9th Science Lesson 13 Notes in English

13] Chemical Bonding

Introduction

- We already know that atoms are **the building blocks of matter**.
- Under normal conditions no atom exists as an independent (single) entity in nature, except noble gases.
- However, a group of atoms is found to exist together as one species.
- Such a group of atoms is called **molecule**.
- Obviously there should be a force to keep the constituent atoms together as the thread holds the flowers together in a garland.
- This attractive force which holds the atoms together is called a **bond**.

Figure. 13.1 Flowers held together by thread

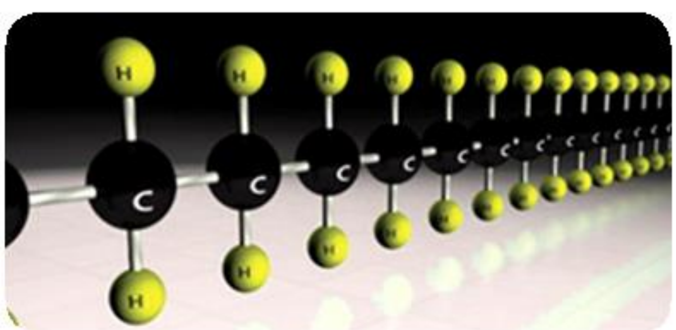


Figure. 13.2 Atoms held together by bond

- **A chemical bond** may be defined as **the force of attraction** between **the atoms** that binds them together as a unit called **molecule**
- In this unit, we will study about **Kossel- Lewis approach** to chemical bonds, **Lewis dot structure** and different types of reactions.

Kossel – Lewis approach to chemical bonds

Octet rule

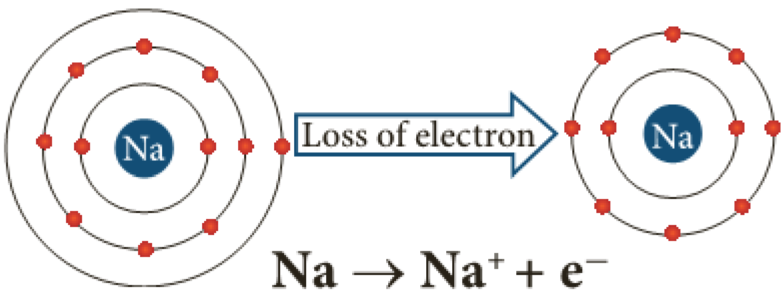
- Atoms of various elements combine together in different ways to form chemical compounds. This phenomenon raised many questions.
  1) Why do atoms combine?
  2) How do atoms combine?
  3) Why do certain atoms combine while others do not?
- To answer such questions different theories have been put forth from time to time and one of such theories which explained the formation of molecules is **Kossel-Lewis theory**.
- Kossel and Lewis gave successful explanation based upon **the concept of electronic configuration of noble gases** about why atoms combine to form molecules.
- Atoms of noble gases have little or no tendency to combine with each other or with atoms of other elements.
- This means that these atoms must be having **stable electronic configurations**.
- The electronic configurations of noble gases are given in Table 13.1.

Table 13.1 The electronic configurations of noble gases

Name of the element	Atomic number	Shell electronic configuration
Helium (He)	2	2
Neon (Ne)	10	2,8
Argon (Ar)	18	2,8,8
Krypton (Kr)	36	2,8,18,8
Xenon (Xe)	54	2,8,18,18,8
Radon (Rn)	86	2,8,18,32,18,8

- **Except Helium**, all other noble gases have eight electrons in their valence shell.
- Even helium has its valence shell completely filled and hence no more electrons can be added.
- Thus, by having **stable valence shell electronic configuration**, the noble gas atoms neither have any tendency to gain nor to lose electrons and hence their valency is zero.
- They are so inert that they even do not form diatomic molecules and exist as **monoatomic gaseous atoms.**
- Based on the noble gas electronic configuration, **Kossel** and **Lewis proposed a theory** in 1916 to explain chemical combination between atoms and this theory is known as **'Electronic theory of valence'** or **Octet rule.**

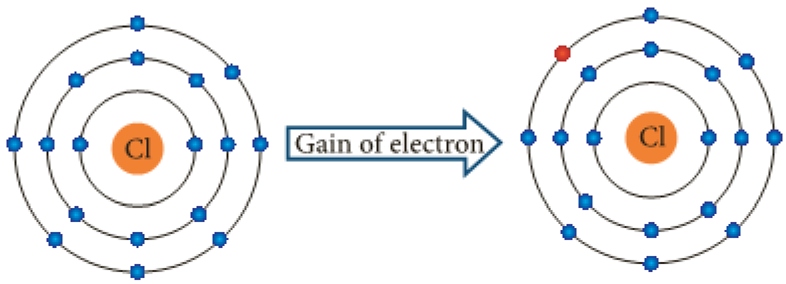
- According to this, atoms of all elements, other than inert gases, combine to form molecules because they have incomplete valence shell and tend to attain a stable electronic configuration similar to noble gases.
- Atoms can combine either by transfer of valence electrons from one atom to another or by sharing of valence electrons in order to achieve the stable outer shell of eight electrons.
- The tendency of atoms to have eight electrons in the valence shell is known as the **'Octet rule'** or the **'Rule of eight'**
- For example, sodium with atomic number 11 will readily loose one electron to attain neon's stable electronic configuration (Figure 13.3).



$$Na \rightarrow Na^{+} + e^{-}$$

Figure. 13.3 Formation of sodium ion

- Similarly, chlorine has electronic configuration 2,8,7.
- To get the nearest noble gas (i.e. argon) configuration, it need one more electron.
- So, chlorine readily gains one electron from other atom and obtains **stable electronic configuration** (Figure 13.4).



$$Cl + e^{-} \rightarrow Cl^{-}$$

Figure. 5.4 Formation of chloride ion

- Thus elements tend to have stable valence shell (eight electrons) either by losing or gaining electrons.
- Which atoms tend to lose electrons?
- Which are tend to gain electrons?

Table 13.2 Unstable electronic configuration

Element	Atomic number	Electron distribution	Valence electrons
Boron	5	2, 3	3
Nitrogen	7	2, 5	5
Oxygen	8	2,6	6
Sodium	11	2, 8, 1	1

- Atoms that have 1, 2, 3 electrons in their valence shell tend to lose electrons whereas atoms having 5, 6, 7 valence electrons tend to gain electrons.

Lewis dot structure

Table 13.3 Lewis dot structure

Element	Atomic number	Electron distribution	Valence electrons	Lewis dot structure
Hydrogen	1	1	1	H•
Helium	2	2	2	He:
Beryllium	4	2, 2	2	•Be•
Carbon	6	2, 4	4	•Ċ• (with dot below)
Nitrogen	7	2, 5	5	•N̈• (with dot below)
Oxygen	8	2,6	6	•Ö: (with dot below)

- When atoms combine to form compounds, their valence electrons involve in bonding.
- Therefore, it is helpful to have a method to depict the valence electrons in the atoms.
- This can be done using **Lewis dot symbol method**.
- The Lewis dot structure or electron dot symbol for an atom consists of the symbol of the element surrounded by dots representing **the electrons of the valence shell of the atom.**
- **The unpaired electron** in the valence shell is represented by a single dot whereas the paired electrons are represented by a pair of dots.
- Symbols other than dots, like crosses or circles may be used to differentiate the electrons of the different atoms in the molecule.

Types of chemical bond

- All the elements have different **valence shell electronic configuration**.
- So the way in which they combine to form compounds also differs.
- Hence, there are different types of **chemical bonding** possible between atoms which make the molecules.
- Depending on **the type of bond**, they show different characteristics or properties.
- Such types of bonding, that are considered to exist in molecules, are categorized as shown Figure 13.5.
- Among these, let us learn about the **Ionic bond, Covalent bond** and **Coordinate bond** in this chapter and other types of bond in the higher classes.

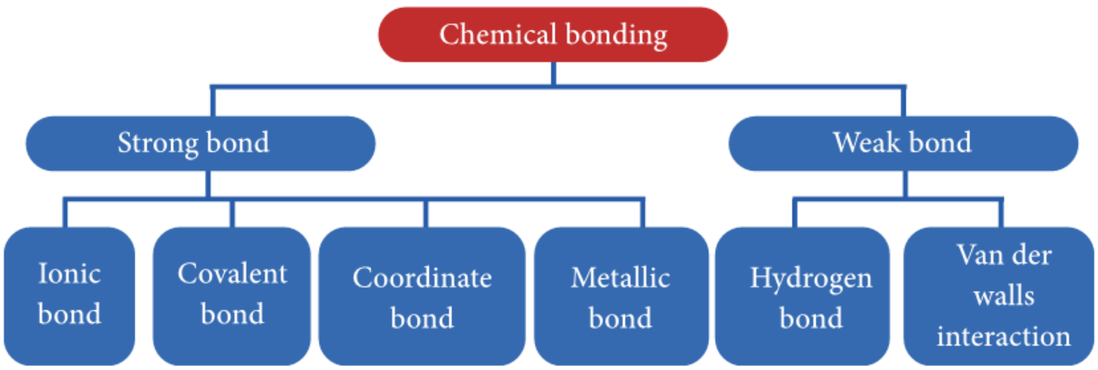


Figure. 13.5 Classification of chemical bond

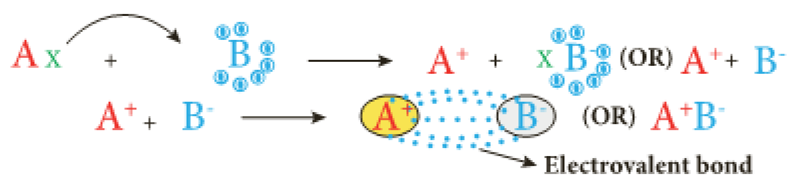
Ionic (or) Electrovalent bond

- An ionic bond is a chemical bond formed by the electrostatic attraction between **positive** and **negative ions.**
- The bond is formed between two atoms when one or more electrons are transferred from the valence shell of one atom to the valence shell of the other atom.
- The atom that loses electrons will form **a cation** (positive ion) and the atom that gains electrons will form **an anion** (negative ion).
- These oppositely charged ions come closer to each other due to electrostatic force of attraction and thus form an ionic bond.
- As the bond is between the ions, it is called **Ionic bond** and the attractive forces being electrostatic, the bond is also called **Electrostatic bond**.
- Since the valence concept has been explained in terms of electrons, it is also called as **Electrovalent bond.**
- Formation of ionic bond Let us consider two atoms A and B. Let atom A has one electron in excess and atom B has one electron lesser than the stable octet electronic configuration.
- If atom A transfer one electron to atom B, then both the atoms will acquire stable octet electronic configuration.
- As the result of this electron transfer, atom A will become positive ion (cation) and atom B will become **negative ion** (anion).

- These oppositely charged ions are held together by electrostatic force of attraction which is called **Ionic bond** or **Electrovalent bond**.
- In general, ionic bond is formed between a metal and non-metal. The compounds containing **ionic bonds** are called **ionic compounds**.
- Elements of Group 1 and 2 in periodic table, i.e. alkali and alkaline earth metals form ionic compounds when they react with non-metals.

Illustration 1 – Formation of ionic bonding in sodium chloride (NaCl)

- The atomic number of Sodium is 11 and its electronic configuration is 2, 8, 1.
- It has one electron excess to the nearest stable electronic configuration of a **noble gas - Neon.**



- So sodium has a tendency to lose one electron from its outermost shell and acquire a stable electronic configuration forming **sodium cation** (Na+).
- The atomic number of chlorine is 17 and its electronic configuration is 2, 8, 7.
- It has one electron less to **the nearest stable electronic configuration of a noble gas - Argon.**

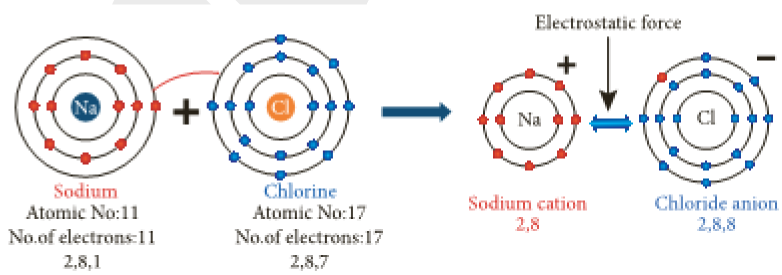


Figure. 13.6 Formation of ionic bond in sodium chloride

- So chlorine has a tendency to gain one electron to acquire a stable electronic configuration forming **chloride anion** (Cl-).
- When an atom of sodium combines with an atom of chlorine, an electron is transferred from sodium atom to chlorine atom forming sodium chloride molecule thus both the atoms attain stable octet electronic configuration.

Illustration 2 – Formation of ionic bond in magnesium chloride (MgCl2)

- The atomic number of magnesium is 12 and the electronic configuration is 2, 8, 2.
- It has two electron excess to the nearest stable electronic configuration of a **noble gas - Neon.**
- So magnesium has a tendency to lose two electrons from its **outermost shell** and acquire a **stable electronic configuration** forming magnesium cation (Mg2+).
- As explained earlier two chlorine atoms will gain two electrons lost by the magnesium atom forming magnesium chloride molecule (MgCl2) as shown in Figure 13.7.

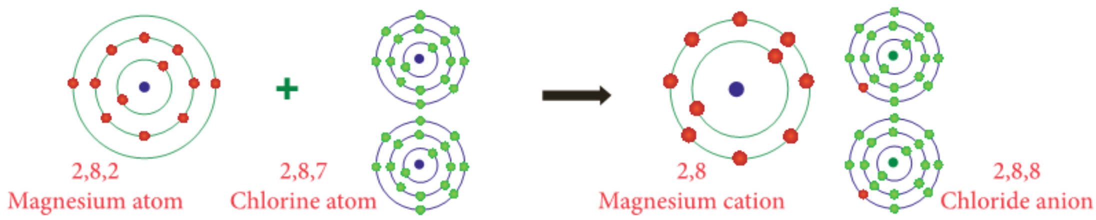


Figure. 13.7 Formation of ionic bond in magnesium chloride

Characteristics of Ionic compounds

- The nature of bonding between the atoms of a molecule is the primary factor that determines the properties of compounds.
- By this way, in ionic compounds the atoms are held together by a **strong electrostatic force** that makes the compounds to have its characteristic features as follows:

Physical state:

- These compounds are formed because of the strong electrostatic force between cations and anions which are arranged in a **well-defined geometrical pattern**.
- Thus ionic compounds are **crystalline solids** at room temperature.

Electrical conductivity:

- Ionic compounds are crystalline solids and so their ions are tightly held together.
- The ions, therefore, **cannot move freely**, and they **do not conduct electricity** in solid state.
- However, in molten state their **aqueous solutions conduct electricity**.

Melting point:

- The strong electrostatic force between **the cations** and **anions** hold the ions tightly together, so very high energy is required to separate them.
- Hence **ionic compounds** have high melting and boiling points.

Solubility:

- Ionic compounds are **soluble in polar solvents** like water.
- They are **insoluble** in **non-polar solvents like benzene** (C6H6), **carbon tetra chloride** (CCl4).

Density, hardness and brittleness:

- Ionic compounds have **high density** and they are **quite hard** because of the **strong electrostatic force** between the ions. But they are **highly brittle**.

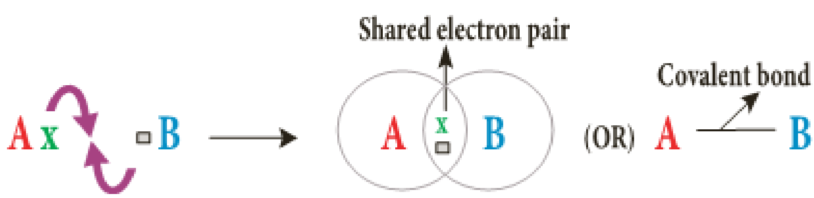
Reactions:

- Ionic compounds **undergo ionic reactions** which are practically **rapid and instantaneous.**

Covalent bond

- Atoms can combine with each other by **sharing the unpaired electrons** in their outermost shell.
- Each of the two combining atoms contributes one electron to the electron pair which is needed for the bond formation and has equal claim on the shared electron pair.
- According to **Lewis concept**, when two atoms form a covalent bond between them, each of the atoms attains **the stable electronic configuration** of the nearest noble gas.
- Since the covalent bond is formed because of the sharing of electrons which become common to both the atoms, it is also called as **atomic bond**.

Formation of Covalent bond



vFigure. 13.8 Schematic representation of covalent bond

- Let us consider two atoms A and B.
- Let atom A has one **valence electron** and atom B has **seven valence electrons**.
- As these atoms approach nearer to each other, each atom contributes one electron and the resulting electron pair fills **the outer shell of both the atoms**.
- Thus both the atoms acquire a completely filled **valence shell electronic configuration** which leads to stability.

More to Know

Covalent bonds are of three types:

1. Single covalent bond represented by a line (—) between the two atoms. Eg. H—H
2. Double covalent bond represented by a double line (=) between the two atoms. Eg. O=O
3. Triple covalent bond represented by a triple line (≡) between the two atoms. Eg. N ≡ N

Illustration 1 – Formation of hydrogen molecule (H2)

- Hydrogen molecule is formed by **two hydrogen atoms.**
- While forming the molecule, both hydrogen atoms contribute one electron each to the shared pair and both **atoms acquire stable** and **completely filled electronic configuration** (**resembly He**).

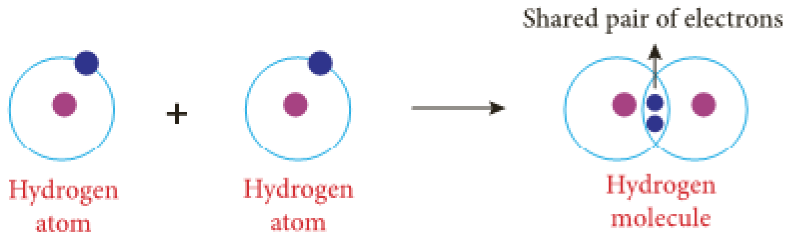


Figure. 13.9 Formation of covalent bond in H_2 molecule

Illustration 2 – Formation of chlorine molecule (Cl2)

- **Chlorine molecule** is formed by two chlorine atoms.
- Each chlorine atom has seven valence electrons (2,8,7).

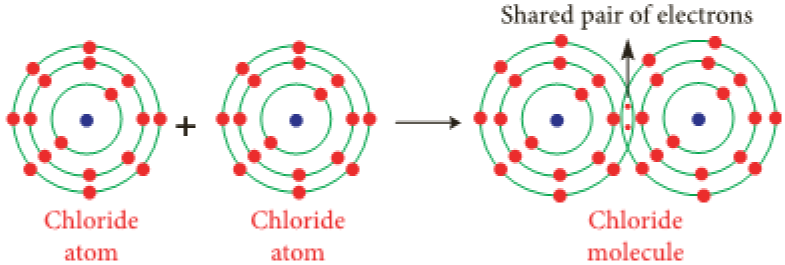


Figure. 13.10 Formation of covalent bond in Cl_2 molecule

- These two atoms achieve a **stable completely filled electronic configuration** (octet) by sharing a pair of electrons.

Illustration 3 – Formation of methane molecule (CH4)

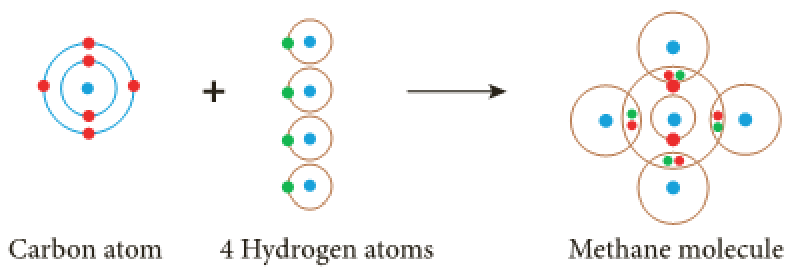


Figure. 13.11 Formation of covalent bond in methane molecule

- **Methane molecule** is formed by the combination of one carbon and four hydrogen atoms.
- The carbon atom has **four valence electrons** (2, 4).
- These four electrons are shared with four atoms of hydrogen to achieve **a stable electronic configuration (octet)** by sharing a pair of electrons.

Illustration 4 – Formation of oxygen molecule (O2)

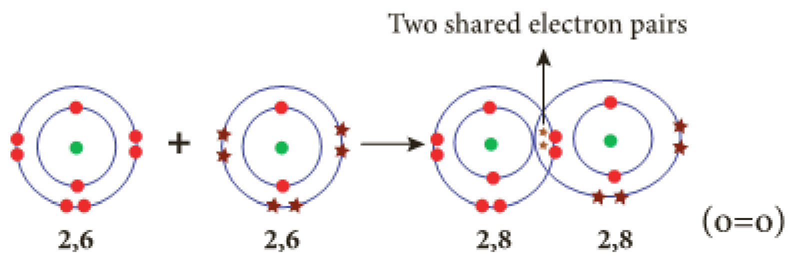


Figure. 13.12 Formation of covalent bond in oxygen molecule

- Oxygen molecule is formed by two oxygen atoms. Each oxygen atom has six valence electrons (2, 6).
- These two atoms achieve a **stable electronic configuration (octet)** by sharing two pair of electrons.
- Hence a double bond is formed in between the two atoms.

Illustration 5 – Formation of nitrogen molecule (N2)

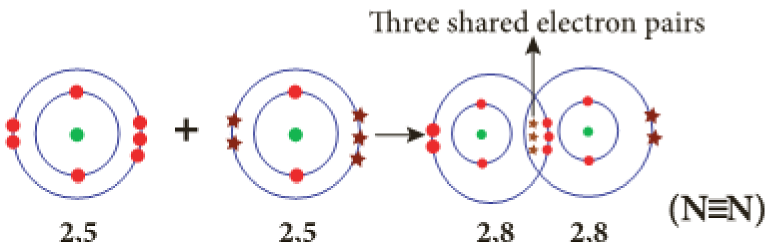


Figure. 13.13 Formation of covalent bond in nitrogen molecule

- **Nitrogen molecule** is formed by two nitrogen atoms. Each nitrogen atom has five valence electrons (2, 5).
- These two atoms achieve a stable completely filled **electronic configuration** (octet) by sharing three pair of electrons.
- Hence a **triple bond** is formed in between the two atoms.

Characteristics of Covalent compounds

- As said earlier, **the properties of compounds** depend on the nature of bonding between their constituent atoms.
- So the compounds containing covalent bonds possess different characteristics when compared to **ionic compounds**.
- **Physical state**: Depending on force of attraction between covalent molecule the bond may be weaker or stronger.
- Thus covalent compounds exists in gaseous, liquid and solid form. Eg. Oxygen-gas; **Water-liquid: Diamond-solid.**
- **Electrical conductivity**: Covalent compounds do not contain charged particles (ions), so they are bad conductors of electricity.
- **Melting point**: Except few covalent compounds (Diamond, Silicon carbide), they have relatively low melting points compared to ionic compounds.
- **Solubility**: Covalent compounds are readily soluble in non-polar solvents like benzene (C6H6), carbon tetra chloride (CCl4). They are insoluble in polar solvents like water.
- **Hardness and brittleness**: Covalent compounds are neither hard nor brittle. But they are soft and waxy.

- **Reactions**: Covalent compounds undergo molecular reactions in solutions and these reactions are slow.

Fajan's Rule:

- As we know, a metal combines with a **non- metal through ionic bond**.
- The compounds so formed are called **ionic compounds**.
- A compound is said to be ionic when the charge of **the cation and anion** are completely separated.
- But in 1923, **Kazimierz Fajans** found, through his **X-Ray crystallographic** studies, that some of the ionic compounds show covalent character.
- Based on this, he formulated a set of rules to predict whether a chemical bond is ionic or covalent.
- Fajan's rules are formulated by considering **the charge of the cation** and **the relative size of the cation and anion.**
- When the size of the cation is small and that of anion is large, the bond is of more covalent character
- Greater the charge of the cation, greater will be the covalent character

Ionic	Covalent
Low positive charge	High positive charge
Large cation	Small cation
Small anion	Large anion

- For example, in sodium chloride, low positive charge (+1), a fairly large cation and relatively small anion make the charges to separate completely.
- So it is **ionic**.
- In **aluminium triiodide**, higher is the positive charge (+3), larger is the anion and thus no complete charge separation.
- So is **covalent**.
- The following picture depicts the relative charge separation of ionic compounds:

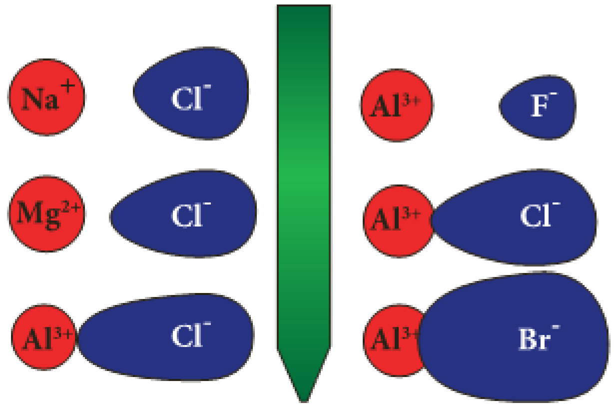


Figure. 13.14 Relatice charge separation of ionic compounds

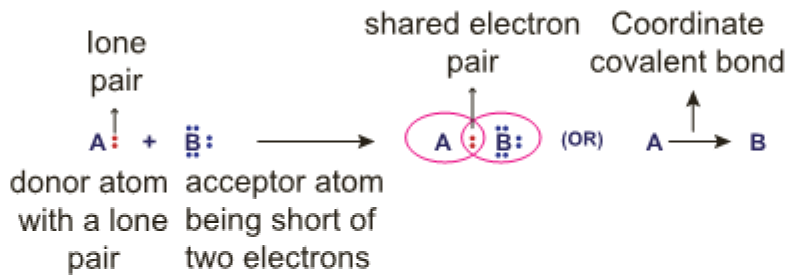
Coordinate covalent bond

- In the formation of normal covalent bond each of the two bonded atoms contribute one electron to form the bond.
- However, in some compounds, the formation of a covalent bond between two atoms takes place by the sharing of two electrons, both of which comes from only one of the combining atoms.
- This bond is called **Coordinate covalent bond** or **Dative bond**.
- Mostly the lone pair of electrons from an atom in a molecule may be involved in the dative bonding.
- The atom which provides the electron pair is called **donor atom** while the other atom which accepts the electron pair is called **acceptor atom**.
- The coordinate covalent bond is represented by an arrow which points from the donor to the acceptor atom.

  an arrow $(\rightarrow)$

- Formation of coordinate covalent bond Let us consider two atoms A and B.
- Let atom A has an unshared lone pair of electrons and atom B is in short of two electrons than the octet in its valence shell.
- Now atom A donates its lone pair while atom B accepts it.
- Thus the lone pair of electrons originally belonged to atom A are now shared by both the atoms and the bond formed by this mutual sharing is called **Coordinate covalent** bond.

  $(A \rightarrow B)$



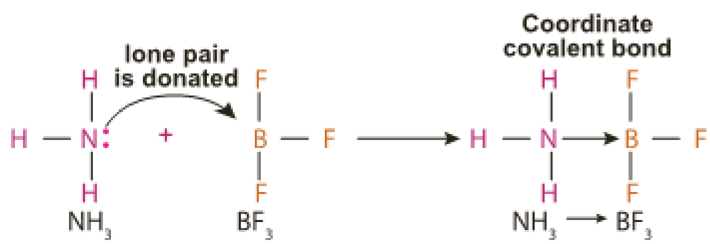
Examples (NH^{4+}, $NH_3 \rightarrow BF_3$)

Illustration 1 – Formation of coordinate covalent bond between

Illustration 1 – Formation of coordinate covalent bond between $NH_3 \rightarrow BF_3$ molecules

- In some cases, the donated pair of electrons comes from a molecule as a whole which is already formed to another acceptor molecule.
- Here the molecule **ammonia (NH3)** gives a lone pair of electrons to **Boron tri fluoride** (BF3) molecule which is electron deficient.
- Thus, a coordinate covalent bond is formed between **NH3 (donor molecule)** and **BF3** (**acceptor molecule**) and is represented by

$NH_3 \rightarrow BF_3$.



Characteristics of coordinate covalent compounds

- The compounds containing coordinate covalent bonds are called coordinate compounds.
- **Physical state**: These compounds exist as gases, liquids or solids.
- **Electrical conductivity**: Like covalent compounds, coordinate compounds also do not contain charged particles (ions), so they are bad conductors of electricity.
- **Melting point**: These compounds have melting and boiling points higher than those of purely covalent compounds but lower than those of purely ionic compounds.

- **Solubility**: Insoluble in polar solvents like water but are soluble in non-polar solvents like benzene, CCl4, and toluene.
- **Reactions**: Coordinate covalent compounds undergo molecular reactions which are slow.

Oxidation, Reduction and Redox reactions

- When an apple is cut and left for sometimes, its surface turns brown.
- Similarly, iron bolts and nuts in metallic structures get rusted.
- Do you know why these are happening? It is because of a reaction called **oxidation**.

Oxidation:

- The chemical reaction which involves **addition of oxygen or removal of hydrogen** or **loss of electrons** is called **oxidation**.

$2\,Mg + O_2 \rightarrow 2\,MgO$ (addition of oxygen)

$CaH_2 \rightarrow Ca + H_2$ (removal of hydrogen)

$Fe^{2+} \rightarrow Fe^{3+} + e^-$ (loss of electron)

Reduction:

- The chemical reaction which involves addition of hydrogen or removal of oxygen or gain of electrons is called **reduction**.

$2\,Na + H_2 \rightarrow 2\,NaH$ (addition of hydrogen)

$CuO + H_2 \rightarrow Cu + H_2O$ (removal of oxygen)

$Fe^{3+} + e^- \rightarrow Fe^{2+}$ (gain of electron)

Redox reactions:

$$2\,PbO + C \rightarrow 2\,Pb + CO_2$$
$$Zn + CuSO_4 \rightarrow Cu + ZnSO_4$$

Oxidation	Addition of oxygen
	Removal of hydrogen
	Loss of electron
Reduction	Removal of oxygen
	Addition of hydrogen
	Gain of electron

- Generally, the oxidation and reduction occurs in the same reaction (simultaneously).
- If one reactant gets oxidised, the other gets reduced.
- Such reactions are called **oxidation-reduction reactions** or **Redox reactions**.

Oxidising agents and Reducing agents

- Substances which have the ability to oxidise other substances are called **oxidising agents.**
- These are also called as **electron acceptors** because they remove electrons from other substances.
- Substances which have the ability to reduce other substances are called **Reducing agents.**
- These are also called as **electron donors** because they donate electrons to other substances.

Example: H_2O_2, MnO_4^-, CrO_3, $Cr_2O_7^{2-}$

Substances which have the ability to reduce other substances are called Reducing agents. These are also called as electron donors because they donate electrons to other substances.

Example: $NaBH_4$, $LiAlH_4$ and metals like Palladium, Platinum.

- Example: NaBH4, LiAlH4 and metals like Palladium, Platinum.

Oxidation reactions in daily life:

- In nature, the oxygen present in atmospheric air oxidises many things, starting from metals to living tissues.
- The shining surface of metals tarnishes due to the formation of respective metal oxides on their surfaces. This is called **corrosion**.
- The freshly cut surfaces of vegetables and fruits turn brown after some time because of the oxidation of compounds present in them.
- The oxidation reaction in food materials that were left open for a long period is responsible for spoiling of food. This is called **Rancidity**.

Oxidation number

- Oxidation number of an element is defined as the total number of electrons that an atom either gains or loses in order to form a chemical bond with another atom.
- Oxidation number is also called **oxidation state**.
- If the oxidation number is positive then it means that the atom loses electron, and if it is negative it means that the atom gains electrons.
- If it is zero then the atom neither gains nor loses electrons.
- The sum of oxidation numbers of all the atoms in the formula for a neutral compound is **ZERO**.
- The sum of oxidation numbers of an ion is the same as the charge on that ion.
- **Negative oxidation number** in a compound of two unlike atoms is assigned to the more **electronegative atom**.

Example:

- Oxidation number of K and Br in KBr molecule is +1 and –1 respectively.
- Oxidation number of N in NH_3 molecule is –3.
- Oxidation number of H is +1 (except hydrides).
- Oxidation number of oxygen in most cases is –2.

Problems on determination of Oxidation Number

ON (Oxidation Number) of neutral molecule is always zero

Illustration 1

Oxidation Number of H and O in H_2O

Let us take ON of H = +1 and ON of O = −2

$2 \times (+1) + 1 \times (-2) = 0$

$(+2) + (-2) = 0$

Thus, ON of H is +1 and ON of O is −2

Illustration 2

Oxidation Number of S in H_2SO_4

Let ON of S be x and we know ON of H = +1 and O = −2

$2 \times (+1) + x + 4 \times (-2) = 0$

$(+2) + x + (-8) = 0$

$x = +6$ Therefore, ON of S is +6

Illustration 3

Oxidation Number of Cr in $K_2Cr_2O_7$

Let ON of Cr be x and we know ON of K = +1 and O = −2

$2 \times (+1) + 2 \times x + 7 \times (-2) = 0$

$(+2) + 2x + (-14) = 0$

$2x = +12$

$x = +6$ Therefore, ON of Cr in $K_2Cr_2O_7$ is +6

Illustration 4

Oxidation Number of Fe in $FeSO_4$

Let ON of Fe be x and we know ON of S = +6 and O = −2

$x + 1 \times (+6) + 4 \times (-2) = 0$

$x + (+6) + (-8) = 0$

$x = +2$ Therefore, ON of Fe in $FeSO_4$ is +2

More to Know

The non-metal

- The number of electrons lost from a metal atom is the valency of the metal and the number of electrons gained by a non- metal is the valency of the non-metal

Valence electrons

- Note that dots are placed one to each side of the letter symbol until all four sides are occupied.

- Then the dots are written two to a side until all valence electrons are accounted for.
- The exact placement of the single dots is immaterial.

Electrovalency

- The number of electrons that an atom of an element loses or gains to form an electrovalent bond is called its **Electrovalency**.

Polar solvents

- Polar solvents contain bonds between atoms with very different electronegativities, such as oxygen and hydrogen.
- Ionic compounds are soluble in polar solvents. Ex: water, ethanol, acetic acid, ammonia
- Non polar solvents contain bonds between atoms with similar electro negativities, such as carbon and hydrogen.
- Covalent compounds are soluble in non- polar solvents. Ex: acetone, benzene, toluene, turpentine

Electronegativity

- Electronegativity is the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.

EXTRA POINTS:

- **Chemical bond**: Force of attraction between the two atoms that binds them together as a unit.
- **Coordinate covalent bond**: Bond formed between atoms by mutual sharing of electrons which are supplied by one atom.
- **Covalent bond**: Bond formed between atoms by the mutual sharing of electrons.
- **Ionic / Electrovalent bond**: Bond formed between cation and anion because of the transfer of electrons from one atom to other atom.
- **Octet rule or Rule of eight**: The tendency of atoms to have eight electrons in the valence shell.
- **Oxidation**: Chemical reaction which involves in the addition of oxygen or removal of hydrogen or loss of electrons.
- **Oxidation number**: The formal charge which an atom has when electrons are counted.
- **Oxidising agents:** Substances which have the ability to oxidise other substances.
- **Redox reaction**: Oxidation and reduction occurs in the same reaction simultaneously.
- **Reducing agents**: Substances which have the ability to reduce other substances.
- **Reduction**: Chemical reaction which involves in the addition of hydrogen or removal of oxygen or gain of electrons.