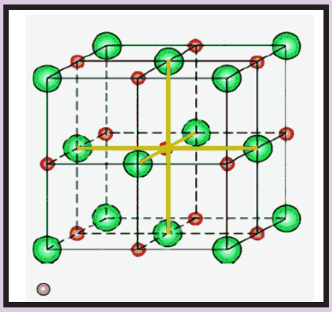
Chapter

10



Chemical Bonding

You have learnt about electron configuration of elements and periodic table in previous lessons. It was learnt that there are over 115 elements.

- How do they usually exist?
- Do they exist as a single atom or as a group of atoms?

In Class IX, we saw that several elements like oxygen, nitrogen and hydrogen exist as diatomic molecules. What is the force that is holding these constituent atoms together in molecules?

- Are there elements which exist as atoms?
- Why do some elements exist as molecules and some as atoms?

In lower classes you have also learnt about the different laws of chemical combinations. Since the formation of chemical compounds takes place as a result of combination of atoms of various elements in different ways, it raises many questions.

- Why do some elements and compounds react vigorously while others are inert?
- Why is the chemical formula for water H_2O and for sodium chloride $NaCl$, why not HO_2 and $NaCl_2$?
- Why do some atoms combine while others do not?

We will try to answer such questions in this chapter.

- Are elements and compounds simply made up of separate atoms individually arranged?
- Is there any attraction between atoms?

Let us take the example of common salt, $NaCl$. When you shake salt from a shaker does it separate into sodium and chlorine? No! This shows that the sodium and chlorine atoms are being held together.

- What is that holding them together?

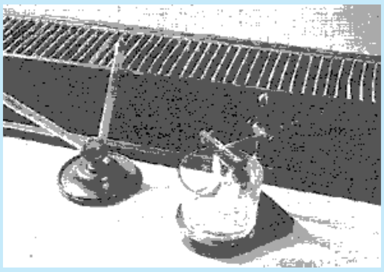
By the late 19th century and early twentieth century, scientists knew about three types of forces – gravitational, magnetic and electrostatic. They also knew about the existence of electrons and protons. So it was believed that electrostatic forces were the cause of attraction between atoms in a molecule. When two atoms come sufficiently close together, the electrons of each atom experience the attractive force of the nucleus of the other atom. But the electrons which are negatively charged and repel each other, and the positively charged nuclei also repel each other. The strength of attraction or repulsion will decide bond formation. If attraction is more than the repulsion then atoms combine. If repulsion is more than attraction then the atoms do not combine. The nucleus and the electrons in the inner shell remain unaffected when atoms come close together. But the electrons in the outermost shell (valence shell) of atoms get affected.

Do you know?

A Votaic pile

Davy's experiment –

Humphry Davy (1778-1819), a professor of chemistry at the Royal Institution in London, constructed a battery of over 250 metallic plates. In 1807, using electricity from this battery, he was able to extract highly reactive metals like potassium and sodium by electrolysis of fused salts.



Experimental set up by Davy

It was seen that the metal part of the compound migrated towards the negative electrode and the non-metal part towards the positive electrode. So it was proposed that metals are responsible for positively charged particles and non-metals are responsible for negatively charged particles. The oppositely charged particles are held together by electrostatic forces in a compound. Do you agree with this explanation? Why? While this explanation could explain bonding in Na*Cl*, K*Cl* etc. it could not explain bonding in carbon compounds or diatomic molecules of elements.

Electrons in valence shell (valence electrons) are responsible for the formation of bonds between atoms.

In previous lessons you studied about exothermic and endothermic reactions. You also studied about reactivity of elements in periodic table.

Some elements are more reactive and some are less reactive;

- Why there is absorption of energy in certain chemical reactions and release of energy in other reactions?
- Where the absorbed energy goes?
- Is there any relation to energy and bond formation between atoms?
- What could be the reason for the change in reactivity of elements?

Lewis symbols *(or)* Lewis dot structures:

Periodic Classification and arrangement of elements in the periodic table according to the electron configuration gave a new thought about chemical bond.

The discovery of noble gases and the understanding of their electronic configurations helped us in explaining the formation of chemical bonds among the atoms of the elements. Noble gases which belong to zero group (18^{th} group or VIII A) are typical gases with almost negligible chemical activity when compared to other elements. They undergo few or no chemical changes. They are more stable and do not form molecules by allowing their atoms to combine among themselves or with the atoms of other elements.

- What could be the reason for this?

Let us investigate.

Consult the periodic table given in the previous chapter and fill the following table:

Table-1

Element	Z	Electronic configuration				Valence elctrons
		K	L	M	N	
Helium (He)	2	2				2
Neon (Ne)	10	2	8			8
Argon(Ar)	18	2	8	8		8
Krypton(Kr)	36	2	8	18	8	8

Look at the second column and third column. It is clear that all the noble gases have eight electrons in the outermost shell, except Helium (He).

The arrangement of electrons in different shells of atoms (electronic configuration) of 18^{th} group elements is shown in the table-1. The valence electrons in the atom of an element is depicted in a short form by *Lewis symbol* or *electron dot structure*. We represent the nucleus and inner shell electrons of the atom by the symbol of the element and electrons in the outer shell by dots or cross marks.

Let us see how? Lewis dot structure of **argon** and **sodium** atoms are written.

Let us start with *argon*. Its atom has eight valence electrons.

First write the symbol of element *argon*. **Ar**

Place the valence electrons around the symbol. Put two dots at a time on each of the four sides of the symbol of the element till all are used up. So, we get,

$$:\overset{\bullet\bullet}{\underset{\bullet\bullet}{\mathbf{Ar}}}:$$

Similarly, for sodium, the number of valence electrons in sodium is one and the symbol is Na. We can also use cross mark for the electrons. The Lewis structure for sodium atom is therefore:

$$\mathbf{Na}^{\times}$$

The Lewis dot structures of the atoms of noble gases are shown below:

$\overset{\bullet\bullet}{\mathbf{He}}$	$:\overset{\bullet\bullet}{\underset{\bullet\bullet}{\mathbf{Kr}}}:$	$:\overset{\bullet\bullet}{\underset{\bullet\bullet}{\mathbf{Ne}}}:$
$:\overset{\bullet\bullet}{\underset{\bullet\bullet}{\mathbf{Xe}}}:$	$:\overset{\bullet\bullet}{\underset{\bullet\bullet}{\mathbf{Ar}}}:$	$:\overset{\bullet\bullet}{\underset{\bullet\bullet}{\mathbf{Rn}}}:$

Activity 1

Write the Lewis structures of the given elements in the table. Also consult the periodic table and fill in the group number of the element.

Table – 2:

Element	Hydrogen	Helium	Beryllium	Boron	Carbon	Nitrogen	oxygen
Group number	1						
Valence electron	1						
Lewis dot structure	$H^{\bullet}$						

Look at the periodic table. Do you see any relation between the number of valence electrons and group numbers? We find that for groups 1-2 and 13 –18, we can use the periodic table to find the number of valence electrons. Group1 has one outer electron, group 2 has two, and group 13 has three and so on.

- What did you notice in Lewis dot structure of noble gases and electronic configurations of the atoms of these elements shown in table - 1.

It was found that the elements which participate in chemical reactions get octet or ns^2 np^6 configuration similar to that of noble gas elements. It is to be noted that octet rule is still a rule not the law, because there are considerable exceptions for this rule.

Electronic theory of valence by Lewis and Kossel

There were number of attempts to explain the chemical bond formation between atoms in terms of electrons but a satisfactory explanation for this concept was given by Kossel and Lewis in 1916. They gave this concept independently. The basis for their theory was valence in terms of electrons. They provided logical explanation of valence on the basis of the lack of chemical activity of noble gases which led to the proposal of *octet rule.*

Observe the practical behaviour of atoms of main group elements (Group IA, IIA, IIIA, IVA, VA, VIA, VIIA and zero or VIIIA group elements). When they are allowed to undergo chemical changes, they try to get octet electronic configuration in the outer shells.

Let us understand this with the following illustrations.

Group IA elements (Li to Cs) try to lose one valence shell electron from their atoms to form corresponding uni-positive ions which get octet in their outer shells.

Example: ${}_{11}Na \rightarrow 2, 8, 1$; ${}_{11}Na^{+} \rightarrow 2, 8$

Group IIA elements (Mg to Ba) try to lose two valence electrons from their atoms during chemical changes and form di– positive ions with the octet in the outer shells.

Example: ${}_{12}Mg \rightarrow 2, 8, 2$; ${}_{12}Mg^{2+} \rightarrow 2, 8$

Group IIIA elements try to lose three valence electrons from their atoms and form corresponding tri positive ions with octates in the outer shells.

Example: ${}_{13}Al \rightarrow 2, 8, 3$; ${}_{13}Al^{3+} \rightarrow 2, 8$

Group VIA elements try to gain two electrons into the valence shells of their atoms during the chemical changes and form corresponding di negative anions which get octet in their outer shells.

Example: ${}_{8}O \rightarrow 2, 6$; ${}_{8}O^{2-} \rightarrow 2, 8$

Group VIIA elements try to gain one electron into the valence shells of their atoms during the chemical changes and form corresponding uni– anions which get octet in their outer shells.

Example: ${}_{9}F \rightarrow 2, 7$; ${}_{9}F^{-} \rightarrow 2, 8$

Group VIIIA elements, the noble gases do not try to lose or gain electrons. Generally, helium and neon do not participate in chemical changes. Even other elements of VIIIA do not gain or lose electrons from their atoms when they participate in a very few chemical changes.

Example: ${}_{10}Ne \rightarrow 2, 8$; No electron gain or loss from the neon atom.

Gain Electrons ↓ **Lose Electrons** ↓

Net charge on the species	–3	–2	–1	Magic octet	+1	+2	+3
	V	**VI**	**VII**	**VIII**	**I**	**II**	**III**
	N	O	F	Ne	Na	Mg	Al
	P	S	Cl	Ar	K	Ca	Ga
	As	Se	Br	Kr	Rb	Sr	In
	Sb	Te	I	Xe	Cs	Ba	Tl
	Bi	Po	At	Ra	Fr	Ra	
	Non metals			**Noble gases**	**Metals**		

- What have you observed from the above conclusions about the main groups?
- Why do atoms of elements try to combine and form molecules?

Noble gases of VIII A possess eight electrons in the valence shells of their atoms. Helium is an exception. Its atom has only two electrons, but its only shell is completely filled. Noble gases with eight electrons in the valence shell in their atoms are highly stable and rarely participate in chemical changes. Therefore it is concluded that any species (atom or ion) with eight electrons in the valence shell is stable.

- Is it accidental that IA to VIIA main group elements during chemical reactions get eight electrons in the outermost shells of their ions, similar to noble gas atoms?

No. It cannot be simply accidental. Eight electrons in the outermost shell definitely gives stability to the ion or atom. Based on the above observations a statement known as " The Octet Rule" is framed.

Octet rule:

It may be stated as "The atoms of elements tend to undergo chemical changes that help to leave their atoms with eight outer-shell electrons."

Lewis depicted the atom in terms of a positively charged kernel (Kernel is the nucleus and all other electrons in the atom except the outer most shell electrons) and the outershell that could accommodate a maximum of eight electrons.

Chemically active elements do not have an octet of electrons in the valence shell of their atoms. Their reactivity arises from their tendency to achieve the octet, by forming bonds either with atoms of their own type or with atoms of other elements.

The force of attraction between any two atoms or a group of atoms that results a stable entity is called a *'chemical bond'*. There are many types of chemical bonds, but here we discuss only about *ionic bond* and *covalent bond.*

Ionic and Covalent bonds with Lewis dot formulae

A. Ionic bond

Kossel proposed the ionic bond (electrostatic bond) based on the following facts.

i. Ionic bond is formed between atoms of two dissimilar elements due to transfer of electrons from the atom of one element to the other.

ii. There are highly reactive metals like alkali metals (IA) and highly reactive non-metals like halogens (VIIA) on the left side and right side of the periodic table respectively.

iii. Noble gases except helium have eight electrons in the valence shells of their atoms. They are chemically inactive and stable.

iv. To attain eight electrons in the outermost shell similar to noble gases metal atoms that have one two or three electrons in the valence shells generally lose those electrons and form stable positive ions called *cations*.

Example: $_{11}Na$ 2, 8, 1 ; Na^+ 2, 8

$_{12}Mg$ 2, 8, 2 ; Mg^{2+} 2, 8

$_{13}Al$ 2, 8, 3 ; Al^{3+} 2, 8

Do you know?

The number of electrons lost from a metal atom is the valence of its element which is equal to its group number.

Ex. Na and Mg have valence 1 and 2 respectively.

v. To attain eight electrons in the outermost shell, non-metal atoms that have 5,6 or 7 valence electrons gain 3,2 or 1 electron respectively and form negative ions known as *anions*.

Example: $_{15}P$ 2, 8, 5 ; P^{3-} 2, 8, 8

$_{16}S$ 2, 8, 6 ; S^{2-} 2, 8, 8

$_{17}Cl$ 2, 8, 7 ; Cl^- 2, 8, 8

Do you know?

The number of electrons gained by a non-metal element for its atom is its valency, which is equal to (8 - its group number).

Ex. The valency of chlorine is (8 – 7) = 1.

vi. Formation of ionic bond

The positive ions (cations) and negative ions (anions) that are formed due to the transfer of electrons from the metal atoms to the non-metal atoms experience the electrostatic forces and get attracted to form chemical bond. As this bond is between charged particles known as ions, it is called *ionic bond*. Sometimes based on the forces being electrostatic, the bond is also called *the electrostatic bond.* As the valence concept has been explained in terms of electrons, it is also called the *electrovalent bond*.

Thus, we can define ionic bond as follows:

The electrostatic attractive force that keeps cation and anion (which are formed from metal atoms and non-metal atoms due to transfer of electrons) together to form a new electrically neutral compounds is called *'ionic bond'*.

- Explain the formation of ionic compounds $NaCl$, $MgCl_2$, Na_2O and AlF_3 through Lewis electron dot symbols (formulae)

Eg-1. Formation of sodium chloride ($NaCl$):

Sodium chloride is formed from the elements sodium and chlorine. It can be explained as follows:

$$Na_{(s)} + 1/2\ Cl_{2(g)} \rightarrow NaCl_{(s)}$$

Cation formation

When sodium (Na) atom loses one electron to get octet electron configuration it forms a cation (Na+) and gets electron configuration that of Neon (Ne) atom.

$${}_{11}Na_{(g)} \rightarrow {}_{11}Na^{+}_{(g)} + e^{-}$$

E.C. 2, 8, 1 → 2, 8

or $[Ne]\ 3s^1$ → $[Ne]$

Anion formation

Chlorine has shortage of one electron to get octet in its valence shell. So it gains the electron from Na atom to form anion and gets electron configuration as that of argon (Ar).

$${}_{17}Cl_{(g)} + e^{-} \rightarrow Cl^{-}_{(g)}$$

Electronic configuration 2, 8, 7 → 2, 8, 8

or $[Ne]\ 3s^2 3p^5$ → $[Ne]\ 3s^2\ 3p^6$ or $[Ar]$

Formation of the compound NaCl from its ions:

Transfer of electrons between 'Na' and 'Cl' atoms, results in the formation of 'Na^+' and 'Cl^-' ions. These oppositely charged ions get attracted towards each other due to electrostatic forces and form the compound sodium chloride ($NaCl$).

$$Na^{+}_{(g)} + Cl^{-}_{(g)} \rightarrow Na^{+}Cl^{-}_{(s)} \text{ or } NaCl$$

Eg-2. Formation of magnesium chloride ($MgCl_2$):

Magnesium chloride is formed from the elements magnesium and chlorine. The bond formation $MgCl_2$ in brief using chemical equation is as follows:

$$Mg_{(s)} + Cl_{2(g)} \rightarrow MgCl_{2(s)}$$

Cation formation

$${}_{11}Mg_{(g)} \rightarrow {}_{11}Mg^{2+}_{(g)} + 2e^{-}$$

Electronic configuration 2, 8, 2 → 2, 8

or $[Ne]\ 3s^2$ → $[Ne]$

Anion formation

$$2Cl_{(g)} + 2e^{-} \rightarrow 2Cl^{-}_{(g)}$$

Electronic configuration 2, 8, 7 → 2, 8, 8

or $[Ne]\ 3s^2 3p^5$ → $[Ne]\ 3s^2\ 3p^6$ or $[Ar]$

The compound $MgCl_2$ formation from its ions:

Mg^{2+} gets 'Ne' configuration and

Each Cl^- gets 'Ar' configuration

$$Mg^{2+}_{(g)} + 2Cl^-_{(g)} \rightarrow MgCl_{2(s)}$$

One 'Mg' atom transfers two electrons one each to two 'Cl' atoms and so formed Mg^{2+} and $2Cl^-$ attract to form $MgCl_2$.

Eg-3. Formation of di sodium monoxide(Na_2O):

Di sodium monoxide formation can be explained as follows:

Cation formation (Na+ formation):

	$2Na_{(g)}$	$\rightarrow$	$2Na^+_{(g)} + 2e^-$
Electronic configuration	2, 8, 1		2, 8
or	[Ne] $3s^1$		[Ne]

Anion formation (O^{2-}, the oxide formation):

	$O_{(g)} + 2e^-$	$\rightarrow$	$O^{2-}_{(g)}$
Electronic configuration	2, 6		2, 8
or	[Ne] $2s^2 2p^4$		[He] $2s^2 2p^6$ or [Ne]

The compound Na_2O formation from its ions is as shown.

$$2Na^+_{(g)} + O^{2-}_{(g)} \rightarrow Na_2O_{(s)}$$

Two 'Na' atoms transfer one electron each to one oxygen atom to form $2Na^+$ and O^{2-}

Each Na^+ gets 'Ne' configuration and O^{2-} gets 'Ne' configuration.

These ions ($2Na^+$ and O^{2-}) attract to form Na_2O.

Eg-4. Formation of aluminium chloride ($AlCl_3$):

Aluminium chloride formation can be explained as follows:

Formation of aluminium ion (Al^{3+}), the cation:

	${}_{13}Al_{(g)}$	$\rightarrow$	$Al^{3+}_{(g)} + 3e^-$
Electronic configuration	2, 8, 3		2, 8
or	[Ne] $3s^2 3p^1$		[Ne]

Formation of chloride ion (Cl-), the anion:

	$3Cl_{(g)} + 3e^-$	$\rightarrow$	$3Cl^-_{(g)}$
Electronic configuration	2, 8, 7		2, 8,8
or	[Ne] $3s^2 3p^5$		[Ne] $3s^2 3p^6$ or [Ar]

Each aluminium atom loses three electrons and three chlorine atoms gain them, one electron each.

The compound $AlCl_3$ is formed from its component ions by the electrostatic forces of attractions.

$$Al^{3+}_{(g)} + 3Cl^{-}_{(g)} \rightarrow AlCl_{3\,(s)}$$

The arrangement of ions in ionic comounds

- How do cations and anions of an ionic compound exist in its solid state?
- Let us explain this with sodium chloride as our example:
- Do you think that pairs of $Na^+ Cl^-$ as units would be present in the solid crystal?

If you think so, it is not correct. Remember that electrostatic forces are non directional. Therefore, it is not possible for one Na^+ to be attracted by one Cl^- and vice-versa. Depending upon the size and charge of a particular ion, number of oppositely charged ions get attracted by it, but, in a definite number. In sodium chloride crystal each Na^+ is surrounded by 6 Cl^- and each Cl^- by six Na^+ ions. Ionic compounds in the crystalline state consist of orderly arranged cations and anions held together by electrostatic forces of attractions in three dimensions. The crystal structure of sodium chloride is given below:

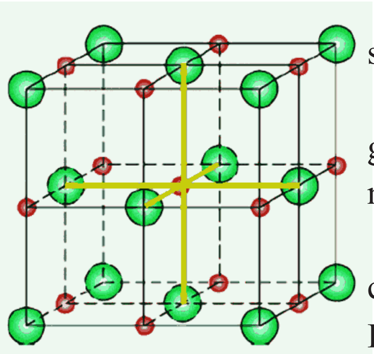


fig-1

NaCl is said to possess face centred cubic lattice crystal structure (see figure-1).

The number of ions of opposite change that surround a given ion of given charge is known as the coordination number of that given ion.

For example, in sodium chloride crystal, the coordination number of Na^+ is 6 and that of Cl^- is also 6.

Factors affecting the formation of cation and anion:

In previous lesson you have learnt about variation of metallic and non metallic character of elements in a group or period of the periodic table. Recall the facts about metallic and non metallic character of elements.

Generally elements of metals have tendency of losing electron to attain the octet in their valence shell. This property is called as the *metallic character* or *electropositivity*. Elements with more electropositive character form cations. Similarly non metals like oxygen ($_8O$), fluorine ($_9F$) and chlorine ($_{17}Cl$) acquire electron configuration of elements of inert gases by gaining electrons this property is called the *non-metallic character* or *electronegativity of the element*. Elements with more electronegative character form anions.

- Can you explain the reasons for all these?

Ionic bond is formed between atoms of elements with electronegativity difference equal to or greater than 1.9.

In ionic bond formation, you have noticed that atoms either lose electrons or gain electrons to attain octet in valence shell. i.e., in an ionic bond, transfer of electrons takes place between combining atoms.

The tendency of losing electrons to form cations (or) gaining electron to form anions depends on the following factors:

i. Atomic size

ii. Ionisation potential

iii. Electron affinity

iv. Electronegativity

The atoms of elements with low ionisation energy, low electron affinity high atomic size and low electronegativity form cations.

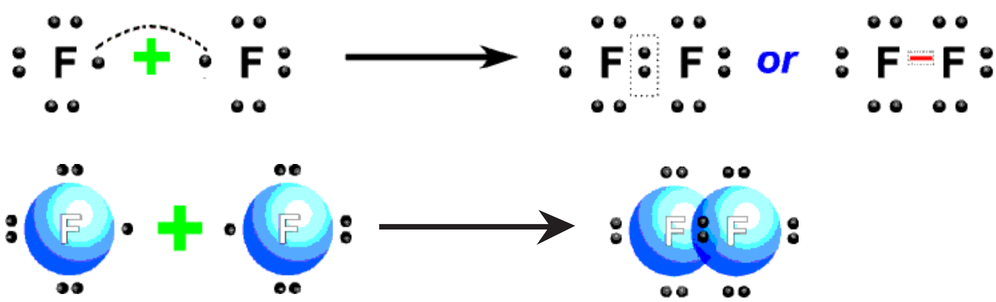
The atoms of elements with high ionisation potential, high electron affinity, small atomic size and hight electronegativity form anions.

B. Covalent bond

G.N. Lewis (1916) proposed that atoms of some elements could achieve an octet in their valence shells without transfer of electrons between them. They can attain octet configuration in their valence shells by sharing the valence electrons with one or more atoms.

The electrons shared between two atoms belong to both the atoms and the sharing of electrons between them leads to the formation of a chemical bond known as *covalent bond*.

For example, take two fluorine atoms which form a stable molecule. Each fluorine atom contributes one electron for bonding and the electron pair that is formed in this way is mutually shared by both the fluorine atoms. Each atom in the F_2 molecule has an octet of valence electrons.



The dots around fluorine atom shows the valence electrons of respective atoms.

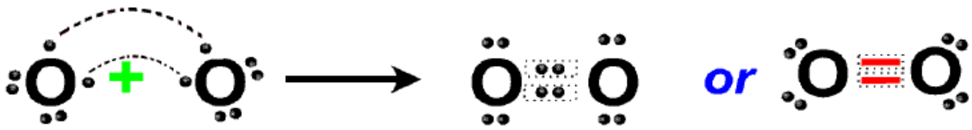
The chemical bond formed between two atoms by mutual sharing of a pair of valence shell electrons so that both of them can attain octet or duplet in their valence shell is called the ***covalent bond***.

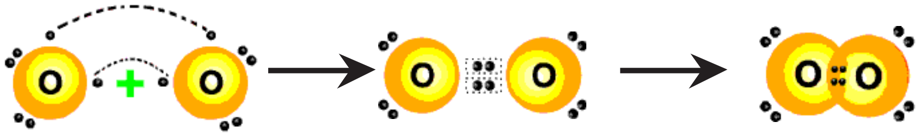
The prefix co – is used to indicate when things are equal or joined. (for example, co exist, cooperate etc). Here each atom contributes one electron from the valence shell for the bonding. Therefore, the name covalent is given (equal valence electrons contribution) for this bond.

Formation of O_2 molecule

The electronic configuration of ${}_8O$ is 2, 6. Oxygen atom has six electrons in its valence shell. It requires two more electrons to get octet in its valence shell. Therefore oxygen atoms come close and each oxygen atom contributes two electrons for bonding. Thus, there exist two covalent bonds between two oxygen atoms in O_2 molecule as there are two pairs of electrons shared between them.

We can say that a double bond is formed between two oxygen atoms in O_2 molecule. Observe the following figures. Both the oxygen atoms have octet in the valence shell.



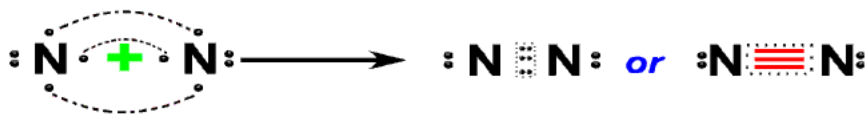


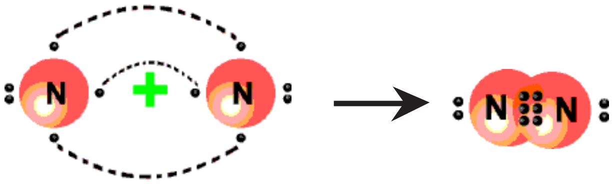
- Can you say what type of bond exists between atoms of nitrogen molecule?

  Let us see

Nitrogen (N_2) molecule

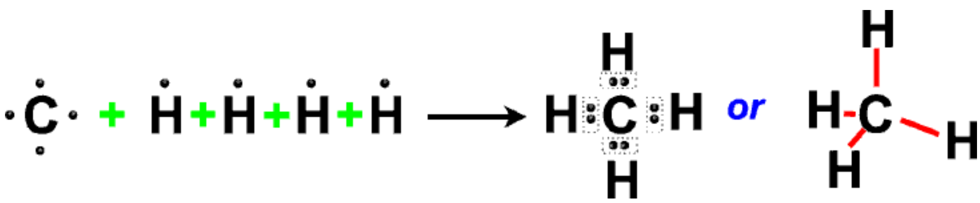
The electronic configuration of 'N' atom is 2,5 and to have octet in the valence shell it requires three more electrons. When two nitrogen atoms approach each other, each atom contributes 3 electrons for bonding. There are six electrons shared between two nitrogen atoms in the form of three pairs. Therefore, there is a triple bond between two nitrogen atoms in N_2 molecule.

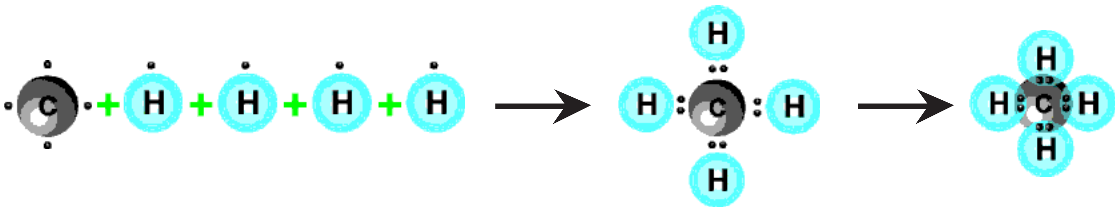




Methane (CH_4) molecule

In the formation of methane, CH_4 molecule, carbon contributes 4 electrons, (one electron to each hydrogen atom) and 4 hydrogen atoms contribute one electron each. Thus in CH_4 molecule, there are four C – H covalent bonds as shown below:



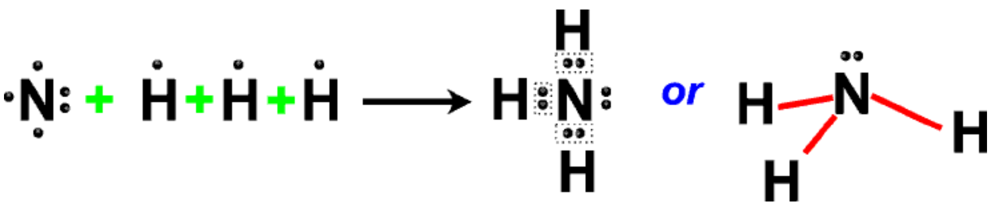


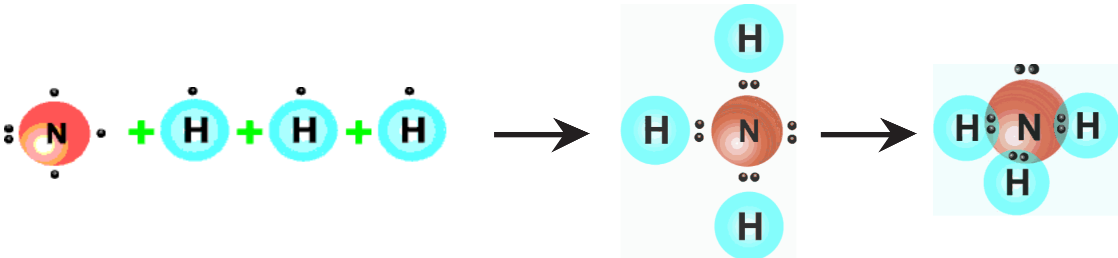
Ammonia (NH_3) molecule

In ammonia molecule, three N – H single covalent bonds are present.

Electron configuration of $_7N$ is 2, 5 and $_1H$ is 1.

Nirogen atom contributes three electrons for bonding. Three hydrogen atoms at the same time contribute one electron each for bonding. Six electrons form three pairs and each pair is shared between nitrogen and one hydrogen atom as shown below:



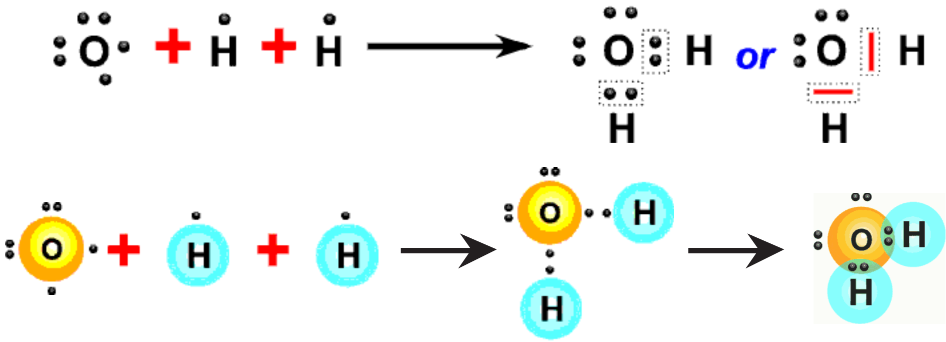


Water (H_2O) molecule

In water molecule (H_2O), there are two O – H single covalents bonds.

Electron configuration of $_8O$ is 2,6 and $_1H$ is 1.

Oxygen atom needs '2' electrons to attain octet in its valence shell. Therefore it shares '2' electron with two hydrogen atoms to form water (H_2O) molecule.



The total number of covalent bonds that an atom of an element forms is called its 'covalency'.

The Bond lengths and Bond energies of covalent bonds

Bond length or bond distance is the equilibrium distance between the nuclei of two atoms which form a covalent bond.

It is generally given in *nm* (nanometer) or Å (Angstrom unit).

Bond energy or Bond dissociation energy is the energy needed to break a covalent bond between two atoms of a diatomic covalent compound in its gaseous state.

Do you know?

- An angstrom (Å) is a unit of length equal to 10^{-10} metre, or 0.1 nanometre, or 100 picometre.
- 1 nanometre = 10^{-9} metre

Draw backs of electronic theory of valence

1) When a covalent bond is formed between any two atoms, irrespective of the nature of the atoms, the bond lengths and bond energies are expected to be the same. This is because any covalent bond between any two atoms is a result of mere sharing of two identical electrons. But, practically it was observed that bond lengths and bond energies are not same when the atoms that form the bond are different. (See table - 3)

- What do you understand from bond lengths and bond energies?

- Are the values not different for the bonds between different types of atoms?

2) The theory could not explain why $Cl\hat{B}eCl$ in $BeCl_2$ is 180^0, $F\hat{B}F$ in BF_3 is 120^o, $H\hat{C}H$ in CH_4 is $109^028'$, $H\hat{N}H$ in NH_3 is $107^018'$ and $H\hat{O}H$ in H_2O is $104^o31'$ etc.

i.e., it fails to explain the shapes of the molecules.

Table – 3

Bond	Bond length (Å)	Bond (dissociation) energy ($KJmol^{-1}$)
H–H	0.74	436
F – F	1.44	159
Cl – *Cl*	1.95	243
Br – Br	2.28	193
I – I	2.68	151
H – F	0.918	570
H – *Cl*	1.27	432
H – Br	1.42	366
H – I	1.61	298
H – O (of H_2O)	0.96	460
H – N (of NH_3)	1.01	390
H – C (of CH_4)	1.10	410

Valence shell electron pair repulsion theory

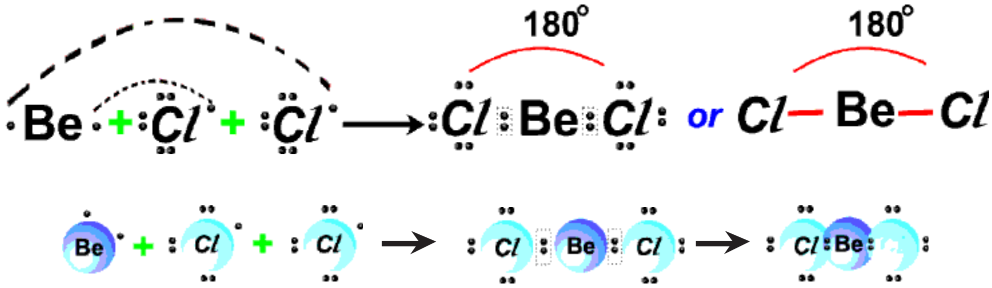
To explain the bond angles in the molecules with three or more than three atoms with all atoms attached to a central atom through covalent bonds a theory called the ***valence – shell – electron – pair repulsion – theory (VSEPRT)*** was proposed by Sidgwick and Powell (1940). It was further improved by Gillespie and Nyholm (1957).

This theory suggests the following points.

1. VSEPRT considers electrons in the valence shells which are in covalent bonds and in lone pairs as charge clouds that repel one another and stay as far apart as possible. This is the reason why molecules get specific shapes.
2. If we know the total number of electron pairs in the valence shell as covalent bonds and lone pairs in the central atom, it will help us to predict the arrangement of those pairs around the nucleus of the central atom and from that the shape of the molecule.
3. Lone pairs occupy more space around the central atom than bond pairs. Lone pair means unshared electron pair or non-bonding electron pair. These are attracted to only one nucleus where as the bond pair is shared by two nuclei. Thus, the presence of lone pairs on the central atom causes slight distortion of the bond angles from the expected regular shape. If the angle between lone pair and bond pair increases at the central atom due to more repulsion, the actual bond angles between atoms must be decreased.

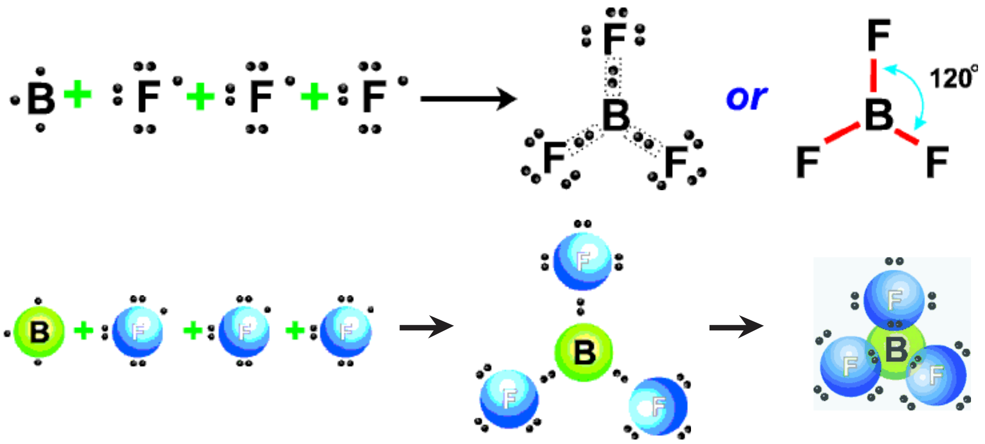
4.i) If two bond pairs are present in two covalent bonds around the nucleus of the central atom without any lone pairs in the valence shell, they must be separated by 180^0 to have minimum repulsion between them. Thus, the molecule would be linear.

Example:



4.ii) If three bond pairs are there in three covalent bonds around the nucleus of the central atom, without any lone pairs they get separated by 120^0 along three corners of a triangle. Therefore, the shape of the molecule is trigonal-planar.

Example:

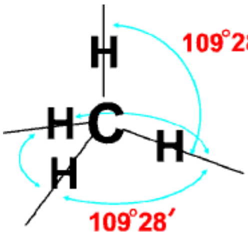


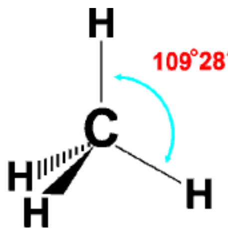
NOTE: Did you notice that in $BeCl_2$ and BF_3 the central atoms Be and B did not possess 8 electrons around them in the valence shell. They have only 4 and 6 electrons respectively. These molecules are known as electron deficient molecules.

4.iii) If there are four bond pairs in the valence shell of the central atom, the four bond pairs will orient along the four corners of a tetrahedron (three dimensional arrangement) and the bond angle expected is $109^0 28'$

.**Example: Methane.**

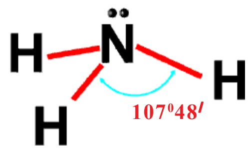
- In methane molecule (CH_4), $H\hat{C}H$ is $109^0 28'$ because of four electron pairs (bonding) around carbon, as shown below:





4.iv) If there are three bond pairs and one lone pair i.e., unshared electron pair, then the lone pair occupies more space around the nucleus of the central atom. The remaining three bond pairs come relatively closer as in NH_3 molecule.

Example: Ammonia

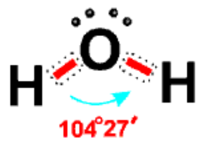


In ammonia (NH_3) molecule, there are three bond pairs in covalent bonds (3 N – H) around the nucleus of the nitrogen atom and one lone pair. Lone pair – bond pair repulsion is greater than bond pair – bond pair repulsion. Therefore, NH_3 which is expected to be tetrahedral with four electron pairs in the valence shell and $H\hat{N}H = 109^0 28'$, it has $H\hat{N}H = 107^0 48'$ due to the more repulsion by lone pair on the bond pairs.

The shape of the NH_3 molecule is triagonal pyramidal with N at the apex of the pyramid.

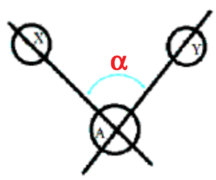
4.v) If there are two bond pairs and two lone pairs of electrons around the nucleus of the central atom in its valence shell, lone pair – lone pair repulsion is greater than lone pair – bond pair repulsion. Therefore, the angle between bond pairs further decreases.

Example: Water



In water molecule, (H_2O) there are four electron pairs around the nucleus of oxygen atom, but, two of them are lone pairs and two bond pairs. Therefore, H_2O molecule gets 'V' shape or bent shape or angular instead of tetrahedral shape as that of CH_4 due to lone pair – lone pair and lone pair – bond pair repulsions. $H\hat{O}H$ is $104^0 31'$.

- What is the bond angle in a molecule?



It is the angle subtended by two imaginary lines that pass from the nuclei of two atoms which form the covalent bonds with the central atom through the nucleus of the central atom at the central atom. (see figure 'α' is the bond angle).

Valence Shell Electron Pair Repulsion Theory (VSEPRT) mainly fails in explaining the strengths of the bonds. This is because; the theory still depends on the Lewis concept of covalent bond formation. It could not say anything extra about the electronic nature of covalent bonds.

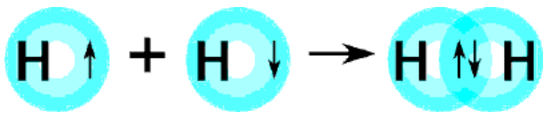
To describe covalent bonding, a quantum mechanical model called *valence bond theory* has been suggested by Linus Pauling (1954).

Valence bond theory

To describe covalent bonding, a quantum mechanical model called *valence bond theory* has been suggested by Linus Pauling (1954). It is explained as follows:

1. A covalent bond between two atoms is formed when the two atoms approach each other closely and one atom overlaps its valence orbital containing unpaired electron, the valence orbital of the other atom that contains the unpaired electron of opposite spin. The so formed paired electrons in the overlapping orbitals are attracted to the nuclei of both the atoms. This bonds the two atoms together.

Eg: In the formation of H_2 molecule, the 1s orbital of one 'H' atom containing an unpaired electron overlaps the '1s' orbital of the other 'H' atom containing unpaired electron of opposite spin giving H-H bond and H_2 molecule.



H atom H atom H_2 molecule

2. The greater the overlapping of the orbitals that form the bond, the stronger will be the bond. This gives a directional character to the bond when other than 's' orbitals are involved.

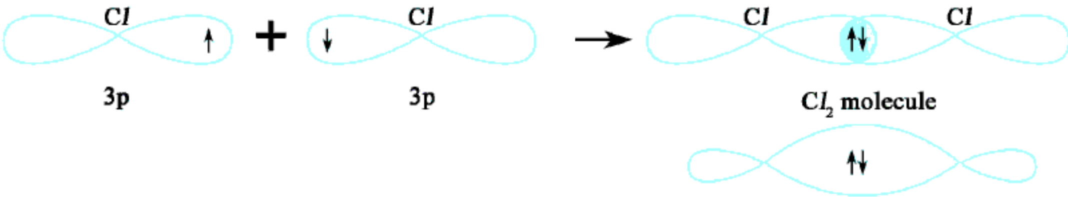
3. Each bonded atom maintains its own atomic orbitals but the electron pair in the overlapping orbitals is shared by both the atoms involved in the overlapping.

4. If two atoms form multiple bonds between them the first bond is due to the overlap of orbitals along the inter-nuclear axis giving a stronger sigma(σ) bond. After formation of (σ) bond the other bonds are formed due to the overlap of orbitals side wise or laterally giving weaker π bonds. The 'σ' bond is stronger because the electron pair shared is concentrated more between the two nuclei due to end-end or head on overlap and attracted to both the nuclei. The π bond overlap gives a weaker bond due to the lateral overlap of 'p' orbitals which is not to greater extent.

Consider *Cl*-*Cl* molecule

$$_{17}Cl \text{ - } 1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p_x^{\,2}\, 3p_y^{\,2}\, 3p_z^{\,1}$$

In the formation of Cl_2 molecule, the $3p_z$ orbital of one chlorine atom containing an unpaired electron overlaps the $3p_z$ orbital of other chlorine atom that contains unpaired electron of opposite spin.

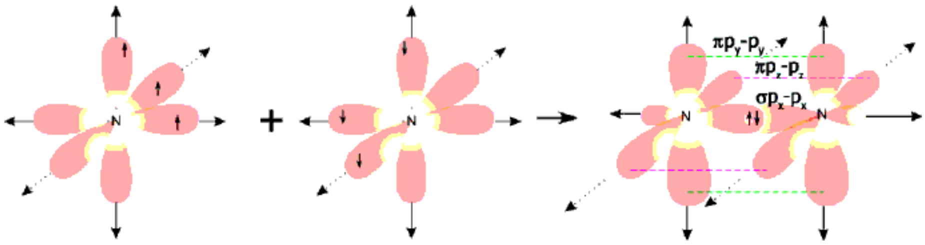


- How is HCl molecule formed?

The '1s' orbital of 'H' atom containing unpaired electron overlaps the '3p' orbital of chlorine atom containing unpaired electron of opposite spin.

Formation of N_2 molecule

$_7$N has electronic configuration $1s^2\ 2s^2\ 2p_x^{\ 1}\ 3p_y^{\ 1}\ 3p_z^{\ 1}$. Suppose that p_x orbital of one 'N' atom overlaps the 'p_x' orbital of the other 'N' atom giving $\sigma\ p_x - p_x$ bond along the inter-nuclear axis. The p_y and p_z orbitals of one 'N' atom overlap the p_y and p_z orbital of other 'N' atom laterally, respectively perpendicular to inter-nuclear axis giving $\pi\ p_y$-p_y and $\pi\ p_z$-p_z bonds. Therefore, N_2 molecule has a triple bond between two nitrogen atoms.



Formation of O_2 molecule

$_8$O has electronic configuration $1s^2\ 2s^2\ 2p_x^{\ 2}\ 3p_y^{\ 1}\ 3p_z^{\ 1}$. If the '$p_y$' orbital of one 'O' atom overlaps the 'p_y' orbital of other 'O' atom along the inter-nuclear axis, a sigma p_y - p_y bond (σp_y - p_y) is formed. p_z orbital of one 'O' atom overlaps the p_z orbital of other 'O' atom laterally, perpendicular to the inter-nuclear axis giving a $\pi\ p_z$- p_z bond. O_2 molecule has a double bond between two oxygen atoms.

Valence bond theory-Hybridisation

Formation of BeCl_2 (Beryllium chloride) molecule

$_4$Be has electronic configuration $1s^2\ 2s^2$. It has no unpaired electrons. It is expected not to form covalent bonds, but it forms two covalent bonds one each with two chlorine atoms. To explain this, an excited state is suggested for Beryllium in which an electron from '2s' shifts to $2p_x$ level.

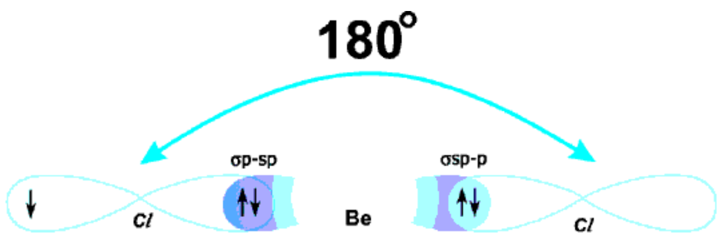
Electronic configuration of $_4$Be $1s^2\ 2s^1\ 2p_x^{\ 1}$ and $_{17}Cl$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^{\ 2}\ 3p_y^{\ 2}\ 3p_z^{\ 1}$

If Be forms two covalent bonds with two chlorine atoms, one bond should be σ2s-3p due to the overlap of '2s' orbital of Be, the '$3p_z$' orbital of one chlorine atom and the other bond should be σ2p-3p due to the overlap of '$2p_x$' orbital of Be atom the 3p orbital of the other chlorine atom. As the orbitals overlapping are different, the bond strengths of two Be-Cl must be different. But, both bonds are of same strength and $Cl\widehat{Be}Cl$ is 180^0. To explain the discrepancies like this a phenomenon called 'hybridisation of atomic orbitals' was proposed by Linus Pauling (1931).

Hybridisation is a phenomenon of intermixing of atomic orbitals of almost equal energy which are present in the outer shells of the atom and their reshuffling or redistribution into the same number of orbitals but with equal properties like energy and shape.

Be atom in its excited state allows its 2s orbital and $2p_x$ orbital which contain unpaired electrons to intermix and redistribute to two identical orbitals. As per Hund's rule each orbital gets one electron. The new orbitals based on the types of orbitals that have undergone hybridisation are called sp orbitals. The two sp orbitals of Be get separated by 180^0. Now, each chlorine atom comes with its $3p_z^{\ 1}$ orbital and overlaps it the sp orbitals of Be forming two identical Be-Cl bonds (σsp-p bonds) . $Cl\widehat{Be}Cl = 180^0$.

Both the bonds are of same strength.



Boron triflouride

Formation of BF_3 molecule

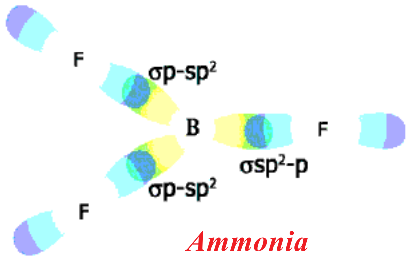
$_5B$ has electronic configuration $1s^2\ 2s^2\ 2p_x^{\ 1}$.

As it has one unpaired electron ($2p_x^{\ 1}$) it should form only one covalent bond to give B-F molecule. But we get practically BF_3 molecule.

To explain this, it is suggested that,

i. Boron (B) first undergoes excitation to get electronic configuration $1s^2\ 2s^1\ 2p_x^{\ 1}\ 2p_y^{\ 1}$.

ii. As it forms three identical B-F bonds in BF_3, it is suggested that excited 'B' atom undergoes hybridisation. There is an intermixing of 2s, $2p_x$, $2p_y$ orbitals and their redistribution into three identical orbitals called sp^2 hybrid orbitals. For three sp^2 orbitals to get separated to have minimum repulsion the angle between any two orbitals is 120^0 at the central atom and each sp^2 orbital gets one electron. Now three fluorine atoms overlap their $2p_z$ orbitals containing unpaired electrons (F_9 $1s^2$ $2s^2$ $2p_x^2$ $2p_y^2$ $2p_z^1$) the three sp^2 orbitals of 'B' that contain unpaired electrons to form three σsp^2-p bonds.



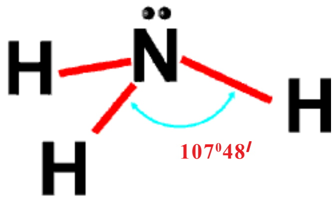
Ammonia

Formation of NH_3 molecule

One nitrogen atom and three hydrogen atoms are present in one ammonia molecule. All the three N-H bonds are of same strength and $H\hat{N}H = 107^0 48'$

$_7N$ has electronic configuration $1s^2$ $2s^2$ $2p_x^1$ $2p_y^1$ $2p_z^1$.

If three hydrogen atoms overlap their 1s orbitals on the three 'p' orbitals of nitrogen atom, they give identical σ p-s bonds but, then the $H\hat{N}H$ should be equal to 90^0 where as $107^0 48'$. To explain the discrepancy in the bond angle 'N' atom is said to undergo sp^3 hybridisation. In this process '2s' and $2p_x$, $2p_y$, $2p_z$ orbitals of nitrogen intermix and redistribute into four identical sp^3 orbitals. One of the four sp^3 orbitals get a pair of electrons and the other three sp^3 orbitals get one electron each. Now hydrogen atoms overlap their 1s orbitals containing unpaired electrons the sp^3 orbitals of 'N' atom containing unpaired electrons to give three σs-sp^3 bonds. $H\hat{N}H$ should be 109^0 $28'$ for sp^3 hybridisation. As there is a lone pair in one of the sp^3 orbitals, there is a greater lone pair – bond pair repulsion which decreases the bond angle $H\hat{N}H$ to $107^0 48'$.



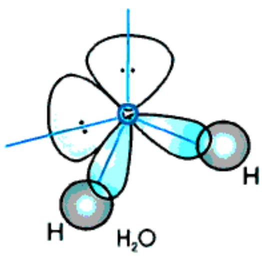
Shape of water molecule

It is found that $H\hat{O}H$ is $104^0 31'$.

The electronic configuration of $_8O$ is $1s^2$ $2s^2$ $2p_x^2$ $2p_y^1$ $2p_z^1$ and $_1H$ is $1s^1$.

Therefore, there should be two σ s-p bonds due to the overlap of 's' orbitals of two hydrogen atoms, the 'p' orbitals of oxygen atom which contain unpaired electrons. $H\hat{O}H$ should be 90^0.

But $H\hat{O}H$ observed is $104^0 31'$. To explain this, sp^3 hybridisation is suggested for the valence orbitals of 'O' atom. One s-orbital (2s) and three 'p' orbitals ($2p_x$, $2p_y$, $2p_z$) intermix and redistribute into four identical sp^3 orbitals. As there are six electrons and two sp^3 orbitals get pairs and two sp^3 orbitals get one electron each. Now, the two sp^3 orbitals of 'O' atom overlap



the 's' orbitals of two hydrogen atoms to give σsp^3-s bonds. Due to the lone pair – lone pair repulsions and lone pair – bond pair repulsions $H\hat{O}H$ decreases from $109^0 28'$ (expected for sp^3 – tetrahedral hybridisation) to $104^0 31'$.

CH_4, C_2H_4 and C_2H_2 molecules and their structures will be explained in carbon and its compounds chapter later in this class.

Properties of ionic and covalent compounds

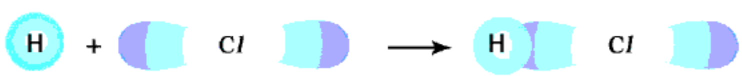
Table -4

S.No	Property	NaC*l* (ionic)	HC*l* (polar covalent)	C_2H_6 (covalent)
1.	Formula mass	58.5	36.5	30.0
2.	Physical appearance	White crystalline solid	Colourless gas	Colourless gas
3.	Type of bond	Ionic	Polar covalent	Covalent
4.	Melting point	801 ^{0}C	-115 ^{0}C	-183 ^{0}C
5.	Boiling point	1413 ^{0}C	-84.9 ^{0}C	-88.63 ^{0}C
6.	Solubility	Soluble in polar solvents like water and insoluble in non -polar solvents	Soluble in polar solvents like water and to some extent in non-polar solvents	Soluble in non-polar solvents but insoluble in polar solvents like water
7.	Chemical activity	Highly reactive in polar solvents and reactions are instantaneous	Moderately reactive	Slow or very slow at room temperature

From the above table we understand that ionic compounds like NaC*l* are solids at room temperature.

Polar compounds like HC*l* possess properties like melting point, boiling point, reactivity, solubility etc, between those of ionic compounds and covalent compounds. If the covalent bond is between atoms of two different elements, the shared electron pair shift more towards the atom of more electronegative element. Thus within the molecule the more electronegative atom bears a partial negative charge and the less electronegative atom bears a partial positive charge. A molecule of this type which is neutral but possesses partial charges on the atoms within the molecule is called a polar molecule and the bond is called a polar covalent bond or partial ionic and partial covalent bond.

Eg: $\quad H + Cl \rightarrow \overset{+\delta}{H} - \overset{-\delta}{Cl}$



In ionic compounds there exist stronger electrostatic forces of attractions between the oppositely charged ions of these compounds. Therefore, they are solids with high melting points and boiling points. Based on the principle "like dissolves in like", they being highly polar, soluble in polar solvents. In chemical reaction of solution of ionic compounds simply rearrangement of ions take place in the solutions, the reactions are instantaneous or very fast.

The forces of attractions among covalent molecules are weak. Therefore, the covalent compounds are gases or liquid at room temperature. They have low melting points and low boiling points. Based on the principle that "like dissolve in like" these covalent compounds are soluble in non-polar solvents because of non-polar nature of solvent molecules. In chemical reactions of covalent compounds there exist bond breaking and bond forming to get products. There fore, these reactions are moderate or very slow.

'Like dissolves in like' means what type of chemical bonds are there in the solute particles that solute could be soluble in that solvent which has the similar type of chemical bonds in its molecules.

Key words

Electrons, noble gases, Lewis dot structures, Octet rule, chemical bond, Ionic bond, covalent bond, cation, anion, electrostatic force, electrovalent, polar solvent, non-polar solvent, formation of molecules, ionic compounds, covalent compounds, electro positive character, electro negative character, polar bonds, bonded pair of electrons, lone pairs, bond length, bond energy, shape of the molecule, linear, tetrahedral, properties of ionic and covalent compounds.

What we have learnt

- Location of elements in the periodic table helps in predicting the type of bonding that will take place between the elements.
- Ions are positively or negatively charged partiCles formed by the loss or gain of electrons.
- The force between any two atoms or a group of atoms that results in the formation of a stable entity is called chemical bond.
- The outer most shell is called valence shell and electrons in this are called valence electrons.
- The gases belongs to 'O' group are called noble gases because they reluctant to combine other atoms. Except helium all other noble gases have an octet of electron configuration in their valence shell.
- The reason why atoms bond can be explained on the basis of the octet rule
- Chemically active elements have an incomplete octet in their valence shell of the atoms.
- The number of valence electrons available in the atoms decides the type of bond
- Elements which have tendency to gain electrons for attaining octet in their valence shell called electro negative character elements. They form anions.
- In the formation of ionic bond the atoms of electro positive elements lose their valence electrons to atoms of electro negative elements so that both of them can attain octet in their valence shell.

- The electrostatic attractive force that keeps cation and anion together to form a new electrically neutral entity is called an ionic bond.
- Ionic compounds are often crystalline solids with high melting points.
- A chemical bond that formed by sharing of valence-shell electrons between the atoms so that both of them can attain octet or duplet in their valence shell is called covalent bond.
- A single covalent bond is formed when two atoms share a pair of electrons
- Each shared pair of electron is equivalent to a covalent bond.
- Electrons are not always shared equally between the atoms in a covalent bond. This leads to bond polarity.
- Formation of various molecules.
- Factors affecting the formation of cation and anion:
- To explain the bond angles in the molecules through covalent bonds the valence – shell – electron – pair repulsion theory (VSEPRT)
- Bond lengths and bond energies in covalent compounds
- Properties of ionic and covalent compounds.

Improve your learning

1. List the factors that determine the type of bond that will be formed between two atoms? (AS1)
2. Explain the difference between the valence electrons and the covalency of an element. (AS1)
3 A chemical compound has the following Lewis notation: (AS1)
   a) How many valence electrons does element Y have?
   b) What is the valency of element Y?
   c) What is the valency of element X?
   d) How many covalent bonds are there in the molecule?
   e) Suggest a name for the elements X and Y. (AS2)

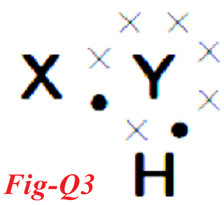


Fig-Q3

4. Why do only valence electrons involve in bond formation? Why not electron of inner shells? Explain. (AS1)
5. Explain the formation of sodium chloride and calcium oxide on the basis of the concept of electron transfer from one atom to another atom. (AS1)
6. A, B, and C are three elements with atomic number 6, 11 and 17 respectively.
   i. Which of these cannot form ionic bond? Why? (AS1)
   ii. Which of these cannot form covalent bond? Why? (AS1)
   iii. Which of these can form ionic as well as covalent bonds? (AS1)
7. How bond energies and bond lengths of molecule helps us in predicting their chemical properties? Explain with examples. (AS1)
8. Predict the reasons for low melting point for covalent compounds when compared with ionic compound. (AS2)
9. Collect the information about properties and uses of covalent compounds and prepare a report? (AS4)
10. Draw simple diagrams to show how electrons are arranged in the following covalent molecules: (AS5)
    a) Calcium oxide (CaO) (b) Water (H_2O) (c) Chlorine (Cl_2)

11. Represent the molecule H_2O using Lewis notation. (AS5)
12. Represent each of the following atoms using Lewis notation: (AS5)
    (a) beryllium (b) calcium (c) lithium
13. Represent each of the following molecules using Lewis notation: (AS5)
    (a) bromine gas (Br_2) (b) calcium chloride ($CaCl_2$) (c) carbon dioxide (CO_2)
    (d) Which of the three molecules listed above contains a double bond?
14. Two chemical reactions are described below. (AS5)
    • Nitrogen and hydrogen react to form ammonia (NH_3)
    • Carbon and hydrogen bond to form a molecule of methane (CH_4).
    For each reaction, give:
    (a) The valency of each of the atoms involved in the reaction. (AS1)
    (b) The Lewis structure of the product that is formed. (AS5)
15. How Lewis dot structure helps in understanding bond formation between atoms? (AS6)
16. What is octet rule? How do you appreciate role of the 'octet rule' in explaining the chemical properties of elements? (AS6)
17. Explain the formation of the following molecules using valence bond theory
    a) N_2 molecule b) O_2 molecule
18. What is hybridisation? Explain the formation of the following molecules using hybridisation
    a) Be Cl_2 b) BF_3

Fill in the blanks

1. Electrons in .... shell are called valence electrons.
2. Except .... gas all other noble gases have octet in their valence shell.
3. Covalency of elements explains about member of .... formed by the atom.
4. Valence bond theory was proposed by ....
5. In ....bonding the valence electrons are shared among all the atoms of the metallic elements.

Multiple choice questions

1) Which of the following elements is electronegative? [ ]
   a) Sodium b) Oxygen c) Magnesium d) Calcium
2) An element ${}_{11}X^{23}$ forms an ionic compound with another element 'Y'. Then the charge on the ion formed by X is [ ]
   a) +1 b) +2 c) -1 d) -2
3) An element 'A' forms a chloride ACl_4. The number electrons in the valence shell of 'A'
   a) 1 b) 2 c) 3 d) 4 [ ]