BASIC CONCEPT OF CHEMISTRY

◆ STOICHIOMETRY

• INTRODUCTION

- Matter is made up of small particles which may be atoms or molecules. As different kinds of matter contain different types of atoms or molecules which have different properties, therefore, different kinds of matter have different properties.
- The word Molecules was introduced by the famous Italian Chemist **Amedo Avogardo**. According to him, a molecule is the smallest particle of an element or a compound which can exist free. A molecule is divisible into its constituent atoms but an atom is not. Thus a hydrogen molecule consists of two hydrogen atoms. Similarly, molecules of chlorine, oxygen, nitrogen, etc. These are called elementary gases. Their molecule contains two atoms each.
- The number of atoms of an element in one molecule of it is called **atomicity**. The atomicity of the elementary gases is 2. The atomicity of ozone, O_3 is 3. Metals and the noble gases are mono–atomic. thus we say Na, K, Ag, Ne, He and not Na_2, K_2 etc.
- A mole of atom is a collection of atoms whose total weight is equal to the gram atomic weight. As equal number of moles of different elements contain equal number of atoms.

• IONS : CATION & ANION

CATION : The positively charged particle is called a cation. When an atom gives up an electron, it acquires an overall positive charge as the number of positively charged protons now exceed the number of negatively charged electrons therefore a cation contains less electrons than a normal atom.

e.g. **MONOVALENT**	**BIVALENT**	**TRIVALENT**
(Valency 1)	**(Valency 2)**	**(Valency 3)**
Hydrogen H^+	Magnesium Mg^{++}	Aluminium Al^{+++}
Sodium Na^+	Zinc Zn^{++}	Chromium Cr^{+++}
Potassium K^+	Copper (II) Cu^{++}	Iron Fe^{+++}
Silver Ag^+	Mercury (II) Hg^{++}	
Copper (I) Cu^+	Iron Fe^{++}	

Mercury (I) Hg^{+}

ANION : The negatively charged particle is called an anion. When an atom gains an electron, it acquires an overall negative charge as the number of negatively charged electrons now exceed the number of positively charged protons.

e.g. **MONOVALENT (Valency 1)**	**BIVALENT (Valency 2)**	**TRIVALENT (Valency 3)**
Chlorine Cl^{-}	Oxygen O^{2-}	Nitrogen N^{3-}
Bromine Br^{-}	Sulphur S^{2-}	(In Sulphides)
Iodine I^{-}		

Illustration –1: Write the valency of the given metals in the following compounds

Compounds	Name of Metal	Valency
Cuprous Chloride	Copper	1
Cupric chloride	Copper	2
Mercurous chloride	Mercury	1
Mercuric chloride	Mercury	2

Illustration –2: Write the formula for following polyatomic ions

Name of the ion	Formula
Hydroxide ion	OH^{-}
Nitrate ion	NO_3^{-}
Chlorate ion	ClO_3^{-}
Nitrile ion	NO_2^{-}
Carbonate ion	CO_3^{2-}
Manganate ion	MnO_4^{2-}
Sulphate ion	SO_4^{2-}
Phosphate ion	PO_4^{3-}
Phosphite ion	PO_3^{3-}

• FORMULA OF COMPOUND

(i) Writing chemical formulae

Knowing symbols of element s and their valencies, it is easy to write the formulae of compounds.

(a) Where the two radicals have the same valency, there is no problem, they are written side by side

Sodium chloride $Na^{I}Cl^{I}$ NaCl

Potassium nitrate $K^{I}NO_3$ KNO_3

(b) When the two radicals have different valencies, they are written side by side and their valency multiplied by suitable natural number. (not zero of fraction but only, 1, 2, 3, …..). So that number of +ve charge equals number of –ve charge

(i) $Al^{+++} \times 2$; $SO_4^{-} \times 3$

$(+3) \times 2 = +6$; $(-2) \times 3 = -6$

Similarly

(ii) $NH_4^+ \times 2; CO_3^{--} \times 1$

$(+1) \times 2 = +2$; $(-2) \times 1 = -2$ [Hence formula is $(NH_4)_2CO_3$]

(ii) Naming of compounds

Binary compounds : These contain only two elements and their names end in ide. The more metallic element comes first.

e.g. : NaCl, HCl, AgCl etc.

Ternary compounds : These contain three elements. if oxygen is one of them their names end in 'ite' or 'ate' according to the parent acid from which they are derived

e.g. : Chloric acid, $HClO_3$ Potassium chlorate, $KClO_3$

Perchloric acid, $HClO_4$ Potassium perchlorate, $KClO_4$

(iii) Empirical And Molecular Formulae

- We have just seen that known the molecular formula of the compound we can calculate percentage composition of the elements.
- Conversely if we know the percentage composition of the elements initially, we can calculate the relative number of atom of each element in molecules of the compound. This gives as the empirical formula of the compound.
- Further if the molecular mass is known then the molecular formula can easily be determined.
- Thus, the empirical formula of a compound is a chemical formula showing the relative number of atoms in the simplest ratio, the molecular formula gives the actual number of atoms of each element in a molecule. The molecular formula is generally an integral multiple of the empirical formula.

  i.e. molecular formula = empirical formula × n ; where $n = \frac{\text{molecular formula mass}}{\text{empirical formula mass}}$

Percentage composition

It is calculated from a knowledge of the formula of a compound and the atomic mass of its constituent atoms.

Example: Calculate the percentage composition of Glucose

Solution: Glucose is $C_6H_{12}O_6$

Molecular Mass = 72 + 12 + 96 = 180

Percentage of Carbon = $\frac{72}{180} \times 100\text{g} = 40.4$

Percentage of Hydrogen = $\frac{12}{180} \times 100 = 6.67$

Percentage of Oxygen = $\frac{96}{180} \times 100 = 53.33$

Deduction Of Molecular Formula : The Steps Are

(i) Divides the percentage composition by the respectively atomic masses of the elements. the atomic ratios are obtained.

(ii) These ratio are then divided by the smallest and expressed in whole numbers. the empirical formula is obtained.

(iii) The empirical formula mass is calculated.

(iv) It is then compared with the actual molecular mass.

(a) If both are same, the empirical formula is the molecular formula.

(b) If they are different, the molecular mass is divided by the empirical formula mass to get a whole number.

(c) The molecular formula is then a multiple of the empirical formula (multiplies by this whole number).

Note : If the vapour density is given, it must be multiplied by 2 to get the molecular mass. A few Illustration will make this clear.

Illustration 3: ***(i)*** Write the name of three binary compounds

Solution: Potassium chloride – KCl
Magnesium oxide– MgO
Zinc Sulphide – ZnS

(ii) Write formula of "ous acid" and "ic acid"

Solution: Sulphuric acid, H_2SO_4
Sulphurous acid, H_2SO_3
Chlorous acid, $HClO_2$

(iii) Acetylene and benzene both have the empirical formula CH. The molecular masses of acetylene and benzene are 26 and 78 respectively. The molecular formula will be.

(A) C_6H_6, C_2H_2 (B) C_7H_7, C_2H_2
(C) C_3H_6, C_4H_4 (D) C_6H_6, C_6H_6

Solution : **(A)**

I: The empirical formula of the compounds is CH
Empirical formula mass
$= (1 \times 10) + 1 = 13$
Molecular mass = 26

II: To calculate the value of 'n'

III: To calculate the molecular formula of the compound.
Molecular formula = n × (Empirical formula of the compound)
$= 2 \times CH = C_2H_2$
Thus the molecular formula is C_2H_2
Similarly for benzene
To calculate the value of 'n'

$$n = \frac{\text{molecular mass}}{\text{empirical formula mass}} = \frac{78}{13} = 6$$

thus the molecular formula is $6 \times CH = C_6H_6$

(iv) Find the percentage Composition of water (H_2O)

Solution: Molecular Mass of water = 18

18 g water contain 2 g of hydrogen

100 g water contain $\frac{2}{8}\times 100g = 11.1g$ hydrogen

18 g of water contain 16 g of oxygen

100 g of water contain $\frac{16}{18}\times 100g$ of oxygen = 88.9 g of oxygen

Percentage composition Hydrogen = 11.1 %

Oxygen = 88.9 %

(v) A compound contains 52% carbon, 13% hydrogen and the rest oxygen. The molecular mass o the compound is 45.5. What is its molecular formula ?

Solution:

Element	Percentage	%/ Atomic mass	Simple ratio
C	52	52/12 = 4.33	4.33/2.18 = 2
H	13	13/1 = 13	13/2.18 = 6
O	35	35/16 = 2.18	2.18/2.18 = 1

Empirical formula = $(C_2H_6O_1)_n$

Empirical formula mass = 12 2 + 1 6 + 16 1

= 24 + 6 + 16 = 46

Molecular mass (given) = 45.5

Molecular formula = $\frac{\text{molecular mass}}{\text{Emprical formula mass}} = n = \frac{45.5}{46} = 0989 = 1$ (whole number)

• ATOMIC MASS:

It is the mass of 1 atom of a substance it is expressed in amu

Atomic mass = RAM × 1 amu

Note: *Relative atomic mass is nothing but the number of nucleons present in the atom*

• GRAM ATOMIC MASS:

The atomic mass of an element expressed in gram is called gram atomic mass of the element.

For example : Atomic mass of 'O' atom = mass of one 'O' atom = 16 amu

gram atomic mas = mass of 6.02×10^{23} 'O' atoms

= 16 amu $\times 6.02\times10^{23}$

= $16\times1.66\times10^{-24}$ g × 6.02×10^{23} = 16 g

• GRAM MOLECULAR MASS:

The molecular mass of a substance expressed in gram is called the gram–molecular mass of the substance.

It is also defined as the mass of one mole. (molar mass)

Illustration – 4: Molecular mass of O_2 molecules = mass of one mole of O_2 molecule

Solution: 2× mass of one O atom

= 2×16 amu = 32 amu

gram molecular mass = mass of 6.02×10^{23} O_2 molecules = 32 amu $\times 6.02\times10^{23}$

= 32 × 1.66 $\times10^{-24}$ gm × 6.02×10^{23} = 32 gm

• MOLE CONCEPT

In plain language, a mole is a unit used by the chemist to count atoms, molecules, ions etc. Just as a dozen means 12 and a score means 20, a mole means a collection of 6.023×10^{23} atoms, molecules, ions or electrons.

- This 'definite number' is called Avogadro number which is equal to 6.023×10^{23}.

  A mole can be defined as, "the amount of a substance containing as many atoms, molecules, ions, electrons or other elementary entities as there are carbon atoms in exactly 12g of C–12".

- ***Gram–molecules and gram–atom are termed as a mole of molecules and a mole of atoms respectively, e.g., 1 gram–molecule of chlorine and 1 gram atom of chlorine are expressed as 1 mole of Cl_2 and 1 mole of Cl respectively.***

- The number of moles can be calculated by a number of formulae depending upon the data given:

- No. of moles of molecules = $\frac{\text{Wt. of substance in g}}{\text{Mol. wt. of substance}}$

- No. of moles of atoms = $\frac{\text{Wt. of element in g}}{\text{At. wt. of the element}}$

- No. of moles of gas = $\frac{\text{Volume of the gas at STP}}{\text{Standard Molar Volume (i.e. 22.4L)}}$

Illustration– 5: 1.80 g of a certain metal burnt in oxygen gave 3.0 g of its oxide. 1.50 g of the same metal heated in steam gave 2.50 g of its oxide. The law show by above data.

(A) law of constnat compostion
(B) law of multiple proportion
(C) law of reciprocal proportion
(D) all

Solution: (A)

In the first sample of oxide,

Wt. of metal = 1.80 g

Wt. of oxygen = (3.0 – 1.80) g = 1.2 g

$\therefore \quad \frac{\text{wt. of metal}}{\text{Wt. of oxygen}} = \frac{1.80\text{g}}{1.2\text{g}} = 1.5$

In the second sample of the oxide,

Wt. of metal = 1.50 g

Wt. of oxygen = (2.50 – 1.50) g = 1 g

$\therefore \quad \frac{\text{wt. of metal}}{\text{Wt. of oxygen}} = \frac{1.50\text{g}}{1\text{g}} = 1.5$

Thus, in both samples of the oxide the proportions of the weights of the metal and oxygen a fixed. Hence the results the law of constant proportion.

◆ CHEMICAL BONDING

• INTRODUCTION

- Atoms are usually not capable of free existence but groups of atoms, of the same or different elements exist as one species. For example H_2, O_2, P_4, S_8, H_2O
- A group of atoms existing together as one species and having characteristic properties is called a **molecule**.
- Obviously, there must be some force which holds these atoms together within a molecule is called a **chemical bond**.
- Various attempts were made to explain the formation of chemical bonds in terms of electrons but it was only in 1916 when Kossel and Lewis succeeded independently gave a satisfactory.
- This theory was based on the electronic concept of atom
- According to it, in the formation of a molecule only the outer shell electrons are involved and they are known as valence electron.

Mechanism of chemical bonds based in stability of Atom

Law of duplet : For an atom to be stable it must have two electrons in its first and outermost shell like that of helium.

Law of octet : For atom to be stable it must possess eight electrons in its outermost shell like that of inert gases other than helium.

Loss and gain of electrons : From the above laws if follows that atoms with incomplete octet of duplets tend to attain the stable electronic configuration of the atoms of their neatest inert gas by losing, gaining or sharing electrons.

- From the study of electronic configuration of the noble gases, it is clear that they have 8 electrons in the outermost shell (called octet) or 2 electrons in case of helium (called duplet). In case of all other elements, the number of electrons in their outermost shell is less than 8 and hence they are chemically reactive. This led to the conclusion, called octet rule.

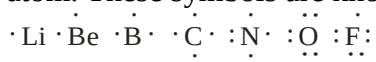
Example : Above statement can be understood with the following table

Noble gas	**Atomic No.**	**Electronic Configuration**
Helium	2	2
Neon	10	2, 8
Argon	18	2, 8, 8
Krypton	36	2, 8, 18, 8
Xenon	54	2, 8, 18, 18, 8
Radon	86	2, 8, 18, 32, 18, 8

Lewis Symbols

- In the formation of a molecule, only the outer shell electrons are involved and they are known as valence electrons.

- G.N. Lewis introduced simple symbols to denote the valence shell electrons in an atom. These symbols are known as Lewis symbols or electron dot symbols.



Lewis Dot Structures (Rules of Representation)

Step – 1 : Calculate the total number of valence electron of the atoms present.

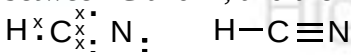
Step – 2 : If the species is an anion, add number of electrons equal to the units of –ve charge and if the species is a cation subtract number of electrons equal to the unit of +ve charge. This gives the total number of electrons to be distributed.

Step – 3 : Select the central atom (which is generally the least electronegative atom) and draw the skeletal structure by intelligent guess to indicate which atom is linked to which other ratio. Also remember that H and F usually occupy terminal position.

Step – 4 : Put one shared pair of electron between every two atoms to represent a single bond between them. Use the remaining pairs of electrons either for multiple bond or to show them as lone pairs, keeping in mind that octet of each atoms completed.

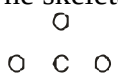
Illustration –6 : ***(i)*** Draw the Lewis dot structure of HCN molecule

Solution: Step 1 : Total number of valence electrons in HCN = 1+ 4 + 5 = 10
($_1$H = 1, $_6$C = 2, 4, $_7$N = 2, 5)
Step 2 : Skeletal structure is
H C N
Step 3 : Putting on shared pair of electrons between H and C and one between C and N, and the remaining as lone pairs, we have

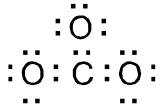


(ii) Draw the Lewis dot structure of CO_3^{2-} ion.

Solution: Step 1 : Total number of valence electron of $CO_3 = 4 + 3 \times 6 = 22$
($_6$C = 2, 4. $_8$O = 2, 6)
Step 2 – Total number of electrons to be distributed in $CO_3^{2-} = 22 + 2$ (for two units – ve charge) = 24
Step 3 – The skeletal structure of CO_3 is



Step 4 – Putting one shared pair of electrons between each C and O and completing the octets of oxygen, we have
In this structure octet of C is not complete. Hence, multiple bonding is required between C and on of the O – atoms. Drawing a double bond between Ca and one O–atom serves the purpose.



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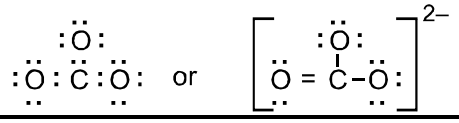


• TYPES OF CHEMICAL BONDS

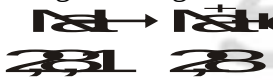
➢ There are mainly three types of Chemical bonds which are as follows :
Electrovalent or Ionic Bond
Covalent Bonds
Co–ordinate Bonds

ELECTROVALENT OR IONIC BOND

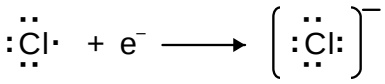
➢ When a bond is formed by complete transfer of electrons from one atom to another so as to complete their octet, duplet is called ionic bond or electrovalent bond.

EXAMPLE :

1. Formation of sodium chloride : Sodium (atomic number = 11) has electronic configuration 2, 8, 1. By losing one electron of its outermost shell, it acquires the inert gas configuration of neon and changes into ion

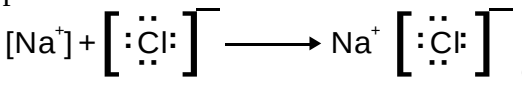


On the other hand, chlorine having electronic configuration 2, 8, 7 accepts one electron released by sodium to complete its octet one electron released by sodium to complete its octet by attaining stable configuration of argon in this process, chlorine is converted into chloride ion



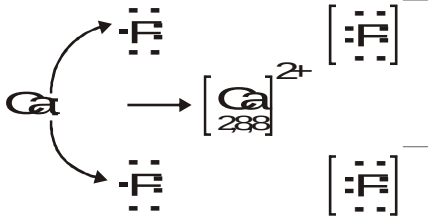
(2,8,7) (2,8,8)

Now, we have two species, one is positively charged sodium ion and the other is negatively. charged chloride ion. As they approach each other, they are held together by strong electrostatic forces of attraction. Thus, formation of sodium chloride takes place.



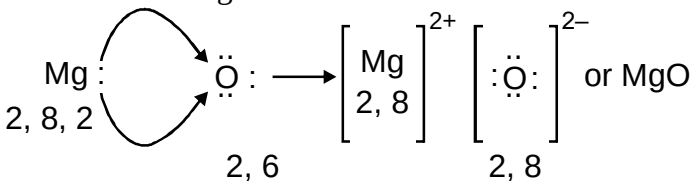
or NaCl

2. Formation of calcium fluoride

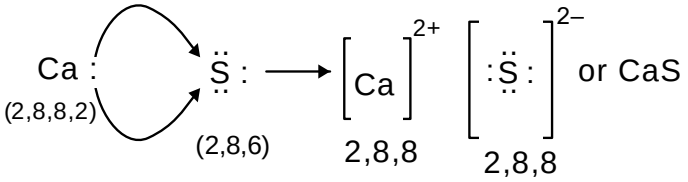


or CaF_2

3. Formation of magnesium oxide



4. Formation of Calcium Sulphide.



Factor Governing the Formation of ionic Bonds

(i) **Ionisation Enthalpy** of any element is the amount of energy required to remove an electron from the outermost shell of an isolated atom in gaseous phase so as to convert in into a gaseous positive ion.

(ii) **Electron affinity or Electron gain enthalpy** of an element is the enthalpy change that takes place when an extra electron is added to an isolated atom in the gaseous phase to form a gaseous negative ion.

Ionic bonds are generally formed between the atoms of the elements placed on the left and on the right side of the periodic table.

(iii) **Lattice Enthalpy (lattice Energy).** The energy released with the requisite number of gaseous positive and negative ions combine to form one mole of the ionic compound is called lattice enthalpy.

The higher the value of lattice enthalpy of the resulting ionic compound, the greater will be the stability of the compound and hence greater will the ease of its formation.

Lattice enthalpy depends upon the following two factors :

(A) **Charge on the ions** : The higher the charge on the ions, greater is the force of attraction and hence larger is the amount of energy released.

For example, lattice enthalpies of some ionic compounds are in the order :

[redacted]

(B) **Size of the ions** : As highly charged species are rare, the other factor i.e., the internuclear distance between the ions becomes more important. If the size of the ions is large, internuclear distance will be more and force of attraction is greater.

For example, lattice enthalpies of some ionic compounds are in the order :

NaCl > NaBr > LiF > LiI,

AgCl > AgI

CaF_2 > $CaCl_2$, BeO > MgO,

General characteristics of ionic compounds

1 **Physical State** : These compounds usually exist in the solid state.

2 **Crystal Structure** : X – ray analysis of the ionic compounds shows that they exist as ions and not as molecules. These ions are arranged in a regular pattern in the three dimensional space to form a lattice.

3 **Solubility** : Electrovalent compounds are soluble in solvents like water which are polar in nature and have high dielectric constant.

4 **Electrical conductivity** : Ionic compounds are good conductors of electricity in solution or in the molten state.

5 **Ionic reactions :** The reaction of the ionic compounds are, in fact, the reactions between the ions produced in the solution.

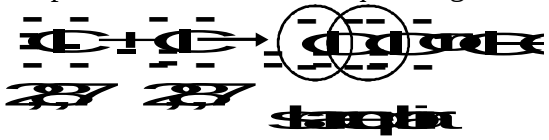
$$Na^+Cl^-(aq) + Ag^+NO_3^-(aq) \longrightarrow AgCl(s) + NaNO_3(aq.)$$

Covalent Bond

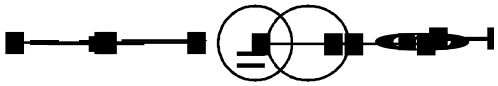
- **Langmuir**, In 1919 improved the Lewis concept by suggesting that when both the atoms taking part in a chemical combination are sort of electrons than the nearest noble gas configuration, they can share their electrons in order to complete their octets.
- The bond formed between the two atoms by mutual sharing of electrons between them so as to complete their octets or duplets in case of elements having only one shell is called covalent bond.

EXAMPLE :

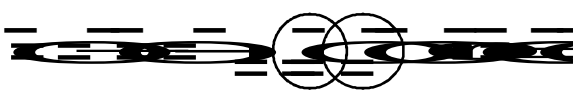
- Formation of chlorine molecules. In this case, two chlorine atoms combine to produce chlorine molecule. Each chlorine atom is short of one electron to attain stable configuration of argon. Each of them contributes one electron to form a common shared pair, which is also termed as ***single bond.***
- Similarly two electron shared by each molecule is termed as ***double bond*** three electron shared by each molecule is termed as ***triple bond***. By doing so, both of them complete their orbits and acquire argon structure.



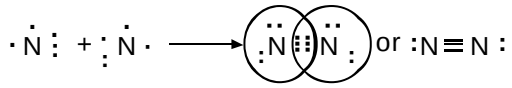
(1) Formation of Hydrogen molecules



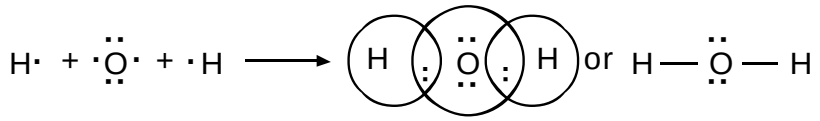
(2) Formation of Oxygen molecule



(3) For motion of Nitrogen molecule



(4) Formation of H_2O molecule



Bonds Pairs and Lone Pairs. The shared pairs of electrons present between the atoms are called bond pair. On the other hand the valence electron not involved in bonding are called non–bonding electron or lone pairs or unshared pairs.

Limitations of the Octet Rule

- Although octet rule is able to explain the formation of a large number of compounds, there are some exceptions to this rule as follows :

(i) Formation of compound involving hydrogen
(ii) Formation of compounds like $BeCl_2$, BF_3, $AlCl_3$ etc.
(iii) Formation of compounds like PCl_5, SF_6, IF_7, H_2SO_4 etc.
(iv) Formation of compounds of noble gases

Characteristics of Covalent Compounds

1. **Physical State** : The covalent compounds exist in all the three states, viz., solid, liquid and gas.
2. **Crystal Structure** : The crystal structure of covalent compounds differs from the ionic compounds. They usually consist of molecules rather than ions.
3. **Melting and Boiling Points :** Covalent compounds have low melting and boiling points.
4. **Solubility** : Covalent compounds are generally soluble in organic Solvents but insoluble in water and other polar solvents.
5. **Electrical conductivity :** Since there are no free ions in covalent compounds to conduct electricity, they are bad conductors of electricity.

Limitations of covalent bond

(i) It could not explain how the atoms are held together

(ii) It could not explain the shapes of molecules containing covalent bonds.

(iii) It could not explain the release of energy during the formation of a covalent bond.

Polar and Non–Polar Covalent Bonds

1. **Non–Polar Covalent Bonds :** If two similar atoms comes close to each other and form a bond by sharing their electrons, the shared electron are equally attracted by the two atoms as the electronegativity of the atoms is same.
2. **Polar Covalent Bonds:** When two dissimilar atoms, having different electronegativity is combine together to form a covalent bond the shared pair of electrons does not lies a

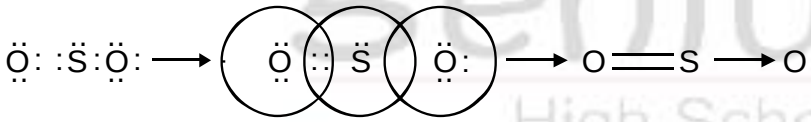
equal distances form the nuclei of both not lie at equal distances from the nuclei of both the bonded atoms but shifts towards the atom having greater electronegativity. Since the more electronegative atom attracts the electrons more strongly, the distribution of electrons gets distorted,. i.e., the electron cloud is displaced more towards the more electronegative atom. As a result, one end of the molecules having more electronegative atom becomes slightly negatively charged while the other end acquires slightly positive charge. Thus, positive and negative poles are developed and this type of bond is called polar covalent bond (or covalent bond with partial ionic character).

It may be noted that in case of symmetrical molecules like CO_2, CCl_4 etc. although there are a number of polar bonds, present, yet the molecules whose are non–polar. This is because the polar bonds cancel the effect of each other.

Ionic bond as an extreme case of polar covalent bond.

Coordinate or Dative Bond

- Perkins in 1921, suggested a third possible way by which atoms can combine and form a molecules to account for the structures of certain compounds such as sulphur trioxide, sulphuric acid, nitric acid. etc.
- When in the formation of a bond, the electron pair (lone pair) is donated by one atom but shared by both the atoms so as to complete their octets, the bond formed is called coordinate bond or dative bond



Bond Characteristics

1. **Bond Length:** The equilibrium distance between the centres of the nuclei of the two bonded atoms is called its bond length.

   In general, for a covalent molecules AB, the bond length (D) is given by

   $$d = r_A + r_B$$

   where r_A and r_B represent covalent radii of atoms A and B respectively.

   In homonuclear diatomic molecule, the covalent radius is equal to half of the bond length.

   For example

   $$r_{Cl} = \frac{d_{Cl-Cl}}{2}$$

   Factors affecting bond length

   (i) **Size of the atoms** : The bond length in creases with increase in the size of the atoms. For example, bond lengths of H—X are in the order

   HI > HBr > HCl > HF

   (ii) **Multiplicity of bond** : The bond length decreases with the multiplicity of the bond. Thus, bond length of carbon–carbon bond are in the order :

$C \equiv C < C = C < C - C$

2. **Bond Enthalpy (Bond Energy)**

The amount of energy required to break one mole of bonds of a particular type so a to separate them into gaseous atoms is called bond dissociation enthalpy or simply bond enthalpy :

Factors affecting bond enthalpy :

(i) **Size of the atoms** : Greater the size of the atoms, greater is the bond length and less is the bond dissociation enthalpy, i.e., less is the bond strength

(ii) **Multiplicity of Bonds** : For the bond between the same two atoms, greater is the multiplicity of the bond, greater is the bond dissociation enthalpy.

$H - H < O = O < N \equiv N$

(iii) Number of lone pairs of electrons present. Greater the number of lone pairs of electrons present on the bonded atoms, greater is the repulsion between the atoms and hence less is the bond, dissociation enthalpy. For example for a few single bonds, we have

3. **Bond Angle** : The angle between the lines representing the direction of the bonds, i.e., the orbitals containing the bonding electrons is called the bond angle.

4. **Bond Order** : The number of bonds present between two atoms is called bond order.

Molecule :	H – H	O = O	N ≡ N	C ≡ C
Bond Order:	1	2	3	3

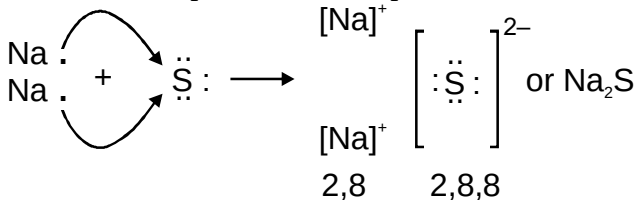
Illustration – 7: ***(i)*** Identify ionic compound from the following

HCl, CO_2, H_2S, $MgCl_2$, N_2, NO_2, H_2SO_4,

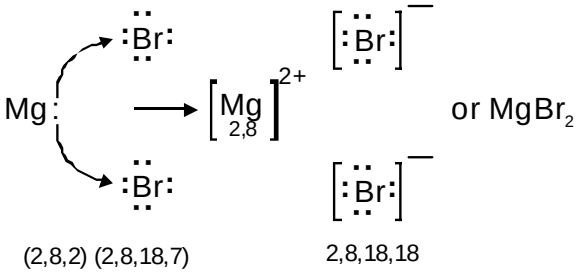
Solution: HCl, $MgCl_2$ are ionic compound

(ii) Explain the formation of sodium sulphide and magnesium bromide.

Solution: Formation of Sodium sulphide can be explain in the following manner

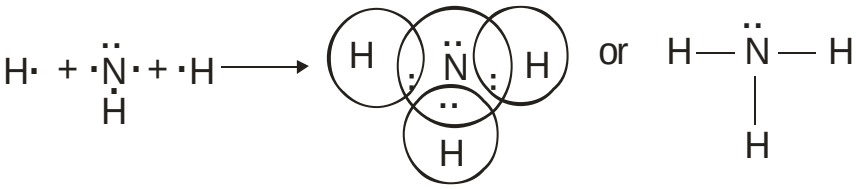


Formation of magnesium bromide can be explain in the following manner

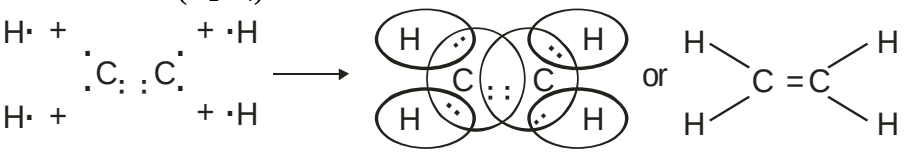


(iii) Explain the formation of ammonia and ethene

Solution: Formation of NH_3 molecule

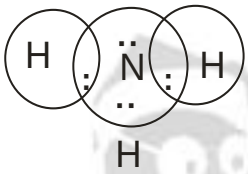


Formation of (C_2H_4) molecule



(vi) Calculate no. of bond pair and lone pair in the following

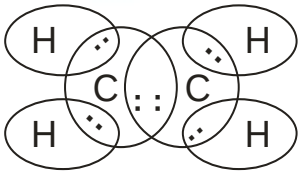
Solution(a): (a) NH_3 (b) C_2H_4



Bond pair = 3

Lone pair = 1

Solution(b):



Bond pair = 6

Lone pair = 0

KEY POINTS

- Atomic Mass: It is the mass of 1 atom of a substance it is expressed in AMU Atomic mass = RAM × 1 amu
- Gram Atomic Mass: The atomic mass of an element expressed in gram is called gram atomic mass of the element.
- Molecules: It is the smallest particle of matter which as free existence. Molecules can be further divided into its constituents atoms by physical and chemical process.
- Molecular Mass: It is the mass of one molecule
- Gram Molecular Mass: The molecular mass of a substance expressed in gram is called the gram–molecular mass of the substance.
- MOLES: Gram–molecules and gram–atom are termed as a mole of molecules and a mole of atoms respectively, e.g., 1 gram–molecule of chlorine and 1 gram atom of chlorine are expressed as 1 mole of Cl_2 and 1 mole of Cl respectively.
- PERCENTAGE COMPOSITION:

  $$\text{percentage mass of an element} = \frac{\text{mass of that elements in the compound}}{\text{molar mass of that compound}} \times 100$$

- EMPIRICAL FORMULA FOR MOLECULAR FORMULA

  $$\text{molecular formula} = \text{empirical formula} \times n \text{ ; where } n = \frac{\text{molecular formula mass}}{\text{empirical formula mass}}$$

- In the formation of a molecule only the outer shell electrons are involved and they are known as valence electrons.
- Atoms are usually not capable of free existence but groups of atoms, of the same or different elements exist as one species.
- Atoms combine together in order to complete their respective octets so as to acquire the stable inert gas configuration
- Ionisation Enthalpy of any element is the amount of energy required to remove an electron from the outermost shell of an isolated atom in gaseous phase
- The higher the value of lattice enthalpy of the resulting ionic compound, the greater will be the stability of the compound.
- Electrovalent compounds are soluble in solvents like water which are polar in nature and have high dielectric constant
- Ionic compounds are good conductors of electricity in solution or in the molten state.
- Covalent bond could not explain how the atoms are held together.
- Covalent bond could not explain the release of energy during the formation of a covalent bond.
- The bond length increases with increase in the size of the atoms
- The bond length decreases with the multiplicity of the bond

ASSIGNMENT-1

1. Calculate the number of g–atom of oxygen in 6.02×10^{24} CO molecules?

2. Calculate the mass of H_2SO_4 present in 0.25 mole of H_2SO_4?

3. 4.6×10^{22} atoms of an element weight 13.8 g. Find the atomic mass of the element?

4*. Vapour density of a gas is 22. What is its molecular mass?
(A) 33 (B) 22
(C) 44 (D) 11

5. The empirical formula of an acid is CH_2O_2, the probable molecular formula of acid may be
(A) CH_2O (B) CH_2O_2
(C) $C_2H_4O_2$ (D) $C_3H_6O_4$

6. Volume of a gas at STP is 1.12×10^{-7} cc. Calculate the number of molecules in it
(A) 3.01×10^{20} (B) 3.01×10^{12}
(C) 3.01×10^{23} (D) 3.01×10^{24}

7. The simplest formula of a compound containing 50% of element X (atomic mass 10) and 50% of element Y (atomic mass 20) is
(A) XY (B) X_2Y
(C) XY_3 (D) X_2Y_3

8. Which of the following has maximum mass?
(A) 0.1 g atom of carbon (B) 0.1 mol of ammonia
(C) 6.02×10^{22} molecules of hydrogen gas (D) 1120 mL of carbon dioxide at S.T.P.

9. Volume at N.T.P. of 0.22 g of CO_2 is same as that of
(A) 0.01 g of hydrogen (B) 0.085 g of NH_3
(C) 320 mg of gaseous SO_2 (D) All the above.

10. The number of moles of SO_2Cl_2 in 13.5 g is
(A) 0.1 (B) 0.2
(C) 0.3 (D) 0.4.

ASSIGNMENT-2

1*. The largest number of molecules is in:
(A) 28 g of CO
(B) 46 g of C_2H_5OH
(C) 36 g of H_2O
(D) 54 g of N_2O_5

2. In covalency :
(A) transfer of electrons takes place
(B) one atom acts as donor and other acts as acceptor
(C) the electrons are shared by one atom
(D) equal sharing of electrons takes place between two atoms

3. Element A has 3 electrons in the outermost orbit and element B has 6 electrons in the outermost orbit. The formula of the compound formed between A and B would be:
(A) A_2B_3
(B) A_2B_6
(C) A_2B
(D) A_3B_2

4. A lone pair of electrons in an atom implies
(A) a pair of valence electrons
(B) a pair electrons
(C) a pair of electrons involved in bonding
(D) a pair of valence electrons not involved in bonding

5. In OF_2, number of bond pairs and lone pairs of electrons are respectively
(A) 2, 6
(B) 2, 8
(C) 2, 10
(D) 2, 9

6. The number of lone pairs on Xe in XeF_2, XeF_4 and XeF_6 respectively are:
(A) 3, 2, 1
(B) 2, 4, 6
(C) 1, 2, 3
(D) 6, 4, 2

7. Which of the following sets which one does NOT contain isoelectronic species?
(A) BO_3^{3-}, CO_3^{2-}, NO_3^-
(B) SO_2^{2-}, CO_3^{2-}, NO_3^-
(C) CN^-, N_2, C_2^{2-}
(D) PO_4^{3-}, SO_4^{2-}, ClO_4^-

8. Among the following, the electron deficient compound is
(A) CCl_4
(B) PCl_5
(C) $BeCl_2$ solid
(D) BCl_3

9. An electrovalent bond is formed between:
(A) two electronegative atoms
(B) two metals
(C) electropositive and electronegative atoms
(D) two electropositive atoms

10*. Most favourable conditions for electrovalent bonding are:
(A) low ionization potential of one atom and high electron affinity of the other atom
(B) high electron affinity and high ionization potential of both the atoms
(C) low electron affinity and low ionization potential of both the atoms
(D) high ionization potential of one atom and low electron affinity of the other atom

ASSIGNMENT-3

1. Both ionic and covalent bonds are present in:
   (A) CH_4 (B) NaOH
   (C) KCl (D) SO_2

2. The types of bond present in NH_4Cl are
   (A) electrovalent, covalent and coordinate (B) only ionic
   (C) only covalent (D) covalent and coordinate

3. Number of bonds in benzene:
   (A) 6 σ and 3π (B) 3σ and 12π
   (C) 12σ and 3π (D) 6σ and 6π

4. The bond that exists between NH_3 and BF_3 is called:
   (A) electrovalent (B) covalent
   (C) coordinate (D) hydrogen

5. Which of the following structure is linear?
   (A) SO_2 (B) CO_2
   (C) CO_3^{2-} (D) SO_3^{2-}

6. The compound having the largest bond angle is **(NTSE 2006)**
   (A) NH_3 (B) PH_3
   (C) AsH_3 (D) SbH_3

7. Which of these is weakest ? **(NTSE 2005)**
   (A) ionic bond (B) Covalent bond
   (C) Metallic bond (D) Vander Waals forces

8. At ordinary temperature and pressure H_2O is a liquid whereas H_2S is a gas. This is because of the presence of
   (A) Covalent bonds (B) Electrovalent bonds
   (C) Sulpher (D) Hydrogen bonds **(NTSE 2004)**

9. The octet rule is obeyed in
   (A) CO_2 (B) PCl_5
   (C) BCl_3 (D) SF_6

10. Which of the following statement is not correct regarding ionic compounds ?
    (A) they are good conducters of electricity at room temp.
    (B) They are more soluble in polar solvents than in non–polar ones.
    (C) they consists of ions
    (D) They generally have high melting and boiling points.

KEY & HINTS

BASIC CONCEPTS OF CHEMISTRY

Assignment – 1

1. 6.02×10^{23} molecules of CO contains 1 mole
   $\therefore$ 1 molecules of CO contains $= \frac{1}{6.02\times10^{23}}$
   $\therefore$ 6.02×10^{24} molecules of CO contains $= \frac{1}{6.02\times10^{23}} \times 6.02\times10^{24} = 10$
   Hence, 10 mole = 10 g–atom oxygen
2. 1 mole of H_2SO_4 = 98 g
   0.25 mole of $H_2SO_4 = 98 \times 0.25$
   = 24.5 g
3. 4.6×10^{22} atoms of an element weight = 13.8 g
   $\therefore$ 6.02×10^{23} atoms of an element weight $= \frac{13.8}{4.6\times10^{22}} \times 6.02\times10^{23} = 180$
4. (C)
   $2 \times$ V.D. = Molecular mass
   $2 \times 22 = 44$ g
5. (B)
   $(CH_2O_2) \times n$ = mol. Formula
   where n = 1, 2, 3
6. (B) 22400 ml = 6.02×10^{23}
   $1.12\times10^{-7}\text{ml} = \frac{6\times10^{23}}{22400} \times 1.12\times10^{-7}$
   $= 3.01 \times 10^{12}$
7. (B)

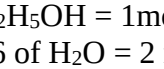
Element	%	% / at. mass	Simple ratio
X	50	50/10 = 5	2
Y	50	50/20 = 2.5	1

Hence the simplest formula = X_2Y

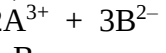
8. (D)
   0.1 g atom of carbon
   $\Rightarrow$ 0.1 mole of carbon = **1.2 g**
   0.1 mol of ammonia $\Rightarrow$ **1.7 g**
   6.02×10^{22} molecules of H_2 gas
   $\Rightarrow$ $\frac{2}{6.02\times10^{23}} \times 6.02\times10^{22} = \frac{2}{10} = 0.2\text{g}$
   1120 ml of CO_2
   $\Rightarrow$ 22400 ml (at S.T.P.) = 44 g
   Hence 1120 ml (at S.T.P.) = $\frac{44}{22400} \times 1120$ **= 2.2g**
9. (D) 0.22 g of CO_2
   $\Rightarrow$ $\frac{0.22}{44}\text{mole} = \frac{1}{200}\text{mole}$
   0.01 g of hydrogen $\Rightarrow$ $\frac{0.01}{2} = \frac{1}{200}\text{mole}$
   0.085 g of $NH_3 \Rightarrow \frac{1}{200}\text{mole}$
10. (A) no. of mole = $\frac{\text{wt.}}{\text{mol.wt.}} = \frac{13.5}{135} = 0.1$

Assignment – 2

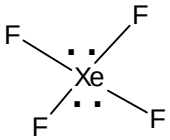
1. (C)
   28 g of CO = 1 mole = N_A 46 g of C_2H_5OH = 1mole = N_A
   36 of H_2O = 2 mole = $2N_A$ 27 g of N_2O_5 = 0.025 mole = 0.025 N_A
2. (D)
3. (A)
   $2A^{3+} + 3B^{2-}$
   A_2B_3
4. (D)
5. (B)



   lone pairs = 8
   bond pairs = 2
6. (A)



   Xe have 3 lone pairs



Xe have 2 lone pairs

XeF_6 Xe have 1 lone pair

7. (B)
SO_2^{2-} have 20 valence electrons where CO_3^{2-} and NO_3^- have 24 valence electrons

8. (D)
In BCl_3, B have only 6 electrons.
$BeCl_2$ solid form dimer

9. (C) Based on theory

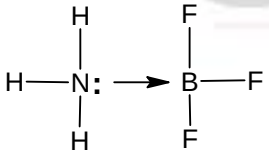
10. (A) Based on theory

Assignment – 3

1. (B)
NaOH showing both ionic and covalent bonds.

2. (A) Electrovalent, covalent and coordinate bonds are present in NH_4Cl

3. (C) Benzene have 9 single bonds and 3 double bonds. Single bond has one sigma bond and double bond have 1 sigma and 1 pie bond.
Total sigma bond = 12, pie bond = 3

4. (C) 

5. (B) O = C = O i.e. linear, but SO_2 is not linear
Because 'S' have 1 lone pair which distorted bond angle

6. (A) Bond angle is in decreasing order $NH_3 > PH_3 > AsH_3 > SbH_3$

7. (D) Bond strength in decreasing order
Ionic bond > Covalent bond > Metallic bond > Vander waals forces

8. (D) H_2O will form inter molecular hydrogen bond while H_2S is not.

9. (A)
In PCl_5, BCl_3, SF_6 here central atom P, B, S have 10, 6, 12 shared electrons.

10. (A)