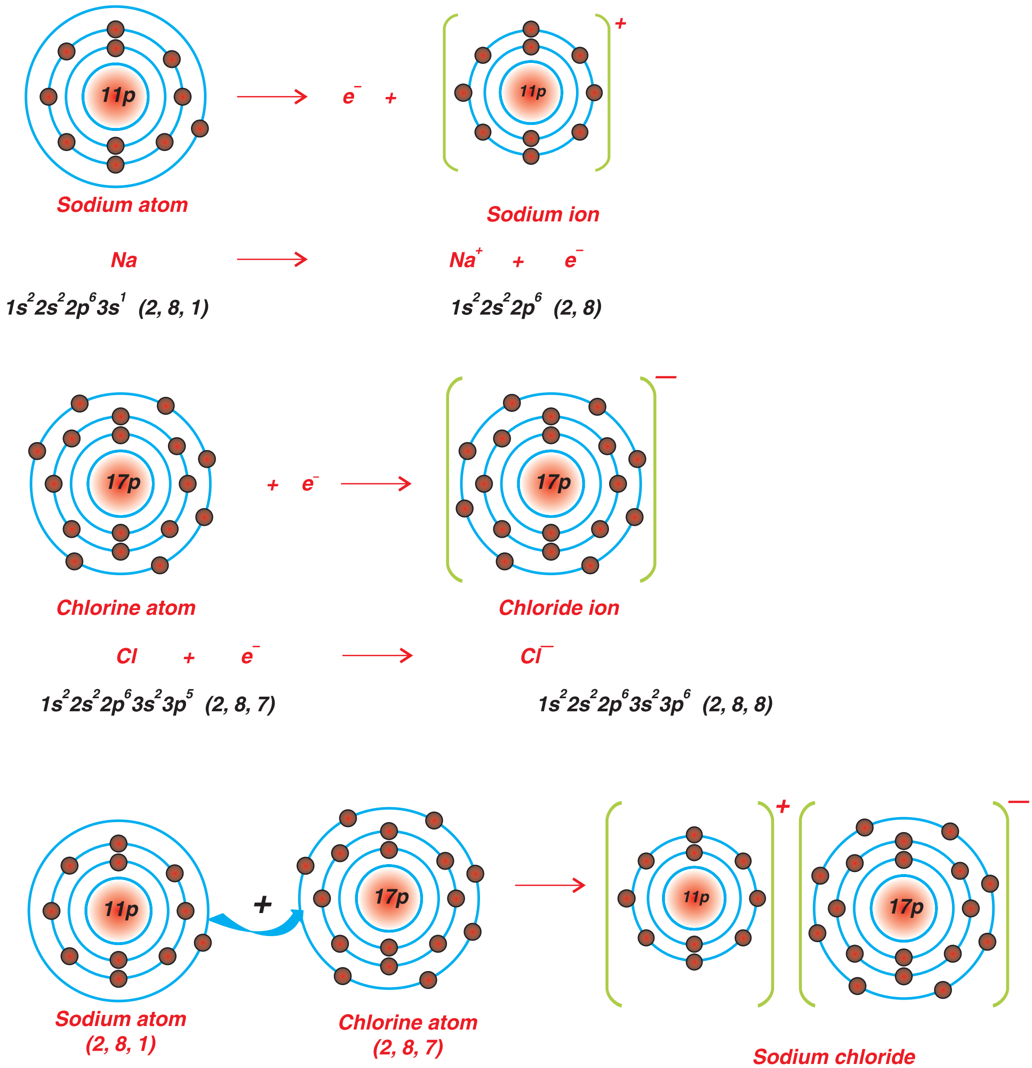


Figure 3.3 Formation of sodium chloride.

Electron-dot notation is often used to represent the outermost shell electron configurations of the elements. These formulas, also called Lewis formulas, consist of the symbol of the element plus dots equal to the number of valence electrons in the atom or ion. Since valence shells contain a maximum of eight electrons, electron-dot symbols contain a maximum of eight dots. Electron-dot formulas of sodium and chlorine are shown below. Sodium is an alkali metal with one valence electron:

Na: $1s^2 2s^2 2p^6 3s^1$

The Lewis symbol for sodium is Na·.

Chlorine is a halogen with seven valence electrons:

Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$

The Lewis symbol for chlorine is :C̈l·.

The formation of ionic bond can also be represented by using electron-dot formulas. Therefore, the Lewis structure for the ionic compound sodium chloride will be

$$\mathrm{Na\cdot} + \mathrm{\cdot\ddot{C}l:} \rightarrow \mathrm{Na^{+}}\left[\mathrm{:\ddot{C}l:}\right]^{-}$$

Activity 3.5

Form a group and perform each of the following task:

Draw Bohr's diagrammatic representation and write the Lewis formula for the following ionic compounds:

a Potassium chloride
b Magnesium oxide
c Calcium chloride
d Potassium sulphide
e Aluminium oxide

Share your ideas with rest of the class.

The following table illustrates the formation of ionic bonds between representative metals and non-metals. Careful observation indicates that, in each case, both the metal and the nonmetal acquire a noble-gas configuration. The compounds formed in each case are electrically neutral as the sum of positive charges equals the sum of negative charges.

Table 3.2 Summary of formula of ionic-compounds.

Metal group	Non-metal group	Formula of Ionic Compound	Examples
IA	VIIA	$MX\ (M^+X^-)$	$NaCl$, KBr
IA	VIA	$M_2X\ (2M^+X^{2-})$	Li_2O, K_2O
IA	VA	$M_3X\ (3M^+X^{3-})$,	Na_3N, K_3P
IIA	VIIA	$MX_2\ (M^{2+}2X^-)$,	$MgCl_2$, CaI_2
IIA	VIA	$MX\ (M^{2+}\ X^{2-})$,	BaS, MgO, MgS
IIA	VA	$M_3X_2\ (3M^{2+}\ 2X^{3-})$	Ca_3N_2, Mg_3P_2
IIIA	VIIA	$MX_3\ (M^{3+}\ 3X^-)$,	AlF_3
IIIA	VIA	$M_2X_3\ (2M^{3+}3X^{2-})$	Al_2O_3
IIIA	VA	$MX\ (M^{3+}\ X^{3-})$	AlN

Note: M = metal; X = non-metal.

Exercise 3.2

1. Draw the Lewis structure for the nitrogen atom, nitrogen molecule and ammonia.
2. Show that the following species have the same number of electrons.
   Na^+, Mg^{2+}, O^{2-}, and Ne

How do you name ionic compounds, for example NaCl?

To name an ionic compound that is formed from a metal and a non-metal, follow the given procedure, by considering the example of NaCl:

1. Write the name of the metal (sodium).
2. Modify the last characters of the name of the non-metal (chlorine) to end it with ide. (Chloride).

Therefore, the name of NaCl is sodium chloride. Similarly, MgO is named as magnesium oxide, Na_3N is named as sodium nitride.

General Properties of Ionic Compounds

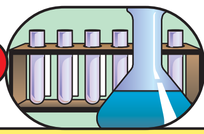
Activity 3.6

Form a group and perform the following task: Collect samples of ionic compounds from your school laboratory and investigate whether the samples are:

a hard or soft b brittle or strong c liquids or solids

What is your generalization about the physical properties of ionic compounds? Share your ideas with the rest of the class.

Experiment 3.1



Investigating the Physical Properties of Ionic Compounds

I Melting Point and Solubility

Objective: To investigate the melting point and solubility of some ionic compounds.

Apparatus: Test -tube, Bunsen burner.

Chemicals: NaCl, $CuCl_2$, ethanol, hexane and benzene.

Procedure:

Perform the following three experiments:

1. Put a few crystals of dry sodium chloride (NaCl) and copper (II) chloride ($CuCl_2$) into separate glass test tubes. Heat strongly on a Bunsen burner. What do you observe?

2. Place about 1 g each of sodium chloride ($NaCl$) and copper (II) chloride ($CuCl_2$) in separate test tubes. Add about 5 mL of water (polar solvent) and shake well.
3. Repeat experiment 2 using the following solvents instead of water. Ethanol (polar solvent), hexane and benzene (non-polar solvents). These solvents are highly flammable and should be kept away from flames.

Observations and analysis

a In experiment 1, do the crystals melt? Do they have high or low melting points?

b In experiment 2 and 3, does each of the solids dissolve in the given solvents? Why? Do they have low or high solubility in the given solvents?

Prepare a table as shown below and fill in the results of the solubility tests.

Substances	Water	Ethanol	Hexane	Benzene
$NaCl$ (s)				
$CuCl_2$ (s)				

II Conductivity

Objective: To investigate the conductivity of some ionic compounds.

Apparatus: Beaker, conductivity cell, graphite or iron rods, bulb, wire, battery.

Chemicals: Sodium chloride, copper (II) chloride, benzene and charcoal.

Procedure:

1. Put some amount of sodium chloride crystals in a beaker.
2. Place two electrodes that are made of iron or graphite in the beaker.
3. Connect the two electrodes to a bulb and a 6-volt battery as shown in Figure 3.5.

   Record your observation.
4. Now add distilled water to the beaker and stir dissolve the salt. Observe the changes.
5. Repeat the experiment using aqueous copper (II) chloride, benzene and charcoal.

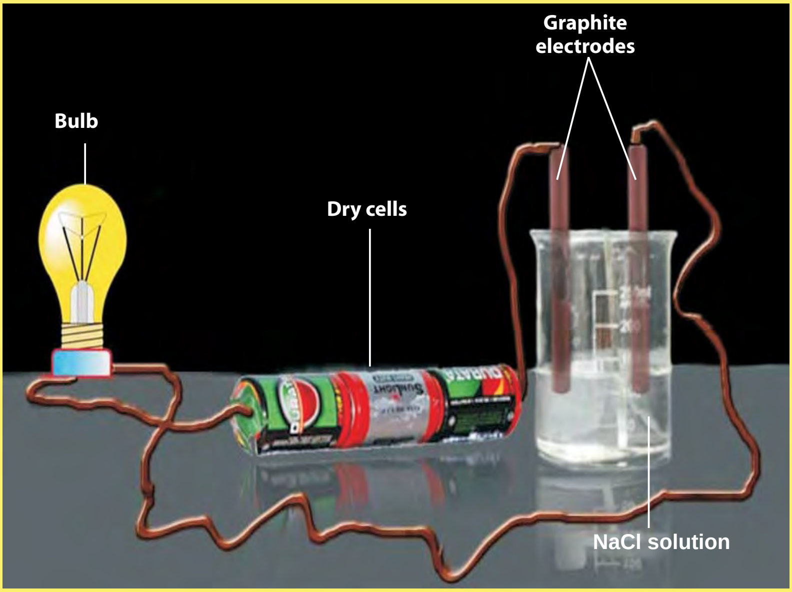


Figure 3.5 Conductivity of sodium chloride solution.

Observations and analysis

a Does sodium chloride conduct electricity in its solid state?

b What do you observe in the bulb when the sodium chloride is dissolved? What does this indicate?

c What can you conclude from this experiment?

d What are your observations for aqueous copper (II) chloride, benzene and charcoal?

Conclusion:

What are your conclusions about:

a melting point,

b solubility in polar and non-polar solvents, and

c electrical conductivity of ionic compounds?

A summary of the general properties of ionic compounds:

1. Ionic compounds do not contain molecules. They are aggregates of positive ions and negative ions. In the solid state, each ion is surrounded by ions of the opposite charge, producing an orderly array of ions called crystal.

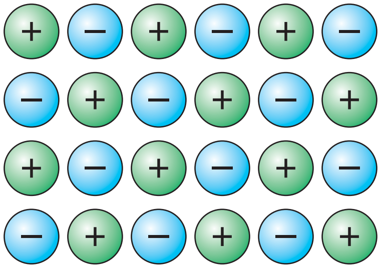


Figure 3.4 Arrangement of ions.

2. At room temperature ionic compounds are hard and rigid crystalline solids. This is due to the existence of strong electrostatic forces of attraction between the ions.
3. Ionic compounds have relatively high melting and boiling points. This is due to the presence of strong electrostatic forces between the ions. These forces can be overcome only by applying very large amounts of energy.
4. Ionic compounds can conduct electric currents when molten or in aqueous solution. This is due to the presence of mobile ions in molten state or in solution. However, ionic compounds do not conduct electricity in the solid state.
5. Ionic compounds are soluble in polar solvents such as water. They are insoluble in non-polar solvents such as benzene.

Exercise 3.3

1. KCl is soluble in water but insoluble in benzene. Explain.
2. Which of the following substances conduct electricity? Give reasons for your answer in each case:

   a NaCl (aq) b NaCl (l) c NaCl (s)
3. Name the ionic compounds formed from the following pairs of elements:

   a calcium and sulphur b sodium and iodine c silver and bromine
4. List the properties of ionic compounds.

3.3 COVALENT BONDING

Competencies

By the end of this section, you will be able to:

- define covalent bonding,
- describe formation of a covalent bond,

- draw Lewis structures or electron-dot formulas of simple covalent molecules,
- give examples of different types of covalent molecules,
- make models of covalent molecules to show single, double and triple bonds using sticks and balls or locally available materials,
- explain polarity in covalent molecules,
- distinguish between polar and non-polar covalent molecules,
- define coordinate (dative) covalent bond,
- illustrate the formation of coordinate covalent bond using appropriate examples,
- explain the general properties of covalent compounds, and
- investigate the properties of given samples of covalent compounds.

Activity 3.7

Form a group and discuss each of the following concepts.

1. What is the difference between the bond when two chlorine atoms combine to form a chlorine molecule (Cl_2) and that formed when sodium combines with chlorine to form sodium chloride (NaCl)?
2. Carbon tetrachloride (CCl_4) is a covalent compound. Would you expect it to be :
   i) a conductor of electricity
   ii) soluble in water.

Share your ideas with the class.

Many molecules are formed when outermost shell or valence electrons are shared between two atoms. This sharing of electrons creates a covalent bond.

Covalent bond formation can be illustrated by the sharing of electrons between two hydrogen atoms to form a molecule of hydrogen.

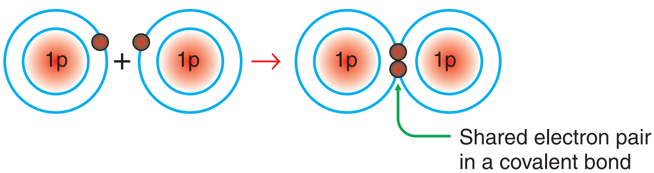


Figure 3.6 Sharing of electrons between hydrogen atoms in H_2 molecule.

In the hydrogen molecule, each hydrogen atom attains the stable electron configuration of helium.

In a covalent bond, each electron in a shared pair is attracted to the nuclei of both atoms as shown in Figure 3.6. The shared electrons spend most of their time between the two nuclei. The electrostatic attraction between the two positively charged nuclei and the two negatively charged electrons hold the atoms in the molecule together. This attractive force between positively charged nuclei of atoms and the shared electrons in a molecule is known as covalent bond.

A molecule of hydrogen chloride is also formed by a pair of electrons shared between the two atoms. Each atom in the molecule attains a stable electron configuration.

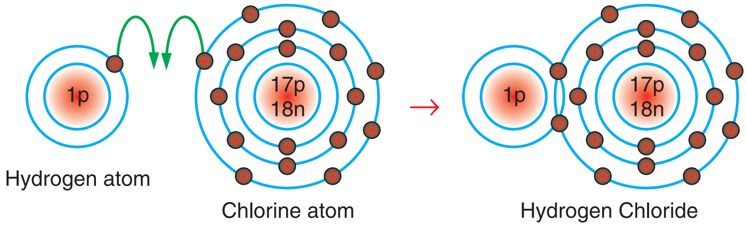


Figure 3.7 Hydrogen and chlorine share a pair of electrons in HCl.

The concept of Lewis formula representation can also be extended to covalent bonds. The Lewis structure for a covalent compound shows the arrangement of atoms in a molecule and all the valence electrons for the atoms involved in the compound. It is conventional to represent the non-bonding (lone pair) electrons by dots and the pair of electrons that are shared between atoms by a dash. For example, consider the hydrogen molecule:

The electron-dot formula of the hydrogen molecule is:

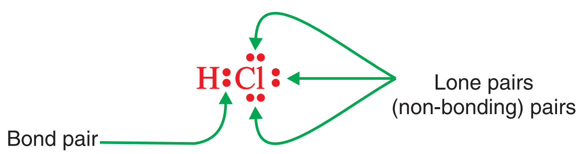
$$H\bullet + \bullet H \rightarrow H{:}H$$

The covalent bond in hydrogen molecule is also written as H – H.

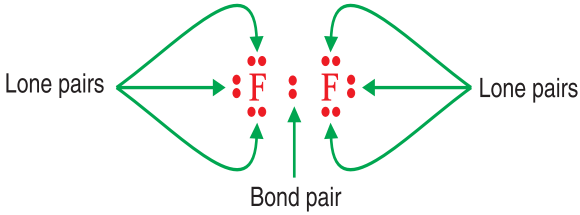
The formation of the covalent bond in hydrogen chloride is shown by the following electron-dot formula. This formula must satisfy the octet rule (for chlorine) and the doublet rule (for hydrogen). As shown in the illustration, these requirements are satisfied. The shared pair belongs to both of the atoms (hydrogen and chlorine) in the hydrogen chloride molecule. The resulting valence electron configuration provides two valence electrons to hydrogen and eight to chlorine.

$$H\bullet + \bullet\underset{\bullet\bullet}{\overset{\bullet\bullet}{Cl}}{:} \rightarrow H{:}\underset{\bullet\bullet}{\overset{\bullet\bullet}{Cl}}{:}$$

The chlorine atom in the molecule has three pairs of electrons, which are not used for bonding. Pairs of electrons that is not used for bonding are called lone-pair electrons. Pairs that are used for bonding are called bonding-pair electrons.

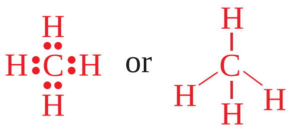


Consider the fluorine molecule, F_2. The electron configuration of fluorine is 2, 7. Thus each fluorine atom has seven valence electrons. Accordingly, there is only one unpaired electron on fluorine. Therefore, the formation of the fluorine molecule is represented as

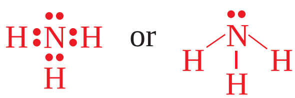


Note that only two valence electrons participate in the formation of fluorine molecule. The others are non-bonding electron (lone pairs). Thus each fluorine atom in fluorine molecule has three lone-pairs of electrons. The resulting molecule is a diatomic molecule. A diatomic molecule consists of two atoms. All the other members of the halogen family form diatomic molecules in the same way as fluorine does.

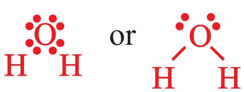
The maximum number of covalent bonds that an atom can form can be predicted from the number of electrons needed to fill its valence shell. For example, each member of Group IVA elements has four electrons in its valence shell, and it needs four more electrons to achieve stable noble-gas electron configuration. Thus, it forms four covalent bonds for carbon in methane, CH_4 as shown below:



Elements of Group VA need three additional electrons to achieve noble gas configuration and they form three covalent bonds as shown below for nitrogen in ammonia NH_3.



Similarly, elements of group VIA form two covalent bonds and Group VIIA elements form single covalent bonds.



Water

Hydrogen chloride

Types of Covalent Bonds

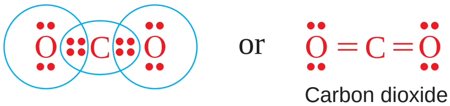
How do you compare the nature and strength of the bonds in H_2, O_2 and N_2?

Atoms can form different types of covalent bonds. These are single bonds, double bonds and triple bonds.

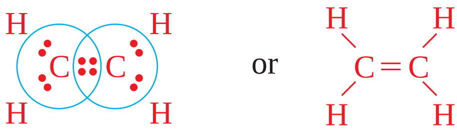
In a single bond two atoms are held together by one electron pair.

How are the covalent bonds in H_2, Cl_2 and HCl formed?

Many covalent compounds are held together by multiple bonds. Multiple bonds are formed when two or three electron pairs are shared by two atoms. If two atoms share two pairs of electrons, the covalent bond is called a double bond. For example, double bonds are found in molecules of carbon dioxide (CO_2) and ethene (C_2H_4).



Carbon dioxide



Ethene

A triple bond is formed when two atoms share three pairs of electrons, as in the nitrogen molecule (N_2).

Nitrogen molecule

The ethyne (acetylene) molecule (C_2H_2) also contains a triple bond. In this case the bond is between two carbon atoms.

Ethyne

Activity 3.8

Form a group and perform the following tasks:

1. Construct a molecular model for each of the following species using locally available materials such as toothpicks to represent bonds, and styrofoam spheres to represent atoms.

   a N_2 b H_2 c O_2 d C_2H_4

2. Based on your models, identify which species contain:
   - *i)* only single bonds
   - *ii)* double bonds
   - *iii)* triple bonds

Present your models and findings to the rest of the class.

Exercise 3.4

1. How many bonding pair and lone pair electrons are found in each of the following molecules?

   a CO_2 b C_2H_4 c N_2 d C_2H_2

2. Consider molecules of carbon disulfide, CS_2, and hydrogen cyanide, HCN.

   a What types of bonds do they contain?

   b Draw their electron-dot formulas.

   c Are there any lone-pair electrons in these molecules?

3. Why is the melting point of ionic compounds higher than that of covalent compounds?

Reading Check

Does the hydrogen atom form covalent as well as ionic bonds? How?

3.3.1 Polarity in Covalent Molecules

Activity 3.9

Form a group and discuss the following idea:

The covalent bonds are formed by sharing of electrons. Compare covalent bonds formed between atoms of the same elements and those formed between atoms of different elements. (Example: H_2 and HCl).

Present your conclusion to the class.

A covalent bond is formed when electron pairs are shared between two atoms. In molecules like H_2, in which the atoms are identical, the electrons are shared equally between the atoms. A covalent bond in which the electrons are shared equally between the two atoms is called a non-polar covalent bond.

$$H - H$$

In other words, a non-polar bond is a covalent bond in which bonding electrons are shared equally between identical atoms, resulting in a balanced distribution of electrical charge.

In contrast, in the covalently bonded HCl molecule, the H and Cl atoms are of different elements; therefore, they do not share the bonding electrons equally.

$$H - \underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}\colon$$

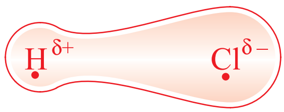
A chemical bond in which shared electrons spend more time in the vicinity of one atom than the other is called a polar covalent bond, or simply a polar bond.

Polarity of bonds is caused by differences in the electronegativity of the two atoms forming the bonds. Electronegativity is the ability of an atom to attract the shared electrons in a chemical bond toward itself.

Elements with high electronegativity have a higher tendency to attract electrons than elements with low electronegativity. For example, in the case of HCl, the electronegativity of the chlorine atom is higher than that of the hydrogen atom. The shared pair of electrons is more strongly attracted to the nucleus of the chlorine atom. As a result, the chlorine atom acquires a partial negative charge (δ^-) whereas the hydrogen atom acquires a partial positive charge (δ^+). The delta is read as "partial" or "slightly."

If a molecule has a positive end and a negative end, it is said to be polar and posses a dipole. Dipole means 'two poles'.

Experimental evidence indicates that, in the HCl molecule, the electrons spend more time near the chlorine atom. We can think of this unequal sharing of electrons as a partial electron transfer or a shift in electron density as shown below:



This unequal sharing of the bonding electron pair results in a relatively higher electron density near the chlorine atom and a correspondingly lower electron density near hydrogen.

Exercise 3.5

1. How many electrons are shared in a:
   a single bond, b double bond, and c triple covalent bond?
2. Draw Lewis structures for:
   a H_2 b Cl_2 c C_3H_6
3. Draw Lewis structures for each of the following molecules:
   a HBr b CO_2 c H_2O
   Also indicate the partial charges using δ^+ and δ^-.

3.3.2 COORDINATE COVALENT BOND

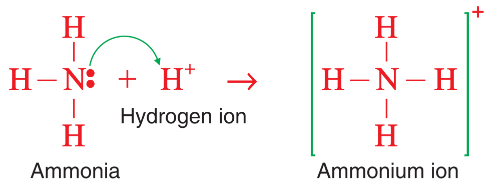
A covalent bond in which one atom donates both electrons of the bond is called a coordinate covalent bond. It is also called a dative bond. Such a bond is hypothetically represented as:

$$A\!: + B \rightarrow A\!:\!B$$

Once formed, a coordinate covalent bond has the same properties as any other covalent bond. The atom that contributes both electrons for the bond is the donor atom, and the atom that shares the electron pair is the acceptor atom.

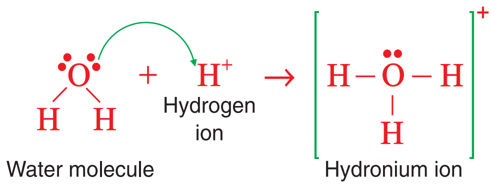
For an atom to act as a donor, it must contain lone pair of electrons in its valence shell and the acceptor atom must have at least one vacant orbital.

For example, the ammonium ion, NH_4^+, is formed by a coordinate covalent bond in which the two non-bonding electrons on NH_3 bond with a hydrogen ion, H^+, which has no electrons to contribute.



In the resulting ion, NH_4^+, the four N – H bonds are identical.

Similarly, a coordinate covalent bond can be formed between a hydrogen ion and a molecule of water, which has two lone pairs of electrons.



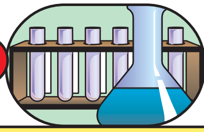
Carbon monoxide, CO, also has a coordinate covalent bond. In order for both carbon and oxygen atoms to attain noble-gas electron arrangements, oxygen donates a pair of electrons to the carbon atom. In the process a coordinate covalent bond is formed between the two atoms.

$$:\!C\!: + :\!\ddot{O}\!: \rightarrow :\!C \overset{\leftarrow}{=} \ddot{O}\!:$$

General properties of covalent compounds

1. Covalent compounds are generally liquids or gases at ordinary temperature. For example : water and ethyl alcohol are liquids. Hydrogen chloride, methane and cabon dioxide are gases. Same covalent compounds are solids (*e.g.* sugar)
2. As compared to ionic compounds, covalent compounds have relatively lower melting points and boiling points.
3. They do not conduct electric current when molten or in aqueous solution, because they consist of molecules rather than of ions.
4. Covalent compounds are insoluble in polar solvents such as water. They are soluble in non-polar solvents such as benzene and carbon tetrachloride.

Experiment 3.2



Investigating the Physical Properties of Covalent Compounds

I. Melting point

Objective: To investigate the effect of heat on covalent compounds.

Apparatus: Test tube, Bunsen burner.

Chemicals: Naphthalene, graphite, iodine.

1. Put a small amount of naphthalene into a dry test tube. Heat it strongly. Naphthalene is toxic. Do not inhale it or get it on your skin.

   a What do you observe? Does the solid melt or vaporize?

   b Is the melting point high or low?

2. Repeat the procedure with graphite, and iodine separately. Iodine vapor is toxic. Do not inhale it.

   Observe and record your observation in a tabular form as shown below.

Substance	Melted, vaporized or nothing happened	High or low melting point
Naphthalene Graphite Iodine		

II. Solubility

Objective: To investigate the solubility of covalent compounds.

Apparatus: Test-tubes, test tube rack.

Chemicals: Naphthalene, graphite, iodine, ethanol, hexane and benzene.

3. Arrange 12 test tubes in three sets (A, B, C) of 4 test tube each. To each test tube of set A, add 1 g of naphthalene. To each test tube of set B add 1 g of graphite and to each test tube of set C add 1 g of iodine.

4. Add about 10 mL of each the following solvents to the four test tubes of each set separately and shake well.

   - Water
   - Ethanol
   - Hexane
   - Benzene

Caution: Ethanol, hexane and benzene are all highly flammable.

Observe and record whether the solids are very soluble, slightly soluble or insoluble.

Substance	Solubility			
	Water	Ethanol	Hexane	Benzene
Naphthalene				
Graphite				
Iodine				

Observations and analysis

Draw general conclusions on the

a melting points,

b solubility in polar and non-polar solvents of the covalent compounds given.

Exercise 3.6

1. Which of the following molecules contain a covalent bond?

   a CaO d SO_2 g MgO

   b HCl e Na_2O h NaH

   c CO_2 f PCl_3 i CH_4

2. Which of the following contain a dative bond?

   a H_3O^+ b NH_3 c NH_2^- d CaO

3. Which of the following molecules are polar ?

   a SO_2 c H_2O e BCl_3 g CH_3Cl

   b CO_2 d CS_2 f CH_4

4. Which of the following are non-polar covalent compounds?

   a O_2 c CH_4 e H_2O g Br_2

   b HCl d O_3 f Cl_2

Activity 3.10

Form a group and perform the following task:

Collect samples of covalent compounds from your school laboratory and investigate whether the samples are:

a liquids or solids   b hard or soft   c brittle or strong

What is your generalization about the physical properties of covalent compounds? Share your ideas with the class.

Reading Check

What is the difference between a coordinate covalent bond and covalent bond?

3.4 METALLIC BONDING

Competencies

By the end of this section, you will be able to:

- explain the formation of metallic bond;
- explain the electrical and thermal conductivity of metals in relation to metallic bonding; and

Activity 3.11

Form a group and discuss the following concepts and present your discussion to the class.

a Metals are solids. They contain large number of atoms in their crystals. What kind of force do you think holds these metal atoms together?

b How do you account for the properties of metals, such as conductivity, malleability, and ductility in terms of the bonds in metals?

The highest energy orbitals of most metals are occupied by very few electrons. In *s*-block metals, for example, one or two valence electrons occupy *s* orbitals in the outermost levels (for example Na and Mg). Furthermore, the *p* orbitals of the outer most level are also occupied partially in *p*-block metals (example Tl, Pb and Bi).

The *d*-block metals contain partially filled $(n–1)d$ levels in their atomic states or principal oxidation states. The bonding in metals is different from that in other types of crystals. The valence electrons of metals are not held by individual atoms. Rather, they are delocalized and mobile (free to move throughout the structure).

The valence electrons form a sea of electrons around the metal ions and these metal ions are organized as a crystal. Metallic bonding results from the attraction between the metal ions and the surrounding sea of electrons.

For example, as illustrated in Figure 3.8, a sodium metal crystal is a lattice-like array of Na^+ ions surrounded by a sea of mobile bonding valence electrons.

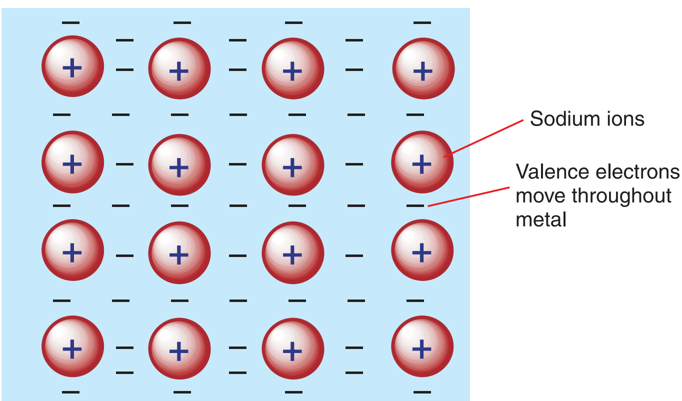


Figure 3.8 Metallic bonding in sodium metal.

The bonding valence electrons move freely throughout the entire crystal. This freedom of movement is responsible for the electrical conductivity of metals.

Properties of Metallic Bonding

Have you ever visited a goldsmith workshop? Why are metals easily shaped into thin sheets and drawn into wires?

The freedom of movement of bonding valence electrons is responsible for the high electrical and thermal conductivity that characterizes the metals. Other properties of metallic bonding contribute to unique properties of metals. For example, most metals are easy to shape due to their malleability and ductility.

Malleability allows a substance such as a metal to be reshaped. By hammering and bending some metals, you can create thin sheets. Ductility allows a substance to be drawn or pulled out into long thin pieces, such as wires.

Metals are malleable and ductile because metallic bonding is the same in all directions throughout the solid.

When we apply a force to metal, its cations swim freely within the sea of electrons without breaking the crystal structure. For example, when you hammer, bend, or pull on a metal to reshape it, you shift its cations around. The force you apply moves the atoms around, for example, around corners in the lattice. This is the basis for malleability and ductility of metals, which allows you to change its shape.

Project work

Model of a Metallic Crystal

Put about one hundred balls (for example, marble balls or balls made from other locally available materials) into a rectangular glass trough. Shake the trough. Allow the balls to settle.

a Draw a two-dimensional diagram to show how the marble balls are now arranged in the trough.

b If the balls represent atoms in a metallic lattice, which species are occupying the 'empty' space between and around them?

Present your model and findings to the class.

Exercise 3.7

1. Describe how a metallic bond is different from those of an ionic bond and a covalent bond.
2. Explain thermal and electrical conductivity in metals.
3. Is metallic bonding responsible to form compounds?

3.5 INTERMOLECULAR FORCES

Competencies

By the end of this section, you will be able to:

- define inter-molecular forces;
- explain hydrogen bonding;
- explain the effects of hydrogen bonding on the properties of substances;
- describe Van der Waal's forces;
- explain dipole-dipole forces;
- give examples of molecules with dipole-dipole forces;
- explain dispersion forces;
- give examples of molecules in which dispersion force is important; and
- compare and contrast the three types of intermolecular forces.

Activity 3.12

Form a group and discuss the following phenomenon: Why do covalent compounds usually exist as gases and liquids?

Share your views with the rest of the class.

Inter-molecular forces are relatively weak forces of attraction that occur between molecules. Inter-molecular forces vary in strength but are generally weaker than the bonds that join atoms in molecules, ions in ionic compounds, and metal atoms in solid metals. Inter-molecular forces acting between molecules include: *dipole-dipole forces*, *London dispersion forces* and *hydrogen bonding*. Dipole-dipole attractions and London forces are collectively called Van der Waal's forces.

A Dipole-Dipole Forces

Dipole-dipole forces are strong inter-molecular forces between polar molecules. A dipole is created by equal but opposite charges separated by a short distance. A polar molecule acts as a tiny dipole because of its uneven charge distribution.

A dipole is represented by an arrow with a head pointing toward the negative pole and crossed tail situated at the positive pole. The dipole created by a hydrogen chloride molecule, which has its negative end at the more electronegative chlorine atom, is as shown below:

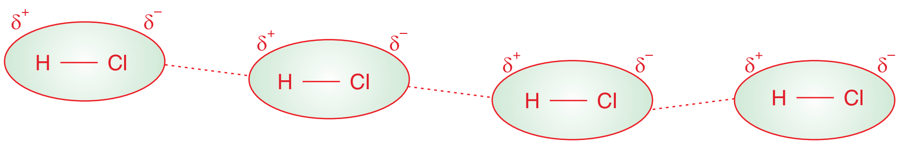


Figure 3.9 Dipole-dipole interactions in HCl molecules.

The negative end in one polar molecule attracts the positive end in an adjacent molecule in a liquid or solid. Dipole-dipole forces occur in molecules such as ethyl alcohol and water.

B London Dispersion Forces

All molecules, including those without dipole moments, exert forces on each other. We know this because all substances, even the noble gases, change from liquid to solid state under different conditions.

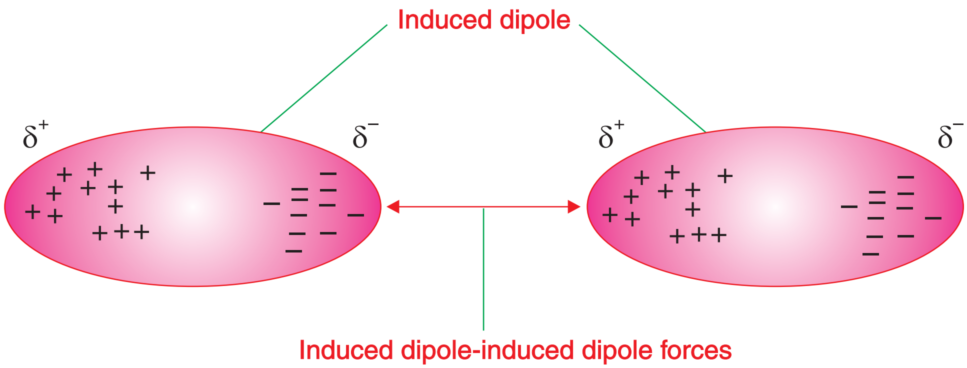


Figure 3.10 Induced dipole-induced dipole forces between non-polar molecules.

London dispersion forces act between all atoms and molecules. They are the only forces that exist between noble gas atoms and non-polar molecules. This fact is reflected in the low boiling points of noble gases and non-polar molecules. Because dispersion forces result from temporary redistribution of the electrons causing induced dipole-dipole interactions, their strength increases with the number of electrons in the interacting atoms or molecules. Hence, dispersion forces increase with atomic number or molar mass. This trend can be seen by comparing the boiling points of gases (helium, He, and argon, Ar), (hydrogen, H_2, and oxygen, O_2), and (chlorine, Cl_2, and bromine, Br_2).

As an illustration, the boiling points of the noble gases are presented in Table 3.4.

Table 3.3 Boiling points of noble gases.

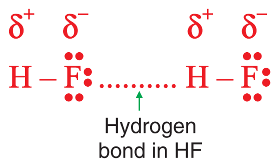
Noble gas	Boiling Point (°C)	Number of electrons
He	-269	2
Ne	-246	10
Ar	-186	18
Kr	-152	36
Xe	-107	54
Rn	-62	86

As you look down the column of noble gases, you note that boiling point increases. This is because the induced dipole-dipole interaction increases.

C Hydrogen Bonding

Hydrogen bonding is a particular type of intermolecular force arising when a hydrogen atom is bonded to highly electronegative elements, fluorine, oxygen and nitrogen. Hydrogen bonding is a particular type of dipole-dipole interactions between polar compounds. In such compounds, large electronegativity differences between the hydrogen and the fluorine, oxygen, or nitrogen atoms make the bonds connecting them highly polar. This polarity gives the hydrogen atom a positive charge. Moreover, the small size of the hydrogen atom allows the atom to come very close to an unshared pair of electrons on an adjacent molecule. Hydrogen bonding is responsible for the unusual high boiling points of some compounds such as hydrogen fluoride (HF), water (H_2O) and ammonia (NH_3).

Hydrogen bonds are usually represented by dotted lines connecting the hydrogen atom to the unshared electron pair of the electronegative atom to which it is attracted. For example, the hydrogen bond in hydrogen fluoride, HF, results when the highly electronegative F atom attracts the H atoms of an adjacent molecule.



Do you think that the intermolecular forces between molecules containing C-H, N-H, and O-H bonds are as strong as the intermolecular forces containing F-H bonds?

Exercise 3.8

1. Which of the following exists predominately in the water (H_2O) molecule?

   a Van der Waal's force
   b hydrogen bond
   c coordinate covalent bond
   d none of these

2. Which of the following has the highest induced dipole interactions in its molecule?

   a He
   b Ar
   c Ne
   d Kr

Critical Thinking

Oxygen ($^{16}_{8}O$) and Sulphur ($^{32}_{16}S$) are in the same group in the periodic table. They form compounds with hydrogen, H_2O and H_2S. However, H_2O is a liquid, whereas H_2S is a gas at room temperature. Give explanation?

Check List

Key terms of the unit

- *Anions*
- *Bonding-pair electrons*
- *Cations*
- *Chemical bond*
- *Conductivity*
- *Coordinate/dative bond*
- *Covalent bond*
- *Delocalized electrons*
- *Dipole*
- *Dipole-dipole interaction*
- *Double-bond*
- *Ductility*
- *Electronegativity*
- *Electrovalent bond*
- *Hydrogen bonding*
- *Inter-molecular forces*
- *Ionic bonding*
- *Ionization energy*
- *London forces*
- *Lone-pair electrons*
- *Malleability*
- *Metallic bond*
- *Mobile electrons*
- *Noble gases*
- *Non-bonding electrons*
- *Octet rule*
- *Polar bond*
- *Polar covalent bond*
- *Polarity*
- *Sea of electrons*
- *Single-bond*
- *Triple-bond*
- *Valence electrons*
- *Van der Waal's forces*

Unit Summary

- A chemical bond is the attractive force that binds atoms together to form a molecule (or a crystal lattice), an ionic or metallic crystal lattice.
- An ionic bond is the electrostatic attraction between oppositely charged ions (cations and anions).
- A covalent bond is formed by a shared pair of electrons.
- A covalent bond in which one pair of electrons is shared is known as a single bond; for example, H_2 written as H – H.
- Atoms can also share more than one pair of electrons to form a multiple bond.
- The sharing of three pairs of electrons forms a triple bond - for example, N_2, written as N≡N.
- A dative or coordinate covalent bond is a bond in which one of the atoms supplies both of the shared electrons to the covalent bond.
- A metallic bond is the electrostatic attraction between positive metal ions and delocalized electrons.

- Inter-molecular forces are forces of attraction between covalently bonded molecules. These include: London forces, dipole-dipole forces, and hydrogen bonding.
- London forces are forces of attractions between nonpolar molecules.
- Dipole-dipole forces are the attractions between dipoles of polar molecules.
- Hydrogen bonding is the attraction of covalently bonded hydrogen to lone pairs on N, O, or F atoms in other molecules or in the same molecule (if the molecule is large enough).

REVIEW EXERCISE ON UNIT 3

Part I: Multiple Choice Type Questions

1. Which of the following contain localized valence electrons?
   a Cu
   b C_2H_6
   c C_6H_6
   d none
2. Which of the following is a compound?
   a Cl_2
   b HCl
   c Na
   d liquid oxygen
3. In the formation of ionic bonding, valence electrons are:
   a shared
   b delocalized
   c transferred
   d not affected
4. What force is responsible for the formation of ice during the freezing of water?
   a ionic
   b covalent
   c dipole-dipole interaction
   d dative bond
5. Which of the following has the highest electrical conductivity in the solid state?
   a water
   b common salt
   c sodium
   d rubber

Part II: Write the missing words in your exercise book

6. Ionic bonding is formed between the atoms of _____ and _____.
7. Covalent bonding is formed between the atoms of _____.
8. Metallic bonding is formed between the atoms of _____.
9. The forces that hold atoms together in molecules of compounds are called _____.

Part III: Short Answer Type Questions

10. Give a simple explanation of the following, using an example:
    - a a covalent bond
    - b an ionic bond
    - c a dative bond
    - d hydrogen bonding
    - e metallic bond
11. Classify the bonds that can be formed between the following pairs of atoms as principally ionic or covalent:
    - a calcium and chlorine
    - b boron and carbon
    - c sodium and bromine
    - d magnesium and nitrogen
    - e sodium and hydrogen
    - f aluminium and oxygen
    - g chlorine and oxygen
    - h iodine and chlorine
12. Give the Lewis structures for the following:
    - a H_2S
    - b CaS
    - c Al_2O_3
    - d HF
    - e N_2
    - f C_2H_4
    - g NH_3
    - h CH_4
    - i CF_4
    - j NO
    - k $CaCl_2$
13. Which of the following molecules can form hydrogen bonding?
    - a H_2S
    - b CO_2
    - c SO_2
    - d CH_4
    - e NH_3
    - f HF
    - g CH_3OH
14. Which of the following substances contain hydrogen bonding?
    - a hydrogen chloride
    - b water
    - c ammonia
    - d methane
15. Classify the following molecules as polar or non-polar:
    - a CH_4
    - b CH_3Cl
    - c C_2H_2
    - d CO_2
    - e H_2O_2
    - f BCl_3
    - g H_2S
    - h HBr
16. What is meant by a polar molecule?
17. Explain the fact that HCl is polar, whereas Cl_2 is a non-polar molecule.
18. What is the difference between intermolecular and intra-molecular attractive forces?