Chemical Bonding

Elementary Idea of Chemical Bonding

Chemical Bonding

Elements are rarely capable of free existence. In a compound, atoms of different elements are held together by bonds. The types of bonds present in a compound are largely responsible for its physical and chemical properties. The different bonds can be classified as **strong** and **weak**.

Why do elements undergo bond formation?

Elements are made of atoms, which comprise of protons, electrons, and neutrons. The protons and the neutrons reside in the nucleus and the electrons revolve around in definite paths called **orbits**. The electrons present in the last shell are called valence electrons. These electrons are responsible for all the chemical reactions of that element.

Every element has a tendency to attain a stable outer octet. To do so, it either gains or loses or shares its electrons; and in this process, it forms the bonds.

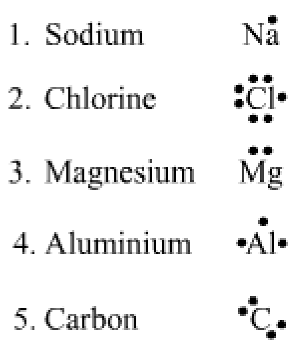
Types of strong bonds:

- Ionic or electrovalent bond
- Covalent bond
- Metallic bond

Types of weak bonds:

- Bonds formed due to van der Waal’s interaction
- Hydrogen bond

This representation of elements with valence electrons as dots around elements is referred to as **Electron Dot structures** for elements. The electron dot structure of some of the elements are:



Formation of Ionic Compounds and Their Properties

We know that common salt is an important dietary mineral essential for animal life. Common salt is chemically known as sodium chloride. The chemical formula of sodium chloride is NaCl. It suggests that it is made up of sodium, which is a reactive metal, and chlorine, which is a non-metal.

Do you know that sodium chloride does not exist as molecules, but aggregates as oppositely charged ions?

An ion is a charged species, which can be negatively charged or positively charged. A negatively charged species is called an 'anion' and a positively charged species is called a 'cation'.

Sodium chloride (NaCl) is formed by the combination of sodium (Na^+) and chloride (Cl^-) ions. Sodium and chloride ions are oppositely charged. Hence, they are held by a strong electrostatic force of attraction in sodium chloride compound. **But why do they react or combine with each other?** This can be explained by considering the formation of sodium chloride.

This representation of elements with valence electrons as dots around elements is referred to as **Electron Dot structures** for elements.

Do you know what type of a compound sodium chloride is? Sodium chloride is an ionic compound.

Ionic compounds:

These are compounds that are formed by the transfer of electrons. In other words, these are compounds that are made up of ions.

The bonding in such compounds is called **ionic bonding or electrovalent bonding**. This type of bonding is also known as **electrostatic bonding** as the forces that hold the ions together are electrostatic in nature. The transfer of electrons always takes place from a metal to a non-metal. Thus, metals and non-metals combine with each other to attain a noble gas configuration.

We know that inert (noble) gases are very stable and almost unreactive. This is because of their stable electronic configuration in which their valence shell is complete. Hence, they do not take part in the formation of ionic compounds. The given table lists some elements with their electronic configurations.

S.No.	Type of element		Element	Symbol	Atomic number	Electronic configuration K L M N	Number of valence electrons
1.	**Noble gases**	1.	Helium	He	2	2	2
		2.	Neon	Ne	10	2, 8	8
		3.	Argon	Ar	18	2, 8, 8	8
2.	**Metals**	1.	Sodium	Na	11	2, 8, 1	1
		2.	Potassium	K	19	2, 8, 8, 1	1
		3.	Magnesium	Mg	12	2, 8, 2	2
		4.	Calcium	Ca	20	2, 8, 8, 2	2
		5.	Aluminium	Al	13	2, 8, 3	3
3.	**Non-metals**	1.	Nitrogen	N	7	2, 5	5
		2.	Phosphorus	P	15	2, 8, 5	5
		3.	Oxygen	O	8	2, 6	6
		4.	Sulphur	S	16	2, 8, 6	6
		5.	Fluorine	F	9		7

		6.	Chlorine	Cl	17	2, 7 2, 8, 7	7

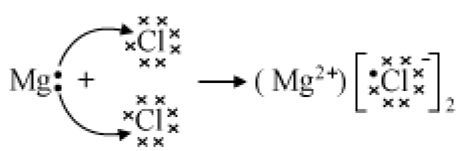
Let us now see the formation of magnesium chloride, which is also an ionic compound.

The atomic number of magnesium is 12. Thus, its electronic configuration is 2, 8, 2. Since it contains two more electrons than a stable noble gas configuration, it loses these two electrons to form Mg^{2+}. On the other hand, the atomic number of chlorine is 17. Thus, its electronic configuration is 2, 8, 7. It requires one more electron to complete its octet. For this, two chlorine atoms accept two electrons that were lost by Mg atom to form two chloride (Cl^-) ions. The chemical equations involved in the process are given below:

$$\begin{array}{ccccc} Mg & \rightarrow & Mg^{2+} & + & 2e^- \\ \text{Magnesium} & & \text{Magnesium ion} & & \text{Electron} \\ (2, 8, 2) & & (2, 8) & & \end{array}$$

$$\begin{array}{ccccc} 2Cl & + & 2e^- & \rightarrow & 2Cl^- \\ \text{Chlorine} & & \text{Electron} & & \text{Chloride ion} \\ (2, 8, 7) & & & & (2, 8, 8) \end{array}$$

The reaction between magnesium and chlorine can be represented as follows:



On the similar basis, the formation of sodium chloride (NaCl) and calcium oxide (CaO) is depicted in the table below:

Compound	**Formation**

Sodium chloride (NaCl)	$\underset{(2,8,1)}{Na} \longrightarrow \underset{(2,8)}{Na^+} + e^-$ $\underset{(2,8,7)}{Cl} + e^- \longrightarrow \underset{(2,8,8)}{Cl^-}$ $Na^+ + Cl^- \longrightarrow [Na^+][Cl^-]$
Calcium oxide (CaO)	$\underset{(2,8,8,2)}{Ca} \longrightarrow \underset{(2,8,8)}{Ca^{2+}} + 2e^-$ $\underset{(2,6)}{O} + 2e^- \longrightarrow \underset{(2,8)}{O^{2-}}$ $Ca^{2+} + O^{2-} \longrightarrow [Ca^{2+}][O^{2-}]$

Potassium oxide (K_2O) is also an ionic compound. It is made of two potassium atoms and one oxygen atom.

Can you draw the Electron Dot structure of potassium and oxygen atoms? Can you show the formation of potassium oxide?

Let us now try to find out the properties of ionic compounds by performing the following activities.

1) Take samples of sodium chloride, potassium iodide, and barium chloride and observe their physical state.

2) After that, take a small amount of a sample on a metal spatula and heat it directly on a flame. Observe what happens to the sample.

3) Now, try to dissolve each sample in water, petrol, and kerosene and observe the solubility of compounds.

4) Now, take a container and fill it with distilled water. Take two electrodes and place them in water. Then, connect the electrodes to a bulb and a battery through electric wires (as shown in **figure 1)**. When the switch is closed, the bulb will not glow as distilled water does not conduct electricity. Now, instead of distilled water, take a solution of an ionic compound and observe.

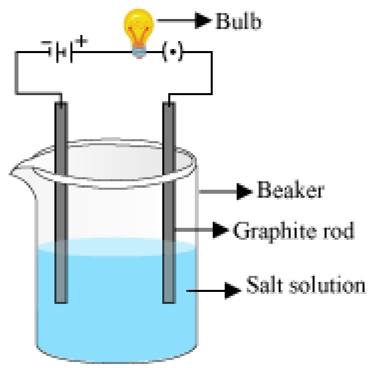


Figure 1: Conductivity of salt

We will observe that

- all compounds are solids
- all have high melting and boiling points
- all samples are soluble in water but insoluble in kerosene and petrol
- the solution of all samples can conduct electricity

When the switch is closed, the bulb starts glowing. This shows that solutions of ionic compounds conduct electricity.

Hence, we can summarize the properties of ionic compounds as follows:

Ionic compounds are hard and brittle crystalline solids: The electrostatic force holding the ions present in ionic compounds are very strong. Therefore, these compounds are quite hard, as they are made up of small crystals.

Ionic compounds have high melting and boiling points: A lot of energy is required to overcome the strong electrostatic force of attraction, which holds the ions present in ionic compounds together. Thus, these compounds have high melting and boiling points.

Table 1: Melting and boiling points of some ionic compounds

Ionic compound	Melting point (K)	Boiling point (K)
$NaCl$	1074	1686
$LiCl$	887	1600
$CaCl_2$	1045	1900

CaO	2850	3120
$MgCl_2$	981	1685

Ionic compounds dissolve only in polar solvents: Ionic compounds are polar in nature due to the presence of opposite charges. Therefore, these compounds dissolve only in polar solvents such as water. These compounds are insoluble in organic solvents such as kerosene, alcohol, and petrol.

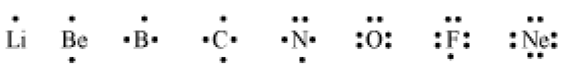
Ionic compounds conduct electricity in a solution or molten state: Ionic compounds consist of small ions, which can conduct electricity.

Chemical Bonding

A chemical bond is an attractive force which holds various constituents (such as atoms, ions) together in different chemical species.

Kossel-Lewis Approach to Chemical Bonding

- Lewis postulated that atoms attain the stable octet when they are chemically bonded.
- **Lewis symbols**
- Notations to represent valence electrons in an atom
- Example:



- Significance of Lewis symbols − The number of dots represents the number of valence electrons.
- Octet rule- Atoms tend to gain, lose or share electrons so as to have eight electrons in their valence shells.
- **Lewis dot structure**

Representation of molecules and ions in terms of the shared pairs of electrons and the octet rule

Steps to writing Lewis dot structure:

- Add the valence electrons of the combining atoms. This will give the total number of electrons required to write the structure.

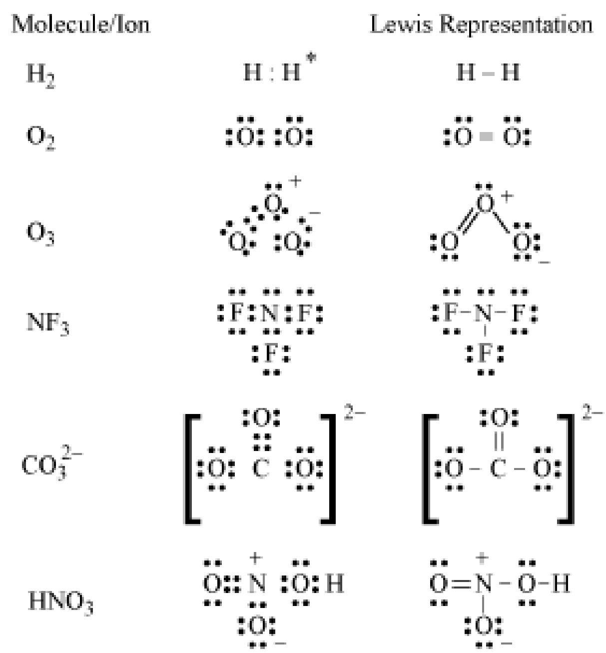
- One negative charge means the addition of an electron. Similarly, one positive charge implies the removal of an electron from the total number of electrons.

- The chemical symbol of the atoms and the skeletal structure of the compound should be known. Then, distribute the total number of electrons as bonding shared pairs between the atoms in proportion to the total bonds.

- The least electronegative atom occupies the central position of the molecule/ion. For example in NF_3, nitrogen occupies the central position whereas the three fluorine atoms occupy the terminal positions.

- When the shared pairs of electrons have been accounted for single bonds, utilise the remaining electron pairs for either multiple bonding or count them as lone pairs. Here, the basic requirement is that each bonded atom gets an octet of electrons.

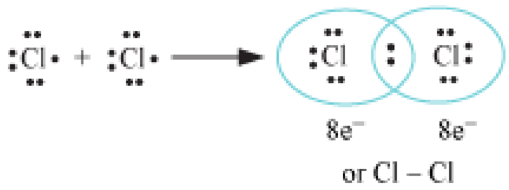
- Lewis representation of some molecules



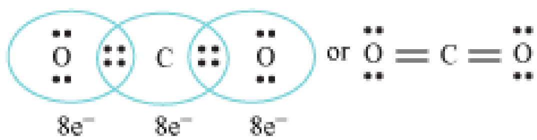
(*- Each hydrogen atom attains the electronic configuration of helium i.e. a duplet of electrons)

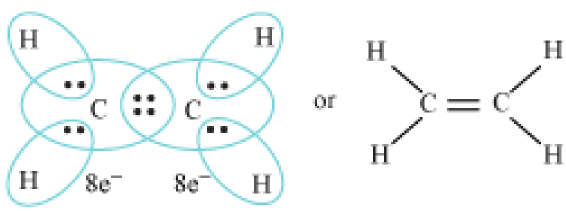
- **Covalent bond**

- Single covalent bond − Sharing of one electron pair

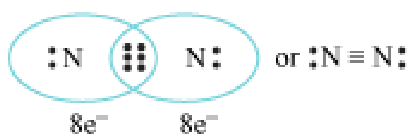


- Double bond − Sharing of two electron pairs

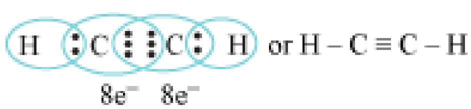




- Triple bond − Sharing of three electron pairs



N_2 molecule



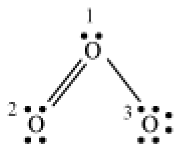
C_2H_2 molecule

- **Formal Charge**

$$\left[\begin{array}{l}\text{Formal charge (F.C)}\\ \text{on an atom in a}\\ \text{Lewis structure}\end{array}\right] = \left[\begin{array}{l}\text{Total number of}\\ \text{valence electrons}\\ \text{in the free atom}\end{array}\right] - \left[\begin{array}{l}\text{Total number of}\\ \text{nonbonding (lone}\\ \text{pair electrons)}\end{array}\right] - \frac{1}{2}\left[\begin{array}{l}\text{Total number of}\\ \text{bonding (shared}\\ \text{electrons)}\end{array}\right]$$

- Example:

Lewis structure of O_3 is



F.C. on the O-1 atom $= 6 - 2 - \frac{1}{2}(6) = +1$

F.C. on the O-2 atom $= 6 - 4 - \frac{1}{2}(4) = 0$

F.C. on the O-3 atom $= 6 - 6 - \frac{1}{2}(2) = -1$

- Smaller the formal charge on the atoms, lower is the energy of the structure.
- The concept of formal charge is based on covalent bonding in which electron pairs are equally shared by neighbouring atoms.
- Limitations of the octet rule:
- Incomplete octet of the central atom

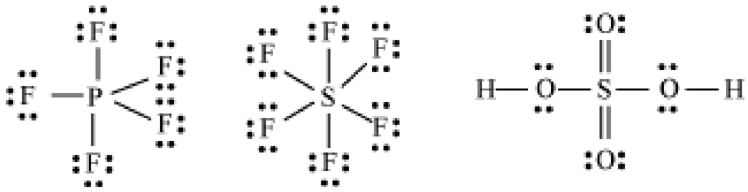
  Examples: LiCl, BeH_2, BCl_3

- Odd electron molecules

  Examples: NO, NO_2

- Expanded octet

  Examples: PF_5, SF_6, H_2SO_4



PF_5 — 10 electrons around the P atom

SF_6 — 12 electrons around the S atom

H_2SO_4 — 12 electrons around the S atom

Some other drawbacks of octet rule:

- It is based upon chemical inertness of noble gases. However, some noble gases can combine to form compounds such as XeF_2, KrF_2, $XeOF_2$, etc.
- It does not account for shape of molecules
- It does not explain the relative stability of molecules

Conditions for Formation of Covalent Bond

- Presence of four or more electrons in the outermost shell of an atom (exception H, Be, B and Al)
- High electronegativity of both the atoms
- High electron affinity for both the atoms
- High ionisation energy of both the atoms
- Electronegativity difference between combining atoms should be zero or very low

Formation of Some Covalently Bonded Molecules

Compound	Molecule	Type and Number of Covalent Bonds
Hydrogen (H_2)	H−H	One single bond
Chlorine (Cl_2)	Cl−Cl	One single bond
Nitrogen (N_2)	N≡N	One triple bond
Water (H_2O)	H−O−H	Two single bonds between O and H
Ammonia (NH_3)	H –N– H | H	Three single bonds between N and H

Carbon tetrachloride (CCl_4)	Cl | Cl –C– Cl | Cl	Four single bonds between C and Cl
Methane (CH_4)	H | H –C– H | H	Four single bonds between C and H

Difference between Properties of Ionic and Covalent Compounds

Ionic Compounds	Covalent Compounds
The constituent particles are ions.	The constituent particles are molecules.
They exist as hard solids.	They exist as gases, liquids or soft solids.
They have high melting and boiling points	They have low melting and boiling points.
They are good conductors of electricity in the aqueous or molten state.	They do not conduct electricity.
They ionise in solution and behave as electrolytes.	Only polar compounds form ions in aqueous solutions.
They undergo dissociation.	They do not undergo dissociation.
They are soluble in water.	They are soluble only in organic solvents.
They undergo fast chemical reactions.	They undergo slow chemical reactions.

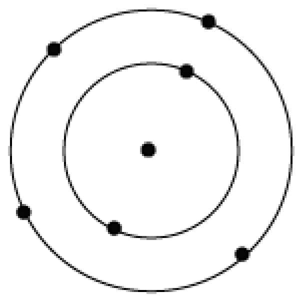
Covalent and Coordinate Bond

We know that a majority of substances used by us daily, from paper and plastics to coal and petrol, are all made up of carbon. Food grains, pulses, medicines, cotton, synthetic fibres, wood, etc. are all made up of carbon. Carbon is also a major part of all living things. In air, it is present as carbon dioxide and comprises around 0.03% of the total atmosphere.

Let us study about carbon and its bonding in its compound in more detail.

Carbon is a non-metal having the symbol '**C**' and atomic number **six**. Since the atomic number of carbon is six, its electronic configuration is 2, 4. This means that carbon contains two electrons in K shell and 4 electrons in L shell (outermost shell). Hence, it has four electrons in its valence shell.

Since carbon has four electrons in its valence shell, it requires four more electrons to complete its octet. Therefore, it is a tetravalent element.



In order to complete its octet i.e., to attain its noble gas configuration and to stabilise itself, carbon can:

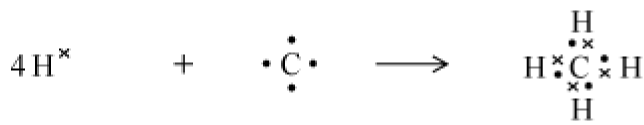
- Either lose four electrons to form $\mathbf{C^{4+}}$ or gain four electrons to form $\mathbf{C^{4-}}$. This, however, requires a lot of energy and would make the system unstable.

- Therefore, carbon completes its octet by sharing its four electrons with the other carbon atoms or with atoms of other elements.

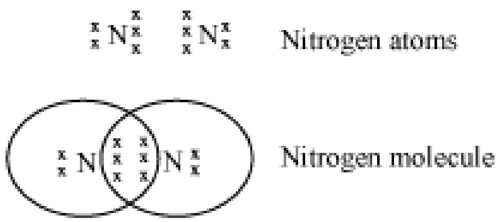
The bonds that are formed by sharing electrons are known as covalent bonds. Covalently bonded molecules have strong intermolecular forces, but intramolecular forces are weak.

Carbon has four valence electrons and requires four more electrons to complete its octet. Therefore, it is capable of bonding with four other atoms of carbon or atoms of other elements having a valency of 1.

For example, the simplest molecule, methane, can be formed with hydrogen (H) atoms that have only one electron in its K shell. To attain the noble gas configuration, it combines with four hydrogen atoms as shown in the figure.



Nitrogen has an atomic number of 7. In order to attain an octet, each nitrogen atom in a molecule of nitrogen contributes three electrons, thereby giving rise to forming three shared pairs of electrons. This is said to constitute a triple bond between the two atoms. The electron dot structure of N_2 and its triple bond can be depicted as follows.:



Now, let us study the properties of carbon covalent compounds.

- Covalent bonds are made by atoms by sharing their electrons. Formation of ions does not take place in this process. In addition, these compounds do not have any extra electrons. Hence, covalent compounds are non-conductors of electricity.

- As all organic compounds contain covalent bonds, they also have low melting and boiling points. This becomes evident from the following data.

Compound	**Melting point (K)**	**Boiling point (K)**
Acetic acid	290	391
Chloroform	209	334
Ethanol	156	351
Methane	90	111

Also, from the above data, it can be inferred that the forces of attraction between the carbon molecules in carbon compounds is not very strong.

- Because of their low melting and boiling points, these compounds mostly exist as liquids or gases at room temperature.

A covalent bond formed between two different atoms, with different electronegativities is known as **polar covalent bond.**
For example when a Covalent bond is formed between H and Cl, it is polar in nature because Cl is more electronegative than H atom. Therefore, electron cloud is shifted towards Cl atom. As a result, a partial negative charge appears on Cl atom and, an equal positive charge on H atom.

$\mathbf{H^{\delta+}}$**----**$\mathbf{Cl^{\delta-}}$

A covalent bond formed between two like atoms, is known as **Non-polar bond**. Since difference of electronegativity is zero, therefore, both atoms attract electron pair equally and no charge appears on any atom, and the whole molecule becomes neutral.
For e.g. **H------H**

Non-polar covalent bond:

Let us now compare the properties of electrovalent compounds with those of covalent compounds.

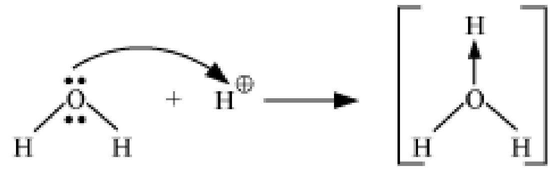
Electrovalent compounds	Covalent compounds
These types of compounds are formed when the atoms containing one to three valence electrons donate electrons. On the other hand, the atoms containing five to seven electrons accept thosethe donated electrons. Thus, both the participating atoms acquire the nearest noble gas configuration.	These types of compounds are formed when the participating atoms mutually contribute one electron each to form an electron pair which that is shared by them in such a manner that both of them attain the nearest noble gas configuration.
They are usually crystalline solids.	They are usually gases or liquids or soft solids.
They are good conductors of electricity.	They are poor conductors of electricity.
They usually have high melting and boiling points.	They usually have low melting and boiling points.
They are soluble in polar solvents but insoluble in organic solvents.	They are soluble in organic solvents but insoluble in polar solvents.

Coordinate Bond

It is formed when the shared pair of electrons is provided by one of the two atoms and shared by both. Examples of coordinate-bonded compounds are:

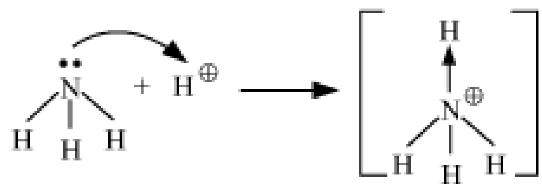
- Hydronium ion
- Ammonium ion

Hydronium ion: The compound giving hydronium ion is water. One lone pair on the oxygen atom of water molecule is shared with the hydrogen ion. In this combination, the hydrogen ion accepts a pair of electrons and distributes the charge over the entire hydronium radical.



Coordinate bond formation in
hydronium ion

Ammonium ion: One lone pair on the nitrogen atom of ammonia molecule is shared with the hydrogen ion. In this combination, the hydrogen ion accepts a pair of electrons and distributes the charge across the entire ammonium radical.



Coordinate bond formation in ammonium ion

Conditions for formation of coordinate bond

- Presence of at least one lone pair of electrons on any of the two atoms. This atom acts like a donor.
- Shortage of a lone pair of electron on the second atom. This atom acts like an acceptor.

Formation of Hydronium (H_3O^+) Ion

- Based on ionisation of water
- One water molecule contains two hydrogen atoms covalently bonded with one oxygen atom.
- Oxygen atom has two lone pairs of electrons for donation.
- Addition of an acid to water, results in dissociation of the water molecules as shown below:

- $H_2O \xrightarrow{HA} H^+ + OH^-$

H^+ ions accepts an electron pair from oxygen atom of another water molecule forming coodinately bonded hydronium ion:

- $H^{+} + H-\ddot{\underset{..}{O}}-H \longrightarrow \left[\begin{array}{ccccc} & & \cdot\cdot & & \\ H & - & O & - & H \\ & & \downarrow & & \\ & & H & & \end{array}\right]^{+}$
- Hydronium ion is also referred to as hydrated proton.
- Similarly, water molecules undergo self-ionisation where one molecule acts as an acid:
  - $H_2O + H_2O \rightleftharpoons H_3O^{+} + OH^{-}$