

## CHEMISTRY

### 1. States of Matter:

Solid, liquid and gaseous states(general characteristics), Change of states, effect of change of temperature and pressure on states of matter, evaporation, mixture, solution, its type and its concentration(percent strength by volume and by mass), separation of components from a mixture using various physical methods, physical and chemical changes, compounds

**Reference :**

1. NCERT class IX text book (chapter 1 and 2)
2. Foundation Science(Chemistry) :Class IX  
A. K. Singh ( Bharati Bhawan)  
(Chapter 1 and 2)
3. A New Approach to ICSE Chemistry(Part-I)  
Goyal Brothers Prakashan  
(Chapter 1 and 3)

### 2. Atoms and Molecules/ Mole concept

Dalton's atomic theory, atoms, molecules, ions, atomic mass, molecular mass, atomic scale units(CU), symbol, valency of ions, formula of ions and compounds, Writing ionic equations, solubility rules, balancing simple equations by hit and trial method, Mole concept, percentage composition of elements in compounds, empirical and molecular formula determination

**Reference:**

1. NCERT class IX text book(chapter 3)
2. Foundation Science(Chemistry) :Class IX  
A. K. Singh ( Bharati Bhawan)(Chapter 3)
3. A New Approach to ICSE Chemistry(Part-I)  
Goyal Brothers Prakashan (Chapter 4)
4. **UTSE study materials given below**

### 3. Atomic Structure :

Fundamental particles of atom(electron, proton, neutron), Discovery of Nucleus by Gold foil experiment, Rutherford's model of atom and its demerits, Bohr's model of atom atomic number, mass number, isotopes, isobars, Distribution of electrons in shells, Valency

**Reference :**

1. NCERT class IX text book(Chapter 4)
2. . Foundation Science(Chemistry) :Class IX  
A. K. Singh ( Bharati Bhawan) (Chapter 4)
3. A New Approach to ICSE Chemistry(Part-I)  
Goyal Brothers Prakashan (Chapter 7)
4. **UTSE study materials given below**

### 4. Periodic Classification

Periodic classification of elements: Salient Features of modern periodic(Long form), Periodic law, acquainting with alkali metals, alkaline earth metals, boron, carbon, nitrogen, oxygen, halogen, noble gas family members, Variation of periodic properties such as atomic size, ionisation energy, electron affinity, electronegativity, metallic and nonmetallic character, valency among normal elements.

**Reference :**

1. A New Approach to ICSE Chemistry(Part-I)  
Goyal Brothers Prakashan (Chapter 8)
2. **UTSE study materials given below**

### 5. Chemical Bonding:

Ionic bond and properties of ionic compounds, covalent bond(nonpolar and polar types), Lewis structures of simple covalent molecules, properties of covalent compounds, Preliminary ideas of metallic bond and intermolecular forces.

**Reference:**

**1. UTSE Study materials given below**

**6. Oxidation-Reduction:**

Oxidation-reduction in terms of loss and gain of electrons, OXIDATION NUMBER(ON) or OXIDATION STATE(OS).

**Reference: 1. UTSE Study materials given below**

**7. Chemical Reactions:**

Types: (a) Redox and (b) Non-redox(Metathesis) types

*Types of metathesis reactions:* Double displacement, Neutralisation and precipitation reactions, reactions of carbonates, sulphites, sulphides and nitrites with dilute HCl/H<sub>2</sub>SO<sub>4</sub>, reaction of nonmetallic oxides with water, reactions of metallic oxides with water, heating of ammonium salts with a base, thermal decomposition of carbonates and bicarbonates, hydrolysis of nitrides, phosphides, carbides, sulphides; amphotericism of certain metallic and nonmetallic oxides and hydroxides.

*Types of some simple redox reactions:* Displacement reactions(metal activity series and halogen activity series), synthesis, analysis or decomposition, combustion and some simple redox reactions.

**Reference: 1. UTSE Study materials given below**

**PHYSICS:**

**Motion** : Motion in one dimension: uniform and non-uniform motion, distance, displacement, velocity, acceleration, distance-time and velocity-time graphs for uniform and uniformly accelerated motion, derivation of equations of motions by graphical method, acceleration due to gravity, free fall motion. **2.**

**Forces** : Contact and non-contact forces, friction-factors affecting friction, sliding and rolling friction, advantages and disadvantages of friction, control of friction. **3. Newtons Laws of Motion** : 1st, 2nd and 3rd laws- inertia of a body, inertial mass, momentum, force and acceleration, conservation of momentum, action and reaction forces, mass and weight. Law of gravitation, **4. Heat** : Concepts of heat and temperature, Temperature scales(Celsius, Fahrenheit and Kelvin), mercury thermometer, clinical thermometer, Specific heat capacities, change of states and Latent heats of fusion and vaporization, calculation of heat lost or gained by method of mixtures, Thermal expansion of solids, liquids and gases(simple idea), coefficient of linear and volume expansion, humidity and relative humidity. **5. Wave motion and Sound** : General concepts of wave motion, Simple harmonic oscillation(simple idea) with graphical representation, types of waves and their examples, amplitude, frequency, wavelength, velocity of a wave and their inter-relationship, Nature of sound and its propagation, range of hearing in humans, effects of medium on sound waves, Reflection of sound: echo, SONAR, pitch, loudness and quality.

**Reference books:**

1. NCERT class IX text book
2. A New Approach to ICSE Physics(Part I) for Class IX : Goyal Brother Prakashan
- 3. Foundation Science: PHYSICS by Prof. H.C. Verma(IIT, Kanpur)  
for class IX (Bharati Bhawan)**

*(The students are strongly requested to get a good concept of each topic for UTSE first from the physics book written by Prof. H.C. Verma(Bharati Bhawan). This is the No.1 book in physics for concept building for high school students).*

**IMPORTANT NOTE:** 1. Previous Years Questions(last three years) and their detailed solution in each topic will be uploaded by December 30, 2010. Model Question for UTSE-2011 will also be uploaded by the same time. The examinees are required to thoroughly prepare the topics from the reference books and study materials given below.

2. Readers are requested to inform us immediately(theuraniumodisha@gmail.com OR theuranium@dataone.in) when they find any mistake in the study materials for verification and rectification.

## (ELECTRONIC CONFIGURATION IN ATOMS)

Each shell is divided into a fixed number of subshells. Shells are designated by shell no. (principal quantum number) 1, 2, 3..... while the subshells are designated by subshell numbers (azimuthal quantum numbers) 0, 1, 2, 3..... The following table gives the shells and subshells with their quantum numbers. The number of subshells (i.e. the number of l values) is equal to shell number (Principal quantum number) n.

Shell Number(n): (Principal Quantum No.)	Subshell No.(l) (Azimuthal Quantum No.)		
n=1	K- Shell	l=0	s-subshell
n=2	L- Shell	l=1	p-subshell
n=3	M-shell	l=2	d-subshell
n=4	N-Shell	l=3	f-subshell
K- Shell(n=1)	1- subshell –	s-subshell (l=0) ( K shell has one subshell i.e s-subshell)	
L- Shell (n=2)	2-subshells –	s-subshell (l=0) & p-subshell (l=1) ( L shell has two subshells i.e s- and p-subshells)	
M- Shell(n=3)	3-subshells -	s-subshell (l=0), p-subshell (l=1) & d-subshell (l=2) (M shell has three subshells i.e s, p and d-subshells)	
N- Shell(n=4)	4-subshells -	s-subshell (l=0), p-subshell (l=1) d-subshell (l=2) f-subshell (l=3) (N shell has four subshells i.e s, p, d and f subshells)	

**ORBITALS:** Each Subshell is made of some orbitals (no. of orbitals =  $2l + 1$ )

Electrons live in orbitals i.e the rooms in which electrons stay.

s-subshell - 1 orbital - s-orbital (s-subshell is the s-orbital itself while K shell is the s-orbital itself)

p-subshell - 3 orbitals -  $p_x$ ,  $p_y$  and  $p_z$  orbitals

d-subshell - 5 orbitals - 5 d orbitals (names not given)

f-subshell - 7 orbitals - 7 f orbitals (names not given)

### PAULI'S EXCLUSION PRINCIPLE:

One orbital can at the most accommodate two electrons with opposite spin i.e if one spins clockwise then the other has to spin anticlockwise. The box diagram below shows an electron pair present in any orbital.



Each s-orbital has two electrons, each p-subshell has  $3 \times 2 = 6$  electrons, each d-subshell has  $5 \times 2 = 10$  electrons while each f-subshell has  $7 \times 2 = 14$  electrons.

### MAXIMUM NUMBER OF ELECTRONS IN EACH SHELL

1(K) SHELL: 2,      2(L) SHELL:  $2 + 6 = 8$ ,      3(M) SHELL:  $2 + 6 + 10 = 18$

4(N) SHELL:  $2 + 6 + 10 + 14 = 32$

Maximum number of electrons in shell =  $2n^2$  (where n = shell number or principal quantum number)

**DEGENERATE ORBITALS:** Orbitals having same energy are called degenerate orbitals.

3- p orbitals are degenerate, 5-d orbitals are degenerate, 7-f orbitals are degenerate

**HUND'S RULE:** All the degenerate orbitals are filled with one electron each with parallel spin and then pairing of electrons occurs

Example : Suppose we have to fill three electrons in p-subshell which contains three orbitals. We have to fill each orbital with one electron with same spin (either all clockwise or all anticlockwise). If there are 4 electrons, then the 4th electron will be paired with the first one and so on.





# Valency

## Modern Definition of Valency

Valency is the power of an atom of an element to combine with other atoms measured by the number of electrons which an atom or radical will lose, gain or share to form a chemical compound.

Valency is of two types:

(i) **Electrovalency** (shown in ionic compounds eg. NaCl, K<sub>2</sub>SO<sub>4</sub> etc) (ii) **Covalency** (shown in covalent compounds eg SO<sub>2</sub>, NH<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub> etc)

## Electrovalency:

### Radicals:

Radical is an atom or a group of atoms which form an ion (positive or negative) by loss or gain of electrons. It is part of an ionic compound. Two radicals form a compound.

There are two radicals

(i) **Basic Radical**: This forms the +ve ion or cation part of the compound. Usually metals e.g Na<sup>+</sup>, K<sup>+</sup>, Ca<sup>2+</sup>, Al<sup>3+</sup>, Ag<sup>+</sup>, Hg<sup>2+</sup>, Fe<sup>3+</sup> etc. (exception: NH<sub>4</sub><sup>+</sup>, H<sup>+</sup>). This part comes from a base.

(ii) **Acid Radical**: This forms the -ve ion or anion part of the compound. Usually it is made up of nonmetals.

(i) **monoatomic acid radical**: Cl<sup>-</sup>, Br<sup>-</sup>, S<sup>2-</sup>, O<sup>2-</sup>, N<sup>3-</sup>, P<sup>3-</sup>, H<sup>-</sup> etc

(ii) **Compound acid radical**: SO<sub>4</sub><sup>2-</sup>, PO<sub>4</sub><sup>3-</sup>, CO<sub>3</sub><sup>2-</sup>, SO<sub>3</sub><sup>2-</sup>, OH<sup>-</sup>, NO<sub>2</sub><sup>-</sup>, NO<sub>3</sub><sup>-</sup> etc. (exceptions: MnO<sub>4</sub><sup>2-</sup>, CrO<sub>4</sub><sup>2-</sup>, AlO<sub>2</sub><sup>-</sup> etc which contain both metal and non-metal)

Charge of Basic Radical = + valency,

Charge of Acid Radical = - valency

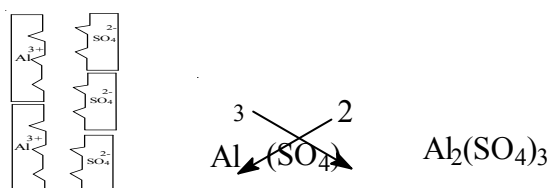
Task: Write the ionic representations of the following radicals: Al, Fe(ous), Cu(ic), Sn(ous), NH<sub>4</sub>, Br, O, SO<sub>4</sub>, NO<sub>2</sub>, MnO<sub>4</sub>, Cr<sub>2</sub>O<sub>7</sub>, C, N, PO<sub>4</sub>

## Chemical Formula

Formula is the simplest representation of a compound containing both basic radical (written in LHS) and acid radical (written in RHS). It represents the simplest ratio of atoms of different elements present in the molecule. In fact this is called the *empirical formula* for ionic compounds.

## Cross-Over Rule:

The charge (valency) of the cation becomes the subscript for the anion and the charge (valency) of the anion becomes the subscript for the cation. ('Crossover rule'). A compound radical having a subscript is written inside a bracket.



This is done for charge balance, i.e net charge of the molecule should be zero.

The hollows (depressions) in the above picture of the basic radical is the site from which electrons have been lost and the mount (bulging out) portions in the acid radical are the sites at which electron have been accepted. There must be a complete matching between the hollows with the mounts.

## Formal Writing For Practice:

Write down the formula of the following:

Set-1 (i) Magnesium Chloride, (ii) Sodium sulphate, (iii) Calcium Nitrate, (iv) Ferric Sulphide, (v) Cupric Oxide, (vi) Potassium sulphite, (vii) Sulphuric acid, (viii) Aluminium Phosphate, (x) Ammonium cyanide

Set-2: Zinc carbonate, Cuprous sulphide, Ammonium sulphate, Potassium oxide, Mercuric Sulphate, Ferrous Sulphide, Sodium carbonate, Calcium Nitrite, Stannous Chloride, Nitric Acid

Set-3: Write the formula of the following.

Potassium Manganate, Ferric sulphate, Mercuric Nitrate, Ammonium bicarbonate, Aluminium Bromide, zinc sulphite, Nitrous acid,

Set-4: Sodium Chromate, Magnesium nitrate, Potassium phosphate, calcium phosphide, Stannous sulphide, Plumbous(lead)carbonate, Ferrous Nitrite, Mercurous hydrogen carbonate, potassium chlorate, Hydrobromic acid

Set-5: Sodium phosphate, Calcium Sulphate, Potassium dichromate, Magnesium bicarbonate, Stannic chloride, Aluminium carbide, Silver nitrite, Cobalt(ous) sulphide, Ferric sulphite, calcium manganate

Set-6: Magnesium sulphite, Sodium cyanide, Ferrous sulphate, Chromic hydroxide, Zinc bicarbonate, Potassium Ferrocyanide, Ammonium dichromate, Arsenous nitrate, cupric(copper)bicarbonate, Ferric ferrocyanide, Cupric ferrocyanide

### Ionisation of compounds:

The basic radical gives the +ve ion and the acid radical gives the -ve ion.



(Note that for compound radical, the individual atoms are not separated, eg.  $\text{Cr}_2\text{O}_7^{2-}$  is the -ve ion from which Cr and O atoms cannot be separated. Only the basic radical is separated from acid radical).

### Solubility of solids in water:

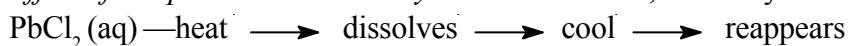
If a substance dissolves in water in large quantity it is called a soluble substance. Eg. NaCl, Sugar,  $\text{NaNO}_3$ ,  $\text{H}_2\text{SO}_4$  etc.

If a substance dissolves in water less than 0.001 mole per litre of water, it is conventionally called an insoluble substance. If more than that dissolves in water we call it soluble substance.

*Definition:* The amount of solute that can dissolve in 100gm water to produce a saturated solution at particular temperature is called solubility of that substance in water.

(Ex. NaCl: 36gm/100gm, Potassium dichromate: 58gm/100gm at room temp)

*Effect of Temperature on Solubility:* For most solutes, solubility increases with temp.



### Solubility Rules

(A)	Soluble Category	Exceptions: (Insoluble)
1.	Inorganic acids (ALL) (HCl, $\text{H}_2\text{SO}_4$ , $\text{HNO}_3$ etc.)	-
2.	$\text{Na}^+$ , $\text{K}^+$ and $\text{NH}_4^+$ salts(ALL)	-
3.	$\text{NO}_3^-$ (Nitrates) (ALL)	-
4.	$\text{SO}_4^{2-}$ (Sulphates)	$\text{Pb}^{2+}$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$ , $\text{Hg}^{2+}$ , $\text{Ag}^+$
5.	$\text{Cl}^-$ (Chloride), $\text{Br}^-$ (Bromide) $\text{I}^-$ (Iodide)	$\text{Hg}_2^{2+}$ (ous), $\text{Ag}^+$ , $\text{Pb}^{2+}$ ,

Task: Indicate whether the following are soluble(S) or insoluble(I) in water.

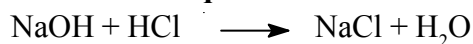
(i)HCl (ii) $\text{NH}_4\text{Cl}$  (iii) $\text{PbSO}_4$  (iv) $\text{Ca}(\text{NO}_3)_2$  (v) $\text{Hg}_2\text{Cl}_2$  (vi) $\text{Na}_2\text{SO}_4$  (vii) $\text{BaSO}_4$ , (viii) $\text{K}_3\text{PO}_4$   
(ix) $\text{H}_3\text{PO}_4$  (x) $\text{AgBr}$  (xi)  $\text{H}_2\text{SO}_4$

(B)	Insoluble Category	Exceptions(Soluble)
1	F-(Fluorides)	$\text{Na}^+$ , $\text{K}^+$ , $\text{NH}_4^+$ , $\text{Ag}^+$
2.	$\text{OH}^-$ (Hydroxides)	$\text{Na}^+$ , $\text{K}^+$ , $\text{NH}_4^+$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$
3.	$\text{CO}_3^{2-}$ (carbonates) $\text{PO}_4^{3-}$ (Phosphates)	$\text{Na}^+$ , $\text{K}^+$ , $\text{NH}_4^+$ , $\text{Be}^{2+}$
4.	$\text{S}^{2-}$ (Sulphides)	$\text{Na}^+$ , $\text{K}^+$ , $\text{NH}_4^+$ , $\text{Mg}^{2+}$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$

Task: Indicate which are insoluble(I) and which are soluble(S) in water.

(i) $\text{Fe}(\text{OH})_3$  (iii) $\text{MgCO}_3$  (iv) $\text{CaCO}_3$  (v) $\text{CaS}$  (vi) $\text{NaOH}$  (viii) $\text{Ca}_3(\text{PO}_4)_2$   
(x) $\text{CuS}$ (viii)  $\text{K}_2\text{S}$  (ix) $\text{HNO}_3$

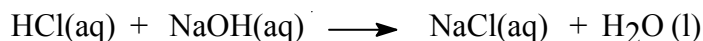
### Chemical and Ionic Equations:



### Ionic equations:

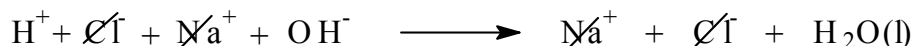
Substance which is soluble in water is given the symbol (aq.) . Aq. stands for aqueous. The solid which is insoluble in water is given the symbol (s) which stands for solid. The pure liquid and gases are given the symbol (l) and (g) respectively. Note that only the substance followed by the symbol (aq.) dissolves in water and undergo ionization.

*Acid-Base Reactions:*



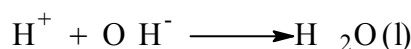
The compound with (aq) symbol is ionized and total ionic equation written.

*Total Ionic Equation(TIE)*

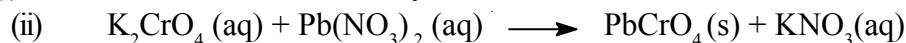
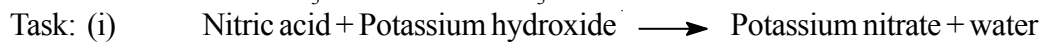
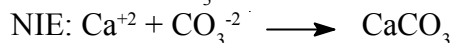
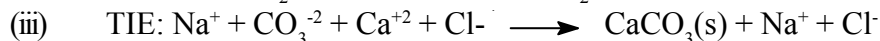
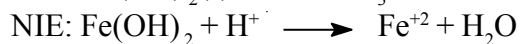
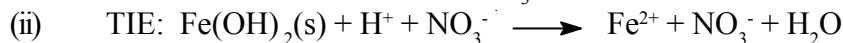
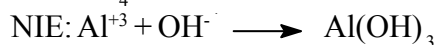
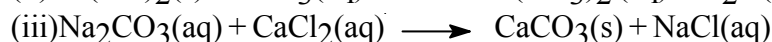
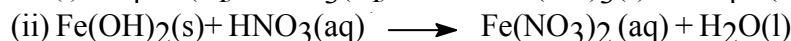
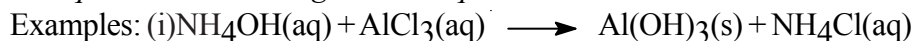


The ions present common to both LHS and RHS are called *spectator ions* which are cancelled to get the net ionic equation.

*Net Ionic Equation(NIE)*



*N. B: Do not write the number of each ion while writing the total and net ionic equations. Balancing the equations at this stage is not required.*



**(PLEASE**

**SEE**

**BELOW**

**FOR**

**MORE )**



## PERCENT COMPOSITION, EMPIRICAL AND MOLECULAR FORMULA

### PERCENTAGE COMPOSITION

The percentage by mass of each element present in a compound can be found out and compared. Take the simplest case of  $H_2O$ . What is the percentage of H and O present in  $H_2O$ ? If in every 18gm(M.M) of  $H_2O$ , there is 16gms of O and 2gm of H, For 100gms of  $H_2O$ , there is  $(16/18) \times 100 = 88.89$ gms of O and  $(2/18) \times 100 = 11.11$ gms of H. So % of O = 88.89% and that of H=11.11%. *Mass of an element present in 100gms of the compound gives the percentage composition of that element.*

(SAQ :

**Practice : 1:** Calculate the percent by mass of O in  $Ca(ClO_3)_2$ .

2: Find the percent of nitrogen in (i) $NH_4NO_3$  and (ii) $(NH_4)_2SO_4$  and indicate which is a better nitrogenous fertilizer?

### EMPIRICAL FORMULA AND MOLECULAR FORMULA:

Empirical formula of a compound is the simplest whole number ratio of atoms which are present in a molecule. For example glucose has the molecular formula  $C_6H_{12}O_6$  but its empirical formula will be obtained by dividing all the coefficients by the highest common factor(HCF) 6. Hence its empirical formula is  $CH_2O$ . Molecular Formula, on the other hand, represents actual number of atoms of each kind present in a molecule. Empirical formula is related to the molecular formula of a compound as follows.

**Molecular Formula = (Empirical Formula)<sub>n</sub>** (Where n is a whole number)

**Practice: 1** The molecular formula of a compound is  $C_7H_{14}$ , find its empirical formula.

2. : The molecular formula of a compound is  $C_4H_8O_2$ . What is its empirical formula?

### Empirical formula Mass:

The mass(relative)obtained from the empirical formula is called empirical formula mass. For example, the empirical formula mass of  $CH_2O$  is  $(12+2+16)=30$ , but the molecular mass is mass obtained from the molecular formula. The molecular mass of  $C_6H_{12}O_6 = 72+12+96=180$ . So if you know the molecular mass of a compound and its empirical formula mass, you can easily find the molecular formula. Look to this example.

*Example: The empirical formula of a compound is  $CH_2O$  and its molecular mass is 180, find the molecular formula.*

**Solution:** The empirical formula mass =  $12+2+16=30$

We know that Molecular Formula = (Empirical Formula)<sub>n</sub>:(Where n is a whole number)

So molecular mass = n X (empirical formula mass)

n = molecular mass/empirical formula mass =  $180/30 = 6$

So molecular formula =  $(CH_2O)_6 = C_6H_{12}O_6$

### Determination of Empirical Formula:

The empirical formula of a compound can be calculated from the percent composition of the compound. Read this example.

*Example: Find the empirical formula of a compound which contains 60% O and 40% S by mass. If its molecular mass is 80, what is its molecular formula.*

**Solution:** Let us take 100gms of the compound. This amount contains 60gms of O and 40gms of S. Let us calculate how many gm. atoms(mole of atoms) of each element present in these amounts and what is their ratio. This is obtained by dividing the mass of the element by the atomic mass of the element. The ratio of gm atoms is same as ratio of the atoms present in the formula, hence it will give the empirical formula. In this case,

the no. of gm. atoms of O=  $60/16=3.75$

the no. of gm. atoms of S=  $40/32= 1.25$

So the ratio of gm. atoms and hence atoms of O and S present in the formula = $3.75:1.25$

But you know that in the formula(empirical or molecular) there are whole number of atoms of the elements. So we have to make this the simplest whole number ratio. This is done by dividing each value by the lowest of the values. In this case the lowest is 1.25.

So whole number ratio =  $3.75/1.25 : 1.25/1.25 = 3 : 1$

Hence the empirical formula =  $O_3S$  or  $SO_3$  : So its empirical formula mass= $32+48=80$

$n = \text{Molecular mass} / \text{empirical formula mass} = 80/80 = 1$   
 So Molecular Formula = (Empirical Formula)<sub>n</sub>,  
 So Molecular Formula = (SO<sub>3</sub>)<sub>1</sub> = SO<sub>3</sub> (same as its empirical formula).

### STEPS FOR FINDING EMPIRICAL FORMULA:

- (i) First the percent composition data of the elements are divided by the respective atomic masses to get ratio of atoms present in the formula
  - (ii) Make this ratio the simplest whole number ratio by dividing all the values by the lowest of the values. If this does not give the whole number ratio, multiply 2, 3 etc. to get the simplest whole number ratio.  
For example, if it comes 1:1.5:2.5, you multiply throughout by 2 to make the ratio 2:3:5.
  - (iii) Then empirical formula is written by placing these whole numbers as coefficients of the respective elements.
  - (iv) Empirical formula mass is then found out. Then divide the molecular mass by the empirical formula mass to find the value of n. Usually the division does not give a whole number value to 'n'. Round off to its nearest whole number to get the value of 'n'.
  - (v) Multiply 'n' in all the coefficients of the empirical formula to get the molecular formula.
- To determine the empirical formula, it is better to calculate in tabular form. The previous example of SO<sub>3</sub> is given in the table for better understanding.

Element	%	At. Mass	Relative number of atoms	Simple Whole No. of atoms	Empirical Formula
O	60	16	60/16=3.75	3.75/1.25=3	SO <sub>3</sub>
S	40	32	40/32=1.25	1.25/1.25=1	

Let us take another example.

*Example:* A compound gave on analysis the following percent composition: K=26.57%, Cr=35.36%, O=38.07%. Derive the empirical formula.

**Solution:** Let us take the help of the table for easy calculation.

Element	%	At. Mass	Relative no. of atoms	Simple Whole No. of atoms	Empirical Formula
K	26.57	39	26.57/39=0.68	0.68/0.68=1X2= 2	K <sub>2</sub> Cr <sub>2</sub> O <sub>7</sub>
Cr	35.36	52	35.36/52=0.68	0.68/0.68=1X2= 2	
O	38.07	16	38.07/16=2.379	2.379/0.68=3.5X2=7	

Remember that fractional values like 3.95 or 3.99 which is close to a whole number is rounded off to the nearest whole number(4). But if it is 1.5 or 2.25 etc. you cannot convert it to the nearest whole number. In that case you have to multiply suitable minimum factor to make it a whole number. For example to convert, 1.5 to a whole number, you have to multiply with 2. To convert 2.25 to a whole number you have to multiply with 4 and so on. Therefore if in the first attempt you did not get a simple whole number ratio of atoms(5th column), you will get a whole number ratio in your second attempt by multiplying suitable lowest factor to make all the figures as whole numbers. In this example, the first ratio was 1:1:3.5 and we got the simple whole number ratio by multiplying with 2 i.e 2:2:7.

Note that the empirical formula, in this example, appears to be the molecular formula as it is the familiar compound potassium dichromate. However for determining molecular formula, it is necessary to have the molecular mass of the compound.

**Practice 1:** Find the empirical formula of a hydrocarbon that on analysis gave the following percent composition: C=85.63%, H=14.37%. If the molecular mass of the compound is 56, what is its molecular formula.

**2:** An oxide of nitrogen contains 30.4% nitrogen. What is its empirical formula?

**3:** Determine the simplest formula of a compound that has Cr=26.52%, S=24.52% and O=48.96%. Could you suggest the formula in proper order i.e in terms of acid and basic radicals with proper valences?

**Periodic Table(Long Form)  
Atomic Number of Elements**

<b>H</b> 1																<b>He</b> 2	
<b>Li</b> 3	<b>Be</b> 4											<b>B</b> 5	<b>C</b> 6	<b>N</b> 7	<b>O</b> 8	<b>F</b> 9	<b>Ne</b> 10
<b>Na</b> 11	<b>Mg</b> 12											<b>Al</b> 13	<b>Si</b> 14	<b>P</b> 15	<b>S</b> 16	<b>Cl</b> 17	<b>Ar</b> 18
<b>K</b> 19	<b>Ca</b> 20	<b>Sc</b> 21	<b>Ti</b> 22	<b>V</b> 23	<b>Cr</b> 24	<b>Mn</b> 25	<b>Fe</b> 26	<b>Co</b> 27	<b>Ni</b> 28	<b>Cu</b> 29	<b>Zn</b> 30	<b>Ga</b> 31	<b>Ge</b> 32	<b>As</b> 33	<b>Se</b> 34	<b>Br</b> 35	<b>Kr</b> 36
<b>Rb</b> 37	<b>Sr</b> 38	<b>Y</b> 39	<b>Zr</b> 40	<b>Nb</b> 41	<b>Mo</b> 42	<b>Tc</b> 43	<b>Ru</b> 44	<b>Rh</b> 45	<b>Pd</b> 46	<b>Ag</b> 47	<b>Cd</b> 48	<b>In</b> 49	<b>Sn</b> 50	<b>Sb</b> 51	<b>Te</b> 52	<b>I</b> 53	<b>Xe</b> 54
<b>Cs</b> 55	<b>Ba</b> 56	<b>La*</b> 57	<b>Hf</b> 72	<b>Ta</b> 73	<b>W</b> 74	<b>Re</b> 75	<b>Os</b> 76	<b>Ir</b> 77	<b>Pt</b> 78	<b>Au</b> 79	<b>Hg</b> 80	<b>Tl</b> 81	<b>Pb</b> 82	<b>Bi</b> 83	<b>Po</b> 84	<b>At</b> 85	<b>Rn</b> 86
<b>Fr</b> 87	<b>Ra</b> 88	<b>Ac**</b> 89	<b>Rf</b> 104	<b>Ha</b> 105	<b>Sg</b> 106	<b>Ns</b> 107	<b>Hs</b> 108	<b>Mt</b> 109	<b>Ds</b> 110	<b>Rg</b> 111	<b>Cn</b> 112	<b>Uut</b> 113	<b>Uuq</b> 114	<b>Uup</b> 115	<b>Uuh</b> 116	<b>Uus</b> 117	<b>Uuo</b> 118

<b>*Ce</b> 58	<b>Pr</b> 59	<b>Nd</b> 60	<b>Pm</b> 61	<b>Sm</b> 62	<b>Eu</b> 63	<b>Gd</b> 64	<b>Tb</b> 65	<b>Dy</b> 66	<b>Ho</b> 67	<b>Er</b> 68	<b>Tm</b> 69	<b>Yb</b> 70	<b>Lu</b> 71
<b>**Th</b> 90	<b>Pa</b> 91	<b>U</b> 92	<b>Np</b> 93	<b>Pu</b> 94	<b>Am</b> 95	<b>Cm</b> 96	<b>Bk</b> 97	<b>Cf</b> 98	<b>Es</b> 99	<b>Fm</b> 100	<b>Md</b> 101	<b>No</b> 102	<b>Lr</b> 103

## Periodic Table(Long Form)

IA 1	IIA 2											IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	0 18
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg	IIIB 3	IVB 4	VB 5	VIB 6	VII B 7	VIII 8 9 10			IB 11	IIB 12	Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn Ga	Ge	As	Se	Br	Kr	
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac**	Rf	Ha	Sg	Ns	Hs	Mt	Ds	Rg	Cn	Uut	Uuq	Uup	Uuh	Uus	Uuo

*Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
**Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

<span style="display: inline-block; width: 15px; height: 15px; background-color: green; border: 1px solid black;"></span> Liquids	<span style="display: inline-block; width: 15px; height: 15px; background-color: orange; border: 1px solid black;"></span> Gases	<span style="display: inline-block; width: 15px; height: 15px; background-color: cyan; border: 1px solid black;"></span> Radioactive	<span style="display: inline-block; width: 15px; height: 15px; background-color: gray; border: 1px solid black;"></span> natural solid	<span style="display: inline-block; width: 15px; height: 15px; background-color: black; border: 1px solid black;"></span> man-made solid
---	--	--	--	--

(From Ds(110) to Uuo(118) the elements are also synthetic and radioactive. The blue and black shades have not been given unfortunately)

## IMPORTANT DATA FOR 118 ELEMENTS

Symbol	Name	At.No	No. of stable isotopes	av. at. mass
Ac	Actinium	89	—	227
Ag	Argentum / Silver	47	2	108
Al	Aluminium	13	1	27
Am	Americium	95	—	243
Ar	Argon	18	3	40
As	Arsenic	33	1	75
At	Astatine	85	—	210
Au	Aurum / Gold	79	1	197
B	Boron	5	2	11
Ba	Barium	56	7	137
Be	Beryllium	4	1	9
Bh	Bohrium	107	—	272
Bi	Bismuth	83	1	209
Bk	Berkelium	97	—	247
Br	Bromine	35	2	80
C	Carbon	6	2	12
Ca	Calcium	20	6	40
Cd	Cadmium	48	8	112
Ce	Cerium	58	4	140
Cf	Californium	98	—	251
Cl	chlorine	17	2	35.5
Cm	Curium	96	—	247
Cn	Copernicium	112	—	285
Co	Cobalt	27	1	59
Cr	Chromium	24	4	52
Cs	Cesium	55	1	133
Cu	Cuprum / Copper	29	2	63.5
Db	Dubnium	105	—	262
Ds	Darmstadtium	110	—	281
Dy	Dysprosium	66	7	162.5
Er	Erbium	68	6	167
Es	Einsteinium	99	—	252
Eu	Europium	63	2	152
F	Fluorine	9	1	19
Fe	Ferrum / Iron	26	4	56
Fm	Fermium	100	—	257
Fr	Francium	87	—	223
Ga	Galium	31	2	69.7
Gd	Gadolinium	64	7	157
Ge	Germanium	32	5	72.6
H	Hydrogen	1	2	1
He	Helium	2	2	4
Hf	Hafnium	72	6	178.5
Hg	Hydrargyrum / Mercury	80	7	200.6
Ho	Holmium	67	1	165
Hs	Hassium	108	—	270
I	Iodine	53	1	127
In	Indium	49	2	115
Ir	Iridium	77	2	192
K	Kalium / Potassium	19	2	39
Kr	Krypton	36	6	83.8
La	Lanthanum	57	2	139
Li	Lithium	3	2	7
Lr	Lawrencium	103	—	262
Lu	Lutetium	71	2	175
Md	Mendelevium	101	—	258
Mg	Magnesium	12	3	24

Mn	Manganese	25	1	55
Mo	Molybdenum	42	7	96
Mt	Meitnerium	109	–	276
N	Nitrogen	7	2	14
Na	Natrium / Sodium	11	1	23
Nb	Niobium	41	1	93
Nd	Neodymium	60	7	144
Ne	Neon	10	3	20
Ni	Nickel	28	5	58.7
No	Nobelium	102	–	259
Np	Neptunium	93	–	237
O	Oxygen	8	3	16
Os	Osmium	76	7	190
P	Phosphorus	15	1	31
Pa	Protactinium	91	–	231
Pb	Plumbum / Lead	82	4	207
Pd	Palladium	46	6	106
Pm	Promethium	61	–	145
Po	Polonium	84	–	209
Pr	Praseodymium	59	1	141
Pt	Platinum	78	4	195
Pu	Plutonium	94	–	244
Ra	Radium	88	–	226
Rb	Rubidium	37	2	85.5
Re	Rhenium	75	1	186
Rf	Rutherfordium	104	–	261
Rg	Roentgenium	111	–	280
Rh	Rhodium	45	1	103
Rn	Radon	86	–	222
Ru	Ruthenium	44	7	101
S	Sulphur	16	4	32
Sb	Stibium / Antimony	51	2	121.8
Sc	Scandium	21	1	45
Se	Selenium	34	6	79
Sg	Seaborgium	106	–	271
Si	Silicon	14	3	28
Sm	Samarium	62	7	150
Sn	Stannum / Tin	50	10	118.7
Sr	Strontium	38	4	87.6
Ta	Tantalum	73	2	181
Tb	Terbium	65	1	159
Tc	Technetium	43	–	99
Te	Tellurium	52	8	127.6
Th	Thorium	90	–	232
Ti	Titanium	22	5	48
Tl	Thallium	81	2	204
Tm	Thulium	69	1	169
U	Uranium	92	–	238
Uuh	Ununhexium	116	–	293
Uuo	Ununoctium	118	–	294
Uup	Ununpentium	115	–	288
Uuq	Ununquadium	114	–	289
Uuu	Ununtrium	113	–	284
V	Vanadium	23	2	51
W	Wolfram / tungsten	74	5	184
Xe	Xenon	54	9	131
Y	Yttrium	39	1	89
Yb	Ytterbium	70	7	173
Zn	Zinc	30	5	65
Zr	Zirconium	40	5	91

N.B: The elements against which \_ sign is given for number of stable isotopes are radioactive and do not have any stable isotopes.

# PERIODIC CLASSIFICATION

## FEATURES OF LONG FORM PERIODIC TABLE:

[Reader is advised to refer the periodic table while studying on periodic classification]

1. There are 7 horizontal rows called **periods** and 18 vertical columns called **groups**. According to old system, the eighteen columns consist of 16 groups written in roman numerical, IA, IIA and ....IB, IIB..... and according to new system, there are 18 groups for the 18 columns numbered in arabic numerical 1, 2, 3..... serially up to 18. The groups in the old and new systems have been written at the top of each group the periodic table.

2. The whole periodic table is divided into the following four portions.

(a) **Left Portion:** This portion consists of two groups: IA and IIA(1 and 2)

**IA(1):** These are called **alkali metals**(except H): Li, Na, K, Rb, Cs, Fr. These metals form alkalis or soluble bases in aqueous solution and hence the name alkali metals.

**IIA(2):** These are called **alkaline earth metals**: Be, Mg, Ca, Sr, Ba and Ra. The oxides of these metals also are bases and hence called alkaline earth metals.

(b) **Right Portion:** This portion consists of six groups IIIA, IVA, VA, VIA, VIIA and zero.

**IIIA(13):** This is called **Boron Family** and consist of B, Al, Ga, In, Tl and Uut.

**IVA(14):** This is called the **Carbon Family** and consist of C, Si, Ge, Sn, Pb and Uuq

**VA(15):** This is called the **Nitrogen Family** and consists of N, P, As, Sb, Bi and Uup.

These are commonly called **pnictogens** (Greek : *pnictogen*- choking. Compounds of P, As, Sb and Bi are called pnictides)

**VIA(16):** This is called the **Oxygen Family** and consists of O, S, Se, Te, Po and Uuh

These are commonly called **chalcogens**.(Greek: *chalcogen* - ore former)

**VIIA(17):** These are called **Halogens** and consist of F, Cl, Br, I and At.(Greek: halogen - sea salt former)

**Zero(18):** This group consist of **noble or inert gases** He, Ne, Ar, Kr, Xe, Rn and Uuo

The elements of left and right portions are commonly called **normal or representative elements**.

(c) **Middle Portion:** This consists of eight groups according to old system. From left to right these groups are IIIB, IVB, VB, VIB, VIIB, VIII, IB and IIB. Note that group VIII has not been given any subgroup but consists of three columns, the Fe, Co and Ni groups which are commonly called group VIII elements. In the new system there are ten groups in the middle portion and the groups starts from group-3 and ends at group-12. Group IIIB in old system is called the group 3 in new system, group IVB is group 4 and so on, up to group IIB which is group 12 according to the new system. All the elements belonging to the middle portions are commonly called **transition elements**. This is because they appear in between group 2(alkaline earth metals) and 13(boron family) and the electrons are filled in the 3rd shell (3d subshell) while the 4th shell (valence shell) remaining filled already(4s). In fact these are called the d-block elements. The modern definition of a transition metal is beyond the scope of this study material. The transition elements are also studied in terms of periods. There are four series of transition elements. Each series consists of ten elements.

**1st transition series**(from left to right): Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu and Zn(from atomic number 21 to 30)

**2nd Transition series:** (from left to right): Y, Zr, Nb, Mo, Tc, Ru, Rh, Pd, Ag, Cd

**3rd Transition series:** (from left to right) : La, Hf, Ta, W, Re, Os, Ir, Pt, Au, Hg

**4th Transition series:** (from left to right) : Ac, Rf, Db, Sg, Bh, Hs, Mt, Ds, Rg, Uub

(d) **Bottom Portion:** This consists of two rows of elements each containing 14 elements. These are commonly called **inner transition elements**. The first row(<sub>58</sub>Ce-<sub>71</sub>Lu) elements are commonly called **Lanthanoids or Lanthanons** (previously called lanthanides). They are so called because they follow just after the element lanthanum(La) of group 3 (III-B).Lanthanum(La-57) and the 14 lanthanoid elements(Ce-Lu) belong to group 3(IIIB) and 6th period(3rd transition series) and similar properties. According to current IUPAC terminology lanthanoid series consists of 15 elements starting from La(57) and ending at Lu(71). So the element lanthanum(57) belongs to the lanthanoid series. These 15 elements along with Sc and Y in group IIIB are commonly called **rare earth metals** as they occur together in the minerals called 'rare earths'. This name is a misnomer as these elements have high relative abundance but their mining and extraction are difficult. The second row(90-103) of elements are called **Actinoids or actinons** (previously called actinides), as they follow the element actinium(89). According to current IUPAC terminology the

actinoid series also consists of 15 elements starting from Ac(89) and ending at Lw(103) which have similar properties.

*(N.B : Transition and inner transition elements are not included in the UTSE course)*

### DUAL POSITION OF HYDROGEN:

Usually the element hydrogen(H) is kept in group 1 (I-A) along with alkali metals. Hydrogen can also be placed in group 17 (VII-A) along with halogens. Hence it has two positions in the table. Hydrogen has some similarities with alkali metals, such as (i) it has one electron in the valence shell ( $ns^1$ ) and it shows oxidation number of +1 like alkali metals and forms  $H^+(H_3O^+)$  ion like  $Na^+$ ,  $K^+$ . (ii) Hydrogen gas can be obtained at the cathode during electrolysis of acidulated water like alkali metals from molten alkali metal halides etc. It has also some similarities with halogens such as (i) it is a diatomic gas like few halogens and it has one electron short of the noble gas configuration(He) like all halogens; (ii) it is a nonmetal like halogens and form  $H^-$ (hydride) ion( $ON = -1$ ) like  $F^-$ ,  $Cl^-$  and form  $H_2$  at anode when molten metal hydride is electrolysed.

### STUDY OF THE PERIODIC TABLE ON THE BASIS OF PERIODS:

There are seven periods.

**1st Period:** This consists of **two** elements H and He with atomic numbers 1 and 2 respectively.

**2nd period:** This consists of **eight** elements from atomic number 3 to 10. The elements from left to right are Li, Be, B, C, N, O, F and Ne. These are commonly called **bridge elements**\*\*\* because they have some similarity with their diagonal elements in the third period. For example Li has some similarity with its diagonal element Mg, Similarly Be has similarity with Al and so on.

**3rd Period:** This consists of **eight** elements from atomic number 11 to 18. The elements are Na, Mg, Al, Si, P, Cl and Ar. These are commonly called **typical elements**. \*\*\*\*\*

**4th Period:** This consists of **eighteen elements**(19-36), out of which two(K, Ca) belong to left portion, ten belong to middle portion(1st series of transition elements from Sc to Zn) and remaining six(Ga, Ge, As, Se, Br and Kr) belong to right portion.

**5th Period:** This also consists of **eighteen elements**(37-54), out of which two(Rb, Sr) belong to left portion, ten belong to middle portion(2nd series of transition elements from Y to Cd) and the remaining six(In, Sn, Sb, Te, I and Xe) belong to the right portion.

**6th period:** This consists of **thirty two** elements(55-86), out of which two(Cs, Ba) belong to left portion, fourteen elements( $_{58}Ce$ - $_{71}Lu$ ) belong to the bottom portion, ten elements( $_{57}La$ ,  $_{72}Hf$ - $_{80}Hg$ ) belong to the middle portion and six elements(Tl, Pb, Bi, Po, At, Rn) belong to the right portion.

**7th Period:** This consists of 32 elements, out of which two(Fr, Ra) are in left portion, fourteen elements( $_{90}Th$ - $_{103}Lr$ ) belong to the bottom portion, ten elements( $_{89}Ac$ ,  $_{104}Rf$ - $_{112}Cn$ ) in the middle portion and the remaining 6 elements in the right portion upto Ununoctium(118). Uut(ununtrium-113), Uuq(ununquadium-114), Uup(ununpentium-115), Uuh(ununhexium-116) and Uuo(ununoctium-118) names have been given on the basis of the latin names for the atomic numbers of these elements on temporary basis as controversies still exist in giving proper names to these elements. [In latin, nil(n)=0, un(u)=1, bi(b)=2, tri(t)=3, quad(q)=4, pent(p)=5, hex(h)=6, sept(s)=7, oct(o)=8, enn(e)=9. The suffix *ium* is added at the end.

### PHYSICAL STATE:

#### **Gaseous Elements:**

There are 11 elements which exist in gaseous state at room temperature. These are inert or noble gases(zero group or group 18) which exist as monoatomic gases in nature(He, Ne, Ar, Kr, Xe and Rn). Other gaseous elements(H, N, O, F, Cl) remain as diatomic molecules( $H_2$ ,  $N_2$ ,  $O_2$ ,  $F_2$  and  $Cl_2$ ).

#### **Liquid Element:**

There are two elements which exist in the liquid state at room temperature. These are mercury(Hg) which is a metal and bromine(Br) which is a nonmetal. Br exists as diatomic molecule( $Br_2$ ). Note that there are four more solid elements which have low melting points close to room temperature and some times considered as liquid elements. These are Rb, Cs, Fr(Group IA) and Ga(group IIIA).

#### **Solid Elements:**

All other elements which are not gases and liquids at room temperature exist in the solid state. Few solid elements like P and S remain in molecular form while the rest remain in atomic form in a network arrangement(details to be learnt in chemical bonding chapter). Phosphorous usually exists as tetra-atomic molecule( $P_4$ ) while sulphur remains as octa-atomic molecule( $S_8$ ) in the solid state.



### **Radioactive elements:**

All the elements above atomic number 83(Bi) i.e **from polonium(Po) onwards are spontaneously radioactive** and emit rays like alpha, beta and gamma(refer the chapter *radioactivity*). Note that there are two elements below atomic number 83 which are also radioactive. These are Technetium(Tc) among the transition elements(middle portion) and promethium(Pm) existing in the bottom portion(lanthanoid).

### **Natural and man-made elements:**

All the elements up to atomic number 92(Uranium) are natural and are available in different forms on earth. Beyond uranium i.e from atomic number 93(Neptunium) and onwards all the elements are man-made and are artificially prepared from natural elements. These are solids at room temperature. These are commonly called **transuranic elements**. Note that all these are radioactive elements. Note also that the elements Tc(43) and Pm(61) which are radioactive as mentioned before are also man-made. Seven elements which are present in trace quantities in nature are believed to be absent in nature for all practical purposes. These are Tc(43), Pm(61), At(85), Rn(86), Fr(87), and the two transuranic elements Np(93) and Pu(94). It is believed that these seven elements were available in nature long time ago. But now they are found in trace quantities on earth and are artificially prepared. **So strictly speaking there are 94 natural elements including Np and Pu.** But for all practical purposes there are now 87 natural elements( 94 - 7).

### **METALS, NONMETALS AND METALLOIDS:**

See the *staircase line* starting just below boron(5) in PT running down on the right side of Al, Ge and Sb. This is the dividing line between metals and nonmetals. Metals lie to the left side of this staircase while nonmetals lie to the right side. The elements which lie in the immediate vicinity of the staircase(except Al), such as B, Si, Ge, As, Sb, Te, Po and At are neither true metals nor true nonmetals and they have some metallic properties and some nonmetallic properties and are called **semi-metals or metalloids**. Here we discuss some of the important properties of metals, nonmetals and their differences.

#### **METALS Versus NONMETALS:**

1. **Melting Points:** Metals have high melting points whereas the nonmetals have low melting points. That is why all metals are solids at room temperature except mercury(Hg) which is a liquid. Nonmetals are usually gases, liquid(bromine) or low melting solids(P<sub>4</sub>, S<sub>8</sub>, I<sub>2</sub> etc). Excepting alkali metals(Li, Na, K, Rb, Cs, Fr) whose melting points are low, metals have fairly high melting points usually above 500°C. Nonmetals however have very low melting points. Take for example, O<sub>2</sub> gas which has melting point of -183°C and phosphorous(solid) has a melting of 44°C. The reason for these observations will be discussed later
2. **Electrical Conductivity:** Metals are good conductors of electricity in solid as well as in liquid states while nonmetals are poor conductors. This is because metals possess free electrons while the nonmetals do not. The only exception of a nonmetal being a good conductor of electricity is **graphite**(one allotrope of carbon). This is because of presence of free electrons in graphite.
4. **Thermal Conductivity:** Metals are good conductors of heat while nonmetals are not. If you heat one end of a metal(say iron), the other end is heated soon. But when you do the same thing to iodine solid, the other end is not heated soon. This is because iodine is a nonmetal and is a poor conductor of heat.
5. **Metallic lustre:** Metals when freshly cut show a shining surface, i.e it has its lustre. For example, when silver is polished and the impurities on the surface are removed, it shines. So also it is true for all metals. But nonmetals do not have a shining surface. This is again due to free electrons present in the metals which scatter light when falls on them.
6. **Tensile Strength:** Metals have high mechanical strength (tensile strength) while nonmetals have low strength. When you take a piece of iron metal and stretch it on two sides, can you pull it? No, because metals have high strength. But if you do the same thing with a nonmetal solid (say phosphorous), you can pull the solid apart and break it by applying force. This is because the metallic bond which bind the atoms in a metal crystal is strong while the weak vander waals forces which bind the nonmetal atoms/molecules are weak.
7. **Malleability and Ductility:** Metals are malleable and ductile while nonmetals are not. Rather the nonmetals are brittle. Malleable means the metals can be hammered or pounded to make thin sheets. Do

you remember the Rutherford's gold foil experiment? Gold was hammered to a thin sheet of 0.00004cm thickness. All metals are malleable to different extents. On the other hand if you hammer iodine solid, it will crumble into pieces because it is brittle. The strong metallic bond present in the metals is responsible for these properties.

8. **Nature of Ions Produced:** Metals produce positive ions easily while nonmetals produce negative ions easily. Na will give Na<sup>+</sup> ions(not Na<sup>-</sup> ions) while Cl gives Cl<sup>-</sup> ion(not Cl<sup>+</sup>) easily.

**MAGIC NUMBERS FOR FINDING ATOMIC NUMBERS OF ELEMENTS:**

**(DRAWING PERIODIC TABLE)**

The student is advised to practise drawing a simple periodic table in a piece of paper as suggested below. Refer the periodic table given for the purpose. First draw seven rows and then eighteen columns. The first two columns belong to left portion and the last six columns belong to right portion. First fill elements group-wise in the left and right portions. Fill group IA(1) with H, Li, Na, K, Rb, Cs, Fr from top to bottom. Remember the proper sequence of these alkali metals. Then fill group IIA(2) with alkaline earth metals Be, Mg, Ca, Sr, Ba and Ra. Note that Be is placed beside Li(not H) and other elements are placed from top to bottom. After filling the left portion, jump to the right portion. Start filling elements of boron family i.e group IIIA(13): B, Al, Ga, In, Tl and Uut. Remember the sequence of elements in each group. Then fill carbon family, i.e IVA(14): C, Si, Ge, Sn, Pb and Uuq followed by nitrogen family i.e groupV(A)(15): N, P, As, Sb, Bi and Uup; oxygen family, i.e group VIA(16): O, S, Se, Te, Po and Uuh; then halogens i.e group VIIA(17): F, Cl, Br, I, At, Uus and finally the noble gases i.e group zero(18): He, Ne, Ar, Kr, Xe, Rn and Uuo. Note that group zero elements start from the 1st period. Helium is placed at the top right corner in the first period horizontal with H and all other elements follow in sequence written above. [After filling group-wise to the left portion and right portion, let us look to 4th period. Fill row-wise all the ten elements of the 1st transition series in between calcium and gallium. The elements in sequence are Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn. Note that elements belonging to 1st transition series are very common and widely available. Then fill the transition elements groups-wise starting from group IIIB(3) to IIB(12). Sc-group: Sc, Y, La, Ac; Ti group: Ti, Zr, Hf, Rf; V group: V, Nb, Ta, Db; Cr group: Cr, Mo, W, Sg; Mn group: Mn, Tc, Re, Bh; Fe group: Fe, Ru, Os, Hs; Co group: Co, Rh, Ir, Mt; Ni group: Ni, Pd, Pt, Ds(Durmstadtium); Cu group(coinage metals): Cu, Ag, Au, Rg(Roentgenium); Zn group: Zn, Cd, Hg, Cn(Copernicium)].

*(N.B: The transition metals kept inside square bracket are not included in the UTSE course)*

Atomic numbers of the elements are obtained by following the magic numbers.

**MAGIC NUMBERS FOR THE LEFT PORTION: 8 + 8 +18 + 18 + 32.**

Before applying magic numbers, first give the atomic numbers of the elements of second period i.e from 3 to 10(from Li to Ne). Then start applying magic numbers to elements group wise. That means you add the first magic number(8) to the atomic number of the first element Li(3) to get the atomic number of the second element, sodium(3+8=11). Then add the second magic number(8) to the atomic number of second element, sodium(11) to get the atomic number of third element and so on(+8+8+18+18+32) successively starting from lithium to get the atomic numbers of all the elements. So the atomic numbers of Na, K, Rb, Cs and Rb will be respectively 11, 19, 37, 55 and 87. Note that H is not to be considered while applying magic numbers. You have to start applying the rules from the element of second period. For groupIIA(2) elements either you add 1 to each of group 1 atomic number to get the atomic number of the corresponding element of group IIA in the same period, or you can apply magic numbers to group IIA starting from Be (+8+8+18+18+32) to get the atomic numbers of all the elements. The atomic numbers of group 2 elements respectively are 4, 12, 20, 38, 56 and 88 for Be, Mg, Ca, Sr, Ba and Ra respectively.

**Magic Numbers for the right portion: 8+18+18+32 + 32**

Here one 8 is missing in the beginning which is compensated by 32 at the end. The atomic numbers of successive elements can be known while starting from the first element of 2nd period(B, C, N, O, F and Ne) and adding +8+18+18+32 +32 in sequence. Let us take group III A(13). The atomic number of Al=5 +8=13, followed by Ga= 13+18=31, then In = 31+18=49, Tl=49+32=81 and finally Uut = 81+32=113. The atomic numbers of all other elements of the groups of the right portion can be known by two ways. The easiest way is by successively adding one to the elements in a given period till the last element(noble gas) is reached. Look to the elements after Al in the same period(3rd). Si is 13+1=14, P=14+1=15, S=16, Cl=17 and finally Ar=18. Similarly you find the atomic numbers to the right of Ga(31) in the 4th period, In(49) in the 5th period

and Tl(81) in the 6th period. The other method is to apply magic numbers to the individual groups elements starting from the first element(of 2nd period) as indicated before. Thus we have got the atomic numbers of all the elements of the left and right portion. The atomic numbers of H and He are already known(1 and 2 respectively).

**1st transition series:** The atomic numbers of the 1st row transition series from Sc to Zn can be obtained by successively adding one to the atomic numbers of elements starting from Ca(20). Sc = 21, Ti = 22, V=23, Cr=24, Mn=25, Fe=26, Co=27, Ni=28, Cu=29, Zn=30. After Zn, the next element is Ga(31) belonging to the right portion.

**Periodic Law:**

*The physical and chemical properties of elements are periodic functions of their atomic numbers. In other words the properties vary in a periodic manner with their atomic numbers.* Similar elements recur after fixed intervals such as 8, 8, 18, 18, 32(left portion) or 8, 18, 18, 32, 32(right portion). The properties also vary in a gradual manner in period from a strong metal to strong nonmetal and finally to a noble gas.

**PERIODIC PROPERTIES**

If we move from one element to the other in a period, we shall find that the properties of the elements are gradually changing and there is often a regular trend in the properties. For example, if you move from left to right in a period the size of the atom gradually decreases. Say for instance in the second period, size of Li atom is the greatest and the atomic size gradually decreases as we move from Li to B-C -N -O to Ne. Similarly if we move from one element to the other in a group, we will find that the properties of elements change in a particular trend. For example, we find that the size of the atom gradually goes on increasing as we move down a group. Take for instance group VA, nitrogen atom has the smallest size in that group and size goes on increasing as we move down from nitrogen to phosphorous, then to arsenic, antimony and finally bismuth. Bismuth has the largest size in that group. We shall discuss seven such properties which show regular trends in the periods and groups. These are called *periodic properties* which change gradually and periodically. These are:

- |  |  |
|--|--|
| (i) Atomic size or the Atomic Radius:      | (ii) Ionisation Energy or Ionization Potential |
| (iii) Electron Affinity or Affinity Energy | (iv) Electronegativity                         |
| (v) Valency                                |  |

**ATOMIC SIZE:**

Size of an individual atom of an element is known from its radius. For nonmetals we consider the covalent radius and for metals we take metallic radius. Both are combinedly called atomic radius. There was another form for radius called Van der Waals radius which is no longer in use now. So we shall not discuss about it.

**Variation across a period:**

In any given period, the atomic size decreases gradually and atoms get thinner and thinner. For example if you look to 3rd period, you will find that the atomic radius of Na is 1.54 Å and as we move to Mg, the atomic radius becomes 1.38 Å, then to Al whose atomic radius is 1.25 Å and the radius continues to decrease till end of the period. This is because, in a period, we are adding successive electrons in the same shell(Na=3s<sup>1</sup>, Mg=3s<sup>2</sup>, Al=3s<sup>2</sup>3p<sup>1</sup>, Si=3s<sup>2</sup>3p<sup>2</sup> and so on). But since the number of protons also increases as we move from element to element, the nuclear pull (attraction) to the electrons increases and therefore the atomic size decreases.

**Variation along a group:**

In any given group the atomic radius gradually increases. In other words the atoms get bulkier and bulkier. Look to any group, say group IIA(alkaline earth metals). The size of the Be is 1.06 Å and that of Mg is 1.38 Å as we go down the group the size further increases. This is because in a group, the valence electrons are present in different higher shells. For example, the valence shell configuration of Be is 2s<sup>2</sup>, while that for Mg is 3s<sup>2</sup> and for Ca 4s<sup>2</sup> and so on. Since the valence electrons are present in higher and higher shells as we move down a group, the size therefore goes on increasing.

**Exceptions:** There are few exception to this general trend. For example in group IIIA, Ga has nearly same size as Al. In transition metal series, the atomic size does not decrease as we move from left to right as expected of a period. The reasons for these deviations will not be discussed is out of scope for UTSE. FWe shall take only the representative(normal) elements(left and right portion).

## IONISATION ENERGY:

When one electron is removed from a neutral atom, it is converted to a positive ion carrying +1 charge. Do you guess from which shell this electron shall be removed? The electron is removed from the valence shell (the last shell) as the electrons are loosely bound to the nucleus in it and can be easily removed. For example, in sodium atom, we remove the one electron present in its valence shell ( $3s^1$ ). Na has 11 electrons and 11 protons (atomic number) and if we remove one electron, the number of electrons becomes 10, but the number of protons is constant (11). So the resulting species acquire +1(11-10) charge. For removing an electron from the hold of the nucleus (nuclear attraction) i.e. to bring about ionization, we need to supply some energy. In other words, the atom has to absorb some energy so as to lose one electron. This amount of energy absorbed by an atom is called *ionization energy* and was formerly called *ionization potential*.

**Definition:** *The amount of energy needed to remove an electron from a gaseous atom is called ionization energy.* Note that the atom has to exist in the gaseous state before measuring the ionization energy. This ionization energy is strictly called **1st ionization energy ( $IE_1$ )** as this amount of energy is spent to remove the 1st electron from an atom. If a second electron is removed from the first formed positive ion i.e. to convert  $X^+$  to  $X^{2+}$ , we have to supply still more energy. This is called the **2nd ionization energy ( $IE_2$ )**. For the third electron removal, we need to supply still more energy called the **3rd ionization energy ( $IE_3$ )** and so on. We shall discuss here only on 1st ionization energy and not the higher orders of ionization energies ( $IE_2, IE_3, \dots$ ). If we talk of ionization energy it means the **1st ionization energy ( $IE_1$ )**.



More the ionization energy, more difficult for an element to form its positive ion. In general, metals have low ionization energy while nonmetals have high ionization energy. Therefore metals form their positive ions easily. For example, between Na and O, Na can form positive ion ( $Na^+$ ) easily by losing an electron as its IE is lower (496 Kjoule/mole) while O cannot form positive ion ( $O^+$ ) easily as the IE of O is very high (1314 Kjoules/mole). Refer table-5 for this.

### **Variation of Ionisation Energy across a Period:**

The ionization energy increases gradually in a given period. There are two deviations from the general expectation in this period. Li has an IE of 520 Kjoule/mole. The next element in the same period, Be has greater IE of 899 Kjoule/mole. But the next element **B has a little lesser IE than Be**. B has an IE of 801 Kjoules/mole. This is the first deviation. Then onwards the IE again increases, C has the IE of 1086, N has IE of 1410 Kjoules/mole. Again we meet another deviation. **O has less IE than N**. O has an IE of 1314 Kjoule/mole. Then onwards the IE increases in F (1681) and then in Ne (2080).

The increase of IE across a period is due to the fact that when nuclear pull increases (electrons are tightly bound to the nucleus) as we move from left to right in a period, it becomes increasingly more difficult to remove an electron. Hence IE increases. Be and N have unexpectedly higher IE than B and O respectively is due their stable filled filled ( $2s^2$ ) and half filled ( $2p^3$ ) configurations. More explanation on this point is beyond the scope of UTSE. Similar type of deviations are also noticed in the 3rd and 4th period.

2nd period:  $Li < B < Be < C < O < N < F < Ne$

3rd period:  $Na < Al < Mg < Si < S < P < Cl < Ar$

### **Variation of Ionisation Energy across a Group:**

Along a group, the ionization energy decreases gradually. As the valence electron goes farther away from the nucleus, it becomes more and more loosely bound to the nucleus and lesser amount of energy is required to remove it. Look to group IA (alkali metal). The IE of Li is 520 Kjoule/mole, that of Na 496 Kjoule/mole and gradually IE decreases as we move down.

### **Metallic and Nonmetallic Character:**

Metals have, in general, low ionization energy compared to nonmetals. The tendency to lose electron is more among metals. Take the case of Na and Cl. Na is a strong metal and has an IE of 496 Kjoules/mole while Cl is strong nonmetal and has a higher value of IE of 1255 Kjoules/mole. Therefore a nonmetal has a little tendency to lose electron. On the other hand it has a strong tendency to gain electron.

## ELECTRON AFFINITY(AFFINITY ENERGY)

When an atom accepts one electron from an outside source it is converted to a negative ion. You know that atoms form their respective positive ions by losing electron(s). Here we shall find that atoms can form their negative ions by accepting electron(s).



**During this process some amount of energy is evolved or released. This amount of energy released when a negative ion is formed from a neutral gaseous atom is called electron affinity.**

More the electron affinity of an element, more is the tendency or ease of forming negative ion.

As we go from left to right in a period the electron affinity increases. That means that the elements acquire greater tendency to form their negative ions. Note that for many elements Electron Affinity(EA) is zero. This means that such element has no tendency to gain electron and hence no tendency to form its negative ion. In 2nd period, the EA of B, C, O and F are 15, 121, 141 and 333 Kjoule/mole respectively. So the tendency to form negative ion(anion) is maximum for F and minimum for B. Similarly look to group VIA. EA of O and S are 141 and 200Kjoule/mole respectively. So the tendency of S to form negative ion is more than that of O. Note that the EA of all alkaline earth metals such as Be, Mg, Ca, Sr and Ba are zero. This is because these elements have full filled valence shell configuration (2) and therefore are stable. They do not have any tendency to gain extra electron. So their EA is zero. Inert gases also have zero EA since all of them have stable octet( $s^2p^6$ )configurations(He also has stable duplet,  $1s^2$  configuration). Nitrogen has also zero EA stable half filled configuration( $2p^3$ ).

Excepting a few elements which have zero EA, EA increases across a period and decreases along a group. Exception: (i) B has lower EA than Li and so also Al has lower EA than Na.

(ii) Al, Si, P, S and Cl have greater EA than B, C, N, O and F respectively in the right portion of PT. (The reasons for these exceptions are beyond the scope of UTSE)

**Nonmetals have higher electron affinity than metals.**

Therefore nonmetals like O, S, F, Cl etc. can form their negative ions more easily than the metals

## ELECTRONEGATIVITY

To know which element is more metallic and which is less metallic i.e more nonmetallic on relative basis, a scale has been introduced. This is called *electronegativity* (EN).

**It is defined as the relative tendency of an atom to attract a shared or bond pair of electron towards itself.** More the tendency of an atom to attract an electron pair, more is its electronegativity. Again more the electronegativity more is its nonmetallic character. For example oxygen atom has greater tendency to attract an electron pair, hence it has higher electronegativity while Na has a very less tendency to attract electron and so its electronegativity value is small. Note that all metals have less tendency to attract electrons and have low electronegativity values, while all nonmetals have greater tendency to attract electrons and have higher electronegativity values. You know that nonmetals are found in the upper right(north east) region of the periodic table and have small atomic sizes. This is because, the nuclear pull to the valence electrons is more. That is again the reason why they have stronger tendency to attract electron and have high electronegativity. We shall discuss more about this electronegativity in the next chapter(chemical bond). One thing is to be noted that EN is a relative scale used for comparison among elements and it does not give the absolute value of the force of attraction of an atom to electron or an energy term like electron affinity.

***Difference between electronegativity and Electron Affinity:***

Electron affinity is the tendency of an element to accept or gain an electron and form negative ion, but electronegativity is the tendency of an element to attract a shared pair electron in a covalent bond towards itself(not accept). Electron affinity is an energy term and corresponds to an absolute value while electronegativity is a relative value used for making comparison among elements and EN values are not absolute. Although elements which have high electron affinity have also high electronegativity, their relative orders are not same in many cases. Some elements have high electronegativity but have lower electron affinity. The details on this aspect is beyond the scope of UTSE.

**Fluorine (F)** has been assigned the **highest electronegativity** of 4. All other elements have lesser electronegativity than it. **Oxygen is second most electronegative** element and has an electronegativity of

**3.5. Nitrogen and chlorine come third** in the merit list. They have the electronegativity of **3** each. Others have EN less than 3.

***Variation of EN across Period:***

EN increases across a period as the nuclear pull increases. Look to the 2nd period, Li has an EN of 1.0 while Be 1.5, B 2.0 and it increases up to F which has the highest EN(4). As the EN increases across a period the metallic character decreases and hence nonmetallic character increases. Li, Be are metals while B is a metalloid (semi-metal) and C, N, O, F are nonmetals. Note that in 4th, 5th, 6th and 7th periods which contain the transition series (middle portion), the EN does not vary regularly. We shall not discuss about them now.

***Variation of EN along a Group:***

The EN is expected to decrease as we move down a group. This is because the atomic size increases and the nuclear pull decreases and thus the tendency of the atom to attract electron should also decrease. Thus it is expected that as we go down a group the metallic character should increase or in other words the nonmetallic character should decrease. Look to group IIA(alkaline earth metals). Be has an EN of 1.5, Mg of 1.2, Ca 1.0 and thereafter the EN remains constant in Sr and then decreases in Ba(0.9) and remains same in Ra. In a gross scale the EN decreases and no doubt, at some places it remains constant. Similarly you can look to group IVA(carbon family). The EN decreases from C(2.5) to Si(1.8) and thereafter the EN remains constant in the rest of the elements Ge, Sn, Pb(all 1.8). The reason for this deviation is not to be discussed here.

**Note that for metals the EN value is usually less than 2 and for nonmetals the EN value is more than 2.** When electronegativity difference between two metals is greater than 1.7, then bond between them will be ionic. If this difference is less than 1.7, then the bond will be covalent. If this difference is zero it is a non-polar covalent bond and if this difference is non-zero then it is polar covalent bond. More about you will get in 'Chemical Bonding' .

**VALENCY**

The maximum valency of elements is the total number of electrons present in the valence shell which would be used in making ionic or covalent bond. From carbon family onwards there is a minimum valency which is equal to (8- group number as per the old name) with a -ve sign. This is because by gaining that many number of electrons it can form negative ions.

Normal Groups:

Group 1(I): Max Val. = +1,                      Group 2(II) : Max Val. = +2,              Group 13(III) : Max Val.= +3

Group 14(IV) : Max. Val. = +4, Min. Val. = -4 [ -(8-4)]

Group 15(V): Max Val. = +5, Min Val. = -3

Group 16(VI) : Max. Val. = +6, Min. Val. = -2;                      Group 17(VII): Max Val. = +7, Min Val. = -1

Group 18(zero): Max. Val. = +8, Min. Valency = 0 (Xe can give  $\text{XeO}_4$  in which valency of Xe = 8)

## MOLE CONCEPT

Mole is a chemical unit( like a dozen, a gross etc. used as family units) which can be defined in two ways.

(i) **Number wise:** *one mole of anything contains Avogadro's Number ( $6.023 \times 10^{23}$ ) of species of that thing.*

(ii) **Mass wise:** *It is the mass of the substance which contains Avogadro's number of species.*

Mole can be determined for atoms, molecules as well as for ions.

(a) *For atoms* : e.g C, Na, H, Fe etc. (b)*For molecules* :e.g  $N_2$ ,  $H_2$ ,  $CO_2$ ,  $H_2SO_4$  etc.

(c) *For ions* : e.g  $SO_4^{2-}$ ,  $NO_3^-$  etc.

(a) **For atoms:**

*The atomic mass expressed in gm(gm atomic mass) is the mass of Avogadro's number of atoms. So one mole of atoms weigh one gm. atomic mass(e.g : one mole of Na atoms weighs 23 gms , one mole of H atoms weigh 1.008 gm and one mole of C weigh 12 gms. and so on)*

*N.B: Some authors write one mole of atoms as one **gm. atom** and according to them the term mole should not be used for atoms. However, we have used both gm. atom and mole for atoms.*

- Task:** 1. Find the mass of  $6.203 \times 10^{23}$  atoms of (a)Phosphorus (b)Calcium  
(c)Helium (d)Boron (Ans: (a) 31g (b) 40g (c) 4g (d) 11g)
2. Find the number of atoms present in (i) 19 gms of Fluorine (ii)39 gms of Potassium  
(iii)63.5 gms of Cu (Ans: each of Avogadro's number of atoms)

### NO. OF ATOMS AND MOLES(GM. ATOMS)PRESENT IN A GIVEN MASS OF AN ELEMENT

*Example: Calculate the number of atoms present in 2.3 gms of Na*

**Solution:**

23 gms(gm. atomic mass) of Na contain  $6.203 \times 10^{23}$  atoms of Na

2.3 gms of Na therefore contain  $\frac{6.023 \times 10^{23}}{23} \times 2.3 = 6.023 \times 10^{22}$  atoms of Na

**Practice :** 1. Calculate the number of atoms present in the following:

- (i) 0.12 gm of C (ii)40 gms of He (iii)4 gms of Oxygen (iv)0.001 gm of Sulphur (v)40mg of Calcium (vi)6.35 kgs of Copper

The atomic masses are as follows: C=12, He=4, O=16, S=32, Ca=40, Cu=63.5

2. Calculate the mass of the following

- (i) 1 million carbon atoms (iii)  $12 \times 10^{30}$  of K atoms  
Atomic Masses: C =12 K=39

### No. of moles(gm. atoms) of atoms of an element present in a given mass of element:

You already know that 1 mole of atoms (also called *1 gm. atom* of an element) contains Avogadro's number of atoms which weigh one gm atomic mass. For example, the atomic mass of Na is 23. So 23 gms of Na = 1 mole of Na atoms =1 gm atom of Na which contains  $6.023 \times 10^{23}$  atoms of Na. So if you are asked to find the number of moles of atoms or gm. atoms of Na present in 2.3 gm of Na, it is 1/10 of a mole =0.1 mole of atoms(gm. atom). Also if you are asked to find the number of atoms present in it, you can find out by the method described earlier. See this example.

*Example: Find the number of moles(gm. atoms) of sodium atoms present in 0.023 gm of Na*

**Solution:**

23 gms of Na = 1 mole of Na atoms(1gm atom of Na)

0.023gm of Na =  $\frac{1}{23} \times 0.023 = 0.001$  mole of Na atoms or 0.001 gm atom of Na

**Practice:** 1. Calculate the number of moles (gm. atoms) of atoms present in the following  
 (i) 1.27 gms of Iodine (ii) 400 mgs of Oxygen (iii) 5.04 gms of H (iv) 0.28 gm of Silicon  
 (Atomic Masses: I=127, O=16, H=1, Si=28)

**(b) For molecule and Ions:**

*Molecular mass expressed in gms (gm. molecular mass) is the mass of Avogadro's number of molecules. Similarly ionic mass expressed in gm. is the mass of Avogadro's number of ions. Hence one mole of molecules will weigh one gm. molecular mass and one mole of ions will weigh one gm. ionic mass. One mole of N<sub>2</sub> gas will weigh 14X2= 28gms, one mole of CO<sub>2</sub> will weigh 44 gms and one mole of SO<sub>4</sub><sup>2-</sup> will weigh 32+4X16=96 gms and so on.*

**Practice 1:** Calculate the mass of 6.023X10<sup>23</sup> molecules of the following:

- (i) CaCO<sub>3</sub> (ii) S<sub>8</sub> (iii) HCl (iv) H<sub>2</sub>SO<sub>4</sub>  
 (Atomic masses: Ca=40, C=12, S=32, Cl=35.5, S=32, O=16, H=1) (Ans. gram. mol. mass)

2: Calculate the mass of 6.023X10<sup>23</sup> number of ions in case of the following

- (i) NO<sub>3</sub><sup>-</sup> (ii) Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup> (Atomic masses: N=14, Cr=52) (Ans. gram ionic mass)

**NUMBER OF MOLECULES PRESENT IN GIVEN MASS OF A SUBSTANCE AND VICE VERSA**

If 36.5 gms (gm molecular mass) of HCl contains 6.023X10<sup>23</sup> molecules of HCl, then x gm of HCl will contain (6.023X10<sup>23</sup>/36.5)X x molecules. So if the mass of a substance is given, we can calculate the number of molecules. We can also get the mass of a substance which will contain a fixed number of molecules. See the following examples.

*Example:* Calculate the number of molecules present in 7 gms of N<sub>2</sub> gas.

**Solution:**

*The molecular mass of N<sub>2</sub> is 28.*

*28 gms of N<sub>2</sub> contains 6.023X10<sup>23</sup> molecules*

$$7 \text{ gms of } N_2 \text{ contains } \frac{6.023 \times 10^{23}}{28} \times 7 = 1.505 \times 10^{23} \text{ molecules.}$$

**Practice :** Find the mass of the following.

- (i) 2.0076 X 10<sup>20</sup> molecules of Na<sub>2</sub>CO<sub>3</sub> (ii) 3.0115 X 10<sup>40</sup> molecules of H<sub>2</sub>SO<sub>4</sub>

**NO. OF MOLES PRESENT IN A GIVEN MASS OF A SUBSTANCE**

Since the gm. molecular mass is the mass of 1 mole of a substance, we can find out the number of moles present in given mass of a substance. Look to the following example.

*Example:* Find the number of moles of sulphuric acid molecules present in 490 gms of H<sub>2</sub>SO<sub>4</sub>.

**Solution:**

*Molecular Mass of H<sub>2</sub>SO<sub>4</sub> = 98*

*98 gms of H<sub>2</sub>SO<sub>4</sub> = 1 mole*

*490 gms of H<sub>2</sub>SO<sub>4</sub> = 490/98 = 5 moles*

**Practice 1.:** Find the number of moles in case of the following:

- (i) 18.6 gms of Phosphorous (P<sub>4</sub>) (ii) 1.47kg of H<sub>2</sub>SO<sub>4</sub>

(iii) 3.55gm of Cl<sub>2</sub> gas

2: Calculate the number of moles and molecules present in

- (i) 0.106 gms Na<sub>2</sub>CO<sub>3</sub> (ii) 48gms of Oxygen (iii) 15.5 gms of P<sub>4</sub>

(iv) 0.49 gms of H<sub>2</sub>SO<sub>4</sub> (v) 2.8 kg of Nitrogen gas

3: Find the mass of the following

(i) 10<sup>23</sup> molecules of hydrogen (ii) 100 million molecules of K<sub>2</sub>SO<sub>4</sub>

(iii) 1.2 X 10<sup>69</sup> molecules of Nitrogen

4: Find the mass of the following.

(i) 0.001mole of Ca<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub> (ii) 2.5moles of H<sub>2</sub>SO<sub>4</sub>

(iii) 1/20 mole of MgCO<sub>3</sub>



## MASS OF ONE ATOM OF AN ELEMENT

Since the mass of Avogadro's number of atoms is gm. atomic mass. So from this the absolute mass of one atom of an element can be calculated. Note that this is the actual mass of an atom which is incredibly minute and small. When you are asked, what is the mass of one oxygen atom, you often say 16gm. But just think how the mass of one tiny oxygen atom could be 16gm?? With 16gm of sugar you can prepare a cup of tea and could one atom of oxygen can weigh an incredibly large 16gm?? No, what we say 16gm; is its gm. atomic mass, which is the mass of Avogadro's number of atoms. So from this the actual mass of one atom in gm can be calculated. See this example.

*Example:* Find the mass of one oxygen atom:

*Solution:*

$6.023 \times 10^{23}$  atoms of oxygen weigh 16 gms

1 atom of oxygen weighs  $16 / (6.023 \times 10^{23}) = 2.66 \times 10^{-23}$  gm.

Do you notice how small the mass of one atom is!!!!

**(Mass of an atom = gm. atomic mass ÷ Avogadro's number)**

**Practice 1:** Find out the mass of one atom of each of the following. The atomic masses are given within brackets.

(i)C(12)

(ii)Al(27)

(iii)H(1)

(iv)S(32)

(v)Cl(35.5)

(vi)Ag(108)

**2:** What is the mass of 1 amu in gm? How this is related with the mass of hydrogen atom?

## MASS OF ONE MOLECULE OF A SUBSTANCE

Since the gm. molecular mass is the mass of Avogadro's number of molecules in the same way as gm atomic mass is the mass of Avogadro's number of atoms, we can calculate the actual mass of one molecule of a substance. See this example.

*Example:* Calculate mass of one molecule of  $CO_2$ .

*Solution:*

Molecular mass of  $CO_2 = 12 + 16 \times 2 = 44$

$6.023 \times 10^{23}$  molecules of  $CO_2$  weigh 44 gms

So 1 molecule of  $CO_2$  weighs  $44 / (6.023 \times 10^{23}) = 7.305 \times 10^{-23}$  gm.

You also found that like the mass of one atom, the mass of one molecule is also very small.

**(Mass of a molecule = gm. molecular mass ÷ Avogadro's number)**

**Practice 1:** Find the mass of one molecule in gm for the following

(i) $NH_3$

(ii) $CaCO_3$

(iii) $(NH_4)_2SO_4$

(iv) $CH_3COOH$

(v)NaCl

## NUMBER OF MOLECULES AND MOLES OF A GAS PRESENT IN A GIVEN VOLUME OF GAS AT A PARTICULAR TEMPERATURE AND PRESSURE

Gram Molar Volume(GMV) at NTP = 22.4 litres

**one mole of any gas occupies 22.4 litres at NTP.**

e.g 32 gms of  $O_2$  gas, 2 gms of  $H_2$  gas, 44 gms of  $CO_2$  gas, 28 gms of  $N_2$  gas etc. each occupies 22.4 litres at NTP ( $0^\circ C$  and 760mm Hg Pressure). So the number of moles and molecules of a gas present in a given volume of gas can be easily calculated. Also we can do the opposite thing. We can calculate the volume of gas provided we know the number of molecules or moles or mass. The examples below will make these more clear.

**Example :** Calculate the number of molecules, moles and mass of oxygen gas present in 224cc of oxygen gas at NTP.

**Solution:**

22400 cc of  $O_2$  gas at NTP contain  $6.023 \times 10^{23}$  molecules of  $O_2$

224 cc of  $O_2$  gas at NTP contain  $6.023 \times 10^{21}$  molecules of  $O_2$

Again 22400 ml of  $O_2$  gas contains 1 mole

224 ml of  $O_2$  contains  $224 / 22400 = 0.01$  mole

1 mole of oxygen weigh 32 gms

$0.01$  mole of oxygen weigh  $32 \times 0.01 = 0.32$  gm

**N.B :** For gas present at conditions other than NTP, the combined gas equation ( $P_1V_1/T_1 = P_2V_2/T_2$ ) is to be used to convert the given volume to NTP or volume at NTP to volume at given conditions depending on the requirement. . If the volume data is given, the conversion is made first and if the volume is to be found out, then the conversion is done at the end.

**Example :** Calculate the volume in cc of nitrogen gas containing  $2.4 \times 10^{20}$  molecules at  $27^\circ\text{C}$  and 800mm pressure.

**Solution:**

Here volume is to be calculated at some other temperature and pressure. So the conversion is to be made at the end.

$6.023 \times 10^{23}$  molecules of  $\text{N}_2$  at NTP occupies 22400 cc

$2.4 \times 10^{20}$  molecules of  $\text{N}_2$  occupies 8.96cc at NTP

But we need to find the volume at the conditions given in the question. Let us apply gas law.

$$\frac{760 \times 8.96}{273} = \frac{800 \times V_2}{(273+23)} \Rightarrow V_2 = 9.22\text{cc}$$

**Example 3:** Calculate the number of molecules and mass in gm. of  $\text{CO}_2$  present in 200 ml of the gas at  $27^\circ\text{C}$  and 800mm pressure.

**Solution:** Since the volume data is given here, the gas equation is to be used first to get the volume at NTP from the volume at given conditions.

$$\frac{800 \times 200}{273+27} = \frac{760 \times V_2}{273} \Rightarrow V_2 = 191.6 \text{ ml (Volume at NTP)}$$

22400 ml of  $\text{CO}_2$  gas at NTP weighs 44 gms

191.6 ml of  $\text{CO}_2$  weighs  $(44/22400)191.6 = 0.376\text{gm} = \underline{376 \text{ mg}}$ .

Again 22400 ml of the gas at NTP contains  $6.023 \times 10^{23}$  molecules at

191.6 ml of the gas contains  $5.1 \times 10^{21}$  molecules .

**Practice 1:**

- Find the volume of 4gms of  $\text{CO}_2$  gas at NTP. How many molecules are present in it?
- Calculate the mass of  $\text{N}_2$  gas present in 560ml of it at NTP. Also find out the number of moles and molecules present in it.
- Find the number of molecules present in 680ml of  $\text{O}_2$  gas at  $27^\circ\text{C}$  and 900mm of Hg pressure. Also calculate the mass in gm and number of moles.
- Calculate the volume of  $\text{H}_2$  at  $27^\circ\text{C}$  and 900 mm pressure if it contains 0.5 mole of hydrogen gas.

### PRACTICE QUESTIONS

- Find the number of gm. atoms(mole of atoms) present in the following. Also calculate the number of atoms present in it. The atomic masses are given inside brackets.
  - 4Kg of Ca (40)
  - 32.7gm of Zn(65.4)
  - 7.09gm of Cl(35.45)
  - 95.4gm of Cu(63.55)
  - 8.62gm of Fe(55.85)
- How many  $\text{H}_2$  molecules are present in 8.5 gm of  $\text{H}_2$ ? How many H atoms are in it?
- Find the mass of one S atom in gm.
- Find the mass of the following:
  - $12.046 \times 10^{24}$  atoms of H(1)
  - $3.0115 \times 10^{30}$  atoms of Zn(65.5)
  - $3.6138 \times 10^{22}$  atoms of Ag(108)
- How many moles and how many molecules of  $\text{H}_2\text{O}$  are present in 48gms of  $\text{H}_2\text{O}$ ?
- What is the mass of 0.004mole of glucose( $\text{C}_6\text{H}_{12}\text{O}_6$ )? What is the mass of one glucose molecule?
- What is the mass of the following. Also find their volumes at NTP.( N stands for Avogadro's number)
  - 3N molecules of  $\text{CO}_2$
  - 0.5N molecules of  $\text{N}_2$
  - 0.0002 N molecules of  $\text{O}_2$
- How many molecules of water and atoms of H and O are present in 9 gms of  $\text{H}_2\text{O}$ ?
- Calculate the number of moles in each of the following. Also find the number of molecules in each case.
  - 10gms of  $\text{CaCO}_3$
  - 3.5kg of  $\text{H}_2\text{SO}_4$
  - 3.16gm of  $\text{KMnO}_4$
- Find the number of  $\text{NO}_3^-$  ions present in 62gm of  $\text{NO}_3^-$ . What is the mass of one  $\text{NO}_3^-$  ion?
- Calculate the number of H atoms, S atoms and O atoms present in 7gms of  $\text{H}_2\text{SO}_4$ .
- Calculate the gm. atoms(mole of atoms) of carbon and oxygen present in 22gms of  $\text{CO}_2$  gas at NTP. Also find the number of carbon and oxygen atoms present in it.
- Calculate the number of oxygen atoms present in 25gm of  $\text{CaCO}_3$ . Also find the number of gm. atoms(mole of atoms) of oxygen present.
- How many H atoms are present in 25gm of  $\text{NH}_4\text{Cl}$ ? Also find the the number of moles of  $\text{NH}_4\text{Cl}$  present.

## CHEMICAL BOND

Why do atoms combine to form a molecule? Why does H atom combine with Cl atom to form HCl molecule? Why does Na atom combine with Cl atom to form NaCl? What is the nature of their linkages? All these questions will be answered in this chapter. When two atoms are tightly bound with each other and establish a contact, we say that a chemical bond has been formed between them. Before this close contact, they were free and independent but after the bonding between them, they are no longer independent and free. They are held together by strong forces of attraction and it is not easy now to separate them from each other. This is called a chemical bond. After the formation of chemical bond the atoms lose their identity and properties as a different species called a molecule is formed.

### Driving Force for Bond Formation:

Two atoms chemically combine by forming a bond between them in order to lower the energy of the system. The driving force for the formation of chemical bond is to acquire stability by way of lowering energy or releasing energy. One of the ways of achieving stability is by acquiring the nearest noble gas configuration. Every element has a tendency to acquire the stable electronic configuration of a noble gas ( $s^2p^6$  called the *octet* or  $s^2$  called the *duplet* of He). This is because the electronic configuration of the noble gas atom is very stable. All other atoms want to make themselves stable by attaining the nearest noble gas configuration. They do so in two different ways.

(i) By way of give and take of electrons between each other: This leads to the formation of *ionic bond*.

(ii) By way of mutual sharing of electrons: This leads to the formation of *covalent bond*.

Note that chemical bonds can also be formed even without the atom attaining the noble gas configuration. The sufficient condition for bond formation is the lowering of energy.

There are three types of bonds between atoms : (a) Ionic bond      (b) Covalent bond  
(c) Metallic bond

### IONIC OR ELECTROVALENT BOND:

The bond between a strong metal and a strong nonmetal is ionic or electrovalent in nature. In this case a give-and-take of electron takes place between them. The metal atom gives the electron while the nonmetal atom takes. We know from *periodic classification* chapter that metals have low ionisation energy and have strong tendency to lose electrons while nonmetals have high electronegativity and have strong tendency to gain electrons. Thus when a metal atom sees a nonmetal atom, their mutual desire get chance to be satisfied. The metal atom gives its loosely bound valence electron and the nonmetal atom takes that electron to its valence shell and are thus converted to positive and negative ions respectively. The oppositely charged ions are now held together tightly by electrostatic forces of attraction. You know from law of electrostatics that unlike charges (+ve and -ve) attract each other while like charges (+ve and +ve or -ve and -ve) repel each other. Due the strong attraction between the +ve and -ve ions formed during bond formation, in this case, they are linked or bonded with each other very firmly. This is called the *ionic or electrovalent bond*. **Note that by losing or gaining electron, each atom (now ion) acquires the stable inert gas configurations.**

Group IA(alkali metals) and IIA(alkaline earth metals) possessing comparatively low ionisation energy (IE) values are strong metals while group VIA(oxygen family) and VIIA(halogens) have high electron affinity(EA) and high electronegativity values and are strong non metals. Thus whenever any of the elements from strong metals(group IA or IIA) unites with any of the elements of strong non-metals(group VIA or VIIA) the bond between them becomes ionic.

### LEWIS STRUCTURES OF IONIC BOND:

#### *Ionic bond formation in NaCl*

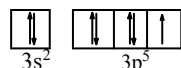
*Electronic Configuration of Na(11):*  $1s^2, 2s^2, 2p^6, 3s^1$

There is one electron (unpaired) in the valence shell of Na. This unpaired valence electron which is loosely bound will be donated by sodium completely to Cl. By that Na will be converted to  $Na^+$ . This is because Na atom has 11 protons(+11) and after losing one electron it will have 10 electrons(-10). So

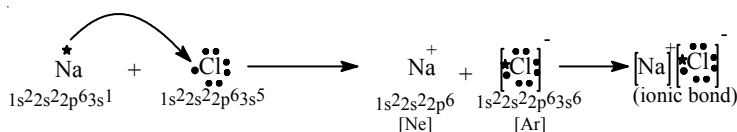
the net charge(+11-10) is +1. Let us see now what happens to Cl.

*Electronic Configuration of Cl(17):*  $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$ .

In its valence shell(3rd), there are three lone pairs and one unpaired electrons. This is known after making the box diagram of its 3s and 3p subshells as explained in the chapter, *atomic structure*.



For writing the Lewis structures for the ionic bond, you have to first find the number of lone pairs and unpaired electrons of each atom participating in the bond formation. The lone pairs are shown by **two closely placed dots** or symbols of a particular type while for unpaired electron a **single dot**(or any other symbol) has to be placed separately, a little away from the lone pairs. Note that for better clarity, the symbols given for electrons should be different for different atoms. In the example below, the electron of Na is given the *star(\*) sign* while that of Cl given *dot sign*. Then electron is transferred from the metal atom to nonmetal atom as shown below.



The unpaired electron of Na atom is transferred to Cl atom, thereby each attains the nearest stable inert gas configurations. Na after losing one electron becomes  $\text{Na}^+$  ion(Ne configuration) while Cl after gaining one electron becomes  $\text{Cl}^-$  ion(argon configuration). The valence shell of Na(3s) has been removed now in  $\text{Na}^+$  while there are four lone pairs in  $\text{Cl}^-$  ion. These two oppositely charged ions attract each other strongly and gives rise to ionic bond.

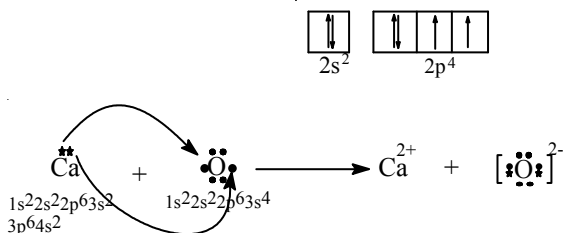
#### ***Ionic bond formation in CaO***

Ca(20):  $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$  .

Hence Ca has one lone pair in the 4s orbital.

O(8):  $1s^2, 2s^2, 2p^4$ . O has two lone pairs and two unpaired electrons in its valence shell(2nd shell).

The box diagram of the valence shell is shown below.



Calcium loses two electrons and these two electrons are gained by oxygen atom. The two unpaired electrons of O now get paired. O atom is thus converted to  $\text{O}^{2-}$  ion by accepting the two additional electrons. Thus the two ion acquire stable inert gas configuration.  $\text{Ca}^{2+}$  has Ar(18) configuration while  $\text{O}^{2-}$  has Ne(10)configuration. The two oppositely charged ions are bound together tightly by strong attractive forces. This is

## **IONIC COMPOUNDS AND THEIR PROPERTIES**

The compound in which the constituent units(ions) are held together by ionic bond is called ionic compound. These compounds are very typical and are different from other compounds. Let us know some of their important characteristics.

### 1. **State:**

All ionic compounds exist in solid states at room temperature. This is because the ionic bond is very strong and the ions stay very close to each other and so ionic compound remains as solid.

### 2. **Melting and Boiling Points:**

Ionic compounds melt and boil at very high temperatures. For example, NaCl which is an ionic solid melts at about  $800^\circ\text{C}$ . Since ionic bond is very strong, a great amount of energy is required to loosen the bonds between the oppositely charged ions, so that enough space is created for the free movement of ions and change into liquid state. Consequently the melting points of ionic solids are very high. For boiling the liquid (molten)ionic compound to gaseous state we need still more energy. For this the ionic bonds have to be completely broken which therefore needs much larger amount of energy. Thus the boiling point is enormously high.

### 3. Solubility:

Ionic compounds dissolve in water and have high degree of solubility in water. In fact, whether an inorganic compound is ionic or not, can be best known when you put a small amount of the compound in water and shake. If it dissolves it is ionic, if not it is covalent. Take for instance, NaCl dissolves in water and hence it is ionic, while AgCl does not dissolve in water, it is not ionic. It is a covalent compound.

To understand why water dissolves ionic compounds we have two explanations. One explanation we take up now and the other, we shall take up a little later. Water is a strongly polar compound (just wait for some time to know what is called **polar compound**) and has high dielectric constant. Dielectric constant is the property of the medium which controls the force of attraction between two ions present inside the medium. The force of attraction between the oppositely charged ions is inversely proportional to the dielectric constant. This, you shall know more clearly in Physics in a law called **Coulomb's law of forces** in higher classes. But this much you must know, at this stage, that more the dielectric constant, less is the force of attraction between the oppositely charged ions. Since water has very high dielectric constant, the force of attraction between positive ion (say  $\text{Na}^+$ ) and negative ion (say  $\text{Cl}^-$ ) of the ionic compound put in water is largely reduced. As a result ions get separated from each other and start moving freely and independent of each other in the solvent (water). The ions mix uniformly throughout the solvent and forms a homogenous solution. The other explanation shall be given after a few pages in this chapter. The ionic compounds are not soluble in non-polar organic solvents such as carbon tetrachloride, benzene etc. What is a non-polar solvent and why the ionic compounds are not soluble in them, we also shall know all these a little later.

### 3. Electrical Conductivity:

*Ionic compounds conduct electricity in the liquid (molten or fused) state or in aqueous solution. They do not conduct electricity in the solid state.*

Let us know what these statements mean. If we make an electrical circuit in which a solid NaCl crystal (common salt crystal) is placed at some place in the path of current in the circuit, will the current pass? Will an electric bulb placed in the circuit glow? The answer is NO. Solid NaCl is a poor conductor of electricity and electric current cannot pass through solid NaCl crystal. However, in stead of solid NaCl crystal if we take NaCl solution in water or liquid NaCl (obtained by melting solid NaCl at about  $800^\circ\text{C}$ ), then the current starts flowing through the circuit and the bulb glows. Thus we found that ionic compounds conduct electricity only in the liquid state and in aqueous solution.

#### **Explanation:**

First of all let us know how electricity is conducted in a conductor. There are two ways for the conduction of electricity. For **metallic conductors** such as Cu, Fe, Al etc. the electric current is carried by **free electrons**. But for aqueous solution of an ionic compound which contains free ions(both +ve and -ve), the current is carried inside the solution by these free ions. This is called an **electrolytic conductor**. For example, in NaCl solution, the current inside the solution is carried by the ions. Positive ions move towards the cathode(-ve electrode) and negative ions move towards the anode(+ve electrode) and thus electricity is conducted through the solution. Same thing will happen if we take a liquid NaCl in stead of NaCl solution. In electrolytic conductor, the current is carried by the **free ions**.

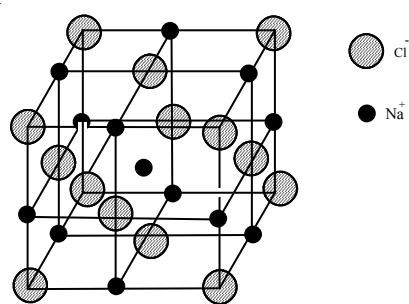
One thing we must remember that in ionic compound, the positive and negative ions exist even in the solid state. It is wrong to write the formula of sodium chloride as NaCl. The correct formula of sodium chloride should be  $\text{Na}^+\text{Cl}^-$ , because in all the states namely solid, liquid and in solution, sodium chloride contains  $\text{Na}^+$  and  $\text{Cl}^-$  ions. We should never think that sodium chloride undergoes ionisation when dissolved in water. Ions are present even before dissolution i.e in the solid state. The ions in the solid state are however very tightly packed forming a crystal and so they are unable to move from their positions. They are rigid at their positions and cannot move from one place to another and thus cannot carry current. However when the ionic solid is put in water, the ions get separated (for the reason already discussed) and freely moving ions are formed which carry the current. In liquid state also, the ions are free (although not as free as in solution) and conduct electricity. In solid state the ions are not free as they are tightly packed and do not conduct electricity.

### 4. Crystal Structure:

Ionic compounds are crystalline in nature. Have you seen a big crystal of common salt(NaCl) or any other crystalline salt such as copper sulphate crystal? Crystals have fixed geometrical shape. Amorphous substances are powder like (like talc-cum powder) in nature. If one sees to the internal structure of an

ionic crystal one finds that hundreds and thousands of positive and negative ions are arranged in a regular, orderly and systematic manner in the same shape as the big crystal that we observe. If you break down a crystal to smaller and smaller units you will arrive at the smallest unit which is called the *unit cell* of the crystal. This is the *motif* or the simplest pattern that is repeated again and again in all possible directions to get the big crystal. This infinite arrangement of ions in the crystal in the systematic way is called a *crystal lattice*. During the formation of this solid crystal lattice from the gaseous ions due to strong forces of attraction, a large amount of energy is released which is called *lattice energy*. A simple unit cell of a NaCl lattice is shown below. The big shaded balls stand for the  $\text{Cl}^-$  ions while the small black balls stand for the  $\text{Na}^+$  ions. We find that  $\text{Cl}^-$  ions are sitting at the corners of an imaginary cube and also at the centres of each face. Since there are six faces of a cube, we find in the unit cell, six  $\text{Cl}^-$  ions are placed at the centres of each face. In between two  $\text{Cl}^-$  ions, we find one  $\text{Na}^+$  ion. In other words  $\text{Na}^+$  ions are placed at the centre of each edge. There are 12 edges and hence we find 12 black balls( $\text{Na}^+$ ) at each edge centre. Besides these, there is one  $\text{Na}^+$  ion at the centre of the cube(called the body centre) in between the  $\text{Cl}^-$  ions present at any two opposite face centres. You imagine that this unit cell(cube) that has been shown in the diagram below is repeated in all sides such that between any two successive  $\text{Cl}^-$  ions you will find one  $\text{Na}^+$  ion and in between any two successive  $\text{Na}^+$  ions you get one  $\text{Cl}^-$  ion, we get an infinite network of ions arranged in a very orderly manner which ultimately results in the big crystal. Hence the shape of the unit cell is same as the shape of the big crystal. Each  $\text{Cl}^-$  ion (shaded ball) and each  $\text{Na}^+$  ion(black ball) shares more than one unit cell and this infinite network of ions is called the crystal lattice. To put in more simple way,  $\text{Na}^+$  and  $\text{Cl}^-$  ions are arranged in an alternate manner in the cubical pattern as shown in the diagram below. Although in the diagram we find a lot of empty space in between the ions, in reality the ions are very closely packed and there is very little space inside them. Due to close packing of ions the ionic compounds are very rigid and have high melting points.

So ionic compound forms crystalline structure which is one of its special properties. Individually an ionic bond which represents the attractive forces between the positive and negative ion has no specific direction and we say the ionic bond is *non-directional*. But collectively the ions give a fixed geometry (shape) to the ionic crystal i.e the ionic solid. This is due to repeated arrangement of positive and negative ions in a regular and orderly manner following a particular geometry while repeating. There are several types of crystal structures for different ionic compounds such as cubic, hexagonal, rhombic, monoclinic, etc. But we shall not discuss about them at this stage. Further details on crystal structures will not be included in this book.



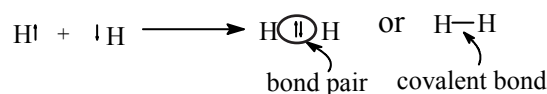
(One unit cell of NaCl crystal lattice)

## COVALENT BOND

We learnt that the bond between a strong metal(like Na, K, Ca etc) and a strong non-metal(like F, Cl, O etc) is ionic in nature. In such compounds, the positive and negative ions are arranged alternately and are packed closely to form the ionic solid. Now the question arises, what kind of bond exists between atoms which are both nonmetals say between H and H in  $\text{H}_2$  molecule, between Cl and Cl in  $\text{Cl}_2$  molecule, O and H in  $\text{H}_2\text{O}$  molecule, N and H in  $\text{NH}_3$  molecule and so on. When both the atoms are nonmetals, no atom has the tendency to lose electron. Rather both have the tendency to gain electrons. In that situation, the bond takes place by mutual sharing. Elements appearing in the right portion of the periodic table which are mostly nonmetals and have high electronegativity form covalent bonds between themselves. Each atom contributes one electron to form a bond pair. **In a bond pair, if one electron spins clockwise, the other spins anticlockwise.** The bond pair sits in between the two atoms. In ionic

bond, this pair is present completely in the nonmetal atom making it a negative ion. But in covalent bond the electron pair called the bond pair is situated in between the two atoms and is shared by both the atoms. The two electrons are counted for the electronic configuration of both the atoms. Thus a bond is established between the two atoms on the basis of sharing. This is called a covalent bond.

### H<sub>2</sub> molecule:



Each H atom contributes one electron (the only electron it possesses) and the two electrons get paired and form the bond pair. This bond pair sits tightly in between the two H atoms and binds the two H atoms to form one H<sub>2</sub> molecule. This is a covalent bond. We represent a covalent bond by a dash (—) symbol in between two atoms. This bond pair is counted for the electronic configuration of both the atoms. Thus after forming covalent bond the configuration of each H atom now becomes 1s<sup>2</sup> i.e. stable He configuration (duplet). Thus we find that by way of covalent bond formation also, the atoms attain the stable configuration of the inert gas.

### Lewis Electron dot and Dash Structures:

We can also write the formation of the covalent bond by showing the electrons as dots in stead of the arrow as shown above. This is called **Lewis electron-dot structure**. When a bond pair is replaced by a dash, it is called **dash structure**. The dot and dash structures in H<sub>2</sub> are as follows.



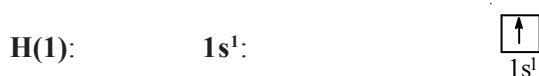
One thing we must bear in mind that for the formation of a covalent bond, each atom should possess at least one unpaired electron. Each atom contributes this unpaired electron to form the bond pair which is shared by both. If an atom has no unpaired electron, there are two possibilities.

(i) It cannot form a bond.

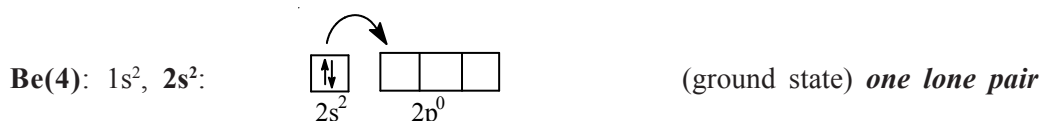
(ii) It breaks down a lone pair and gets two unpaired electrons and then forms two covalent bonds.

So to conclude again we emphasize here that the atom should have unpaired electron for establishing covalent bond with other atoms (either with identical atom or different atom). **The number of unpaired electrons is called the valency or more correctly the covalency of the element.**

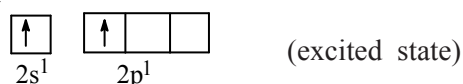
Before we take up more examples on the Lewis structures of covalent molecules, first let us know how to find the number of lone pairs and unpaired electrons for common atoms. Do you remember what are lone pairs? *The pairs of electrons present in the valence shell (last shell) of an atom are called the lone pairs.* So we have to write the electronic configuration and then **box diagram** of the valence shell and arrange electrons in them following the Hund's rule.



Hydrogen has one unpaired electron.



Be has one lone pair and no unpaired electron in its ground state. **Ground state** means the most stable state in which the atom usually remains before bonding. Since Be has no unpaired electron, we may expect that it would not form any bond with any atom. But we know that Be has a valency of 2 and therefore it has to use two unpaired electrons. How can we account for this? So the lone pair has to be broken. One of the electrons of 2s orbital jumps to one vacant p orbital so as to form two unpaired electrons. This is called the **excited state** of Be atom. This is the higher energy state which the atom attains before bond formation.



So Be has now two unpaired electrons for bonding.

**B(5):**  $1s^2, 2s^2, 2p^1:$



B has one lone pair and one unpaired electron. But B has a valency of 3. How to account for it? One electron has to jump from the 2s orbital to a vacant p orbital, thus producing three unpaired electrons.



So B has now three unpaired electrons for bonding.

**C(6):**  $1s^2, 2s^2, 2p^2:$

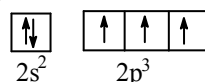


C has one lone pair and two unpaired electrons in the ground state. But the valency of carbon is 4. So one electron from 2s orbital jumps to the vacant 2p orbital and thus produces four unpaired electrons.



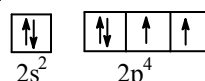
Thus C has four unpaired electrons for bonding.

**N(7):**  $1s^2, 2s^2, 2p^3$



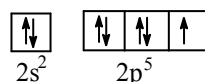
N has one lone pair and three unpaired electrons. Since valency of N is three, so the three unpaired electrons solve the purpose. There is no necessity of breaking the lone pair as we were doing before. In fact it is not possible in case of N. Why it is so, we shall know it later.

**O(8):**  $1s^2, 2s^2, 2p^4$



Thus O has two lone pairs and two unpaired electrons. The two unpaired electrons satisfy the valency of oxygen(2). Moreover breaking of lone pairs to get more unpaired electrons is not possible like N.

**F(9):**  $1s^2, 2s^2, 2p^5$



F has thus three lone pairs and one unpaired electrons. Valency of F(1) is satisfied by one unpaired electron. Breaking down of lone pairs is not possible like N and O.

So let us make a summary table showing the number of lone pairs and unpaired electrons in the common elements **which are used for bonding**. The element belonging to the same group have same number valence electrons. Therefore they also have same number of lone pairs and unpaired electrons. For example, Al belongs to boron family and it has one lone pair and one unpaired electron in the ground state and three unpaired electrons in the excited state like B for bonding. Similarly Si has same configuration as C(group IVA), P same as N(group VA), S same as O(group VIA), Cl same as F(group VIIA)

**TABLE TO BE USED ONLY FOR BONDING**

Name of the Element	Number of lone pairs	Number of unpaired electrons
Be	0	2
B, Al	0	3



C, Si	0	4
N, P	1	3
O, S	2	2
F, Cl, Br, I	3	1

Now let us return to Lewis structures. For that let us know the rules.

### RULES:(for simple diatomic molecules)

1. First write the symbols of the atoms which will form the covalent bond side by side.
2. Place the lone pairs and unpaired electrons around the atom. The electrons are shown by dots or any other similar signs. The dots representing the lone pairs are closely placed whereas the dots for the unpaired electrons are kept separately and a little away from the lone pairs. Look to the example for forming Cl<sub>2</sub> molecule from two Cl atoms.



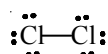
You can clearly see the difference between lone pairs and unpaired electrons.

3. Bring the atoms close to each other so that the two unpaired electrons of the two atoms come close and form the bond pair. Give a oval shaped mark over the bond pair to show its difference from the lone pairs. Thus we get the Lewis dot structure of Cl<sub>2</sub>.

**Lewis Electron dot structure:**



**Lewis Dash Structure:**

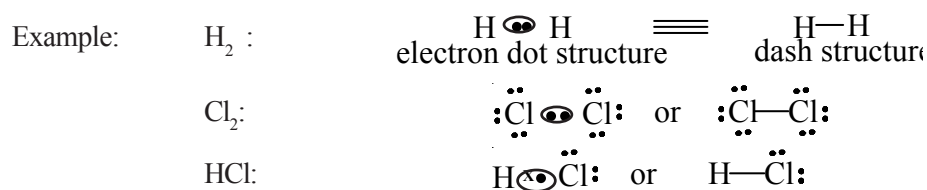


Note that in both *dot and dash structures*, the lone pairs on each atom have to be shown as before. The readers here are made cautious not to forget placing the lone pairs in the Lewis dot or dash structures. The bond pair is counted for the electronic configuration of both the atoms. Each Cl atom now acquires the stable electronic configuration of the inert gas(Ar=18).

### TYPES OF COVALENT BOND ACCORDING TO THE NUMBER OF BOND PAIRS SHARED:

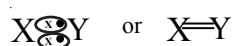
There are three types of covalent bonds according to number of bond pairs shared between two atoms.

- (i) **Single Bond:** If each atom contributes one unpaired electron to produce only one bond pair, a single bond results. (Oval shape marks are not given anywhere in this material, readers are advised to give such marks for the bond pairs in the hard copies, if they take)



In H<sub>2</sub> and Cl<sub>2</sub>, only one bond pair is formed between the atoms, hence these are single bonds. In HCl, each of H and Cl contributes one electron and one bond pair is formed between the atoms and hence a single bond results. The electron of H has been shown by a different symbol(cross in stead of dot)merely to show that bond pair contains one electron from H and one electron from Cl. Although you can give a dot symbol for the electron of each atom, it looks more convincing if you make different symbols for the electrons of atoms of different elements. For same element(Cl<sub>2</sub>, H<sub>2</sub>) however you may show the electrons by the same symbol(dot), but for dissimilar atoms, it is better to show the electrons by different symbols. But remember that this is not a binding rule.

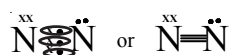
- (ii) **Double Bond:** If two electron pairs are shared between two atoms by the contribution of two unpaired electron from each atom then we get a double bond.



Each atom contributes two unpaired electrons. One unpaired electron of one atom(say X) gets paired with the one unpaired electron of the other atom(say Y) while the other unpaired electron of first atom(X) gets paired with the other unpaired electron the second atom(Y) to produce two bond pairs. These two bond pairs are jointly shared by the two atoms. This is called a double bond. We shall see several

examples of double bonds later.

(iii) **Triple Bond:** If three electron pairs are shared between two atoms by the contribution of three unpaired electrons from each atom, then a triple bond is formed. Take the case of  $N_2$  molecule. Nitrogen atom has three unpaired electrons and one lone pair (see the table given before). So three unpaired electrons of one N atom pair with three unpaired electrons of the other N atom to form three bond pairs. Thus a triple bond is formed.



Note that the electrons of the two N atoms have been shown in different symbols to clearly indicate that each N atom contributes three unpaired electrons. You can also show the electrons by the same symbol (dot), when you practise, as there is no strict rule for giving the symbols for the electrons of atoms. Also see that each N atom carries a lone pair in the Lewis structure.

## LEWIS STRUCTURES OF SOME MORE COVALENT COMPOUNDS:

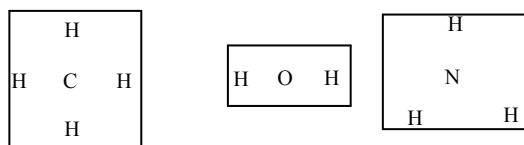
### RULES:

When the molecule contains more than two atoms, say for example  $CH_4$ ,  $H_2O$ ,  $NH_3$  etc. then you have to first find out the **central atom** among them. For simple molecules like the above you can make a correct guess. In  $CH_4$ , C is the central atom to which all the four H atoms are attached. Similarly in  $H_2O$ , O and in  $NH_3$ , N are the central atoms. But let us learn some rules how to know the central atom.

(i) H cannot be the central atom. The least electronegative (least nonmetallic) among the rest is to be selected as the central atom.

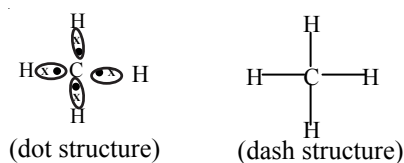
In case of  $CH_4$ ,  $H_2O$  and  $NH_3$ , there is no other choice than to make C, O and N as central atoms respectively as H atom cannot be the central atom. But for molecule like  $HNO_3$ , which is the central atom? H cannot be the central atom, then between N and O which is less electronegative? Please refer the electronegativity table in the *periodic classification* chapter for comparing the electronegativity values. Since N is less electronegative, it is taken as the central atom.

(ii) Write the central atom first and place all other atoms in a symmetrical manner surrounding the central atom. See the case of  $CH_4$ ,  $H_2O$  and  $NH_3$ .

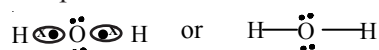


(iii) Next, the unpaired electrons and lone pairs for each atom are placed in a very symmetrical manner. Care should be taken to place the unpaired electron of one atom opposite to the unpaired electron of the other atom with which it will form covalent bond. Refer table given before to know the number of unpaired electrons and lone pairs available in each of the common atoms for bonding. Then highlight the bond pairs formed by the pairing of unpaired electrons by giving a oval shaped mark for each bond. The lone pairs are kept as such without being used up. Thus the Lewis electron dot structure is complete. You can also draw the *Lewis dash structures* by replacing the dots of the bond pairs by dashes. Single dash (single bond) is used for a single bond pair, double dash (double bond) for the two bond pairs and triple dash (triple bond) is used for three bond pairs between the atoms. For sometime you practise both dot and dash structures. After becoming familiar with them, you may stop writing the dot structures and write only the dash structures which look more clean.

**$CH_4$ :** C has four unpaired electrons and no lone pair and H has one unpaired electron.

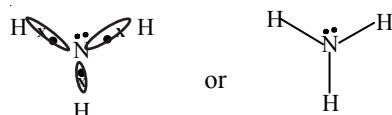


**$H_2O$ :** O has two lone pairs and two unpaired electrons. It is the unpaired electrons which form the covalent bonds. Two unpaired electrons form two single bonds with two H atoms as shown below.

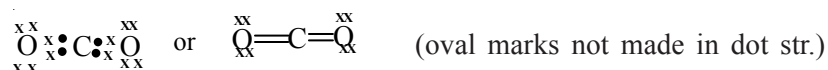


You find that the two lone pairs of O atom have been shown in both the dot and dash structures.

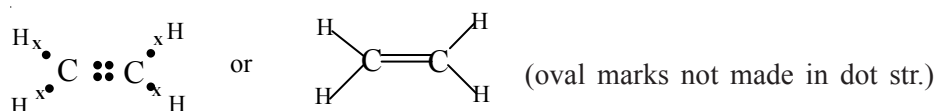
$\text{NH}_3$ : N atom has one lone pair and three unpaired electrons. These three unpaired electrons form three single bonds with three H atoms as shown below.



$\text{CO}_2$ : C has four unpaired electrons and no lone pair in the excited state while each O has two lone pairs and two unpaired electrons. The central atom is C as it is less electronegative than O. The two O atoms are to be connected with C on either side of it. The two unpaired electrons of each O has to match with two unpaired electrons of C by means of a double bond on either side of C.



$\text{C}_2\text{H}_4$ (Ethylene): C is the central atom and there are two carbon atoms which are to be placed first. Then four H atoms are to be symmetrically distributed about the two carbon atoms. Obviously two H atoms are to be attached to each carbon atom. Each C has four unpaired electrons while each H has one unpaired electron. Two of the unpaired electrons of each carbon will match with the unpaired electrons of two H atoms to form two C-H single bonds on either side. Thus each carbon atom is left with two unpaired electrons. They will hand shake with each other with a double bond.



$\text{C}_2\text{H}_2$ (Acetylene): In this case the two C atoms will have to use three unpaired electrons each to form a triple bond between them.



#### (iv) **Octet Rule:**

You already know that by sharing electrons and by forming covalent bonds each atom attains the stable electronic configuration of the nearest inert gas. Carbon(6) in the above example, attains the configuration of Ne(10) and H attains the configuration of He(2). If you consider the valence shell only(which is drawn in the Lewis structure while the inner electrons are not shown), each atom except H will attain the stable octet( $s^2p^6$ ) configuration of an inert gas. H will attain the stable duplet( $s^2$ ) configuration. So if you count the dots enjoyed by each atom(other than H), you should find the total electrons inclusive of the lone pairs will be equal to eight in the Lewis structures. In the above example, each carbon is enjoying 8 electrons(six from the triple bond and two from the one C-H single bond). H is enjoying two electrons. In all the previous examples that we have seen so far e.g  $\text{CH}_4$ ,  $\text{H}_2\text{O}$ ,  $\text{NH}_3$ ,  $\text{CO}_2$ ,  $\text{C}_2\text{H}_4$ , all the atoms other than H namely C, O, N enjoy eight electrons(octet) each. This is called the octet rule. During bond formation, the central atom and other atoms excepting H will have eight electrons in the valence shell(Lewis structure). As a matter of first choice, the central or any other atom(except H) cannot violet the octet rule. However if it is impossible for the central atom to attain octet configuration, it will violet the octet rule. In such cases the atom can enjoy more or less than eight electrons. We shall look to those cases a little later.

### **TYPES OF BONDS ON THE BASIS OF ELECTRON SHARING:**

On the basis of the nature of sharing of electrons by the atoms, there are three kinds of covalent bonds.

- (i) Non-polar covalent bond                      (ii) Polar covalent bond      (iii) Coordinate covalent bond or Dative bond

#### **Non-Polar Covalent Bond:**

When the two atoms are identical, the bond is called non polar covalent bond.

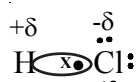


**This is a case of equal contribution and equal sharing.** The three bond pairs are equally shared by the two identical atoms. We shall know more about the non-polar covalent bond after we know about polar covalent bond.

### Polar Covalent Bond:

When the two atoms are different, the bond is called polar covalent bond.

Take the case of HCl.



Each atom contributes one electron to form the bond pair. But the bond pair is not equally shared by the two dissimilar atoms. You know that Cl is more electronegative than H (Cl=3 and H=2.1). So Cl atom will attract the bond pair a little more than the H atom. Thus the bond pair is slightly tilted or shifted to the side of the Cl. Note that electron pair does not entirely go to Cl. In that case, it would form  $\text{H}^+$  ion and  $\text{Cl}^-$  ion and make the ionic bond instead of covalent bond. The electron pair will remain in between the two atoms and will be shared by the two atoms but the sharing will be unequal. The electron pair instead of sitting exactly at the mid point between the two atoms, sits a little closer to Cl atom i.e a little away from H atom. Therefore Cl atom will gain a little additional share over H's electron while H atom loses a little share over its own electron. This results in the development of a fractional positive charge (+ $\delta$ ) on H atom and a fractional negative charge (- $\delta$ ) on the Cl atom. In general, the less electronegative element (like H) will have + $\delta$  charge and more electronegative element (like Cl) will have - $\delta$  charge over them. Had the charges been +1 and -1 in stead of fractional (+ $\delta$  and - $\delta$ ), the bond would have been ionic. But in this case the bond is covalent but having polar or a little ionic character. This separation of charges within a neutral molecule i.e development of partial ionic character in the covalent molecule is called polarity and the bond is called a **polar covalent bond**. *This is a case of equal contribution but unequal sharing of electrons.*

In  $\text{H}_2\text{O}$  molecule, the two O-H bonds are polar covalent bonds. There is + $\delta$  charge on the less electronegative H atom and - $\delta$  charge on more electronegative O atom.

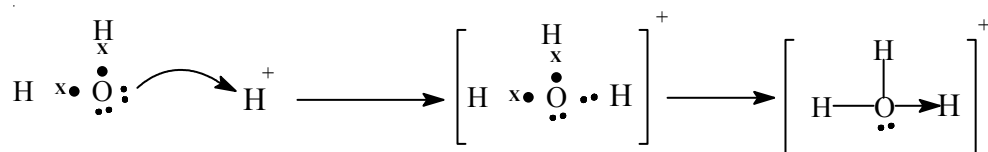


In  $\text{NH}_3$  also, the three N-H bonds are polar, with N having - $\delta$  charge and H having + $\delta$  charge. So in general, the bond between any two dissimilar atoms is polar in nature with + $\delta$  remaining on the less electronegative atom and - $\delta$  remaining on the more electronegative atom.

### (iii) Coordinate Covalent Bond (Dative Bond):

If out of the two atoms, one atom contributes two electrons (electron pair) and the other atom does not contribute any electron, then the bond formed between them will be called coordinate covalent or dative bond. It is a case of one-sided contribution but both-sided sharing. *Dative* means *donated* i.e out of the two electrons contributed by one atom, as if, one electron is completely donated to the other atom because it has no electron to contribute. See this example.

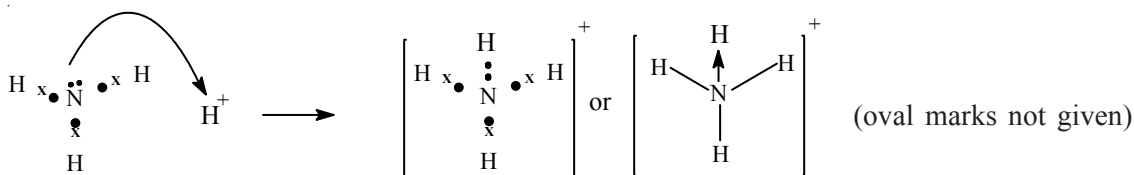
(i)  $\text{H}_3\text{O}^+$  (hydronium ion): (oval marks not given below)



When water molecule reacts with hydrogen ion ( $\text{H}^+$ ), we get hydronium or hydroxonium ion ( $\text{H}_3\text{O}^+$ ). This is formed when one lone pair of O atom in water forms a coordinate covalent bond with  $\text{H}^+$  ion.  $\text{H}^+$  has no electron to contribute. For forming a bond with O atom, it has to seek the donation of electron by O atom which is called the **donor atom**. H atom will be called the **acceptor atom**. Out of the two lone pairs present on O atom, it uses one lone pair and forms a new bond with H. This lone pair is converted to the bond pair. As if out of the two electrons, it completely donates one electron to  $\text{H}^+$  ion thus the positive charge on H is neutralised. By losing one electron completely, O now carries the +ve charge which is customarily shown above the bracket symbol. Thus a dative bond is formed between O and H. A dative bond is shown by an arrow mark ( $\longrightarrow$ ) as shown in the figure from the donor atom to the acceptor atom.

(ii)  $\text{NH}_4^+$  ion:

$\text{NH}_3$  reacts with  $\text{H}^+$  ion to form ammonium ion ( $\text{NH}_4^+$ ) by forming a dative bond between N and  $\text{H}^+$ . The lone pair of N takes part in the formation of this dative bond.



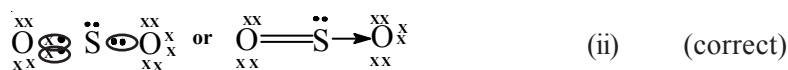
**Practice:** Indicate which of the following forms non-polar and which polar covalent bonds  
 HF, O<sub>2</sub>, Br<sub>2</sub>, HBr, I<sub>2</sub>, N<sub>2</sub>, H<sub>2</sub>S

### LEWIS STRUCTURES OF SOME MORE COVALENT MOLECULES:

We know that while drawing the Lewis structures, we should see that the central atom and all other atoms excepting H contain 8 electrons (octet) in the valence shell. For that purpose we have to make a dative bond in place of double bond or vice versa whenever situation demands. Note that a dative bond is considered to be a single bond having two electrons for sharing. Look to the Lewis structure of SO<sub>2</sub> and SO<sub>3</sub>.

#### SO<sub>2</sub>:

Both O and S belong to the same family. Each of them (see table) has two lone pairs and two unpaired electrons. Since S is less electronegative, it is the central atom. The O atoms have to be placed on either side of S. Then the lone pairs and unpaired electrons are distributed on each atom. First of all, we have to establish a double bond between the S and O oxygen atom by employing two unpaired electrons of each. Now we do not have any unpaired electron left on S to form covalent bond with the other oxygen. We have now two possibilities.



(i) One of the lone pairs of S has to be broken and make two additional unpaired electrons and form a double bond with the other O atom using its two unpaired electrons.

(ii) The other O atom has to make pairing of the two unpaired electrons and form the third lone pair, thus giving a room free for forming a dative bond with S. S uses one of its lone pairs and forms a dative bond with O atom which now has three lone pairs. Note that O has four orbitals (one s and three p) in its valence shell and hence it has to make one orbital free for accommodating the lone pair of S in forming the dative bond. Therefore it has to make pairing of its two unpaired electrons to form the third lone pair.

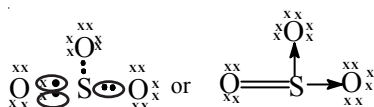
Out of the two possibilities let us find which one will be correct.

Look to the S atom in the 1st case. Count the electrons it is sharing. It is sharing 10 electrons (four bond pairs and one lone pair). This violates the octet rule which says that usually the central atom should share a total of eight electrons in the valence shell. In the second possibility, a dative bond has been made in place of a double bond. In this case the central S atom shares eight electrons (three single bonds and one lone pair). Thus the second Lewis structure is correct.

(CAUTION: Young readers of class IX and X should obey octet rule wherever possible for writing Lewis structures. However, in higher classes they will know that octet rule is violated in many cases)

#### SO<sub>3</sub>:

In SO<sub>3</sub>, there are three O atoms which surround the central S atom. There will be one more dative bond as compared to SO<sub>2</sub> which will be formed with the help of the remaining lone pair of S.



If we make two double bonds in place of two dative bonds, S would share 12 electrons. In this structure S shares 8 electrons. *Note that wherever there is dative bond with O, that O atom has three lone pairs.*

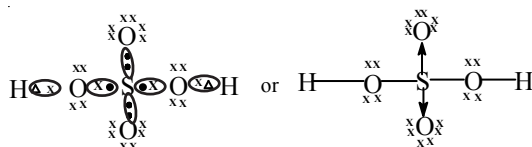
### LEWIS STRUCTURES OF MOLECULES CONTAINING MORE THAN THREE ATOMS:

**Rule:** *H atom cannot be directly attached to the central atom.*

Take the examples of the common mineral acids:  $\text{H}_2\text{SO}_4$ ,  $\text{H}_3\text{PO}_4$ ,  $\text{HNO}_3$

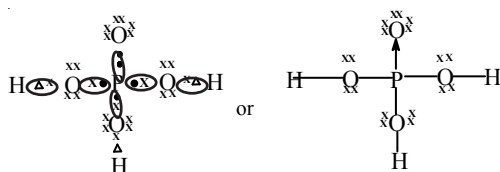
#### $\text{H}_2\text{SO}_4$ :

S is the central atom in this case. First, we have to keep four O atoms distributed around S. Then keep two H atoms symmetrically adjacent to two O atoms on either side of S but away from it. This is because, H atom cannot be directly attached to the central atom according to rule. So that they have to be attached to O atoms. Then place the lone pairs and unpaired electrons on each atom. S and O each contain two lone pairs and two unpaired electrons while H has one unpaired electron only. After placing the electrons, make planning to establish bonds between the atoms. Towards the side where H atom has been placed, S has to establish single bond with O by using one unpaired electron. So S uses two unpaired electrons on either side of it to form two single bonds with the two O atoms to which H atoms are to be connected. Each such O atom uses one unpaired electron for S and one for the H atom on the other side to form single bonds with S and H respectively. Now we have only two lone pairs on S atom. If we break the lone pairs and make double bonds with the remaining O atoms, we will violate the octet rule. S will share 12 electrons. In such case, we must think of dative bonds. S uses two lone pairs separately to establish two dative bonds with the two O atoms. Thus S shares now eight electrons. The O atoms which are bonded to S atom by dative bonds have now three lone pairs each.



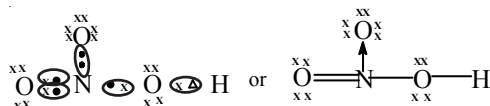
#### $\text{H}_3\text{PO}_4$ :

P is the central atom. Four O atoms are placed symmetrically surrounding P. Three H atoms are placed beside the three O atoms. One O atom is left without H. P has one lone pair and three unpaired electrons like nitrogen atom. O has two lone pairs and two unpaired electrons. Place the electrons and plan for establishing bonds. P has to use the unpaired electrons in the direction of H atom to make three single bonds with the three O atoms. These three O atoms will make single bonds with P as well as with H atoms by using two of unpaired electrons each. P has now one lone pair which it will use for making dative bond with the remaining O atom. It cannot make double bond by breaking the lone pair because by that it will violate the octet rule.



#### $\text{HNO}_3$ :

N is the central atom. Three O atoms are to be placed first surrounding N. Then place one H atom beside any one O atom. After placing the valence electrons (N has one lone pair and three unpaired electrons while for O and H you know), first make a double bond between N and one free O (not connected with H) by using two unpaired electrons of each. N has one more unpaired electron which is to be used to make a single bond with O in the direction of H. That O atom will have to form single bond with N as well as with H on either side of it by using two of its unpaired electrons. Now we have only one lone pair on N which is to be used to make a dative bond with the remaining O. If you make a double bond, the octet rule will be violated.

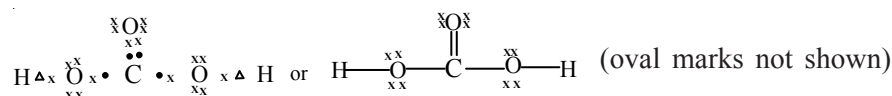


## LEWIS STRUCTURES OF NEGATIVE IONS

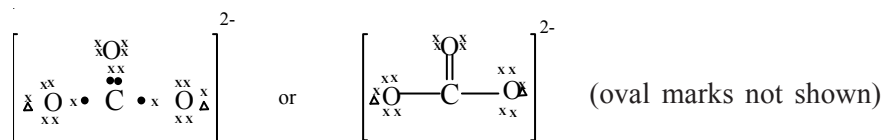
For writing the Lewis structures of negative ions like  $\text{SO}_4^{2-}$ ,  $\text{NO}_3^-$ ,  $\text{CO}_3^{2-}$  etc. we shall have to first write the Lewis structures of the corresponding acids and then remove (wipe out) the H atoms from that structure. Note that while rubbing of the H atoms, leave the electron of H atom on the O atom, so that a new lone pair is formed. The presence of additional electron coming from H atom forms a negative charge on O atom.

### $\text{CO}_3^{2-}$ :

First let us draw Lewis structure of  $\text{H}_2\text{CO}_3$ . C is the central atom and has four unpaired electrons for bonding. First we have to make a double bond between C and O by using two unpaired electrons each. Then we make single bonds between C and the remaining two O atoms (to be connected with H atoms) by using the remaining unpaired electrons of C and one unpaired electrons of each of the two O atoms. These two O atoms also make single bonds with H atoms by using each of its remaining unpaired electrons with unpaired electrons of H atoms. Note that we cannot make a dative bond in place of a double bond in between C atom and O atom. That will violate the octet rule as in that case C will share only six electrons.



Now the two H atoms have to be removed with the creation of a new lone pair and consequently -1 charge on each oxygen. In stead of writing -1 on each O, we shall write the total charge of 2- above the bracket as shown below.



Note that the electron of H indicated by a symbol of triangle(D) is present on each O atom. This is responsible for the negative charge of -2 for the two additional electrons left by H atoms.

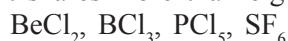
Similarly we can write the structure of  $\text{SO}_4^{2-}$  and  $\text{NO}_3^-$  ion first writing the structures of  $\text{H}_2\text{SO}_4$  and  $\text{HNO}_3$  respectively and then rubbing off the H atoms as explained before. Refer the structures of these two acids given earlier. The structures of their ions would be as follows. We give only the dash structures.



Likewise the reader is supposed to practise writing Lewis structures of some more ions.

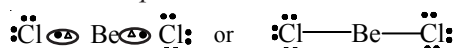
## LEWIS STRUCTURES OF MOLECULES VIOLATING THE OCTET RULE:

There are many molecules in which the central atom cannot, by any means, share eight electrons. It shares either less than eight electrons, in such case we call it to be an *electron deficient molecule* or in other cases it shares more than eight electrons. See the following cases.



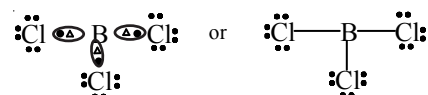
### $\text{BeCl}_2$ :

Be has two electrons in the valence shell for bonding (see table). It forms two single bonds with two chlorine atoms. Each Cl atom uses its one unpaired electron for the purpose. Thus Be shares only four electrons. This is a violation of octet rule. Note that when there is no other way, octet rule is violated. *So you should not remain under the impression that octet rule can never be violated.*



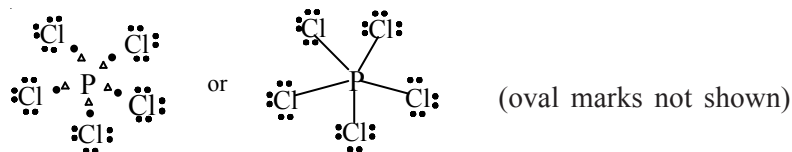
### $\text{BCl}_3$ :

B has three electrons in the valence shell for bonding (see table). Each Cl has one unpaired electron for bonding. So there will be three single bonds and B will share only six electrons. This is violation of octet rule.



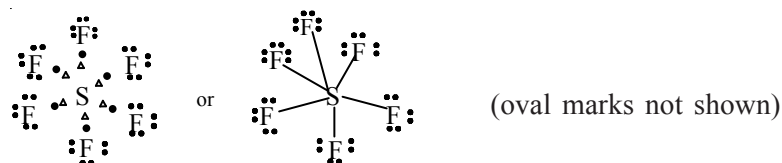
### PCl<sub>5</sub>:

P has one lone pair and three unpaired electrons in its valence shell. But it needs five unpaired electrons to make five single bonds with five Cl atoms. So it has to break the lone pair and produce two more unpaired electrons. Now it has five unpaired electrons to make five single bonds. In PCl<sub>5</sub>, P thus shares 10 electrons in stead of 8. This is again a violation of octet rule.



### SF<sub>6</sub>:

S has two lone pairs and two unpaired electrons in its valence shell like O. In SF<sub>6</sub>, S needs six unpaired electrons to make six single bonds with six F atoms. So it has to break both of its lone pairs and produce four more unpaired electrons. Thus there are six unpaired electrons for the purpose.



### Practice 1:

- Draw the Lewis dash structures of the following. N<sub>2</sub>O<sub>3</sub>, N<sub>2</sub>O<sub>5</sub> (there is one bridging O between N atoms)
- Why it is possible for P and S atom to break their lone pair and produce additional unpaired electrons? (Ans: P and S has vacant 3d orbitals which N and O do not have)

## PROPERTIES OF COVALENT SUBSTANCES

Substance which contains covalent bonds is called a covalent substance. This may be a compound or an element. Covalent substances exhibit different properties as compared to ionic compounds. These are discussed below.

### 1. Physical State:

Covalent compounds are usually gases, liquids or soft and low melting solids at room temperature. Say for example, at room temperature CO<sub>2</sub>, CO, NH<sub>3</sub>, H<sub>2</sub>, N<sub>2</sub>, O<sub>2</sub>, PH<sub>3</sub>, HCl, HBr, HI, SO<sub>2</sub>, SO<sub>3</sub>, H<sub>2</sub>S, C<sub>2</sub>H<sub>4</sub>, C<sub>2</sub>H<sub>2</sub> etc are all gases while Br<sub>2</sub>, CCl<sub>4</sub>, H<sub>2</sub>O, benzene(C<sub>6</sub>H<sub>6</sub>), chloroform(CHCl<sub>3</sub>) etc. are liquids and I<sub>2</sub>, P<sub>4</sub>, S<sub>8</sub> etc. are soft solids.

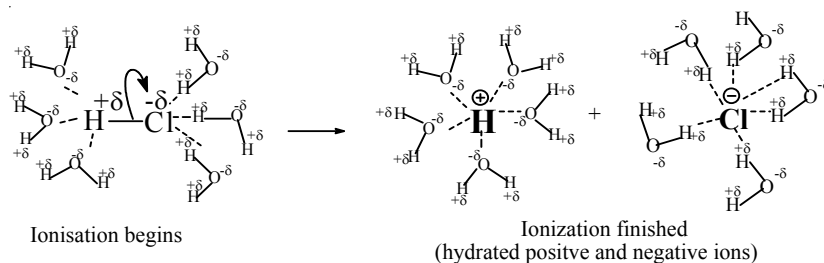
### 2. Melting and Boiling Points:

The covalent substances have low melting and boiling points. For example the boiling point of NH<sub>3</sub> is -33°C and its melting point is still less. The same is the case with other gases. The melting and boiling points of covalent liquids and solids are also low. Take for example, sulphur (S<sub>8</sub>) melts at 113°C while an ionic compound like NaCl melts nearly at 800°C. Because their melting and boiling points are low, covalent substances remain mostly as gases or liquids. The reason for this will be discussed later.

### 3. Electrical Conductivity:

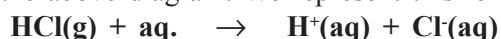
Covalent substances do not conduct electricity in any state. However some covalent compounds do conduct electricity in solution as they undergo ionisation when dissolved in water. For example, HCl(gas) is a covalent compound, but when HCl is dissolved in water, it undergoes ionisation to give H<sup>+</sup> and Cl<sup>-</sup> ions. Due to presence of free ions in solution, it conducts electric current. All mineral acids like H<sub>2</sub>SO<sub>4</sub>, HNO<sub>3</sub>, H<sub>3</sub>PO<sub>4</sub> etc. are covalent compounds and they do not possess free ions in pure state. But when they dissolve in water, undergo ionisation producing ions. That is why they conduct electricity. Let us know now how the covalent compounds like HCl, H<sub>2</sub>SO<sub>4</sub> etc. undergo ionisation in water.

#### Ionisation of HCl(g) in water:





HCl is a covalent molecule and has polar covalent bond. H is the +ve pole and carries + $\delta$  charge while Cl is the negative pole and carries - $\delta$  charge. H<sub>2</sub>O is also a strongly polar compound which is acting as a solvent. H<sub>2</sub>O has two O-H bond dipoles. O is the negative pole having - $\delta$  charge and H is the positive pole carrying + $\delta$  charge. When H-Cl molecule enters in water, large number of water molecules start surrounding each H-Cl molecule in such a manner that the -ve poles of H<sub>2</sub>O molecules face the +ve pole of HCl and the +ve poles of H<sub>2</sub>O molecules face the -ve pole of the same HCl molecule. Water is the solvent and is present in large amount. Large number of H<sub>2</sub>O molecules face the two poles of each HCl molecule as explained and as shown in the above diagram. Due to strong attraction between water molecules and each pole of HCl, the two atoms H and Cl are pulled apart from each other with the bonding pair of electrons shifting to the more electronegative atom chlorine. Ultimately the water molecules succeed in breaking the bond between H and Cl and bringing about ionization. The ions H<sup>+</sup> and Cl<sup>-</sup> do not remain in the naked state. They are jailed inside the solvent cage. Water molecules surround the +ve ion(H<sup>+</sup>) with their negative poles facing the +ve ion and also water molecules surround the -ve ion(Cl<sup>-</sup>) with their positive poles facing the -ve ion. These are called the hydrated ions(hydrated H<sup>+</sup> ion and hydrated Cl<sup>-</sup> ion) as shown in the above diagram. We represent this ionisation as follows.



The process of surrounding of the water molecules to an ion is called **hydration of ion**. During hydration of ions, some amount of energy is liberated. This is called **hydration energy**. We shall study more about this in higher classes. These hydrated ions are responsible to conduct electricity in aqueous solution. Ionisation of polar covalent molecules like HCl, HBr, H<sub>2</sub>SO<sub>4</sub>, HNO<sub>3</sub> etc. in water is due to hydration of ions. The strongly polar solvent molecules are responsible to break the covalent bond and produce hydrated ions.

**Practice:** In the light of hydration of ions, explain why ionic compounds are soluble in water.

Answer: The positive poles of water molecules pull the +ve ion of ionic compound say Na<sup>+</sup> while the negative poles of water molecules pull the -ve ion of the ionic compound say Cl<sup>-</sup>, thus the ions get separated.

### (3) **Solubility:**

Covalent compounds which are non-polar or weakly polar in nature are not soluble in strongly polar solvent like water. They are soluble in non-polar solvents like carbon tetrachloride(CCl<sub>4</sub>), benzene(C<sub>6</sub>H<sub>6</sub>) or weakly polar solvents like chloroform(CHCl<sub>3</sub>), acetone etc. Take for example, I<sub>2</sub> is a non-polar covalent solid and is very poorly soluble in water, but highly soluble in non-polar solvents like CCl<sub>4</sub>. H<sub>2</sub>, N<sub>2</sub>, O<sub>2</sub>, Cl<sub>2</sub> which are non-polar gases are very poorly soluble in water. CCl<sub>4</sub> which is a non-polar compound (a liquid) does not mix with water. Benzene which also a non-polar liquid does not mix with water. However the covalent compounds which are polar in nature like HCl, HBr, NH<sub>3</sub>, SO<sub>2</sub>, glucose(C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>), ethyl alcohol(C<sub>2</sub>H<sub>5</sub>OH) etc. are soluble in water. This is because water is also a polar compound. Some of the polar compounds like HCl, HBr etc. dissolve due to ionisation (accompanied by hydration), some other covalent molecules like glucose, ethyl alcohol dissolve due to some other factor. This, we shall discuss a little later. So now we know that not only ionic compounds are soluble in water but also polar covalent compounds are soluble in water. As a matter of fact, polar covalent compounds can be considered as partially ionic and behave similar to ionic compounds so far as solubility in water is concerned. All these observations have resulted a famous slogan in the principle of solubility:

#### **Like dissolves like.**

This means that if the solvent is polar(H<sub>2</sub>O) then it will dissolve polar compounds and ionic compounds.(Note that ionic compounds are 100% polar while polar covalent compounds are partially ionic and so ionic character is common in both). If the solvent is non-polar (CCl<sub>4</sub> type), then it will dissolve non-polar compounds like H<sub>2</sub>, N<sub>2</sub>, I<sub>2</sub>, P<sub>4</sub> etc. We shall discuss about its reason a little later.

#### **Practice 1:**

(i) NH<sub>3</sub> is a covalent molecule but aqueous solution of NH<sub>3</sub> conducts electricity. Explain.(Ans: Forms NH<sub>4</sub><sup>+</sup> and OH<sup>-</sup> in aq. solution)

(ii) CO<sub>2</sub> is weakly soluble in water while HCl is highly soluble in water. Why? (Ans: CO<sub>2</sub> produces H<sub>2</sub>CO<sub>3</sub> in water which is a weak acid and forms small number of ions. HCl is a strong acid and dissociates in water forming large number of ions)

(iii) Indicate which among the following are weakly soluble and which highly soluble in water. I<sub>2</sub>, HI, SO<sub>2</sub>, H<sub>2</sub>, H<sub>2</sub>SO<sub>4</sub>, CCl<sub>4</sub>, N<sub>2</sub>. (Ans: highly soluble : HI, H<sub>2</sub>SO<sub>4</sub>, SO<sub>2</sub>

(4) **Shape:**

Covalent molecule has a fixed shape. This is because covalent bond is *directional* in nature. This is unlike an ionic bond which is non-directional. There is a fixed bond angle between two covalent bonds within a molecule. For example, the two O-H bonds in  $\text{H}_2\text{O}$  make an angle of  $104.5^\circ$ . This makes the shape of  $\text{H}_2\text{O}$  V-shaped(or bent). Any two N-H bonds in  $\text{NH}_3$  make an angle of  $107^\circ$ . Thus the shape of  $\text{NH}_3$  is pyramidal. N is sitting at the apex of an imaginary pyramid and the three H atoms are sitting at the three corners of the triangular base. The two C=O bonds in  $\text{CO}_2$  make an angle of  $180^\circ$ . Thus the shape of  $\text{CO}_2$  is linear. Thus all the covalent molecules have fixed shapes and fixed bond angles. This has been revealed by actual experimental techniques. Shape of molecules topic is beyond UTSE syllabus.

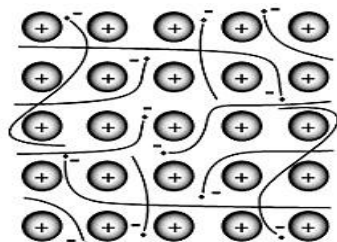
(5) **Isomerism:**

The covalent compounds exhibit a phenomenon called isomerism. This you shall know in organic chemistry in higher classes and hence beyond the scope of UTSE.

### METALLIC BONDING:

This the third type of bonding which is comparable with ionic and covalent bonding. The bonding or interaction present between the atoms of a metal is called metallic bonding.

The valence electrons of metal atoms are loosely bound to their nuclei and are delocalised from their assigned orbitals. In other words, valence electrons are not associated with any specific atom, but wander freely among adjacent metal atoms. These electrons constantly interchange between different atoms. The delocalisation of valence electrons is possible due to overlapping of the valence shell atomic orbitals of neighbouring atoms. For example, in Na metal, the 3s orbital of each Na atom overlaps with 3s orbitals of eight neighbouring Na atoms(since Na has a bcc lattice- refer solid state for the details). Thus infinite number of overlappings of atomic orbitals occur in the bulk of the metal lattice to form infinite number of molecular orbitals having very close energy values. Thus the valence electrons are free to move about all the molecular orbitals extending the whole metallic structure.



The phase of the metal consisting of the delocalised valence electrons is the mobile phase called the '**electron sea**' or '**electron gas**'. Each metal atom releases its valence electron(s) to the electron sea or gas, but its core electrons( for example  $2s^2 2p^6$  for Na) remain static around its nucleus. Thus metal atoms devoid of their valence electrons remain as positive ions or cations. These are called **positive cores** or **kernels** which constitute the static phase. The 'electron sea' constituting the delocalised electrons is free to flow around the regular arrangement of metal cations, but cannot escape completely without getting extra energy from somewhere else. But the positive ions(cores) remain fixed at their positions as long as any external stress is not applied.

Metallic bonding is the electrostatic force of attraction between the static phase(positive ions or atomic cores) and the mobile phase(electron gas or sea). This strong force binds all the atoms together to form a gigantic network. Thus metal consists of a lattice of positive ions(not neutral atoms as we generally believe) immersed or suspended in a sea of electrons(free electrons). These free electrons move in a random manner in all possible directions.

These free or delocalised electrons present in the metals are responsible to explain all the properties of metals such as

- (a) electrical
- (b) thermal conductivity
- (c) malleability and ductility
- (d) metallic lustre
- (e) high melting and boiling point
- (f) strength
- (g) hardness

## INTERMOLECULAR FORCES

In ionic compounds we found that the positive and negative ions are arranged in an alternate manner in a regular order in all directions through infinite distance. This is an infinite network structure and no single and isolated ionic molecule (say NaCl) exists. But if we consider the covalent substances the situation is different. Let us take the case of  $\text{CCl}_4$  which is a liquid at room temperature. The bonds within each  $\text{CCl}_4$  molecule is covalent. There are four covalent (single) bonds in  $\text{CCl}_4$  and it has tetrahedral shape. But in a bottle containing a few millilitres of carbon tetrachloride, there are thousands and thousands of molecules. What kind of relationship exists between these molecules? Are the molecules entirely free and have no attraction between themselves? No doubt, they are independent of each other and exist separately and there is no real chemical bond (ionic or covalent) between the molecules unlike ionic compounds in which ions are interconnected in an infinite network. Between the covalent molecules, however, there exists a very weak attraction and molecules are associated with each other in a very temporary manner. In some compounds this *cohesive association* is extremely weak and in some cases the associations are relatively stronger. But in no case these attractions are comparable with real chemical bonds such as ionic or covalent bond. These are called the intermolecular forces of attractions which exist only in covalent substances as the covalent substances exist in molecular form. Depending on the nature of associations, the intermolecular forces are of two types.

(i) Hydrogen Bonding

(ii) Vander Waals' Forces.

### HYDROGEN BONDING:

This is not true bonding and is wrong to call it a bonding. The reader should not consider this as a real bonding. This is a form of weak intermolecular force of attraction between covalent molecules, although it is strongest among all intermolecular forces of attraction.

#### Conditions:

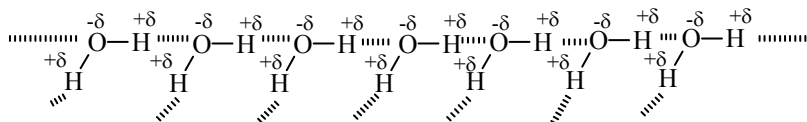
The compound in order to exhibit this type of intermolecular forces should satisfy the following conditions.

- (i) The compound should have H atom and one of the most electronegative and small size atoms such as F, O or N.
- (ii) H should be directly attached (bonded) to one of the most electronegative atoms cited above (F, O or N).

Water, ammonia and hydrogen fluoride satisfy the above requirements and hence are associated by means of hydrogen bonding.

#### Hydrogen bonding in Water ( $\text{H}_2\text{O}$ ):

In a sample of water,  $\text{H}_2\text{O}$  molecules are strongly associated by dipole-dipole attractions. We know that  $\text{H}_2\text{O}$  molecule is strongly polar in nature and contains two O-H polar covalent bonds, in which there is  $+\delta$  on H and  $-\delta$  on O atoms. There is attraction between +ve pole(H) of one water molecule and -ve pole(O) of another neighbouring  $\text{H}_2\text{O}$  molecule. This is called dipole-dipole attraction and this is continued throughout the bulk of the water.



All the  $\text{H}_2\text{O}$  molecules present in that sample are associated in a gigantic network by these dipole-dipole attractions, i.e. +ve and -ve poles of adjacent molecules attracting each other. The hydrogen bonding interactions are shown by broken lines between opposite poles of neighbouring molecules.

#### Hydrogen bonding in Hydrofluoric acid (HF):

Similarly, HF is strongly polar molecule, because of H-F is a dipole. The H pole carries  $+\delta$  and F pole carries  $-\delta$  charge. There is attraction between neighbouring opposite poles throughout the bulk of the substance like water.



#### Hydrogen bonding in Ammonia ( $\text{NH}_3$ ):

Ammonia contains three N-H bonds which are polar in nature. N carries  $-\delta$  and H carries  $+\delta$  charge.

The attractions between opposite poles in neighbouring molecules gives rise to hydrogen bonds similar to HF and H<sub>2</sub>O.

### Consequences of Hydrogen Bonding:

#### (i) *Melting and Boiling Points: Physical State:*

Compounds which exhibit hydrogen bonding have higher boiling points than those which do not exhibit hydrogen bonding. This is because we have to supply large amount of energy to break this innumerable hydrogen bonds. It must be noted that individually the hydrogen bonding interaction is weak as it is not a true bonding, but collectively in presence of thousands and thousands of hydrogen bonding interactions, their cumulative (joint) effect becomes very strong. So the boiling points of such compounds are relatively higher than compounds which do not exhibit hydrogen bonding.

(a) H<sub>2</sub>O boils at 100°C while H<sub>2</sub>S boils at -67°C. Therefore water at room temperature is a liquid while H<sub>2</sub>S is a gas. There is H-bonding interaction present in water but not in H<sub>2</sub>S. H<sub>2</sub>S, no doubt, exhibits intermolecular interaction (about which we shall study soon) but the interactions are much weaker than the H-bonding interactions present in water.

(b) HF has the boiling point of +19°C while HCl has the boiling point of -85°C. In cold seasons, HF remains as a liquid while HCl is always a gas. The high boiling point of HF is due to the presence of H-bonding in HF.

(c) NH<sub>3</sub> has the b.p of -33°C while PH<sub>3</sub> has a b.p of -88°C. This is due to the presence of H bonding in NH<sub>3</sub>.

#### (ii) *Solubility in Water:*

Compounds which exhibit H-bonding are soluble in water. This is because water also exhibits H-bonding. If both solute and solvent exhibit H-bonding, then the solute molecules exchange with solvent molecules and H-bonding interactions are also exchanged. This results in mixing of the solute with the solvent. Just like you make friendship with people of your nature and temperament and not with people who have completely different nature. In the same way the nature of intermolecular association within the solute molecules should be same as that within solvent molecules. Then only the exchange between them is possible and dissolution is possible.

(a) Glucose and sucrose (table sugar) are covalent compounds, but are highly soluble in water, whereas CCl<sub>4</sub> is also a covalent liquid but it is not miscible with water. This is because glucose and sucrose themselves exhibit H-bonding interactions within their molecules (they possess a lot of -OH bonds in their structures) and hence are soluble in water which also exhibit H-bonding. However CCl<sub>4</sub> does not exhibit H-bonding and so it cannot dissolve in water.

(b) Ethyl alcohol (C<sub>2</sub>H<sub>5</sub>OH) which is used as liquor is miscible with water in all proportions although it is a covalent compound. It is because ethyl alcohol exhibits H-bonding like water, so they are miscible.

### VAN DER WAALS FORCES:

Covalent compounds which do not exhibit H-bonding interactions, are associated with still weaker intermolecular forces of attraction. These very weak intermolecular attractions are commonly called *Van der Waals forces*. This is divided into two types.

#### (a) *Dipole-dipole Interaction:*

This exists in polar compounds which do not exhibit H-bonding. This is analogous to H-bonding interaction but is weaker than it. For example, HCl, H<sub>2</sub>S, PH<sub>3</sub> etc. which are analogous to HF, H<sub>2</sub>O and NH<sub>3</sub> respectively in structure are associated in an analogous manner as described in H-bonding. But the attractions are much weaker and are not called H-bonding. They are simply called dipole-dipole attraction.



Due to smaller size of F, O and N (so that the molecules can come very close to each other) as well as their higher electronegativity (so that the magnitude of + $\delta$  and - $\delta$  charges are more) the H-bonding interactions in HF, H<sub>2</sub>O and NH<sub>3</sub> are much stronger than normal dipole-dipole attractions present in HCl, H<sub>2</sub>S, PH<sub>3</sub> etc.. These normal dipole-dipole interactions belong to the Van der Waals forces.

(IMPORTANT NOTE : Nearly 70% of the intermolecular forces present in molecules showing dipole-dipole attractions like HCl, PH<sub>3</sub>, H<sub>2</sub>S etc. are belong to London forces type which is described below)

(b) **London Forces:**

These are the weakest among all intermolecular forces. This type of interaction exists in non-polar molecules like  $I_2$ ,  $H_2$ ,  $N_2$ , benzene,  $Br_2$ ,  $CCl_4$  etc. Since non-polar molecules do not carry any fractional charges ( $+\delta$  and  $-\delta$ ) on their atoms, we may think that there cannot be any attraction between the molecules. This is not so. Although these are non-polar, they acquire small polarity on temporarily basis. This temporary polarity constantly fluctuates in its direction. Due to rapid rotation, vibration, translation (movement) of molecules (particularly in the gaseous state), the electron cloud of the non-polar molecules are temporarily shifted to one side. As a result a  $+\delta$  charge develops on the side in which electron density is less and  $-\delta$  charge develops on the side in which the electron density is more. This is called a **temporary dipole**. These temporary dipoles continuously change their directions .

The attractions between the opposite poles of the neighbouring temporary dipoles are very weak and are called London forces which are named after its discoverer London. Note that some authors address the intermolecular forces among nonpolar compounds as Van der Waals forces. This is not untrue. London forces are merely one kind of forces under Van der Waals category. Dipole-dipole attractions are relatively stronger while London forces are weaker. But both belong to Van der Waals type of forces. It has been already said that more than 70% of contribution in polar molecules under Van der Waals category comes from London type of forces and rest 30% is from actual dipole-dipole attraction.

**Note that Van der Waals forces increase with increase in size of the molecules.**

Example:

(i)  $F_2$ ,  $Cl_2$ ,  $Br_2$ ,  $I_2$ : If we consider the halogens which are covalent substances, we find that the melting points increases smoothly from  $F_2$  to  $I_2$ . While  $F_2$  and  $Cl_2$  are gases and  $Br_2$  is a liquid and  $I_2$  is a solid at room temperature. This is because the size of halogen increases from F to I.

(ii) The boiling point increases smoothly in the order:  $H_2S < H_2Se < H_2Te$ . This is because size increases from S to Te in group VIA and so the Van der Waals forces increases and hence the boiling point increases. Note that  $H_2O$  has not been included in the comparison with the other hydrides of the oxygen family members as shown above. This is because,  $H_2O$  exhibits H-bonding while others exhibit Van der Waals forces.

**Practice 1**

(i) Why do covalent substances have usually lower melting and boiling points compared to ionic compounds ?

(Ans: While we melt or boil an ionic compound, the ionic bond is weakened or broken which needs much greater amount of energy. But when we melt or boil a covalent substance the covalent bond is not weakened or broken. In stead the weak intermolecular forces (H-bonding or Van der Waals forces) are further weakened. Hence energy required is less.)

**(PLEASE**

**SEE**

**BELOW**

**FOR**

**MORE)**

## OXIDATION AND REDUCTION

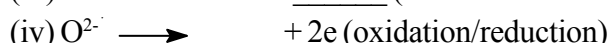
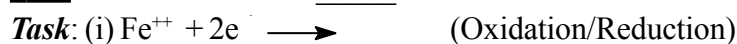
### Loss of Electron (Oxidation) : Cation Formation

Metals lose electrons easily and form cations(+ve ions)



### Gain of Electron(Reduction): Anion formation

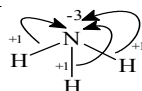
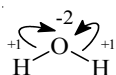
Non-metals gain electron easily to form anions(-ve ion)



### OXIDATION NUMBER(ON) or OXIDATION STATE(OS):

Oxidation Number is the valency of an atom in a molecule or ion which is assigned the sign either + or -. It may be (i) electrovalency or (ii) covalency. When the proper sign is associated with valency it becomes oxidation number(ON). It is mostly a theoretical concept particularly for covalent compounds and ions.

**Definition:** It is the theoretical charge which an atom will possess when all the atoms in a molecule or ion are separated from each other in such a manner that the the bond pairs are transferred to the more electronegative atom. If the two atoms are identical, then each will take back its own electron.



### RULES FOR DETERMINING ON:

(1) The O.N of any element in the free or uncombined state is **zero**. (Zn, Na, Ca, Fe, H<sub>2</sub>, O<sub>2</sub>, P<sub>4</sub>, etc.)

(2) The O.N of element existing as simple monoatomic radical in the compound is equal to the charge of the ion (e.g. Na<sup>+1</sup>, Cl<sup>-1</sup>). The charge of the ion = valency with appropriate sign. For basic part(radical) +ve and acid part(radical) -ve signs are used with the respective valences)

**Task:** What is ON of the underlined element:

(i) CaCl<sub>2</sub>, (ii) NH<sub>4</sub>Cl, (iii) FeSO<sub>4</sub>, (iv) Al2(SO<sub>4</sub>)<sub>3</sub>, (v) ZnS, (vi) Sn(NO<sub>3</sub>)<sub>2</sub>, (vii) Mg<sub>3</sub>N<sub>2</sub>, (viii) HgI<sub>2</sub>

(3) The ON of oxygen(O) is always -2 excepting in peroxides in which it is -1, in superoxides in which it is -1/2 and in OF<sub>2</sub> it is +2.

**Task:** Find the ON of oxygen in the following: ZnO, O<sub>2</sub>, H<sub>2</sub>O, H<sub>2</sub>O<sub>2</sub>, Al<sub>2</sub>O<sub>3</sub>

(4) The ON of hydrogen(H) is always +1 excepting metal hydrides(NaH, CaH<sub>2</sub> etc) in which it is -1.

**Task:** Find the ON of hydrogen in the following: H<sub>2</sub>O, NH<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>, NaH, CaH<sub>2</sub>, AlH<sub>3</sub>

(5) In a neutral compound the sum of O.Ns of all the atoms of all the elements is zero.

**Example:** Let us take the molecule SO<sub>3</sub>. What is the ON of S in this molecule?

Let us assume the ON of S to be x. The sum of the ONs of all the atoms in the molecule is zero as per this rule.

$x + 3(-2) = 0$ , So  $x = +6$ . For one oxygen atom the ON is -2 and therefore for 3 oxygen atom it is  $3(-2) = -6$ ; The ON of S in SO<sub>3</sub> is +6.

**Task:** Find the ON of underlined elements. (i) SO<sub>2</sub>, (ii) HNO<sub>3</sub>, (iii) H<sub>2</sub>SO<sub>4</sub>, (iv) KMnO<sub>4</sub>, (v) P<sub>2</sub>O<sub>5</sub>, (vi) K<sub>2</sub>CrO<sub>7</sub>, (vii) KClO<sub>3</sub>

(6) The sum of ONs of all the atoms in a polyatomic ion is equal to the charge of the ion, while the charge of the ion is equal to the valency of the ion with the appropriate sign(acid radical -ve and basic radical +ve).

*Example:*  $\text{SO}_4^{2-}$ . Since the valency of sulphate radical is 2, its charge will be -2 as it is an acid radical. Let us find the ON of S in the ion. According to the above rule the sum of the ONs of all the atoms should be equal to the charge of the ion.

$$x + 4(-2) = -2 \Rightarrow x = +6. \text{ So the ON of S in sulphate ion is } +6.$$

NB: Always ON is determined for one atom, but for calculation purpose all atoms are considered

**Task:** Find the ON of the underlined atoms

i)  $\underline{\text{N}}\text{H}_3$ , (ii)  $\underline{\text{K}}\underline{\text{N}}\text{O}_3$ , (iii)  $\underline{\text{N}}\text{O}_2^-$ , (iv)  $\underline{\text{N}}\text{O}_2$ , (v)  $\underline{\text{N}}_2$ , (vi)  $\underline{\text{N}}_2\text{O}_5$ , (vii)  $\underline{\text{K}}\underline{\text{Mn}}\text{O}_4$ , (viii)  $\underline{\text{Cr}}_2\underline{\text{O}}_7^{2-}$ ,  
(ix)  $\underline{\text{S}}_2\underline{\text{O}}_3^{2-}$ , (x)  $\underline{\text{Cl}}\underline{\text{O}}_3^-$ , (xi)  $\underline{\text{C}}\underline{\text{O}}_2$  (xii)  $\underline{\text{N}}\underline{\text{O}}_3^-$  (xiii)  $\underline{\text{S}}\underline{\text{O}}_3$  (xiv)  $\underline{\text{S}}\underline{\text{O}}_3^{2-}$

Home task: (xi)  $\underline{\text{N}}\underline{\text{H}}_4^+$ , (xii)  $\underline{\text{N}}\underline{\text{O}}_3^-$ , (xiii)  $\underline{\text{N}}\underline{\text{H}}_4\underline{\text{N}}\underline{\text{O}}_3$ , (xiv)  $\underline{\text{Mn}}\underline{\text{O}}_2$ , (xv)  $\underline{\text{Cr}}_2(\underline{\text{S}}\underline{\text{O}}_4)_3$ ,  
(xvi)  $\underline{\text{H}}\underline{\text{Cl}}\underline{\text{O}}$ , (xvii)  $\underline{\text{S}}\underline{\text{O}}_3^{2-}$ , (xviii)  $\underline{\text{H}}\underline{\text{Cl}}$ , (xix)  $\underline{\text{Sb}}\underline{\text{Cl}}_3$ , (xx)  $\underline{\text{P}}\underline{\text{H}}_3$  (xxi)  $\text{Pb}(\underline{\text{N}}\underline{\text{O}}_3)_2$   
(xxii)  $\text{Ca}(\underline{\text{O}}\underline{\text{Cl}})\underline{\text{Cl}}$

(Note that the ON of an atom in a covalent molecule ( $\text{NH}_3$ ,  $\text{SO}_2$ ,  $\text{NO}_2$  etc) or ions ( $\text{NO}_3^-$ ,  $\text{CO}_3^{2-}$  etc) is usually the covalency of that atom. Note also that in valency the +/- signs are deleted. In ON, the appropriate signs are given).

### Types of Reactions:

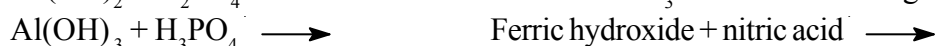
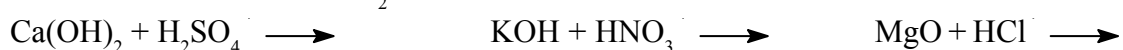
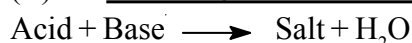
(i) **METATHESIS REACTIONS:** Reactions in which Oxidation Number of each atom remains unchanged.

(ii) **REDOX REACTIONS:** The reaction in which changes in ON of elements take place.

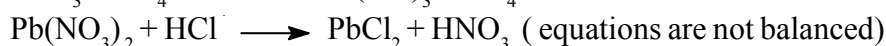
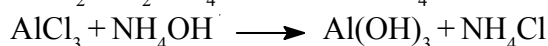
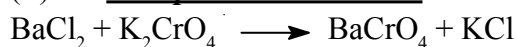
### **TYPES OF METATHESIS REACTIONS**

#### 1. **Double Displacement Reactions:**

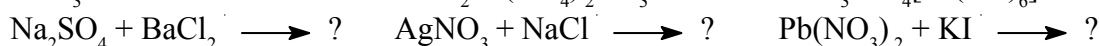
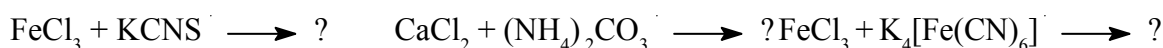
##### (A) Neutralization Reactions:



##### (B) Precipitation Reactions

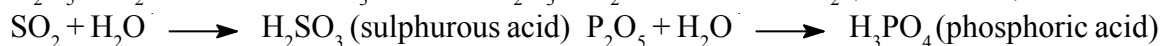


#### **Practice:**



#### 2. **Nonmetallic Oxides with water**

Most oxides of nonmetals form acids when react with water. Therefore they are called acidic oxides (acid anhydrides) (exceptions: CO, NO and  $\text{N}_2\text{O}$  are neutral)

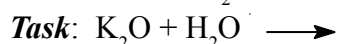


Trick to know the acid anhydride from the formula of acid: Remove all H atoms in the form of water. If required multiply a suitable minimum factor to do that. Example:  $2 \times \text{H}_3\text{PO}_4 - 3 \text{H}_2\text{O} \longrightarrow \text{P}_2\text{O}_5$

**Practice:** Find the anhydrides of  $\text{HClO}_4$  (perchloric acid),  $\text{HClO}_3$  (chloric acid),  $\text{HClO}_2$  (chlorous acid) and  $\text{HOCl}$  (hypochlorous acid). Ans:  $\text{Cl}_2\text{O}_7$ ,  $\text{Cl}_2\text{O}_5$ ,  $\text{Cl}_2\text{O}_3$ ,  $\text{Cl}_2\text{O}$  respectively.

#### 3. **Reaction of metallic oxide with water**

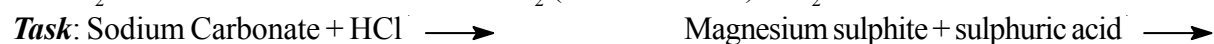
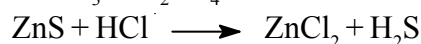
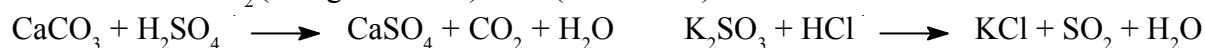
Metallic oxides which are soluble in water form their hydroxides.



#### 4. **Reactions of Carbonate, Sulphite, Sulphide, Nitrite with dilute acids:**



Nitrite  $\longrightarrow$   $\text{NO}_2$  (nitrogen dioxide) +  $\text{NO}$  (nitric oxide)

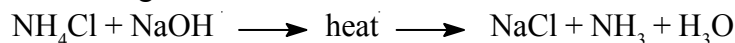


N.B: bicarbonate behaves similar to carbonates



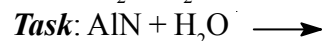
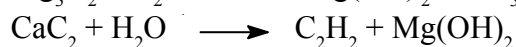
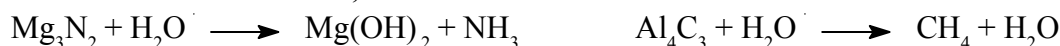
### 5. Heating an Ammonium salts with a base(alkali):

Ammonia gas is evolved



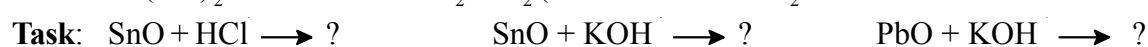
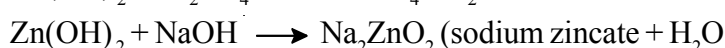
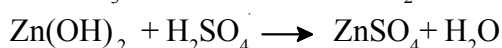
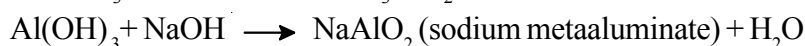
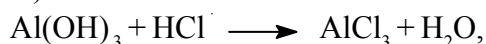
**Task:** ammonium sulphate + calcium hydroxide  $\xrightarrow{\text{heat}}$

### 6. Reaction of metal nitrides, and carbides with water



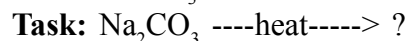
### 7. Amphoterism:

Oxides and hydroxides of Al, Zn, Sn, Pb, Cr and a few others react both with acids and strong bases (NaOH and KOH) to form salt and water. These are amphoteric substances.



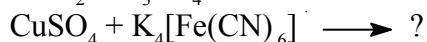
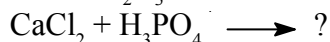
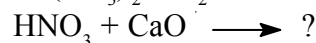
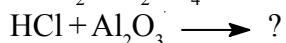
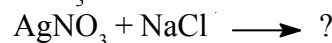
(Ans:  $\text{SnCl}_2 + \text{H}_2\text{O}$ ,  $\text{K}_2\text{SnO}_2 + \text{H}_2\text{O}$ ,  $\text{K}_2\text{PbO}_2 + \text{H}_2\text{O}$ ,  $\text{K}_2\text{PbO}_4 + \text{H}_2\text{O}$ ;  $\text{K}_2\text{SnO}_3 + \text{H}_2\text{O}$ )

### 8. Thermal Decomposition of carbonates and bicarbonates:



#### Double Displacement Reactios(For practice)

(IMPORTANT NOTE: While writing an equation, do not think about balancing it. If you predict the products logically with correct formula then you can balance it any time later. If the prediction of products is wrong, then you cannot balance the equation. Often students give more emphasis on the coefficients of balanced equations. They somehow try to memorise the equation. This is a wrong practice. Never do that. Every reaction has a logic. Why  $A + B$  gives  $C + D$  and not  $E + F$ , has some reason. Once you know the logic, then chemical reactions will be a game for you.)



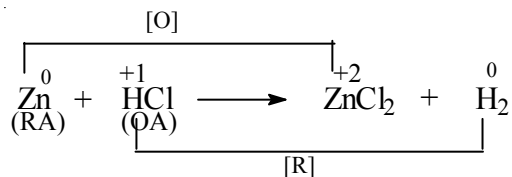
### Oxidation and Reduction Reactions(Redox Reaction)

Now you shall be told about one of the most interesting type of reaction, called *redox reaction*. This type of reaction is also called *oxidation-reduction reaction*, which involves both oxidation as well as reduction. You already know some primitive definitions of oxidation and reduction in terms of addition of oxygen and hydrogen or removal of hydrogen and oxygen. Now you will be knowing the most modern definition of oxidation and reduction.



**Oxidation:** A process in which ON of certain element is increased is called oxidation. This takes place by loss of electron which has been discussed before.

**Reduction:** A process in which ON of certain element is decreased is called reduction. This takes place by gain of electron which has been discussed before. Look to the following example.



In this example, the ON of Zn was zero (uncombined state) in LHS and it became +2 in ZnCl<sub>2</sub> in RHS. Thus the ON of Zn has increased from 0 to +2. This process is oxidation. We say that Zn is oxidised to ZnCl<sub>2</sub>. On the other hand, the ON of H in HCl is +1 in LHS and in H<sub>2</sub> it is 0 in the RHS. Thus the ON of H has decreased from +1 to 0. This is the reduction process. We say that HCl is reduced to H<sub>2</sub>. We found also that wherever there is oxidation, there is a reduction. Whenever there is a loss there is definitely a gain. Just like, if you lose a 100 rupee note on the road, it is your loss, but it is the gain of a thief who stole it. Like loss and gain, oxidation and reduction also go simultaneously. There cannot be an oxidation without having a reduction and vice versa.

The substance which is oxidised is called **Reducing Agent (RA)** and the substance which is reduced is **Oxidising Agent (OA)**. In the above reaction, Zn is oxidised, hence it is the reducing agent (RA) and HCl is reduced, hence it is the oxidising agent (OA). Try these SAQs.

**Practice:** Show by indicating the ON which is oxidised and which reduced. Also indicate which is Oxidising Agent (OA) and which Reducing Agent (RA). Clue elements which have undergone the changes in ON have been shown by sign of asterisk. The equations are not balanced. Don't also try to balance it.

- (i)  $\text{Na}^* + \text{H}_2^* \text{O} \longrightarrow \text{Na}^* \text{OH} + \text{H}_2^*$
- (ii)  $\text{N}_2^* + \text{H}_2^* \longrightarrow \text{N}^* \text{H}_3^*$
- (iii)  $\text{P}_4^* + \text{Cl}_2^* \longrightarrow \text{P}^* \text{Cl}_3^*$
- (iv)  $\text{N}^* \text{H}_3 + \text{Cu}^* \text{O} \longrightarrow \text{Cu}^* + \text{N}_2^* + \text{H}_2\text{O}$
- (v)  $\text{H}_2^* + \text{I}_2 \longrightarrow \text{H}^* \text{I}^*$
- (vi)  $\text{Mg}^* + \text{H}_2^* \text{SO}_4 \longrightarrow \text{MgSO}_4 + \text{H}_2^*$
- (vii)  $\text{N}_2^* + \text{O}_2^* \longrightarrow \text{N}^* \text{O}^*$
- (viii)  $\text{Cu}^* + \text{HNO}_3^* \longrightarrow \text{Cu}^* (\text{NO}_3)_2 + \text{N}^* \text{O}_2 + \text{H}_2\text{O}$
- (ix)  $\text{Mn}^* \text{O}_2 + \text{HCl}^* \longrightarrow \text{Mn}^* \text{Cl}_2 + \text{H}_2\text{O} + \text{Cl}_2^*$
- (x)  $\text{KMn}^* \text{O}_4 + \text{H}_2\text{SO}_4 + \text{Fe}^* \text{SO}_4 \longrightarrow \text{K}_2\text{SO}_4 + \text{Mn}^* \text{SO}_4 + \text{Fe}_2^* (\text{SO}_4)_3 + \text{H}_2\text{O}$

## Some Simple Redox Reactions

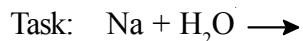
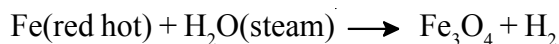
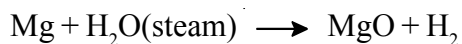
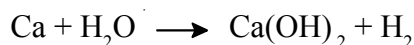
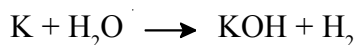
**(IMPORTANT NOTE: For all simple redox reactions, always show the ON of the elements which have undergone changes)**

**(A) Displacement Reactions:** (in acidic medium)

**Metal Activity Series:** Li, K, Ca, Na, Mg, Al, Mn, Zn, Cr, Fe, Co, Ni, Sn, Pb, H, Sb, Bi, As, Cu, Hg, Ag, Pd, Pt, Au

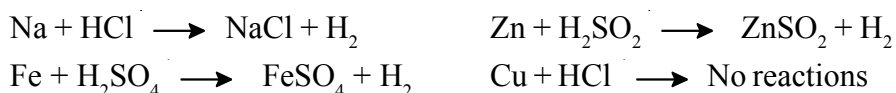
**(i) Reaction of metals with water** (hydrogen displacement from water) :

Most active metals (lying above H in the series) like K, Ca, Na, Mg, Al, Zn, Fe react with water and produce H<sub>2</sub> gas. This is displacement of H from H<sub>2</sub>O. Li, K, Ca and Na react violently with cold water while others react with steam.

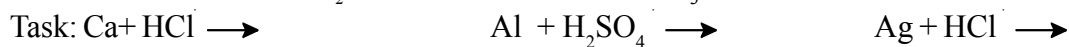


**(ii) Reactions of metals with dilute mineral acids:**

Metals lying above H in the activity series displace H<sub>2</sub> gas from dilute acids like dil. HCl, H<sub>2</sub>SO<sub>4</sub>



**IMPORTANT:**  $\text{HNO}_3$  does not give  $\text{H}_2$  with most metals. In stead it gives  $\text{NO}_2$ ,  $\text{NO}$ ,  $\text{N}_2\text{O}$ ,  $\text{N}_2$  depending on the concentration of  $\text{HNO}_3$  with most metals as it is a strong oxidising agent. (Exception:  $\text{Mg}$  and  $\text{Mn}$  which give  $\text{H}_2$  with very very dilute  $\text{HNO}_3$ )



**(iii) Metal Displacement Reactions:**

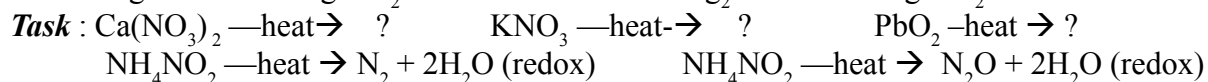
A metal occupying higher position in the activity series can displace another metal lying at a lower position from its compound.



**(B) Thermal Decomposition:**

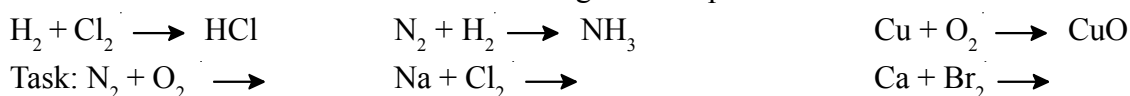
On heating many compounds break down to form a mixture of products. In this case same substance gets oxidized as well as reduced. These are called **disproportionation reactions**.

$\text{Pb}(\text{NO}_3)_2 \xrightarrow{\text{heat}} \text{PbO} + \text{NO}_2 + \text{O}_2$  (All nitrates excepting  $\text{NaNO}_3$  and  $\text{KNO}_3$  react in the similar manner). In this case N changes from +5 to +4 and O changes from -2 to 0.



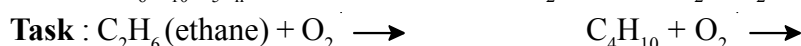
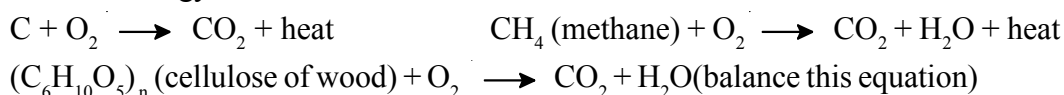
**[C] Synthesis Reactions:**

Two or more elements react with each other to give a compound.



**[D] Combustion Reactions:**

A substance burning in air or oxygen is called combustion reaction. A lot of heat is liberated in these reactions. Hydrocarbons and other organic compounds burn to give  $\text{CO}_2$  and water. Carbon or carbon oxide also burns in oxygen to evolve a lot of heat.

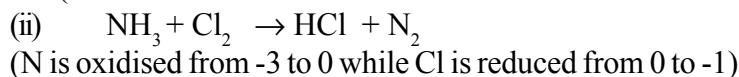
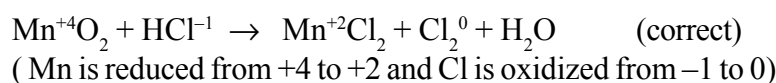
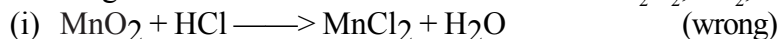


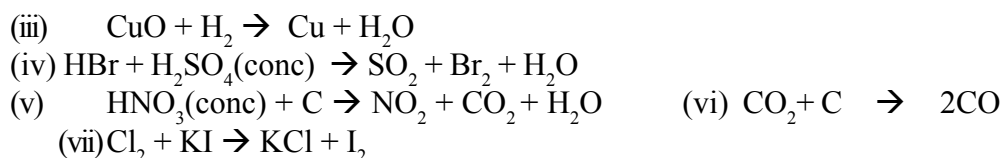
(Incomplete combustion give  $\text{CO}$ ,  $\text{HCHO}$ ,  $\text{CH}_3\text{OH}$  etc.)

**[E] Some simple Redox Reactions**

**Some common OAs :**  $\text{MnO}_2$ ,  $\text{PbO}_2$ ,  $\text{Pb}_3\text{O}_4$ ,  $\text{KMnO}_4$ ,  $\text{K}_2\text{Cr}_2\text{O}_7$ ,  $\text{F}_2$ ,  $\text{Cl}_2$ ,  $\text{Br}_2$ ,  $\text{O}_2$ ,  $\text{HNO}_3$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{O}_3$

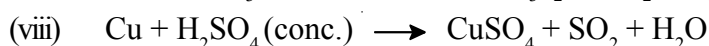
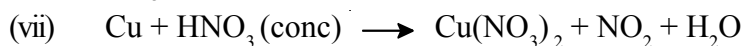
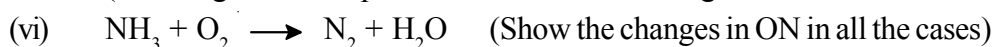
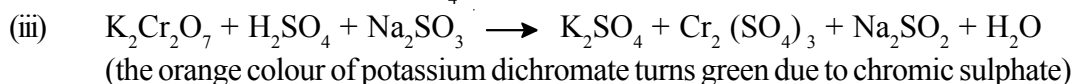
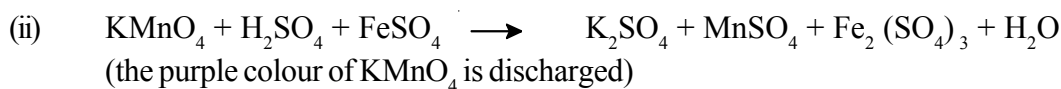
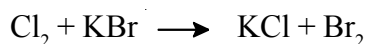
**Some common RAs :**  $\text{H}_2\text{S}$ ,  $\text{FeSO}_4$ ,  $\text{NH}_3$ ,  $\text{HI}$ ,  $\text{HBr}$ ,  $\text{HI}$ , all metals,  $\text{I}_2$ ,  $\text{P}_4$ ,  $\text{S}$ ,  $\text{H}_2$ ,  $\text{CO}$ ,  $\text{C}$ ,  $\text{SnCl}_2$   
 Some agents which can act both as RA and OA :  $\text{H}_2\text{O}_2$ ,  $\text{SO}_2$ ,  $\text{HNO}_2$ ,





**[F] Reactivity among halogens:**  $\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$

$\text{F}_2$  can displace Cl, Br, I from their salts while  $\text{Cl}_2$  can displace Br and I from their salts,  $\text{Br}_2$  can displace only I from its salt and I cannot displace any halogen



(Some important hints for prediction: Usually the ON increases from -ve state to 0 and decreases from +ve state to 0. (Exception : Cl decreases from +ve state to -1, Mn from +7 to +2 in acidic medium, Cr from +6 to +3, Pb and Mn from +4 to +2, Fe from +3 to +2, Hg from +2 to +1 and then to 0, Cu from +2 to +1

Sn increases from +2 to +4, Fe from +2 to +3, Mn from +2 to +7, Cr from +3 to +6 etc)

**Predict the following redox reactions.**

1.  $\text{H}_2\text{O}_2 + \text{H}_2\text{S} \rightarrow$       2.  $\text{H}_2\text{O}_2 + \text{NaOCl} \rightarrow$       3.  $\text{SO}_2 + \text{H}_2\text{S} \rightarrow$
4.  $\text{SO}_2 + \text{H}_2\text{O}_2 \rightarrow$       5.  $\text{HI} + \text{HIO}_3 \rightarrow$       6.  $\text{HI} + \text{H}_2\text{O}_2 \rightarrow$
7.  $\text{Cu} + \text{conc. H}_2\text{SO}_4 \rightarrow$       8.  $\text{Zn} + \text{conc. HNO}_3 \rightarrow$       9.  $\text{PbO} + \text{C} \rightarrow$
10.  $\text{Cl}_2 + \text{FeSO}_4 + \text{H}_2\text{SO}_4 \rightarrow$       11.  $\text{SnCl}_2 + \text{HgCl}_2 \rightarrow$       12.  $\text{KMnO}_4 + \text{H}_2\text{SO}_4 + \text{H}_2\text{O}_2 \rightarrow$
13.  $\text{H}_2\text{SO}_4(\text{conc}) + \text{NaI} \rightarrow$       14.  $\text{H}_2\text{SO}_4(\text{conc}) + \text{NaBr} \rightarrow$       16.  $\text{H}_2\text{SO}_4(\text{conc}) + \text{NaCl} \rightarrow$

**A FEW PRACTICE REACTIONS:**

1.  $\text{NH}_3 + \text{CO}_2 + \text{H}_2\text{O} \rightarrow$       2.  $\text{NaCl} + \text{NH}_4\text{HCO}_3 \rightarrow$
3.  $\text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O} \rightarrow$       4.  $\text{Ca}(\text{OH})_2 + \text{Cl}_2 \rightarrow$
5.  $\text{Ca}(\text{OCl})\text{Cl} + \text{H}_2\text{O} \rightarrow$       6.  $\text{Ca}(\text{OCl})\text{Cl} + \text{HCl} \rightarrow$
7.  $\text{CH}_4 + \text{H}_2\text{O}(\text{steam}) \xrightarrow{\text{high temp. \& Pr}}$
8.  $\text{CuO} + \text{H}_2 \rightarrow$       9.  $\text{Na} + \text{H}_2 \rightarrow$
10.  $\text{Fe}_2\text{O}_3 + \text{CO} \rightarrow$       11.  $\text{MnO}_2 + \text{Al} \rightarrow$
12.  $\text{FeS}_2 + \text{O}_2 \rightarrow$       13.  $\text{PbO} + \text{C} \rightarrow$
14.  $\text{Mg} + \text{N}_2 \rightarrow$       15.  $\text{P}_4 + \text{O}_2 \rightarrow$
16.  $\text{NH}_4\text{NO}_2 \xrightarrow{\text{heat}}$       17.  $\text{KClO}_3 \xrightarrow{\text{heat}}$
18.  $\text{HI} + \text{H}_2\text{O}_2 \rightarrow$

**Task:** Predict the following Reactions, Indicate which is metathesis and which is redox. In redox reactions, indicate which is oxidised and which is reduced by assigning ONs.

1.  $\text{FeS} + \text{H}_2\text{SO}_4(\text{dil}) \longrightarrow$
2.  $\text{CO}_2 + \text{H}_2\text{O} \longrightarrow ?$
3.  $\text{Cr}(\text{OH})_3 + \text{H}_2\text{SO}_4 \longrightarrow ?$
4.  $\text{Na} + \text{H}_2\text{O} \longrightarrow ?$
5.  $\text{CaO} + \text{H}_2\text{O} \longrightarrow ?$
6.  $\text{BaCO}_3 + \text{HCl} \longrightarrow ?$
7.  $\text{Mg}_3\text{N}_2 + \text{H}_2\text{O} \longrightarrow$
8.  $\text{NH}_4\text{Cl} + \text{CaO} \xrightarrow{\text{heat}} ?$
9.  $\text{Cu} + \text{AgNO}_3 \longrightarrow ?$
10.  $\text{Cl}_2 + \text{KI} \longrightarrow ?$
11.  $\text{NH}_3 + \text{HCl} \longrightarrow ?$
12.  $\text{N}_2\text{O}_5 + \text{H}_2\text{O} \longrightarrow ?$
13.  $\text{Mg} + \text{H}_2\text{SO}_4(\text{dil.}) \longrightarrow$
14.  $\text{CaCO}_3 \xrightarrow{\text{heat}} ?$
15.  $\text{Na}_2\text{SO}_3 + \text{dil. H}_2\text{SO}_4 \longrightarrow ?$
16.  $\text{Al} + \text{HCl} \longrightarrow ?$
17.  $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \longrightarrow ?$
18.  $\text{Al}_4\text{C}_3 + \text{H}_2\text{O} \longrightarrow ?$
19.  $\text{CrCl}_3(\text{aq}) + \text{NH}_4\text{OH}(\text{aq}) \longrightarrow ?$
20.  $\text{PbO}_2 \xrightarrow{\text{heat}} ?$
21.  $\text{SO}_3 + \text{NaOH} \longrightarrow ?$
- 22(a).  $\text{Fe} + \text{H}_2\text{SO}_4(\text{dil.}) \longrightarrow$
- 22(b).  $\text{Cu} + \text{Zn}(\text{NO}_3)_2 \longrightarrow ?$
23.  $\text{Mg} + \text{HCl} \longrightarrow ?$
24.  $\text{MnCl}_2(\text{aq}) + \text{H}_2\text{S}(\text{aq}) \longrightarrow ?$
25.  $\text{H}_2 + \text{I}_2 \longrightarrow ?$
26.  $\text{C}_2\text{H}_6 + \text{O}_2 \longrightarrow ?$
27.  $\text{Ca} + \text{H}_2\text{O} \longrightarrow ?$
28.  $\text{Al} + \text{Cl}_2 \longrightarrow ?$
29.  $\text{Pb}(\text{NO}_3)_2 \xrightarrow{\text{heat}} ?$
30.  $\text{Ca}(\text{OH})_2 + \text{H}_2\text{SO}_4 \longrightarrow ?$
31.  $\text{NH}_4\text{NO}_2 \xrightarrow{\text{heat}} ?$
32.  $\text{Mg} + \text{SO}_2 \longrightarrow ?$
33.  $\text{K}_2\text{CO}_3 + \text{H}_2\text{SO}_4(\text{dil}) \longrightarrow ?$
34.  $\text{PbO} + \text{HBr} \longrightarrow ?$
35.  $\text{AlP} + \text{H}_2\text{O} \longrightarrow ?$
36.  $\text{KClO}_3 \xrightarrow{\text{heat}} ?$
37.  $(\text{NH}_4)_2\text{SO}_4 + \text{KOH} \longrightarrow ?$
38.  $\text{Al}(\text{OH})_3 + \text{NaOH} \longrightarrow ?$
39.  $\text{Hg}_2(\text{NO}_3)_2(\text{aq}) + \text{HCl}(\text{aq}) \longrightarrow ?$
40.  $\text{Fe}_2\text{O}_3 + \text{H}_2\text{SO}_4 \longrightarrow ?$
41.  $\text{KNO}_3 \xrightarrow{\text{heat}} ?$
42.  $\text{Fe}(\text{red hot}) + \text{H}_2\text{O}(\text{steam}) \longrightarrow ?$
43.  $\text{Hg}(\text{NO}_3)_2(\text{aq}) + \text{H}_2\text{S}(\text{aq}) \longrightarrow ?$
44.  $\text{Ag} + \text{H}_2\text{SO}_4(\text{dil}) \longrightarrow ?$
45.  $\text{Al} + \text{Fe}_2\text{O}_3 \xrightarrow{\text{heat}} ?$
46.  $\text{SO}_2 + \text{H}_2\text{O} \longrightarrow ?$
47.  $\text{Br}_2 + \text{KCl} \longrightarrow ?$
48.  $\text{CaCl}_2(\text{aq}) + (\text{NH}_4)_2\text{C}_2\text{O}_4(\text{aq}) \longrightarrow ?$
49.  $\text{MnO}_2 + \text{HBr} \xrightarrow{\text{heat}} ?$
50.  $\text{Fe} + \text{Cl}_2 \longrightarrow ?$
51.  $\text{Al} + \text{N}_2 \longrightarrow ?$
52.  $\text{PbO} + \text{H}_2 \longrightarrow ?$
54.  $\text{Zn} + \text{O}_2 \longrightarrow ?$

## ANSWERS TO TASK ON CHEMICAL REACTIONS

### ANSWERS

1.  $\text{FeSO}_4 + \text{H}_2\text{S}(\text{M})$
2.  $\text{H}_2\text{CO}_3(\text{M})$
3.  $\text{Cr}_2(\text{SO}_4)_3 + \text{H}_2\text{O}(\text{M})$
4.  $\text{NaOH} + \text{H}_2(\text{R})$
5.  $\text{Ca}(\text{OH})_2(\text{M})$
6.  $\text{BaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}(\text{M})$
7.  $\text{NH}_3 + \text{Mg}(\text{OH})_2(\text{M})$
8.  $\text{NH}_3 + \text{CaCl}_2 + \text{H}_2\text{O}(\text{M})$
9.  $\text{Cu}(\text{NO}_3)_2 + \text{Ag}(\text{R})$
10.  $\text{KCl} + \text{I}_2(\text{R})$
11.  $\text{NH}_4\text{Cl}(\text{M})$
12.  $\text{HNO}_3(\text{M})$
13.  $\text{MgSO}_4 + \text{H}_2(\text{R})$
14.  $\text{CaO} + \text{CO}_2(\text{M})$
15.  $\text{Na}_2\text{SO}_4 + \text{SO}_2 + \text{H}_2\text{O}(\text{M})$
16.  $\text{AlCl}_3 + \text{H}_2(\text{R})$
17.  $\text{AgCl} + \text{NaNO}_3(\text{M})$
18.  $\text{CH}_4 + \text{Al}(\text{OH})_3(\text{M})$
19.  $\text{Cr}(\text{OH})_3 + \text{NH}_4\text{Cl}(\text{M})$
20.  $\text{PbO} + \text{O}_2(\text{R})$
21.  $\text{NaHSO}_4 + \text{H}_2\text{O}$  or  $\text{Na}_2\text{SO}_4 + \text{H}_2\text{O}(\text{M})$
- 22(a).  $\text{FeSO}_4 + \text{H}_2(\text{R})$
- 22(b) No reaction
23.  $\text{MgCl}_2 + \text{H}_2(\text{R})$
- (24)  $\text{MnS} + \text{HCl}(\text{M})$
25.  $\text{HI}(\text{R})$
26.  $\text{CO}_2 + \text{H}_2\text{O}(\text{R})$
27.  $\text{Ca}(\text{OH})_2 + \text{H}_2(\text{R})$
28.  $\text{AlCl}_3(\text{R})$
29.  $\text{PbO} + \text{NO}_2 + \text{O}_2(\text{R})$
30.  $\text{CaSO}_4 + \text{H}_2\text{O}(\text{M})$
31.  $\text{N}_2 + 2\text{H}_2\text{O}(\text{R})$
32.  $\text{MgO} + \text{S}(\text{R})$
33.  $\text{K}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}(\text{M})$
34.  $\text{PbBr}_2 + \text{H}_2\text{O}(\text{M})$
35.  $\text{Al}(\text{OH})_3 + \text{PH}_3(\text{M})$
36.  $\text{KCl} + \text{O}_2(\text{R})$
37.  $\text{NH}_3 + \text{K}_2\text{SO}_4 + \text{H}_2\text{O}(\text{M})$
38.  $\text{NaAlO}_2 + \text{H}_2\text{O}(\text{M})$
39.  $\text{Hg}_2\text{Cl}_2(\text{or HgCl}) + \text{HNO}_3(\text{M})$
40.  $\text{Fe}_2(\text{SO}_4)_3 + \text{H}_2\text{O}(\text{M})$
41.  $\text{KNO}_2 + \text{O}_2(\text{M})$
42.  $\text{Fe}_3\text{O}_4 + \text{H}_2(\text{R})$
43.  $\text{HgS} + \text{HNO}_3(\text{M})$
44. No reaction
45.  $\text{Al}_2\text{O}_3 + \text{Fe}(\text{R})$
46.  $\text{H}_2\text{SO}_3(\text{M})$
47. No reaction
48.  $\text{CaC}_2\text{O}_4 + \text{NH}_4\text{Cl}(\text{M})$
49.  $\text{MnBr}_2 + \text{Br}_2 + \text{H}_2\text{O}(\text{M})$
50.  $\text{FeCl}_3(\text{R})$
52.  $\text{AlN}(\text{R})$
53.  $\text{Pb} + \text{H}_2\text{O}(\text{R})$
54.  $\text{ZnO}(\text{R})$