

## SLO 5: Structure of Molecules

### 5.1 Formation of Chemical Bond

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#### 5.1.1 Explain Duplet And Octet Rules That Help Atoms Achieve Stable Electron Configurations By Forming Bonds

**Duplet Rule:**

This rule applies primarily to first five elements. Atoms achieve stability by having two electrons in their outermost shell, mimicking the electron configuration of helium ( $1s^2$ ).

**Octet Rule:**

This rule states that atoms of elements tend to achieve eight electrons in their valence shell, resembling the stable configuration of noble gases. Atoms can do this by gaining, losing, or sharing electrons.

#### 5.1.2 Define A Chemical Bond And Its Types

A chemical bond is defined as a force of attraction between atoms that holds them together in a substance. The valence electrons, which are involved in chemical bonding, are termed as bonding electrons.

**Ionic Bond**

This is a type of bond formed by the electrostatic attraction between oppositely charged ions, resulting from the transfer of electrons from a metal (low electronegativity) to a non-metal (high electronegativity).

**Covalent Bond**

This bond involves the sharing of electron pairs between atoms, usually non-metals with similar electronegativities, to achieve stable configurations. It can be single (one pair shared, e.g.,  $H_2$ ), double (two pairs, e.g.,  $O_2$ ), or triple (three pairs, e.g.,  $N_2$ ).

**Coordinate Covalent Bond (Also Called Dative Bond):**

A subtype of covalent bond where both electrons in the shared pair are provided by one atom (the donor), while the other atom (the acceptor) contributes an empty orbital. Once formed, it is indistinguishable from a regular covalent bond.

**Metallic Bond**

This bond occurs between metal atoms, where valence electrons are delocalized and form a sea of electrons surrounding positively charged metal ions. The attraction between the ions and the mobile electrons holds the structure together.

#### 5.1.3 Exemplify The Formation Of

**Cations From An Atom Of A Metallic Element**

Cations are positively charged ions formed when atoms lose one or more valence electrons. Metallic elements, with low ionization energies, readily lose electrons to achieve stable configurations.

Example: Sodium (Na), a metallic element in group 1, has an electron configuration of  $[Ne] 3s^1$ . It loses one electron to form  $Na^+$  ( $[Ne]$ ), a cation. This occurs in reactions like  $2Na + Cl_2 \rightarrow 2NaCl$ , where sodium transfers its electron to chlorine.

**Anions From An Atom Of A Non-Metallic Element**

Anions are negatively charged ions formed when atoms gain one or more electrons to complete their valence shells. Non-metallic elements, with high electron affinities, attract electrons easily.

Example: Chlorine (Cl), a non-metallic element in group 17, has an electron configuration of  $[\text{Ne}] 3s^2 3p^5$ . It gains one electron to form  $\text{Cl}^-$ , an anion. This is seen in the same reaction:  $\text{Na} + \text{Cl} \rightarrow \text{Na}^+ + \text{Cl}^-$ , where chlorine accepts the electron from sodium.

#### 5.1.4 Explain The Nature Of Bonding Based On The Electronegativity Difference Of Bonded Atoms Using Linus Pauling Scale

Electronegativity is a measure of an atom's ability to attract electrons in a bond, quantified on the Linus Pauling scale (ranging from about 0.7 for francium to 4.0 for fluorine). The difference in electronegativity ( $\Delta\text{EN}$ ) between two bonded atoms determines the bond's nature: ionic, polar covalent, or nonpolar covalent.

- Nonpolar Covalent Bond ( $\Delta\text{EN} \approx 0$  to 0.4): When atoms have very similar electronegativities, electrons are shared equally. Example: H-H in  $\text{H}_2$  (both EN = 2.1,  $\Delta\text{EN} = 0$ ), resulting in a symmetric electron distribution.
- Polar Covalent Bond ( $\Delta\text{EN} \approx 0.5$  to 1.6): Electrons are shared unequally, creating partial charges ( $\delta^+$  and  $\delta^-$ ). The atom with higher EN pulls electrons closer. Example: H-Cl in HCl (EN of H = 2.1, Cl = 3.0,  $\Delta\text{EN} = 0.9$ ), where chlorine is  $\delta^-$  and hydrogen is  $\delta^+$ .
- Ionic Bond ( $\Delta\text{EN} > 1.7$ ): Extreme difference leads to electron transfer rather than sharing, forming ions. Example: Na-Cl in NaCl (EN of Na = 0.9, Cl = 3.0,  $\Delta\text{EN} = 2.1$ ), resulting in  $\text{Na}^+$  and  $\text{Cl}^-$ .

These thresholds are approximate; bonds can have mixed character (e.g., highly polar covalent nearing ionic). Pauling scale helps predict bond polarity and compound properties, like solubility or melting point.

## 5.2 Ionic Bond

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### 5.2.1 Explain The Formation Of An Ionic Bond

An ionic bond forms when one or more electrons are transferred from one atom (typically a metal) to another atom (typically a non-metal), creating oppositely charged ions that are held together by strong electrostatic forces. This process occurs because atoms aim to achieve a stable electron configuration, often resembling that of a noble gas.

#### Process:

A metal atom with low electronegativity loses one or more valence electrons, forming a positively charged cation. A non-metal atom with high electronegativity gains those electrons, forming a negatively charged anion. The resulting oppositely charged ions attract each other, forming an ionic bond.

#### Example:

In sodium chloride (NaCl), sodium (Na) loses one electron to form  $\text{Na}^+$ , and chlorine (Cl) gains one electron to form  $\text{Cl}^-$ . The electrostatic attraction between  $\text{Na}^+$  and  $\text{Cl}^-$  creates the ionic bond.

### 5.2.2 Explain The General Characteristics Of Ionic Compounds

1. High Melting and Boiling Points: The strong ionic bonds require significant energy to break, resulting in high melting and boiling points (e.g., NaCl melts at  $801^\circ\text{C}$ ).
2. Solubility in Water: Most ionic compounds dissolve in water because water molecules can surround and separate the ions (e.g., NaCl is highly soluble).
3. Electrical Conductivity: Ionic compounds conduct electricity when molten or dissolved in water, as the ions are free to move and carry charge. In solid form, they do not conduct because the ions are fixed in the lattice.

4. Hard and Brittle: The rigid lattice structure makes ionic compounds hard, but they shatter when struck because like-charged ions can align and repel each other.
5. Crystal Structure: Ionic compounds form crystalline solids with a repeating lattice pattern (e.g., NaCl forms a cubic lattice).

### 5.2.3 Identify A Compound As Having Ionic Bond

- Ionic compounds are mostly crystalline solids.
- Ionic compounds in solid state have negligible electrical conductance but they are good conductors in solution and in the molten form. It is due to presence of free ions in them.
- Ionic compounds have high melting and boiling points. For example, sodium chloride has melting point 800 °C and a boiling point 1413 °C.
- As ionic compounds are made up of positive and negative ions, there exist strong electrostatic forces of attraction between oppositely charged ions. A great amount of energy is required to break these forces.
- They dissolve easily in polar solvents like water. Water has high dielectric constant that weakens the attraction between ions.

### 5.2.4 Draw Electron Dot And Cross Structure Of A Binary Ionic Compound For Example, NaCl, MgO, K<sub>2</sub>O and CaCl<sub>2</sub>

Practice work. Do it yourself.

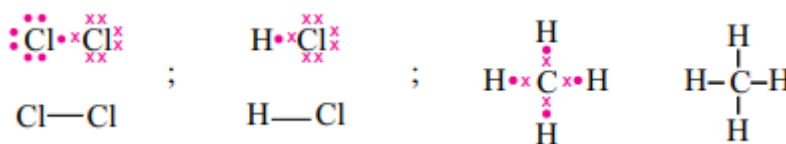
## 5.3 Covalent Bond

### 5.3.1 Explain The Formation Of A Covalent Bond Between Two Non-Metallic Elements

When two atoms approach each other, attractive forces develop between electrons of one atom and nucleus of the other atom. Simultaneously, repulsive forces between electrons of the two atoms as well as between their nuclei are also created. When the attractive forces dominate due to decrease in distance between those two atoms, a chemical bond is formed between them. The formation of hydrogen, chlorine, nitrogen and oxygen gases are few examples of this type of bonding.

### 5.3.2 Exemplify Single, Double And Triple Covalent Bonds

When one electron is contributed by each bonded atom, one bond pair is formed and it forms a single covalent bond. While drawing the structure of such molecules the single bond pair is indicated by a line between those two atoms. A few examples of molecules with single covalent bonds are hydrogen (H<sub>2</sub>), chlorine (Cl<sub>2</sub>), hydrochloric acid (HCl) and methane (CH<sub>4</sub>).



When each bonded atom contributes two electrons, two bond pairs are shared and a double covalent bond is formed. These bond pairs are indicated as double line between those atoms in the structure of such molecules. The molecules like oxygen (O<sub>2</sub>) and ethene (C<sub>2</sub>H<sub>4</sub>) show such type of double covalent bonds.





carbon and hydrogen in methane, or nitrogen and hydrogen in ammonia. If the difference of electronegativities between two elements is more than 1.7 the bond between them will be predominantly ionic bond and if it is less than 1.7, the bond between two atoms will be predominantly covalent.

## 5.4 Coordinate Covalent Bond

### 5.4.1 Explain The Formation Of Coordinate Covalent Bond

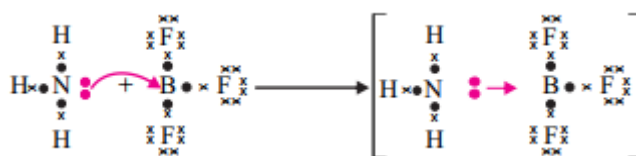
Coordinate covalent or dative covalent bonding is a type of covalent bonding in which the bond pair of electrons is donated by one bonded atom only. The atom which donates the electron pair is called donor and the atom which accepts the electron pair is called acceptor. A small arrow is usually used to indicate the atom and pair of electron being donated. The head of arrow is towards the acceptor atom.

### 5.4.2 Draw Electron Dot And Cross Structure Of Coordinate Covalent Compounds, For Example, Ammonium Ion, Oxonium (Hydronium) Ion, Aluminum Tetrachloride Anion, Adduct (Addition Product) Of Ammonia And Boron Trifluoride

Practice work. Do it yourself.

### 5.4.3 Compare The Formation And Characteristics Of Coordinate Covalent Compounds With Covalent Compounds

In the formation of  $\text{BF}_3$  (boron trifluoride) molecule, three valence electrons of boron atom pair up with three electrons, one from each three fluorine atoms. The boron atom even after this sharing of electrons (covalent bond formation), remains short or deficit of two electrons in its outermost shell. Now if a molecule with a lone pair approaches this molecule, it accepts lone pair from that donor and forms a coordinate covalent bond. The lone pair on nitrogen of  $\text{NH}_3$  makes it a good donor molecule to form a coordinate covalent bond.

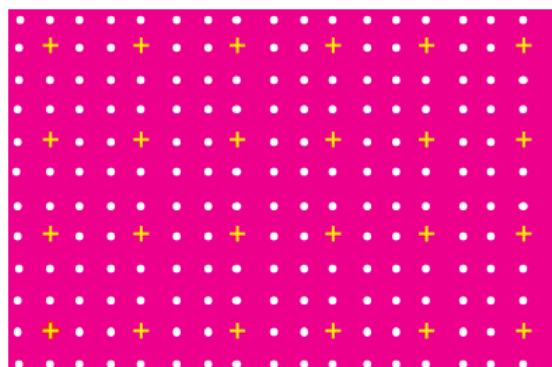


## 5.5 Metallic Bond

### 5.5.1 Explain The Formation Of Metallic Bonding (Electron-Sea Model)

The metallic bond is defined as a bond formed between metal atoms (positively charged ions) due to mobile or free electrons. In case of metals, the hold of nucleus over the outermost electrons is weak because of large sized atoms and greater number of shells in between nucleus and valence electrons. Furthermore, because of low ionization potentials, metals have the tendency to lose their outermost electrons easily.

Resultantly, these loose or free electrons of all metal atoms move freely in the spaces between atoms of a metal. None of these electrons is attached to any particular atom. Either they belong to a common pool, or belong to all the atoms of that metal. Nuclei of metal atoms appear submerged in sea of these free mobile electrons. These mobile electrons are responsible for holding the atoms of metals together forming a metallic bond.



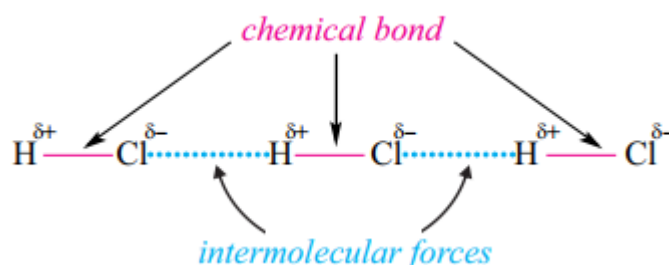
### 5.5.2 Explain The Structure Of Metals And The Mobility Of Their Electrons

- Malleability: The property by virtue of which a metal can be rolled into sheets
- Ductility: The property by virtue of which a metal can be drawn into wires.
- Melting And Boiling Points: They have usually high melting and boiling points.
- Lustre: They show metallic luster.
- Tensile Strength: Their tensile strength is very high.
- Electrical And Thermal Conductivity: They are good conductors of heat and electricity in solid and liquid state due to mobile electrons.

## 5.6 Intermolecular Forces

### 5.6.1 Explain Weak Intermolecular Forces, Including Van Der Waals Forces (London Dispersion Forces And Dipole-Dipole Interactions) And Hydrogen Bonding, With Their Effect On The Melting And Boiling Points Of Compounds.

The forces that hold atoms in a compound are chemical bonds. In addition to these strong bonding forces, relatively weak forces also exist in between the molecules, which are called intermolecular forces. The bonding and intermolecular forces of hydrochloric acid are shown.



It requires about 17 kJ energy to break these intermolecular forces between one mole of liquid hydrogen chloride molecules to convert it into gas. Whereas, about 430 kJ energy's required to break the chemical bond between hydrogen and chlorine atoms in 1 mole of hydrogen chloride.

#### Dipole - Dipole Interaction

All intermolecular forces, which are collectively called van der Waals forces, are electrical in nature. They result from the attractions of opposite charges which may be temporary or permanent. The unequal sharing of electrons between two different types of atoms makes one end of molecule slightly positive and other end slightly negatively charged. As shared pair of electron is drawn towards more electronegative atom, it is partially negatively charged, as chlorine in hydrogen chloride. The other end automatically becomes partially positively charged.

When partial positive and partial negative charges exist at different positions in a molecule, the adjacent molecules will arrange themselves in such a way that negative end of that molecule comes near to positive end of other molecule. It results in a net forces of attraction between oppositely charged ends of two adjacent molecules. These attractive forces are called dipole – dipole interactions as represented in HCl

