LLesson 1 : Types of Chemical Bonding

lesson Objects :

Dear students , this is the first lesson for the second unit.

At the end of the unit , you will be able to:

- List the favourable conditions for the formation of ionic bonds.
- Explain the formation of ionic bonding
- Give examples of ionic compounds
- Define lattice energy.
- Calculate lattice energy of ionic crystals from given data using the Born-Haber cycle
- Discuss the exceptions to the octet rule.
- Describe the properties of ionic bonding
- Carryout anactivity to demonstrate the effect of electricity on ioniv compounds (PbI_2 and NaCl).
- Carryout an investigation into the melting point and solubility of some ionic compounds (NaCl and $CuCl_2$.
- List types of chemical Bonding.
- Define covalent bond.
- Understand the metallic bonding
- Lewis structures or electrondotformulasof some covalent moleculecule.
- Illustrate the formation of coordinate covalent bonding using example draw resonance structures of some covalent molecules and polyatomic ion.
- Discuss the exceptions to the octet rule in covalent bond in distinguish between polar and non polar covalent molecule describe the properties of covalent molecules ,electricity and some solvents on covalent compounds (naphthalene, graphite, iodine and ethanol).

Brainstorming questions

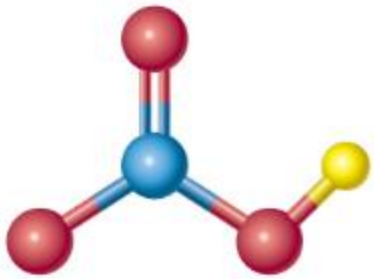
- why do atoms combine ?
- why molecules are stable than atoms ?

key terms/Concepts

- Ionic Bond
- Covalent Bond
- Metallic Bonding
- Lattice energy

An ionic bond is formed when one or more electrons are transferred from a metal atom (which becomes a positively charged ion or cation) to a non-metal atom (which becomes a negatively charged ion or anion).

2.1 Introduction



What is chemical bonding?

- Chemical bondis a strong force of binding between two or many atoms in compounds, molecules or ions.

Why do atoms combine? Atoms are bonded together to:

- Attain noble gas electronic configuration.
- To stability.
- have lowestpossible energy.

2.2. Types of Chemical Bonding

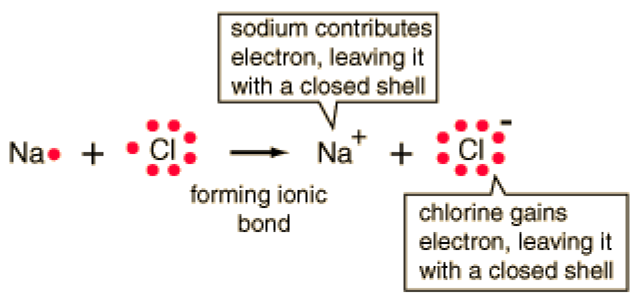
Chemical bonds are basically classified into three types consisting of

- Ionic or electrovalent bond.
- Covalent bond and.
- Metallic bond.

Mostly, valence electrons in the outer energy level of an atom take part in the chemical bonding.

2.2.1 Ionic Bonding

- The electrostatic force of attraction between oppositely charged ions results in the formation of ionic bond or electrovalent bond.
- Example:Consider the formation of some ionic bandings b/n Na+and Cl- ions:



- Examplesof ionic compounds include: NaCl, MgO, MgBr2, CaC2, HCl, Li3N…

Factors Affecting The Formation Of Ionic Bonds
Ionization energy(IE)
Electronaffinity (EA)
Lattice Energy (U)
Lattice Energy formation of Ionic bond

You need know the lattice energy of solid lithium fluoride. Finding the lattice energy of anionicsolidbyeperimentisdifficult.However,this quantity can be found indirectly using the Born–Haber cycle. The reasoning is based on Hess’s law, which states that an overall reaction’s enthalpy change is the sum of the enthalpy changes for the individual reactions that make it up:
$\Delta H_{total} = \Delta H_1 + \Delta H_2 + \Delta H_3 + \ldots$
Consider the Born-Haber cycle for the formation ofNaCl. We think that solid sodium chloride can be formed from the elements by two different routes,as shown in the following figure. In one route, NaCl(s) is formed directly from Na(s) and ½Cl (g); ΔH =−411 kJmol-1

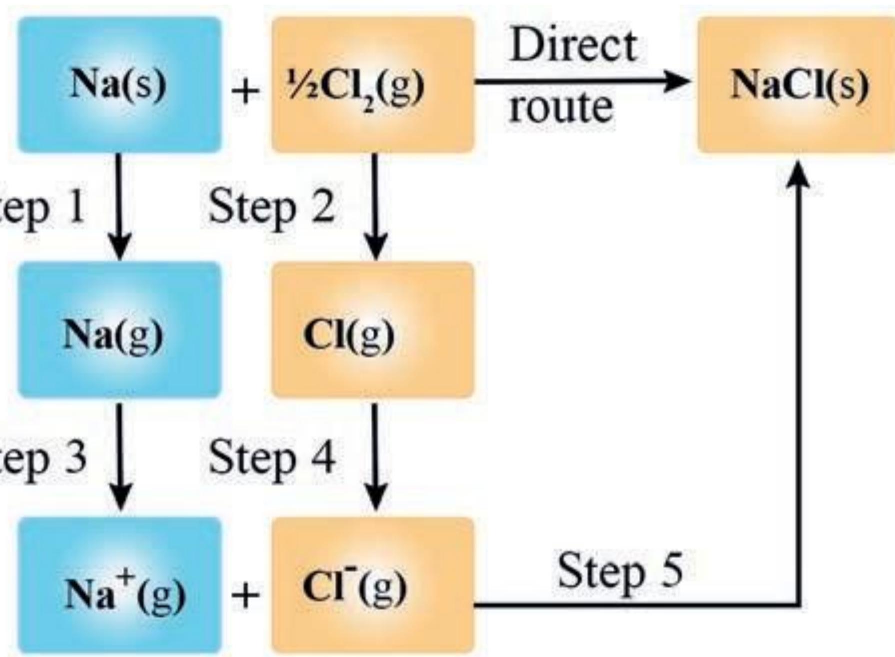


Figure :Born–Haber cycle for NaCl
The second route consists of the following five steps,along with the enthalpy change for each.
Step1: Metallic sodium is vaporized to a gas of sodium atom:
Na (s)→ Na (g) ΔH°step1= + 108 kJmol–1
Step 2: Chlorine molecules are dissociated to atoms:
½Cl2(g)→Cl (g) ΔH°step2=½bondenergy of Cl2
=½ (240 KJ)= + 120 kJmol–1
Step3:Sodium atoms are ionizedtoNa+ions:
Na(g)→Na+(g)+e– ΔH°step3=IE1=+ 496 kJmol–1

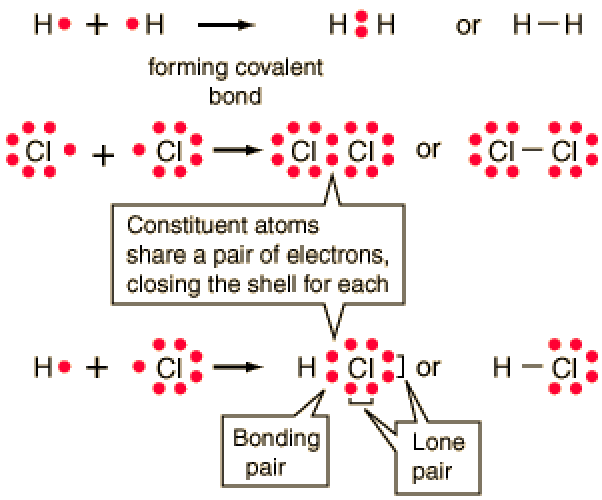
Step 4:Formation of chloride ion:
$Cl(g)+e^{–}\rightarrow Cl^{–}(g)$ $\Delta H°_{step4}=EA=–349$ $kJmol^{–1}$
Step 5:Formation of$NaCl(s)$ from ions.The ions $Na^{+}$and $Cl^{−}$combine to give solid sodium chloride whose enthalpy changes (the lattice energy)is unknown:
$Na^{+}(g)+Cl^{–}(g)\rightarrow NaCl\ (s)\Delta H°=U$(lattice energy)=?
We know the enthalpy formation ($\Delta_f H°$) of NaCl (Direct route) and equals −411 $kJmol^{-1}$. therefore, we can calculate the lattice energy using Hess's low: Solving for U NaCl gives:

NaCl f step1 step2 step3 step4
$=–411kJmol^{–1}–[108kJmol^{–1}+120kJmol^{–1}+496kJmol^{–1}+(–349kJmol^{–1})]$
$=–786$ $kJmol^{–1}$
Note:Ionic solid sexist only because the lattice energy exceeds the energy required for the electron transfer.
General Properties of Ionic Compounds

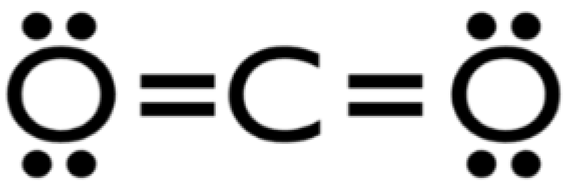
- Ionic compounds usually exist in the form of crystalline solids.
- Ionic compounds have high melting point and high boiling point due to strong attraction of force between ions.
  Ionic compounds are generally soluble in water and other polar solvents having high dielectric constants.
- Ionic compounds are good conductors of electricity in the solutions or in their molten states.
- In ionic – compounds, each ion is surrounded by oppositely charged ions uniformly distributed all around the ion. i.e., Ionic bonding is non-directional

2.2.2 The Covalent Bond

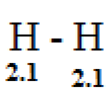
- Covalent bonds involve sharing of a pair of valence electrons by two atoms.
- When two atoms have a small difference in their tendencies to lose or gain electrons, we observe electron sharing and covalent bonding.
- This type of bonding most commonly occurs between nonmetal atoms (although a pair of metal atoms can sometimes form a covalent bond).Example:



Multiple covalent bonds are common for certain atoms depending upon their valence configuration.
For example, in ethylene (CO_2), a double covalent bond results from the sharing of two sets of valence electrons.



Atomic nitrogen (N_2) is an example of a triple covalent bond.
The polarity of a covalent bond is defined by any difference in electro-negativity between the two atoms participating in the covalent bond formation.
Bond polarity describes the distribution of electron density around two bonded atoms.
For two bonded atoms with similar electronegativities, the electron density of the bond is equally distributed (shared) between the two atom, this is a non-polar covalent .bond. eg, H_2,



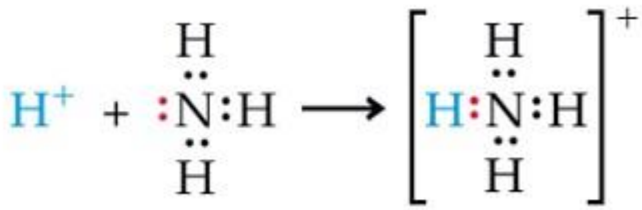
The electron density of a covalent bond is shifted towards the atom with the largest electro-negativity.

Polar covalent molecules.

eg. HCl, HBr……

$$\overset{\delta^{+}}{\underset{2.1}{\mathrm{H}}}\text{-}\overset{\delta^{-}}{\underset{2.8}{\mathrm{Br}}}$$

Coordinate covalent Bond



A coordinate covalent bond (also called a dative bond) is formed when one atom donates (contribute) both of the electrons from single atom
The atom which gives both electron called doner. And the opposite is Acceptor.

Properties of covalent compounds

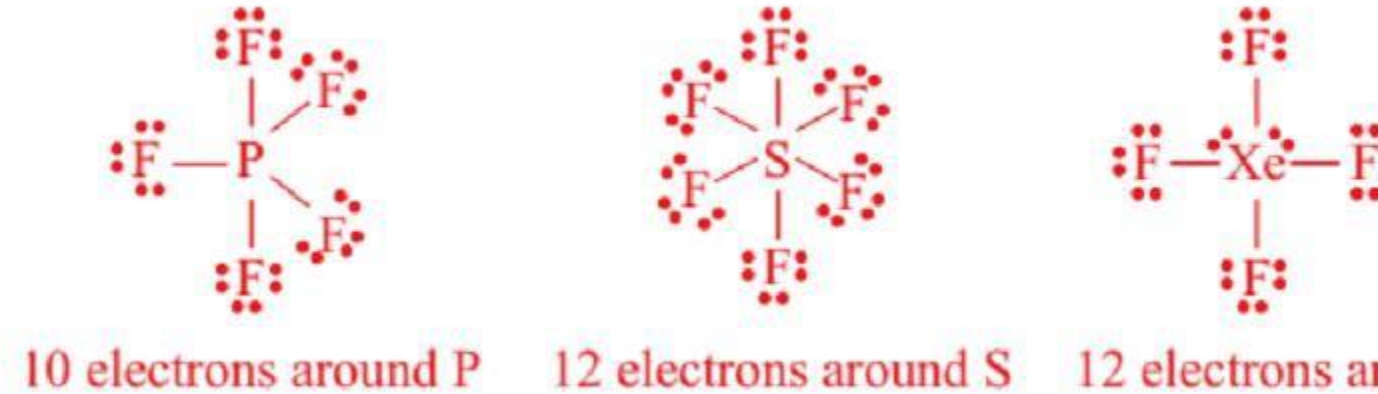
- The covalent compounds do not exist as ions but they exist as molecules.
- The m.pt. and b.pt. of covalent compounds are generally low
- Covalent compounds are generally insoluble or less soluble in water and other polar solvents. However, these are soluble in non-polar solvents.
- They are poor conductors of electricity in the fused or dissolved state.
- Molecular reactions are quite slow because energy is required to break covalent bonds.
- Since the covalent bond is localized in between the nuclei of atoms, it is directional in nature

The Octet Rule

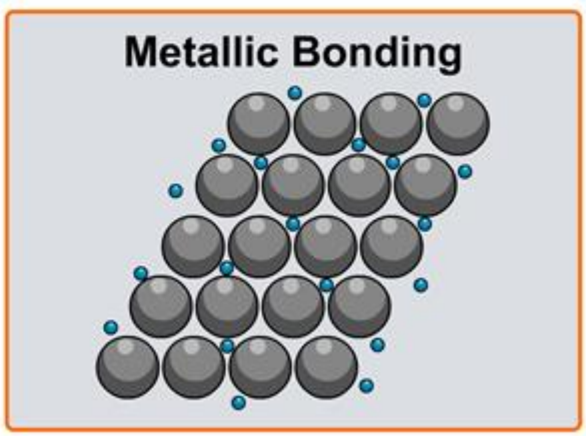
States that an atom tends to gain, lose or share electrons until there are eight electrons in its valence shell (the nearest noble gas configuration at its row).
The octet rule guides us in allotting electrons to the atoms in a Lewis structure; in a few cases, however, we set the rule aside.
Example: Na+ = 1s22s22p6 = 2:8 ,Configuration of Ne.
Cl- = 1s22s22p63s23p6 = 2:8:8 ,Configuration of Ar.
Note: – Hydrogen follows a special configuration called Dublet rule (2 electrons in its outer most shell).
Exceptions to the Octet Rule

1) Electron deficient molecules
Cpds containing either Be or B are often electron deficient, i.e., they have fewer than 8 electrons around Be or B atom.
Example: BH3, BeCl2, BF3, etc.

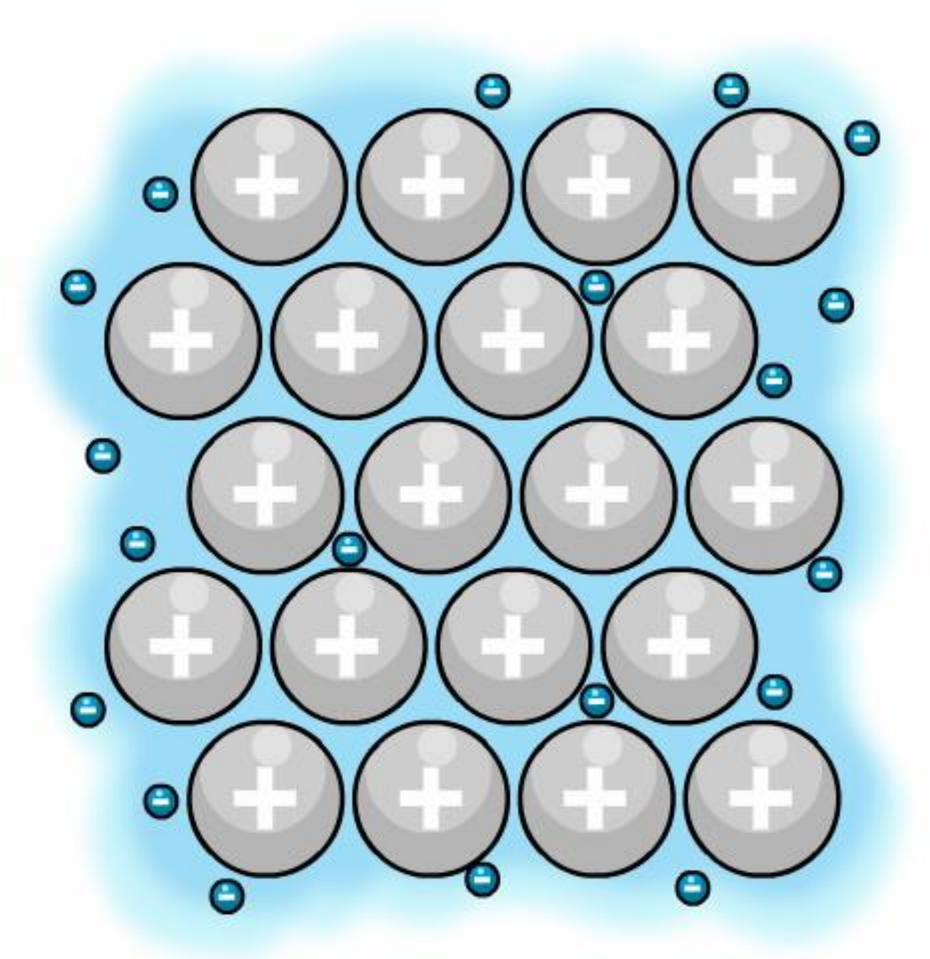
More than Octet:



2.2.3 Metallic Bonding



The atoms in a pure metal are in tightly-packed layers, which form a regular lattice structure.
Electrostatic forces of attraction between the positively charged neuclei and the negatively charged electrons hold the lattice together.
A metal is therefore a seen as a rigid framework of cations immersed in a ‘sea’ of electrons that serve as the cement holding the three-dimensional cationic network together – Metallic bonding.



The metal atoms become positively charged ions and are attracted to the sea of electrons. This attraction is called metallic bonding.

Properties of Metals

- conduct heat and electricity
- high Melting and boiling Point
- malleable and ductile
- strong, not brittle

Lesson 1: Summary

Chemical bonding refers to the attractive forces that hold atoms together in molecules, ions, or crystals. There are three main types of chemical bonds: covalent bonds, ionic bonds, and metallic bonds. the following are detailed summary of each type:

1. Covalent Bonds:

- Covalent bonds form when atoms share one or more pairs of electrons to achieve a stable electron configuration.
- **Electron Sharing**: Each atom contributes one or more electrons to the shared pair, allowing both atoms to achieve a noble gas configuration.
- **Types of Covalent Bonds**:
- **Single Covalent Bond**: Involves the sharing of one pair of electrons (e.g., H_2).
- **Double Covalent Bond**: Involves the sharing of two pairs of electrons (e.g., O_2).
- **Triple Covalent Bond**: Involves the sharing of three pairs of electrons (e.g., N_2).

Properties:

- Covalent compounds can be solids, liquids, or gases.
- They generally have lower melting and boiling points compared to ionic compounds.
- They may be polar or non-polar depending on the electronegativity difference between bonded atoms.
- **Examples**: H_2O (water), CH_4 (methane), CO_2 (carbon dioxide).

2. Ionic Bonds:

- Ionic bonds form between ions of opposite charges (cation and anion) through electrostatic attraction.
- **Electron Transfer**: One atom (typically a metal) loses electrons to become a cation, while another atom (typically a non-metal) gains those electrons to become an anion.
- **Formation**: The resulting ions attract each other to form a stable crystal lattice structure.

Properties:

- Ionic compounds are typically crystalline solids at room temperature.
- They have high melting and boiling points due to strong electrostatic attractions.
- They are brittle and often dissolve in water, where they conduct electricity.
- **Examples**: NaCl (sodium chloride), $MgCl_2$ (magnesium chloride), $CaCO_3$ (calcium carbonate).

3. Metallic Bonds:

- Metallic bonds occur between atoms within metals and alloys.
- **Electron Delocalization**: Metal atoms release their valence electrons into a “sea” of delocalized electrons, which move freely throughout the metal lattice.

Properties:

- Metals are good conductors of electricity and heat due to the mobility of electrons.
- They are malleable and ductile, meaning they can be hammered into thin sheets or drawn into wires.
- Metals have a characteristic luster due to the reflection of light by delocalized electrons.
- **Examples**: Copper (Cu), Iron (Fe), Aluminum (Al).

Comparison:

- **Bond Strength**: Ionic bonds are typically stronger than covalent bonds, and metallic bonds are stronger than both.
- **Nature of Attraction**: Ionic bonds involve electrostatic attraction, covalent bonds involve electron sharing, and metallic bonds involve electron delocalization.
- **Physical Properties**: Each type of bond contributes distinct physical properties to substances, influencing their behavior in different environments.