

# CHEMISTRY WORK BOOK

## CLASS - XI



State Council of Educational Research and Training  
Govt. of Tripura

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# **CHEMISTRY WORK BOOK**

Class - XI

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রতন লাল নাথ  
মন্ত্রী  
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শিক্ষার প্রকৃত বিকাশের জন্য, শিক্ষাকে যুগোপযোগী করে তোলার জন্য প্রয়োজন শিক্ষাসংক্রান্ত নিরন্তর গবেষণা। প্রয়োজন শিক্ষা সংশ্লিষ্ট সকলকে সময়ের সঙ্গে সঙ্গে প্রশিক্ষিত করা এবং প্রয়োজনীয় শিখন সামগ্রী, পাঠ্যক্রম ও পাঠ্যপুস্তকের বিকাশ সাধন করা। এস সি ই আর টি ত্রিপুরা রাজ্যের শিক্ষার বিকাশে এসব কাজ সুনামের সঙ্গে করে আসছে। শিক্ষার্থীর মানসিক, বৌদ্ধিক ও সামাজিক বিকাশের জন্য এস সি ই আর টি পাঠ্যক্রমকে আরো বিজ্ঞানসম্মত, নান্দনিক এবং কার্যকর করবার কাজ করে চলেছে। করা হচ্ছে সুনির্দিষ্ট পরিকল্পনার অধীনে।

এই পরিকল্পনার আওতায় পাঠ্যক্রম ও পাঠ্যপুস্তকের পাশাপাশি শিশুদের শিখন সক্ষমতা বৃদ্ধির জন্য তৈরি করা হয়েছে ওয়ার্ক বুক বা অনুশীলন পুস্তক। প্রসঙ্গত উল্লেখ্য, ছাত্র-ছাত্রীদের সমস্যার সমাধানকে সহজতর করার লক্ষ্যে এবং তাদের শিখনকে আরো সহজ ও সাবলীল করার জন্য রাজ্য সরকার একটি উদ্যোগ গ্রহণ করেছে, যার নাম 'প্রয়াস'। এই প্রকল্পের অধীনে এস সি ই আর টি এবং জেলা শিক্ষা আধিকারিকরা বিশিষ্ট শিক্ষকদের সহায়তা গ্রহণের মাধ্যমে প্রথম থেকে দ্বাদশ শ্রেণির ছাত্র-ছাত্রীদের জন্য ওয়ার্ক বুকগুলো সুচারুভাবে তৈরি করেছেন। ষষ্ঠ থেকে অষ্টম শ্রেণি পর্যন্ত বিজ্ঞান, গণিত, ইংরেজি, বাংলা ও সমাজবিদ্যার ওয়ার্ক বুক তৈরি হয়েছে। নবম দশম শ্রেণির জন্য হয়েছে গণিত, বিজ্ঞান, সমাজবিদ্যা, ইংরেজি ও বাংলা। একাদশ দ্বাদশ শ্রেণির ছাত্র-ছাত্রীদের জন্য ইংরেজি, বাংলা, হিসাবশাস্ত্র, পদার্থবিদ্যা, রসায়নবিদ্যা, অর্থনীতি এবং গণিত ইত্যাদি বিষয়ের জন্য তৈরি হয়েছে ওয়ার্ক বুক। এইসব ওয়ার্ক বকের সাহায্যে ছাত্র-ছাত্রীরা জ্ঞানমূলক বিভিন্ন কার্য সম্পাদন করতে পারবে এবং তাদের চিন্তা প্রক্রিয়ার যে স্বাভাবিক ছন্দ রয়েছে, তাকে ব্যবহার করে বিভিন্ন সমস্যার সমাধান করতে পারবে। বাংলা ও ইংরেজি উভয় ভাষায় লিখিত এইসব অনুশীলন পুস্তক ছাত্র-ছাত্রীদের মধ্যে বিনামূল্যে বিতরণ করা হবে।

এই উদ্যোগে সকল শিক্ষার্থী অতিশয় উপকৃত হবে। আমার বিশ্বাস, আমাদের সকলের সক্রিয় এবং নিরলস অংশগ্রহণের মাধ্যমে ত্রিপুরার শিক্ষাজগতে একটি নতুন দিগন্তের উন্মেষ ঘটবে। ব্যক্তিগত ভাবে আমি চাই যথাযথ জ্ঞানের সঙ্গে সঙ্গে শিক্ষার্থীর সামগ্রিক বিকাশ ঘটুক এবং তার আলো রাজ্যের প্রতিটি কোণে ছড়িয়ে পড়ুক।

(রতন লাল নাথ)

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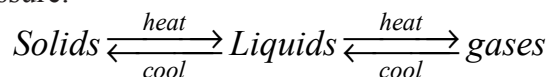
## SOME BASIC CONCEPTS OF CHEMISTRY

### Chapter at a glance :-

- The branch of science which deals with preparations, properties, structure and the changes of matter under different conditions and the laws which govern these changes is called Chemistry. Chemistry is further divided into three main branches i.e. physical, inorganic and organic for better understanding.
- Many years before the development of modern science Indian Scientists were familiar with many scientific matters. They utilised this knowledge in their daily life. Acharya Kanad was the first to introduce the finest indivisible particle i.e. 'paramanu' (atom). He also said that like matter, paramanu (atom) also have variety.
- In National Economy Chemistry plays an important significant role.
  - For the fulfilment of increasing crisis of crops and fruits chemistry provided chemical fertilizers, insecticides, pesticides, fungicides etc.
  - Chemistry has also developed many kinds of life saving drugs. e.g. – Cis-platin and Taxol (used in the treatment of cancer), AZT (Azidothymidine) used in the treatment of AIDS.
- Anything around us that occupies some space and has mass is called matter. eg – book, pen, water, air etc.

Matter can exist in three physical states viz solid, liquid and gas.

- Solids: have definite volume and definite shape.
- Liquids: have definite volume but not the definite shape. They take the shape of the container.
- Gases: Have neither definite volume nor definite shape. They completely occupy the container.
- These three states of matter are inter-convertible by changing the conditions of temperature and pressure.



- At the microscopic or bulk level, matter can be classified as mixtures or pure substances. Pure substances are made up of same components and have fixed composition. e.g.: Copper, Silver etc. Mixtures contain two or more substances at any ratio. It is further classified in two forms, homogeneous and heterogeneous.

The mixture in which the components completely mix with each other and its composition is uniform throughout is called homogeneous mixtures. e.g.: air etc.

Homogenous Mixture has no visible boundaries of separation between the various constituents.

The mixture in which the composition is not uniform throughout and sometimes the different components can be observed is called heterogeneous mixtures. e.g. mixture of salt and sugar, grain and pulses along with some dirt, etc.

Heterogeneous mixture has visible boundaries of separation between the various constituents.

6. Pure substances can further be classified into elements and compounds.

Elements consist of only one type of particles and these particles may be atoms or molecules.

e.g.: Sodium, Copper, Silver, Oxygen, Hydrogen.

Elements cannot be decomposed into simpler substances by ordinary physical or chemical methods.

Compound is a pure substance containing two or more elements combined together in a definite proportion by mass.

e.g.: Water, ammonia, sugar etc.

Compounds can be decomposed into its constituent elements by chemical methods.

7. The properties of a substance can be classified into physical properties and chemical properties.

Physical properties are those properties which can be measured or observed without changing the identity or composition of the substance. e.g.: colour, odour, melting point, boiling point, density etc.

Chemical properties are those properties in which a chemical change in the substances takes place. e.g. - acidity, basicity, combustibility etc.

8. The International System of Units abbreviated as SI units was established in 1960.

The Seven Basic Physical Quantities and their SI units are

Basic physical Quantity	Symbol for Quantity	Name of SI Unit	Symbol for SI Unit
Length	l	Metre	m
Mass	m	Kilogram	Kg
Time	t	Second	S
Electric Current	I	Ampere	A
Thermodynamic Temperature	T	Kelvin	K
Amount of Substances	n	Mole	mol
Luminous Intensity	I <sub>v</sub>	Candela	cd



## 9. Definition of SI Base Units :

**Table 1.2 Definitions of SI Base Units**

<b>Unit of length</b>	<b>metre</b>	The <i>metre</i> is the length of the path travelled by light in vacuum during a time interval of $1/299\,792\,458$ of a second.
<b>Unit of mass</b>	<b>kilogram</b>	The <i>kilogram</i> is the unit of mass; it is equal to the mass of the international prototype of the kilogram.
<b>Unit of time</b>	<b>second</b>	The <i>second</i> is the duration of $9\,192\,631\,770$ periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the caesium-133 atom.
<b>Unit of electric current</b>	<b>ampere</b>	The <i>ampere</i> is that constant current, which if maintained in two straight parallel conductors of infinite length of negligible circular cross-section and placed 1 metre apart in vacuum, would produce between these conductors a force equal to $2 \times 10^{-7}$ newton per metre of length.
<b>Unit of thermodynamic temperature</b>	<b>kelvin</b>	The <i>kelvin</i> , unit of thermodynamic temperature, is the fraction $1/273.16$ of the thermodynamic temperature of the triple* point of water.
<b>Unit of amount of substance</b>	<b>mole</b>	<ol style="list-style-type: none"> <li>1. The <i>mole</i> is the amount of substance of a system, which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12; its symbol is 'mol'.</li> <li>2. When the mole is used, the elementary entities must be specified and these may be atoms, molecules, ions, electrons, other particles, or specified groups of such particles.</li> </ol>
<b>Unit of luminous intensity</b>	<b>candela</b>	The <i>candela</i> is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency $540 \times 10^{12}$ hertz and that has a radiant intensity in that direction of $1/683$ watt per steradian.

10. Chemistry is associated with experimental measurement. Every experimental measurement has some amount of uncertainty associated with it.

**Precision** refers to the closeness of various measurements for the same quantity.

**Accuracy** is the agreement of a particular value to the true value of the result.

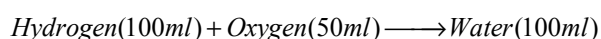
11. **Significant figures** are meaningful digits which are known with certainty.

There are certain rules for determining the number of significant figures. These are –

- All non zero digits are significant.
- Zeros preceding to first non-zero digit are not significant.
- Zeros between two non zero digits are significant.
- Zeros at the end or right of a number are significant provided they are on the right side of the decimal point.
- Exact numbers have an infinite number of significant figures.

12. The combination of elements to form compounds is governed by the following five basic laws.

- Law of conservation of mass:** This law was put forth by Antonie Lavoisier in 1789. According to this law, matter can neither be created nor be destroyed. The total mass of the reactants must be equal to the total mass of the products for a chemical reaction.
- Law of Definite proportions:** This Law was given by Joseph Proust. According to this law- ‘A given compound always contains exactly the same proportion of elements by weight.’
- Law of Multiple proportions:** This Law was proposed by Dalton in 1803. According to this law – ‘If two elements can combine to form more than one compound, the masses of one element that combine with the fixed mass of the other element are in the ratio of small whole numbers.’
- Gay Lussac’s Law of Gaseous volume:** This Law was given by Gay Lussac in 1808. According to this law- When gases combine or are produced in a chemical reaction, they do so in a simple ratio by volume, provided all gases are at same temperature and pressure.



- Avogadro Law:** This law was proposed in 1811 by Amedeo Avogadro. According to this law – ‘equal volumes of gases at the same temperature and pressure should contain equal number of molecules.’ Avogadro was the first to make a distinction between ‘atoms’ and ‘molecules’.

13. In 1808, Dalton published a atomic theory. The main postulates of his atomic theory are –

- Matter consist of indivisible atoms.
- All the atoms of a given element have identical properties including identical mass.  
Atoms of different elements differ in mass.
- Compounds are formed when atoms of different elements combine in a fixed ratio.
- Chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.

Daltons theory can explain all the laws of chemical combinations.

14. The average relative mass of an atom of any element as compared to the mass of an atom of

carbon ( $^{12}\text{C}$ ) taken as 12 - is called **atomic mass**.

$$\text{Atomic mass} = \frac{\text{mass of an atom}}{\frac{1}{12} \text{ mass of an } ^{12}\text{C} \text{ atom}}$$

- 15.** Atomic mass unit (amu) is defined as a mass exactly equal to one-twelfth the mass of one  $^{12}\text{C}$  atom.

$$1 \text{ amu} = 1.66056 \times 10^{-24} \text{ g}$$

Nowadays 'amu' has been replaced by 'u' which is known as **unified mass**.

- 16. Average atomic mass:** Many naturally occurring elements exist as more than one isotope. When we take into account of the existence of these isotopes and their relative abundance (percent occurrence) the average atomic mass of that element can be computed.

For example – the relative abundance (%) of  $^{12}\text{C}$ ,  $^{13}\text{C}$  and  $^{14}\text{C}$  are 98.892%, 1.108% and  $2 \times 10^{-10}\%$  and their atomic mass (amu) are 12, 13.00335 and 14.00317.

Therefore, the average atomic

$$\begin{aligned} \text{mass of Carbon} &= (0.98892) (12\text{u}) + (0.01108) (13.00335\text{u}) + (2 \times 10^{-12}) (14.00317\text{u}) \\ &= 12.011 \text{ u} \end{aligned}$$

- 17. Molecular mass:** It is the sum of atomic masses of the elements present in molecule. For example:

$$\begin{aligned} \text{Molecular mass of CH}_4 \text{ (Methane)} &= (12.011\text{u}) + 4(1.008\text{u}) \\ &= 16.043\text{u} \end{aligned}$$

- 18. Formula Mass:** Some substances do not contain discrete molecules as their constituent units. For such compound instead of molecular mass, formula mass is used.

It is the sum of atomic masses of all atoms in a formula unit.

For Example :

$$\begin{aligned} \text{Formula mass of Sodium Chloride (NaCl)} &= \text{Atomic mass of Na} + \text{Atomic mass of Chlorine} \\ &= 23.0\text{u} + 35.5 \text{ u} \\ &= 58.5\text{u} \end{aligned}$$

- 19.** One mole is the amount of substance that contains as many particles or entities (atom, molecule or ions) as there are atoms in exactly 12g (or 0.012 Kg) of the  $^{12}\text{C}$  isotope.

One mole of any substance always contain the same number of entities.

This number is called Avogadro constant OR Avogadro Number and denoted by  $N_A$ .

The value of  $N_A = 6.022 \times 10^{23}$

- 20.** The mass of one mole of a substance in grams is called its molar mass. For example,

$$\text{Molar mass of water} = 18.02 \text{ g}$$

21. One mole of any gas at STP/NTP occupies 22.44 Lt. i.e volume of any gas at STP is 22.44 Lt.

$$22. \text{ Mass \% of an element} = \frac{\text{mass of element in the compound} \times 100}{\text{molar mass of the compound}}$$

23. The formula representing the simplest whole number ratio of various atoms present in a compound is called **empirical formula**.

The formula representing the exact number of various atoms present in a compound is called **molecular formula**.

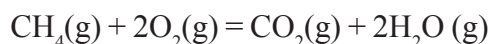
For example:

Empirical formula of di-chloroethane is  $\text{CH}_2\text{Cl}$ . But its molecular formula is  $\text{C}_2\text{H}_4\text{Cl}_2$ . Relation between empirical formula and molecular formula is –

$$\text{Molecular formula} = (\text{Empirical Formula}) \times (n)$$

Where  $n = 1, 2, 3, \dots$  etc whole numbers.

24. A balanced chemical equation has the same number of atoms of each elements of both sides of the equation. The balanced equation for the combustion of methane is –



The coefficient 2 for  $\text{O}_2$  and  $\text{H}_2\text{O}$  are called Stoichiometric coefficients. Similarly the coefficient for  $\text{CH}_4$  and  $\text{CO}_2$  for this reaction is 1.

25. Many a time the reactants are not present in the amount as required by a balanced chemical reaction. Hence, the reactant which gets consumed first, limits the amount of product formed and is therefore called the **limiting reagent**.

26. The concentration of a solution can be expressed in any of the following –

- Mass percent or Weight percent (W/W %)
- Mole fraction
- Molarity
- Molality

$$27. \text{ Mass percent of Solute} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

28. Mole fraction is the ratio of number of moles of a particular component to the total number of moles of the solution/mixture.

29. Molarity is defined as the number of moles of the solute in 1 litre of the solution.

Molarity is the most widely used unit and is denoted by M.

$$\text{Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in litre}}$$

30. Molarity is defined as the number of moles of solute present in 1 kg of solvent. It is denoted by  $m$ .

$$\text{Molality } (m) = \frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$$

A. Choose the correct answer (MCQ) :-

(1 Mark for each question)

1. The finest undivisible particle concept for matter was first launched by -

- a) Acharya Kanad                      b) Chakrapani  
c) Avagadro                              d) Berjeleous

2. The compound Mercury Sulphide was discovered by -

- a) Chakrapani                              b) Nagarjun  
c) Berjeleous                              d) Acharyaa Kanad

3. The compound used in the treatment of AIDS is -

- a) Cisplatin                                b) Taxol  
d) AZT                                        d) None of these

4. The number of basic unit in SI system is -

- a) 5                      b) 6                      c) 7                      d) 8

5. The SI Unit of density is -

- a) Kg-m<sup>-2</sup>            b) g-m<sup>-3</sup>            c) Kg-m<sup>-3</sup>            d) g-m<sup>-2</sup>

6. How many significant number is present in 2.0034 -

- a) 2                      b) 3                      c) 4                      d) 5

7. Which one of the following will have largest number of atoms -

- a) 1g Au(S)    b) 1g Na(S)    c) 1gm Li(S)    d) 1g of Cl<sub>2</sub>(g)

8. The value of  $1 \text{amu}$  is -

- a)  $16.605 \times 10^{-24}$                       b)  $16.605 \times 10^{-23}$   
c)  $16.605 \times 10^{-22}$                       d)  $1.6605 \times 10^{-23}$

9. Which has maximum number of significant figures -  
a)  $3.1 \times 10^6$     b) 0.315    c) 0.000020    d) 0.032
10. The value of  $36^\circ\text{C}$  in fahrenheit scale is -  
a) 97.8 F    b) 98.8 F    c) 96.8 F    d) 96.6 F

**B. Assertion and Reason type Question :- (Each question marks - 1)**

- a) Both assertion and reason are correct and reason is the correct explanation of the assertion  
b) Both assertion and reason are correct but reason is not the correct explanation of the assertion.  
c) Assertion is correct but reason is incorrect  
d) Assertion is incorrect but reason is correct.
1. Assertion (A) : Both 17g of  $\text{NH}_3$  and 16g of  $\text{CH}_4$  contain same number of molecules.  
Reason (R) : Equal moles of two compounds contain same number of molecules.
2. Assertion (A) : 1 mole  $\text{CO}_2$  and  $\text{SO}_2$  have same volume at same temperature and pressure.  
Reason (R) : At N.T.P molar volume of any gas is 22.4 L.
3. Assertion (A) : Molarity is the number of moles of solute in one litre of the solution.  
Reason (R) : Molarity is independent on temperature.
4. Assertion (A) : Molecular formula =  $n \times$  empirical formula.  
Reason (R) : Molecular formula and empirical formula are always same.
5. Assertion (A) : 18g of water contains  $6.022 \times 10^{23}$  molecules  
Reason (R) : 1 mole of any substance contains Avogadro's number of molecules.
6. Assertion (A) : Mole fraction depends on temperature.  
Reason (R) : Mole fraction of any substance can't be greater than one.
7. Assertion (A) : Molality increases with increase in temperature.  
Reason (R) : Molality is expressed in Moles/kg Unit.

**C. Very short answer type question :-**

**(Each question marks- 2)**

1. Cisplatin is used for the treatment of which disease?
2. What is the SI Unit of density?
3. How many particles contains in 1 mole substance?
4. How many electrons are present in 16 gm of  $\text{CH}_4$ ?
5. How many significant figures are present in 0.001420?
6. Round off the number 52.216 upto three significant figure.
7. Express the scientific notation of 0.0052.
8. How much copper can be obtained from 100gm of copper sulphate ( $\text{CuSO}_4$ )?
9. What do you mean by significant figures?
10. What is the SI unit of Mass?
11. What is the SI unit of Luminous intensity?
12. What is the SI unit of Temperature?
13. What will be the mass of one  $^{12}\text{C}$  atom in g?
14. What will be the number of atoms present in 1.4g of  $\text{N}_2$  gas?
15. What is limiting reagent?
16. Define the term 'molarity'?

17. Define the term 'molality'?
18. In which temperature scale negative value is not possible?
19. What is the atomicity of Na?
20. Write the name of the gas which burns in air with a pop sound?

**D. Short answer type questions :-**

**( Each question marks-2)**

1. Write the statements of Law of definite proportion? Explain it with an example.
2. Define Density. Write its SI unit.
3. Calculate the molecular masses of the following -  
a)  $H_2O$                       b)  $CH_4$
4. Calculate the mass percent of different elements present in Sodium Chloride (NaCl).
5. Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% hydrogen by mass.
6. How much copper can be obtained from 100gm of Copper Sulphate ( $CuSO_4$ )?
7. What do you mean by significant figures?
8. If the density of methanol is  $0.793 \text{ Kg L}^{-1}$ , What is the volume needed for making 2.5L of its 0.25m solution?
9. How many significant figures are present in the following?  
a) 0.0017                      b) 8003
10. Round up the following upto three significant figures -  
a) 84. 326                      b) 0.0598



11. If two litres of  $N_2$  is mixed with two litres of  $H_2$  at STP, then what will be the volume of  $NH_3$  formed?
12. What is limiting reagent ? Is it true that reactant which is present in lesser amount in a reaction, consumed first?
13. Define the following with an example -
  - a) Empirical formula
  - b) Molecular formula
14. Define the following with an example -
  - a) Molecular mass
  - b) Formula mass
15. What do you mean by Molar mass? What is the molar mass of water?
16. State Avogadro Law. Write the number of entities present in 1 mole of any substance.
17. How many moles of methane are required to produce 22g  $CO_2(g)$  after combustion?
18. A solution is prepared by adding 2 g of a substance A to 18g of water. Calculate the mass percent of the solute?
19. What mass of Calcium oxide will be obtained by heating 2 mol of  $CaCO_3$ ?
20. Calculate the mass of one  $^{12}C$  atom in g?

**E. Short Answer type Questions :-**

**( Each question marks - 3)**

1. Define Empirical and Molecular formula. What is the relation between Empirical and Molecular formula.
2. A compound contains 4.07% Hydrogen, 24.27% Carbon and 71.65% Chlorine. Its molar mass is 98.96g. What is the empirical and molecular formula of the compound?
3. Calculate the molarity of NaOH solution prepared by dissolving 4g NaOH in enough water to form 250 ml NaOH solution?
4. Calculate the mass of Sodium Acetate ( $\text{CH}_3\text{COONa}$ ) required to make 500ml 0.375m aqueous solution. Molar mass of  $\text{CH}_3\text{COONa}$  is  $82.0245 \text{ mol}^{-1}$ .
5. How many significant figures are present in the following:
  - a) 0.0048
  - b) 234,000
  - c) 8008
6. In three moles of ethane ( $\text{C}_2\text{H}_6$ ), Calculate the following:
  - a) Number of moles of carbon atoms.
  - b) Number of moles of Hydrogen atoms.
  - c) Number of molecules of ethane.
7. Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal is as shown below:
$$1\text{Pa} = 1\text{Nm}^{-2}$$
If mass of air at sea level is  $1034 \text{ g cm}^{-2}$ , calculate the pressure in pascal.
8. If the speed of light is  $3.0 \times 10^8 \text{ ms}^{-1}$ , Calculate the distance covered by light in 2.00 ns.
9. Write the main postulates of Dalton's atomic theory.
10. What do you mean by balanced chemical reaction? Why it is necessary to balance a equation?

**F. Long Answer type Questions :-****(Each questions marks -5)**

1. The following data are obtained when di-nitrogen and di-oxygen react together to form different compounds.

Mass of di-nitrogen	Mass of di-oxygen
i) 14 g	i) 16 g
ii) 14 g	ii) 32 g
iii) 28 g	iii) 32 g
iv) 28 g	iv) 80 g

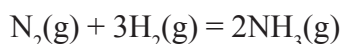
Which law of chemical combination is obeyed by the above experimental data? Give its statement.

2. In a reaction,  $A + B_2 = AB_2$

Identify the limiting reagent, if any, in the following reaction mixtures.

- 300 atoms of A + 200 molecules of B
- 2 mol of A + 3 mols of B
- 100 atoms of A + 100 molecules of B
- 5 mol of A + 2.5 mole of B
- 2.5 mol of A + 5 mole of B

3. Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical reaction,



- Calculate the mass of ammonia produced if  $2.00 \times 10^3$  g of dinitrogen reacts with  $1.00 \times 10^3$  g of dihydrogen.
- Will any of two reactants remain unreacted?
- If yes, which one will remain unreacted and what would be its mass?

4. A welding fuel gas contains carbon and hydrogen gas only. Burning a small sample of it in oxygen gives 3.38g of  $CO_2$ , 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weight 11.6 g. Calculate -

- Empirical formula of the fuel gas.
- Molar mass of the fuel gas.
- Molecular formula of the fuel gas.

5. Calcium carbonate reacts with aqueous HCl to give CaCl<sub>2</sub> and CO<sub>2</sub> according to the following reaction -



What mass of CaCO<sub>3</sub> is required to react completely with 25ml of 0.75 (M) HCl?

6. Chlorine is prepared in laboratory by treating Manganese-di-oxide (MnO<sub>2</sub>) with aqueous HCl according to the following reaction -



How many grams of HCl will be required for the reaction with 5.0 g of MnO<sub>2</sub>?

7. If 4g NaOH dissolves in 36g of water, calculate the mole fraction of each components in the solution. Also determine the molarity of the solution (specific gravity of solution is 1 g-l<sup>-1</sup>).

## Solution

### A. Choose the correct answer (MCQ)

1. a
2. b
3. c
4. c
5. c
6. d
7. c
8. b
9. b
10. d

### B. Assertion and Reason :

1. a
2. a
3. c
4. c
5. a
6. d
7. d

**Chapter at a glance :**

1. According to early Indian and Greek philosophers(400B.C), the continued subdivisions of matter would ultimately yield atoms which would not be further divisible.

The word 'atom' has been derived from the Greek word 'a-tomio' means 'uncuttable' or 'non-divisible'.

The atomic theory of matter was first proposed by John Dalton in 1808 on a firm scientific basis. His theory, called Dalton's Atomic Theory, regarded the atom as the ultimate particle of matter. Towards the end of nineteenth and beginning of twentieth century scientists established that atom can further be divided in sub-atomic particles, i.e. electrons, protons and neutrons.

2. The basic rule regarding the behavior of charged particles is, 'like charges repel each other and unlike charges attract each others'.
3. In mid 1850 s' many scientists mainly Faraday began to study electrical discharges in partially evacuated tubes known as Cathod Ray Discharge tubes.

A cathod ray tube is made of glass containing two thin pieces of metal, called electrodes, sealed in it. The electrical discharges through the gases could be observed only at very low pressures and at very high voltages. When sufficiently high voltage is applied across the electrodes, current starts flowing through a stream of particles moving in the tube from negative electrode (cathode) to the positive electrode (anode). These were called cathode rays or cathode ray particles.

This flow of current from cathode to anode was further checked by making a hole in the anode and coating the tube behind anode with zinc sulphide.

**The results of these experiments are summarised below:**

- i. The cathode rays start from cathode and move towards anode.
  - ii. The cathode rays themselves are not visible but after passing through anode when these rays strike the zinc sulphide coating, a bright spot on the coating is developed (same thing happens in a television set).
  - iii. In the absence of electrical or magnetic field, these rays travel in straight lines.
  - iv. In the presence of electrical or magnetic field cathode rays behaves like negatively charge particles, suggesting that the cathode rays consist of negatively charged particles called electrons.
  - v. The characteristics of cathode rays do not depend upon the material of electrodes and the nature of the gas.
4. J. J. Thomson measured the ratio of electrical charge ( $e^-$ ) to the mass of electron ( $m_e$ ) by using cathode ray tube in 1897 by applying electrical and magnetic field perpendicular to each other as well as to the path of electrons.

He observed that the path of the negatively charged particles is deviated from their path in the presence of electrical or magnetic field.

**The amount of deviation depends upon the following:-**

- i. Greater the magnitude of the charge on the particle, greater is the deflection.
- ii. Lighter the mass of the particle greater is the deflection.
- iii. Increase the voltage across the electrodes or the strength of the magnetic field, greater is the deflection.

Thomson was able to determine the value of  $\frac{e}{m_e}$  as:

$$\frac{e}{m_e} = 1.758820 \times 10^{11} \text{ c kg}^{-1}$$

5. By Millikan's oil drop experiment Millikan determine the charge on the electrons. He found that charge on a electron is  $1.6 \times 10^{19} \text{ c}$ .

At present the value of charge on electron is  $-1.6022 \times 10^{19} \text{ c}$ .

$$\begin{aligned} \text{Mass of electron (} m_e \text{), } \therefore \frac{e}{m_e} &= \frac{1.6022 \times 10^{19} \text{ c}}{1.758820 \times 10^{11} \text{ c kg}^{-1}} \\ &= 9.1094 \times 10^{-31} \text{ kg.} \end{aligned}$$

6. Positively charge particle also known as canal rays was discovered with the help of modified cathode ray tube.

The smallest and lightest positive ion was obtained from hydrogen and was called proton.

**Characteristics of positively charged particles are:-**

- i. It depends upon the nature of gas present in the cathode ray tube.
  - ii. The charge to mass ratio of the particles is found to depend on the gas from which these originate.
  - iii. Some of the positively charged particles carry a multiple of the fundamental unit of electrical charge.
  - iv. The behaviour of these particles in the magnetic or electric field is opposite to that of electron or cathode rays.
7. In 1932, Chadwick discovered electrically neutral particle neutron by bombarding a thin sheet of beryllium by  $\alpha$  – particles.

**8. Properties of Fundamental particles :**

Name	Symbol	Absolute Charge/ C	Relative Charge	Mass/Kg	Mass/u	Approx. Mass /u.
Electron	e	$-1.6022 \times 10^{19}$	-1	$9.10939 \times 10^{-31}$	0.00054	0
Proton	p	$+1.6022 \times 10^{-19}$	+1	$1.67262 \times 10^{-27}$	1.00727	0
Neutron	n	0	0	$1.67493 \times 10^{-27}$	1.00867	0

## 9. Thomson Model of Atom :

J. J. Thomson in 1898 proposed that an atom possesses a spherical shape (radius approximately  $10^{-10}$ ) in which the positive charge is uniformly distributed. The electrons are embedded into it in such a manner as to give the most stable electrostatic arrangement.

This model was also known as plum pudding, raising pudding or watermelon model.

An important feature of this model is that the mass of atom is assumed to be uniformly distributed over the atom.

## 10. Rutherford's Nuclear Model of Atom :

Rutherford performed an experiment by bombarding very thin (100nm) gold foil with  $\alpha$  – particles. The thin gold foil had a circular fluorescent zinc sulphide screen around it. This experiment was named as  $\alpha$  – particles scattering experiment.

He observed that —

- i. Most of the  $\alpha$  – particles passed through the gold foil undeflected.
- ii. a small fraction of the  $\alpha$  – particles was deflected by small angles.
- iii. a very few  $\alpha$  – particles (1 in 20,000) bounced back i.e. were deflected by nearly  $180^\circ$ .

On the basis of these observations, Rutherford drew the following conclusions regarding the structure of atom.

- i. Most of the space in the atom is empty.
- ii. The positively charge particles are concentrated in a very small volume in the atom.
- iii. The volume of the nucleus is negligibly small as compared to the total volume of the atom.

On the basis of above observations and conclusions Rutherford proposed his nuclear Model of Atom. According to this model:

- i. Positively charge and most of the mass of the atom was densely concentrated in extremely small region and is called nucleus by Rutherford.
- ii. The nucleus is surrounded by electrons that move around the nucleus with a very high speed in circular paths, called orbits.
- iii. Electrons and the nucleus are held together by electrostatic forces of attraction.

## 11. Drawbacks of Rutherford Model :

Rutherford nuclear model of an atom is like a small scale solar system with the nucleus playing the role of the sun and the electrons being similar to the lighter planets.

- i. The similarity between the solar system and nuclear model suggests that electron should move around the nucleus in well defined orbit. However when a body is moving in an orbit it must undergoes acceleration. According to the electromagnetic theory of Maxwell, a charged particle when accelerated should emit electromagnetic radiation. The orbit will thus continue to shrink. Calculations shows that an electron will take only  $10^{-8}$  sec time to spiral into the nucleus. But this does not happen.
- ii. Another serious drawback of Rutherford Model is that it says nothing about the electronic structure of atoms i.e. how the electrons are distributed around the nucleus and what are the



energies of these electrons.

## 12. Atomic Number and Mass Numbers :

The total number of protons present in the nucleus of an atom is called the atomic number of that atom. It is denoted as  $Z$ .

$$\begin{aligned}\text{Atomic Number (Z)} &= \text{Number of protons in the nucleus of an atom} \\ &= \text{Number of electrons in a neutral atom.}\end{aligned}$$

The sum of total number of protons and total number of neutrons present in the nucleus of an atom is called mass number of that atom.

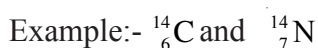
It is denoted as  $A$ .

Mass Number ( $A$ ) = Number of Protons( $Z$ ) + Number of Neutron ( $N$ )

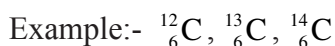
i.e.  $A = Z + n$

## 13. Isobars and Isotopes :

Atoms with same mass number( $A$ ) but different atomic number ( $Z$ ) are called isobars.



Atoms with same atomic number ( $Z$ ) but different mass number ( $A$ ) are called isotopes.



All isotopes of a given element show same chemical behaviour as chemical properties of atoms are controlled by the number of electrons, which are determined by the number of protons in the nucleus.

## 14. Radioactivity :

Henry Becquerel (1852 - 1908) observed that there are certain elements which can emit radiation on their own and named this phenomenon as radioactivity and the elements known as radioactive elements.

This kind of radiation consist of three types of rays i.e.  $\alpha$ ,  $\beta$  and  $\gamma$  rays.

$\alpha$  - rays : when combined with two electrons yielded helium gas. It is expressed as  ${}^4_2\text{He}^{+2}$

$\beta$  -rays : are negatively charge particles similar to electrons.

$\gamma$  rays : are high energy radiation like X-rays and are neutral in nature and do not consist of particles.

The increasing penetrative power of these three radiations is

$$\alpha < \beta \ll \gamma$$

## 15. Bohr's Model for hydrogen atom :

Two developments played a major role in the formulation of Bohr's Model of atom. These were:

- Dual Character of electromagnetic radiation which means that radiation possess both wave like and particle like properties. and
- Experimental results regarding atomic spectra which can be explained only by assuming quantized electronic energy levels in atoms.

### Postulates of Bohr's model for hydrogen :

- i. The electrons in the hydrogen atom can move around the nucleus in a circular path at fixed radius and energy. These paths are called orbit. These orbits are arranged concentrically around the nucleus.
- ii. The energy of an electron in the orbit does not change with time. However, the electron can move from one orbit to another, by loosing or gaining energies. The energy change does not take place in a continuous manner.
- iii. The frequency of radiation absorbed or emitted when transition occurs between two stationary states, that differ in energy by  $\Delta E$  is given by,

$$\gamma = \frac{\Delta E}{h} = \frac{E_2 - E_1}{h}$$

where  $E_1$  and  $E_2$  are the energies of the lower and higher allowed energy state respectively. This expression is known as Bohr's frequency rule.

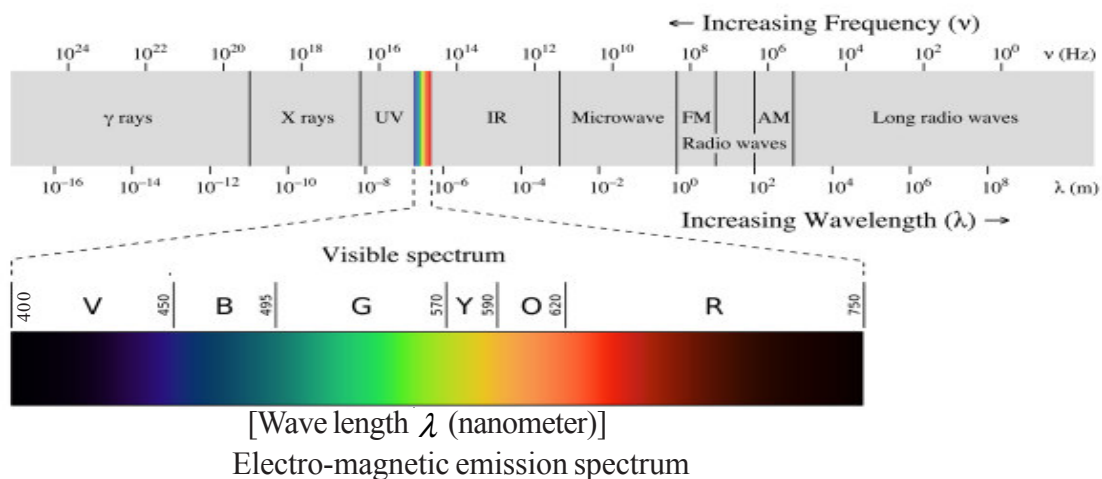
- iv. The angular momentum of an electron (in a given stationery state) can be expressed by the following equation,

$$m_e v r = n \cdot \frac{h}{2\pi}, \quad n = 1, 2, 3 \dots \text{etc.}$$

Thus an electron can move only in those orbits for which its angular momentum is integral multiple of  $\frac{h}{2\pi}$ . i.e. certain fixed orbit are allowed for electron to move.

## 16. Properties of electromagnetic radiation :

- i. The oscillating electric and magnetic fields produced by oscillating charged particles are perpendicular to each other and both are perpendicular to the direction of propagation of the wave.
- ii. Electromagnetic waves do not require medium and can move in vaccum.
- iii. There are many types of electromagnetic radiations, which differ from one another in wavelength or frequency. These constitute what is called a electromagnetic spectrum.
- iv. Different kinds of units are used to represent electromagnetic spectrum.



### 17. Frequency, wavelength and wave number :

Frequency is defined as the number of waves that pass a given point in one second. It is denoted by  $\gamma$  and its SI unit is hertz (Hz).

The SI unit of wavelength ( $\lambda$ ) is meter ( $m$ ).

Wave number ( $\tau$ ) is defined as the number of wavelengths per unit length.

Its SI unit is  $m^{-1}$ .

However commonly used unit is  $cm^{-1}$ .

The frequency ( $\nu$ ), wavelength ( $\lambda$ ) and velocity of light ( $c$ ) are related by the following equation:

$$c = \gamma \lambda$$

### 18. Black body radiation :

The ideal body which emits and absorbs all frequencies is called a black body and the radiation emitted by such a body is called black body radiation.

The exact frequency distribution of the emitted radiation from a black body depends only on its temperature.

First concrete explanation for the phenomenon of the black body radiation was given by Max Planck in 1900.

Planck suggested that atoms and molecules could emit or absorb energy only in discrete quantities and not in a continuous manner.

The smallest quantity of energy that can be emitted or absorbed in the form of electromagnetic radiation is called quantum.

The energy ( $E$ ) of a quantum of radiation is proportional to its frequency ( $\nu$ ) and is expressed by the following equation.

$$E \propto \nu \quad \text{or} \quad E = h\nu$$

Here  $h$  is Planck constant and has the value  $6.626 \times 10^{-34} \text{ Js}$

### 19. Photoelectric Effect :

In 1887, H. Hertz performed a very interesting experiment in which electrons (or electric current) were ejected when certain metals (for example potassium, rubidium, caesium etc) were exposed to a beam of light. This phenomenon is called photoelectric effect.

**Observations regarding photoelectric effect are —**

- i. There is no time lag between the striking of light beam and the ejection of electrons from the metal surface.
- ii. The number of electrons ejected is proportional to the intensity or brightness of light.
- iii. For each metal, there is a characteristic minimum frequency ( $\gamma_0$ ) below which photoelectric effect is not observed. This is known as threshold frequency ( $\gamma_0$ ).
- iv. The kinetic energy of ejected electrons increases with the increase of frequency of light used.
- v. The number of electrons ejected should depend on the brightness of the light but the kinetic

energy of the ejected electrons does not depend on the brightness of the light.

- vi. Kinetic energy of the ejected electron is proportional to the frequency of the electromagnetic radiation.
- vii. Kinetic energy of the ejected electron is given by the following equation,

$$h\nu = h\nu_0 + \frac{1}{2}m_e v^2$$

Here,  $m_e$  = mass of electron and

$v$  = velocity associated with ejected electron

## 20. Diffraction and interference :

Diffraction is the bending of wave around a obstacle.

Interference is the combination of two waves of the same or different frequencies to give a wave whose distribution at each point in space is the algebraic or vector sum of disturbances at that point resulting from each interfering wave.

## 21. Values of work function ( $W_0$ ) for a few metals.

Metal	Li	Na	K	Mg	Cu	Ag
$W_0/\text{ev}$	2.42	2.3	2.25	3.7	4.8	4.3

## 22. Spectrum :

The colour band that can be found by splitting a beam of light is called spectrum.

When a ray of white light is passed through a prism it spread out into a series of colored bands. The spectrum of white light ranges from violet at  $7.50 \times 10^{14}$  Hz to red at  $4 \times 10^{14}$  Hz.

This type of spectrum is called continuous spectrum.

The spectrum of radiation emitted by a substance that has absorbed energy is called an emission spectrum.

Emission spectrum consist of bright colored lines separated by dark spaces.

Emission spectrum is line spectrum or discontinuous spectrum. In case of emission spectrum  $n_f > n_i$  and energy is released.

An absorption spectrum is like the photographic negative of an emission spectra.

Absorption spectrum is obtained when light is first passed through the substance and the transmitted light is analysed with a spectroscope.

Absorption Spectrum consist of dark lines. It is a continuous spectrum. In case of absorption spectrum  $n_f > n_i$  and energy is absorbed.

## 23. Line spectrum of hydrogen :

Of all the elements hydrogen atom has the simplest line spectrum.

The hydrogen spectrum consist of several series of lines.

Balmer showed in 1885 that if spectral lines are expressed in terms of wave number ( $\bar{\nu}$ ), the visible lines of hydrogen spectrum obey the following formula :

$$\bar{\nu} = 109,677 \left( \frac{1}{2^2} - \frac{1}{n^2} \right)$$

where,  $n$  is an integer equal to or greater than 3.

The series of lines described by this formula are called Balmer Series.

Rydberg noted that all series of lines in hydrogen spectrum could be described by the following expression,

$$\bar{\nu} = 109,677 \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \text{cm}^{-1}$$

Where,  $n_1 = 1, 2, 3, \dots$

$n_2 = (n_1+1), (n_1+2), \dots$

The value  $109,677 \text{ cm}^{-1}$  is called the Rydberg constant for hydrogen.

#### The Spectral lines for atomic hydrogen :

Series	$n_1$	$n_2$	Spectral Region
Lyman	1	2, 3 . . .	Ultraviolet
Balmer	2	3, 4 . . .	Visible
Paschen	3	4, 5 . . .	Infrared
Brackett	4	5, 6 . . .	Infrared
Pfund	5	6, 7 . . .	Infrared

24. Bohr's theory can also be applied to the ions containing only one electron. For example  $\text{He}^+$ ,  $\text{Li}^{2+}$ ,  $\text{Be}^{3+}$  . . . etc.

The energies of stationary states associated with these kinds of ions are given by the expression,

$$E_n = -2.18 \times 10^{-18} \left( \frac{Z^2}{n^2} \right) \text{J}$$

and radii by the expression,

$$r_n = \frac{52.9(n^2)}{Z} \text{pm}$$

where  $Z$  is the atomic number.

25. The most important property associated with the electron is the energy of its stationary state. It is given by the expression,

$$E_n = -R_H \left( \frac{1}{n^2} \right)$$

Where  $R_H$  is called Rydberg constant and its value is  $-2.18 \times 10^{-18} \text{ J}$

The energy of the lowest state i.e. ground state is,

$$E_1 = 2.18 \times 10^{-18} \left( \frac{1}{1^2} \right) = -2.18 \times 10^{-18} \text{ J}.$$

$$\text{Similarly } E_2 = -0.545 \times 10^{-18} \text{ J}.$$

$$\Delta E = R_H \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = 2.18 \times 10^{-18} \text{ J} \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

Here i stand for initial orbit and f stand for final orbit.

26. The frequency ( $\gamma$ ) associated with the absorption or emission of photon can be evaluated by using the following equation,

$$\Delta E = h\gamma$$

$$\begin{aligned} \therefore \gamma &= \frac{\Delta E}{h} = \frac{R_H}{h} \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \\ &= \frac{2.18 \times 10^{-18} \text{ J}}{6.626 \times 10^{-34} \text{ Js}} \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \\ &= 3.29 \times 10^{15} \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \end{aligned}$$

### 27. Dual behaviour of matter :

De Broglie in 1924 proposed that matter, like radiation, should also exhibit dual behaviour. According to De Broglie every object in motion has a wave character. Because of the large masses wave properties of ordinary objects can't be detected. But the wave-length associated with electrons and other sub-atomic particles can however be detected experimentally by the following equation,

$$\lambda = \frac{h}{mv} = \frac{h}{p}$$

Where  $m$  = mass of particle,  $v$  = velocity of the particle and  $p$  = momentum of the particle.

### 29. Heisenberg's uncertainly principle :

According to Heisenberg's uncertainly Principle, it is impossible to determine simultaneously the exact position and exact momentum (or velocity) of an electron (or any other moving particle).

Mathematically it can be expressed by the following equation,

$$\Delta x \times \Delta p_x \geq \frac{h}{4\pi}$$

or,

$$\Delta x \times \Delta v_x \geq \frac{h}{4\pi m} \quad [ \because \Delta p_x = \Delta(mv_x) ]$$

Where  $\Delta x$  is the uncertainty in position and  $p_x$  (or  $\Delta v_x$ ) is the uncertainty in momentum (or velocity) of the particle.

**Significance of uncertainty principle :**

- i. It rules out existence of definite paths and trajectories of electrons and other similar particles.
- ii. This principle is significant only for motion of microscopic objects and is negligible for that of macroscopic object.

**29. Quantum mechanical model of atom :**

The branch of Science that takes into account the dual behaviour of matter is called quantum mechanics.

Quantum mechanics deal with the study of the motions of the microscopic objects that have both observable wave like and particle like properties.

Quantum mechanics was developed independently in 1926 by Heisenberg and Schrodinger. Schrodinger equation is written as,

$$H\psi = E\psi, \text{ Where } H \text{ is a mathematical operator called Hamiltonian.}$$

**Important features of quantum mechanical model of atom :**

- i. The energy of electrons in atoms is quantized.
- ii. The existence of quantized electronic energy levels is a direct result of the wave like properties of electrons and are allowed solutions of Schrodinger equation.
- iii. Both the exact position and velocity of an electron in an atom cannot be determined simultaneously
- iv. An atomic orbital is the wave function  $\psi$  for an electron in an atom.
- v. The probability of finding an electron at a point within an atom is proportional to the square of the orbital wave function i.e.  $|\psi|^2$  at that point.

$|\psi|^2$  is known as probability density and is always positive.

From the value of  $|\psi|^2$  at different points within an atom, it is possible to predict the region around the nucleus where electrons will most probably be found.

**30. Orbitals and Quantum Numbers :**

An orbital is the region in space around the nucleus where the probability of finding an electron is maximum.

An orbital of smaller size means there is more chance of finding the electron near the nucleus.

Atomic orbitals are precisely distinguished by three quantum number and these are principal quantum number ( $n$ ), azimuthal quantum number ( $l$ ) and magnetic orbital quantum number ( $m$ )

**31. Principal quantum number (n) :**

It determines the size and to large extent the energy of the orbital. The principal quantum number also identifies the shell.

With the increase in the value of ' $n$ ' the number of allowed orbitals increases and are gave by ' $n^2$ '.

$n$	=	1	2	3	4	. . .
shell	=	K	L	M	N	. . .

Size of an orbital increases with increase of  $n$  value.

### 32. Azimuthal quantum numbers ( $l$ ):

It is also known as orbital angular momentum or subsidiary quantum number. It defines the three dimensional shape of the orbital.

For a given value of  $n$ , the possible values of  $l$  are  $0, 1, 2, \dots, (n - 1)$

The number of subshells in a principal shell is equal to the value of  $n$ . Each subshell is assigned an azimuthal quantum number ( $l$ ):

Value for $l$ :	0	1	2	3	4	5	. . .
Notation for Subshell :	$s$	$p$	$d$	$f$	$g$	$h$	. . .

### 33. Magnetic orbital quantum number ( $m_l$ ):

It gives information about the special orientation of the orbital with respect to standard set of co-ordinate axis.

For any sub-shell (defined by  $l$  value)  $2l + 1$  values of  $m_l$  are possible

$$m_l = -l, -(l - 1), -(l - 2) \dots 0 \dots (l - 2), (l - 1), l$$

The following chart gives the relation between the sub-shell and the number of orbitals associated with it.

Value of $l$	0	1	2	3	4	5
Sub-shell notation	$s$	$p$	$d$	$f$	$g$	$h$
Number of orbitals	1	3	5	7	9	11

### 34. Spin quantum number ( $m_s$ ):

Spin quantum number refers to orientation of spin of the electron. An orbital cannot hold more than two electrons and these two electrons should have opposite spins.  $+\frac{1}{2}$  or  $-\frac{1}{2}$ .

### 35. Energies of orbitals :

Energies of orbitals increases as follows :

$$1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4f < \dots$$

The orbitals having the same energy are called degenerate.

The main reason for having different energies of the sub-shells is the mutual repulsion among the electrons in a multi electron atoms.

Due to presence of electrons in the inner shells, the electron in the outer shell will not experience the full positive charge on the nucleus. This is known as shielding of the outer shell electrons from the nucleus by the inner shell electrons.

The increasing order of shielding capacity of the orbitals is

$$s > p > d > f$$



The lower the value of  $(n + 1)$  for an orbital, the lower is its energy. If two orbitals have the same value of  $(n + 1)$ , the orbital with lower value of  $n$  will have the lower energy.

Energies of the orbitals in the same sub shell decreases with increase in the atomic number ( $Z_{\text{eff}}$ ). For example,  $E_{2s}(\text{H}) > E_{2s}(\text{Li}) > E_{2s}(\text{Na}) > E_{2s}(\text{K})$

### 36. Aufbau Principle :

The filling of electrons into the orbitals of different atoms takes place according to the Aufbau principle.

The principle states - “ In the ground state of the atoms, the orbitals are filled in order of increasing energies.”

The increasing order of energies of orbitals are follows:

$$1S < 2S < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 4f < 5d < 6p < 7s \dots$$

### 37. Pauli's Exclusion Principle :

Its states that - ‘Only two electrons may exist in the same orbital and these electrons must have opposite spin’.

The maximum no. of electrons in the shell with principal quantum number  $n$  is equal to  $2n^2$ .

### 38. Hund's rule of Maximum Multiplicity :

It states - pairing of electrons in the orbitals belonging to the same subshell does not take place until each orbital belonging to that subshell has got one electron each i.e. it is singly occupied.

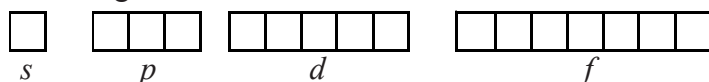
### 39. Electronic configuration of atoms :

The distribution of electrons into orbitals of an atom is called its electronic configuration.

The electronic configuration of different atoms can be represented in two ways,

i)  $s^a, p^b, d^c, \dots$  notation

ii) Orbital diagram



### 40. Stability of completely filled and half filled subshells :

The completely filled and completely half filled subshells attain an extra stability due to the following reasons.

#### i. Symmetrical distribution electrons :

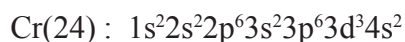
It is well known that symmetry leads to stability. The completely filled or half filled subshells have symmetrical distribution of electrons in them and are therefore more stable.

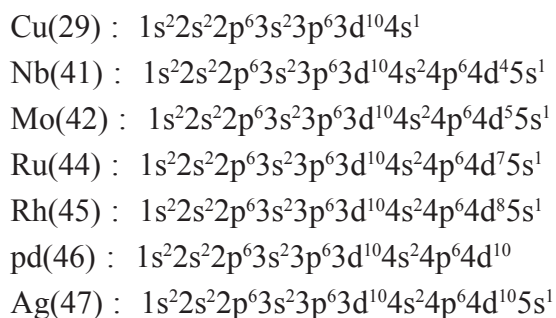
#### ii. Exchange Energy :

Electrons in the degenerate orbitals of a subshell tend to exchange their positions and the energy released due to this exchange is called exchanged energy.

The No. of exchanges that can take place is maximum when the subshell is either half filled or completely filled. Larger the exchange energy greater is the stability.

### 41. Elements with exceptional electronic configurations :





**A. Choose the correct answer (MCQ)**

**[Each question mark 1]**

- Mass of an electron is —
 

a) $9.1094 \times 10^{-31} \text{ kg}$ .	b) $1.758820 \times 10^{11} \text{ kg}$
c) $1.6022 \times 10^{19} \text{ kg}$	d) $1.67262 \times 10^{-27} \text{ kg}$
- Mass of a proton is
 

a) $9.10939 \times 10^{-31} \text{ kg}$	b) $1.67262 \times 10^{-27} \text{ kg}$
c) $1.67493 \times 10^{-27} \text{ kg}$	d) $1.75882 \times 10^{12} \text{ kg}$
- Neutron was discovered by
 

a) Chadwick	b) Millikan
c) J. J. Thomson	d) Rutherford
- $^{14}_6\text{C}$  and  $^{14}_7\text{N}$  are -
 

a) Isotope	b) Isobar
c) Isotone	d) None of these
- Whose penetrative power is maximum -
 

a) $\alpha$ particle	b) $\beta$ particle
c) $\gamma$ particle	d) photon
- The energy of an electron in the orbit -
 

a) does not change with time	b) increase with time
c) decreases with time	d) none of these

7. The SI Unit of wavelength is -  
 a) cm                      b) m                      c) nm                      d) nano centimeter
8. The SI Unit of wavenumber is -  
 a) cm                      b)  $\text{cm}^{-1}$                       c) m                      d)  $\text{m}^{-1}$
9. The frequency of the emitted radiation from a black body depends on -  
 a) mass of the body                      b) Temperature  
 c) Pressure                      d) None of these
10. The increasing order of shielding capacity of the orbital is -  
 a)  $s < p < d < f$                       b)  $s > p > d > f$   
 c)  $s > p = d > f$                       d)  $s < p < d = f$
11. In Bohr's theory, the radius  $r$  of orbit is proportional to -  
 a)  $n$                       b)  $n^{-1}$                       c)  $n^2$                       d)  $n^{-2}$
12. Presence of three unpaired electron in Nitrogen atom can be explained by -  
 a) uncertainty principle                      b) Aufbau's rule  
 c) Pauli's rule                      d) Hund's rule
13.  $n = 2, l = 0, m_l = 0, m_s = +\frac{1}{2}$ , this set of quantum number for outermost electron associated with the element -  
 a) Hydrogen                      b) Lithium  
 c) Beryllium                      d) Magnesium
14. The number of radial nodes possible of 3d orbital is -  
 a) 1                      b) 0                      c) 2                      d) 3
15. A subshell with  $n=5, l = 2$  can accommodate a maximum of -  
 a) 10 electrons                      b) 12 electrons  
 c) 36 electrons                      d) 18 electrons

16. Which of the following is not possible -

- a)  $n = 3, l = 3, m = 2$                       b)  $n = 3, l = 2, m = 0$   
c)  $n = 1, l = 0, m = 0$                       d)  $n = 4, l = 3, m = -3$

17. For 4s orbital, the magnetic quantum number has the value -

- a) -1                      b) +1                      c) 2                      d) 0

18. How many electrons in Argon have  $m = 0$  ?

- a) 10                      b) 12                      c) 8                      d) 6

19. Which of the following electronic configuration is not possible according to Hund's rule?

- a)  $1s^2 2s^2$                       b)  $1s^2 2s^1$                       c)  $1s^2 2s^2 2p_x^2$                       d)  $1s^2 2s^2 2p_x^1 2p_y^1$

20. Which of the following sets of the quantum number is correct for an electron in 4f orbital?

- a)  $n = 4, l = 3, m = +1, s = +\frac{1}{2}$                       b)  $n = 4, l = 4, m = -4, s = -\frac{1}{2}$   
c)  $n = 4, l = 3, m = +4, s = +\frac{1}{2}$                       d)  $n = 3, l = 2, m = -2, s = -\frac{1}{2}$

21. The radius of the first Bohr orbit of Hydrogen atom is  $0.59 \text{ \AA}$ . The radius of the third orbit of  $\text{He}^+$  will be -

- a)  $8.46 \text{ \AA}$                       b)  $0.705 \text{ \AA}$                       c)  $2.38 \text{ \AA}$                       d)  $1.59 \text{ \AA}$

22. If uncertainty in position and momentum are equal, then uncertainty in velocity is -

- a)  $\frac{1}{2m} \sqrt{\frac{h}{\pi}}$                       b)  $\sqrt{\frac{h}{2\pi}}$                       c)  $\frac{1}{m} \sqrt{\frac{h}{\pi}}$                       d)  $\sqrt{\frac{h}{\pi}}$

23. The correct set of four quantum numbers of outermost electron of potassium ( $Z=19$ ) is -

- a)  $3, 1, 0, \frac{1}{2}$                       b)  $4, 0, 0, \frac{1}{2}$                       c)  $3, 0, 0, \frac{1}{2}$                       d)  $4, 1, 0, \frac{1}{2}$

24. Maximum number of electrons in a subshell of an atom is determined by -

- a)  $2l + 1$                       b)  $4l - 2$                       c)  $2n^2$                       d)  $4l + 2$

25. Which one of the following ion has electronic configuration [Ar] 3d<sup>6</sup>?

- a) Fe<sup>3+</sup>                      b) Co<sup>3+</sup>                      c) Ni<sup>3+</sup>                      d) Mn<sup>3+</sup>

26. If the energies of two photons are in the ratio of 3 : 2, their wavelength will be in the ratio of -

- a) 9 : 4                      b) 2 : 3                      c) 1 : 2                      d) 3 : 2

27. Which of the following statement is correct ?

- a) Size of an orbital increases with increase of n value.  
 b) Size of an orbital decreases with increase of n value.  
 c) Size of an orbital increases with increase of l value.  
 d) Size of an orbital decreases with increase of l value.

28. Which of the following statement is incorrect?

- a) The lower the value of (n + l) for an orbital, the lower is its energy.  
 b) Energies of the orbitals in the same subshell decreases with increase in the atomic number.  
 c) The orbitals having the same energy are called degenerate  
 d) None of these.

29. Which of the following does not represents the mathematical expression for the Heisenberg uncertainty principle?

- a)  $\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$                       b)  $\Delta x \cdot \Delta v \geq \frac{h}{4\pi m}$   
 c)  $\Delta E \cdot \Delta t \geq \frac{h}{4\pi}$                       d)  $\Delta E \cdot \Delta x \geq \frac{h}{4\pi}$

30. Who proposed dual nature of electron?

- a) De - Broglie                      b) Einstein  
 c) Bohr                      d) Heisenburg

31. The orbital diagram in which Aufbau principle is violated is -

- a)  $\uparrow\downarrow$     $\uparrow$   $\uparrow$     $\square$                       b)  $\uparrow\downarrow$     $\uparrow$   $\uparrow$   $\uparrow$   
 c)  $\uparrow$     $\uparrow\downarrow$     $\square$     $\square$                       d)  $\uparrow\downarrow$     $\uparrow\downarrow$   $\uparrow\downarrow$   $\uparrow$

32. The electronic configuration of  $\text{Cu}^{2+}$  ion is -

- a)  $[\text{Ar}]3d^84s^1$                       b)  $[\text{Ar}]3d^94s^0$   
c)  $[\text{Ar}] 3d^74s^2$                       d)  $[\text{Ar}] 3d^84s^0$

**B. Assertion -Reason type questions :**

**[Each question mark 1]**

The Question given below consist of an Assertion and the Reason. Use the following key to choose the appropriate answer :

- a) If both assertion and reason are correct and reason is the correct explanation of the assertion.  
b) If both assertion and reason are correct, but reason is not the correct explanation of the assertion.  
c) If assertion is correct but reason is incorrect.  
d) If assertion is incorrect but reason is correct

- Assertion : Hydrogen has one electron in its orbit but it produces several spectral lines.  
Reason : There are many excited energy levels available.
- Assertion : The configuration of carbon can't be  $1s^2 2s^2 2p_x^2$ .  
Reason : According to Pauli's exclusion principle an orbital can have maximum of two electrons.
- Assertion : Black body is an ideal body that emits and absorbs radiation all frequencies.  
Reason : The frequency of radiation emitted by a body goes from a lower frequency to higher frequency with an increase in temperature.
- Assertion : All isotopes of a given element show the same type of chemical behaviour.  
Reason : The chemical property of an atom are controlled by the number of electrons in the atom.
- Assertion : The 19th electron in potassium atom enters into 4s orbital and not in the 3d orbital.  
Reason :  $(n + 1)$  rule is followed for determining the orbital of lowest energy state.
- Assertion : Absorption spectrum consist of dark lines.  
Reason : It is an continuous spectrum.
- Assertion : Emission spectrum consist of dark lines.  
Reason : Emission spectrum is a line spectrum.

8. Assertion : According to De-Broglie wave properties of ordinary objects can't be detected.  
Reason : Because of the large volume of ordinary object.
9. Assertion : The energy of electron in atoms is quantized.  
Reason : Both the exact position and velocity of an electron in an atom cannot be determined simultaneously.
10. Assertion : Cathode rays are not visible.  
Reason : Cathode ray is the flow of electron from cathode to anode in a discharge tube.

**C. Very short answer type questions :**

**( Each question mark 1)**

- Write the complete symbol of the atom with atomic number 17 and atomic mass 35.
- How many protons, electrons and neutrons are present in  ${}_{12}^{25}\text{Mg}$  nuclei.
- Calculate the number of electrons which will together weigh one gram.
- How many electrons in an atom may have the following quantum number?

$$n = 4, m_s = -\frac{1}{2}$$

- An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol.
- An ion with mass number 37 possesses one unit of negative charge. If the ion contains 11.1% more neutrons than the electrons, find the symbol of the ion?
- Among 2s and 3s orbital which orbital will experience the larger effective nuclear charge?
- How many unpaired electrons are present in Fe?
- How many sub-shell are associated with  $n = 4$ ?

10. Which of the following are isoelectronic species :  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Mg}^{2+}$ ,  $\text{Ca}^{2+}$ ,  $\text{S}^{2-}$ , Ar.
11. How many electrons are present in  $\text{O}_2^+$  ion ?
12. Calculate the total number of electrons present in one mole of methane?
13. State the basic rule regarding the behaviour of charge particles.
14. Write the value of  $\frac{e}{m_e}$ .
15. What is the mass of one electron ( $m_e$ )?
16. Write the name of element from which the smallest and lightest positive ion was obtained?
17. Name the sub-atomic particle which was discovered by Chadwick in 1932 ?
18. According to Rutherford's model of atom by which force electrons and nucleus are held together?
19. Define the term isobar?
20. Define the term wave-number?
21. Write the relation between frequency, wave-length and velocity of light?
22. Define the term blackbody?
23. What is the value of Plank's constant?
24. What do you mean by 'interference'?
25. Write the expression for the Balmer series of Hydrogen spectrum.
26. What do you mean by 'threshold frequency'?



27. Define the term 'photo-electric effect'?
28. Write the value of 'Rydberg constant'?
29. Write the expression of Rydberg equation for hydrogen like ions.
30. A certain particle carries  $2.5 \times 10^{-16} \text{C}$  of static charge. Calculate the number of electrons present in it.
31. The diameter of Zinc atom is  $2.6 \text{ \AA}$ . Calculate the radius of Zinc atom in pm.
32. Which of the following has lowest frequency?  
x-rays,  $\gamma$ -rays, microwaves.
33. Which series of hydrogen spectrum lies in the visible region?
34. Which transitions between orbits corresponds to second line in the Balmer series?
35. Write the statement of Pauli's exclusion principle.
36. What sub-shells are possible in  $n = 3$  energy level ?
37. How many orbitals are possible in  $n = 3$  energy level?
38. Using the s, p and d notations, describe the orbital with the following quantum number -  
a)  $n = 1, l = 0$                       b)  $n = 3, l = 2$                       c)  $n = 3, l = 1$
39. Write the electronic configuration of the following ions -  
a)  $\text{O}^{2-}$                                       b)  $\text{F}^-$
40. Explain the meaning of the symbol  $4d^6$ ?
41. What physical meaning is attributed to the square of the absolute value of wave function  $|\psi|^2$ ?

42. Which of the following orbitals has the highest energy?

5p, 5d, 5f, 6s, 6p

43. Which of the following orbitals are degenerate?

$3d_{xy}$ ,  $4d_{xy}$ ,  $3d_{z^2}$ ,  $3d_{yz}$ ,  $4d_{yz}$ ,  $4d_{z^2}$

44. If (n+l) value of two orbitals are same then which orbital has higher energy?

45. Name the phosphorescent material that was used by Faraday in cathod ray discharged tube?

**D. Short answer type questions :**

**(Each question marks = 2)**

1. Write the main characteristics of cathode rays.
2. Write the main drawback's of Rutherford model of atom.
3. Define atomic number and mass number.
4. What is isobars and isotopes? Give one example of each.
5. What is radioactivity? Write the symbol of  $\alpha$  - particle.
6. Write two properties of electromagnetic radiation?
7. Define the term wave number? What is its SI Unit?
8. Write the equation which expresses the relation between frequency of light, wave length and wave number? What is the conventional unit of wave number?
9. Define the term diffraction and interference?
10. What is absorption and emission spectrum? Write one difference between them.
11. Define the term orbit and orbital?

12. Write difference between orbit and orbital.
13. Write the statement for Heisenberg's uncertainty principle. What is the significance of uncertainty principal?
14. Why completely filled and half filled subshells are more stable?
15. Write the electronic configuration of Cu(29) and Ag(47).
16. The number of electrons, protons and neutrons in a species are equal to 18, 16 and 16 respectively. Assign proper symbol to the species.
17. Calculate the mass and charge of one mole of electrons.
18. Yellow light emitted from a sodium lamp has a wave-length  $\lambda$  of 580 nm. Calculate the frequency and wave number of this light.
19. Wave length range of the visible spectrum extends from violet (400nm) to red (750nm). Express these wavelengths in frequencies (Hz).
20. A 100 watt bulb emits electromagnetic light of wavelength 400nm. Calculate the number of photons emitted per second by the bulb.
21. In the Rydberg equation a spectral line corresponds to  $n_1 = 3$  and  $n_2 = 5$ . Calculate the wavelength and frequency of the spectral line.
22. What transition in the hydrogen spectrum would have the same wavelength as the Balmar Transition,  $n = 4$  to  $n = 2$  of  $\text{He}^{2+}$  spectrum?
23. Calculate the energy associated with the first orbit of  $\text{He}^+$ . What is the radius of this orbit?
24. Between 2s and 3s subshell whose energy is more? and why?
25. Between 3p and 3d subshell whose energy is more? and why?

26. Calculate the wave number for the longest wavelength transition in Balmer series of atomic hydrogen.
27. Symbols  ${}^{79}_{35}\text{Br}$  and  ${}^{79}\text{Br}$  can be written, whereas symbols  ${}^{79}_{35}\text{Br}$  and  ${}_{35}\text{Br}$  are not acceptable. Justify.
28. Energy of an electron in H atom in ground state 13.6 eV. What is the value in first excited state?
29. Which transitions between Bohr orbits corresponds to
- Second line in the Balmer series and
  - First line in Brackett series of hydrogen spectrum.
30. Give the number of electrons in the species  $\text{H}_2^+$ ,  $\text{H}_2$  and  $\text{O}_2^+$

**E. Long answer type questions :**

**(Each Question marks 3)**

- Calculate the wavelength, frequency and wave number of light wave whose period is  $2.0 \times 10^{-10} \text{ s}$ .
- Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency ( $\gamma_0$ ) and work function ( $w_0$ ) of the metal.
- What is the maximum number of emission lines when the excited electron of a H atom in  $n = 6$  drops to the ground state?
- An atomic orbital has  $n = 3$ . What are the possible values of  $l$  and  $m_l$ ?
  - List the quantum numbers ( $m_l$  and  $l$ ) of electrons for 3d orbital.
  - Which of the following orbitals are possible?  
1p, 2s, 2p and 3f
- Which atoms are indicated by the following configurations?  
 (i)  $[\text{He}] 2s^1$                       (ii)  $[\text{Ne}] 3s^2 3p^3$
  - What is the lowest value of  $n$  that allows g orbitals to exist?
- How many electrons in an atom may have the following quantum numbers?  
 (i)  $n = 4, m_s = -\frac{1}{2}$                       (ii)  $n = 3, l = 0$
  - Among the following pair of orbitals which orbital will experience the larger effective nuclear charge?  
 (i) 4d and 4f                      (ii) 3d and 3p

7. a) What is photoelectric effect?  
 b) State Aufbau principle. Write Electronic configuration of element with atomic number 24.
8. The ejection of photo electron from the silver metal in the photoelectric effect experiment can be stopped by applying the voltage of 0.35v when the radiation 256.7 nm is used. Calculate the work function for silver metal.
9. If the velocity of electrons in Bohr's first orbit is  $2.19 \times 10^6 \text{ ms}^{-1}$ . Calculate the De Broglie wavelength associated with it?
10. a) Show the distribution of electrons in oxygen atom using orbital diagram.  
 b) Calculate the total number of angular nodes and radial nodes present in 3P orbital.
11. Wave lengths of different radiators are given below:  
 $\lambda (A) = 300\text{nm}$                        $\lambda (B) = 300\mu\text{m}$   
 $\lambda (C) = 3\text{nm}$                           $\lambda (D) = 30 \text{ \AA}$   
 Arrange these radiations in the increasing order of their energies.
12. a) Which of the following will not show deflection from the path on passing through an electric field?  
 Proton, cathod rays, electron, neutron  
 b) What is the experimental evidence in support of idea that electronic energies in an atom are quantized?
13. Calculate the energy required for the process,  $\text{He}^+(\text{g}) \rightarrow \text{He}^{2+}(\text{g}) + \text{e}^-$   
 The ionisation energy for the H-atom in the ground state is  $2.18 \times 10^{-18}\text{J-atom}^{-1}$ .
14. The longest wavelength doublet absorption transition is observed at 589 and 589.6 nm. Calculate the frequency of each transition and energy difference between two excited states.
15. Calculate the Kinetic energy of moving electron which has a wavelength of 4.8 pm (mass of electron =  $9.11 \times 10^{-31}\text{Kg}$ ,  $h=6.63 \times 10^{-34}\text{Js}$ )

**F. Very long answer type questions :**

**(Each question marks 5)**

1. Choose the correct alternative -
  - i) The reciprocal of wavelength is called wave function / wave number.
  - ii)  ${}^{37}_{17}\text{Cl}$  and  ${}^{35}_{17}\text{Cl}$  are isotope / isobar.
  - iii) Kinetic energy of ejected electron is proportional / inversely proportional to the frequency of the electromagnetic radiation.
  - iv) The element helium (He) / hydrogen(H) was discovered in the Sun by spectroscopic method.
  - v) Angular momentum is integral multiple of  $\frac{h}{4\pi}$  /  $\frac{h}{2\pi}$
  
2.
  - a) Write the limitations of Bohr model of atom.
  - b) Is there any difference between the orbital angular momentum of 3d and 4d electrons? explain.
  - c) If the velocity of a electron is zero, can we apply uncertainty principle for that electron? - Explain.
  
3.
  - a) Calculate the energy difference between third and fourth orbits of H-atom.
  - b) Calculate the total mass of protons present in 34mg of  $\text{NH}_3$  gas at STP.
  - c) How many nodes are present in 3s orbital?

## Solution

- A.**
- |      |       |       |       |
|------|-------|-------|-------|
| 1. a | 9. b  | 17. d | 25. b |
| 2. b | 10. b | 18. a | 26. b |
| 3. a | 11. c | 19. c | 27. a |
| 4. b | 12. d | 20. a | 28. d |
| 5. c | 13. b | 21. c | 29. d |
| 6. a | 14. b | 22. a | 30. a |
| 7. b | 15. a | 23. b | 31. c |
| 8. d | 16. a | 24. d | 32. b |
- B.**
- |      |       |
|------|-------|
| 1. a | 6. a  |
| 2. b | 7. d  |
| 3. b | 8. c  |
| 4. a | 9. a  |
| 5. a | 10. d |
- C.**
- |                         |             |                           |                            |
|-------------------------|-------------|---------------------------|----------------------------|
| 1. Cl                   | 2. 12,12,13 | 3. $1.098 \times 10^{27}$ |                            |
| 30. $=1.56 \times 10^3$ |             | 31. 130 pm                | 34. 4th orbit to 2nd orbit |
| 36. s, p and d          |             | 37. 9 orbitals            | 42. 5f                     |
| 45. ZnS.                |             |                           |                            |
- D.**
- |   |   |
|---|---|
| 16. ${}_{16}^{32}\text{S}^{-2}$   | 17. $5.486 \times 10^{-7}\text{Kg}$ , $9.647 \times 10^4\text{c}$ |
| 18. $5.17 \times 10^{14}\text{S}^{-1}$ , $1.72 \times 10^6\text{m}^{-1}$    |   |
| 19. $7.5 \times 10^{14}\text{Hz}$ , $4.0 \times 10^{14}\text{Hz}$           |   |
| 20. $2.01 \times 10^{20}$   |   |
| 21. $\lambda = 1282\text{nm}$ , $\gamma = 2.34 \times 10^{14}\text{S}^{-1}$ |   |
| 22. From $n = 2$ to $n = 1$   |   |
| 23. $-8.72 \times 10^{-18}\text{J}$ , $0.02645\text{ nm}$                   |   |
| 26. $1.523 \times 10^6\text{ m}^{-1}$                                       |   |
| 28. $-3.4\text{ eV}$  |   |
| 29. i) 4th to 2nd    ii) 5th to 4th   |   |
| 30. 1, 2, 15  |   |

- E.
1.  $5 \times 10^9 \text{S}^{-1}$ ,  $6.0 \times 10^{-2} \text{m}$ ,  $16.66 \text{m}^{-1}$
  2.  $4.41 \times 10^{14} \text{S}^{-1}$ ,  $2.92 \times 10^{-19} \text{J}$
  3. 15 lines
  4. I)  $l = 0, 1, 2$   
 $l = 0, m_l = 0$   
 $l = 1, m_l = -1, 0, +1$   
 $l = 2, m_l = -2, -1, 0, +1, +2$   
II)  $n = 3, l = 2$   
 $l = 2, m_l = -2, -1, 0, +1, +2$   
III) 2S, 2P
  5. a) Li, P                                b)  $n = 5$
  6. a) i) 16 electrons                    ii) 2 electrons  
      b) i) 4d                                ii) 3P
  8. 4.48 eV
  9. 332 pm
  10. angular -1, radial -1.
  11.  $B < A < C = D$
  13.  $8.72 \times 10^{-18} \text{J}$
  14.  $\gamma_1 = 5.093 \times 10^{14} \text{S}^{-1}$ ,  $\gamma_2 = 5.083 \times 10^{14} \text{S}^{-1}$ ,  $\Delta E = 3.313 \times 10^{-22} \text{J}$
  15.  $1.05 \times 10^{-14} \text{J}$



### Chapter at a glance :-

1. Periodic table is an arrangement of elements in which elements with similar properties are placed together in the same vertical column.
2. The elements are arranged in rows (period) and columns (group). From the position of an element in the periodic table, it is possible to know its properties.
3. At present, 118 elements are known.
4. **Dobereiner triads** : Dobereiner classified the elements in groups of three elements (triads) having similar properties and according to him, the middle elements of each of the triad had an atomic weight very close to the arithmetic mean of the other two elements.
5. **De Chancourtois classification** : De Chancourtois arranged the known elements in order of increasing atomic weights and made a cylindrical table of elements to display the periodic recurrence of properties.
6. **Newlands law of octaves** : According to Newlands, when the elements are arranged in the increasing order of their atomic weights, the properties of every eighth elements are similar to the first one.
7. **Lothar Meyer classification** : Lothar Meyer proposed that, when the elements are arranged in the increasing order of their atomic weights, similarities in physical and chemical properties appear at regular intervals.
8. **Mendeleev's periodic law** : According to Mendeleev's periodic law, the properties of the elements are a periodic function of their atomic weights.
9. Mendeleev's proposed that some of the elements were still undiscovered and, therefore, left several gaps in his table. For example, both gallium and germanium were unknown at that time. He left the gap under aluminium and a gap under silicon, and called these elements Eka-Aluminium and Eka-Silicon.
10. **Modern periodic law** : Modern periodic law can be stated as: the physical and chemical properties of the elements are periodic functions of their atomic numbers.
11. Modern periodic table consists of seven periods and eighteen groups.
12. Elements of the same group have similar valence shell electronic configuration and, therefore, exhibit similar chemical properties.
13. **IUPAC nomenclature of elements with atomic number > 100** : IUPAC recommended an nomenclature for naming elements with atomic numbers more than 100. This nomenclature is based on the Latin words for their numbers (Root name). For example, nil for 0, un for 1, bi for 2, tri for 3 and so on. Then the names are derived by combining together the roots in the order of digits from atomic number and ium is added at the end. For example, IUPAC name for the element with atomic number 106 is un (1) + nil (0) + hex (6) + ium → unnilhexium and its symbol is Unh

- Notation for IUPAC nomenclature of elements :

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	p
6	hex	h
7	Sept	s
8	oct	o
9	enn	e

14. The long form of the periodic table are divided into four blocks. There are s, p, d and f- blocks.

15. General electronic configuration for :

*s*-block :  $ns^{1-2}$  (Elements of group 1 and 2)

*p*-block :  $ns^2np^{1-6}$  (Elements of group 13 to 18)

*d*-block :  $(n-1)d^{1-10} ns^{1-2}$  (Elements of group 3 to 12)

*f*-block :  $(n-2)f^{1-14} (n-1)d^{0-1} ns^2$  (Elements of group 3)

16. In modern periodic table,

- First period contains 2 elements.
- Second and third periods contain 8 elements each.
- Fourth and fifth periods contain 18 elements each.
- Sixth and seventh periods contain 32 elements each.

17.

Group number	Name
1	Alkali metals
2	Alkaline earth metals
3-12	Transition metals
13	Boron family
14	Carbon family
15	Pnictogens or Nitrogen family
16	Chalcogens or Oxygen family
17	Halogen family
18	Noble gas or inert gas

All s-block elements are metals (except-Hydrogen),

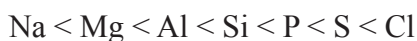
18. P-block elements are mainly non-metals, **d-block elements** are known as transition elements and f-block elements are known as inner-transition elements (i.e Lanthanoids and Actinoids)
19. The two rows of elements at the bottom of the periodic table, called the Lanthanoids, Ce (Z=58) to Lu (Z=71) and Actinoids, Th (Z=90) to Lr (Z=103)
20. Mercury (Hg) is the liquid metal, Bromine (Br) is the liquid non-metal and arsenic (As), antimony (Sb) etc are metalloids.
- 21. Prediction of period, group and block of an element :**
- (i) Period – Principal quantum number of the valence shell.
- (ii) Group –
- For s-block elements : Number of valence electrons (ns electrons)
- For p-block elements : 10 + number of valence electrons (ns and np electrons)
- For d-block elements : (n-1) d electrons + ns electrons.
- For f-block elements : Group is 3.
- (iii) Block – Subshell which receives the last electron.
- 22. Atomic radius :** It is the distance from the centre of the nucleus to the outermost shell of electrons in an atom.
- 23.** Atomic radius decreases while going from left to right in a period and increases while going from top to bottom in a group.
- 24. Covalent radius :** It is one half of the distance between the nuclei of the two similar atoms covalently bonded to each other.
- 25. Metallic radius :** It is one half of the distance between the nuclei of the two neighbouring atoms of a metal in a metallic lattice.
- 26. Van der waals radius :** It is one half of the internuclear distance between the nuclei of two adjacent atoms belonging to the two neighbouring molecules of the same substance in the **solid state**.
- Van der waals radius > Metallic radius > Covalent radius.
- 27. Ionic radius :** It is defined as the effective distance from the centre of the nucleus of an ion upto which it has influence on its electron cloud.
- 28.** Cations are smaller than the parent atoms and anions are larger than the parent atoms.
- 29.** Isoelectronic ions have same number of electrons. For example,  $N^{3-}$ ,  $O^{2-}$ ,  $F^{-}$ ,  $Na^{+}$ ,  $Mg^{2+}$ ,  $Al^{3+}$  are isoelectronic ions. All these ions have 10 electrons.
- 30.** For isoelectronic ions, size decreases with increase in nuclear charge. For example, the order of ionic radii is  $N^{3-} > O^{2-} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+}$
- 31. Ionization enthalpy :** It is the energy required to remove an electron from isolated gaseous atom. It is expressed in units of  $\text{KJmol}^{-1}$  or electron volt (eV) per atom or  $\text{kcal mol}^{-1}$ .
- 32. Factors affecting ionization enthalpy :**
- (i) Nuclear charge

- (ii) Size of the atom.
  - (iii) Screening effect of the inner electrons.
  - (iv) Penetration effect of electrons
  - (v) Electronic configuration.
33. (i) With the increase in nuclear charge, ionization enthalpy increases.
- (ii) Ionisation enthalpy increases with decrease in atomic size.
  - (iii) With the increase in screening effect of the inner electrons, ionization enthalpy decreases.
  - (iv) Ionisation enthalpy increases with increases in penetration power of the electrons. In a given shell the penetration power decreases in the order :  $s > p > d > f$
  - (v) Ionisation enthalpy increases with the stability of the electronic configuration.
34. In general, Ionisation enthalpy (IE) increases in moving from left to right in a period and decreases in moving from top to bottom in a group.
35. The order of ionization enthalpies of the second period elements is :  $Li < B < Be < C < O < N < F < Ne$
36. The order of successive ionization enthalpies is :  $IE_3 > IE_2 > IE_1$
37. **Electron gain enthalpy:** The enthalpy change when an electron is added to an isolated gaseous atom to convert it into a negative ion is known as electron gain enthalpy.
- The process of adding an electron to the atom can be either endothermic or exothermic.
- It is expressed in  $KJ\ mol^{-1}$  unit and is represented as  $\Delta_{eg} H$
38. Halogens have very high negative electron again enthalpy values because they can attain stable noble gas electronic configuration by picking up an electron.
39. **Factors affecting electron gain enthalpy:**
- i. Nuclear Charge:** As the nuclear charge increases, the electron gain enthalpy become more negative.
  - ii. Atomic Size:** The electron gain enthalpy become less negative with increase in the atomic size.
  - iii. Electronic Configuration:** The elements having stable electronic configurations will have less negative electron gain enthalpy.
40. In general, electron gain enthalpy increases in a period from left to right and decreases on moving down a group.
41. The electron gain enthalpies of halogens are in the order  $F < Cl > Br > I$
42. **Electronegativity:** A qualitative measure of the ability of an atom in a chemical compound to attract shared electrons towards itself is called electronegativity.
43. Fluorine is the most electronegative element.
44. **Factors affecting electronegativity:**
- i. Atomic radius:** Electronegativity increases with the decreasing atomic radius.
  - ii. Effective nuclear charge:** Electronegativity increases with the increasing effective nuclear charge.

iii. **Oxidation state:** In general, the electronegativity increases as the positive oxidation state of the atom increases.

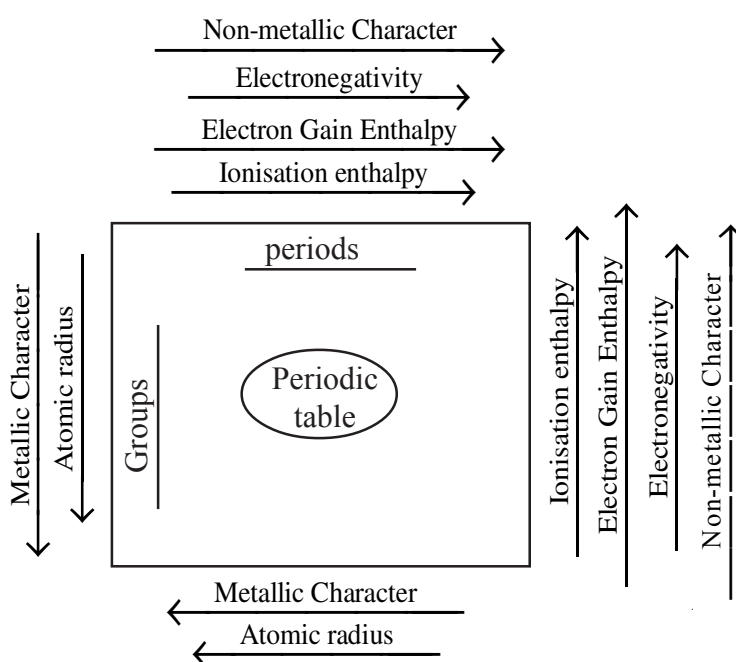
iv. **Hybridisation:** Electronegativity increases with the increasing s-character of the hybrid orbital of the atom.

45. Electronegativity generally increases across a period from left to right and decreases down a group in the periodic table.
46. **Metallic and non-metallic character :** Metallic character increases on moving down the group and decreases along the period. Reverse is true for non-metallic character.
47. **Valency:** it increases from 1 to 7 along a period with respect to oxygen and with respect to hydrogen it first increases from 1 to 4 and then decreases to 1. In a group, there is no change in valency. Noble gases have valency equal to zero.
48. **Oxidising and reducing property :** Oxidising property increases from left to right in a period, but reducing property decreases. For example, the order of oxidizing property is:



Reducing property increases from top to bottom in a group, but oxidizing property decreases. For example, the order of oxidizing property is:  $\text{F} > \text{Cl} > \text{Br} > \text{I}$

49. The basic strength of alkaline earth metal hydroxides in water increases down the group, i.e.,  $\text{Be}(\text{OH})_2 < \text{Mg}(\text{OH})_2 < \text{Ca}(\text{OH})_2 < \text{Sr}(\text{OH})_2 < \text{Ba}(\text{OH})_2$
50. The solubility of alkaline earth metal hydroxide in water increases as we go down the group.
51. Oxides formed by the elements on the left of a period are basic while those of the elements on the right of a period are acidic in nature. Oxides of elements in the centre are amphoteric or neutral.
52. Diagonal relationship occurs in Li with Mg, Be with Al and B with Si.



**A. Select the correct Answer for the following Questions :- (MCQ) (Each question Mark1)**

- Which property did Mendeleev use to classify the elements in his periodic table?
  - Atomic Weight
  - Atomic Number
  - Melting Point
  - Boiling Point
- Predict the position of an element in the periodic table having the electronic configuration  $1S^22S^22P^63S^23P^63d^54S^2$ 
  - Period - 4, Group - 6
  - Period - 4, Group - 7
  - Period - 5, Group - 7
  - Period - 6, Group - 5
- The correct order of ionic radii of the species  $Na^+$ ,  $Mg^{2+}$ ,  $F^-$ ,  $O^{2-}$  is -
  - $F^- < O^{2-} < Mg^{2+} < Na^+$
  - $Mg^{2+} < Na^+ < F^- < O^{2-}$
  - $O^{2-} < F^- < Na^+ < Mg^{2+}$
  - $O^{2-} < F^- < Mg^{2+} < Na^+$
- The maximum number of elements in 3rd period is
  - 18
  - 8
  - 32
  - 24
- The order of screening effect of electrons of S, P, d and f orbitals of a given shell of an atom on its outer shell electrons is
  - $S > P > d > f$
  - $P < d < s < f$
  - $f > d > S > p$
  - $f > P > S > d$
- Which of the following has largest negative electron gain enthalpy?
  - F
  - Cl
  - Br
  - I
- Which one of the following processes proceeds with the absorption of energy?
  - $Cl(g) + e^- \rightarrow Cl^-(g)$
  - $O(g) + e^- \rightarrow O^-(g)$
  - $O^-(g) + e^- \rightarrow O^{2-}(g)$
  - $F(g) + e^- \rightarrow F^-(g)$

8. Which of the following grouping represents a set of isoelectronic species -  
 a)  $\text{Na}^+$ ,  $\text{Mg}^{2+}$ ,  $\text{Ca}^{2+}$ , b)  $\text{F}^-$ ,  $\text{Na}^+$ ,  $\text{N}^{3-}$   
 c)  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Ca}^{2+}$  d)  $\text{Al}^{3+}$ ,  $\text{Cl}^-$ ,  $\text{K}^+$
9. Which of the following have no unit?  
 a) Atomic radius b) Electronegativity  
 c) Electron gain enthalpy d) Ionisation Enthalpy
10. Electronic configuration of four elements A, B, C and D are given below:  
 A)  $1\text{S}^22\text{S}^22\text{P}^6$  B)  $1\text{S}^22\text{S}^22\text{P}^4$   
 C)  $1\text{S}^22\text{S}^22\text{P}^63\text{S}^1$  D)  $1\text{S}^22\text{S}^22\text{P}^5$   
 Which of the following is the correct order of increasing tendency to gain electron?  
 a)  $\text{A} < \text{B} < \text{C} < \text{D}$  b)  $\text{A} < \text{C} < \text{B} < \text{D}$   
 c)  $\text{D} < \text{A} < \text{B} < \text{C}$  d)  $\text{D} < \text{B} < \text{C} < \text{A}$
11. The correct order of metallic character is :  
 a)  $\text{Al} < \text{Mg} < \text{B} < \text{K}$  b)  $\text{B} < \text{Al} < \text{Mg} < \text{K}$   
 c)  $\text{K} > \text{Mg} > \text{Al} > \text{B}$  d)  $\text{Mg} > \text{Al} > \text{K} > \text{B}$
12. In the modern periodic table, the period indicates the value of:  
 a) atomic number b) atomic mass  
 c) azimuthal quantum number d) principal quantum number
13. Which of the following statement is correct?  
 a) Helium has the highest first ionisation enthalpy in the periodic table.  
 b) Chlorine has less negative electron gain enthalpy than fluorine.  
 c) Chlorine is the highest electronegative element in the periodic table.  
 d) In any period, atomic radius of alkali metals is the lowest.
14. The correct order of ionisation enthalpy of C, N, O, F is:  
 a)  $\text{C} < \text{O} < \text{N} < \text{F}$  b)  $\text{F} < \text{N} < \text{C} < \text{O}$   
 c)  $\text{C} < \text{N} < \text{O} < \text{F}$  d)  $\text{F} < \text{O} < \text{N} < \text{C}$
15. In which of the following, the order is not correct according to the property indicated against it?  
 a)  $\text{Al}^{3+} < \text{Mg}^{2+} < \text{Na}^+ < \text{F}^-$   
 Increasing ionic size

- b)  $I < Br < F < Cl$   
 Increasing negative electron gain enthalpy
- c)  $B < C < N < O$   
 Increasing first ionisation enthalpy
- d)  $Li < Na < K < Rb$   
 increasing metallic radius.

16. Which of the following represents the correct order of decreasing ionic radii?

- a)  $Na^+ > Mg^{2+} > Al^{3+} > O^{2-} > F^-$   
 b)  $Na^+ > F^- > Mg^{2+} > O^{2-} > Al^{3+}$   
 c)  $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$   
 d)  $Al^{3+} > Mg^{2+} > Na^+ > F^- > O^{2-}$

17. The correct order of negative electron gain enthalpy of the halogen atoms is -

- a)  $Cl > F > Br > I$                       b)  $I > Br > Cl > F$   
 c)  $F < Cl > Br > I$                       d)  $Cl > Br > I > F$

18. The tendency towards complex formation is maximum in -

- a) S-Block elements                      b) P-Block elements  
 c) d-Block elements                      d) None of these

19. Which of the following oxide is most acidic?

- a) MgO                      b) BeO                      c) CaO                      d) BaO

20. The family of elements with the highest ionisation enthalpy is -

- a) Halogens                      b) Noble gases  
 c) Alkali metals                      d) Alkaline earth metals.

21. The atom having smallest atomic radius among the following is -

- a) Na                      b) K                      c) Rb                      d) Li

22. General electronic configuration of d-block elements is -

- a)  $(n-1)d^{1-10}ns^{1-2}$                       b)  $(n+1)d^{1-10}ns^{0-2}$   
 c)  $(n-1)d^{1-10}ns^{0-2}$                       d)  $(n-1)d^0ns^{0-2}$



23. Which of the following is the correct order of size of the given species -
- a)  $I^+ > I^- > I$
  - b)  $I > I^- > I^+$
  - c)  $I^- > I > I^+$
  - d)  $I > I^+ > I^-$
24. The diagonal partner of Be is -
- a) Li
  - b) Al
  - c) Mg
  - d) Si
25. The group of elements in which the differentiating electron enters the antipenultimate shell of atoms are called -
- a) S - Block Elements
  - b) P - Block Elements
  - c) d - Block Elements
  - d) f - Block Elements
26. Among the following elements, the most electronegative is -
- a) Chlorine
  - b) Nitrogen
  - c) Fluorine
  - d) Oxygen
27. The energy released when an isolated gaseous atom takes up an electron is called -
- a) Ionisation energy
  - b) Electronegativity
  - c) Electron gain enthalpy
  - d) Solvation energy
28. Which of the following electronic configuration will have the highest ionisation enthalpy?
- a)  $[\text{Ne}] 3S^2 3P^2$
  - b)  $[\text{Ne}] 3S^2 3P^3$
  - c)  $[\text{Ar}] 3d^{10} 4S^2 4P^3$
  - d)  $[\text{Ne}] 3S^2 3P^1$
29. The ionisation energy is -
- a) Lowest in noble Gases
  - b) Highest in the alkali metals
  - c) Lowest in the alkali metals
  - d) Lowest in the alkaline earth metals.
30. Long form of the periodic table is based on the properties of elements as a function of
- a) Atomic number
  - b) Atomic size
  - c) Atomic mass
  - d) Atomic volume

31. Which of the following is wrong?

- a) Ionic size:  $N^{3-} > O^{2-} > F^-$   
b) Electron gain enthalpy :  $Cl > F > Br$   
c) Atomic Size:  $O > F > N$   
d) Ionisation enthalpy:  $N > O > F$

32. According to modern periodic law, the properties of elements are the periodic functions of their

- a) Atomic mass                                  b) Density  
c) Atomic Number                                d) Mass Number

33. The number of periods and groups in the long periodic table are

- a) 7 and 18                                        b) 7 and 9  
c) 6 and 10                                        d) 8 and 18

34. The correct decreasing order of atomic radius among Na, K, Mg and Rb is -

- a)  $Rb > Na > K > Mg$                         b)  $Mg > Rb > K > Na$   
c)  $K > Rb > Na > Mg$                         d)  $Rb > K > Na > Mg$

35. Which of the following statement is not correct?

- a) Ionisation energies of noble gases are high.  
b) The  $IE_2$  is always more than  $IE_1$   
c) Within a group, there is a gradual increase in ionisation energy due to increase in nuclear charge.  
d) Ionisation energy of Be is more than B.

**B. Assertion and Reason type questions :**

**(Each question mark 1)**

**The questions given below consists of an Assertion and the Reason. Use the following Key to choose the appropriate answer.**

- a) Both assertion and reason are correct and reason is the correct explanation of the assertion.  
b) Both assertion and reason are correct, but reason is not the correct explanation of the assertion.  
c) Assertion is correct but reason is incorrect.  
d) Assertion is incorrect but reason is correct.

1. Assertion (A) : Inert gases have zero electron affinity.

Reason (R) : Due to stable electronic configuration, inert gases have very little tendency to gain additional electrons.

2. Assertion (A) : Mg<sup>2+</sup> ion is smaller than Na<sup>+</sup> ion.  
Reason (R) : For iso-electronic species, greater the effective nuclear charge per electron, lesser is the size.
  
3. Assertion (A) : F atom has less negative electron gain enthalpy than Cl atom.  
Reason (R) : Additional electrons are repelled more effectively by 3P-electrons in Cl than by 2P-electrons in F atom.
  
4. Assertion (A) : According to Mendeleev, periodic properties of elements is a periodic function of their atomic number.  
Reason (R) : Atomic number is equal to the number of protons.
  
5. Assertion (A) : Atomic Size increases along a period.  
Reason (R) : Effective nuclear charge increases as the atomic number increases and results in the increased attraction of electrons to the nucleus.
  
6. Assertion (A) : Helium is placed in group 18 along with P-block elements.  
Reason (R) : Helium shows similiar properties like P-Block elements.
  
7. Assertion (A) : The order of ionic radii is N<sup>3-</sup> > O<sup>2-</sup> > F<sup>-</sup>  
Reason (R) : These are iso electronic species. So, size decreases with increasing atomic number.
  
8. Assertion (A) : The elements having 1S<sup>2</sup>2S<sup>2</sup>2P<sup>3</sup> and 1S<sup>2</sup>2S<sup>2</sup>2P<sup>6</sup>3S<sup>2</sup>3P<sup>3</sup> configuration belong to same group  
Reason (R) : These have same number of valence electrons.
  
9. Assertion (A) : Ionisation enthalpy decreases as we go down a group.  
Reason (R) : Size of atoms decreases on going down a group.
  
10. Assertion (A): Ionisation enthalpy of nitrogen is greater than that of oxygen.  
Reason (R) : Nitrogen has half filled stable configuration.

**C. Very short answer type questions :-**

**(Each question mark 1)**

1. Who developed the long form of periodic table?

2. Write the IUPAC name of the element with atomic number 103.
3. How many periods and groups are present in the long form of the periodic table?
4. Write the general electronic configuration of P-Block elements.
5. Give one example of representative element.
6. What are lanthanides?
7. Give one example of transition element.
8. What are transition elements?
9. Give one example of metalloid?
10. Predict the position of the element in the periodic table whose electronic configuration is  $[\text{Ar}]3d^14s^2$ .
11. How many elements are present in the third period of the periodic table?
12. What do you mean by iso electronic species?
13. Define electron gain enthalpy.
14. Which element has the highest electron gain enthalpy?
15. Arrange the following elements in order of increasing ionisation enthalpy.  
B, C, N, O
16. Arrange the following in order of increasing ionic size.  
 $\text{Al}^{3+}$ ,  $\text{Na}^+$ ,  $\text{F}^-$ ,  $\text{Mg}^{2+}$
17. Give example of one cation which is iso electronic with Ne.
18. Give one example of amphoteric oxide.

19. Name the element present in the third period and sixteenth group of the periodic table.
20. Arrange S, P, d, f orbitals in order of decreasing penetration power.
21. Which element is most electronegative in the periodic table?
22. Which important property did Mendeleev use to classify the elements in his periodic table?
23. State modern periodic law.
24. Name one inner transition element.
25. Which element has diagonal relationship with Mg?

**D. Short answer type question :-**

**(Each question marks 2)**

1. What is the basic difference between the Mendeleev's periodic law and the modern periodic law?
2. Why  $\text{Na}^+$  ion is smaller in size than Na atom?
3. Why oxygen has lower ionisation enthalpy than nitrogen?
4. What are the differences between electron gain enthalpy and electronegativity?
5. How does ionisation enthalpy vary in a period and in a group?
6. Why halogens have the highest negative electron gain enthalpies?
7. Zinc is not a transition element - Explain.

8.
  - i) State Mendeleev's periodic law.
  - ii) Write one drawback in Mendeleev's periodic table.
  
9. Consider the following species:  $O^{2-}$ ,  $Na^+$ ,  $N^{3-}$ ,  $Mg^{2+}$ ,  $Na^+$ 
  - i) What is common in them?
  - ii) Arrange these species in increasing order of ionic radii.
  
10. The electron gain enthalpy of chlorine is  $349 \text{ KJmol}^{-1}$ . How much energy in KJ is released when 1g of chlorine is converted completely to  $Cl^-$  ions in the gaseous state?
  
11. Assign the position of the following elements having outer electronic configuration:
  - i)  $(n-1)d^2ns^2$ ,  $n = 4$
  - ii)  $(n-2)f^7(n-1)d^1ns^2$ ,  $n = 6$ , in the periodic table.
  
12. Explain: ionisation enthalpy decreases in a group from top to bottom.
  
13. How does the metallic and non-metallic character vary on moving from left to right in a period?
  
14. What is diagonal relationship? Give example.
  
15. Why do elements in the same group have similar physical and chemical properties?
  
16. Write two properties of d-block elements.
  
17. Write the general outer electronic configuration of d and f-block elements.

18. Give two examples of metalloids.
19. What do you mean by d-block elements? Give one example.
20. Write the IUPAC name and symbol of the element with atomic number 112.
21. Why anions are larger in radii than their parent atom?
22. What are the various factors on which ionisation enthalpy does depend?
23. Arrange the elements P, O, N and S in the order of
  - i) Increasing non-metallic character.
  - ii) Increasing first ionisation enthalpy.
24.
  - i) Define electronegativity.
  - ii) Arrange N, O, F and P in order of increasing electronegativity.
25. What are actinoids? Give one example.

**E. Short answer type questions :**

**(Each question marks 3)**

1.
  - i) Explain why the electron gain enthalpy of fluorine is less negative than that of chlorine.
  - ii) Arrange F, Cl, Br, I in order of increasing electron gain enthalpy.
2. The first ionisation enthalpy of sodium is lower than that of magnesium but its second ionisation enthalpy is higher than that of magnesium - explain.
3.
  - i) Why electron gain enthalpy values of noble gases are positive?
  - ii) Which is more metallic: Na or K?

4. Discuss the factors on which electron gain enthalpy does depend?
5.
  - i) State Newland's law of octaves.
  - ii) Define Van der Waal's radius.
6.
  - i) Explain why Be has higher  $\Delta_f H$  than B.
  - ii) Which is more basic :  $\text{Al}(\text{OH})_3$  or  $\text{Mg}(\text{OH})_2$ ?
7.
  - i) How would you react to the statement that electronegativity of N on Pauling scale is 3.0 in all its compounds?
  - ii) Give one example of liquid metal.
8.
  - i) How would you justify the presence of 18 elements in the 5th period of the periodic table?
  - ii) Which block elements are strong reducing agents?
9. Discuss the factors affecting electronegativity and the trend in its variation in the periodic table.
10. What do you mean by exothermic reaction and endothermic reaction? Give one example of each type.

**F. Long answer type questions :**

**(Each question marks 5)**

1. Discuss the factors affecting electron gain enthalpy and the trend in its variation in the periodic table.
2. Write the electronic configurations of the elements with atomic number 9 , 17, 19, 22 and 36. Predict the following from these configurations:
  - i) Which of them has the highest electronegativity?
  - ii) Which of them has the highest electron gain enthalpy?
  - iii) Which of them has the lowest ionisation potential?
  - iv) Which of them has zero electron gain enthalpy?
  - v) Which of them is a transition element?



3. i) All transition elements are d-block elements, but all d-block elements are not transition elements - Explain.  
 ii) Electronegativity of elements increases on moving from left to right in the periodic table - Explain.  
 iii) Which block of elements consists of metals, non-metals and metalloids?

4. a) Energy of an electron in the ground state of the hydrogen atom is  $-2.18 \times 10^{-18} \text{J}$ . Calculate the ionisation enthalpy of atomic hydrogen in terms of  $\text{J mol}^{-1}$ .  
 b) The first ( $\text{IE}_1$ ) and second ( $\text{IE}_2$ ) ionisation enthalpies ( $\text{KJ mol}^{-1}$ ) of three elements A, B and C are given below:

	A	B	C
$\text{IE}_1$	403	549	1142
$\text{IE}_2$	2640	1060	2080

Identify the element which is likely to be (i) a non-metal ii) an alkali metal iii) an alkaline earth metal.

5. i) Using the periodic table, predict the formulas of compounds which might be formed by the following pairs of elements:  
 a) Silicon and Bromine      b) Aluminium and Sulphur  
 ii) By reacting with water show that  $\text{Na}_2\text{O}$  is a basic oxide and  $\text{Cl}_2\text{O}_7$  is an acidic oxide.

## Solution

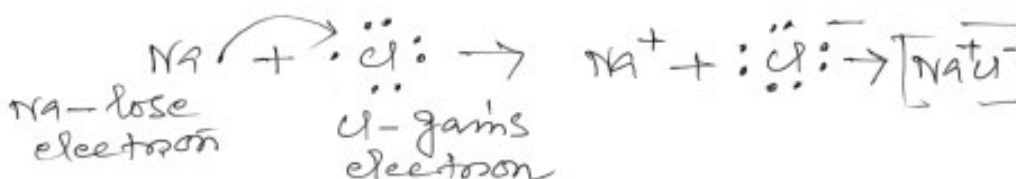
- A:**
- |         |         |         |
|---------|---------|---------|
| 1. (a)  | 16. (c) | 31. (d) |
| 2. (b)  | 17. (c) | 32. (c) |
| 3. (b)  | 18. (c) | 33. (a) |
| 4. (b)  | 19. (b) | 34. (d) |
| 5. (a)  | 20. (b) | 35. (c) |
| 6. (b)  | 21. (d) |         |
| 7. (c)  | 22. (c) |         |
| 8. (b)  | 23. (c) |         |
| 9. (b)  | 24. (b) |         |
| 10. (a) | 25. (d) |         |
| 11. (c) | 26. (c) |         |
| 12. (d) | 27. (c) |         |
| 13. (a) | 28. (b) |         |
| 14. (a) | 29. (c) |         |
| 15. (c) | 30. (a) |         |

**D:** 10. 9.83 Kj

**F:** 4 (a),  $13.13 \times 10^5 \text{ J mol}^{-1}$

## Chemical Bonding

- 1. Chemical bond:** Chemical bond is nothing but the force of attraction between the two oppositely charged ions, two identical or different atoms or two molecules. Such as attraction between  $\text{Na}^+$  &  $\text{Cl}^-$ , or attraction between water molecules.
- 2. Type of bonds:** i) Ionic bond or Electrovalent bond. ii) Co-valent bond, iii) Co-ordinate bond or Co-ordinate co-valent bond. iv) H - bond, v) Vanderwall force of attraction etc.
- 3. Necessity of Chemical bond :** Tendency to attain the noble gas-configuration, minimum energy and stability.
- 4. Lewis symbols:** Valence electron of atom and ion are expressed by dot (.) symbol. Such as  $\text{Na}\cdot$ ,  $\ddot{\text{O}}:$   $:\ddot{\text{O}}:$  etc.
- 5. Ionic bond:** Ionic bond is formed between metal and non metal. Such as  $\text{NaCl}$ ,  $\text{CaO}$ ,  $\text{K}_2\text{S}$  etc. This bond is formed by lossing and gaining electron (s) in between metal and non-metal. Like.



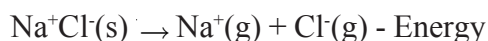
In ionic bond strong electro-static force of attraction are acting between the two oppositely charged ions.

- 6. Factors on which the formation of ionic bond depend:**
  - Atoms those have low IE and high electron gain enthalpy can easily form ionic bond.
  - High lattice energy also goes to formation of ionic bond.
- 7. Lattice enthalpy :** The amount of energy released during the formation of one mole of solid from the constituent gaseous ion. Such as,



This released energy is the Lattice Energy, denoted by symbol U.

On the other hand Lattice Energy can also be defined as the energy required for the separation of one mole of ionic solid into gaseous ionic constituents.



In both cases magnitude of energy are same. Higher the Lattice Energy stability of the molecule will be more.

- 8. Factors on which lattice energy depends :**
  - Lattice enthalpy  $\propto$  Products of the charge of ions

ii. Lattice enthalpy  $\propto \frac{1}{\text{internuclear distance between the two oppositely charged ions}}$

Lattice enthalpy of  $\text{Al}_2\text{O}_3$  is more than  $\text{Na}_2\text{O}$ , Lattice enthalpy of  $\text{MgO}$  is more than  $\text{CaO}$ .

### 9. Some important properties of ionic compound:

- Ionic compounds are non-directional. Ionic compounds usually exist in the form of crystalline solid.
- Melting and Boiling point of ionic compounds are generally more.
- Ionic compounds are generally soluble in water and other polar solvents those have high dielectric constant.
- Ionic compounds has isomorphism property.
- Rate of reaction between ionic compounds in aqueous medium is more.
- Ionic compounds can conduct electricity either in aqueous medium or in fused state.

**10. Covalent bond:** This type of bond is formed by mutual sharing of electron(s) in between the non-metals. Such as -

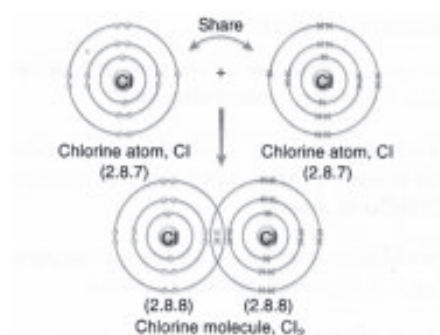
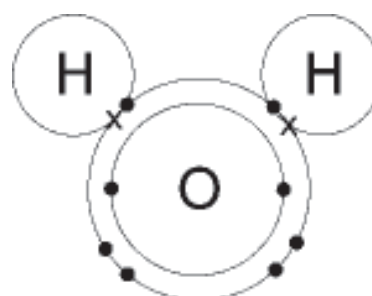


Figure Formation of a chlorine molecule,  $\text{Cl}_2$



water molecule

### 11. Types of Covalent Bond:

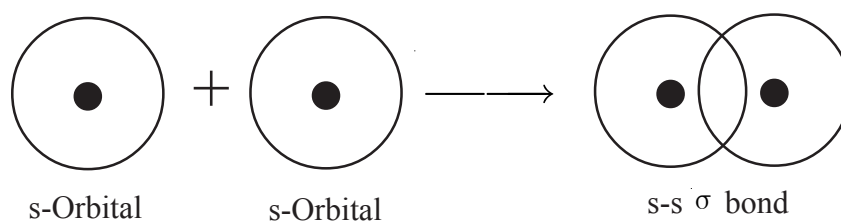
On the basis of overlapping :

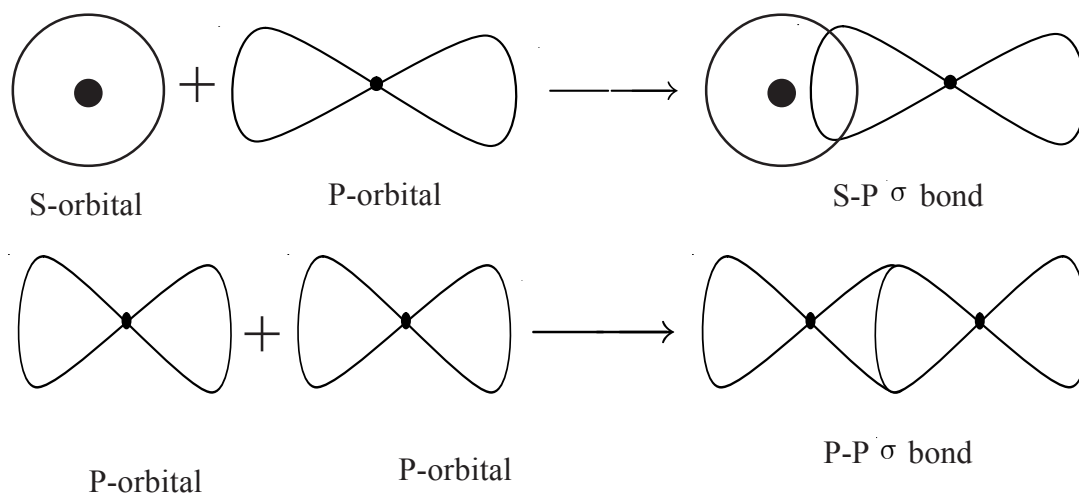
- Sigma ( $\sigma$ ) bond
- Pi ( $\pi$ ) bond

On the basis of the difference of electronegativity -

- Polar covalent bond
- Non-polar covalent bond

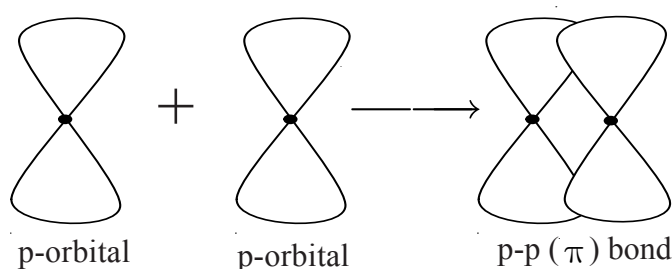
**12. Sigma ( $\sigma$ ) bond:** During face to face overlapping sigma ( $\sigma$ ) bond is formed, such as -



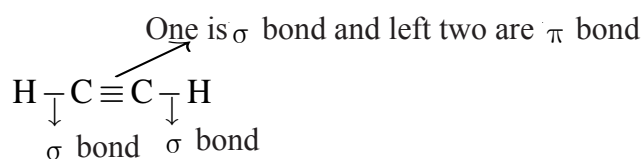


Without sigma ( $\sigma$ ) bond a co-valent molecule does not exist and in between two atoms there are only one sigma bond.

**13. Pi ( $\pi$ ) bond:** During lateral overlapping a  $\pi$  bond is formed.



$\pi$  bond is weaker than sigma bond. If there are more than one bonds in between two atoms then there will be pi - bond. Such as in  $C_2H_2$ ,



So in  $C_2H_2$  compound total  $\sigma$  bond = 3  
 and total  $\pi$  bond = 2

**14. Polar and non-polar co-valent bond:-**

Polar covalent bond:

If electronegativity difference between two hetero atom in a molecule is more, then it is polar covalent bond. Such as  $\overset{\delta+}{\text{H}}-\overset{\delta-}{\text{F}}, \overset{\delta+}{\text{H}}-\overset{2\delta-}{\text{O}}-\overset{\delta+}{\text{F}}$

### Non-polar co-valent bond :

This type of bond will form between two identical atom or two hetero atom having less electro negativity difference. Like,  $\text{Cl}_2$ ,  $\text{CH}_4$ ,  $\text{CO}_2$  etc.

### 15. Rules for writing Lewis structure:

- Firstly calculate the total number of valence electrons of all the atoms, such as in the  $\text{CH}_4$  molecules there are 8 valence electron available for bonding (4 from carbon and 4 from 4 hydrogen) atoms.
- For anions, each negative charge would be added with the total no. of valence electrons and for cation each positive charge would be subtracted from the total no. of valence electron.  
For example,  $\text{NO}_2^-$  ion, total no. of electrons available for bonding  $17 + 1 = 18$ .  
For  $\text{NH}_4^+$ , Total No. of electron available for bonding  $= 9 - 1 = 8$
- Now you have to write the skeletal structure of the molecule and distribute the total no. of electrons as bonding shared pairs between the atoms in proportion to the total bonds.
- Least electronegativity atom-will be the central atom of a molecule / ions. Such as  $\text{NF}_3$  and  $\text{CO}_3^{2-}$ , 'N' and 'C' atoms are the central atom. Generally 'H', 'N' and 'O' are terminal atoms.
- After distributing the shared pairs of electrons for single bonds the remaining electron pairs are used either for multiple bond or they constitute lone pair.

For example, for drawing the single bond between  $\downarrow \text{NO}_2^-$  ion structure

Step - 1 : Total No of electrons = 18

Step - 2 : The skeletal structure of  $\text{NO}_2^-$  ion is



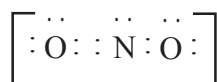
Step - 3 : Now draw the single bond between N & O.



Step - 4 : Now you have to complete the octets of atom



In this structure  
'N' has no octet so



### 16. Lewis dot representation of some molecules



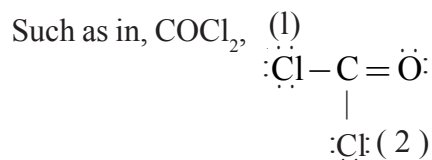
**17. Formal Charge :** Formal charge help us to determine the lowest energy structure among the Lewis dot structures.

$$\text{Formal charge (F. C.) on an atom in a Lewis structure} = V - \frac{S}{2} - U$$

Here, V = No. of valence electrons (S)

U = No. of non bonding electrons or unshared electrons

S = Total no. of shared electrons



$$\text{FC of 'Cl' (1)} = 7 - 6 - \frac{2}{2} = 0$$

$$\text{FC of 'Cl' (2)} = 7 - 6 - \frac{2}{2} = 0$$

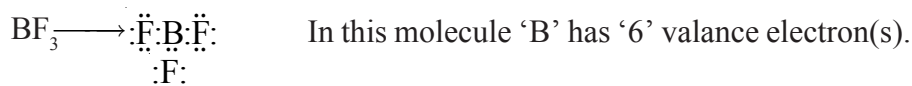
$$\text{FC of 'O' atom} = 6 - 4 - \frac{4}{2} = 0$$

$$\text{FC of 'C' atom} = 4 - \frac{8}{2} - 0 = 0$$

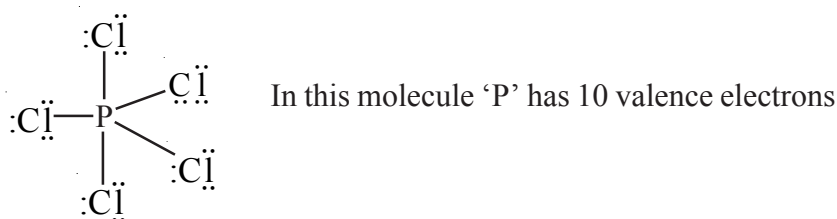
More stable structure is the lowest energy structure and the lowest energy structure is the one which has lowest formal charges on the atoms.

### 18. Limitation of the Octet rule :

a) Incomplete octet of central atom :



b) Expanded Octet :



c) Odd-electron molecule :



### Other Limitation of Octet theory :

Octet rule is based on the inertness of noble gases. But some noble gases can combine with  $F_2$ ,  $O_2$  etc and to form  $XeF_2$ ,  $XCOF_4$  etc.

### 19. Some important properties of Covalent compound:

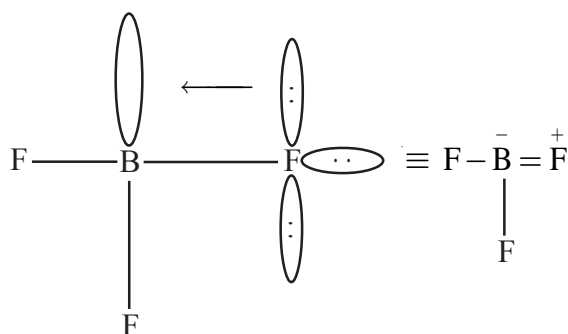
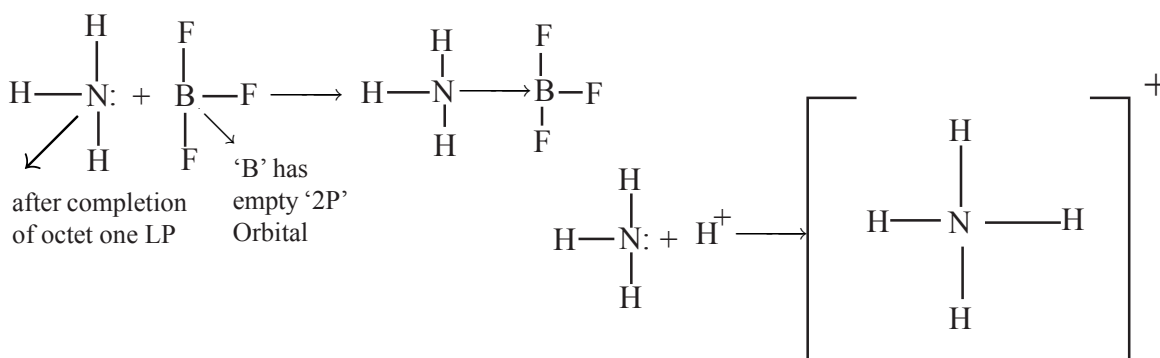
1. Covalent compound exist as molecule. Weak inter molecular force of attraction are acting between the molecules.
2. Covalent compound has low melting point and low boiling point.
3. Covalent compounds are insoluble or less soluble in water. They are generally soluble in organic solvent. Such as  $CCl_4$ ,  $CS_2$  etc.
4. Generally covalent compounds are non conductor or poor conductor.
5. Covalent compound take part in the chemical reaction in molecular form. Hence rate of reaction is slow.

### 20. Co-ordinate covalent bond:

Co-ordinate covalent bond is expressed by ( $\longrightarrow$ ) Co-ordinate covalent bond is also known as dative bond or semipolar bond.

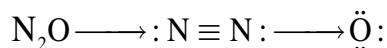
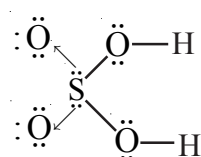
21. For the formation of co-ordinate bond one atom should have free lone pair after completion of octet and another atom should have vacant orbital. Besides, size of the orbital should be same.

For example,

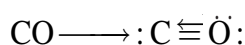




Structure of  $\text{H}_2\text{SO}_4$



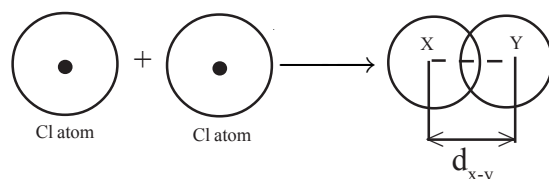
Due to presence of co-ordinate bond in between N & O,  $\text{N}_2\text{O}$  can produce [O] more easily. Thus  $\text{N}_2\text{O}$  is the better supporter of combustion than  $\text{O}_2$ .



In presence of dative bond-in between C & O, the net dipole moment of CO is less than the expected dipole moment.

**20. Bond length:** Actually bond length is the internuclear distance between the two homonuclear covalent molecule and also two-heteronuclear molecule.

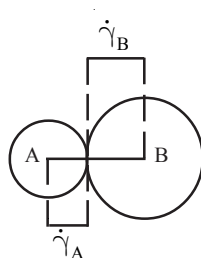
For example -



$d_{x-y}$  is the co valent distance

$d_{x-y}$  is the bond length of covalent molecule  $\text{Cl}_2$

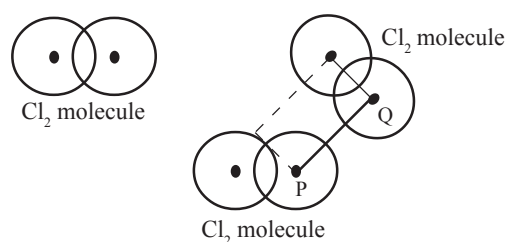
In case of heteronuclear molecule,  $d_{A-B}$  is the bond length of heteronuclear molecule.



$$\therefore d_{A-B} = r_A + r_B$$

**Vander Waal bond distance/ bond length :-**

Inter nuclear distance of two adjacent atom of two nearest molecule is the vander waal bond length. For example  $\text{Cl}_2$  molecule.



$d_{p-q}$  is the vander waal distance and  $1/2 d_{p-q}$  is the vander waal radius.

Co-valent radius is always less than vander waal radius.

- Note:**
- Size of the atom are small then bond length will be less.
  - Bond order i.e. no. of bonds in between two atoms is more, bond length is less.

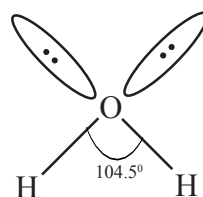
Average bond length for some single, double or triple bonds are given below in a chart.

Bond Type	Co valent bond length (Pm)
O - H	96
C - H	107
C - C	154
C = O	121
C = C	133
C $\equiv$ C	120

Bond length of some common molecule (in pm)

H <sub>2</sub>	74
F <sub>2</sub>	144
Cl <sub>2</sub>	199
Br <sub>2</sub>	228
I <sub>2</sub>	267
N <sub>2</sub>	109
O <sub>2</sub>	121

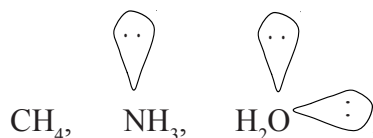
- 21. Bond Angle :** Bond angle is the angle between the orbitals containing bonding electron pairs around the central atom in a molecule. Bond angle is expressed in degree and bond angle also help us in determining the shape of the molecule. Such as -



Bond angle depends upon the following factors -

- If state of hybridisation of central atom are same and number of LP on the central atom are different then with increasing number of lp bond angle decreases.

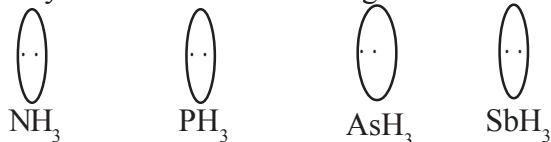
Such as,



Here all the central atom are SP<sup>3</sup> hybridised but no. of lp on 'C' atom is nil and on 'O' atom are two. Therefore order of Bond angle.



ii. State of hybridisation and number of lp on central atom are same then with increasing electro negativity of central atom bond angle increases.

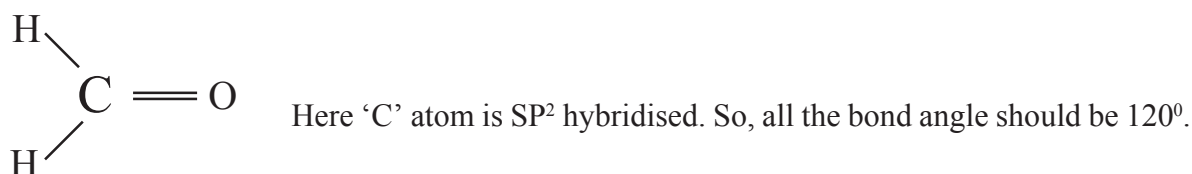


Order of EN of central atom  $\text{N} > \text{P} > \text{As} > \text{Sb}$

Therefore, order of bond angle



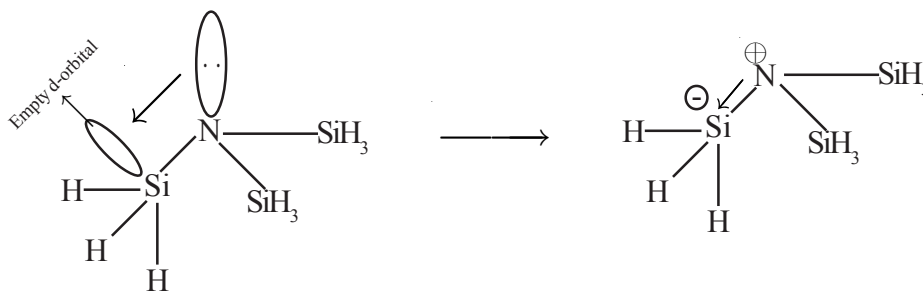
iii. In presence of multiple bond between the atoms bond angle increases then the expected value. For example,



But practically it is different. Value of  $\text{H} - \text{C} = \text{O}$  bond angle is more than  $120^\circ$  and

iv. If dative bond is formed between two atoms in a molecule then bond angle will be more. For example,

Bond angle of  $\text{N}(\text{SiH}_3)_3$  is more than  $\text{N}(\text{CH}_3)_3$



No dative bond will be formed in between N & C due to lacking of empty orbital in C atom.

## 22. Bond Enthalpy or Bond Energy :

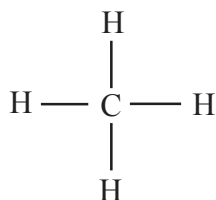
Energy required for the breaking of one mole of bond of a particular type between two atom in gaseous state. Unit of bond energy is  $\text{KJmol}^{-1}$  or  $\text{KCal mol}^{-1}$

Such as  $\text{H}-\text{H}(\text{g}) \longrightarrow \text{H}(\text{g}) + \text{H}(\text{g}) - 435.8\text{KJmol}^{-1}$ . This reaction also indicates the heat of atomisation.

$$\Delta_a H^\circ = 435.8\text{KJmol}^{-1}$$

**Note :** More is the bond dissociation enthalpy, more thermally stable is the molecule.

In case of poly atomic molecule, bond energy of two identical atom are not same. Such as in CH<sub>4</sub> molecule,

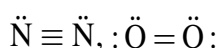


Bond enthalpy of four C - H bond are not same. Let the bond energy of '4' C-H bonds are X<sub>1</sub>, X<sub>2</sub>, X<sub>3</sub> and X<sub>4</sub> KJmol<sup>-1</sup> respectively.

$$\text{Hence average bond enthalpy} = \frac{X_1 + X_2 + X_3 + X_4}{4} \text{ KJmol}^{-1}$$

**Some important information about bond energy :**

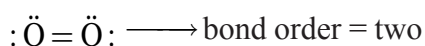
1. Lesser is the bond length, more is the bond energy. Such as H - H bond energy is more than I - I bond energy.
2. If bond length is small but no. of lp around the atom is more i.e. electron cloud around the atom is more, then bond energy is less due to more repulsive force. Such as : $\ddot{\text{F}} - \ddot{\text{F}}:$ , : $\ddot{\text{Cl}} - \ddot{\text{Cl}}:$ . Bond energy of Cl<sub>2</sub> molecule is more than F<sub>2</sub>, though bond length of F<sub>2</sub> is less than Cl<sub>2</sub>. Similarly N - N single-bond is weaker than P - P single bond.
3. If bond order is more, bond energy is also more. Such as,



Bond energy of N<sub>2</sub> is more than O<sub>2</sub>. One important relation between BL, BE and BO is

$$\text{BE} \propto \frac{1}{\text{BL}} \propto \text{BO} \quad [\text{BL} = \text{Bond Length}; \text{BE} = \text{Bond Energy}; \text{BO} = \text{Bond Order}]$$

**23. Bond Order :** Number of bonds in between two atoms in a co-valent molecule. Such as



In case of, iso electronic molecule / ions bond order is same. Such as

Molecule / ion	No. of electron (S)	bond order
F <sub>2</sub>	18	1
O <sub>2</sub> <sup>2-</sup>	18	1

An important information is, more is the Bond Order the molecule is more stable.

Bond order of some important molecule

CO - bond order - 3

NO<sup>+</sup> - bond order - 3

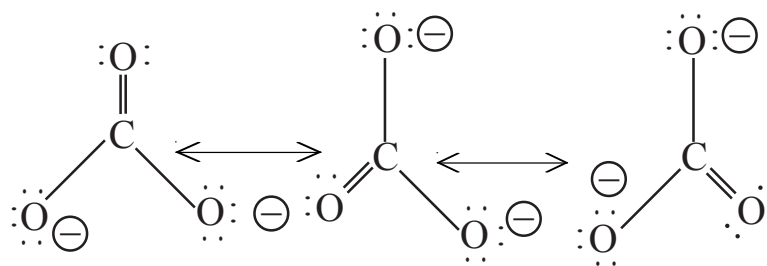
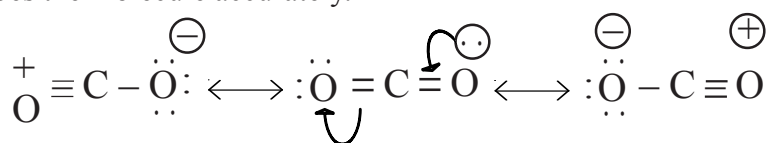
N<sub>2</sub><sup>+</sup> - bond order - 2.5

O<sub>2</sub><sup>+</sup> - bond order - 2.5

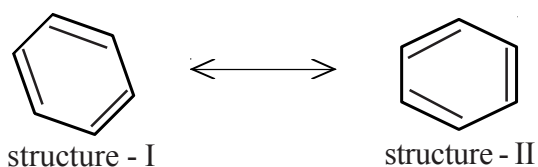
CO<sup>+</sup> - bond order - 2.5

- 23. Resonance :** According to the concept of resonance, whenever a single Lewis structure can not describe a molecule perfectly, a number of structures with similar energy position of nuclei, bonding and non bonding pairs of electrons are taken as the canonical structures of the hybrid which describes the molecule accurately.

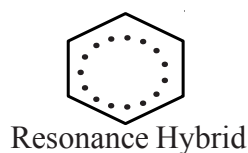
For example:



- 24. Resonance Hybrid :** The structure by which all the resonating structures can be expressed is the resonance hybrid. Such as



I and II can be expressed by resonance hybrid



- 25. Dipole moment:** Dipole moment is the vector quantity by which we can measure the degree of polarity of polar covalent molecule.

In case of heteroatomic molecule, dipole moment is measured by  $\mu = e \times d$ , where  $\mu$  = dipole moment,  $e$  = charge,  $d$  = internuclear distance between the two atoms. In case of polyatomic ion or molecule

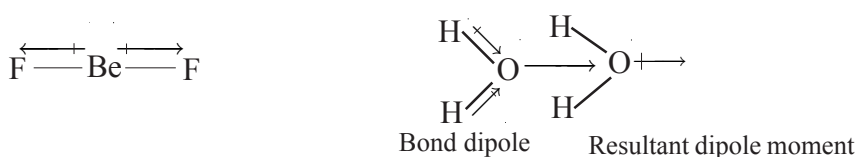
$\mu = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos \theta}$ , Here  $\mu_1$  = one bond moment,  $\mu_2$  = Another bond moment,  $\theta$  = Angle between the two moments.

### Unit of Dipolemoment :

CGS system = Debye (D) [1D =  $10^{-18}$ esu x cm]

SI system = Coulomb m.

Dipole moment of some important molecule may be represented as ,



### Important informations :

- Value of  $\mu \propto \frac{1}{\theta}$ , such as,
 

$\theta = 60^\circ$

meta di chloro benzene  
 $\theta = 120^\circ$

para dichloro benzene  
 $\theta = 180^\circ$

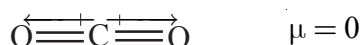
Therefore order of dipole moment

$$i > ii > iii$$

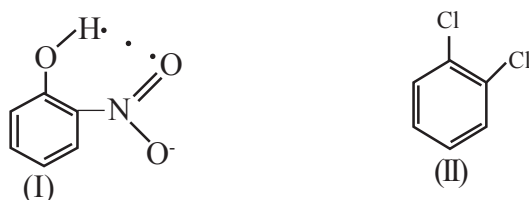
- We can also calculate the percentage of ionic character with the help of dipole moment value.

$$\% \text{ of ionic character} = \frac{\mu_{\text{obs}}}{\mu_{\text{cal}} / \mu_{\text{theo}}} \times 1000$$

- There are some molecule where polar bond is present but as a whole dipole moment is zero. Such as  $\text{CO}_2$ .



- Among 'O<sup>-</sup>' nitrophenol and O<sup>-</sup>-dichloro benzene formar has more dipole moment



In (I), due to intra molecular H-bond  $\theta < 60^\circ$ , So (I) has more dipole moment than (II).

## 26. Fajan's Rule:

Fajan's rule is applicable for the determination of the co-valent character or ionic character of a molecule.

Following points goes in favour of ionic character

1. Greater the size of cation, more will be ionic character.
2. Smaller the size of anion, ionic character of the compound will be more.
3. If cation has 8 electrons in the outer most shell, then ionic character of the compound will be more. Such as,

- (i) NaCl and KCl,  
Size of  $K^+ > Na^+$   
Hence KCl is more ionic than NaCl.
- (ii) NaCl and NaF  
Size of  $F^- < Cl^-$   
Therefore NaF is more ionic.
- (iii)  $CaCl_2$  and  $SnCl_4$   
 $Ca^{2+}$  has 8 electrons in the valence shell  
 $Sn^{4+}$  has 18 electrons in the valence shell  
Hence  $CaCl_2$  is more ionic.

The following points goes in favour of co-valent compound.

- (i) If charge of cation is more, the covalent character of the compound will be more.

Among NaCl and  $AlCl_3$ ,  $Al^{3+}$  has more charge. So  $AlCl_3$  is the more covalent compound

- (ii) Larger the size of anion and also charge on anion is more, the compound covalent character will have more.

Among  $Na_2O$  and  $Na_2S$ ,  $Na_2S$  is the more covalent compound.

- (iii) If cation has '18' electrons in the outer most shell, this goes in favour of more covalent character of the compound.

## 27. Valence Shell Electron Pair Repulsion (VSEPR) Theory

**Main postulate of VSEPR Theory :**

- i. The shape of a molecule depends upon the number of valence shell electron pairs (bonded or non bonded) around the central atom.
- ii. Pairs of electrons (bonded or non bonded) repel each other, since electrons are negatively charged.
- iii. These pairs of electrons tend to occupy such position in the molecule that minimum repulsion and maximum stability occur.
- iv. The valence shell is taken as a sphere with the electron pairs localising on the spherical surface at maximum distance from one another.
- v. A multiple bond is treated as if it is a single electron pair and the two or three electron pairs of a multiple bond are treated as a single super pair.

vi. Where two or more resonance structures can represent a molecule, the VSEPR model is applicable to any such structure.

**Order of repulsion force :**

lp(lonepair) - lp repulsion > lp-bp(bond pair) > bp-bp

For the prediction of geometrical shapes of molecule with the help of VSEPR Theory it is convenient to divide molecules into two categories as

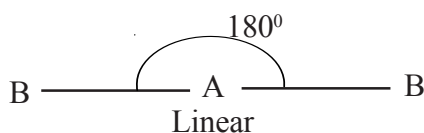
- Molecules in which the central atom has no lone pair and
- Molecules in which the central atom has one or more than one lone pairs.

According to VSEPR Theory shape of the molecule of the following type having no lp on the central atom.

- AB<sub>2</sub>- Linear
- AB<sub>3</sub>- Trigonal planer
- AB<sub>4</sub>- Tetrahedral
- AB<sub>5</sub>- Trigonal bi pyramidal/Octahedral
- AB<sub>6</sub>- Octahedral
- AB<sub>7</sub>- Pentagonal bipyramical

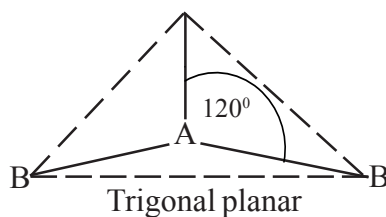
**Note :** 'A' is the central atom  
'B' is the atoms around the central atom

Such as AB<sub>2</sub> -



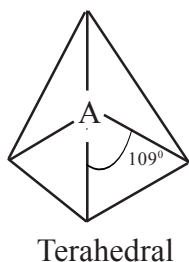
eg-BeCl<sub>2</sub>

AB<sub>3</sub>-



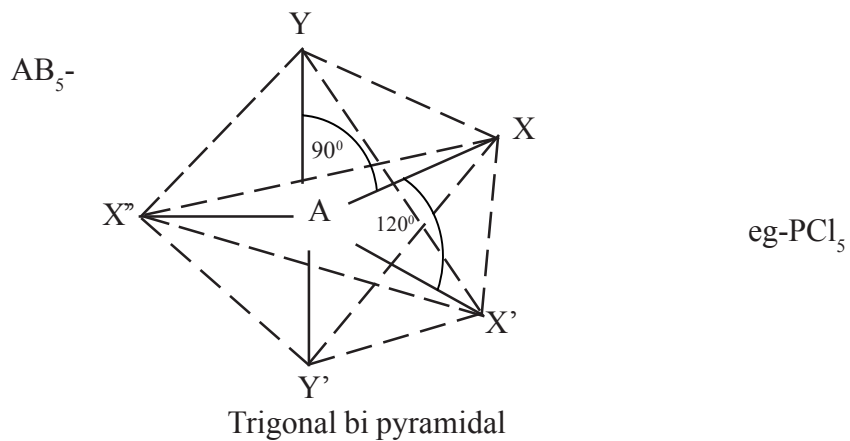
eg-BF<sub>3</sub>

AB<sub>4</sub>-



eg-CH<sub>4</sub>

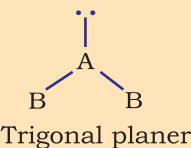
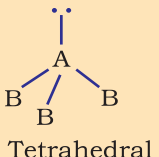
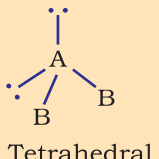
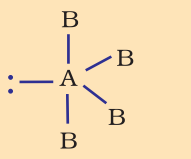
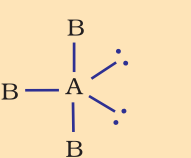
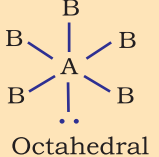
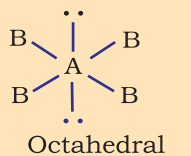




**Table 4.6 Geometry of Molecules in which the Central Atom has No Lone Pair of Electrons**

Number of electron pairs	Arrangement of electron pairs	Molecular geometry	Examples
2	<p>Linear</p>	B—A—B Linear	$BeCl_2$ , $HgCl_2$
3	<p>Trigonal planar</p>	<p>Trigonal planar</p>	$BF_3$
4	<p>Tetrahedral</p>	<p>Tetrahedral</p>	$CH_4$ , $NH_4^+$
5	<p>Trigonal bipyramidal</p>	<p>Trigonal bipyramidal</p>	$PCl_5$
6	<p>Octahedral</p>	<p>Octahedral</p>	$SF_6$

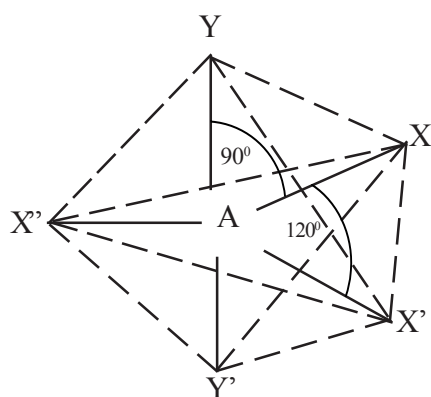
**Table 4.7 Shape (geometry) of Some Simple Molecules/Ions with Central Ions having One or More Lone Pairs of Electrons(E).**

Molecule type	No. of bonding pairs	No. of lone pairs	Arrangement of electron pairs	Shape	Examples
$AB_2E$	2	1	 <p>Trigonal planer</p>	Bent	$SO_2O_3$
$AB_3E$	3	1	 <p>Tetrahedral</p>	Trigonal pyramidal	$NH_3$
$AB_3E_2$	2	2	 <p>Tetrahedral</p>	Bent	$H_2O$
$AB_4E$	4	1	 <p>Trigonal bi-pyramidal</p>	See saw	$SF_4$
$AB_3E_2$	3	2	 <p>Trigonal bi-pyramidal</p>	T-shape	$ClF_3$
$AB_5E$	5	1	 <p>Octahedral</p>	Square pyramid	$BrF_5$
$AB_4E_2$	4	2	 <p>Octahedral</p>	Square planer	$XeF_4$

### Important information about the geometry of trigonal bipyramidal and Octahedral

In trigonal bi pyramidal and octahedral geometry there are two points.

1. Equatorial Point
2. Axial Point

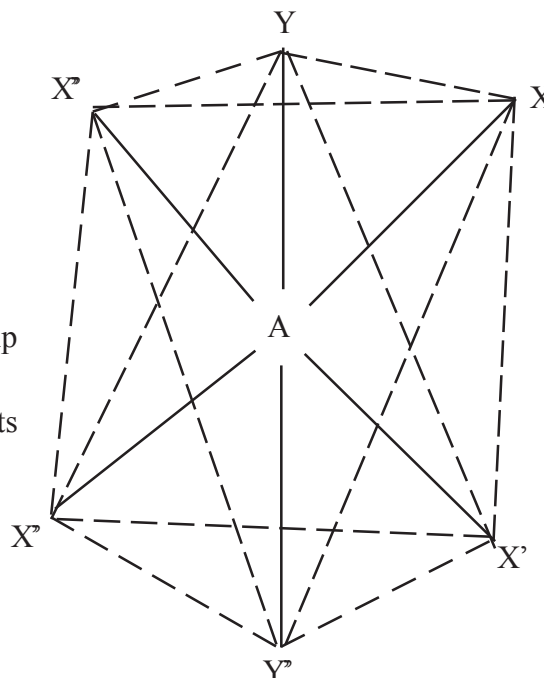


X, X' and X'' are the equatorial Points  
Y, Y' are axial points.

If there are lp in trigonal bi pyramidal geometry, lp goes to equatorial points.

Similarly, X, X', X'' and X''' are the equatorial points and Y, Y' are axial points.

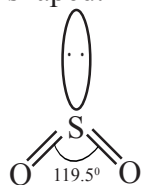
In Octahedral geometry, lp must go to axial point.



### Shapes of molecules according to VSEPR Theory containing lp and bond pairs

AB<sub>3</sub> - Type - one lp + 2bp - such as SO<sub>2</sub> molecule geometry gets distorted and it become V shaped or angular shaped.

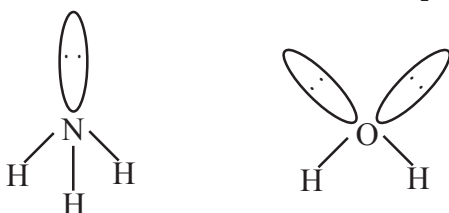
Fig: -



AB<sub>4</sub> - Type - i) One lp + 3bp - such as NH<sub>3</sub> molecule

ii) 2lp + 2bp - such as H<sub>2</sub>O molecule

Fig: -



Already we learnt that, AB<sub>4</sub> type is tetra hedral geometry, but due to lp geometry get disturbed. According to 1st one, shape of the molecule become pyramidal and for 2nd one, molecule become bent or angular.

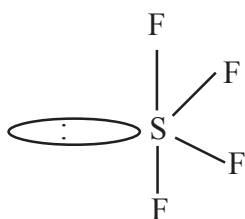
AB<sub>5</sub> - Type - i) 1lp + 4bp - Such as SF<sub>4</sub> molecule

ii) 2 lp + 3bp - ClF<sub>3</sub>, molecule

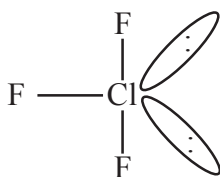
iii) 3lp + 2bp - ICl<sub>2</sub>, I<sub>3</sub> molecule etc.

AB<sub>5</sub> - Type is trigonal bi pyramidal geometry.

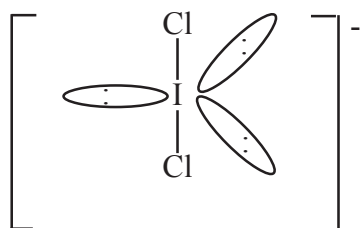
For one lp, geometry become see- saw shape.



For two lp, geometry become T-shaped



For three lp, geometry become linear  $\text{ICl}_2$

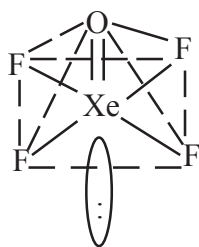


$\text{AB}_6$  - Type is Octahedral geometry

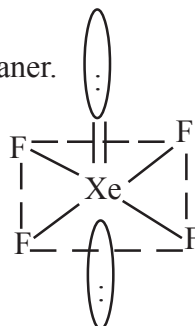
There are i) 1 lp + 5 bp - such as  $\text{XeOF}_4$  molecule

ii) 2 lp + 4 bp - such  $\text{XeF}_4$  molecule

Due to presence of one lp, geometry become square pyramidal.



Due to presence of two lp, geometry become square planer.



## 28. Valence bond theory :

Lewis failed to explain the following features of a chemical compound.

- i) Formation of chemical bond
- ii) Difference of bond dissociation enthalpy and bond length like -  $\text{F}_2$ ,  $\text{H}_2$  etc.
- iii) Shape of poly atomic molecules.

Similarly the VSEPR Theory, also has limited application.

To overcome these drawbacks two important theories based on quantum mechanical principles are established. These are VBT (Valence Bond Theory) and MOT (Molecular Orbital Theory).

Valence bond theory is based on atomic orbitals, electronic configuration of elements, overlapping of atomic orbitals and the principles of variation and superposition of atomic orbitals.

Consider two hydrogen atoms X & Y approaching each other having nuclei  $N_x$  and  $N_y$ , electrons present in hydrogen atoms are  $e_x$  and  $e_y$ .

When the two atoms are far away there is no interaction between them. When the atoms approach to each other then attractive and repulsive forces begin to operate.

Attraction force takes place due to

- i)  $N_x - e_x$  and  $N_y - e_y$
- ii)  $N_x - e_y$  and  $N_y - e_x$

Similarly repulsion force takes place due to

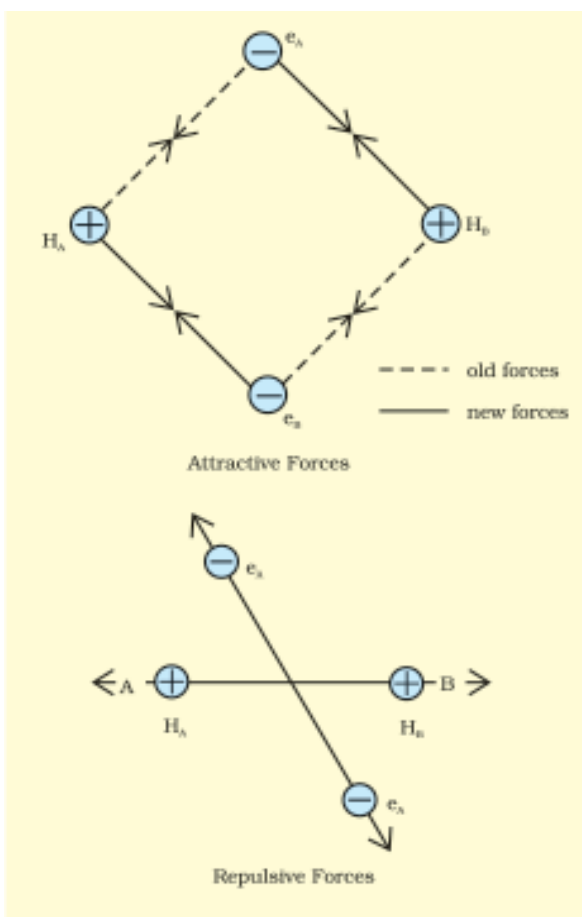
$$e_x - e_y \quad \text{and} \quad N_x - N_y$$

During attraction force, atoms comes in contact to each other and form chemical bond where as due to repulsion, atoms tends to stay away from each other.

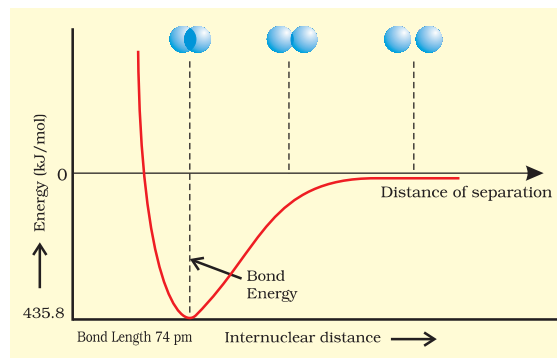
Experimental result revealed that magnitude of new attractive force is more than the repulsive force.

As a result Potential energy decreases when two atoms approaches to each other and form a chemical bond and attains minimum energy and we get a definite bond length of  $H_2$  molecule which is 74 pm.

Since the magnitude of attractive force is more than repulsive force, hence energy get released during the bond formation and  $H_2$  molecule become more stable than the isolated gaseous Hydrogen atoms.



**Fig. 4.7** Forces of attraction and repulsion during the formation of  $H_2$  molecule.



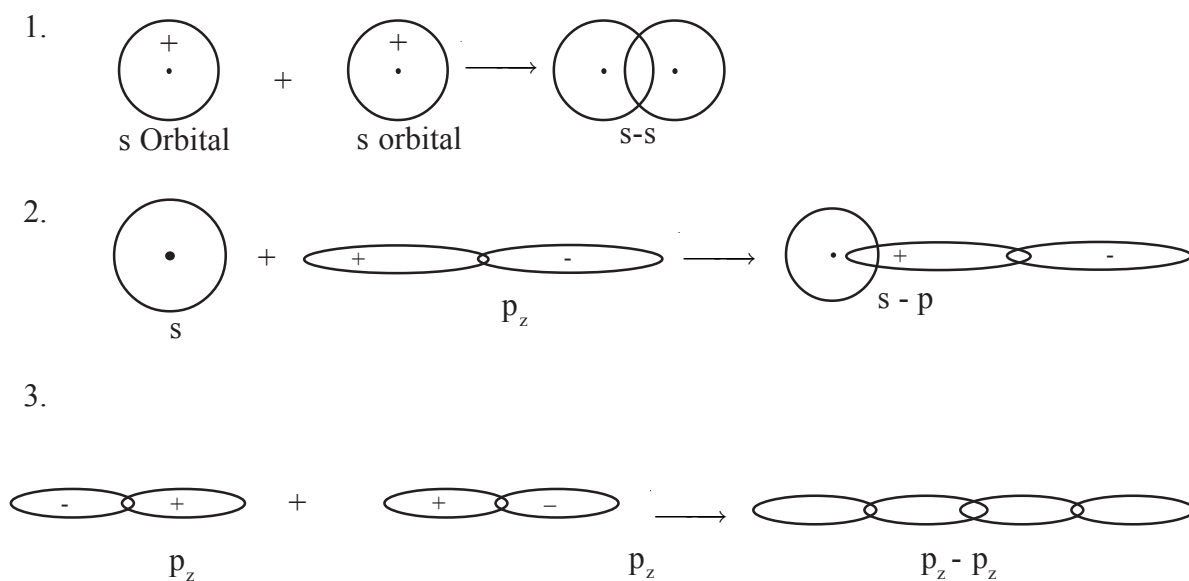
**Fig. 4.8** The potential energy curve for the formation of  $H_2$  molecule as a function of internuclear distance of the H atoms. The minimum in the curve corresponds to the most stable state of  $H_2$ .

### 29. Importance / Significance of VBT :

VBT help us to explain

- i) Shape of the molecule
- ii) The formation and directional properties of bonds in polyatomic molecules like  $\text{CH}_4$ ,  $\text{NH}_3$  etc in terms of overlap and hy-bridisation of atomic orbitals.

### 30. Overlapping of atomic orbitals :



### 31. Type of Overlapping :

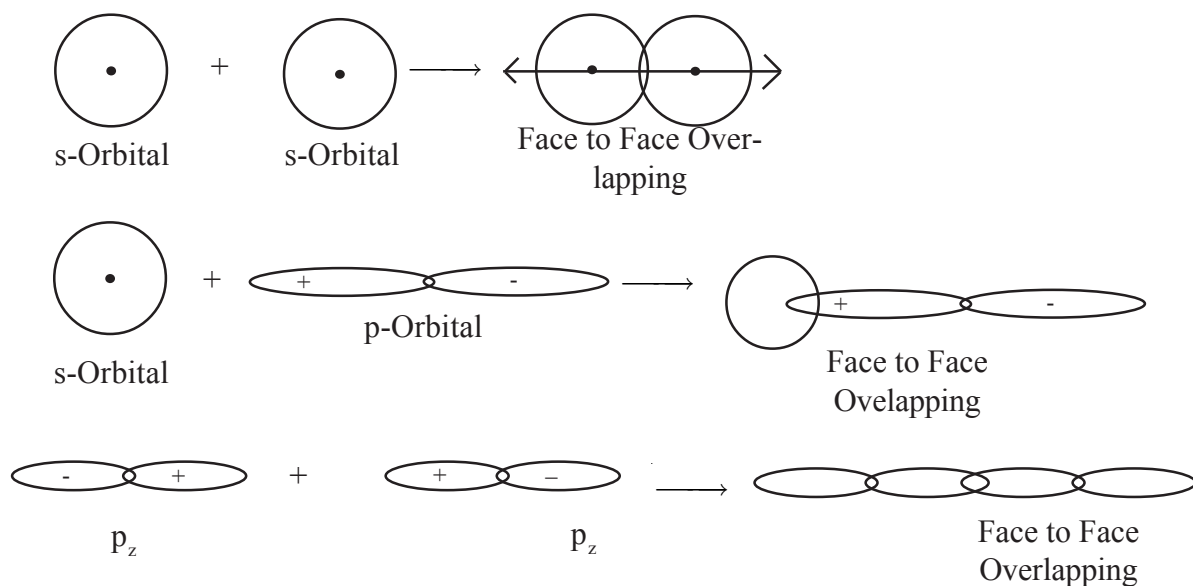
There are two types of overlapping

- i) Face to Face Overlapping or End to End Overlapping
- ii) Lateral Overlapping

#### Face to Face Overlapping :

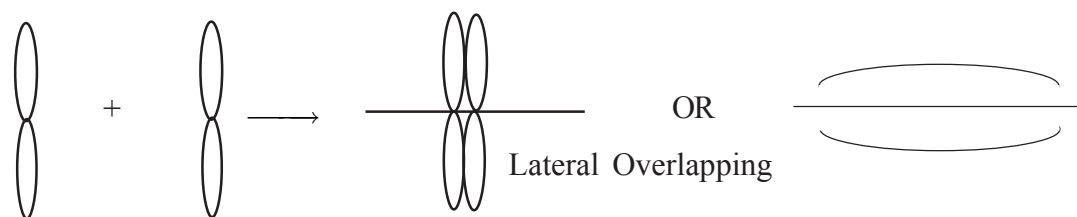
During overlapping, if nucleus of two atomic orbitals are at  $180^\circ$  angle and two atomic orbitals are in internuclear axis then, it will be face to face overlapping.

The diagram of Face to Face overlapping are given below:



### Lateral Overlapping :

During Lateral Overlapping, the atomic orbitals are parallel to each other and the orbitals are perpendicular to the internuclear axis.



## 32. Nature of Covalent bond on the basis of Overlapping :

There are two types of Co-Valent bond.

- I. Sigma ( $\sigma$ ) bond
- II. Pi ( $\pi$ ) bond

**Sigma ( $\sigma$ ) bond :** This type of bond is formed due to Face to Face Overlapping.

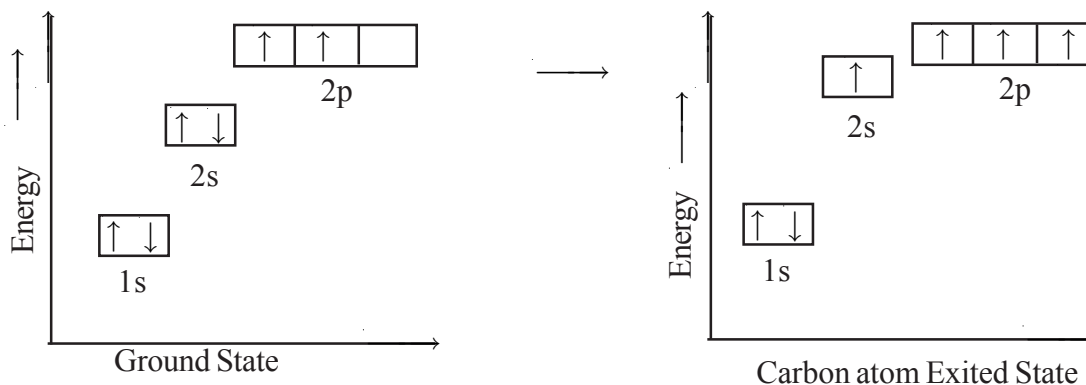
**Pi ( $\pi$ ) bond :** During Lateral overlapping  $\pi$  bond is formed.  $\pi$  bond is weaker than  $\sigma$  bond.

An important information is s-orbital always form sigma bond and p-orbital can form  $\pi$  as well as  $\sigma$  bond.



### 33. Hybridisation :

Pure atomic orbitals can exchange the energy among them and form hybrid orbitals. Thus produced hybrid orbitals are identical in all respect. The phenomenon by which hybrid orbitals are formed is known as Hybridisation.



#### Important information :

- I. Hybrid orbitals are identical in all respect i.e. energy and shape.
- II. No. of atomic orbitals. = No. of hybrid orbitals
- III. In the formation of stable bond, hybrid orbitals are more effective than the pure orbitals.
- IV. Hybrid orbitals are directed in space in such a way that the repulsion force between the electron pairs are minimum and we get a stable structure.

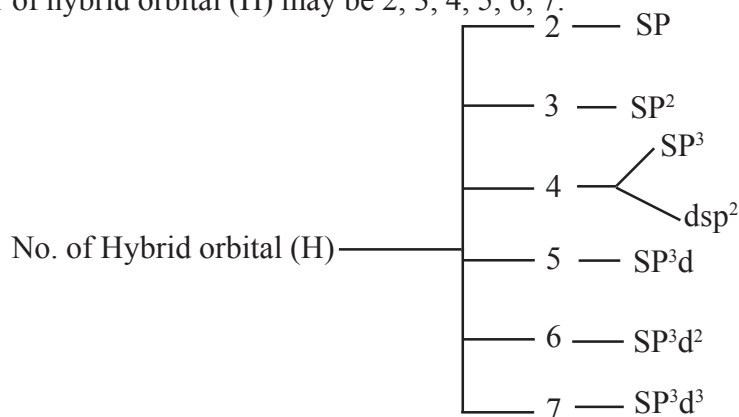
### 34. Important condition for Hybridisation :

- I. The orbitals present in the valence shell in an atom, get hybridised.
- II. The orbitals undergoing hybridisation should have almost equal energy.
- III. Promotion of electron is not essential condition prior to hybridisation.
- IV. It is not necessary that only half filled orbitals can take part in the hybridisation. To some extent even filled orbitals of valence shell can also take part in the hybridisation

### 35. Rules for finding out the state of hybridisation of central atom in a compound.

From the number of orbitals participating in the hybridisation of central atom(H) of a molecule, we can assume the hybridisation state and geometry of the molecule.

Number of hybrid orbital (H) may be 2, 3, 4, 5, 6, 7.



### 1st Rule

$$\text{No. of hybrid orbital (H)} = \frac{1}{2}[V + M - C + A]$$

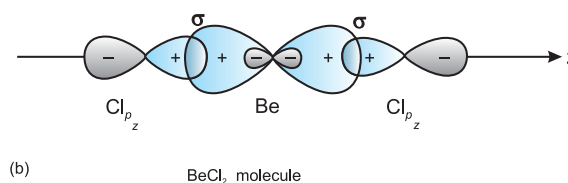
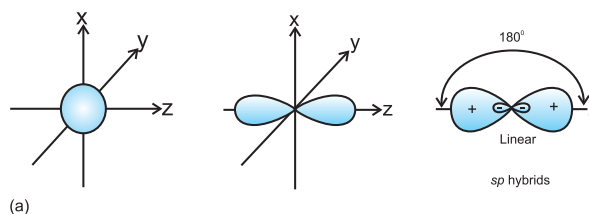
where V = No. of Valence electrons of central atom  
 M = No. of mono valent atom (Except central atom)  
 C = Charge of cation.  
 D = Charge of anion.

**Example of Hybridisation of some compounds are given below :**

1.  $\text{BeCl}_2$  'Be' is the central atom.

$$\text{No. of hybrid orbital (H)} = \frac{1}{2}[2 + 2 - 0 + 0] = 2$$

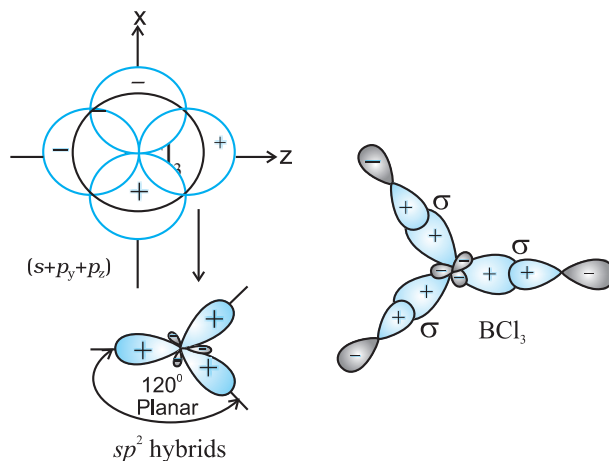
Therefore, 'Be' is '-sp' hybridised.



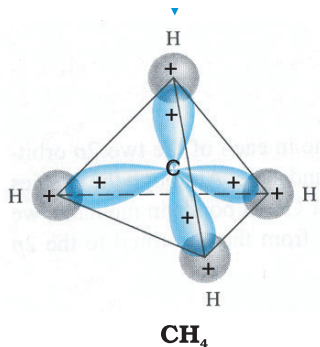
2.  $\text{BCl}_3$  - 'B' is the central atom.

$$(\text{H}) = \frac{1}{2}[3 + 3 - 0 + 0] = 3$$

Therefore, 'B' is -sp<sup>2</sup> hybridised

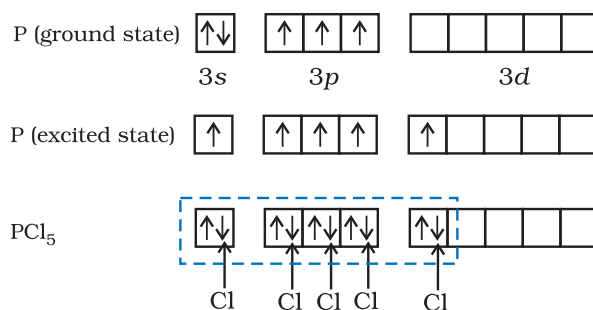


3.  $\text{CH}_4$ ,  $H = \frac{1}{2}[4 + 4 - 0 - 0] = 4 \longrightarrow \text{SP}^3$  ('C' has no 'd' orbital)



4.  $\text{PCl}_5$ , 'P' is the central atom.

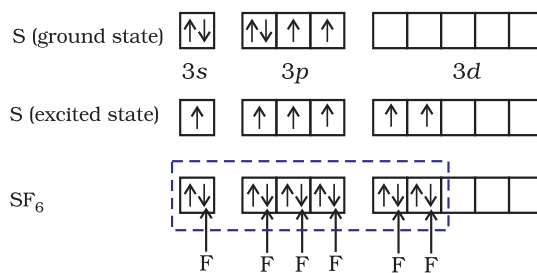
$H = \frac{1}{2}[5 + 5 - 0 + 0] = 5 \longrightarrow \text{SP}^3\text{d}$



*sp<sup>3</sup>d hybrid orbitals filled by electron pairs donated by five Cl atoms.*

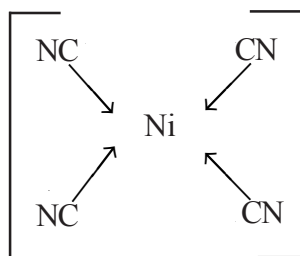
5.  $\text{SF}_6$ , '-s' is the central atom,  $H = \frac{1}{2}[6 + 6 - 0 + 0]$

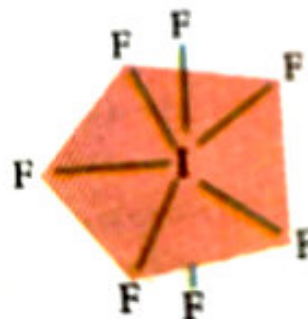
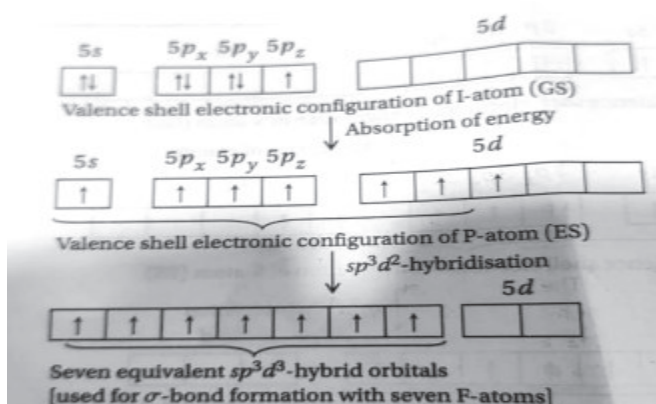
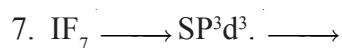
$= 6 \longrightarrow \text{SP}^3\text{d}^2$



6.  $[\text{Ni}(\text{CN})_4]^{2-} \longrightarrow \text{dsp}^2$

*sp<sup>3</sup>d<sup>2</sup> hybridisation*

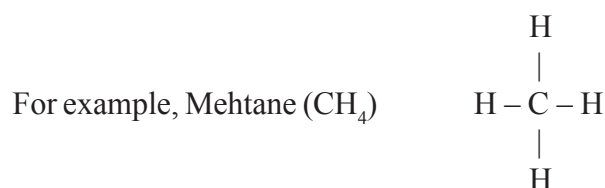




**2nd rules for finding out the state of hybridisation.**

**Generally this rule is applicable for organic chemistry**

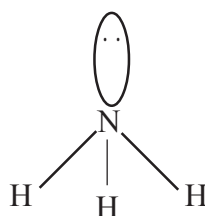
No. of Hybrid orbital (H) = No. of Sigma ( $\sigma$ ) + No. of localised lp.



No. of Hybrid orbital (H) =  $4 + 0 = 4$

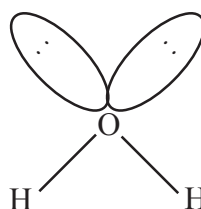
In Ammonia, ( $\text{NH}_3$ )

No. of hybrid orbital (H) =  $3\sigma + 1\text{lp} = 4$



In  $\text{H}_2\text{O}$ ,

No. of Hybrid orbital (H) =  $2\sigma + 2\text{lp} = 4$



In  $\text{H} - \ddot{\text{Cl}} :$

No. of hybrid orbital (H) in HCl molecule =  $1\sigma + 3\text{lp} = 4$ .



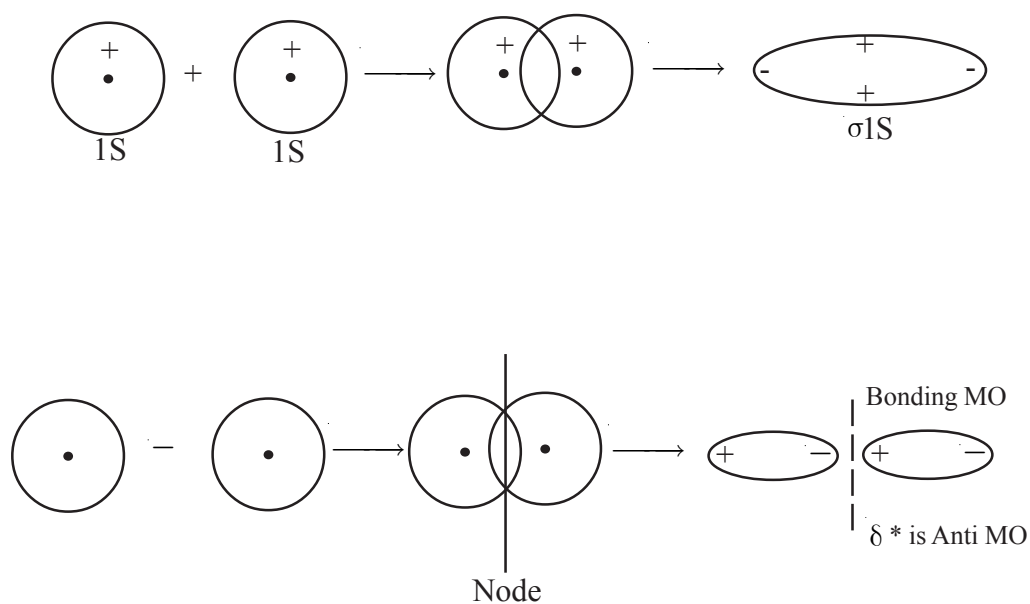
Sl. No.	No. of Hybrid Orbital (H)	No. of Ip	Bond angle	Geometry	Shape of the molecule	Example
1	H = 2 SP Hybridised	0	180°	Linear	Linear	BeCl <sub>2</sub> , CO <sub>2</sub>
2	H = 3 SP <sup>2</sup> Hybridised	0	120°	Trigonal Planer	Trigonal Planer	BF <sub>3</sub> , BH <sub>3</sub>
		1		V – Shaped	V – Shaped	SO <sub>2</sub>
3	H = 4 SP <sup>3</sup> Hybridised	0	109°28'	Regular Tetrahedral	Regular Tetrahedral	CH <sub>4</sub>
		1	107°	Pyramidal	Pyramidal	NH <sub>3</sub>
		2	104.5°	V-Shaped	V – Shaped	H <sub>2</sub> O
4	H = 5 SP <sup>3</sup> d Hybridisation	0	120°, 90°	Trigonal Bipyramidal	Trigonal Bipyramidal	P <sub>2</sub> Cl <sub>5</sub>
		1			See – saw shaped	SF <sub>4</sub>
		2			'T' – Shape	ClF <sub>3</sub>
		3			Linear	I <sub>3</sub>
5	H = 6 SP <sup>3</sup> d <sup>2</sup> Hybridisation	0	90°	Octahedral or Square Bipyramidal	Octahedral	SF <sub>6</sub>
		1			Square Pyramidal	XeOF <sub>4</sub>
		2			Square Planer	XeF <sub>4</sub>
6	H = 7 SP <sup>3</sup> d <sup>3</sup> Hybridisation	0	72°, 90°	Pentagonal Bipyramidal	Pentagonal Bipyramidal	IF <sub>7</sub>
		1			Distorted Octahedral	XeF <sub>6</sub>
7	H = 4, dSP <sup>2</sup>	0	90°	Square Planer	-	[Ni(CN) <sub>4</sub> ] <sup>2-</sup>

### Molecular Orbital Theory (MOT) :

VBT failed to explain the magnetic character of the molecule, stable existence of molecule, fractional bond order of the molecule. With the help of Molecular orbital theory above criteria of the molecule could be explained and this is why MOT (Molecular Orbital theory) was established. F. Hund and R. S. Mulliken established this MOT. [\*AO = Atomic Orbital].

#### Salient Features of MOT :

1. Just like atomic orbital a molecule has molecular orbital where electrons are present.
2. Molecular orbitals are formed by the combination of atomic orbital of proper symmetry and comparable energies.
3. No. of combined atomic orbitals is equal to the No. of molecular orbitals. Thus formed molecular orbitals are
  - i. Bonding Molecular Orbitals
  - ii. Anti-bonding Molecular Orbital
4. Bonding MO has lower energy and hence greater stability and anti-bonding MO has more energy and stability is less.
5. An electron in an atomic orbital is under the influence of only one nucleus but in case of molecule, an electron is under the influence of two or more nuclei. This is why atomic orbital is monocentric while molecular orbital is polycentric.
6. Just like atomic orbital molecular orbitals also follow the Aufbau Principle, Pauli's exclusion principle and Hund's rule.
7. Molecular orbitals are denoted by  $\sigma$ ,  $\sigma^*$ ,  $\pi$ ,  $\pi^*$ ,  $\delta$ ,  $\delta^*$  etc. Out of these  $\sigma$ ,  $\pi$  &  $\delta$  are the bonding MO and  $\sigma^*$ ,  $\pi^*$ ,  $\delta^*$  are the anti-bonding MO.



## Formation of Molecular orbitals (Linear Combination of Atomic orbitals LCAO)

According to wave mechanics, (1) the atomic orbitals of H atom can be expressed by  $\psi_A$  &  $\psi_B$  ()  
molecular orbital can be expressed by  $\psi_{MO}$

### 14. Probability of finding electron ( $\psi^2$ )

According to LCAO method, combination of atomic orbitals can take place in two ways.

i)  $\psi_{MO} = \psi_A + \psi_B$

ii)  $\psi_{MO} = \psi_A - \psi_B$

In case of (i), it is constructive interference and we get bonding MO and in case of (ii), It is destructive interference & we get anti bonding MO.

During constructive interference probability of finding electron in MO.

$$\psi_{MO}^2 = \psi_A^2 + \psi_B^2 + 2\psi_A \psi_B \text{ . (Squaring the both-side)}$$

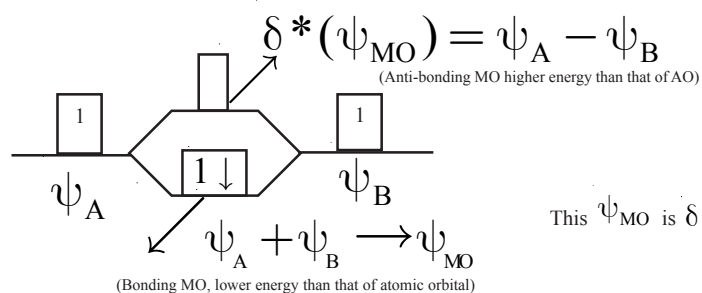
$$\psi_{MO}^2 > \psi_A^2 + \psi_B^2$$

i.e. probability of finding electron in MO is more than the atomic orbital.

During destructive interference, probability of finding electron is-

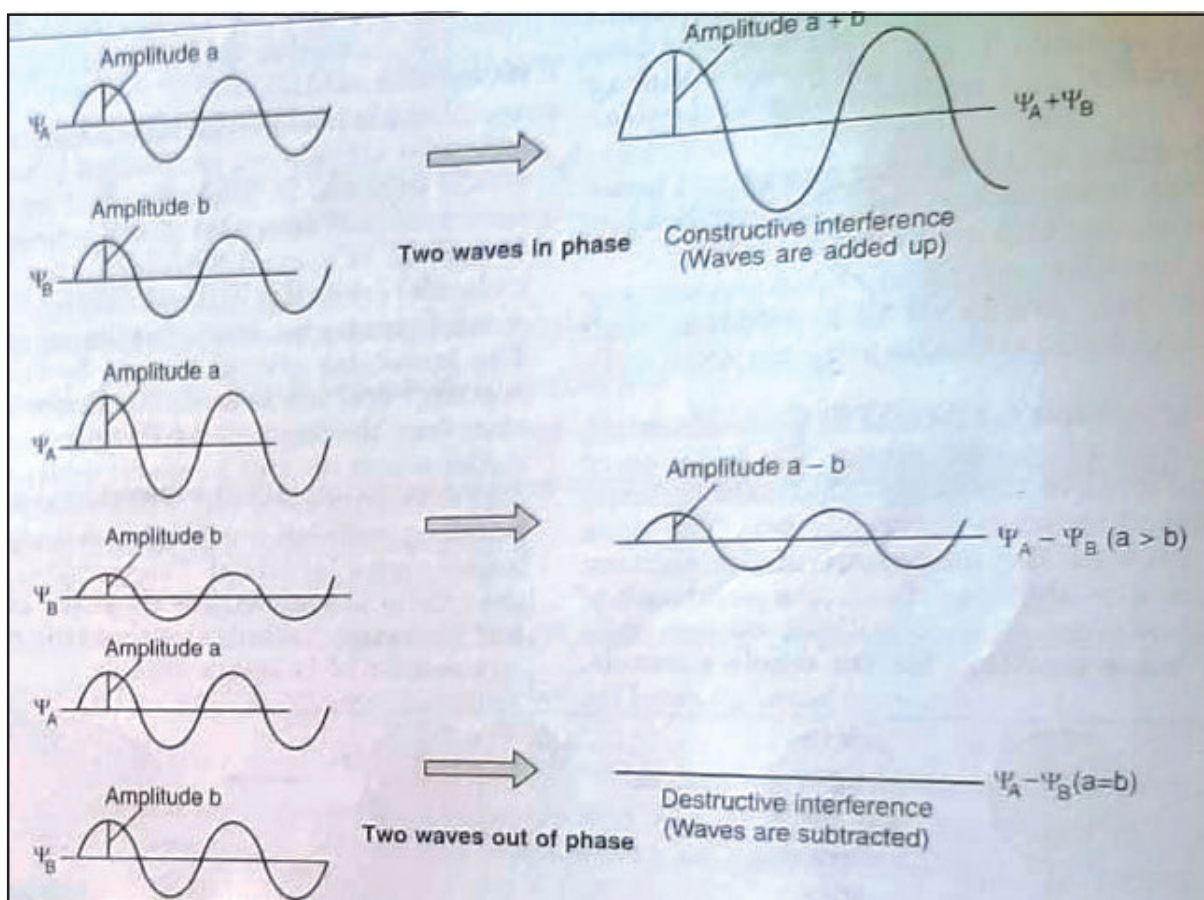
$$\psi_{MO}^2 = \psi_A^2 + \psi_B^2 - 2\psi_A \psi_B$$

i.e. probability of finding electron in MO is less than the atomic orbitals.



The formatin of bonding and anti bonding MO can be explained in terms of constructive and destructive interference of the electron wave of the combining atoms.



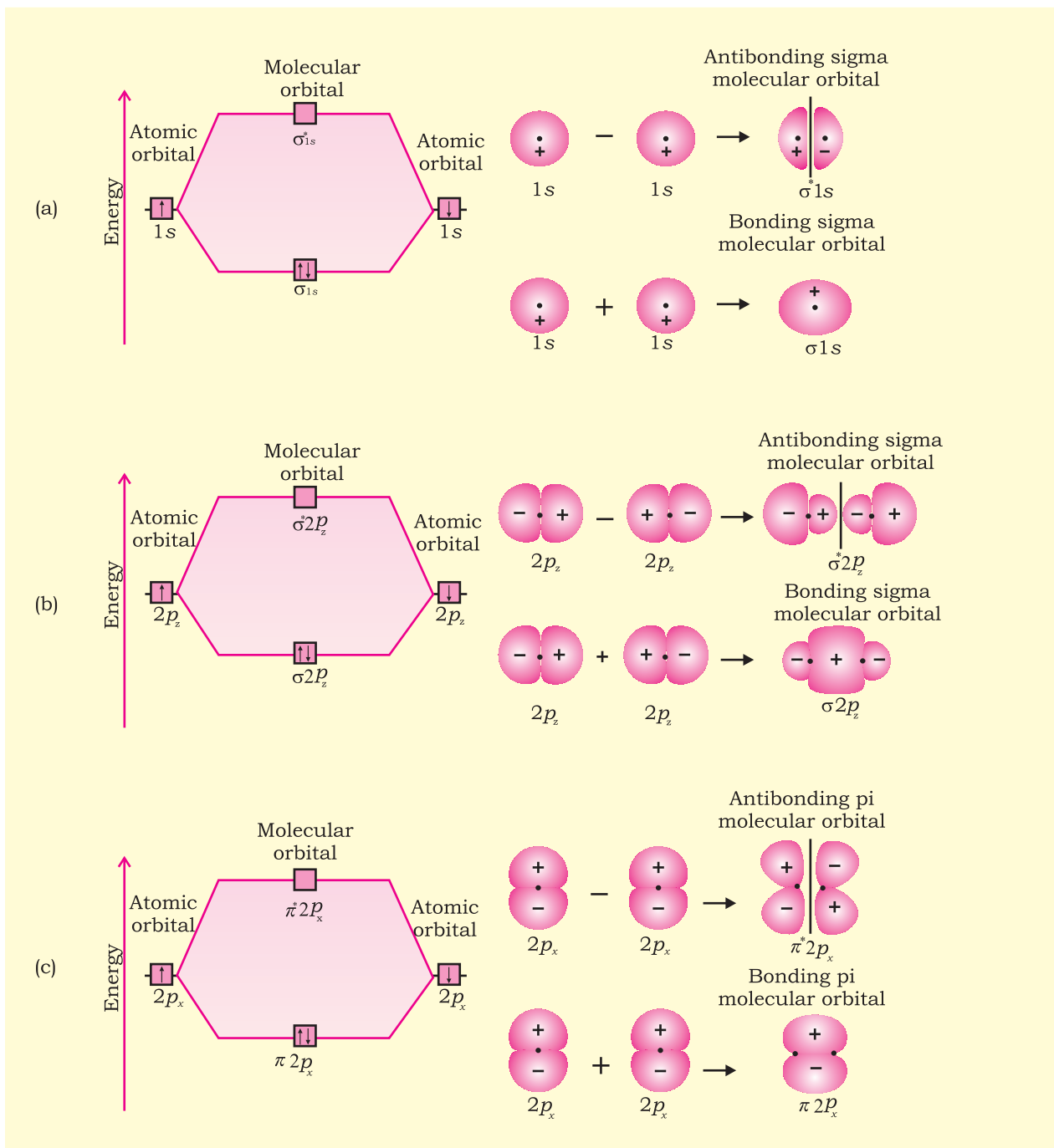


\* Conditions for the combination of Atomic orbitals

- 1) The combining atomic orbital must have the same or nearly the same energy.  
'IS' orbital combine with another 'IS' orbital not 2S.
- 2) The combining atomic orbitals must have the same symmetry about the molecular axis.  
2Pz orbital can combine with another 2Pz orbital but not with 2Px or 2Py orbital.
- 3) The combining atomic orbitals must overlap to the maximum extent.

### Type of Molecular Bond (MO)

- I.  $\delta$  (Sigma) MO:  $\delta$  MO are symmetrical around the bond axis while pi ( $\pi$ ) MO are not symmetrical
- II. If internuclear axis is taken to be in the Z-direction; it can be seen that a linear combination of 2Pz-orbitals of two atoms produce two sigma MO designated as  $\delta 2p_z$  and  $\delta^* 2p_z$ .



III. Energy level of MO for molecule having electron No. upto 14 are  $\delta 1s < \delta^* 1s < \delta 2s < \delta^* 2s < \pi 2p_x = \pi 2p_y < \delta 2p_z < \pi^* 2p_x = \pi^* 2p_y < \delta^* 2p_z$

Energy level of MO, having electron No. 16 onwards  $\delta_{1s} < \delta_{1s} < \delta_{2s} < \delta_{2s} < \pi_{2p_z} < \pi_{2p_x} = \pi_{2p_y} < \delta_{py}^* < \pi_{2p_x}^* = \pi_{2p_y}^* < \delta_{2p_z}^*$

**Note:** From MOT we can explain the following.

1. Bond order.
2. Magnetic Character of the molecule
3. Existence of a molecule
4. Stability of the molecule

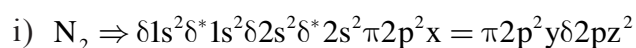
$$1. \text{ Bond Order} = \frac{\text{No. of bonding electron} - \text{No. of anti bonding electron}}{2}$$

i.e. 
$$\text{BO} = \frac{N_b - N_a}{2}$$

- a)  $N_b > N_a$  molecule is stable and exist.
- b)  $N_b = N_a$ , molecule does not exist.
- c)  $N_b < N_a$  molecule does not exist.

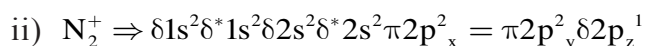
Higher the value of Bond order, more will be the stability of the molecule.

\*Electronic configuration of some important molecules and ions, according to energy level of MOT

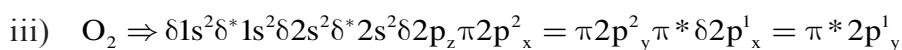


From the Electronic Conf<sup>x</sup> we can calculate the bond order (BO)

$$\text{BO} = \frac{10 - 4}{2} = 3 \text{ (molecule is diamagnetic)}$$

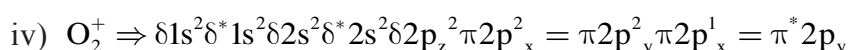


$$\text{BO} = \frac{9 - 4}{2} = 2.5 \text{ (Molecule is paramagnetic)}$$



$$\pi 2p_x^1 = \pi 2p_y^1$$

$$\text{BO} = \frac{10 - 6}{2} = 2 \text{ (Molecule is paramagnetic)}$$

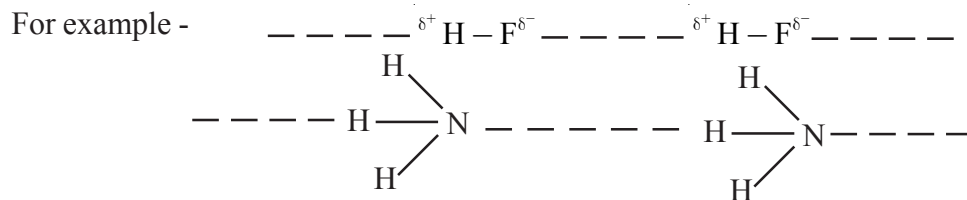


$$\text{BO} = \frac{10 - 5}{2} = 2.5 \text{ (Molecule is Paramagnetic)}$$

**Another trick for finding out the BO and magnetic character :**

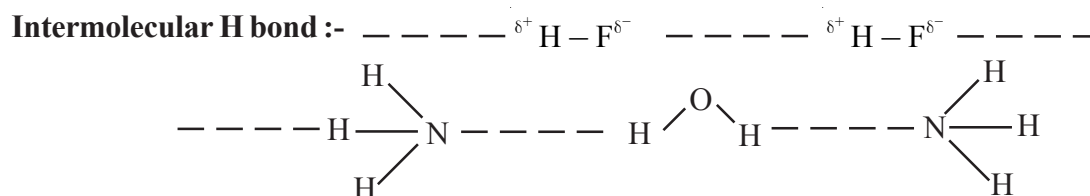
No. of electron in a molecule	BO	Magnetic Character	Example
10	1	Paramagnetic	$B_2$
11	1.5	Paramagnetic	$C_2^+$
12	2	Diamagnetic	$C_2$
13	2.5	Paramagnetic	$N$
14	3	Diamagnetic	$N_2^+$
15	2.5	Paramagnetic	$NO$
16	2	Paramagnetic	$O_2$
17	1.5	Paramagnetic	$O_2^-$
18	1	Diamagnetic	$O_2^{2-}$

Hydrogen -Bond : (i) Polar Co-valent-bond such as H-F, H-O, H-N i.e. H-atom linked with F, O or N any one of them are able to form H-Bond.



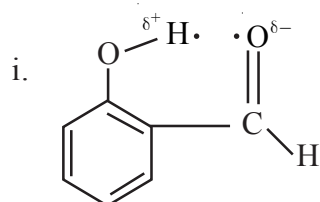
- \* 2. A more polar covalent bond will form stronger H bond. Such as, Among  
 i) F - H .....O & (ii) O - H.....F  
 1st one can form strong H - bond.
- \* 3. Hydrogen bond can be classified as -  
 a) Inter molecular H bond  
 b) Intra molecular H bond

Example:

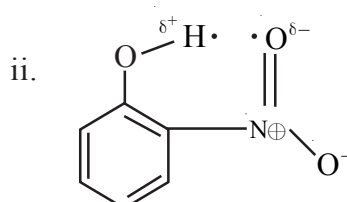


I & II are the example of intermolecular H bond. Due to intermolecular H bond molecules exist in associated form.

**Intramolecular H bond :-**



O-Hydroxy Benzaldehyde



O - Nitro Phenol

Due to intra moleuclar H bond molecules exist in discrete form.

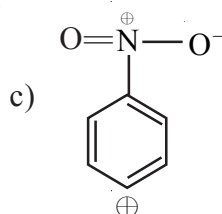
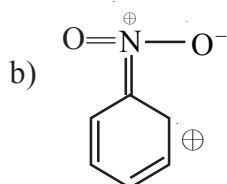
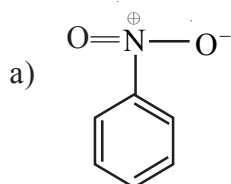
4. In H bond weak electrostatic force of attraction are acting between the molecules. i.e.  $\delta^+$  - of one polarmolecule and  $\delta^-$  of another polar molecules.
5. H bond is stronger than the vander waal force of attraction
6. With the help of H bond we can take the idea about the following properties of molecules.
- Physical state of a molecule :- such as  $H_2O$  is liquid but  $H_2S$  is gas.
  - BP of molecule: Such as BP of HF molecule is more than HCl molecule due to H bond etc.
  - Solubility of Covalent compound: For example,  $C_2H_5OH$  molecule can form H bond with  $H_2O$  molecule. Hence  $C_2H_5OH$  is soluble in water.

**A. Choose the correct answer (MCQ) :**

**(Each question mark 1)**

1. Non-directional compound is -  
a)  $\text{SiCl}_4$       b)  $\text{MgO}$       c)  $\text{SF}_6$       d)  $\text{POCl}_3$
2. Which one shows maximum covalent character  
a)  $\text{LiCl}$       b)  $\text{MgCl}_2$       c)  $\text{BaCl}_2$       d)  $\text{KCl}$
3. Among the following which one do not obey the octet rule  
a)  $\text{CH}_4$       b)  $\text{NH}_3$       c)  $\text{BF}_3$       d)  $\text{NF}_3$
4. Correct order of Lattice energy of compounds  $\text{Na}_2\text{O}$ ,  $\text{CaO}$ ,  $\text{MgO}$ ,  $\text{Al}_2\text{O}_3$  are  
a)  $\text{Na}_2\text{O} < \text{MgO} < \text{CaO} < \text{Al}_2\text{O}_3$   
b)  $\text{Na}_2\text{O} < \text{CaO} < \text{MgO} < \text{Al}_2\text{O}_3$   
c)  $\text{CaO} < \text{Na}_2\text{O} < \text{MgO} < \text{Al}_2\text{O}_3$   
d)  $\text{Al}_2\text{O}_3 < \text{CaO} < \text{MgO} < \text{Na}_2\text{O}$
5. Among the following correct order of Bond enthalpy -  
a)  $\text{F}_2 > \text{Cl}_2 > \text{Br} > \text{I}_2$       b)  $\text{Cl}_2 > \text{F}_2 > \text{Br}_2 > \text{I}_2$   
c)  $\text{Cl}_2 > \text{Br}_2 > \text{F}_2 > \text{I}_2$       d)  $\text{I}_2 > \text{F}_2 > \text{Cl}_2 > \text{Br}_2$
6. Among the following correct order of Bond Angle is  
a)  $\text{NH}_3 > \text{PH}_3 > \text{SbH}_3 > \text{AsH}_3 > \text{BiH}_3$   
b)  $\text{PH}_3 > \text{NH}_3 > \text{AsH}_3 > \text{SbH}_3 > \text{BiH}_3$   
c)  $\text{NH}_3 > \text{PH}_3 > \text{AsH}_3 > \text{SbH}_3 > \text{BiH}_3$   
d)  $\text{BiH}_3 > \text{AsH}_3 > \text{SbH}_3 > \text{NH}_3 > \text{PH}_3$

7. Which one is unlikely structure in the following -



d) all of these

8. Which one has zero dipole moment -
- a)  $\text{N}_2\text{O}$       b)  $\text{SO}_2$   
c)  $\text{CH}_3\text{Cl}$       d)  $\text{BI}_3$

9. In the following, planar molecule is -  
 a)  $\text{SF}_6$                                       b)  $\text{ClF}_3$   
 c)  $\text{SF}_4$                                         d)  $\text{CH}_4$
10. Among the following carbonate compounds which one can't produce  $\text{CO}_2(\text{g})$  on heating -  
 a)  $\text{Li}_2\text{CO}_3$                                     b)  $\text{CaCO}_3$   
 c)  $\text{ZnCO}_3$                                      d)  $\text{Na}_2\text{CO}_3$
11. Geometry and state of hybridisation of the  $\text{PCl}_5$  molecule is  
 a)  $\text{SP}^3\text{d}^2$ , Octahedral                      b)  $\text{SP}^3\text{d}$ , See-Saw  
 c)  $\text{SP}^3\text{d}$ , Trigonal bi pyramidal        d)  $\text{SP}^3\text{d}^2$  T-Shape
12. Shape and state of hybridisation of the molecule  $\text{XeF}_4$  is -  
 a)  $\text{SP}^3\text{d}$ , planar                              b)  $\text{SP}^3\text{d}^2$ , Octahedral  
 c)  $\text{SP}^3\text{d}^2$ , Square Planar                d)  $\text{SP}^3\text{d}^3$ , distorted Octahedral
13. In  $\text{SP}^3\text{d}$  hybridisation, which d orbital is involved ?  
 a)  $\text{dx}^2\text{-y}^2$                                     b)  $\text{dz}^2$                                         c)  $\text{dxy}$                                         d)  $\text{dyz}$
14. No. of  $\delta$  bond and  $\pi$  bond in tetra cyano ethene molecule is -  
 a)  $9\delta$  and  $9\pi$                                 b)  $5\delta$ ,  $1\pi$   
 c)  $9\delta$  and  $8\pi$                                 d)  $5\delta$  and  $5\pi$
15. No. of lp and bp in  $\text{SF}_4$  molecule around central atom are -  
 a) 1lp, 4bp                                    b) 2lp, 4bp                                    c) 3lp, 1bp                                    d) 2lp, 2bp
16. Which one is correct bond order in the following species  
 a)  $\text{O}_2^+ > \text{O}_2 > \text{O}_2^- > \text{O}_2^{2-}$                       b)  $\text{O}_2^{2-} > \text{O}_2^+ > \text{O}_2 > \text{O}_2^-$   
 c)  $\text{O}_2 > \text{O}_2^+ > \text{O}_2^- > \text{O}_2^{2-}$                       d)  $\text{O}_2 > \text{O}_2^- > \text{O}_2^{2-} > \text{O}_2^+$
17. The paramagnetic species is -  
 a)  $\text{BaO}_2$     b)  $\text{KO}_2$                                     c)  $\text{Na}_2\text{O}$                                         d) None

18. Among the following strongest H-bond is -  
 a) F - H.....O                      b) O - H .....F  
 c) N - H .....F                      d) O - H .....O
19. A water molecule can form maximum No. of Hydrogen bond -  
 a) 4            b) 3            c) 2            d) 1
20. Correct order of BP of compounds:  $\text{H}_2\text{O}_2$ ,  $\text{H}_2\text{O}$ , HF and  $\text{NH}_3$   
 a)  $\text{H}_2\text{O}_2 > \text{H}_2\text{O} > \text{HF} > \text{NH}_3$       b)  $\text{H}_2\text{O}_2 > \text{H}_2\text{O} > \text{HF} > \text{NH}_3$   
 c)  $\text{HF} > \text{H}_2\text{O}_2 > \text{H}_2\text{O} > \text{NH}_3$       d)  $\text{H}_2\text{O}_2 > \text{HF} > \text{H}_2\text{O} > \text{NH}_3$
21. Number of 2c-2e and 3c-2e bonds in  $\text{B}_2\text{H}_6$  molecule are  
 a) 3, 3            b) 4, 2            c) 2, 4            d) 2, 2
22. Number of  $\text{P}\pi - \text{P}\pi$  and  $\text{P}\pi - \text{d}\pi$  bonds in  $\text{SO}_2$  molecule are  
 a) 1, 1            b) 2, 2            c) 3, 1            d) 1, 2
23. In  $\text{PO}_4^{3-}$  ion, the formal charge on the oxygen atom of P - O bond is  
 a) +1            b) 0.75            c) - 0.75            d) -1
24. Which bond angle  $\theta$  would result in the maximum dipole moment for the tri atomic molecule  $\text{AB}_2$   
 a)  $\theta = 90^\circ$       b)  $\theta = 120^\circ$       c)  $\theta = 150^\circ$       d)  $\theta = 180^\circ$
25. The co valency of Nitrogen in  $\text{N}_2\text{O}_5$  molecule is -  
 a) 5            b) 4            c) 3            d) 2
26. Both  $\text{BF}_3$  and  $\text{NF}_3$  are covalent compounds  $\text{BF}_3$  is a non polar but  $\text{NF}_3$  is polar. The reason is that  
 a) Boron is solid and nitrogen is gas in free state.  
 b)  $\text{BF}_3$  is planar but  $\text{NF}_3$  is pyramidal m-shape  
 c) Boron is metalloid while nitrogen is a non metal  
 d) Atomic size of boron is smaller than that of nitrogen
27. The Electronegativity of H & Cl are 2.1 & 3 respectively. The correct statement(s) about the nature of HCl is are —  
 a) 17% ionic            b) 83% ionic            c) 50%            d) 100%

28. Which of the following have identical bond order?  $\text{CN}^-$ ,  $\text{O}_2^-$ ,  $\text{NO}^+$ ,  $\text{CN}^+$

- I)  $\text{CN}^-/\text{O}_2^-$     II)  $\text{O}_2^-/\text{NO}^+$     III)  $\text{NO}^+/\text{CN}^+$     IV)  $\text{NO}^+/\text{CN}^-$

29. In Which of the following pairs, the two species are not isosturctural?

- a)  $\text{PCl}_4$  and  $\text{SiCl}_4$     b)  $\text{PF}_5$  and  $\text{BrF}_5$   
c)  $\text{AlF}_6^{3-}$  and  $\text{SF}_6$     d)  $\text{CO}_3^{2-}$  and  $\text{NO}_3$

30. Correct order of BP among the hybrides of Gr-16 elements is

- a)  $\text{H}_2\text{O} > \text{H}_2\text{Te} > \text{H}_2\text{Se} > \text{H}_2\text{S}$   
b)  $\text{H}_2\text{Te} > \text{H}_2\text{O} > \text{H}_2\text{Se} > \text{H}_2\text{S}$   
c)  $\text{H}_2\text{O} > \text{H}_2\text{S} > \text{H}_2\text{Se} > \text{H}_2\text{Te}$   
d)  $\text{H}_2\text{Se} > \text{H}_2\text{O} > \text{H}_2\text{S} > \text{H}_2\text{Te}$

31. In which of the following pairs of molecules / ions, both the species are not likely to exist -

- a)  $\text{H}_2^-$ ,  $\text{He}_2^{2+}$     b)  $\text{H}_2^+$ ,  $\text{He}_2^{2-}$   
c)  $\text{H}_2^-$ ,  $\text{He}_2^{2-}$     d)  $\text{H}_2^{2+}$ ,  $\text{He}_2$

32. The percentage of p-character of the hybrid orbitals in graphite and diamond are respectively -

- a) 33 & 25    b) 50 & 75  
c) 67 & 75    d) 33 & 75

33. Match the following :

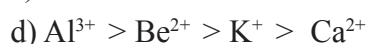
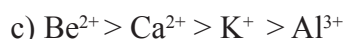
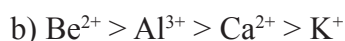
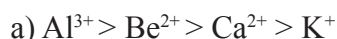
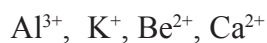
<b>List - I</b> <b>(Molecule)</b>	<b>List - II</b> <b>(No. of lp on central atom)</b>
A. $\text{NH}_3$	i) Two
B. $\text{H}_2\text{O}$	ii) Three
C. $\text{XCF}_2$	iii) Zero
D. $\text{CH}_4$	iv) Four
	v) one

**The correct answer is:**

- a) A - IV, B - I, C - III, D - IV  
b) A - III, B - I, C - II, D - V  
c) A - V, B - I, C - II, D - III  
d) A - I, B - V, C - III, D - IV



34. Correct increasing order of polarising power of the cationic species,



35. The ONO angle is maximum in -



### B. Assertion and Reason

(Each question 1 mark)

The questions given below consist of an assertion & reason (R) use the following key to choose the appropriate answer :

A) If both A & R are correct and R is the correct explanation of A.

B) If both A & R are correct but R is not correct explanation for A.

C) If A is correct but R is incorrect.

D) A is incorrect but R is correct.

1. Assertion (A) :  $\text{K}_2\text{SO}_4$  is soluble in water while  $\text{BaSO}_4$  is insoluble.

Reason (R) : Lattice energy of  $\text{BaSO}_4$  is more than the hydration energy.

2. A : Bond order of  $\text{H}_2^+$  and  $\text{H}_2^-$  are same

R :  $\text{H}_2$  and  $\text{H}_2^-$  are iso electronic to each other.

3. A : The shape of  $\text{SF}_4$  molecule is see saw

R : State of hybridisation of 'S' atom in  $\text{SF}_4$  molecule is  $\text{SP}^3\text{d}$ .

4. A : Dipole moment value (M) of CO is less than the expected value

R : Bond order in 'CO' molecule is 3

5. A :  $\text{O}_2$  is the better supporter of combustion than  $\text{N}_2\text{O}$ .

R :  $\text{N}_2\text{O}$  is linear in shape.

6. A : H - N - H Bond angle in  $\text{NH}_3$  is more than  $\text{NH}_4^+$

R : State of hybridisation of 'N' in  $\text{NH}_4^+$  is  $\text{SP}^3$ .

7. A : BP of H<sub>2</sub>O is more than HF.  
 R : HF is the ionic compound.
8. A : Ionic Compounds shows isomorphism property.  
 R : Ionic compounds are soluble in water.
9. A : Maximum co-valency of Phosphorus is +5  
 R : Solid PCl<sub>5</sub> exist as PCl<sub>6</sub><sup>-</sup> and PCl<sub>4</sub><sup>+</sup> ions.
10. A : SF<sub>6</sub> exist but SCl<sub>6</sub> does not exist.  
 R : 'S' has vacant 'd-orbital'.

**C. Very short answer type questions**

**(Each question 1 mark)**

- Name the 'd'-orbital which is involved in the dSP<sup>2</sup> hybridisation.
- Among NH<sub>3</sub> and NF<sub>3</sub>, which one has more Dipole moment?
- Among BF<sub>3</sub> and BI<sub>3</sub> which one is more acidic ?
- What is the shape of AsF<sub>5</sub> molecule using VSEPR model?
- Draw the Lewis dot structure of CaO.
- What is the bond order of N<sub>2</sub><sup>+</sup> molecule.
- Which one is stronger acid in the following pair: H<sub>2</sub>SO<sub>4</sub> and HClO<sub>4</sub>
- Which type H-bond is exist as associated form?
- Is there any change in the hybridisation of B&N atoms as a result of the following reactions?  

$$\text{NH}_3 + \text{BF}_3 \rightarrow \text{NH}_3\text{BF}_3$$

10. Give the example of a molecule having bond order zero.

**D. Short answer type questions:**

**(Each question 2 marks)**

1. Why are the axial bond longer as compared to equatorial bond ?
2. X - axis is the inter nuclear axis. - Why  $2p_y$  and  $2p_y$  do not form (sigma) bond?
3. Why  $C_2H_5Cl$  has more dipole moment than vinyl Chloride?
4. Define Electronegativity?
5. Why  $CH_4$  is not square planar?
6. Why  $\pi$  (pi) bond is weaker than  $\delta$  (sigma) bond?
7.  $KHF_2$  exist but  $KHCl_2$  does not exist. Why?
8. Among O - nitrophenol and P - nitro phenol, which one is more volatile? Explain?
9. Is H bond weaker or stronger than Vander Waal s forces ? Why?
10. What is the change in hybridisation of Al-atom in the following reaction.



**E. Long answer type questions:**

**(Each question 3 marks)**

1. a) What is Lattice energy?  
b) Arrange  $Na^+$ ,  $Li^+$ ,  $K^+$  in the increasing order of heat of hydration?
2. Write down the geometrical structure of  $I_3$ ,  $XeF_6$ ,  $BrF_5$ .
3. a) What is H - bond?  
b) Write the resonance structure of  $NO_2$  molecule.



3. a) Give the example of a compound where ionic, covalent and Co-ordinate bonds are present.  
b) Solubility of AgCl in water is less but NaCl is more soluble - Why?  
c) How many water molecule(s) are connected by H - bond in  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  molecule?
4. a) On the basis of hybridisation discuss the structures of -  
i)  $\text{PCl}_5$                       ii)  $\text{ICl}_2$                       iii)  $\text{XeO}_2\text{F}_2$   
b) Is it correct to say that bond order always increases when an electron is lost?
5. a) Sketch the bond moments and resultant dipole moment in the following molecules  
 $\text{POCl}_3$ ,                       $\text{NH}_3$ ,                       $\text{CO}_2$   
b) Name the molecule having polar bonds but dipole moment is zero.  
c)  $\text{N}_2$  and  $\text{N}_2^+$  which one is stable?

## Solution

**A.**

1-b, 2-a, 3-c, 4-b, 5-c, 6-b, 7-d, 8-d, 9-b, 10-d,

11-c, 12-c, 13-b, 14-a, 15-a, 16-a, 17-b, 18-b, 19-a, 20-a,

21-b, 22-a, 23-c, 24-a, 25-b, 26-b, 27-a, 28-d, 29-b, 30-a,

31-d, 32-c, 33-d.

**B.**

1-a, 2-c, 3-a, 4-a, 5-d, 6-d, 7-c, 8-b, 9-d, 10-b

## Chapter -5th

# States of Matter

### Chapter at a glance :-

1. Chemical properties of a substance do not change with the change of its physical state; but rate of chemical reactions depend upon the physical state. Many times for experimental calculation we require knowledge of the state of matter. Therefore, it becomes necessary for a chemist to know the physical laws which govern the behavior of matter in different states. In this chapter, we will deal with the physical states of matter particularly liquid and gaseous states.
2. It is necessary to understand the nature of intermolecular forces, molecular interactions and effect of thermal energy on the motion of particles because a balance between these determines the state of a substance.

### 3. Intermolecular Forces :-

Intermolecular forces are the forces of attraction and repulsion between interacting particles (atoms and molecules). These are the weak forces and generally affect the physical properties and some chemical properties of matter.

4. Attractive intermolecular forces are known as Vander Waals forces. We will now learn about different types of Vander Waals forces.

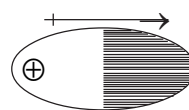
#### (i) Dispersion forces or London forces :-

Atoms and non-polar molecules are electrically symmetrical and have no dipole moment because their electronic charge cloud is symmetrically distributed. But a dipole may develop momentarily even in such atoms and molecules.

Suppose we have two atoms 'A' and 'B' in the close vicinity of each other. It may so happen that momentarily electronic charge distribution in one of the atoms, say 'A', becomes unsymmetrical i.e., the charge cloud is more on one side than the other. This results in the development of instantaneous dipole on the atom 'A' for a very short time. This instantaneous dipole distorts the electron density of the other atom 'B', which is close to it and as a consequence a dipole is induced in the atom 'B'. The force of attraction between two temporary dipoles is known as London force.



Atom 'A' with instantaneous dipole, more electron density on the right hand side.



Atom 'B' with induced dipole.



Atom 'A' more electron density on the left hand side.



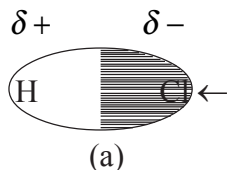
Atom 'B' with induced dipole.

**Fig :** London forces between atoms.

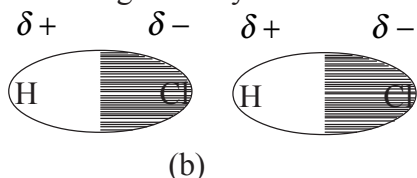
These forces are important only at short distances ( $\sim 500\text{pm}$ ) and their magnitude depends on the polarisability of the particle.

**(ii) Dipole-dipole forces :-** Dipole-dipole forces act between the molecules possessing permanent dipole. Ends of the dipoles possess “partial charges” and these charges are shown by Greek Letter delta ( $\delta$ ). Partial charge are always less than the unit electronic charge ( $1.6 \times 10^{-19}\text{C}$ ).

The polar molecules interact with neighbouring molecules. Dipole-dipole interaction between two *HCl* molecules are shown below.



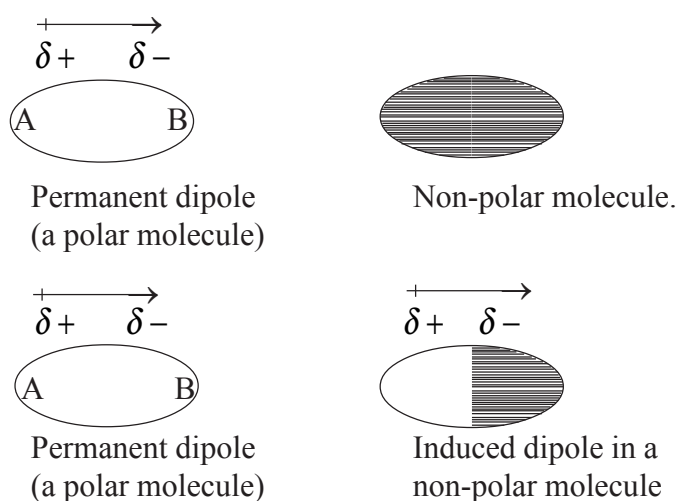
More charge density towards chlorine.



**Fig :** Distribution of electron cloud in *HCl* a polar molecule. Dipole-dipole interaction between two *HCl* molecules.

**(iii) Dipole-Induced Dipole Forces :-**

This type of attractive forces operate between the polar molecules having permanent dipole and the molecules lacking permanent dipole. Permanent dipole of the polar molecule induces dipole on the electrically neutral molecule by deforming its electronic cloud. Thus an induced dipole is developed in the other molecule. Induced dipole moment depends upon the dipole moment present in the permanent dipole and the polarisability of the electrically neutral molecule. High polarisability increases the strength of attractive interactions.

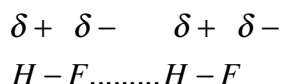


**Fig. :** Dipole-induced dipole interaction between permanent dipole and induced dipole.



**(iv) Hydrogen bond :-**

This is the special case of dipole-dipole interaction. This is found in the molecules in which highly polar  $N-H$ ,  $O-H$  or  $H-F$  bonds are present. Hydrogen bonding is regarded as being limited to  $N$ ,  $O$  and  $F$ . Hydrogen bonds are powerful force in determining the structure and properties of many compounds. Following diagram shows the formation of hydrogen bond.



5. Molecules also exert repulsive forces on one another. When two molecules are brought into close contact with each other, the repulsion between the electron cloud and that between the nuclei of two molecules comes into play. Magnitude of the repulsion rises very rapidly as the distance separating the molecules decreases.

**6. Thermal Energy :-**

Thermal energy is the energy of a body arising from motion of its atoms or molecules. It is directly proportional to the temperature of the substance. It is the measure of average kinetic energy of the particles of the matter and is thus responsible for movement of particles.

**7. Intermolecular forces Vs thermal interactions :-**

The intermolecular forces tend to keep the molecules together but thermal energy of the molecules tend to keep them apart. Three states of matter are the result of balance between intermolecular forces and the thermal energy of the molecules.

- (a) In a solid, the intermolecular forces predominate over the thermal energy, that's why particles are held together in rigid and close packed structure.
- (b) In Liquids, the intermolecular forces are no longer strong enough, but these are still enough so that particles remain in each other's environment.
- (c) In gases, the thermal energy predominate over intermolecular forces, thus, the gas molecules acquire the unrestricted and independent mobility in the vapour state.

**8. The gaseous state :-**

This is the simplest state of matter. Throughout our life we remain immersed in the ocean of air which is a mixture of gases. We spend our life in the lowermost layer of the atmosphere called troposphere, which is held to the surface of the earth by gravitational force. The thin layer of atmosphere is vital to our life. It shields us from harmful radiations and contains substances like dioxygen, dinitrogen, carbon dioxide, water vapour etc.

**9. The gaseous state is characterized by following physical properties.**

- (i) Gases are highly compressible.
- (ii) Gases exert pressure equally in all directions.
- (iii) Gases have much lower density than the solids and liquids.
- (iv) The volume of gases are not fixed. They assume volume and shape of the container.
- (v) Gases mix evenly and completely in all proportions without any mechanical aid.

## 10. Gas Laws :-

The gas laws which we will study now are the result of research carried on for several centuries on the physical properties of gases.

### (i) Boyle's Law :-

At constant temperature, the pressure of a fixed amount of gas varies inversely with its volume. This is known as Boyle's law.

Mathematical expression :

$$P \propto \frac{1}{V} \quad \text{at constant temperature } T,$$

$$\text{Or, } P = K \cdot \frac{1}{V} \quad \text{Here, } P = \text{Pressure,}$$

$$\text{Or, } \boxed{PV = K} \quad \begin{array}{l} V = \text{Volume} \\ K = \text{Constant.} \end{array}$$

If a fixed amount of gas at constant temperature  $T$ , with volume  $V_1$  at pressure  $P_1$  undergoes expansion, so that volume becomes  $V_2$  and pressure becomes  $P_2$ , then according to Boyle's Law :-

$$P_1 V_1 = P_2 V_2 = \text{Constant,}$$

$$\text{Or, } \frac{P_1}{P_2} = \frac{V_2}{V_1}$$

- Experiments of Boyle, in a quantitative manner prove that gases are highly compressible because when a given mass of a gas is compressed, the same number of molecules occupy a smaller space. This means that gases become denser at high pressure. A relationship can be obtained between density and pressure of a gas by using Boyle's Law :

According to Boyle's Law,  $P = K \cdot \frac{1}{V} \rightarrow (1)$

$$\text{Again, density } (d) = \frac{\text{mass}(m)}{\text{volume}(v)}$$

$$\text{Or, } V = \frac{m}{d}$$

Now, from equation (1),  $P = k \cdot \frac{1}{m/d}$

$$\text{Or, } P = k \cdot \frac{d}{m}$$

$$\text{Or, } P = \frac{k}{m} \cdot d$$

$$\text{Or, } P = k^1 \cdot d \quad \left| \text{Here, } \frac{k}{m} = \text{Constant} = k^1 \right.$$

$\therefore$  We can write,  $p \propto d$

**(ii) Charles' Law :-**

For a fixed mass of gas at constant pressure volume of a gas increases on increasing temperature and decreases on cooling. Charles' and Gay Lussac found that for each degree rise in temperature,

volume of a gas increases by  $\frac{1}{273.15}$  of the original volume of the gas at  $0^\circ\text{C}$ . Thus if volumes of

the gas at  $0^\circ\text{C}$  and at  $t^\circ\text{C}$  are  $V_0$  and  $V_t$  respectively, then,

$$V_t = V_0 + \frac{t}{273.15} V_0$$

$$\text{Or, } V_t = V_0 \left( 1 + \frac{t}{273.15} \right)$$

$$\text{Or, } \boxed{V_t = V_0 \left( \frac{273.15 + t}{273.15} \right)}$$

At this stage, we define a new scale of temperature such that  $t^\circ\text{C}$  on new scale is given by  $T = 273.15 + t$  and  $0^\circ\text{C}$  will be given by  $T_0 = 273.15$ . This new temperature scale is called the Kelvin temperature scale or absolute temperature scale.

If a given quantity of gas is held at constant pressure, its volume is directly proportional to the absolute temperature.

i.e.,  $V \propto T$ ,

$$\text{Or, } V = KT$$

$$\text{Or, } \frac{V}{T} = K$$

= Constant

Here,  
V = Volume  
T = Temp.

For two different type of gases,

$$\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$

Here,  
 $V_1 = \text{Volume of 1st gas.}$   
 $T_1 = \text{Temp. of 1st gas.}$   
 $V_2 = \text{Volume of 2nd gas.}$   
 $T_2 = \text{Temp. of 2nd gas.}$

### Graph of Volume Vs temperature :

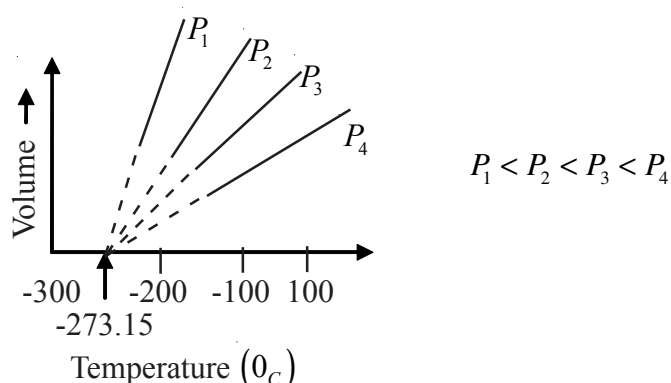


Fig. : Volume Vs Temperature ( $0_c$ ) graph

### (iii) Gay Lussac's Law :-

The mathematical relationship between pressure and temperature was given by Joseph Gay Lussac and is known as Gay Lussac's Law. It states that at constant volume, pressure of a fixed amount of a gas varies directly with the temperature.

Mathematically,

$$P \propto T$$

Or,  $P = KT$

Here,

$K = \text{Constant}$

$P = \text{Pressure}$

$T = \text{Temperature}$

### (iv) Avogadro Law :-

It states that equal volumes of all gases under the same conditions of temperature and pressure contain equal number of molecules. Mathematically we can write,

$$V \propto n$$

Where  $n$  is the number of moles of the gas.

Or,  $V = k.n$

## 11. Ideal gas equation :

The three laws which we have learnt till now can be combined together in a single equation which is known as ideal gas equation.

At Constant  $T$  and  $n$ ,  $V \propto \frac{1}{p}$  (Boyle's Law)

At Constant  $P$  and  $n$ ,  $V \propto T$  (Charle's Law)

At Constant  $P$  and  $T$ ,  $V \propto n$  (Avogadro's Law)

Thus,  $V \propto \frac{nT}{P}$

Or,  $V = R \cdot \frac{nT}{P}$

Or,  $PV = RnT$

Where  $R$  is Proportionality constant.

$\therefore$  We can write,  $PV = nRT$

Or,  $R = \frac{PV}{nT}$

$R$  is called gas constant. It is same for all gases. Therefore it is also called universal gas constant.

### 12. Values of Gas Constant 'R' :-

Value of 'R' for one mole of an ideal gas can be calculated under the conditions as follows-

$$R = \frac{10^5 \text{ Pa} \times 22.71 \times 10^{-3} \text{ m}^3}{1 \text{ mol} \times 273.15 \text{ K}}$$

$$\therefore R = \frac{PV}{nT}$$

Here,

$$P = 10^5 \text{ Pa}$$

$$V = 22.71 \times 10^{-3} \text{ m}^3$$

$$T = 273.15 \text{ K}$$

Or,  $R = 8.314 \text{ Pa m}^3 \text{ k}^{-1} \text{ mol}^{-1}$

Or,  $R = 8.314 \text{ Jk}^{-1} \text{ mol}^{-1}$

At STP,  $R = 0.0821 \text{ L.atm k}^{-1} \text{ mol}^{-1}$

And,  $R = 1.98 \text{ Cal k}^{-1} \text{ mol}^{-1}$

### 13. Relation between density and molecular mass of a gaseous substance :-

Ideal gas equation, can be rearrange as follows,

$$\frac{n}{v} = \frac{P}{RT} \quad \rightarrow (1)$$

Here,  $n$  = number of moles =  $\frac{W}{M}$

$W$  = given Mass

$M$  = Molecular Mass

Again, Volume ( $V$ ) =  $\frac{\text{mass}(W)}{\text{density}(d)}$

Now, from equation (1), 
$$\frac{\frac{W}{M}}{\frac{W}{d}} = \frac{P}{RT}$$

Or, 
$$\frac{d}{M} = \frac{P}{RT}$$

Or, 
$$d = \frac{PM}{RT}$$

**14. Dalton's Law of Partial Pressure :-**

Dalton's law states that the total pressure exerted by the mixture of non-reactive gases is equal to the sum of the partial pressure of individual gases.

**15. Partial Pressure :-**

In a mixture of gases, the pressure exerted by the individual gas is called partial pressure.

Mathematically,

$$P_{Total} = P_1 + P_2 + P_3 + \dots \dots \dots \text{(at constant } T, V)$$

Here,  $P_{Total}$  = Total pressure exerted by the gas.

$P_1, P_2, P_3$  etc. are partial pressures of gases.

**16. Aqueous tension :-**

Pressure exerted by saturated water vapour is called aqueous tension.

$$P_{Dry\ gas} = P_{Total} - \text{Aqueous tension}$$

**17. With the help of ideal gas equation and Dalton's law of partial pressure we can easily determine the relation between partial pressure and mole fraction.**

Mathematically,

$$P_i = P \times x_i$$

Here,  
 $P_i$  = Partial Pressure  
 $P$  = Total Fraction  
 $x_i$  = Mole Fraction

**18. Kinetic energy and Molecular Speed :-**

Molecules of gases remain in continuous motion. While moving they collide with each other and with the walls of the container. This results in change of their speed and redistribution of energy. So the speed and energy of all the molecules of the gas at any instant are not the same.

(i) Average Speed : ( $U_{av}$ )

If there are  $n$ -number of molecules in a sample and their individual speeds are  $u_1, u_2, \dots, u_n$  then average

speed of molecules  $U_{av}$  can be calculated as follows –

$$U_{av} = \frac{u_1 + u_2 + \dots + u_n}{n}$$

(ii) Most Probable Speed : ( $U_{mp}$ )

The speed which is possessed by maximum number of gas molecules at a given temperature are known as most probable speed. This is very close to the average speed of the molecules. On increasing the temperature most probable speed increases.

(iii) Root mean square is the direct measure of the average kinetic energy of gas molecules. If we take the square root of the mean of the square of speeds then we get a value of speed which is different from most probable speed and average speed. This speed is called root mean square speed and is given by the expression as follows –

$$U_{rms} = \sqrt{\bar{u}^2}$$

Here, Mean square of speed,  $\bar{u}^2 = \frac{u_1^2 + u_2^2 + \dots + u_n^2}{n}$

**19.** Root mean square speed, average speed and the most probable speed have following relationship

$$U_{rms} > U_{av} > U_{mp}$$

The ratio between the three speeds is given below :-  $U_{rms} : U_{av} : U_{mp} = 1 : 1.128 : 1.224$

**20.** On the basis of postulates of kinetic theory of gases, the following gas equation is derived.

$$PV = \frac{1}{3} mnu^2$$

Here, P = Pressure exerted by the gas

V = Volume of the gas

**21. The different velocities possessed by the gas molecules are :**

(i) Average velocity ( $U_{av}$ ) =  $\sqrt{\frac{8RT}{\pi M}}$

(ii) Most probable velocity ( $U_{mp}$ ) =  $\sqrt{\frac{2RT}{M}}$

(iii) Root mean square velocity ( $U_{rms}$ ) =  $\sqrt{\frac{3RT}{M}}$

Here, M = Molecular Mass.

R = Gas Constant

T = Temperature.

## 22. Kinetic Molecular theory of gases :

Assumptions or postulates of the kinetic molecular theory of gases are given below. These postulates are related to atoms and molecules which cannot be seen, hence it is said to provide a microscopic model of gases.

(i) Gases consist of large number of identical particles that are so small and so far apart on the average that the actual volume of the molecules is negligible in comparison to the entire volume of the container. They are considered as point masses. This assumption explains the great compressibility of gases.

(ii) The intermolecular forces of attraction between gaseous molecules are negligible.

(iii) Particles of gas are always in constant and random motion.

(iv) Particles of a gas move in all possible directions in straight lines. During their random motion, they collide with each other and with the walls of the container.

(v) Collisions of gas molecules are perfectly elastic. This means that total energy of molecules before and after the collision remains same. If there were loss of kinetic energy, the motion of molecules will stop and gases will settle down.

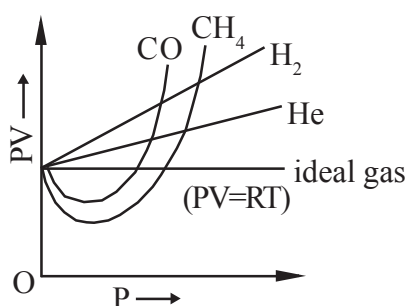
(vi) At any particular time, different particles in the gas have different speeds and hence different kinetic energies.

(vii) If a molecule has variable speed, then it must have a variable kinetic energy. Under these circumstances, we can talk only about average kinetic energy.

Kinetic theory of gases allows us to derive theoretically, all the gas laws studied in the previous sections.

## 23. Behaviour of Real gases : (Deviation from ideal gas behaviour)

Ideal gases are hypothetical gases whose pressure, volume and temperature behaviour can be completely accounted by ideal gas equation. If we plot a graph PV vs P of gases at constant temperature, PV will be constant.



**Fig :** Plot of PV vs P for real gas and ideal gas.

It can be seen easily that at constant temperature PV vs P plot for real gases is not a straight line. There is a significant deviation from ideal behaviour. Two types of curves are seen. In the curves of H<sub>2</sub> and He, as the pressure increases the value of PV also increases. In second type of plot first there is a negative deviation from ideal behaviour, the PV value decreases with increasing in pressure



and reaches to a minimum value characteristic of a gas.

24. It is found that real gases do not follow, Boyle's Law, Charles Law and Avogadro Law perfectly under all conditions.
25. Real gases show deviations from ideal gas Law because molecules interact with each other. At high pressure molecules of gases are very close to each other.
26. The pressure exerted by the gas is lower than the pressure exerted by the ideal gas.

$$P_{\text{ideal}} = P_{\text{real}} + \frac{an^2}{V^2}$$

Observed Pressure      Correction term

27. The volume occupied by the molecules also becomes significant because instead of moving in volume  $V$ , these are now restricted to volume  $(V - nb)$  where  $nb$  is approximately the total volume occupied by the molecules themselves.
28. This Van der Waal's equation explains the behaviour of a real gas correctly and is given as,

$$\left( P + \frac{n^2 a}{V^2} \right) (V - nb) = nRT$$

Here,  $a$  and  $b$  are called Van der Waals' constants.

29. The deviation from ideal behaviour can be measured in terms of compressibility factor  $Z$ , which is the ratio of product  $PV$  and  $nRT$ . Mathematically,

$$Z = \frac{PV}{nRT}$$

- \* For ideal gas  $Z = 1$  at all temperature and pressure.
- \* In case of real gas,  $PV \neq nRT$ ,  $Z \neq 1$
- \* At high pressure all the gases have  $Z > 1$ . (gas is less compressible)
- \* At intermediate pressures, most gases have  $Z < 1$ . (gas is more compressible)

### 30. Liquefaction of gases :-

For a certain gas to liquefy, there is maximum limit of temperature and minimum limit of pressure. These limits are called critical constants.

- (i) Critical Temperature ( $T_c$ )
- (ii) Critical Pressure ( $P_c$ )
- (iii) Critical Volume ( $V_c$ )

\* Relation between critical constant and Van der Waals constant :

(a)  $T_c = \frac{8a}{27Rb}$

$$(b) P_c = \frac{a}{27b^2}$$

$$(c) V_c = 3b$$

### 31. Liquid State :-

Intermolecular forces are stronger in liquid state than in gaseous state. Molecules in liquids are so close that there is very little empty space between them and under normal conditions liquids are denser than gases. In the following sections we will look into some of the physical properties of the liquids such as vapour pressure, surface tension and viscosity.

**(i) Vapour Pressure :-** If an evacuated container is partially filled with a liquid, a portion of liquid evaporates to fill the remaining volume of the container with vapour. Initially the liquid evaporates and pressure exerted by vapours on the walls of the container increases. After some time it becomes stable vapour pressure at this stage is known as equilibrium vapour pressure or saturated vapour pressure. The process of vapourisation is temperature dependent.

**(ii) Surface tension:-** It is well known fact that liquids assume the shape of the container. Surface tension is defined as the force acting per unit length perpendicular to the line drawn on the surface of liquid. It is denoted by Greek letter  $\gamma$ .

The magnitude of surface tension of a liquid depends on the attractive forces between the molecules. When the attractive forces are large, the surface tension is large. Increase in temperature increases the kinetic energy of the molecules and effectiveness of intermolecular attraction decreases, so surface tension decreases as the temperature is raised.

**(iii) Viscosity :-** Viscosity is a measure of resistance to flow which arises due to the internal friction between layers of fluid as they slip past one another while liquid flows.

Viscosity of liquids decreases as the temperature rises because at high temperature molecules have high kinetic energy and can overcome the intermolecular forces to slip past one another between the layers.

### 32. Graham's Law of diffusion :-

Graham's law states that the rate of diffusion of a gas is inversely proportional to the square root of density.

$$\text{i.e., } r \propto \frac{1}{\sqrt{d}}; \quad \left| \begin{array}{l} \text{Here, } r = \text{Rate of diffusion} \\ d = \text{density} \end{array} \right.$$

$\therefore$  For two different type of gases.

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}$$

\* Relation between molecular mass and rate of diffusion :

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \quad \left| \begin{array}{l} \text{Here, } M_1 \text{ and} \\ M_2 \text{ are the} \\ \text{Molecular Mass.} \end{array} \right.$$

\* Relation between volume, time and rate of diffusion :

$$r = \frac{V}{t} \quad \left| \begin{array}{l} \text{Here, } V = \text{Volume} \\ t = \text{time} \end{array} \right.$$

[A] Multiple Choice Questions :- (MCQ)

(Each question mark 1)

Choose the correct options :-

- Dipole - induced dipole interactions are present in which of the following pairs?  
(a)  $HCl$  and  $He$  atoms (b)  $SiF_4$  and  $He$  atoms.  
(c)  $H_2O$  and Alcohol (d)  $Cl_2$  and  $CCl_4$
- What is the density of  $N_2$  gas at  $227^\circ C$  and  $5 \text{ atm}$  pressure? ( $R = 0.082 \text{ L.atm k}^{-1} \text{ mol}^{-1}$ )  
(a)  $1.40 \text{ g/ml}$  (b)  $2.81 \text{ g/ml}$   
(c)  $3.41 \text{ g/ml}$  (d)  $0.29 \text{ g/ml}$
- A certain gas takes three times as long to effuse out as helium. Its molecular mass will be –  
(a)  $27 u$  (b)  $26 u$  (c)  $64 u$  (d)  $9 u$
- Which of the following mixtures of gases does not obey Dalton's Law of partial pressure?  
(a)  $Cl_2$  and  $SO_2$  (b)  $CO_2$  and  $He$   
(c)  $O_2$  and  $CO_2$  (d)  $N_2$  and  $D_2$
- At constant temperature, in a given mass of an ideal gas —  
(a) The ratio of pressure and volume always remains constant.  
(b) Volume always remains constant.  
(c) Pressure always remains constant.  
(d) The product of pressure and volume always remains constant.
- By what factor does the average velocity of a gaseous molecule increase when the temperature (in kelvin) is doubled?  
(a) 2.0 (b) 2.8 (c) 4.0 (d) 1.4
- The temperature of a gas is raised from  $27^\circ C$  to  $927^\circ C$ . The root mean square speed of the gas-

- (a) remains same      (b) gets  $\sqrt{\frac{927}{27}}$  times  
(c) gets halved      (d) gets doubled.

8. Internal energy and pressure of a gas per unit volume are related as —

- (a)  $P = \frac{2}{3} E$       (b)  $P = \frac{3}{2} E$       (c)  $P = \frac{1}{2} E$       (d)  $P = 2E$

9. Maximum deviation from ideal gas is expected from —

- (a)  $CH_4$       (b)  $NH_3(g)$       (c)  $H_2(g)$       (d)  $N_2(g)$

10. Van der Waal's real gas, acts as an ideal gas at which conditions?

- (a) High temperature, low pressure,  
(b) Low temperature, high pressure,  
(c) High temperature, high pressure,  
(d) Low temperature, low pressure.

11. In Van der Waal's equation of state for a non ideal gas, the term that accounts for intermolecular forces is —

- (a)  $(V - b)$  (b)  $(RT)^{-1}$  (c)  $\left(P + \frac{a}{v^2}\right)$  (d)  $RT$

12. Which is not true in case of an ideal gas?

- (a) It cannot be converted into a liquid.  
(b) There is no interaction between the molecules.  
(c) All molecules of the gas move with same speed.  
(d) At a given temperature,  $PV$  is proportional to the amount of the gas.

13. A gas such as carbon monoxide would be most likely to obey the ideal gas law at —

- (a) Low temperature and high pressures.  
(b) High temperature and high pressures.  
(c) Low temperature and low pressures.  
(d) high temperature and low pressures.

14. The volume occupied by 1.8g of water vapour at  $374^{\circ}\text{C}$  and 1 bar pressure will be —  
(Use  $R = 0.083 \text{ bar L k}^{-1}\text{mol}^{-1}$ )  
(a) 96.66 L (b) 55.87 L (c) 3.10 L (d) 5.37 L
15. Two gases  $A$  and  $B$  having the same volume diffuse through a porous partition in 20 and 10 seconds respectively. The molecular mass of  $A$  is  $49u$ . Molecular mass of  $B$  will be —  
(a)  $50u$  (b)  $12.25u$  (c)  $6.50u$  (d)  $25.00u$
16. The pressure exerted by 6.0g of methane gas in a  $0.03\text{m}^3$  vessel at  $129^{\circ}\text{C}$  is (Atomic masses :  $C = 12.01, H = 1.01$  and  $R = 8.314 \text{ Jk}^{-1}\text{mol}^{-1}$ )  
(a) 215216 Pa (b) 13409 Pa (c) 41648 Pa (d) 31684 Pa.
17. If a gas expands at constant temperature, it indicates that —  
(a) Kinetic energy of molecules remains the same.  
(b) Number of the molecules of gas increases.  
(c) Kinetic energy of molecules decreases.  
(d) Pressure of the gas increases.
18. The average kinetic energy of an ideal gas, per molecule in S.I. units, at  $25^{\circ}\text{C}$  will be —  
(a)  $6.17 \times 10^{-20} \text{ J}$  (b)  $7.16 \times 10^{-20} \text{ J}$   
(c)  $61.7 \times 10^{-20} \text{ J}$  (d)  $6.17 \times 10^{-21} \text{ J}$
19. The ratio of masses of oxygen and nitrogen in a particular gaseous mixture is 1 : 4. The ratio of number of their molecules is —  
(a) 1 : 4 (b) 7 : 32 (c) 1 : 8 (d) 3 : 16
20. At very high pressures, the compressibility factor of one mole of a gas is given by —  
(a)  $1 + \frac{Pb}{RT}$  (b)  $\frac{Pb}{RT}$  (c)  $1 - \frac{Pb}{RT}$  (d)  $1 - \frac{b}{VRT}$
21. If ' $Z$ ' is a compressibility factor, Van der Waals equation at low pressure can be written as :

$$(a) Z = 1 + \frac{RT}{Pb} \quad (b) Z = 1 - \frac{a}{VRT}$$

$$(c) Z = 1 - \frac{Pb}{RT} \quad (d) Z = 1 + \frac{Pb}{RT}$$

22. The temperature at which oxygen molecules have the same root mean square speed as helium atoms have at  $300K$  is :

(Atomic masses :  $He = 4u, O = 16u$ )

- (a)  $300K$       (b)  $600K$     (c)  $1200K$    (d)  $2400K$

23. Assuming ideal gas behaviour, the ratio of density of ammonia to that of hydrogen chloride at same temperature and pressure is : (Atomic wt. of  $Cl = 35.5u$ )

- (a) 1.46      (b) 1.64    (c) 0.46    (d) 0.64

24. Which of the following will increase with the increase in temperature?

- (a) Surface Tension (b) Viscosity      (c) Molality      (d) Vapour Pressure.

25. During the evaporation of liquid —

- (a) The temperature of the liquid will rise.  
 (b) The temperature of the liquid will fall.  
 (c) May rise or fall depending on the nature.  
 (d) The temperature remains unaffected.

26. The density of a gas is  $1.964 \text{ gdm}^{-3}$  at  $273K$  and  $76 \text{ CmHg}$ . The gas is —

- (a)  $CH_4$       (b)  $C_2H_6$     (c)  $CO_2$     (d)  $Xe$

27. Diffusion of helium gas is four times faster than —

- (a)  $CO_2$       (b)  $SO_2$       (c)  $NO_2$     (d)  $O_2$

28. If two moles of an ideal gas at  $546K$  occupy volume  $44.8L$ , then pressure must be —

- (a) 2 atm      (b) 3 atm    (c) 4 atm    (d) 1 atm.

29. The rms speed of hydrogen is  $\sqrt{7}$  times the rms speed of nitrogen. If ' $T$ ' is the temperature of the

gas, then —

- (a)  $T_{H_2} = T_{N_2}$                       (b)  $T_{H_2} > T_{N_2}$   
(c)  $T_{H_2} < T_{N_2}$                       (d)  $T_{H_2} = \sqrt{7T_{N_2}}$

30. Compressibility factor of an ideal gas is —

- (a) Equal to 2                      (b) Equal to 1                      (c) Greater than 1 (d) Always less than 1

**[B] Assertion and Reason Type Questions :- (MCQ)**

**(Each question mark-1)**

- (a) Both Assertion (A) and Reason (R) are correct and Reason is the correct explanation of the Assertion.  
(b) Both A and R are correct but R is not the correct explanation of A.  
(c) A is correct but R is incorrect.  
(d) A is incorrect and R is correct.
- 1) Assertion (A) : Compressibility factor for hydrogen varies with pressure with positive slope at all pressures.  
Reason (R) : Even at low pressure, repulsive forces dominate in hydrogen gas.
- 2) Assertion (A) : At high pressure, the compressibility factor  $Z$  is  $\left(1 + \frac{Pb}{RT}\right)$ .  
Reason (R) : At high pressure, Van der Waal's equation is modified as  $P(V-b) = RT$
- 3) Assertion (A) : A spherical water drops become flattened on flatter surface.  
Reason (R) : It become flat due to gravity.
- 4) Assertion (A) : Meniscus of a liquid disappears at the critical temperature.  
Reason (R) : Density of liquid and its gaseous phase become equal at the critical temperature.
- 5) Assertion (A) : Viscosity of a liquid decreases on increasing the temperature.  
Reason (R) : Evaporation of liquid increases with rise in temperature.
- 6) Assertion (A) : The graph between  $P$  Vs  $\frac{1}{V}$  is a straight line.  
Reason (R) : For adiabatic process;  $P \propto \frac{1}{V}$

- 7) Assertion (A) : Critical temperature of  $CO_2$  is  $304K$ , it cannot be liquefied above  $304K$ .  
Reason (R) : At a certain temperature,  $Volume \propto \frac{1}{Pressure}$
- 8) Assertion (A) : The value of Van der Waal's constant 'a' for ammonia is larger than that of nitrogen gas.  
Reason (R) : Molecular weight of ammonia is smaller than nitrogen gas.
- 9) Assertion (A) : Effusion rate of oxygen is smaller than nitrogen.  
Reason (R) : Molecular size of nitrogen is smaller than oxygen.
- 10) Assertion (A) : Liquids tend to have maximum number of molecules at their surface.  
Reason (R) : Small liquid drops have spherical shape.

**[C] Very Short Answer Type Questions :-**

**(Each question mark - 1]**

1. Write Van der Waal's equation for  $n - mole$  of a gas.
2. What is meant by compressibility factor?
3. Write the units of Van der Waal's constant 'a' and 'b'.
4. Define critical pressure.
5. What is the effect of temperature on viscosity?
6. What is absolute zero temperature?
7. What would be the SI unit for the quantity  $PV^2 T^2/n$ ?
8. Write the value of universal Gas constant 'R' at C.G.S unit.
9. Write the relation between partial pressure and mole fraction.



10. What is hydrogen bond?
11. Why helium is used in balloons in place of hydrogen?
12. What is critical volume?
13. Define ideal gas.
14.  $1 \text{ atm} = \text{_____ Pa}$ .
15. Define aqueous tension?
16. What is the value of compressibility factor for an ideal gas?
17. At what temperature is the kinetic energy of a gas molecule half of its value at  $327^{\circ}\text{C}$ ?
18. What is the ratio of kinetic energies of 3g. of hydrogen and 4g. of oxygen at  $T(K)$ ?
19. Define Van der Waal's forces.
20. What is the value of ' $R$ ' at  $L \text{ atm } k^{-1} \text{ mol}^{-1}$  ?

**[D] Short Answer Type Questions :-**

**( Each question marks - 2)**

1. In terms of Charles' Law explain why  $-273^{\circ}\text{C}$  is the lowest possible temperature.
2. Calculate the total number of electrons present in 1.4g of dinitrogen gas.
3. State Dalton's Law of partial pressure.
4. What will be the minimum pressure required to compress  $500 \text{ dm}^3$  of air at 1 bar to  $200 \text{ dm}^3$  at  $30^{\circ}\text{C}$ ?

5. Calculate the volume occupied by 8.8g of  $CO_2$  at  $31.1^\circ C$  and 1 bar pressure.  
[ $R = 0.83 \text{ bar L k}^{-1} \text{ mol}^{-1}$ ]
6. What is critical temperature & critical pressure?
7. What is Boyle temperature?
8. Why liquids diffuse slowly as compared to gases?
9. Why is moist air lighter than dry air?
10. State Charle's Law. Give its mathematical expression.
11. Derive the relation between average velocity, root mean square velocity and most probable velocity.
12. The average velocity of gas molecules is  $200 \text{ ms}^{-1}$  calculate its *r.m.s* velocity to the same temperature.
13. Derive the relation between pressure and density.
14. Derive the relation between temperature and density.
15. Explain the following terms :
  - (a) dipole - dipole interaction.
  - (b) Dipole - induced dipole interaction.

**[E] Long Answer Type Questions :-**

**(Each question marks - 3)**

1. Prove that,  $PV = nRT$ . ( $n$  = number of moles of gas).
2. Using the equation of state  $PV = nRT$ ; show that at a given temperature density of a gas is proportional to gas pressure  $P$ .
3. Explain the physical significance of Van der Waals parameters.
4. Critical temperature for carbon dioxide and methane are  $31.1^\circ\text{C}$  and  $-81.9^\circ\text{C}$  respectively. Which of these has stronger intermolecular forces and why?
5. 2.9g of a gas at  $95^\circ\text{C}$  occupied the same volume as 0.184g of dihydrogen at  $17^\circ\text{C}$ , at the same pressure. What is the molar mass of the gas?
6. Density of a gas is found to be  $5.46\text{ g / dm}^3$  at  $27^\circ\text{C}$  at 2 bar pressure. What will be its density at STP?
7. A mixture of dihydrogen and dioxygen at 1 bar pressure contains 20% by weight of dihydrogen. Calculate the partial pressure of dihydrogen.
8. What do you mean by ideal gas and real gas? Why do real gases deviate from ideal behaviors?
9. Name two phenomena that can be explained on the basis of surface tension.
10. 2g. of hydrogen diffuse from a container in 10 minutes. How many gram of oxygen would diffuse through the same container in the same time under similar conditions?
11. At identical temperature and pressure, the rate of diffusion of hydrogen gas is  $3\sqrt{3}$  times that of a hydrocarbon having molecular formula  $C_nH_{2n-2}$ . What is the value of 'n'?

12. If two molecules  $A$  and  $B$  have mass  $100\text{kg}$  and  $64\text{kg}$  and rate of diffusion of  $A$  is  $12 \times 10^{-3}$ , then what will be the rate of diffusion of  $B$ ?
13. If average velocity of a sample of gas molecules at  $300\text{K}$  is  $5\text{ cm s}^{-1}$ , what is RMS velocity of the same sample of gas molecules at the same temperature? [Given,  $U_{mp} : U_{av} : U_{rms} = 1 : 1.128 : 1.224$ ]
14. Why does viscosity of liquids decrease as the temperature is raised?
15. Explain the following :
  - (a) Surface tension.
  - (b) Viscosity.
16. Describe the conditions required for liquifaction of gases.
17. Determine the relation between partial pressure and mole fraction.

**[F] Very Long Answer Type Questions :-**

**(Each question marks 5)**

1. (a) What will be the pressure of the gaseous mixture when  $0.5\text{L}$  of  $\text{H}_2$  at  $0.8\text{ bar}$  and  $2.0\text{L}$  of dioxygen at  $0.7\text{ bar}$  are introduced in a  $1\text{L}$  vessel at  $27^\circ\text{C}$  ?  
 (b) What are the conditions under which gases deviate from ideality?
2. (a) Why do gases deviate from the ideal behaviour?  
 (b) Calculate the temperature of  $4.0\text{ mol}$  of a gas occupying  $5\text{ dm}^3$  at  $3.32\text{ bar}$ .  
 $[R = 0.083\text{ bar dm}^3\text{K}^{-1}\text{ mol}^{-1}]$   
 (c) What is 'torr'?
3. (a) A vessel of  $120\text{ml}$  capacity contains a certain amount of gas at  $35^\circ\text{C}$  and  $1.2\text{ bar}$  pressure. The gas is transferred to another vessel of volume  $180\text{ml}$  at  $35^\circ\text{C}$ . What would be its pressure?  
 (b) Describe the conditions required for liquifaction of gases.

4. (a)  $34.05\text{ml}$  of phosphorus vapour weight  $0.0625\text{g}$  at  $546^\circ\text{C}$  and  $0.1\text{ bar}$  pressure. What is the molar mass of phosphorus?  
(b) Write the units of Van der Waal's constant ' $a$ ' and ' $b$ '.
5. (a) How much time would it take to distribute one Avogadro number of wheat grains, if  $10^{10}$  grains are distributed each second?  
(b) Pay load is defined as the difference between the mass of displaced air and the mass of the balloon. Calculate the pay load when a balloon of radius  $10\text{m}$ , mass  $100\text{kg}$  is filled with helium at  $1.66\text{ bar}$  at  $27^\circ\text{C}$ . [Density of air =  $1.2\text{ kg m}^{-3}$  and  $R = 0.083\text{ bar dm}^3\text{K}^{-1}\text{mol}^{-1}$ ]

## Solutions

[A] 1. a 2. c 3. b 4. a 5. d 6. d 7. d 8. a 9. b 10. a 11. c 12. c 13. d 14. d 15. b  
16. c 17. a 18. d 19. b 20. a 21. b 22. d 23. c 24. d 25. b 26. c 27. b 28. a 29. c  
30. b

[C] 2.  $4.2154 \times 10^{23}$  Nos. of electrons 4. 2.5 bar 5. 5.05L

[D] 5.  $40 \text{ g mol}^{-1}$  6.  $3 \text{ g dm}^{-3}$  7. 0.8 bar 10. 8g 11.  $\text{C}_4\text{H}_6 (n=4)$  12.  $15 \times 10^{-3}$  13.  
5.43 cm/s

[F] 1. (a) 1.8 bar 2. (b) 50k 3. (a) 0.8 bar 4. (a)  $1294.8 \text{ g mol}^{-1}$  5. (a)  $1.90956 \times 10^6 \text{ years}$  (b)  
3811.1kg

**Chapter at a glance :-**

1. The branch of science which deals with study of different forms of energy and their interconversion is called Thermodynamics.
2. A system in thermodynamics refers to that part of universe in which observation are made.
3. The remaining part of Universe which is not part of system constitutes the surroundings. The surroundings include everything other than the system.
4. The wall (real or imaginary) that separates the system from the surroundings is called boundary.
5. **Types of the system :**

Types of the system	Exchange of energy	Exchange of matter
Open	Yes	Yes
Closed	Yes	No
Isolated	No	No

6. **Intensive properties :** The properties of the system whose value are independent on the amount of substance  
**E.g.:** Temperature, Surface Tension.
7. **Extensive properties :** The properties of the system whose value depends on the amount of substance.  
**Eg.:** Mole number, Volume, Internal energy.
8. State variables are the measurable properties of system required to describe the state of the system. Examples are temperature, pressure, volume etc.
9. **Various types of processes :**

Types of process	Definition
Isothermal	Process in which temperature of system remains constant
Adiabatic	Process in which there is no transfer of heat between the system and surroundings
Isobaric	Process in which pressure of system remains constant.
Isochoric	Process in which volume of system remains constant.
Reversible Process	A process is said to be reversible, if a change is brought out in such the process could, at any moment, be reversed by an infinitesimal change.
Irreversible Process	A Process is said to be irreversible, if a change is brought out in such a way that the process could not, at any moment, be reversed by an infinitesimal change.
Cyclic process	The process in which the initial and final state of the system is same then it is called cyclic process. In cyclic process work done, $w = 0$

**10. Work done in a adiabatic process and sign convention :**

Work done ( $W_{ad}$ )	Sign of W
By the system	-
On the system	+

**11. Heat changes and sign convention :**

Heat (q) transferred	Sign of q
From surroundings to the system	+
From system to the surroundings	-

**12.** First law of thermodynamics is also called as law of conservation of energy i.e. energy can neither be created nor destroyed. It also states that the energy of an isolated system is constant.

Mathematical expression of first law is –

$$\Delta U = q + w$$

Where,  $\Delta U$  = Internal energy change

$q$  = amount of heat absorbed

$w$  = Work done.

**13. a)** For an isothermal reversible process, the mathematical expression for work done is

$$Q = -W_{rev} = -2.303 nRT \log \frac{V_f}{V_i} = -2.303 nRT \log \frac{P_i}{P_f}$$

where,  $R$  = Universal Gas Constant,  $n$  = number of moles

$V_i$  = initial volume,  $V_f$  = final volume

$P_i$  = initial pressure,  $P_f$  = final pressure

$T$  = absolute temperature

**b)** For isothermal irreversible change, the mathematical expression of work done is –

$$Q = -W_{rev} = p_{ex} (v_f - v_i), \text{ where, } p_{ex} = \text{external pressure}$$

Expansion of gas in vacuum ( $p_{ex} = 0$ ) is called free expansion. No work is done during free expansion of an ideal gas whether the process is reversible or irreversible.

**c)** For adiabatic change,  $q=0$  and  $\Delta U = w_{ad}$

**14.** The enthalpy of a system may be defined as the sum of the internal energy and the product of its pressure and volume. It is denoted by the symbol  $H$  and is given by mathematically as follows

$$H = U + PV$$

where  $H$  = Enthalpy,  $U$  = Internal energy,

$P$  = Pressure and  $V$  = Volume.



15. The relation between enthalpy change ( $\Delta H$ ) and internal energy change ( $\Delta U$ ) of system is –

$$\Delta H = \Delta U + P\Delta V$$

$$\Delta H = \Delta U + \Delta n_g RT$$

Where  $\Delta n_g$  = difference between the total sum of mole number of gaseous products and mole number of gaseous reactants.

16. **Enthalpy change and sign convention :**

Type of Reaction	Definition	$\Delta H$
Exothermic reaction	Reactions in which heat is evolved during the reaction	–
Endothermic reaction	Reactions in which heat is absorbed during the reaction.	+

17. Specific heat capacity is the quantity of heat required to raise the temperature of one unit mass of a substance by one degree Celsius (or one Kelvin).

At constant pressure as  $q_p = C_p \Delta T = \Delta H$

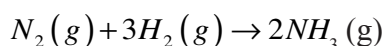
At constant volume as  $q_v = C_v \Delta T = \Delta U$

For one mole of an ideal gas, the relation between  $C_p$  and  $C_v$  is  $C_p - C_v = R$ , For 'n' mole of gas,  $C_p - C_v = nR$

18. The enthalpy change accompanying a reaction is called the reaction enthalpy ( $\Delta_r H$ )

$$\Delta_r H = (\text{Sum of enthalpies of products}) - (\text{Sum of enthalpies of reactants})$$

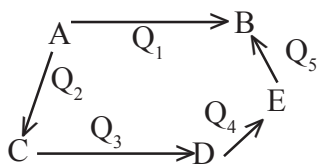
For example, In case of the following reaction, the value of enthalpy of Reaction is,



$$\Delta_r H = 2 \times \Delta H_{NH_3} - \Delta H_{N_2} + 3 \times \Delta H_{H_2}$$

19. **Hess's Law of Constant Heat Summation:** If a reaction takes place in several steps then its standard reaction enthalpy is the sum of the standard enthalpies of the intermediate reactions into which the overall reaction may be divided at the same temperature.

If a reaction occur by using the following steps,



According to Hess's law,

$$Q_1 = Q_2 + Q_3 + Q_4 + Q_5$$

Hess's Law is used for the determination of –

- Enthalpy of formation
- Enthalpy of combination
- Enthalpy of atomization
- Enthalpy of phase transition.

20. Born-Haber cycle is used to determine lattice enthalpy of ionic compounds since they cannot be determined by experiment directly.

21. A process that has a natural tendency of occurrence in a particular direction and is reversible only by application of some external agency is known as spontaneous process.

A process which can neither take place by itself nor by initiation is called non spontaneous process

22. Entropy(s) is the measure of randomness of a system. It is a state function. For reversible process entropy change is given by

$$\Delta S = \frac{q_{\text{rev}}}{T} \text{ where}$$

$q_{\text{rev}}$  is heat absorbed or released during the reaction

T is the temperature of the reaction

At constant pressure  $q_{\text{rev}} = \Delta H$

$$\Delta S = \frac{\Delta H}{T}$$

For spontaneous process  $\Delta S(\text{total}) = \Delta S(\text{sys}) + \Delta S(\text{surr}) > 0$

Entropy increases from solid to gas. Entropy is maximum for gases as they have maximum disorderness.

23. According to second law of thermodynamics entropy of the universe always increases during a spontaneous change.

24. Gibbs function or Gibbs energy is denoted by G.

$G = H - TS$ , where H = Enthalpy, T = Temperature, S = Entropy.

The criteria for spontaneous reaction in relation to  $\Delta G$  at constant pressure and constant temperature.

If  $\Delta G < 0$ , process is spontaneous

If  $\Delta G = 0$ , process is equilibrium

If  $\Delta G > 0$ , Process is non-spontaneous

25. Relation between Gibbs energy change ( $\Delta_r G^\ominus$ ) and equilibrium constant (K)  $\Delta_r G^\ominus = -RT \ln K$

$$= \Delta_r G^\ominus = -2.303RT \log K$$

26. Effect of temperature on spontaneity of Reactions :

$\Delta_r H^\ominus$	$\Delta_r S^\ominus$	$\Delta_r G^\ominus$	Description
-	+	-	Reaction spontaneous at all temperature
-	-	- (at low T)	Reaction spontaneous at low temperature
-	-	+ (at high T)	Reaction non spontaneous at high temperature
+	+	+ (at low T)	Reaction non spontaneous at low temperature
+	+	- (at high T)	Reaction spontaneous at high temperature
+	-	+ (at all T)	Reaction non spontaneous at all temperature

**A. Choose the correct answer : ( MCQ)****(Each question 1 mark)**

1. For a reversible process, the change of entropy for system and its surroundings are  $\Delta s_1$  and  $\Delta s_2$  respectively, Now,
- a)  $\Delta s_1 + \Delta s_2 > 0$                       b)  $\Delta s_1 + \Delta s_2 < 0$   
c)  $\Delta s_1 + \Delta s_2 = 0$                       d)  $\Delta s_1 + \Delta s_2 \geq 0$
2. The Molar Heat capacity of which one of the following gas is independent on temperature -  
a) Ar                      b) H<sub>2</sub>                      c) N<sub>2</sub>                      d) CO<sub>2</sub>
3.  $2\text{NH}_3(\text{g}) \longrightarrow \text{N}_2(\text{g}) + 3\text{H}_2(\text{g})$  for this reaction -  
a)  $\Delta H < 0, \Delta S > 0$                       b)  $\Delta H > 0, \Delta S > 0$   
c)  $\Delta H > 0, \Delta S < 0$                       d)  $\Delta H < 0, \Delta S < 0$
4. Which of the following is correct for isothermal free expansion of an ideal gas -  
a)  $\Delta U = 0$                       b)  $\Delta S = 0$                       c)  $\Delta S > 0$                       d)  $\Delta H > 0$
5. The enthalpy of formation of gaseous N<sub>2</sub>O and NO at 298K are 82 and 90KJmol<sup>-1</sup> respectively. The enthalpy of reaction -  
 $\text{N}_2\text{O}(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$   
a) - 8 KJ                      b) + 98 KJ                      c) - 74KJ                      d) + 8 KJ
6. Hess's law is an application of  
a) 1st law of thermodynamics                      b) 2nd law of thermodynamics  
c) Entropy change                      d)  $\Delta H = \Delta U + P\Delta V$
7. The enthalpy of neutralisation of HCl and NaOH is -  
a) zero                      b) -57.3 KJ  
c) +57.3KJ                      d) cannot be predicted
8. Internal energy of a molecule is -  
a) Kinetic energy only                      b) Potential energy only  
c) Kinetic as well as potential energy                      d) Unpredictable
9. For the two reactions given below:  
 $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}) + x_1\text{kJ}$   
 $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g}) + x_2\text{kJ}$   
a)  $x_1 > x_2$                       b)  $x_1 < x_2$   
c)  $x_1 = x_2$                       d)  $x_1 + x_2 = 0$
10. Which of the following is not a state function?  
a) T                      b) H  
c) q                      d) S

11. Which one is true about endothermic reaction?  
 a)  $H(P) > H(R)$  b)  $H(P) < H(R)$   
 c)  $H(P) \geq H(R)$  d) Sign of  $\Delta H$  is -ve
12. The difference between  $C_p$  and  $C_v$  is equal to -  
 a)  $2 \text{ Kcal mol}^{-1}$  b)  $1.99 \text{ cal mol}^{-1}$   
 c)  $8.314 \text{ J K}^{-1} \text{ mol}^{-1}$  d)  $8.314 \text{ erg K}^{-1} \text{ mol}^{-1}$
13. Which one is true about cyclic process ?  
 a)  $\Delta U = 0$  ;  $\Delta H = 0$   
 b) It is also reversible process  
 c)  $\Delta U > 0$  but  $\Delta H < 0$   
 d)  $\Delta H = 0$  but  $\Delta U < 0$
14. Which compound is most stable on the basis of the  $\Delta_f H^0$  values given in parenthesis?  
 a)  $C_3H_8(g)$  [ $-103.8 \text{ KJ mol}^{-1}$ ] b)  $C_4H_{12}(g)$  [ $-124.7 \text{ KJ mol}^{-1}$ ]  
 c) *n-hexane* [ $-167.2 \text{ KJ mol}^{-1}$ ] d) *n-octane* [ $-208.4 \text{ KJ mol}^{-1}$ ]
15. Which of the following is true for the reaction ?  
 $H_2O(l) \rightleftharpoons H_2O(g)$  at  $100^\circ\text{C}$  at  $100^\circ\text{C}$  and 1 atm pressure  
 a)  $\Delta S = 0$  b)  $\Delta H = T\Delta S$  c)  $\Delta H = \Delta U$  d)  $\Delta H = 0$
16. In which of the following process, a maximum increase in entropy is observed?  
 a) Dissolution of Salt in water b) Condensation of water vapour  
 c) Sublimation of Naphthalene d) Melting of ice
17. Which of the following is correct mathematical equation?  
 a)  $\Delta U = \Delta Q + W$  b)  $\Delta W = \Delta U + \Delta Q$   
 c)  $\Delta U = \Delta W + \Delta Q$  d) None of these
18. Which statement is correct at constant T, for a reaction?  
 $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$   
 a)  $\Delta H = \Delta U$  b)  $\Delta H < \Delta U$   
 c)  $\Delta H > \Delta U$  d)  $\Delta H$  is independent of physical states of CO and  $CO_2$
19.  $\Delta H$  and  $\Delta U$  are related as -  
 a)  $\Delta H = \Delta U$  b)  $\Delta H = \Delta U - nRT$   
 c)  $\Delta H = \Delta U + RT$  d)  $\Delta H = \Delta U + \Delta n_g RT$

20. For the equilibrium change  $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{O}(\text{g})$  at 1 atm and 298K

- a) Standard Gibb's Free Energy Change  $(\Delta G^0) < 0$
- b) Gibb's free energy change  $(\Delta G) < 0$
- c) Standard Gibb's free energy change  $(\Delta G^0) > 0$
- d) Standard Gibb's free energy change  $(\Delta G^0) = 0$

**B. Assertion Reason Type Questions:**

**(Each question mark 1)**

In the following questions a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below for each question.

- a) Both A and R are true and R is the correct explanation of A.
- b) Both A and R are true but R is not the correct explanation of A
- c) A is true but R is false
- d) A is false but R is true

1. Assertion (A) : Combustion of all organic compounds is an endothermic reaction.  
Reason (R) : The enthalpies of all elements in their standard state are zero.
2. Assertion (A) : Both internal energy and enthalpy are state functions.  
Reason (R) : Internal energy and enthalpy are extensive properties.
3. Assertion (A) : The temperature of a gas does not change when it undergoes an adiabatic expansion.  
Reason (R) : During an adiabatic process, heat energy is not exchanged between a system and its surroundings.
4. Assertion (A) : For cyclic process,  $\Delta U = 0$ .  
Reason (R) : Cyclic process represents equilibrium state.
5. Assertion (A) : For adiabatic process,  $dq = 0$ .  
Reason (R) : There is no exchange of heat in adiabatic process.
6. Assertion (A) :  $C_p$  is greater than  $C_v$ .  
Reason (R) : For monoatomic gases, the ratio  $C_p/C_v$  is 1.67.
7. Assertion (A) : For the reaction,  $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$   
$$\Delta H = \Delta U$$
  
Reason (R) : For this reaction,  $\Delta n_g$  is zero.
8. Assertion (A) : A liquid crystallises into a solid and is accompanied by decrease in entropy.  
Reason (R) : In crystals, molecules organise in an ordered manner.
9. Assertion (A) : Enthalpy of combustion is always negative.  
Reason (R) : Combustion reactions proceeds with absorption of heat.
10. Assertion (A) : Both  $\Delta H$  and  $\Delta U$  are state function.

Reason (R) : A state function depends only on the initial and final states of the system and does not depend on the path of the system.

**C. Very short answer type questions :**

**(Each question 1 marks)**

1.  $H = U + PV$ , In this expression 'P' indicates the pressure of system or surroundings?
2. Is the entropy of Universe Constant?
3. Arrange the following in increasing order of entropy: water, water vapour and ice.
4. What will be the value of  $\Delta G$  for melting of ice?
5. For work done by the system, is there any change in internal energy?
6. What will be the value of entropy for a perfectly crystalline substance at absolute 0 Kelvin?
7. Give an application of third law of thermodynamics?
8. What is the value of  $C_p - C_v$  for 'n' mole of ideal gas?
9. Write the relation between standard free energy change ( $\Delta G^0$ ) and equilibrium constant (K).
10. In Isochoric process, which state variable kept as constant?
11. What will be the value of  $\Delta S$ , for an isolated system when  $\Delta U = 0$  ?
12. Give an example of spontaneous reaction where enthalpy change is positive.
13. Define Molar heat capacity of a substance.
14. What do you mean by Gibb's free energy?
15. Is there any heat absorption or emission occurs when an ideal gas expands against zero pressure?

**D. Short answer type questions :**

**(Each question 2 Marks)**

1. "Heat capacity is an extensive property of matter." - Explain.
2. Write the effect of temperature on entropy of the substance.
3. Give an example of each of spontaneous and non-spontaneous process.
4. Write the physical significance of entropy.
5. Explain :  $X(g) + X(g) \rightarrow X_2(g)$  is an exothermic reaction.
6. Derive thermal relation,  $\Delta H = \Delta U + \Delta n_g RT$  for a gaseous reaction in a closed container.
7. What are the necessary condition for a system to be spontaneous and equilibrium with respect to free energy change?
8. What will be the value of "q" and "w", if heat released by the system and work done on the system is 90 joule and 120 joule respectively.
9. The internal energy of a system increases with the increases in temperature - Explain.
10. Why the entropy of ice is less than water?
11. What is the relation between the enthalpy of reaction and bond enthalpy?
12. Explain how is enthalpy related to spontaneity of a reaction?
13. Why is the difference between  $\Delta H$  and  $\Delta U$  not significant for solids or liquids?

14. Give two examples of reactions which are driven by enthalpy change.
15. "Change in internal energy is a state function while work is not." – Why?
16. With the help of first law of thermodynamics and  $H = U + PV$ . Prove that,  $\Delta H = q_p$ .
17. The  $\Delta H$  and  $\Delta S$  for  $2Ag_2O(s) \rightarrow 4Ag(s) + O_2(g)$  are given  $+61.17 \text{ KJmol}^{-1}$  and  $+132 \text{ JK}^{-1}\text{mol}^{-1}$  respectively. Above what temperature will the reaction be spontaneous?

**E. Short answer type questions :**

**(Each question 3 Marks)**

1. Derive the relationship between  $\Delta U$  and  $\Delta H$  for gases.
2. Which one is greater between  $C_p$  and  $C_v$  and why?
3. In a process 800 J of heat is absorbed by a system and 390 J of work is done by the system. What is the change of internal energy for the process?
4. At  $25^\circ\text{C}$  temperature, the standard Gibb's free energy change for a reaction is 5.4 KJ. Calculate the equilibrium constant for this reaction at this temperature.
5. Calculate free energy change at  $45^\circ\text{C}$  temperature for the following reaction:



$$\Delta S^0 = -209\text{J.K}^{-1}$$

6. 1 mole of an ideal gas is compressed isothermally and reversibly from 20 litre to its half at 400k. Calculate the workdone for the process.
7. What is bond enthalpy? Why is average value accepted for the calculation of bond enthalpy?
8. What do you mean by atomisation enthalpy of an element? Write its unit. Mention whether the atomisation enthalpy of element is positive or negative?
9. The reaction :  $\text{C}(\text{graphite}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 393.5 \text{ KJmol}^{-1}$  represents the formation of  $\text{CO}_2$  and also combustion of carbon. Write the  $\Delta H^0$  values of the two processes.
10. It has been found that 221.4 J is needed to heat 30g of ethanol from  $15^\circ\text{C}$  to  $18^\circ\text{C}$ . Calculate -
  - a) Specific heat capacity and
  - b) Molar heat capacity of ethanol.

**F. Long Answer Questions :**

**(Each question 5 Marks)**

1. a) Give an application of Hess's law of constant heat summation.  
b) Calculate the lattice enthalpy of NaCl at standard condition with the help of Born Haber cycle.
2. a) What will be the enthalpy change for the following reaction:  $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr}(\text{g})$ . Given: Bond enthalpy of  $\text{H}_2$ ,  $\text{Br}_2$  and  $\text{HBr}$  is  $435 \text{ KJmol}^{-1}$ ,  $192 \text{ KJmol}^{-1}$  and  $368 \text{ KJmol}^{-1}$  respectively.  
b) Enthalpies of formation of  $\text{CO}(\text{g})$ ,  $\text{CO}_2(\text{g})$ ,  $\text{N}_2\text{O}(\text{g})$  and  $\text{N}_2\text{O}_4(\text{g})$  are -110, -393, 81 and 9.7  $\text{KJmol}^{-1}$  respectively.

Find the value of  $\Delta_r H$  for the reaction :  $N_2O_4(g) + 3CO(g) \rightarrow N_2O(g) + 3CO_2(g)$

3. a) Identify the state functions and path functions out of the followings:  
enthalpy, entropy, heat, temperature, work, free energy.
- b) The enthalpy of vapourisation of  $CCl_4$  is  $30.5 \text{ KJ mol}^{-1}$ . Calculate the heat required for the vapourisation of 284g of  $CCl_4$  at constant pressure. (Molar mass of  $CCl_4 = 154 \text{ g. mol}^{-1}$ ).
4. a) What is lattice enthalpy?
- b) Use the following data to calculate  $\Delta H_{\text{lattice}}^0$  for  $NaBr$ .

Given,

$$\Delta H_{\text{sub}}^0 \text{ for sodium Metal} = 108.4 \text{ KJ mol}^{-1}$$

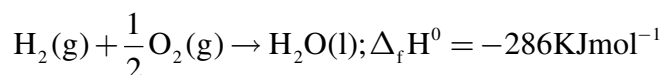
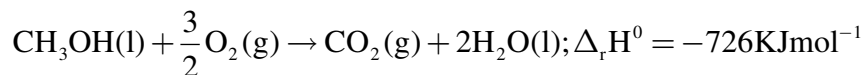
$$\Delta H_i^0 \text{ of sodium} = 496 \text{ KJ mol}^{-1}$$

$$\Delta H_{\text{eg}}^0 \text{ of bromine} = -325 \text{ KJ mol}^{-1}$$

$$\Delta H_{\text{bond dissociation}}^0 \text{ of bromine} = 192 \text{ KJ mol}^{-1}$$

$$\Delta H_f^0 \text{ for } NaBr(s) = -360.1 \text{ KJ mol}^{-1}$$

5. a) For the reaction,  $2A(g) + B(g) \rightarrow 2D(g)$   $\Delta U^\theta = -10.5 \text{ KJ}$  and  $\Delta S^\theta = -44.1 \text{ JK}^{-1}$ . Calculate  $\Delta G^\theta$  for the reaction and predict whether the reaction may occur spontaneously.
- b) The equilibrium constant for a reaction is 10. What will be the value of  $\Delta G^\theta$ ?
- $R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}$ ,  $T = 300 \text{ K}$ .
6. a) Calculate the standard enthalpy of formation of  $CH_3OH(l)$  from the following data:



- b) In a process, 701J of heat is absorbed by a system and 394 J of work is done by the system. What is the change in internal energy for the process?

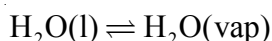


## Solution

- A.** 1-c      2-a      3-b      4-a      5-d      6-a      7-b  
8-c      9-a      10-c      11-a      12-c      13-a      14-d  
15-b      16- c      17-d      18-b      19-d      20-d
- B.** 1-d      2-b      3-d      4-c      5-a  
6-b      7-a      8-a      9-a      10-a

**Chapter at a glance:**

1. When a liquid evaporates in a closed container, molecules with relatively higher kinetic energy escape the liquid surface into the vapour phase and number of liquid molecules from the vapour phase strike the liquid surface and are retained in the liquid phase. A time will come when the number of molecules leaving the liquid equals the number of molecules returning to liquid from vapour. We say the system has reached equilibrium state at this stage. It may be represented by,



At equilibrium the rate of evaporation is equal to the rate of condensation. The mixture of reactants and products in the equilibrium state is called an equilibrium mixture.

2. Equilibrium can be established for both physical process and chemical reactions. Chemical equilibria are important in numerous biological and environmental process. For example, equilibrium involving  $\text{O}_2$  molecules and the protein hemoglobin play a crucial role in the transport and delivery of  $\text{O}_2$  from our lungs to our muscles.
3. When the reactants in a closed vessel of a particular temperature react to give product the concentrations of the reactants keep on decreasing, while those of products keep on increasing for some time after which there is no change in the concentrations of either of the reactants or products. This stage of the system is the dynamic equilibrium. From outside it appears that nothing is happening at equilibrium but in reality on molecular level, there is a continuous change.

Equilibrium is a state at which the concentrations of reactants and products do not change with time.

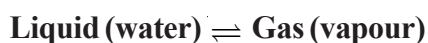
4. The state of chemical equilibrium in a chemical reaction may be classified in three groups.
  - i. The reactions that proceed nearly to completion and only negligible concentrations of the reactants are left.
  - ii. The reactions in which only small amounts of products are formed and most of the reactants remain unchanged at equilibrium state.
  - iii. The reactions in which the concentrations of the reactants and products are comparable at equilibrium.
5. The equilibrium involving ions in aqueous solutions is called ionic equilibrium.
6. Equilibrium for phase transformation processes are some examples of physical equilibrium.

**a) Solid - liquid Equilibrium:**

For this system -

- i. Both the opposing processes occur simultaneously at equilibrium.
- ii. Both the processes occur at the same rate, so that the amount of ice and water remains constant.

For any pure substance at atmospheric pressure, the temperature at which the solid and liquid phases are at equilibrium is called the normal melting point or normal freezing point of the substances.

**b) Liquid-Vapour Equilibrium :**

At equilibrium,

**rate of evaporation = rate of condensation.**

The pressure exerted by the water molecules at a given temperature remains constant and is called the equilibrium vapour pressure of water or vapour pressure of water.

Vapour pressure of water increases with temperature.

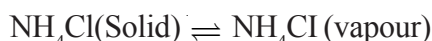
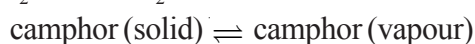
**Evaporation depends on -**

- i. The nature of the liquid*
- ii. Amount of the liquid and*
- iii. the temperature*

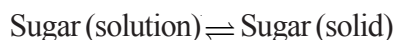
For any pure liquid at atmospheric pressure (1.013 bar) the temperature at which the liquid and vapours are at equilibrium is called boiling point of the liquid. Boiling point depend on the atmospheric pressure.

**c) Solid - Vapour equilibrium:**

Some examples are -

**7. Equilibrium involving dissolution of solid in liquids:**

In a saturated solution, a dynamic equilibrium exist between the solute molecules in the solid state and in the solution.



For this equilibrium,

**The rate of dissolution of sugar = rate of crystallisation of sugar.****8. Equilibrium involving dissolution of gases in liquids:**

In a Soda water bottle there is an equilibrium between the molecules in the gaseous state and the molecules dissolved in the liquid under pressure. i.e.



This equilibrium is governed by Henry's law, which states that, 'the mass of a gas dissolved in a given mass of a solvent of any temperature is proportional to the pressure of the gas above the solvent'.

**9. Some Features of physical equilibrium:**

Process	Conclusion
Liquid $\rightleftharpoons$ Vapour $\text{H}_2\text{O(l)} \rightleftharpoons \text{H}_2\text{O(g)}$	$P_{\text{H}_2\text{O}}$ constant at given temperature.
Solid $\rightleftharpoons$ liquids $\text{H}_2\text{O(s)} \rightleftharpoons \text{H}_2\text{O(l)}$	Melting point is fixed at constant pressure

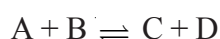
Solute(s) $\rightleftharpoons$ Solute (solution) Sugar(s) $\rightleftharpoons$ Sugar (solution)	Concentration of solute in solution is constant at a given temperature
Gas(g) $\rightleftharpoons$ Gas (aq) CO <sub>2</sub> (g) $\rightleftharpoons$ CO <sub>2</sub> (aq)	[gas(aq)/gas(g)] is constant at a given temperature. [CO <sub>2</sub> (aq)/CO <sub>2</sub> (g)] is constant at a given temperature.

### 10. General characteristics of physical equilibrium:

- Equilibrium is possible only in a closed system at a given temperature.
- Both of opposing processes occur at the same rate and there is a dynamic but stable condition.
- All measurable properties of the system remain constant.
- When equilibrium is attained for a physical process, it is characterised by constant value of one of its parameters at a given temperature.
- The magnitude of such quantities at any stage indicates the extent to which the reaction has proceeded before reaching equilibrium.

### 11. Equilibrium in chemical processes:

Let us consider the following reversible reaction



With passage of time, there is accumulation of the products C and D and depletion of the reactants A and B. This leads to a decrease in the rate of forward reaction and increase in the rate of reverse reaction. Eventually, the two reactions occur at the same rate and the system reaches a state of equilibrium and is called Chemical equilibrium.

Chemical equilibrium can be reached from either direction.

Chemical reactions reach a state of dynamic equilibrium in which the rates of forward and reverse reactions are equal and there is no net change in composition.

### 12. Law of chemical equilibrium and equilibrium constant:

A mixture of reactants and products in the equilibrium state is called an equilibrium mixture.

In 1864, Cato Maximilian Guldberg and Peter Waage proposed that, the concentrations in an equilibrium mixture are related by the following equilibrium equation,

$$K_c = \frac{[C][D]}{[A][B]}$$

Where  $K_c$  is the equilibrium constant and the expression on the right side is called the equilibrium constant expression.

The equilibrium equation is also known as the law of mass action because in the early days of chemistry concentration was called 'active mass'.

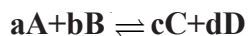
The subscript 'c' in  $K_c$  indicates that  $K_c$  is expressed in concentrations of **mol-L<sup>-1</sup>**.

### 13. Law of Chemical equilibrium:

At a given temperature, the product of concentrations of the reaction products raised to the respective stoichiometric coefficient in the balanced chemical equation divided by the product of concentrations of

the reactants raised to their individual stoichiometric coefficients has a constant value. This is known as the Equilibrium Law or Law of Chemical Equilibrium.

The equilibrium constant for the following general reaction



is expressed as,

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Here [A], [B], [C] and [D] expressing the molar concentration of A, B, C and D respectively.

Equilibrium constant for the reverse reaction is the inverse of the equilibrium constant for the reaction in the forward direction.

$cC + dD \rightleftharpoons aA + bB$  for this reaction equilibrium constant,

$$K'_c = \frac{[A]^a [B]^b}{[C]^c [D]^d} = \frac{1}{K_c}$$

Thus  $K'_c = \frac{1}{K_c}$

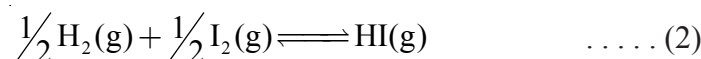
#### 14. For the following, equation



$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

If we change the stoichiometric coefficients in a chemical equation by multiplying throughout by a factor then we must make sure that the expression for equilibrium constant also reflects that change.

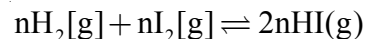
If the above reaction is written as



The equilibrium constant,

$$K''_c = \frac{[HI]}{[H_2]^{1/2}[I_2]^{1/2}} = K_c^{1/2}$$

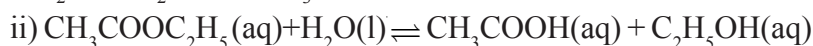
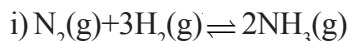
On multiplying the equation (1) by n, we get,

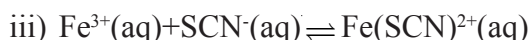


The equilibrium constant for the reaction is equal to  $K_c^n$ .

#### 15. Homogeneous Equilibrium

In a homogeneous system all the reactants and products are in the same phase. For example,





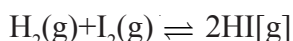
### 16. Equilibrium constant for homogeneous gaseous systems:

For reactions involving gases, it is usually more convenient to express the equilibrium constant in terms of partial pressure.

At constant temperature, the pressure of the gas is proportional to its concentration i.e.

$$p \propto [\text{gas}]$$

For the reaction in equilibrium



We can write either,

$$K_c = \frac{[\text{HI}(\text{g})]^2}{[\text{H}_2(\text{g})][\text{I}_2(\text{g})]}$$

or

$$K_p = \frac{(P_{\text{HI}})^{-2}}{(P_{\text{H}_2})(P_{\text{I}_2})}$$

$$\text{Since, } P_{\text{HI}} = [\text{HI}(\text{g})]RT$$

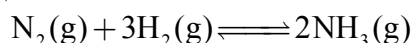
$$P_{\text{I}_2} = [\text{I}_2(\text{g})]RT$$

$$P_{\text{H}_2} = [\text{H}_2(\text{g})]RT$$

Therefore,

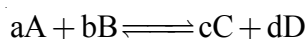
$$K_p = \frac{(P_{\text{HI}})^2}{(P_{\text{H}_2})(P_{\text{I}_2})} = \frac{[\text{HI}(\text{g})]^2 (RT)^2}{[\text{H}_2(\text{g})]RT \cdot [\text{I}_2(\text{g})]RT} = \frac{[\text{HI}(\text{g})]^2}{[\text{H}_2(\text{g})][\text{I}_2(\text{g})]} = K_c$$

But for all reaction  $K_p$  is not equal to  $K_c$ . For Example take the reaction,



$$K_p = \frac{[\text{NH}_3(\text{g})]^2 [\text{RT}]^2}{[\text{N}_2(\text{g})]RT \cdot [\text{H}_2(\text{g})]^3 (\text{RT})^3} = K_c \cdot (\text{RT})^{-2}$$

Similarly, for a general reaction,



$$K_p = \frac{(P_c^c)(P_D^d)}{(P_A^a)(P_B^b)} = \frac{[\text{C}]^c [\text{D}]^d (\text{RT})^{(c+d)}}{[\text{A}]^a [\text{B}]^b (\text{RT})^{(a+b)}} = \frac{[\text{C}]^c [\text{D}]^d}{[\text{A}]^a [\text{B}]^b} (\text{RT})^{(c+d)-(a+b)} = K_c \cdot \text{RT}^{\Delta n}$$

[Where,  $\Delta n$  = number of moles of gaseous products - number of moles of gaseous reactant]

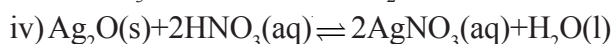
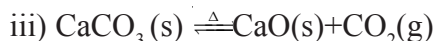
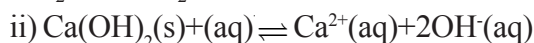
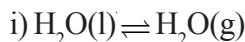
It is necessary that while calculating the value of  $K_p$ , pressure should be expressed in **bar** as standard state is 1 **bar**.

$$1 \text{ Pa} = 1 \text{ Nm}^{-2} \text{ and } 1 \text{ bar} = 10^5 \text{ Pa}$$

### 17. Equilibrium constant for heterogeneous systems

In a heterogeneous system reactants and products are present in more than one phase.

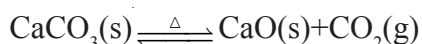
Some examples of heterogeneous system,



Heterogeneous equilibria often involve pure solids or liquids.

The molar concentration of a pure solid or liquid is constant i.e. independent of the amount present.

For the following heterogeneous reaction,



$$K_c = \frac{[\text{CaO}(\text{s})][\text{CO}_2(\text{g})]}{[\text{CaCO}_3(\text{s})]}$$

Since,  $[\text{CaCO}_3(\text{s})]$  and  $[\text{CaO}(\text{s})]$  are both constant therefore the modified equilibrium constant

$$K_c = [\text{CO}_2(\text{g})] \text{ or } K_p = P_{\text{CO}_2}$$

### 18. Important features of equilibrium constant:

- i) Expressin for Equilibrium constant is applicable only when concentrations of reactants and products have attained constant value at equilibrium state.
- ii) The value of equilibrium constant is independent of initial concentrations of the reactants and products.
- iii) Equilibrium constant is temperature dependent having one unique value for a particular reaction represented by a balanced equation at a given temperature.
- iv) The equilibrium constant for the reverse reaction is equal to the inverse of the equilibrium constant for the forward reaction.
- v) The equilibrium constant  $K$  for a reaction is related to the equilibrium constant of corresponding reaction, whose equation is obtained by multiplying or dividing the equation for the original reaction by a small integer.

### 19. Applications of equilibrium constant:

- i) Can predict the extent of a reaction on the basis of its magnitude.
- ii) Can predict the direction of the reaction
- iii) Can calculate equilibrium concentrations.

20. It is important to note that an equilibrium does not give any information about the rate at which the equilibrium is reached. The magnitude of  $K_c$  and  $K_p$  is directly proportional to the concentrations of the products. This implies that a high value of  $K$  is suggestive of a high concentration of products and vice-versa.

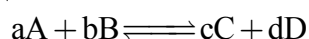
If  $K_c > 10^3$  products predominate over reactants.

If  $K_c < 10^{-3}$  reactant predominate over products.

If  $K_c$  is in range of  $10^{-3}$  to  $10^3$ , appreciable concentrations of both reactants and products are present.

21. The equilibrium constant helps in predicting the direction in which a given reaction will proceed at any stage. For this purpose, we calculate the reaction quotient. It is defined in the same way as the equilibrium constant  $K_c$  except that the concentrations in  $Q_c$  are not necessarily equilibrium values.

For a general reaction,



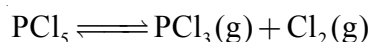
$$Q_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

22. i) if  $Q_c < K_c$ , net reaction goes from left to right.  
ii) if  $Q_c > K_c$ , net reaction goes from right to left.  
iii) if  $Q_c = K_c$ , no net reaction occurs.

**23. Calculation of equilibrium constant and equilibrium concentration:**

In case of a problem in which we know the initial concentration but do not know any of the equilibrium concentrations, then we can calculate it easily by the following process-

Let, a gaseous reaction,



Initial amount: a moles                      0                      0

at equilibrium: (a-x)                      x                      x

Here, x = degree of dissociation.

let, the volume = V

total pressure = P

Now, 
$$K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]}$$

$$\therefore \text{concentration}(C) = \frac{\text{No. of moles (n)}}{\text{Volume}(V)}$$

or, 
$$K_c = \frac{\frac{x}{v} \times \frac{x}{v}}{\frac{a-x}{v}}$$

$$\therefore K_c = \frac{x^2}{(a-x)v}$$

For the calculation of  $K_p$ :

$$\begin{aligned} \text{Total no. of moles at equilibrium} &= a - x + x + x \\ &= a + x \end{aligned}$$



$$\text{Now, } K_p = \frac{P_{\text{PCl}_3} \times P_{\text{Cl}_2}}{P_{\text{PCl}_5}}$$

$$\text{or, } K_p = \frac{P \times \frac{x}{a+x} \times P \times \frac{x}{a+x}}{P \times \frac{a-x}{a+x}} \quad \because \text{ Partial pressure} = \text{Total pressure} \times \text{mole fraction}$$

$$\text{or, } K_p = \frac{x^2 P}{(a+x)(a-x)}$$

With the help of above two process we can easily calculate the values of  $K_p$ ,  $K_c$  and also equilibrium concentration of reactants and products.

### 8. Relationship between Equilibrium Constant and Gibb's free energy ( $\Delta G$ ):

At equilibrium, standard Gibb's Free energy  $\Delta G^0$  and equilibrium constant ( $K_c$ ) are related as -

$$\Delta G^0 = -RT \ln K_c$$

$$\Delta G^0 = -2.303RT \log K_c \quad [\because \ln = 2.303 \log]$$

and, for gaseous reaction,

$$\Delta G^0 = -2.303RT \log K_p$$

### 9. Le-Chatelier's Principle:

It states that a change in any of the factors that determine the equilibrium conditions of a system will cause the system to change in such a manner so as to reduce or to counteract the effect of the change. This is applicable to all physical and chemical equilibria.

#### \* Factors affecting Equilibrium :

**I. Temperature :** If temperature is increased, then an endothermic reaction will proceed faster in the forward direction.

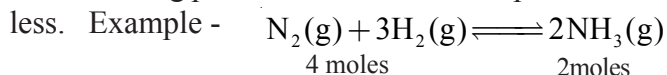
However, If the reaction is exothermic then the reaction will proceed to the backward direction, i.e. the rate of reverse reaction increases.

**II. Concentration :** If at the equilibrium point, the concentration of product is increased, the reaction will proceed to the backward direction.

If product is removed, then the reaction will proceed to the forward direction.

**III. Pressure :** Effect of pressure depends on numbers of moles of gaseous reactant and gaseous product. If the reaction has the same number of moles in reactant and products. Then there is no effect of pressure.

On increasing pressure the reaction will proceed in the direction where total number of moles are less.



For the above reaction if we increase pressure reaction will proceed in forward direction.

#### IV. Addition of Catalyst :

After addition of any Catalyst, rate of reaction increases, but the value of equilibrium constant will not be affected.

#### V. Addition of inert gas :

At constant volume: if the volume of reactant and products are same, then by adding inert gas there is no change in volume, there is no effect of adding inert gas.

At constant pressure: if volume changes, the reaction will proceed in a direction where the number of moles is less.

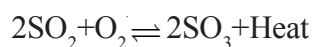
#### A. Choose the correct answer (MCQ) :

(Each question 1 mark)

- In the reaction,  $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2\text{HCl}(\text{g})$   
a)  $K_p \neq K_c$       b)  $K_p = K_c$       c)  $K_p > K_c$       d)  $K_p < K_c$
- Standard free energy change for an equilibrium is zero, the value of  $K_p$  is -  
a) Zero      b) 1      c) 2      d) 100
- $\text{A} + \text{B} \rightleftharpoons \text{C} + \text{D}$   
Initially moles of A and B are equal. At equilibrium moles of C are three times that of A. The equilibrium constant of the reaction will be -  
a) 1      b) 3      c) 4      d) 9
- For an equilibrium reaction if the value of  $K_c \gg 1$ , then the reaction favoured more towards  
a) backward      b) forward  
c) equilibrium will be obtained      d) reaction will stop
- For the following reaction in gaseous phase -  
 $\text{CO}(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g})$   $K_p / K_c$  is  
a)  $(RT)^{\frac{1}{2}}$       b)  $(RT)^{-\frac{1}{2}}$       c)  $(RT)$       d)  $(RT)^{-1}$
- Two moles of  $\text{PCl}_5$  is heated in a closed vessel of 2L capacity. When the equilibrium is attained 40% of it has been found to be dissociated. What is the  $K_c$  in  $\text{mol}/\text{dm}^3$ ?  
a) 0.532      b) 0.266      c) 0.133      d) 0.25

7. If a system is at equilibrium; the rate of forward to the reverse reaction is -  
 a) Less            b) Equal            c) High            d) None of these.
8. For a reaction, if  $K_p > K_c$  the forward reaction is favoured by  
 a) low pressure      b) high pressure    c) high temperature    d) low temperature.
9. A liquid is in equilibrium with its vapour at its boiling point. On the average, the molecules in two phases have equal.  
 a) Inter molecular forces    b) Potential energy  
 c) Kinetic energy              d) Total energy
10. If equilibrium constant is  $2.6 \times 10^8$  at 298K then find the value of  $\Delta G^0$   
 a) -48.014KJ            b) -63.2KJ            c) -23.3KJ    d) +40KJ
11. An equilibrium constant of  $10^{-4}$  for a reaction means, the equilibrium is  
 a) largely towards backward direction  
 b) largely towards forward direction  
 c) equally poised  
 d) never established.
12. For the reaction,  $\text{SO}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightleftharpoons \text{SO}_3(\text{g})$  if we write  $K_p = K_c(\text{RT})^x$ , then x is equal to  
 a) -1                      b)  $-\frac{1}{2}$                       c)  $\frac{1}{2}$                       d) 1
13. If in the reaction,  $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$ , x is that part of  $\text{N}_2\text{O}_4$ , which dissociates; then the total number of molecules at equilibrium will be -  
 a) 1                      b) 3                      c) (1+x)              d) (1+x)<sup>2</sup>
14.  $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$  reaction in lime Kiln goes to completion, because  
 a) CaO does not react to  $\text{CO}_2$  to give  $\text{CaCO}_3$   
 b) Backward reaction is very low  
 c)  $\text{CO}_2$  formed escapes out  
 d) None of the above.

15. For the given equilibrium,



the equilibrium reaction proceeds in forward direction when -

- a) Oxygen is removed                      b)  $\text{SO}_3$  is added  
c) Heat is added                              d) Oxygen is added

**B. Assertion and Reason type Questions :**

**(Each question 1 mark)**

The question given below consists of an assertion and reason. Use the following key to choose the appropriate answer

- a) Both A and R are correct and R is the correct explanation of A  
b) Both A and R are correct but R is not the correct explanation of A  
c) A is correct but R is incorrect  
d) Both A and R are incorrect.

16. Assertion (A) :  $K_p$  can be equal to or less than or even greater than value of  $K_c$

Reason (R) :  $K_p = K_c (RT)^{\Delta n}$  Relation between  $K_p$  and  $K_c$  depends on the change in the number of moles of gaseous reactants and products.

17. Assertion (A) : For the Reaction,  $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$ , increase in pressure favours the formation of  $\text{NO}_2$ .

Reason (R) : The reaction is exothermic.

18. Assertion (A) : For the reaction,  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ , unit of  $K_c$   $\text{L}^2 \text{mol}^{-2}$

Reason (R) : For this reaction equilibrium constant is  $K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$

19. Assertion (A) : For any chemical reaction at a particular temperature, the equilibrium constant is fixed and is a characteristic property.

Reason (R) : Equilibrium constant is independent of temperature.

20. Assertion (A) : An aqueous solution of ammonium acetate can act as a buffer.

Reason (R) : Acetic acid is a weak acid and  $\text{NH}_4\text{OH}$  is a weak base

21. Assertion (A) : The ionisation of hydrogen sulphide in water is low in the presence of hydrochloric acid.

Reason (R) : Hydrogen sulphide is a weak acid.

22. Assertion (A) : In the dissociation of  $\text{PCl}_5$  at constant pressure and temperature addition of helium at equilibrium increases the dissociation of  $\text{PCl}_5$ .

Reason(R) : Helium removes  $\text{Cl}_2$  from the field of action.

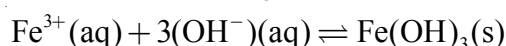
23. Assertion (A) : A solution containing a mixture of acetic acid and sodium acetate maintains a constant value of  $\text{P}^{\text{H}}$  on addition of small amounts of acid or alkali.

Reason (R) : A solution containing a mixture of acetic acid and sodium acetate acts as a buffer solution around  $\text{P}^{\text{H}4.75}$

**C. Very short answer type questions :**

**(Each question 1 mark)**

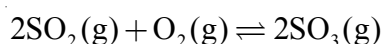
1. What is reversible reaction?
2. Give one example of Solid-Liquid equilibrium.
3. Give one example of liquid-Vapour equilibrium.
4. At what temperature solid and liquid are in equilibrium under 1 atm pressure?
5. Write the relation between  $K_p$  and  $K_c$ .
6. Write the relation between  $K_p$  and  $K_c$  for the reaction given below,  $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$
7. What is the value of  $K_c$  for the reaction  $2\text{Mg}(\text{s}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{MgO}(\text{s})$ ?
8. What is the value of Gibb's free energy at equilibrium conditions?
9. Write the units of  $K_c$  for the reaction,  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$ .
10. Write the expression of  $K_c$  for the following reaction -



**D. Short answer type questions :**

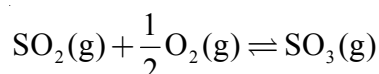
**(Each question : 2 Marks)**

1. Write the characteristics of chemical equilibrium.
2. Why equilibrium constant for a chemical reaction is not effected by catalyst?
3. State Le-Chatelier's principle.
4. Chemical equilibrium is a dynamic equilibrium. Explain.
5. Explain why pure liquids and solids can be ignored while writing the equilibrium constant expression?
6. At 450K,  $K_p = 2.0 \times 10^{10}/\text{bar}$  for the given reaction at equilibrium.



What is  $K_c$  at this temperature?

7. Consider the following gaseous equilibria with equilibrium constants  $K_1$  and  $K_2$  respectively.

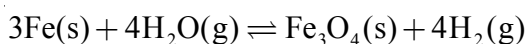


What is the relation between equilibrium constant  $K_1$  and  $K_2$ ?

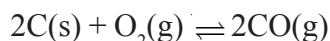
8. Final pressure is higher than initial pressure of a container filled with an ideal gas at constant temperature. What will be the value of equilibrium constant?
9. Three moles of  $\text{PCl}_5$ , three moles of  $\text{PCl}_3$  and two moles of  $\text{Cl}_2$  are taken in a closed vessel. If at equilibrium the vessel has 1.5 moles of  $\text{PCl}_5$ , then calculate the number of moles of  $\text{PCl}_3$  at equilibrium?
10. A sample of  $\text{HI}(\text{g})$  is placed in flask at pressure of 0.2 atm. At equilibrium the partial pressure of  $\text{HI}(\text{g})$  is 0.04 atm. What is  $K_p$  for the given equilibrium?



11. How do rates of evaporation and condensation change initially?
12. What is the initial effect of the change on vapour pressure?
13. Write the expression for equilibrium constant  $K_c$  for the reaction,



14. Define reaction quotient.
15. If,  $Q_c > K_c$ . What would be the type of reaction?
16. What happen when temperature increases for an exothermic equilibrium reaction?
17. What is the effect of increasing the volume on the system described below?



18. What is physical equilibrium? Give one example.
19. What is the unit of equilibrium constant for the reaction,  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$
20. Why the addition of inert gas does not change the equilibrium at constant volume?

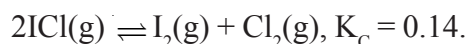
#### E. Short answer type questions :

(Each question = 3 Marks)

1. A vessel at 1000K contains  $\text{CO}_2$  with a pressure of 0.5 atm. Some of the  $\text{CO}_2$  is converted into  $\text{CO}$  on the addition of graphite. If the total pressure at equilibrium is 0.8 atm, then calculate the value of  $K_p$  for the reaction?
2. 5 moles of  $\text{SO}_2$  and 5 moles of  $\text{O}_2$  are allowed to react. At equilibrium, it was found that 60% of  $\text{SO}_2$  is used up. If the total pressure of the equilibrium mixture is one atmosphere, then calculate the partial pressure of  $\text{O}_2$ ?
3. In the reaction,  $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$ .  
In a 2L flask 0.4 moles of each  $\text{H}_2$  and  $\text{I}_2$  are taken. At equilibrium 0.5 moles of  $\text{HI}$  are formed. What will be the value of equilibrium constant  $K_c$ ?
4. For the following three reactions I, II and III equilibrium constants are given -
  - I.  $\text{CO}(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}_2(\text{g}) + \text{H}_2(\text{g}), K_1$
  - II.  $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + 3\text{H}_2(\text{g}), K_2$
  - III.  $\text{CH}_4(\text{g}) + 2\text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}_2(\text{g}) + 4\text{H}_2(\text{g}), K_3$

What is the relation between  $K_1$ ,  $K_2$  and  $K_3$ ?

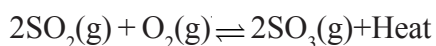
5. State Law of mass action. Derive the mathematical expression of an equilibrium constant for a reversible reaction.
6. Determine relation between  $K_p$  and  $K_c$ .
7. What is the equilibrium concentration of each of the substances in the equilibrium when the initial concentration of ICl was 0.78M?



8.  $\text{A} + \text{B} \rightleftharpoons \text{C} + \text{D}$

If initially the concentration of A and B are both equal but at equilibrium, concentration of D will be twice of that of A, then what will be the equilibrium constant of the reaction?

9. Equimolar concentration of  $\text{H}_2$  and  $\text{I}_2$  are heated to equilibrium in a 2L flask. At equilibrium, the forward and backward rate constants are found to be equal. What percentage of initial concentration of  $\text{H}_2$  has reacted at equilibrium.
10. For the given equilibrium reaction -



Explain the effect of temperature and pressure.

#### F. Long answer type questions :

(Each question Marks - 5)

1. (a) At equilibrium, the concentration of

$$\text{N}_2 = 3 \times 10^{-3} \text{ M}$$

$$\text{O}_2 = 4.2 \times 10^{-3} \text{ M}$$

$$\text{and NO} = 2.8 \times 10^{-3} \text{ M}$$

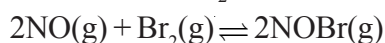
in a sealed vessel at 800K and 1atm pressure. What will be the value of  $K_p$  for the given reaction?

b) Write the relation between Gibb's free energy and equilibrium constant.

2. a) In the equilibrium,  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$ , the  $\text{N}_2\text{O}_4$  is 50% dissociated at 60°C. What will be the value of  $K_p$  at this temperature and one atmosphere pressure?

b) On which factor does the concentration of solute in a saturated solution depends?

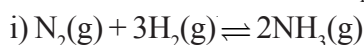
3. a) Nitric oxide reacts with  $\text{Br}_2$  and gives nitrosyl bromide as per reaction given below:



When 0.087 mol of NO and 0.0437 mol of  $\text{Br}_2$  are mixed in a closed container at constant temperature, 0.0518 mol of NOBr is obtained at equilibrium. Calculate equilibrium amount of NO and  $\text{Br}_2$ .

b) Define dynamic equilibrium.

4. a) Write the relation between  $K_p$  and  $K_c$  for the following reactions -



- ii)  $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$
- b) Mention the general characteristics of equilibria involving physical processes.
5. a) What is Homogeneous equilibrium reaction? Give an example.
- b) Using Le-Chatelier's principle, predict the effect of -
- Decreasing the temperature
  - Increasing the pressure.
- in the following equilibrium system -
- $$\text{N}_2(\text{g}) + \text{O}_2(\text{g}) + \text{heat} \rightleftharpoons 2\text{NO}(\text{g})$$

### Solution

- A.** 1. b      2. b      3. d      4. b      5. b  
 6. b      7. b      8. a      9. C      10. a  
 11. a      12. b      13. c      14. c      15. d
- B.**
16. a      17. b      18. a      19. c      20. c  
 21. b      22. d      23. a

#### D. Short answer type question

6.  $7.47 \times 10^{11} \text{ M}^{-1}$       7.  $K_1^2 = \frac{1}{K_2}$       9. 4.5 10. 4 .0

#### E. Short answer type question

1. 1.8atm      2. 0.41 atm      3. 11.11      4.  $K_3 = K_1 \cdot K_2$   
 7.  $[\text{I}_2] = [\text{Cl}_2] = 0.167 \text{ molar}$ ,  $[\text{ICI}] = 0.446 \text{ molar}$ .  
 8. 4  
 9. 33%

#### F. Long answer type question

1. a) 0.622      2. a) 2      3. a) 0.0352 mol NO and 0.0178 mol  $\text{Br}_2$ .



# Ionic Equilibrium

## Part - II

Equilibria that involves ion only are called ionic equilibrium. In the following sections we will study the equilibria involving ions.

1. It is well known that the aqueous solution of sugar does not conduct electricity. However, when common salt (sodium chloride) is added to water it conducts electricity. There are two types of substances-

- ◆ **Non-electrolytes :-** A solution of substance like glucose, sugar etc, that does not dissociates into ions. when dissolve in water. They are poor conductors of electricity.
- ◆ **Electrolytes :-** Substances dissociates into ions when dissolved in water. Two types of electrolytes are there - a) Strong electrolyte b) Weak electrolyte.

**a) Strong electrolyte :** The strong electrolyte is a solution / solute that completely or almost completely ionised or dissociates in aqueous solution. Example - HCl, NaCl, CaCl<sub>2</sub>, H<sub>2</sub>SO<sub>4</sub> etc.

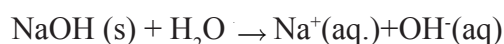
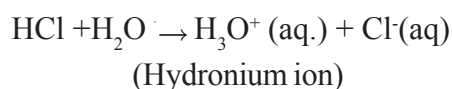
**b) Weak electrolyte :** A weak electrolyte is a solution / solute that partially (less than 5%) ionised or dissociates in aqueous solution. Example - CH<sub>3</sub>-COOH, HCOOH, AgCl etc.

2. **Concepts of Acids and Bases :**

**a) Arrhenius concepts :**

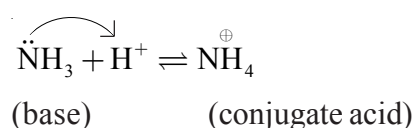
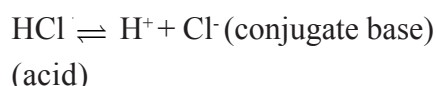
- 1 Acids are the compounds that produce H<sup>+</sup> (as H<sub>3</sub>O<sup>+</sup>) in aqueous solution. Eg - HCl, H<sub>2</sub>SO<sub>4</sub>, HNO<sub>3</sub>, H<sub>2</sub>CO<sub>3</sub>.
- 1 Bases are the compounds that produce OH<sup>-</sup> ion in aqueous solution. Eg - NaOH, KOH, Ca(OH)<sub>2</sub>.

Ionic reactions of acid and base -



**b) Bronsted-Lowry concepts :**

- 1 Acid means proton donor substance or species. Eg - HCl, H<sub>2</sub>CO<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>.
- 1 Base mean proton acceptor substance or species. Eg - NH<sub>3</sub>.
- 1 Conjugate acid and conjugate base :



Conjugate acid = Species + H<sup>+</sup>

Conjugate base = Species - H<sup>+</sup>

### c) Lewis Concepts :

- 1 Any ionic or molecular species that can accept a pair of electrons in the formation of a co-ordinate bond are known as Lewis acid. Eg.  $\rightarrow$   $\text{BF}_3$ ,  $\text{AlCl}_3$ .
- 1 Any ionic or molecular species that can donate a pair of electrons to form a co-ordinate bond are known as Lewis base. Eg.  $\rightarrow$   $\text{NH}_3$ .

### 3. Ostwald's dilution Law :

The degree of dissociation ( $\alpha$ ) of weak electrolyte is inversely proportional to the square root of the molar concentration of the solution.

$$\text{i.e., } \alpha = \sqrt{\frac{K}{C}}, \text{ Here, } K = \text{dissociation constant ; } C = \text{Molar concentration.}$$

$$\text{For any weak electrolyte, (Acid), } \alpha = \sqrt{\frac{K_a}{C}}, \text{ Here } K_a = \text{dissociation constant of acid.}$$

$$\text{For any weak electrolyte, (Base), } \alpha = \sqrt{\frac{K_b}{C}}, \text{ Here } K_b = \text{dissociation constant of base.}$$

### 4. Ionic product of water :

Water is a weak electrolyte and poor conductor of electricity, water undergoes ionisation to a small extent.

The equilibrium constant for the self-ionisation of water is called ionic product of water. It is denoted by  $K_w$ .

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$\text{At } 25^\circ\text{C } K_w = 10^{-14}$$

### 5. $\text{P}^{\text{H}}$ and $\text{P}^{\text{OH}}$ Scale :

$\text{P}^{\text{H}}$  is the negative logarithm of hydrogen ion concentration.

$$\text{i.e. } \text{P}^{\text{H}} = -\log_{10} [\text{H}^+]$$

Similarly, we can write,  $\text{P}^{\text{OH}} = -\log_{10} [\text{OH}^-]$

#### 1 Relation between $\text{P}^{\text{H}}$ and $\text{P}^{\text{OH}}$ :

We know that at  $25^\circ\text{C}$

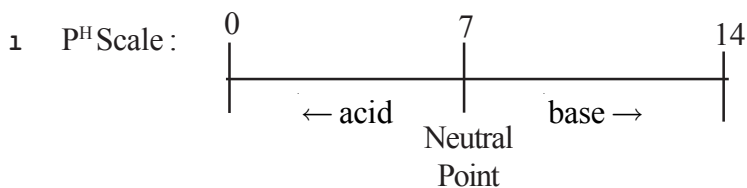
$$K_w = [\text{H}^+][\text{OH}^-] = 10^{-14}$$

$$\text{or, } \log[\text{H}^+] + \log[\text{OH}^-] = \log 10^{-14}$$

$$\text{or, } -\log[\text{H}^+] - \log[\text{OH}^-] = -\log 10^{-14}$$

$$\text{or, } \text{P}^{\text{H}} + \text{P}^{\text{OH}} = 14 \log 10 \text{ [ } \because \log_{10} 10 = 1 \text{ ]}$$

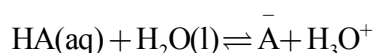
$$\text{or } \text{P}^{\text{H}} + \text{P}^{\text{OH}} = 14$$



Here,  $P^H \propto \frac{1}{\text{Acidity}}$

## 6. Acid Base Equilibria and ionisation constant :

1 Ionisation of weak acid :



at Equil<sup>m</sup>:  $c(1 - \alpha)$        $c\alpha$        $c\alpha$

$\therefore$  Ionisation constant,  $K_a = \frac{c\alpha^2}{1 - \alpha}$       Here,  $\alpha$  = degree of dissociation

After calculation we can get,  $pK_a = -\log K_a$ .

$\therefore$  pH of weak acid,  $pH = \frac{1}{2} pK_a - \frac{1}{2} \log c$

and  $P^{OH}$  of weak base,  $P^{OH} = \frac{1}{2} pK_b - \frac{1}{2} \log c$

## 7. Common ion effect :

The decrease in solubility of a weak electrolyte on adding a strong electrolyte having a common ion is called common ion effect.

Eg: Mixture of  $\text{CH}_3\text{COOH} + \text{CH}_3\text{COO}^-\text{Na}^+$ , [Here, common ion  $\text{CH}_3\text{COO}^-$ ]

## 8. Buffer solution :

A solution that resists change in pH when a small amount of acid or base is added. A buffer solution is acidic, basic.

Acidic buffer,  $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$

Basic buffer -  $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$

The pH of buffer solution can be calculated by using Henderson-Hasselbatch Equation.

For acidic buffer,  $P^H = pK_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$

For basic buffer solution,  $P^{OH} = pK_b + \log \frac{[\text{Salt}]}{[\text{Base}]}$

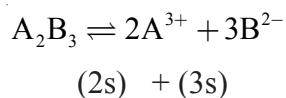
## Buffer Capacity :

It is defined as the number of moles of acid or base added in one litre of a buffer solution to change its  $p^H$  by one unit.

9. Solubility and Solubility product: At a constant temperature the maximum amount of solute which dissolves in a given amount of solvent to produce a saturated solution is known as solubility of that solute.

Solubility product: It is mainly applicable for sparingly soluble salt. (eg.-  $\text{CaSO}_4$ ,  $\text{AgCl}$ )

Calculation of Solubility product: For a sparingly soluble salt  $\text{A}_2\text{B}_3$ ,



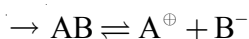
let solubility = S g.mol/Lt.

$$\begin{aligned} \therefore \text{Solubility Product / } K_{\text{SP}} &= [\text{A}^{3+}]^2 [\text{B}^{2-}]^3 \\ &= (2s)^2 \times (3s)^3 \\ &= 4 \times s^2 \times 27 \times s^3 \\ &= 108s^5 \end{aligned}$$

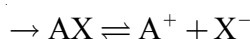
### Effect of common ion in solubility and solubility product :

We can explain it with an example -

Weak electrolyte



Strong electrolyte with common ion



$\therefore$  Solubility product of 'AB',

$$K_{\text{SP}} = [\text{A}^{+}][\text{B}^{-}]$$

$$K_{\text{SP}} = (S_1 + x)S_1$$

Here,  $S_1$  = Solubility of sparingly soluble salt 'AB' after addition of a strong electrolyte with common ion.

**N.B.:-** Presence of common ion decrease the solubility of weak electrolyte.

### 10. Hydrolysis of salts :

Salts formed by the reaction between acids and bases in definite proportions, undergo ionisation in water. The cations/ anions formed on ionisation of salts either exist as hydrated ions in aqueous solutions or interact with water to reform corresponding acids/bases depending upon the nature of salts. The  $\text{P}^{\text{H}}$  of the solution gets affected by this interaction.

We now consider the hydrolysis of the salts of the following types and calculate the  $\text{P}^{\text{H}}$  of mixture by the following process-

#### i) For salts of weak acid and strong base :

$$\text{Hydrolysis constant : } K_{\text{h}} = \frac{K_{\text{w}}}{K_{\text{a}}}$$

$$\text{degree of hydrolysis: } h = \sqrt{\frac{K_{\text{h}}}{C}}$$

$$= \sqrt{\frac{K_{\text{w}}}{K_{\text{a}} \cdot C}}$$

$$\therefore P^H = \frac{1}{2}(P^{K_a} + P^{K_w} + \log C)$$

$$\text{or, } P^H = 7 + \frac{1}{2}(P^{K_a} + \log C)$$

**(ii) For Salts of strong acid and weak base :**

$$\text{Hydrolysis constant, } K_h = \frac{K_w}{K_b}$$

$$\text{degree of hydrolysis, } h = \sqrt{\frac{K_h}{C}}$$

$$= \sqrt{\frac{K_w}{K_b \cdot C}}$$

$$\therefore P^H = \frac{1}{2}(P^{K_w} - P^{K_b} - \log C)$$

$$\text{or, } P^H = 7 - \frac{1}{2}(P^{K_b} + \log C)$$

**(iii) For the salts of weak acid and weak base :**

$$\text{Hydrolysis constant, } K_h = \frac{K_w}{K_a \times K_b}$$

$$\text{Degree of hydrolysis, } h = \sqrt{K_h}$$

$$= \sqrt{\frac{K_w}{K_a \times K_b}}$$

$$\therefore P^H = \frac{1}{2}(P^{K_w} + P^{K_a} - P^{K_b})$$

$$\text{or, } P^H = 7 + \frac{1}{2}(P^{K_a} - P^{K_b})$$

## 11. Relation between ionic product and solubility product :-

- i. If ionic product and solubility product are equal, then the solution is in equilibrium.
- ii. If ionic product < solubility product, then the solution is unsaturated and more salt can be dissolved in it.
- iii. If ionic product > solubility product, then the solution is containing more salt than it can

dissolve, that's why precipitation starts and continues till ionic product becomes equal to solubility product.

**A. Choose the correct answer (MCQ) :**

**(Each question mark : 1)**

- Which of the following cannot act both as Bronsted acid and as Bronsted base?  
a)  $\text{HCO}_3^-$  b)  $\text{NH}_3$  c)  $\text{HCl}$  d)  $\text{HSO}_4^-$
- Which of the following compounds or ions is most likely to behave as a Lewis base?  
a)  $\text{BF}_3$  b)  $\text{PF}_3$  c)  $\text{CO}$  d)  $\text{F}^-$
- Which is the strongest acid in the following?  
a)  $\text{HClO}_4$  b)  $\text{H}_2\text{SO}_3$  c)  $\text{H}_2\text{SO}_4$  d)  $\text{HClO}_3$
- Which of the following molecules acts as a Lewis acid ?  
a)  $(\text{CH}_3)_2\text{O}$  b)  $(\text{CH}_3)_3\text{P}$  c)  $(\text{CH}_3)_3\text{N}$  d)  $(\text{CH}_3)_3\text{B}$
- In  $\text{HS}^-$ ,  $\text{I}^-$ ,  $\text{R-NH}_2$ ,  $\text{NH}_3$  order of proton accepting tendency will be -  
a)  $\text{I}^- > \text{NH}_3 > \text{R-NH}_2 > \text{HS}^-$  b)  $\text{NH}_3 > \text{R-NH}_2 > \text{HS}^- > \text{I}^-$   
c)  $\text{R-NH}_2 > \text{NH}_3 > \text{HS}^- > \text{I}^-$  d)  $\text{HS}^- > \text{R-NH}_2 > \text{NH}_3 > \text{I}^-$
- Conjugate acid of  $\text{NH}_2^-$  is -  
a)  $\text{NH}_4\text{OH}$  b)  $\text{NH}_4^+$  c)  $\text{NH}_2^-$  d)  $\text{NH}_3$
- The strongest conjugate base is -  
a)  $\text{SO}_4^{2-}$  b)  $\text{Cl}^-$  c)  $\text{NO}_3^-$  d)  $\text{CH}_3\text{COO}^-$
- The  $\text{pH}$  of 0.01M  $\text{NaOH}(\text{aq.})$  solution will be -  
a) 7.01 b) 2 c) 12 d) 9
- Which of the following salts will give highest  $\text{pH}$  in water?  
a)  $\text{KCl}$  b)  $\text{NaCl}$  c)  $\text{Na}_2\text{CO}_3$  d)  $\text{CuSO}_4$
- The  $\text{pH}$  value of a 10M solution of  $\text{HCl}$  is -  
a) equal to 1 b) Equal to 2 c) less than 0 d) Equal to 0

11. Which of the following pairs constitutes a buffer?  
 a) HCl and KCl                      b)  $\text{HNO}_2$  and  $\text{NaNO}_2$   
 c) NaOH and NaCl                    d)  $\text{HNO}_3$  and  $\text{NH}_4\text{NO}_3$
12. Solution of 0.1(N) $\text{NH}_4\text{OH}$  and 0.1 (N)  $\text{NH}_4\text{Cl}$  has pH=9.25. Then find out  $\text{P}^{\text{Kb}}$  of  $\text{NH}_4\text{OH}$ .  
 a) 9.25    b) 4.75                      c) 3.75                      d) 8.25
13. The hydrogen ion concentration of a  $10^{-8}$  M HCl aqueous solution at 298K ( $\text{K}_w=10^{-14}$ ) is-  
 a)  $1.0 \times 10^{-8}$  M                      b)  $1.0 \times 10^{-6}$  M  
 c)  $1.0525 \times 10^{-7}$  M                    d)  $9.525 \times 10^{-8}$  M
14. The solubility of an aqueous solution of  $\text{Mg}(\text{OH})_2$  be x, then its  $\text{K}_{\text{sp}}$  is -  
 a)  $4x^3$                       b)  $108x^5$                       c)  $27x^4$                       d)  $9x$
15. The solubility in water of a sparingly soluble salt  $\text{AB}_2$  is  $1.0 \times 10^{-5}$  mol  $\text{L}^{-1}$ . Its solubility product will be-  
 a)  $4 \times 10^{-10}$                       b)  $1 \times 10^{-15}$                       c)  $1 \times 10^{-10}$                       d)  $4 \times 10^{-15}$
16. NaOH is a strong base. What will be the pH of  $5.0 \times 10^{-2}$  M NaOH Solution? ( $\log 2 = 0.3$ )  
 a) 14.00    b) 13.70    c) 13.00    d) 12.70
17.  $\text{P}^{\text{Ka}}$  of a weak acid (HA) and  $\text{P}^{\text{Kb}}$  of a weak base (BOH) are 3.2 and 3.4 respectively. The pH of their salt (AB) solution is -  
 a) 7.2                      b) 6.9                      c) 7.0                      d) 1.0
18. Assuming that the degree of hydrolysis is small, the pH of 0.1 M solution of sodium acetate ( $\text{K}_a = 1.0 \times 10^{-5}$ ) will be-  
 a) 5.0                      b) 6.0                      c) 8.0                      d) 9.0
19. The  $\text{P}^{\text{Ka}}$  of a weak acid (HA) is 4.5. The  $\text{P}^{\text{OH}}$  of an aqueous buffer solution of HA in which 50% of the acid is ionized is -  
 a) 7.0                      b) 4.5                      c) 2.5                      d) 9.5
20. The pH of a 0.1 molar solution of the acid HQ is 3. The value of the ionisation constant;  $\text{K}_a$  of the acid is:  
 a)  $3 \times 10^{-1}$                       b)  $1 \times 10^{-3}$                       c)  $1 \times 10^{-5}$                       d)  $1 \times 10^{-7}$

21. Solubility of  $M_2S$  Salt is  $3.5 \times 10^{-6}$  then find out solubility product.  
 a)  $1.7 \times 10^{-6}$     b)  $1.7 \times 10^{-16}$     c)  $1.7 \times 10^{-18}$     d)  $1.7 \times 10^{-12}$
22. What is the pH of the resulting solution when equal volumes of 0.1 M NaOH and 0.01M HCl are mixed?  
 a) 2.0                    b) 7.0                    c) 1.04                    d) 12.65
23. If 50ml of 0.1M HBr is mixed with 50ml of 0.2M NaOH. Find pH of resulting mixture.  
 a) 2.7                    b) 12.7                    c) 10.7                    d) 1.3
24. Which one of the following has pH value of 10.5?  
 a) Lemon juice            b) Blood            c) Milk                    d) Lime water
25. 0.005 M acid solution has 5 pH. The percentage ionisation of acid is -  
 a) 0.8%    b) 0.6%    c) 0.4%    d) 0.2%
26.  $P^{K_a}$  of  $CH_3COOH$  is 4.74. The pH of 0.01 M  $CH_3COONa$  will be  
 a) 3.28    b) 2.38                    c) 8.37                    d) 3.87
27. Which will have the minimum value of  $P^{K_b}$ ?  
 a)  $OH^-$                     b)  $CH_3COO^-$                     c)  $Cl^-$                     d)  $HCOO^-$
28. Which of the following salts will give highest pH in water?  
 a)  $MgCl_2$     b)  $NaCl$     c)  $Na_2CO_3$     d)  $FeSO_4$
29. Equimolar solutions of the following substances were prepared separately. Which one will record the highest pH value?  
 a)  $BaCl_2$     b)  $AlCl_3$     c)  $LiCl$                     d)  $BeCl_2$
30. Hydrolysis of sodium acetate will give -  
 a) acidic solution    b) basic solution  
 c) neutral solution    d) normal solution
31. Which is a basic salt ?  
 a)  $PbS$                     b)  $2PbCO_3 \cdot Pb(OH)_2$   
 c)  $PbSO_4$                     d)  $PbCO_3$



32. A solution of  $\text{FeCl}_3$  in water act as acidic solution due to -  
 a) hydrolysis of  $\text{Fe}^{3+}$  b) acidic impurities  
 c) dissociation d) ionisation
33. Relation between hydrolysis constant and dissociation constant are given. Which is the correct formula for  $\text{MgCl}_2$ ?  
 a)  $K_h = \frac{K_w}{K_a}$  b)  $K_h = \frac{K_w}{K_b}$   
 c)  $K_h = \frac{K_w}{K_a \times K_b}$  d)  $K_h = \frac{K_h}{K_b}$
34. HCl is completely neutralised by NaOH solution. The resulting solution contains the following species.  
 a)  $\text{Na}^+$ ,  $\text{Cl}^-$ ,  $\text{H}_3\text{O}^+$  b)  $\text{Na}^+$ ,  $\text{Cl}^-$ ,  $\text{H}_2\text{O}$   
 c)  $\text{Na}^+$ ,  $\text{Cl}^-$ ,  $\text{OH}^-$  d)  $\text{H}^+$ ,  $\text{H}_3\text{O}^+$
35. A solution has  $\text{pH}=5$ , it is diluted 100 times, then it will become  
 a) neutral b) basic c) unaffected d) more acidic
36. Which of the following pairs constitutes a buffer?  
 a)  $\text{HNO}_2$  and  $\text{NaNO}_2$  b)  $\text{NaOH}$  and  $\text{KOH}$   
 c)  $\text{NO}_3^-$  and  $\text{NH}_4\text{NO}_3$  d)  $\text{HCl}$  and  $\text{KCl}$
37. Which will have the minimum value of  $\text{p}^{K_a}$ ?  
 a)  $\text{OH}^-$  b)  $\text{CH}_3\text{COO}^-$  c)  $\text{CN}^-$  d)  $\text{HCOO}^-$
38. When ammonium chloride is added to ammonia solution, the pH of the resulting solution will be -  
 a) increased b) seven c) decreased d) unchanged
39. pH of the solution containing  $1.04 \times 10^{-12} \text{M}$  KOH is -  
 a) - 11.9 b) +11.9 c) 7 d) 2.30
40. What is the value of  $K_{sp}$  of  $\text{Fe}(\text{OH})_3$ ? (Solubility =  $S \text{g.mol/L}$ )  
 a)  $S^2$  b)  $4s^3$  c)  $27s^4$  d)  $64s^4$

41. Which is not an example of common ion effect?  
 a)  $\text{NaCl} + \text{AgCl}$                       b)  $\text{H}_2\text{S} + \text{HCl}$   
 c)  $\text{CH}_3\text{COOH} + \text{NaOH}$                 d)  $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$
42. The solubility product of  $\text{BaCl}_2$  is  $4 \times 10^{-9}$ . Its solubility in mo/L is -  
 a)  $4 \times 10^{-3}$       b)  $4 \times 10^{-9}$     c)  $1 \times 10^{-3}$     d)  $1 \times 10^{-9}$
43. If the solubility of calcium fluoride in pure water is  $x$  mol/L; its solubility product is -  
 a)  $\sqrt{2}x$               b)  $2x^2$             c)  $4x^3$             d)  $x^2$

**B. Assertion and Reason type Questions :**

**(Each question 1 mark)**

The question given below consists of an assertion and reason. Use the following key to choose the appropriate answer

- a) Both assertion and reason are CORRECT and reason is the correct explanation of the assestion.  
 b) Both assertion and reason are CORRECT but reason is NOT THE CORRECT EXPLANATION of the assertion.  
 c) Assertion is CORRECT but, reason is INCORRECT.  
 d) Assertion is INCORRECT but, reason is CORRECT.

- Assertion : Buffer solutions are composed of strong acids and strong bases.  
Reason : It maintain the pH to a constant value of 7.4.
- Assertion :  $\text{NaCl}$  is precipitated when  $\text{HCl}$  gas is passed in a saturated solution of  $\text{NaCl}$ .  
Reason :  $\text{HCl}$  is strong acid.
- Assertion : The pH of acid rain is less than 5.6.  
Reason : Carbon dioxide present in the atmosphere dissolves in rain water and becomes carbonic acid.
- Assertion : The ionisation of hydrogen sulphide in water is low in the presence of  $\text{HCl}$ .  
Reason :  $\text{H}_2\text{S}$  is a weak acid.
- Assertion : The aqueous solution of  $\text{CH}_3\text{COONa}$  is alkaline in nature.  
Reason : Acetate ion undergoes anionic hydrolysis.
- Assertion : The pH of a buffer solution containing equal moles of acetic acid and sodium acetate is 4.8 ( $p^{K_a}$  of acetic acid is 4.8)  
Reason : The ionic product of water at  $25^\circ\text{C}$  is  $10^{-14} \text{ mol}^2\text{L}^{-2}$

**C. Very Short Answer type Questions :**

**(Each question marks - 1)**

1. Give one example of dibasic acid.
2. Give one example of diacidic base.
3. Write the conjugate base of  $\text{H}_2\text{PO}_4^-$  ion.
4. What is pH?
5. What is the pH value of 0.001 (N) HCl solution?
6. What is the pH value of acid rain?
7. pH of black coffee is 5.0 at 25°C. Is black coffee acidic or basic?
8. Calculate the pH value of 0.01 (N) NaOH solution?
9. Give one example of acidic buffer solution?
10. Give one example of basic buffer solution?
11. Which conjugate base is stronger  $\text{CN}^-$  or  $\text{Br}^-$ ?
12. If pH of a solution is 7. Calculate its pOH value.
13. Calculate the pH value of 0.0005 (M)  $\text{H}_2\text{SO}_4$  solution.
14. Define common ion effect.
15. Define Buffer solution.
16. What is the pH value of human blood?
17. What is the pH value of soft drinks?
18. Define solubility product.
19. Define weak electrolyte.
20. Between 0.1 M HCl and 0.1 M KCl which one is a good conductor of electricity.
21. What happens to the pH of buffer solution on dilution?
22. Give one example of Bronsted base.
23. What is Lewis acid? Give examples.
24. What is Lewis base? Give examples.
25. What is buffer capacity?
26. What is the nature of aqueous solution of  $\text{NH}_4\text{Cl}$ ?

**D. Short answer type questions :**

**(Each questions 2 marks)**

1.  $\text{BF}_3$  does not have proton but still acts as an acid. Explain.
2. pH of 0.08 mol  $\text{dm}^{-3}$  HOCl solution is 2.85. Calculate its ionisation constant.
3. Define ionic product of water.
4. Define Solubility product.
5. Explain common ion effect with an example.
6. What is the pH of  $10^{-3}\text{M}$   $\text{H}_2\text{SO}_4$ ?
7. A 0.01 M ammonia solution is 5% ionised. Calculate its pH.

8. What is the pH of 0.1M  $\text{NH}_3$  solution? [ $K_b=1.8 \times 10^{-5}$ ]
9. 50ml of 0.1M HCl and 150ml of 0.5 M NaOH are mixed. Calculate the pH of resulting solution.
10. Calculate the solubility of AgCl in 0.2M NaCl Solution. ( $K_{sp}$  of AgCl is  $1.20 \times 10^{-10}$ )
11. Explain conjugate acid and conjugate base with an example.
12. Define Buffer Capacity.
13. Write Henderson-Hasselbalch equation.
14. The Solubility product of  $\text{BaCl}_2$  is  $4 \times 10^{-9}$ . Calculate its solubility in mol/L.
15. If the pH value is 4.5 for a solution then what is the value of  $\text{H}^+$  concentration?
16. 0.023g of Sodium metal is reacted with 100ml of water. Calculate the pH value of the resulting solution?

**E. Short answer type questions :**

**(Each question marks 3)**

1. pH of a saturated solution of  $\text{Ca}(\text{OH})_2$  is 9. Calculate the solubility product ( $K_{sp}$ ) of  $\text{Ca}(\text{OH})_2$  solution.
2. Calculate the molar solubility of  $\text{CaF}_2$  ( $K_{sp} = 5.3 \times 10^{-11}$ ) in 0.1 M solution of NaF.
3. The solubility of  $\text{BaSO}_4$  in water is  $2.42 \times 10^{-3} \text{g L}^{-1}$  at 298K. Calculate its solubility product. (Given molar mass of  $\text{BaSO}_4=233 \text{g mol}^{-1}$ )
4. At 25°C, the dissociation constant of a base, BOH is  $1.0 \times 10^{-12}$ . What is the concentration of hydroxyl ions in 0.01M aqueous solution of the base?
5. What is the pH of millimolar solution of ammonium hydroxide which is 20% dissociated?
6. What is  $[\text{H}^+]$  in mol/L of a solution that is 0.20M in  $\text{CH}_3\text{COONa}$  and 0.10M in  $\text{CH}_3\text{COOH}$ ? [ $K_a$  for  $\text{CH}_3\text{COOH} = 1.8 \times 10^{-5}$ ]
7. What will be the resultant pH when 200ml of an aqueous solution of HCl (pH=2) is mixed with 300ml of an aqueous solution of NaOH (pH=12)?
8. The  $\text{P}^{K_a}$  of a weak acid HA and  $\text{P}^{K_b}$  of weak base BOH are 4.60 and 4.80 respectively. What is the pH of 0.1 M solution of the salt, BA formed from the acid HA and base BOH?
9. If 150ml of 0.1M HBr is mixed with 250ml 0.2 M NaOH, find the pH value of resulting mixture?
10. State Ostwald dilution Law. Determine the mathematical expression of it.
11. Explain the function of acidic buffer solution.
12. Prove that,  $\text{pH}+\text{pOH}=14$ ,
13. A monoprotic acid in 1.00M solution is 0.01% ionised. Calculate the dissociation constant of this acid.
14. The  $\text{P}^{K_a}$  of a weak acid, HA is 4.80. The  $\text{P}^{K_b}$  of a weak base, BOH is 4.78. What is the pH of an aqueous solution of the corresponding salt, BA.
15. HX is a weak acid ( $K_a = 10^{-5}$ ). It forms a salt NaX[0.1M] on reacting with caustic soda. Calculate the degree of hydrolysis of NaX.
16. 0.1 mole of  $\text{CH}_3\text{NH}_2$  ( $K_b=5 \times 10^{-4}$ ) is mixed with 0.08 mole of HCl and diluted to 1L. What will be the  $\text{H}^+$  ion concentration in the solution?

17. The solubility product of  $\text{Ag}_2\text{CrO}_4$  in water at 298K is  $3.2 \times 10^{-11}$ . What will be the concentration of  $\text{CrO}_4^{2-}$  ions in the Saturated solution of  $\text{Ag}_2\text{CrO}_4$ ?
18. Solid  $\text{Ba}(\text{NO}_3)_2$  is gradually dissolved in a  $1.0 \times 10^{-4} \text{ M}$   $\text{Na}_2\text{CO}_3$  solution. At what concentration of  $\text{Ba}^{2+}$  will a precipitate begin to form? ( $K_{\text{sp}}$  for  $\text{BaCO}_3 = 5.1 \times 10^{-9}$ )

**F. Long answer type questions :**

**(Each question marks 5)**

- a) Determine the solubility and solubility product with an example.

b) The solubility product constant of  $\text{Ag}_2\text{CrO}_4$  and  $\text{AgBr}$  are  $1.1 \times 10^{-12}$  and  $5.0 \times 10^{-13}$  respectively. Calculate the ratio of the molarities of their saturated solutions.
- a) Equal volumes of 0.002 M solutions of sodium iodate and cupric chlorate are mixed together. Will it lead to precipitation of copper iodate? (For Cupric Iodate,  $K_{\text{sp}} = 7.4 \times 10^{-8}$ ).

b) Define conjugate acid and conjugate base with a proper example.
- a) The ionisation constant of chloroacetic acid is  $1.35 \times 10^{-3}$ . What will be the pH of 0.1M acid and its 0.1M sodium salt solution?

b) Ionic product of water at 310K is  $2.7 \times 10^{-14}$ . What is the pH of neutral water at this temperature?
- a) What is the minimum volume of water required to dissolve 1g of calcium sulphate at 298K? (For calcium sulphate;  $K_{\text{sp}}$  is  $9.1 \times 10^{-6}$ )

b) Common ion affect influences the solubility for a sparingly soluble salt. Explain with an example.

## Solution

### A. Choose the correct answer : (MCQ)

- |       |       |       |       |       |       |
|-------|-------|-------|-------|-------|-------|
| 1. c  | 2. b  | 3. c  | 4. d  | 5. c  | 6. d  |
| 7. d  | 8. c  | 9. d  | 10.   | 11. b | 12. b |
| 13. b | 14. a | 15. d | 16. d | 17. b | 18. d |
| 19. d | 20. c | 21. b | 22. d | 23. b | 24. d |
| 25. d | 26. c | 27. a | 28. d | 29. a | 30. b |
| 31. b | 32. b | 33. b | 34. b | 35. d | 36. a |
| 37. d | 38. d | 39. b | 40. c | 41. c | 42. c |
| 43. c |       |       |       |       |       |

### B. Assertion/Reason :-

- |      |      |      |      |      |
|------|------|------|------|------|
| 1. d | 2. b | 3. b | 4. b | 5. c |
|------|------|------|------|------|

### D. Short answer type questions :-

- |                           |          |          |                                    |                            |
|---------------------------|----------|----------|------------------------------------|----------------------------|
| 6) 1.7                    | 7) 10.69 | 8) 11.13 | 10) $6.0 \times 10^{-10} \text{M}$ | 14) $10^{-3} \text{mol/L}$ |
| 15) $3.16 \times 10^{-5}$ | 16) 12   |          |                                    |                            |

Marks - 2

### E. Short answer type questions :-

- |                                      |                                      |  |
|--------------------------------------|--------------------------------------|--|
| 1) $0.5 \times 10^{-15}$             | 2) $5.3 \times 10^{-9} \text{mol/L}$ | 3) $1.08 \times 10^{-10} \text{mol}^2 \text{L}^{-2}$ |
| 4) $1.0 \times 10^{-7} \text{Mol/L}$ | 5) 10.3                              | 6) $9 \times 10^{-6}$                                |
| 7) 11.3                              | 14) 7.01                             | 15) 0.01%  |
| 16) $8 \times 10^{-11} \text{M}$     | 17) $2 \times 10^{-4} \text{M}$      | 18) $5.1 \times 10^{-5} \text{M}$                    |

Marks - 3

### F. Long answer type questions :-

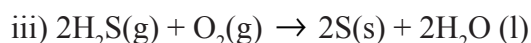
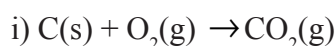
- |                                |                         |   |
|--------------------------------|-------------------------|---|
| 1) b) 91.9                     | 2) a) does not give ppt | 3) pH=1.94 (Acidic Sol <sup>n</sup> ), pH = |
| 2.87 (Basic Sol <sup>n</sup> ) | 4) a) 2.43 L water.     |   |

Marks - 5

**Chapter at a glance :-****1. Redox = Oxidation + Reduction**

Oxidation :- According to classical idea -

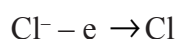
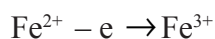
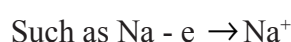
Addition of oxygen or any electro negative atom and removal of hydrogen or any electropositive atom. Such as -



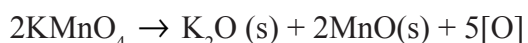
Reduction :- Reverse of Oxidation i.e. addition of Hydrogen or any electropositive atom and removal of oxygen or any electronegative atom.

**2. Electronic Concept :-**

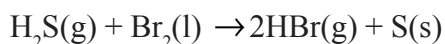
Oxidation = Loss of electron (s)



Reduction = Gain of electron (s)

**3. Oxidant or Oxidising Agent :-** According to classical concept substance which can supplies oxygen or any other electronegative atom they are oxidising agent.

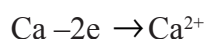
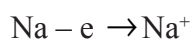
On the basis of electronic concept substance which can gain electron(s) they are oxidising agent.

**4. Reductant or Reducing Agent :-** On the basis of classical concept Hydrogen supplier or any electro positive element supplier are reductant

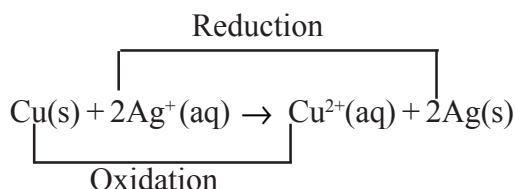
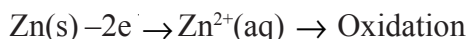
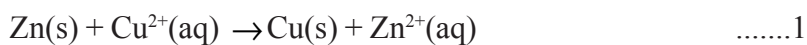
Here H<sub>2</sub>S is the Reducing agent.

According to electronic concept, substance which can lose electron(s).

Such as -



**5. Oxidation and Reduction takes place simultaneously :-**



Zn - Reductant  $\text{Cu}^{2+}$  Oxidant

From the above two reactions 1 & 2 order of reducing character are



**6. Oxidation Number :-** The oxidation number is defined as the charge which appear on an atom of the element when all the atom attached to it are removed in the form of their ions oxidation number and oxidation state are all most to be same.

**7. Rules for finding out the oxidation No. :-**

I. The oxidation number of an element in free state is zero. Like  $\text{O}_2$ , Na,  $\text{P}_4$  etc.

II. The oxidation number of alkali metal and alkaline earth metal is equal to their valency or charge.

Example :  $\text{Na}^+$ ,  $\text{Ca}^{2+}$ .

III. Generally oxidation No. of Oxygen is  $-2$ . Exception in peroxide ( $\text{O}_2^{2-}$ ) it is  $-1$  and in  $\text{OF}_2$  it is  $+2$ .

IV. Generally oxidation no. of Hydrogen is  $+1$ . Exception in metallic hydride it is  $-1$ .

V. Sum of the oxidation no. of atoms in a compound is equal to zero but in case of polyatomic ion it is equal to its net charge.

Example :  $\text{H}_2\text{S}^*\text{O}_4$ , Let the oxidation No. of 'S' be x

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

$$x = 6$$

In case of polyatomic ion,  $\text{Cr}_2\text{O}_7^{2-}$

$$2x - 14 = -2$$

$$x = 6$$

VI. In case of non-metals,

Highest Oxidation No. = Total No. of Valence electron

Least Oxidation No. = No. of unpaired electron with negative sign such as Nitrogen

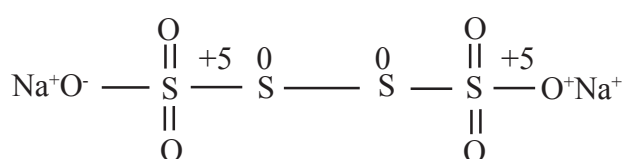
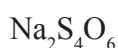
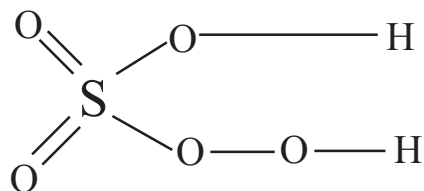
highest oxidation No. is  $+5$  and least oxidation No. is  $-3$ .

Note ' $\text{F}_2$ ' and ' $\text{O}_2$ ' do not obey this rule.

VII. To find out the actual oxidation no. must we have to draw the structure of the compound. if more than one atom are there of whose oxidation no. is changed then structure is must.

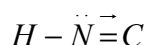


Such as oxidation no. of S in  $H_2SO_5$  (Per Mono Sulphuric Acid) is +8 but it is not possible. According to structure it is +6



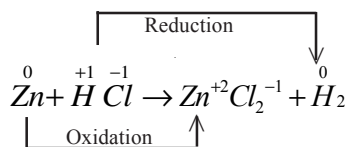
Oxidation no. of 'S' are +5, 0, 0, +5

VIII. If electron goes to higher electronegative to lower electronegative then this/these electron(s) is not considered during the determination of oxidation no. such as



Oxidation No. of 'N' is -3 carbon and hydrogen are +2 and +1 respectively.

- IX. Oxidation no. of metals and non metals in amalgum and alloys are zero. Such as  ${}^0Na/Hg$
- X. In interstitial compound oxidation no. of metal is zero. Such as in  $PdH_2$  oxidation no. of 'Pd' is zero.
- XI. In  $CuH_2$  oxidation No. of 'Cu' is -2
- XII. Oxidation and reduction in terms of oxidation No.



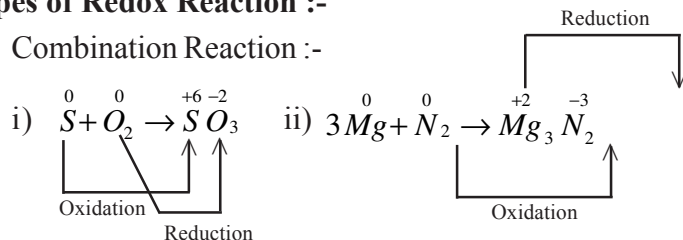
## 9. Oxidation No and Nomenclature | Stock notations :-

The metals having variable oxidation state in a compound are distinguished from one another by placing a roman numeral indicating the oxidation state of the metal with in first bracket ( ) after the symbol.



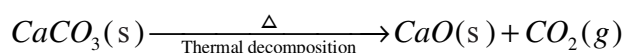
## 10. Types of Redox Reaction :-

a. Combination Reaction :-

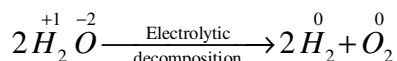
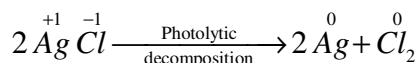
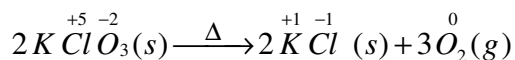


Generally combination reactions are redox reaction but all combination reactions are not redox reaction like,  $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$

b. Decomposition :- All decomposition are not redox reaction. Like,

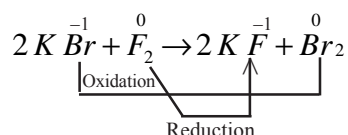
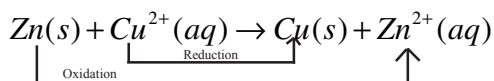


But decomposition of  $\text{KClO}_3$  is the redox reaction:



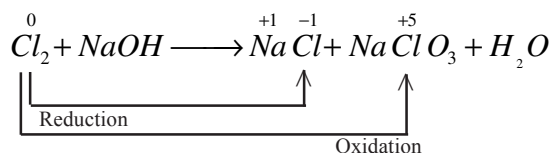
c) Displacement Reaction :-

Displacement reactions are Redox reaction.:



d) Disproportionation Reaction :-

A reaction in which same reactant get oxidised as well as reduced is the disproportionation reaction.

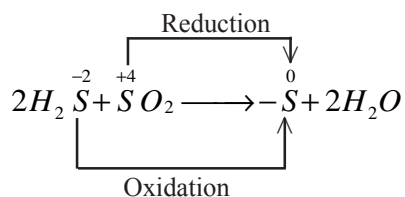


Here,  $\text{Cl}_2$  is the both oxidant and reductant.

e) Comproportionation Reaction :

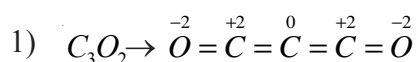
A redox reaction in which oxidised and reduced product are same this reaction is called

comproportionation reaction.

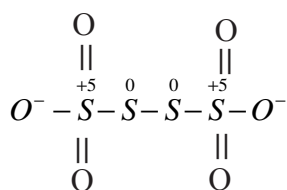


Here, 'S' is the oxidised as well as reduced product.

### 11. Fractional Oxidation State :-



Oxidation No. of 'C' atoms are +2, 0, +2.      Average Oxidation No. of Carbon is =  $\frac{4}{3}$



Oxidation No. of 'S' atoms are +5, 0, 0, +5.      Average Oxidation No. of S =  $\frac{10}{4} = 2.5$

### 12. Balancing of Redox Reaction :-

There are two methods

- i) Oxidation number method
- ii) Ion electron method

i) Oxidation number. Method :

Step - 1: Firstly, you have to write the skeleton equation.

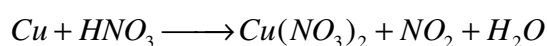
Step - 2: Now, you are to indicate the oxidation number of atoms above its symbol and mark the elements which undergo a change in the oxidation number.

Step - 3: Now calculate the increase or decrease in Oxidation No. per atom and identify the Oxidation and Reduction.

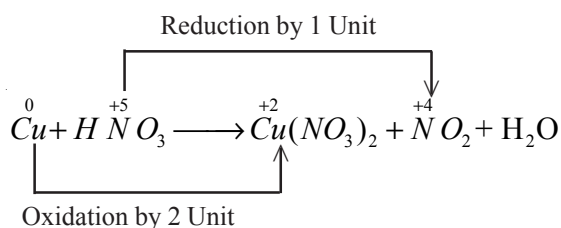
Step - 4: Multiply the formule of the oxidising and reducing agent by suitable integer so as to equalise the total increase or decrease in oxidation number calculated in step-3.

Step - 5: Now you have to balance all the atoms.

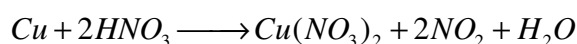
Example : # Balancing of chemical equation by oxidation number method



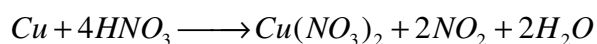
Step - 1, 2 and 3 :



Step - 4:



Step - 5:



ii) Ion Electron method in Acid medium :

Step - 1: Firstly, you have to write the oxidation half and reduction half reaction.

Step - 2: In this case change in oxidation number is expressed by electron number.

In case of oxidation half electron No. must be mentioned in the RHS and in case of reduction half electron number would be mentioned in the left hand side.

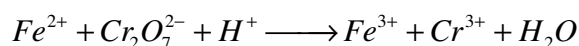
Step - 3: For balancing the oxygen atom we have to add water molecule and then balance H-atoms by adding  $\text{H}^+$  ion on the other side.

Step - 4: You have to equalize the electron number by multiplying the equations by suitable integer and after that two equations (oxidation half and reduction half) would be added.

Step - 5: You have to check the charge of the final equation.

**Balancing of chemical reaction by ion electron method in acidic medium:**

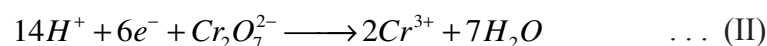
Example:



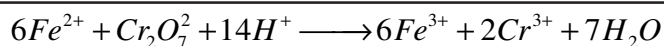
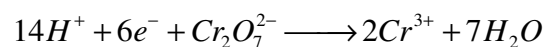
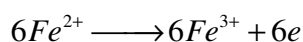
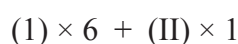
Oxidation half



Reduction half



Equation



Basic medium :

Step - 1 and Step - 2 are same

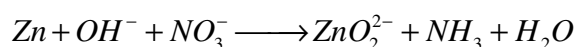
Step - 3: After mentioning the electron number we must have to balance the charge by  $\text{OH}^-$  ion.

Now you have to balance the hydrogen atom by adding  $\text{H}_2\text{O}$  molecule on the otherside.

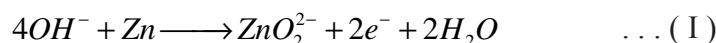
Step - 4 and step - 5 are same as before.

Example :

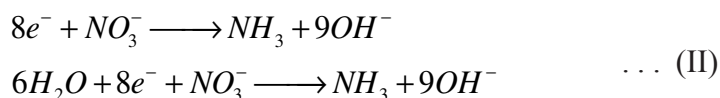
#Basic Medium



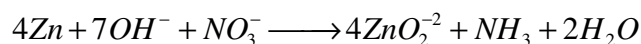
Oxidation half:



Reduction half



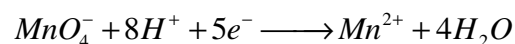
Equation - (I)  $\times$  4 + (II), we get



### 13. Equivalent mass of an oxidant and Reductant :-

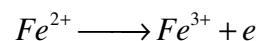
$$\text{Equivalent mass of an oxidant} = \frac{\text{Molecular weight}}{\text{No. of electrons gained / Change in Oxidation number}}$$

Such as,



$$\text{Equivalent wt. of KMnO}_4 = \frac{158}{5} = 31.5$$

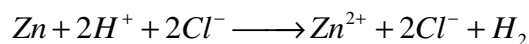
$$\therefore \text{Equivalent mass of reductant} = \frac{\text{Molecular mass}}{\text{No. of electrons released / Change in Oxidation number}}$$



$$\therefore \text{Equivalent mass of FeSO}_4 = \frac{\text{Molecular mass}}{1}$$

**14. Spectator ion :-** Substance those are present in the solution but do not take part in the reaction and also can be canceled during writing the net ionic equation are called spectator ions.

Such as



$\text{Cl}^-$  is the spectator ion.

### 15. Redox Titration :-

Three types of Redox titration (i) Permanganometry, (ii) Dichrometometry, (iii) Iodimetry and Iodometry

i) Permanganometry :-  $\text{KMnO}_4$  is the oxidant and  $\text{KMnO}_4$  it self-indicator. Titration of  $\text{KMnO}_4$

$\text{COOH}$

and  $\text{Fe}^{2+}$  is carried out at room temp while with  $\text{COOH}$ .  $2\text{H}_2\text{O}$ , titration is done at  $60^\circ\text{--}70^\circ\text{C}$ .

$\text{COOH}$

This reaction is carried out in presence of conc.  $\text{H}_2\text{SO}_4$ .

ii) Dichrometometry :-  $\text{K}_2\text{Cr}_2\text{O}_7$  is the Oxidant.  $\text{K}_2\text{Cr}_2\text{O}_7$  is used as primary standard Solution. During dichrometometry titration diphenyl amine or diphenyl benzidine is used as indicator.

iii) Iodimetry and Iodometry :- Both type of redox titrations involve iodine. Iodometry deals with the titration of Iodine liberated in chemical reactions while iodimetry involves titration with a standard solution of Iodine.

**16. Electrode potential :-** Electrode potential is the tendency of an electrode to either lose or gain electrons when it is in contact with the solution of its ions. Hence electrical potential set up between the metal and the solution is called electrode potential.

Electrode potential are two types -

1. Oxidation Electrode potential ( $E_{\text{ox}}$ )
2. Reduction Electrode potential ( $E_{\text{Red}}$ )

Relation between two electrodes are

$$E_{\text{ox}} = -E_{\text{Red}}$$

**17. Standard Electrode Potential :-** Standard electrode potential is denoted by  $E^\circ$ .

When an electrode is suspended in one molar concentrated solution at 298K the electrode potential which is developed is called standard electrode-potential. Standard condition for gas electrode (Inert electrode) is 1 atm pressure.

The standard hydrogen electrode is represented as  $\text{Pt}, \text{H}_2(1\text{atm}) / \text{H}^+(1\text{M})$  or  $\text{H}^+(1\text{M}) / \text{H}_2(1\text{atm}), \text{Pt}$

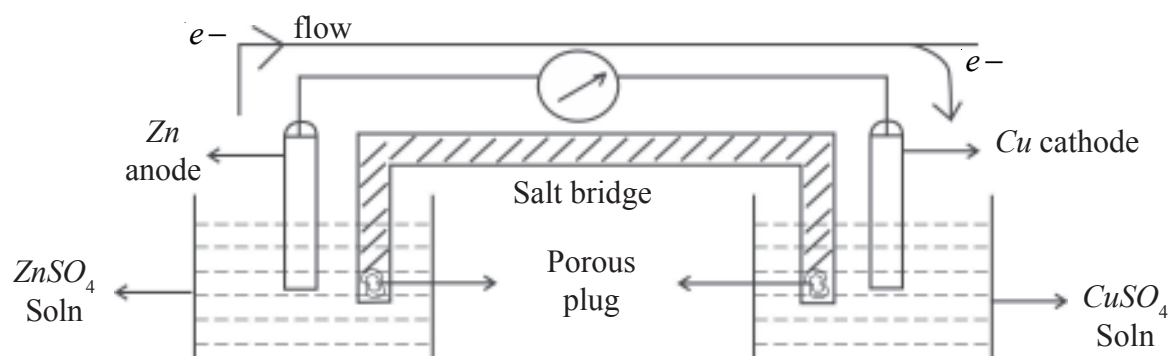
**18. Emf of the cell :-**

$$\begin{aligned} \text{Emf of the cell } (E_{\text{Cell}}) &= E_{\text{Cathode}} - E_{\text{Anode}} \\ &= E_{\text{right}} - E_{\text{left}} \end{aligned}$$

**19.  $E_{\text{Cell}}$  depends on the following factors :-**

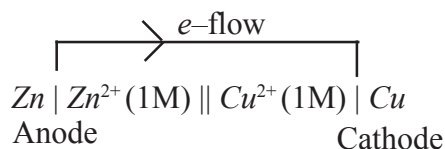
- I. The nature of reactants
- II. Concentration of the solution
- III. Temperature

**20. Electro-Chemical cell :-**



This is Galvanic cell or Daniel Cell

According to Latest cell conventions, cell notation of a Daniel Cell is —



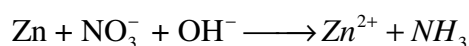
**21. Function of Salt Bridge :-**

- I. It helps to complete the electrical circuit.
- II. It helps to maintain the electrical neutrality.
- III. A salt bridge provides a path for mobility of charges between half cells.

**A. Choose the correct answer (MCQ) :**

**(Mark of each question - 1)**

1. Metal which can easily release electron(s) in aqueous solution, is -  
 a) Zn      b) Mg      c) Fe      d) Cu
2. Oxidation No. of 'P' in Barium Hypophosphite is -  
 a) +2      b) +1      c) +3      d) -3
3. Oxidation No. of 'S' in Permonosulphuric acid is -  
 a) +4      b) 0      c) +1      d) +6
4. Consider the given reaction,



The balanced equation for the given equation is -

- a)  $4\text{Zn} + \text{NO}_3^- + \text{H}_2\text{O} \longrightarrow 4\text{Zn}^{2+} + \text{NH}_3 + 10\text{OH}^-$
- b)  $4\text{Zn} + \text{NO}_3^- + 10\text{H}^+ \longrightarrow 4\text{Zn}^{2+} + \text{NH}_3 + 3\text{H}_2\text{O}$
- c)  $4\text{Zn} + \text{NO}_3^- + 7\text{H}_2\text{O} \longrightarrow 4\text{Zn}^{2+} + \text{NH}_3 + 10\text{OH}^-$
- d) None

5. Oxidation No. of C in carbon sub oxide are -

- a) +2, 0, +2
- b) 0, 2, 0
- c) 2, 2
- d) 2, 4

6. Oxidation number of 'Fe' in  $\text{Fe}_3\text{O}_4$  is/are:

- a) 2,3
- b)  $8/3$
- c) 0, 3
- d) 2, 0

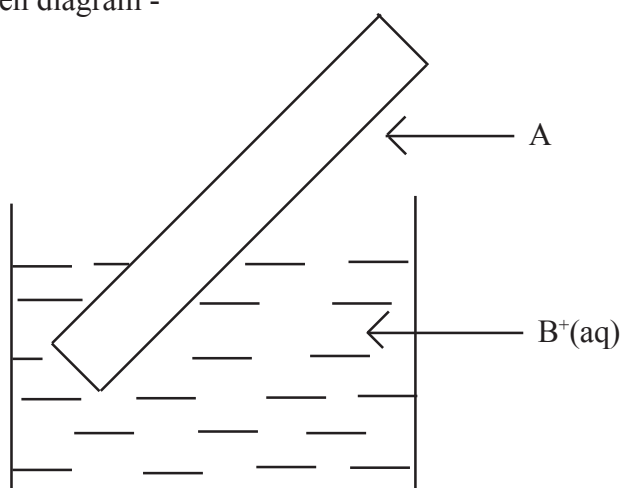
7. Oxidation number of 'Pd' in  $\text{PdH}_2$  is -

- a) 2
- b) 0
- c) -2
- d) None

8. Required number of moles of  $\text{KMnO}_4$  for one mole of  $\text{FeC}_2\text{O}_4$  in acidic medium -

- a)  $1/5$
- b)  $3\frac{5}{3}$
- c)  $3/5$
- d)  $2/5$

9. Consider the given diagram -



If on dipping A in  $\text{B}_2\text{SO}_4(\text{aq})$  the colour of the solution changes from colourless to blue then A & B respectively are -

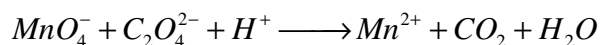
- a) Ag & Cu
- b) Cu & Ag
- c) Cu & Au
- d) Fe & Cu



10. In case of ethyl alcohol the oxidation number of carbon atoms are -

- a) -1, -3      b) +1, +3      c) -1, 0      d) 3, 0

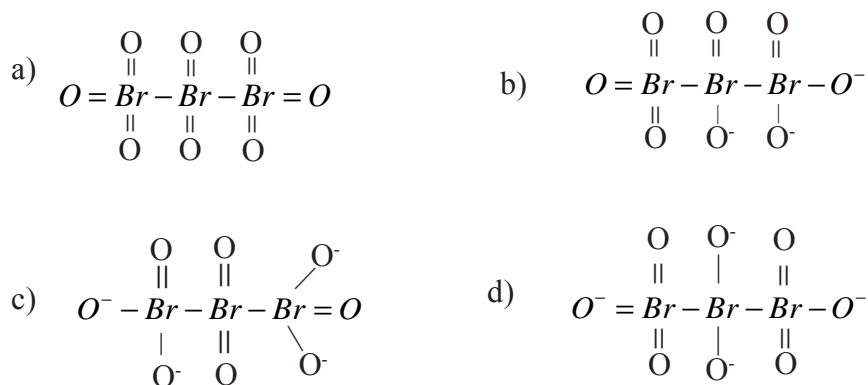
11. For the redox reaction -



The correct co-efficients of the reactants for the balanced equation are -

	$\text{MnO}_4^-$	$\text{C}_2\text{O}_4^{2-}$	$\text{H}^+$
a)	16	5	2
b)	2	5	16
c)	5	16	2
d)	2	16	5

12. The correct structure of tribromo octaoxide is -



13. Which one is can act as both oxidising agent and reducing agent in the following:

- a)  $\text{HClO}_4$       b)  $\text{H}_2\text{S}$       c)  $\text{H}_3\text{PO}_3$       d)  $\text{NH}_3$

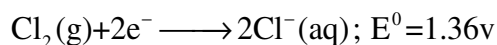
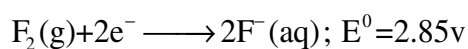
14. Oxidation numbers of A, B, C are +2, +5 & -2 respectively. Possible formula of compound is -

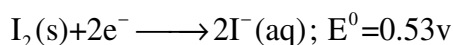
- a)  $\text{A}_2(\text{BC}_2)_2$  b)  $\text{A}_3(\text{BC}_4)_2$  c)  $\text{A}_2(\text{BC}_3)_2$  d)  $\text{A}_3(\text{B}_2\text{C})_2$

15. Which element undergoes disproportionation in water -

- a)  $\text{Cl}_2$       b)  $\text{F}_2$       c) K      d) Ca

16. Standard reduction potentials of the half reactions are given below: -

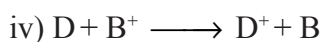
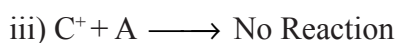
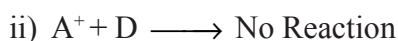
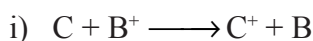




The strongest oxidising and reducing agent respectively are -

- a)  $\text{F}_2$  &  $\text{I}^-$     b)  $\text{Br}_2$  &  $\text{Cl}^-$     c)  $\text{Cl}_2$  &  $\text{Br}^-$     d)  $\text{Cl}_2$  &  $\text{I}_2$

17. Given the following reactions involving A, B, C and D



The correct arrangement of A, B, C, D in order of their decreasing ability as reducing agent.



18. Redox reaction have no concern with -



19. The oxidation no. of 'N' in Ammonium nitrate is /are -



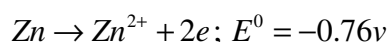
20. The oxidation no. of 'C' in HNC is -



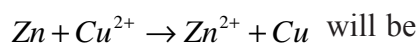
21. Standard reduction electrode potential of three metal x, y and z are -1.2v, 0.5v and -3.0v respectively. The reducing power of these metals will be



22. Electrode potential for the following half cell reactions are -



The EMF for the cell reaction



- a) 1.1v      b) -1.1v      c) 0.32v      d) -1.2v

23. Standard reduction potential at 25°C of  $E^{\circ}\text{Li}^+/\text{Li}$ ,  $E^{\circ}\text{Ba}^{2+}/\text{Ba}$ ,  $E^{\circ}\text{Na}^+/\text{Na}$  and  $E^{\circ}\text{Mg}^{2+}/\text{Mg}$  are – 3.05, –2.90, –2.71 and –2.37 volt respectively. Which one of the following is the strongest oxidising agent.

- a)  $\text{Ba}^{2+}$       b)  $\text{Mg}^{2+}$       c)  $\text{Na}^+$       d)  $\text{Li}^+$

24. Which of the following statement is correct for an electro chemical cell -

- a)  $\text{H}_2$  is anode and Cu is cathode  
b)  $\text{H}_2$  is cathode and Cu is Anode  
c) Reduction occurs at  $\text{H}_2$  electrode  
d) Oxidation occurs at Cu electrode

25. Which of the following statement is wrong for a Fuel cell?

- a) A fuel cell is a galvanic cell  
b) One of the reactant in a fuel cell is a traditional fuel.  
c) The cell reactants in a fuel cell are continuously supplied from a external source.  
d) Modern fuel cells can be easily regenerated using house hold current.

### B. Assertion and Reason type questions :

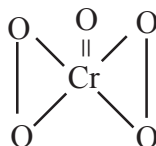
(Each question 1 mark)

Each of the question that follow two statements Assertion (A) and Reason (R) are given. Study both the statements carefully and then mark your answers, according to the codes given below:

- a) Both 'A' and 'R' are true and 'R' as the correct explanation of 'A'  
b) Both 'A' and 'R' are true but 'R' is not the correct explanation of 'A'  
c) 'A' is true but 'R' is false.  
d) 'A' is false but 'R' is true.

1. Assertion (A) : Oxidation number of chromium (Cr) in  $\text{CrO}_5$  compound is +6

Reason (R) : The structure of  $\text{CrO}_5$  is



2. Assertion (A) : 'Cl<sub>2</sub>' can take part in the disproportionation reaction in basic medium.  
Reason(R) : 'Cl<sub>2</sub>' can act as both oxidising and reducing agent.
3. Assertion (A) : Oxidation number of carbon in HCN is +2.  
Reason (R) : Carbon always shows an oxidation state of +4.
4. Assertion (A) : H<sub>2</sub>O<sub>2</sub> can act as both oxidising agent and reducing agent.  
Reason (R) : Oxidation number of oxygen in H<sub>2</sub>O<sub>2</sub> is -1
5. Assertion(A) : Zn(S) + Cu<sup>2+</sup>(aq) → Zn<sup>2+</sup>(aq) + Cu(s). This is comproportionation reaction.  
Reason (R) : In comproportionation reaction oxidised and reduced product are same.
6. Assertion(A) : All single displacement reaction is a redox-reaction.  
Reason(R) : Double displacement reaction is also redox reaction.
7. Assertion(A) : All thermal decomposition reaction reaction is a redox reaction.  
Reason (R) : Acidbase neutralisation reaction can never be a redox reaction.

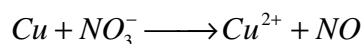
**C. Very Short answers question :**

**(Each question carry 1 marks)**

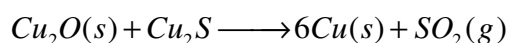
1. Name a solid oxidising agent.
2. Name the strongest oxidising element.
3. What is the Oxidation no. of 'S' in sodium thiosulphate?
4. What is the equivalent weight of KMnO<sub>4</sub> in acidic medium?
5.  $\text{H}_3\text{PO}_4 + \text{Ca}(\text{OH})_2 \longrightarrow \text{CaHPO}_4 + 2\text{H}_2\text{O}$   
What is the equivalent weight of H<sub>3</sub>PO<sub>4</sub> in the given reaction?
6. Name a compound of sulphur, which act as both oxidising agent and reducing agent.
7. Name the compound of 'Fe' where 'Fe' shows +1 Oxidation state.
8. Is single displacement reaction be a redox reaction?
9. What is he oxidation number of 'C' of -COOH group in CH<sub>3</sub>COOH?
10. Among  $E^0_{\text{Zn}^{2+}|\text{Zn}}$  &  $E^0_{\text{Cu}^{2+}|\text{Cu}}$  which one have -ve  $E^0$  value?

**D. Short answer type questions :****(Each Question carry two marks) :**

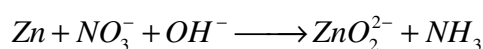
1. Why  $\text{Cu}_2\text{O}$  act as both oxidant and reductant?
2. Balance the chemical equation in acidic medium:



3. Justify that the reaction is a redox reaction.



4. Balance the chemical equation in basic medium.



5. What is standard electrode potential?
6. Why KCl is used as an electrolyte in salt bridge?
7. Mention the function of salt bridge.
8. How many moles of  $\text{KMnO}_4$  are required to oxidise one mole of  $\text{FeSO}_4$  in acidic medium.
9. Balance the following redox reaction by oxidation number method.



10. What is redox titration?

**E. Long answer type questions :****(Each question carry 3 marks)**

1. What is Electrode-Potential? Name the indicator which is used in dichrometry titration?
2. What is the function of Salt-Bridge?
3. What is the condition of standard electrode potential? What is the relation between oxidation electrode potential and reduction electrode potential?
4. How can you measure the Emf of the cell from the electrode potential value? Mention the factors on which Emf of the cell depends?
5. How can you represent the standard hydrogen electrode? Mention the change of energy in electrochemical cell?
6. Mention the difference between Emf and cell potential.

7. Mention the stock notation of the following compound.
  - i. Mercurous (I) - Chloride
  - ii. Stannic (IV) - Oxide
  - iii. Ferrous (II) - Sulphate
  
8. What is the spectator ion? Give the example.
  
9. Define Oxidation and Reduction with example.
  
10. i. Why  $\text{SO}_2$  act as both oxidising and reducing agent?  
 ii. Name a liquid Oxidising agent?
  
11. 2.5 mole  $\text{N}_2\text{H}_4$  losses 25 mole of electrons and form a compound x. In compound x all the nitrogen atom are appeared. Assume that oxidation no. of Hydrogen remain same. What is the Oxidation Number of Nitrogen in the new compound x.

**F. Long answer type questions :**

**(Each question carry 5 mark)**

1. i. Explain with chemical equation that oxidation and reduction takes place simultaneously.  
 ii. Why no indicator is used in the permanganometry titration? 3+2=5
  
2. i. Give the example of a comproportionation reaction?  
 ii. Mention a chemical reaction which is not a redox.  
 iii. What is the oxidation No. of 'S' in sodium tetra thionate? 2+2+1=5
  
3. i. What is Oxidation Number?  
 ii. Balance the equation by oxidation No. method  

$$\text{KMnO}_4 + \text{H}_2\text{SO}_4 + \text{H}_2\text{O}_2 \longrightarrow \text{K}_2\text{SO}_4 + \text{MnSO}_4 + \text{O}_2 + \text{H}_2\text{O}$$

or

$$\text{Na}_2\text{S}_2\text{O}_3 + \text{I}_2 \longrightarrow \text{NaI} + \text{Na}_2\text{S}_4\text{O}_6$$
 iii. What is the Oxidation Number of 'Fe' in hydrated nitroso ferrous sulphate. 2+2+1=5

## Solution

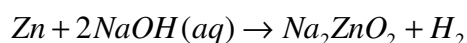
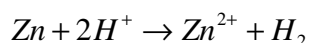
A.

- |       |       |       |       |
|-------|-------|-------|-------|
| 1. b  | 2. b  | 3. c  | 4. d  |
| 5. b  | 6. a  | 7. b  | 8. c  |
| 9. b  | 10. a | 11. b | 12. a |
| 13. c | 14. b | 15. a | 16. a |
| 17. d | 18. d | 19. a | 20. a |
| 21. d | 22. a | 23. c | 24. a |
| 25. a |       |       |       |

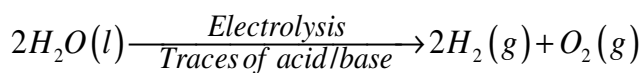
**Chapter at a glance :-**

- Hydrogen is the most abundant (70%) and lightest element of universe having only one electron. It has three isotopes protium ( ${}^1_1H$ ), Deuterium ( ${}^2_1H$ ) and radioactive Tritium ( ${}^3_1H$ ). Deuterium (D) forms heavy water ( $D_2O$ ) which is used in nuclear reactors as moderator. Due to resemblance to alkali metals and halogens, the position of hydrogen in periodic table is not clear.
- Preparation of Hydrogen :**

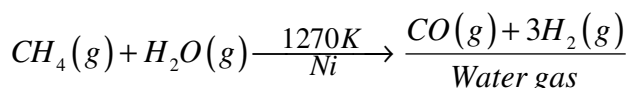
- It can be prepared by the reaction of granulated zinc with dilute hydrochloric acid or aqueous alkali.



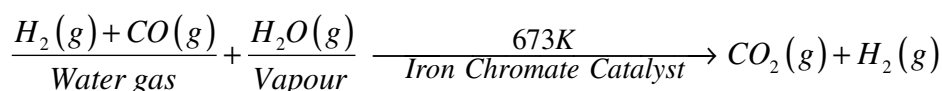
- Electrolysis of water (acidic or alkaline medium) using platinum electrodes gives hydrogen.



- At a very high temperature methane reacts with water vapour in presence of nickel dust to form water gas.



The mixture of  $CO$  and  $H_2$  is also called synthetic gas or 'Syngas'.



This reaction is known as water-gas shift reaction.

**3. Properties of Hydrogen :**

Due to highest bond dissociation enthalpy of  $H-H$  single bond, it is comparatively inert. Due to 1s electronic configuration it accomplishes reactions by loss of electron, gain of electron or by sharing electrons. Dihydrogen is a colourless, tasteless, combustible gas. It does not dissolve in water.

- d-block elements like Pt, Pd, Ni etc. reacts with dihydrogen to form interstitial hydride - it is known as occlusion.
- Hydrogenation of vegetable oils using nickel as catalyst gives edible fats (margarine and vanaspati ghee). Through this process oxidation of unsaturated vegetable oils can be stopped. As a result

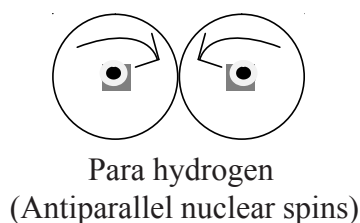
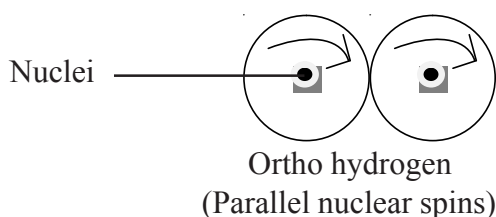


vegetable oils become free from rancidity and tasteless.

**6. Uses of Hydrogen :**

Dihydrogen is used in preparation of nitric acid and nitrogenous fertilizers, vanaspati oil, many bio chemical compounds, metallic hydrides etc.

7. Atomic hydrogen and oxy-hydrogen torches are used in welding and cutting of metals. It is used as a rocket fuel in space research. Research works are going on to use dihydrogen as fuel in modern sophisticated vehicles.
8. Molecules of hydrogen in which the spins of nuclei of both hydrogen atoms are in the same direction are known as ortho hydrogen and the molecules of hydrogen in which the spins of nuclei of both hydrogen atoms are in opposite direction are known as para hydrogen.



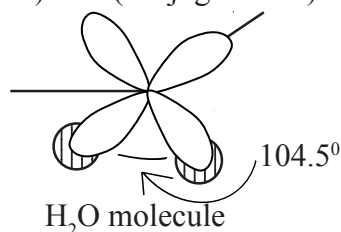
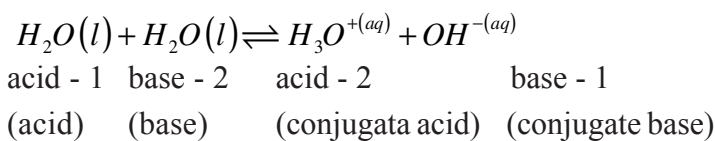
**9. Generally hydrogen forms three types of hydrides :**

- i) Ionic hydride (like  $LiH$ ,  $NaH$ ,  $KH$ ,  $NaBH_4$ ,  $LiBH_4$ )
- ii) Covalent or molecular hydride (like  $CH_4$ ,  $NH_3$ ,  $H_2O$ ,  $PH_3$  etc.)
- iii) Metallic or interstitial hydride (like  $ScH_2$ ,  $CrH$ ,  $LaH$ ,  $YbH$  etc.)

Depending on relative number of electrons and bonds in their lewis structures covalent hydride are further classified as —

- a) Electron deficient molecular hydrides (like  $B_2H_6$ ,  $(AlH_3)_n$  etc.)
- b) Electron precise molecular hydrides (like  $CH_4$ )
- c) Electron rich molecular hydrides (like  $NH_3$ ,  $H_2O$ ,  $HF$ )

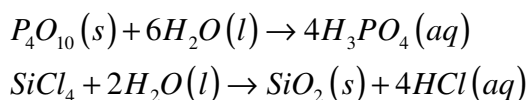
10. Water is a tasteless, colourless and very important substance for whole biosphere. It is a good solvent. Water has a higher specific heat, thermal conductivity, surface tension, dipole moment and dielectric constant. In the gas phase water is a bent molecule with a bond angle of  $104.5^\circ$ . The crystalline form of water is ice. Density of ice is less than that of water, therefore an ice cube floats on water. Water is amphoteric, water shows auto-protolysis —



11. Due to presence of extensive hydrogen bonding between water molecules water shows unusual properties in condensed phase (liquid and solid phase).

In  $CuSO_4 \cdot 5H_2O$  the water molecule outside co-ordination sphere is hydrogen-bonded and other four molecules of water are co-ordinated with central atom.

12. Due to high dielectric constant, water has very strong hydrating tendency, as a result along with ionic compounds some covalent compounds are hydrolysed in water. Example -

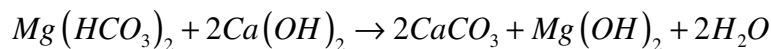
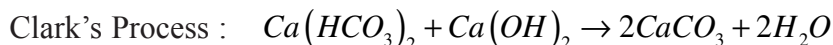
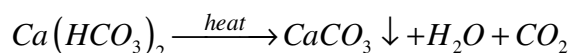
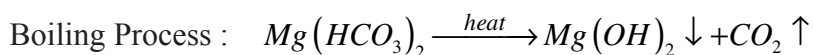


13. Examples of some Hydrate compound —

\*  $[Cr(H_2O)_6]^{3+} 3Cl^-$  – water molecule is attached as co-ordinated bond.

\*  $BaCl_2 \cdot 2H_2O$  – water molecule resides in interstitial position of crystal lattice.

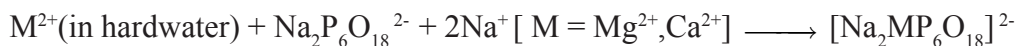
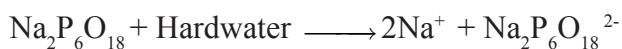
14. Due to presence of bicarbonates, chlorides and sulphates of calcium and magnesium water becomes hard. Among this due to presence of calcium and magnesium hydrogen carbonate water becomes temporary hard which can be softened by boiling or by Clark's process.



Permanent hardness of water cannot be removed by boiling. To remove permanent hardness (1) Washing soda method (2) Calgon's method (3) Ion-exchange method and (4) Synthetic resins or organic ion exchangers (ion exchange resins) etc. are used. The reaction involved in this process are as follows -

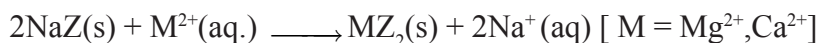
**Washing Soda method:**



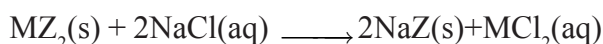
**Calgon's method:**

Ion exchange method:

Hydrated  $\text{NaAlSiO}_4$  is known as zeolite / permutit generally expressed as NaZ.

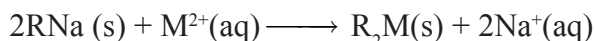


Permutit / Zeolite can be regenerated for further use by treating with an aqueous sodium chloride solution.

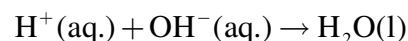
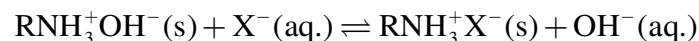
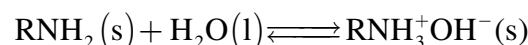
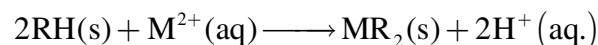


Synthetic resins method:

Now a days this method is used extensively for softening of water. Water insoluble large organic molecule with  $-\text{SO}_3\text{H}$  group are the main part of cation exchange resins. Ion exchange resin ( $\text{RSO}_3\text{H}$ ) react with  $\text{NaCl}$  and gets converted into  $\text{RNa}$ . The resin exchanges  $\text{Na}^+$  ions will  $\text{Ca}^{2+}$  and  $\text{Mg}^{2+}$  ion present in hard water to make the water soft.



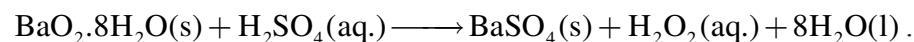
After that this water is passed through anion exchange resins to get pure demineralised or ion-free water.



### 15. Hydrogen peroxide ( $\text{H}_2\text{O}_2$ ):

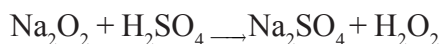
Preparation:

i. Acidifying barium peroxide and removing excess water by evaporation under reduced pressure gives hydrogen peroxide.



ii. By Marck's Process:

In this process, Sodium peroxide is mixed with 20%  $\text{H}_2\text{SO}_4$  solution to prepare  $\text{H}_2\text{O}_2$ .

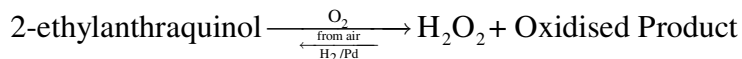


iii. By electrolysis of  $\text{H}_2\text{SO}_4$  Sol<sup>n</sup>: In this method a 50% solution of sulphuric acid is electrolysed at high current density to peroxodisulphuric acid which on further hydrolysis with water gives hydrogen peroxide.





iv. In industry 2-Ethylanthraquinol is self-oxidised in presence of  $\text{O}_2$  (from air) to get  $\text{H}_2\text{O}_2$ .



#### 16. Properties:

Pure hydrogen peroxide is almost colourless (Pale blue) and syrupy liquid. It is unstable in nature. When stored for long time or heated it decomposes to give water and oxygen.

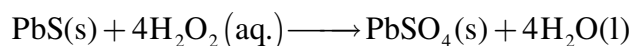
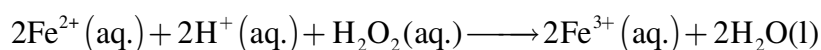
The presence of metal dust, alkali or glass acts as catalyst in the decomposition process.



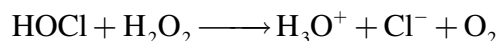
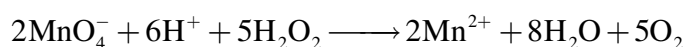
That is why it is stored in plastic bottle or glass bottle covered with wax and in dark place.

17. Hydrogen peroxide is weakly acidic. It is strong oxidising and weak reducing agent. Some examples are as follows -

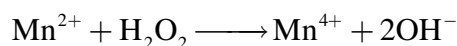
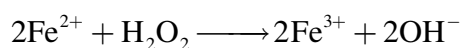
Oxidation in acidic medium:



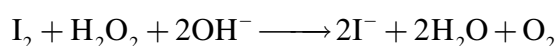
Reduction in acidic medium:



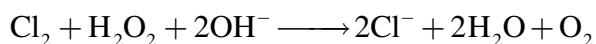
Oxidation in alkaline medium:



Reduction in alkaline medium:



18. Hydrogen peroxide reduces chlorine and for this action it is known as **Antichlor**.

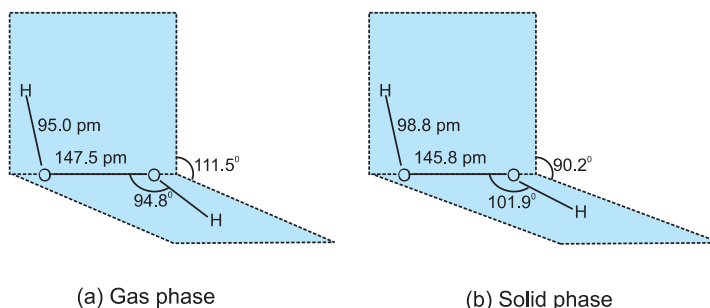


Hydrogen peroxide breaks up to produce nascent oxygen and is used as bleaching agent.

19. Structure of  $\text{H}_2\text{O}_2$ :

Hydrogen peroxide has a non-planar structure in which two hydrogen atoms are placed almost

perpendicular to each other. The structure of hydrogen peroxide in gaseous and solid phase are as follows - (a) Gaseous Phase. (b) Solid Phase



**Fig :** Structure of  $\text{H}_2\text{O}_2$  in (a) Gas phase (b) Solid phase

20. Uses:

Hydrogen peroxide is used as disinfectant and used as antiseptic under the name of Perhydrol. It is used as a bleaching agent for textiles, paper, pulp, leather, oils etc.

21. Prolonged electrolysis of water gives heavy water ( $\text{D}_2\text{O}$ ). Heavy water is harmful for health. It is used as moderator in nuclear reactors, as a tracer compound in many reactions, in NMR spectroscopy and in the preparation of deuterium.

22. Strength of hydrogen peroxide is expressed in volume of 30%  $\text{H}_2\text{O}_2$  solution is represented as '100 vol'  $\text{H}_2\text{O}_2$ , '30 vol' represents at STP 1 lit of  $\text{H}_2\text{O}_2$  solution will give 30 lit of oxygen.

**A. Find out the correct answer:- (MCQ)**

**(Each question 1 mark)**

1. Which one is an Interstitial hydride?

- a) LiH      b)  $\text{B}_2\text{H}_6$       c)  $\text{BeH}_2$       d) LaH

2. Which of the following is not correct regarding hydrogen?

- a) It is colourless, odourless, tasteless gas.  
 b) It has a very low solubility in water.  
 c) It is highly reactive gas  
 d) It forms more compounds than any other element.

3. The adsorption of hydrogen on palladium is known as -

- a) Hydrogenation      b) reduction      c) dehydrogenation      d) Occlusion

4. Temporary hardness of water is due to presence of -

- a)  $\text{MgSO}_4$       b)  $\text{MgCl}_2$       c)  $\text{MgCO}_3$       d)  $\text{Mg}(\text{HCO}_3)_2$

5. The compound that can act both as an oxidising agent and as a reducing agent is:  
a)  $K_2Cr_2O_7$       b)  $H_2S$       c)  $BaO_2$       d)  $H_2O_2$
6. Blackened oil paintings can be restored into original form by the action of -  
a)  $Cl_2$       b)  $BaO_2$       c)  $H_2O_2$       d)  $MnO_2$
7. In which of the following hydrides, hydrogen exists in negative oxidation state-  
a)  $NaH$       b)  $HBr$       c)  $CaH_2$       d)  $HF$
8. Hard water may be softened by using -  
a) Sodium aluminium silicates      b) Calgon  
c) Ion exchange resin      d) All the above
9. Metals which liberates  $H_2(g)$  from acids are:  
a)  $Mg$       b)  $Fe$       c)  $Zn$       d) All the above
10. Heavy water is -  
a) Less Hard      b) Oxygen is  $SP^2$  hybridised.  
c) Not stronger than  $H_2O$       d) Used as moderator and coolant in nuclear reactors
11. Para and ortho hydrogen differ in:  
a) Atomic Number      b) Atomic mass  
c) Spins of protons      d) Number of neutrons.
12. The structure of  $H_2O$  is -  
a) Planer      b) Non-planar      c) Spherical      d) Linear
13. Heavy water is obtained by -  
a) Boiling of water      b) Fractional distillation of water  
c) Prolonged electrolysis of water      d) Heating  $H_2O_2$
14. 30 Volume  $H_2O_2$  means -  
a) 30%  $H_2O_2$   
b) 30  $cm^3$  of the solution, contain 1 gm of  $H_2O_2$   
c) 1  $cm^3$  of solution liberates 30  $cm^3$  of  $O_2$  at STP  
d) 30  $cm^3$  of the solution contains 1 mole of  $H_2O_2$

15. What is false about  $\text{H}_2\text{O}_2$  -
- acts as both oxidising and reducing agent
  - two -OH bonds lies in the same plane.
  - Pure  $\text{H}_2\text{O}_2$  is pale blue coloured.
  - Can be oxidised by  $\text{O}_3$
16. Hydrogen can be prepared by the action of dil  $\text{H}_2\text{SO}_4$  on-
- Copper
  - Iron
  - Lead
  - Mercury
17. Both temporary and permanent hardness of water can be removed by -
- Boiling
  - Boiling with  $\text{Ca}(\text{OH})_2$
  - Adding  $\text{NaCl}$
  - adding  $(\text{NaPO}_3)_6$
18. De-ionized water is produced by:
- Permutit process
  - Calgon process
  - ion-exchange process
  - Clark Process
19. The oxide that gives  $\text{H}_2\text{O}_2$  on treatment with a dilute acid is -
- $\text{PbO}_2$
  - $\text{Na}_2\text{O}_2$
  - $\text{MnO}_2$
  - $\text{TiO}_2$
20. Water has low volatility because:
- Its molecules are joined by inter molecular hydrogen bonds.
  - Its enthalpy of formation is high.
  - It has low molecular weight.
  - It is a covalent compound.

**B. Assertion - Reason type questions:**

**(Each question 1 mark)**

In the following question a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below for each question:

- Both A and B are true and R is the correct explanation of A.
  - Both A and B are true but R is not the correct explanation of A.
  - A is true but R is False.
  - A is false but R is true.
- Assertion: Ortho and Para-hydrogen are nuclear spin isomers.  
Reason: They have same nuclear spin.
  - Assertion: Presence of soluble bicarbonates of calcium, magnesium and iron in water causes temporary hardness.

- Reason: By adding calculated quantity of lime stone to water, temporary hardness can be removed.
3. Assertion:  $\text{H}_2\text{O}_2$  acts as oxidising or a reducing agent.  
Reason:  $\text{H}_2\text{O}_2$  contains Oxygen in the intermediate oxidation state of -1.
  4. Assertion: Nascent hydrogen is more reactive and a powerful reducing agent than ordinary hydrogen.  
Reason: When Zinc and Sulphuric acid are added to yellow  $\text{FeCl}_3$  solution latter is rapidly reduced to colourless solution.
  5. Assertion: The boiling point and heat of vaporisation of water are abnormally higher than those of  $\text{H}_2\text{S}$ ,  $\text{H}_2\text{Se}$  etc.  
Reason: Water molecules have intermolecular hydrogen bonding. These are absent in molecules of  $\text{H}_2\text{S}$ ,  $\text{H}_2\text{Se}$  etc.
  6. Assertion: Some metals like platinum and palladium can be used as storage media for hydrogen.  
Reason: Platinum and palladium can absorb large volumes of hydrogen.
  7. Assertion: Permanent hardness of water is removed by treatment with washing soda.  
Reason: Washing Soda react with soluble magnesium and calcium carbonate to form insoluble carbonates.
  8. Assertion:  $\text{H}_2\text{O}_2$  is better oxidising agent than water.  
Reason: Due to presence of -O-O- bond i.e. peroxide linkage generally  $\text{H}_2\text{O}_2$  is unstable and breaks up to give  $\text{O}_2$ .
  9. Assertion: Water is known to be a universal solvent.  
Reason: Due to high dielectric constant water can dissolve large no of compounds.
  10. Assertion: Hydrogen generally forms covalent compounds.  
Reason: Due to high ionisation enthalpy of hydrogen covalent compounds are easy to form than ionic compounds.

**C. Very short Answer type questions:**

**[Mark of each question=1 marks]**

1. Name one electro positive non-metal.
2. Which isotope of hydrogen is radioactive?
3. Name of compound in which oxidation number of Hydrogen is - 1.
4. Name the isotope of hydrogen which finds use in nuclear reactor.
5. Which one is more reactive Nascent Hydrogen or Normal Hydrogen?
6. Give one example of interstitial hydride.



7. Name the substance used in nuclear reactor as moderator.
8. Give example of electron deficient hydride.
9. Give example of ionic hydride.
10. Give example of molecular hydride.
11. What is permutit?
12. Dihydrogen is not preferred in balloons?
13. How is heavy water obtained from ordinary water?
14. What type of elements form interstitial hydrides?
15. Why pure water does not conduct electricity?
16. What is Calgon?
17. Draw the structure of  $\text{H}_2\text{O}_2$ .
18. What do you mean by "30 Volume" of  $\text{H}_2\text{O}_2$ .
19. Name the metal found in bathing soap.
20. Which is more powerful 3%  $\text{H}_2\text{O}_2$  or "3-Volume"  $\text{H}_2\text{O}_2$ .

**D. Short Answer type questions:**

**(Each question 2 Marks)**

1. Why  $\text{H}^+$  does not exist in free state?
2. What do you mean by hydrogen economy?
3. Give examples of electron deficient and electron precise compounds of hydrogen.
4. Give example of oxidising and reducing properties of hydrogen.
5. Among  $\text{NH}_3$ ,  $\text{H}_2\text{O}$  and  $\text{HF}$ , Which would you expect to have highest magnitude of hydrogen bonding and why?
6. Write chemical reaction to show amphoteric nature of water.

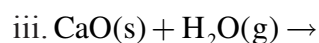
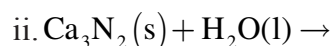
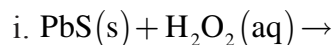
7. What do you mean by non-stoichiometric hydride?
8. Write the names of isotopes of hydrogen.
9. Compare the structure of  $\text{H}_2\text{O}$  and  $\text{H}_2\text{O}_2$ .
10. Why hard water is harmful for steam boiler?
11. Why  $\text{H}_2\text{O}_2$  is stored in plastic or wax coated glass bottle in dark place?
12. What is meant by de-ionised water?
13. What are the reason behind temporary and permanent hardness of water?
14. Write chemical reaction to show amphoteric nature of water.
15. What is auto-protolysis of water?
16. Why ice-float above water?
17. How does  $\text{H}_2\text{O}_2$  behaves as a bleaching agent?
18. How can living beings survive under the ice in winter season?
19. Why demineralised water is not useful for drinking?
20. Soft-water lathers with soap but not hard water?

**E. Short Answer type question:**

**(Each question 3 Marks)**

1. Arrange the following as indicated below:
  - i.  $\text{CaH}_2$ ,  $\text{BeH}_2$  and  $\text{TiH}_2$  according to increasing order of electrical conductance.
  - ii.  $\text{LiH}$ ,  $\text{NaH}$  and  $\text{CsH}$  according to ionic character.
  - iii.  $\text{H-H}$ ,  $\text{D-D}$  and  $\text{F-F}$  according to increasing bond dissociation enthalpy.
2. Write down the chemical equation for the preparation of dihydrogen from (i)  $\text{H}_2\text{O}$ , (ii)  $\text{NaOH}$ , (iii) Acid.
3.
  - i. What is understood by 'hydride gap'?
  - ii. Why do lakes freeze from top towards bottom?
  - iii. How can saline hydrides remove traces of water from organic compounds?
4. Complete and balance the following reaction:
  - i.  $\text{CO}_2(\text{g}) + \text{H}_2(\text{g}) \xrightarrow[\text{Catalyst}]{\text{Cobalt}}$
  - ii.  $\text{Ca}_3\text{N}_2(\text{s}) + \text{H}_2\text{O}(\text{l}) \longrightarrow$
  - iii.  $\text{MnO}_4^{-1}(\text{aq}) + \text{H}_2\text{O}_2(\text{aq}) \longrightarrow$
5. How many types of hydride exist? Give example for each case.
6. Explain the following term -
  - i. Syn-gas
  - ii. Fuel-cell
  - iii. Hydrogen economy

7. Complete the following reaction:



8. Express the strength of 20 Volume hydrogen peroxide in percentage strength.

9. a. What is Merck's Perhydrol?

b. What is meant by "20-Volume"  $\text{H}_2\text{O}_2$ .

10. a. How decomposition of solution of  $\text{H}_2\text{O}_2$  prevented?

b. Write about use of bleaching action of  $\text{H}_2\text{O}_2$ .

**F. Long Answer type question:**

**(Each question 5 Marks)**

1. With suitable chemical reaction show that  $\text{H}_2\text{O}_2$  act as -

i. Acidic nature

ii. Oxidising agent

iii. Reducing Agent.

iv. Bleaching action

v. Antichlor

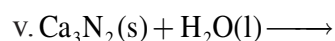
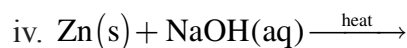
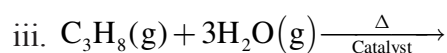
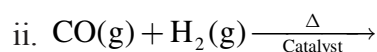
2. How you can prepare de-ionised water? How you can regenerate cation and anion exchange resins?

3. i. What do you mean by "100 volume" hydrogen peroxide?

ii. Describe the structure of common form of ice.

iii. What is the difference between the terms hydrolysis and hydration?

4. Complete the following reactions:



5. i. How temporary hardness of water can be removed using Clark's method?

ii. What is the permutit method of softening of hard water?

6. a. Calculate the strength in volumes of a solution containing 30.36g/ litre of  $\text{H}_2\text{O}_2$ .

b. What are the advantages of hydrogen as a fuel for rockets?

7. a. What causes the temporary and permanent hardness of water?

b. Discuss the method of softening of hard water by synthetic ion exchange resins.

## Solution

### A. Answer the MCQ :

- |        |        |        |        |
|--------|--------|--------|--------|
| 1. d.  | 2. c.  | 3. d.  | 4. d.  |
| 5. d.  | 6. c.  | 7. a.  | 8. d.  |
| 9. d.  | 10. d. | 11. c. | 12. b. |
| 13. c. | 14. d. | 15. b. | 16. b. |
| 17. d. | 18. c. | 19. a. | 20. a. |

### B. Assertion & Reason :

- |     |     |     |     |      |
|-----|-----|-----|-----|------|
| 1-c | 2-b | 3-a | 4-a | 5-a  |
| 6-a | 7-a | 8-a | 9-a | 10-a |

## Chapter at a glance :-

- s-Block elements: The elements of group- 1 and Group - 2 of the periodic table are called the s-block elements. As the last electron goes into s-orbital they are named so..
- The group-I elements (Li, Na, K, Rb, Cs, Fr) are called alkali metals and the group - 2 elements (Be, Mg, Ca, Sr, Ba, Ra) are called alkaline earth metals.
- The general electronic configuration of s-block elements is [noble gas]  $ns^1$  for alkali metals and [noble gas]  $ns^2$  for alkaline earth metals.
- Both the oxides and hydroxides of s-block elements are basic in nature. The alkaline earth metal oxide and hydroxides are less basic than the corresponding alkali metals.
- Alkali metals have lowest ionization enthalpies and largest atomic radius in their respective periods.
- Alkali metals behave as strong reducing agents due to their low ionization enthalpies.
- Metallic character and reactivity of s-block elements increase downwards the group. Due to high reactivity they do not occur free in nature. Compounds formed by s-block elements are mostly ionic in nature except the compounds of Li and Be.
- The first ionization enthalpies of the alkaline earth metals are higher than those of the corresponding group - 1 metals. This is due to their small size as compared to the corresponding alkali metals.
- The hydration enthalpies of both group - 1 and group-2 metal ions decrease with increase in ionic size downwards the group.
- All alkali metals impart characteristic colour to the flame.
 

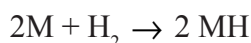
Li - Carmine red,	Na - golden yellow,	K - pale violet
Rb - Reddish violet	Cs - Sky blue	
- Among alkaline earth metals some impart colour during flame test .
 

Ca - brick red	Sr - Crimson	Ba - apple green
----------------	--------------	------------------
- In flame the electrons are excited to higher energy levels and when they drop back to the ground state energy is emitted in the form of visible light. Be and Mg do not impart colour to the flame test due to their small size and high ionization enthalpy. The required high excitation energy is not available in energy of the flame.
- Lithium and beryllium exhibit some difference in properties from those of the other members of the respective group. It is due to their exceptionally small size & high polarizing power (i.e, charge/ radius ratio).
- There are similarities in properties of first elements of both the groups (Li of Gr - I and Be of Gr - 2) with the second element of next group (Mg of Gr-2 and Al of Gr-3) respectively. This type of diagonal similarity is known as diagonal relationship.

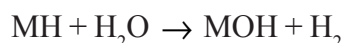
- Alkali metals dissolve in liquid ammonia giving deep blue solutions which contain ammoniated cations and ammoniated electrons
- The alkaline earth metals dissolve in liquid ammonia to give deep blue black solutions forming ammoniated cations & electrons. The solution formed is electrically conductive, paramagnetic in nature and reductive as well. The solvated electrons get absorbed in the visible region and so the solution becomes deep blue black in colour.
- The atomic and ionic radii of the alkaline earth metals are smaller than those of the corresponding alkali metals in the same periods. This is due to the increased nuclear charge in these elements. This causes increase in melting point, boiling point. They have less reactivity than the alkali metals.
- The alkaline earth metals burn in Oxygen to form monoxide (MO). They form salts of Oxoacids like carbonates, sulphates, nitrates like alkali metals.
- Li forms monoxide  $\text{Li}_2\text{O}$ , even with excess of Oxygen. Na forms peroxide  $\text{Na}_2\text{O}_2$ , while K, Rb, Cs form superoxides on reaction with excess oxygen.

#### Reaction with Dihydrogen :

- Alkali metals combine with  $\text{H}_2$  forming ionic hydrides.



- Hydrides of alkali metals are attacked by water to give back hydrogen.



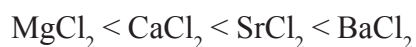
- $\frac{\text{LiH, NaH, KH, RbH, CsH}}{\text{Thermal stability decreases,}} \rightarrow$   
Basic property increases.

- Except Be all the alkaline metals form  $\text{MH}_2$  type hydrides ( $\text{MgH}_2$ ,  $\text{CaH}_2$ ,  $\text{SrH}_2$ ,  $\text{BaH}_2$ ) on heating directly with  $\text{H}_2$ .
- $\text{BeH}_2$  is prepared by action of  $\text{BeCl}_2$  with  $\text{LiAlH}_4$ .  
$$2\text{BeCl}_2 + \text{LiAlH}_4 \rightarrow 2\text{BeH}_2 + \text{LiCl} + \text{AlCl}_3$$
- $\text{BeH}_2$  and  $\text{MgH}_2$  are covalent, others are ionic.

#### Halides :

- Alkali metals react directly with halogen to form ionic halide MX (M - alkali metal ion, X - halide ion)
- Ionic nature of MX increases from  $\text{LiCl}$  to  $\text{CsCl}$ .
- Due to polarization of  $\text{Cl}^-$  ion by small  $\text{Li}^+$  ion,  $\text{LiCl}$  is covalent in nature. So its tendency of hydrolysis is more.
- K, Rb, Cs halides react with more halogens to give poly halides.  
$$\text{KI} + \text{I}_2 \rightarrow \text{KI}_3$$
  
$$\text{CsBr} + \text{Br}_2 \rightarrow \text{CsBr}_3$$
- Alkaline earth metals (M) react with halogen(X) to form halides ( $\text{MX}_2$ ).

- Order of ionic nature of the halides are

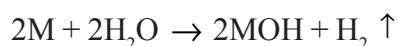


Beryllium halides are essentially covalent and soluble in organic solvents.

- The tendency of hydrolysis of these halides decrease from  $\text{BeCl}_2$  to  $\text{BaCl}_2$  due to decrease in covalent nature.

#### Reaction with water :

- Alkali metals react vigorously with  $\text{H}_2\text{O}$  forming hydroxides with liberation of  $\text{H}_2$  gas.



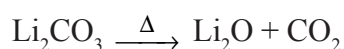
- Alkali earth metals react slowly with  $\text{H}_2\text{O}$  to give metal hydroxides &  $\text{H}_2$  Gas.



- Reactivity with  $\text{H}_2\text{O}$  increases from Li to Cs. Li - is least reactive towards water, Na reacts vigorously, K reacts producing a flame, Rb & Cs reacts explosively.
- Be does not react with water, Mg reacts only with hot water, Ca, Sr, Ba reacts with cold water but not as like as alkali metals.

#### Carbonates :

- All the alkali metals form  $\text{M}_2\text{CO}_3$  type carbonates, where as the alkaline earth metals form  $\text{MCO}_3$  type carbonates.
- Except  $\text{Li}_2\text{CO}_3$ , all the other  $\text{M}_2\text{CO}_3$  type carbonates are stable towards heat.

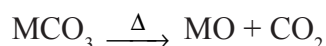


The order of thermal stability is -



Thermal stability of  $\text{M}_2\text{CO}_3 \propto \frac{1}{\text{Polarisation power}}$

- All the carbonates of alkaline earth metals decompose on heating.



The thermal stability increases with increasing cationic size.



#### Sulphates :

- Alkali metals form  $\text{M}_2\text{SO}_4$  type sulphates & alkaline earth metals form  $\text{MSO}_4$  type sulphates. All these sulphates are ionic in nature. The order of ionic nature is –



and



- Solubility of alkaline earth metal carbonates and sulphates in water decrease downwards the group. The size of anions being much larger as compared to cations, the lattice enthalpy will remain almost constant within a particular group, solubility will decrease as found for alkaline earth metal carbonates & sulphates.

### Nitrates :

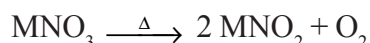
- Alkali metals form  $MNO_3$  type and alkaline earth metals form  $M(NO_3)_2$  type nitrates.

- Thermal stability increases from  $LiNO_3$  to  $CsNO_3$ .

$LiNO_3$  on thermal decomposition gives Lithium Oxide,  $NO_2$  &  $O_2$ .

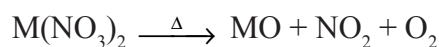


- Other nitrates on thermal decomposition give nitrates & oxygen.



Where  $M = Na, K, Rb, Cs$ .

- All alkaline earth metals nitrates on heating gives metal oxide,  $NO_2$ ,  $O_2$ .



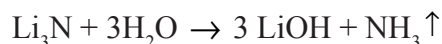
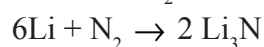
$M = Be, Mg, Ca, Sr, Ba$

- Thermal stability increases from  $Be(NO_3)_2$  to  $Ba(NO_3)_2$ , but these are less stable than alkali metal nitrates due to smaller atomic size of group - 2 elements than gr - 1 elements.

- $Be(NO_3)_2$  forms a layer of  $BeO$  on its surface so reaction stops.

### Nitrides :

- Among group - 1 elements, only  $Li$  reacts directly with  $N_2$  to form nitride which gives  $NH_3$  on reacting with  $H_2O$ .



- Among group - 2 elements only  $Be, Mg$  and  $Ca$  burns in  $N_2$  to give  $M_3N_2$  ( $Be_3N_2, Mg_3N_2, Ca_3N_2$ )

#### ■ Reaction with acids :

Both group - 1 and group - 2 elements react with acids to produce corresponding salt and hydrogen gas.

#### ■ Formation of amalgam :

Both gr - 1 and gr - 2 elements give amalgam with  $Hg$ . These elements also form alloys with other metals.

- **Some important compounds of Na** are  $NaCl, Na_2CO_3, NaHCO_3, NaOH$  etc.  $Na_2CO_3$  can be prepared by solvay process and  $NaOH$  can be prepared by Castner-Kellner process commercially.

- Sodium Chloride is used as common salt for domestic purposes. It is also used for the preparation of  $Na_2O_2, NaOH$  and  $Na_2CO_3$ .

- Sodium hydroxide is used in the manufacture of soap, paper, artificial silk and a number of chemicals. It is used in petroleum refining, purification of bauxite, for the preparation of pure fats, oils and as a laboratory reagent.

- Sodium bi-carbonate or baking soda is used for baking cakes, pastries etc.

- Sodium ions present in our body fluids participate in the transmission of nerve signals, in regulating



the flow of water across cell membranes and in the transport of sugars and amino acids into cells.

- Potassium also plays important role in sodium-potassium pump in our body  
Calcium ion plays important roles in neuro muscular function, interneuronal transmission, cell membrane integrity and blood cogulation.
- Magnesium present in Chlorophyll of green plants helps in absorption of light for photosynthesis.
- Compounds of Calcium like Quick lime, slaked lime, Calcium carbonate, Portland cement are industrially important. Portland cement is one of the most important building materials used for constrution of buildings, roads, dams and bridges.

**A. Choose the correct answer (MCQ) :**

**(Each question 1 mark)**

1. Lithium shows diagonal relationship with  
a) Beryllium    b) Magnesium  
c) Boron        d) Calcium
  
2. Which of the following readily forms nitride?  
a) K                      b) Mg  
c) Ba                     d) Na
  
3. Which alkali metal has lowest melting point?  
a) Na                     b) K  
c) Rb                     d) Cs
  
4. Which of the following alkali metal forms hydrated salt?  
a) Li                      b) Na  
c) K                      d) Cs
  
5. Alkali metals react with water vigorously to form hydroxides and dihydrogen. Which of the following alkali metals react with water least vigorously?  
a) Li                      b) Na  
c) K                      d) Cs
  
6. Which of the following carbonates is most stable thermally?  
a)  $\text{MgCO}_3$               b)  $\text{CaCO}_3$   
c)  $\text{SrCO}_3$                 d)  $\text{BaCO}_3$

7. Solvay's Process is used for the manufacture of  
a) NaOH      b)  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$   
c)  $\text{K}_2\text{CO}_3$       d)  $\text{Na}_2\text{O}_2$
8. Alkali metals give a \_\_\_\_\_ when dissolved in liquid ammonia  
a) Deep Blue solution      b) Colourless  
c) Red Color      d) None of the above.
9. Which of the following metals is not manufactured by electrolysis?  
a) Na      b) Mg  
c) Al      d) Fe
10. CsOH is  
a) Strongly basic      b) weakly basic  
c) slightly acidic      d) amphoteric
11. The salt that is added to table salt to make it flow freely in rainy season is  
a) KCl      b) KI  
c)  $\text{Ca}_3(\text{PO}_4)_2$       d)  $\text{Na}_3\text{PO}_4$
12. Which of the following alkaline earth metals does not impart characteristic colour to the flame?  
a) Be      b) Ca  
c) Ba      d) Sr
13. Which of the following has largest solubility in water?  
a)  $\text{Mg}(\text{OH})_2$       b)  $\text{Ca}(\text{OH})_2$   
c)  $\text{Ba}(\text{OH})_2$       d)  $\text{Sr}(\text{OH})_2$
14. The Alkali metals are low melting which of the following alkali metal is expected to melt if the room temperature is  $30^\circ\text{C}$ ?  
a) Na      b) K  
c) Rb      d) Cs
15. Which of the following metal hydroxide is least basic?  
a)  $\text{Mg}(\text{OH})_2$       b)  $\text{Ca}(\text{OH})_2$

- c)  $\text{Sr}(\text{OH})_2$                       d)  $\text{Ba}(\text{OH})_2$
16. Among the following metal halides, the one which is soluble in ethanol is
- a)  $\text{BeCl}_2$                       b)  $\text{MgCl}_2$   
c)  $\text{CaCl}_2$                       d)  $\text{SrCl}_2$
17. In the synthesis of  $\text{Na}_2\text{CO}_3$  the recovery of  $\text{NH}_3$  is done by treating  $\text{NH}_4\text{Cl}$  solution with  $\text{Ca}(\text{OH})_2$ . The by product obtained in the process is
- a)  $\text{NaCl}$                       b)  $\text{CaCl}_2$   
c)  $\text{NaOH}$                       d)  $\text{NaHCO}_3$
18. The order of decreasing ionisation enthalpy in alkali metal is
- a)  $\text{Na} > \text{Li} > \text{K} > \text{Rb}$                       b)  $\text{Rb} > \text{Na} > \text{K} > \text{Li}$   
c)  $\text{Li} > \text{Na} > \text{K} > \text{Rb}$                       d)  $\text{K} < \text{Li} < \text{Na} < \text{Rb}$
19. Which of the following is not known?
- a)  $\text{KO}_3$                       b)  $\text{KO}_4$   
c)  $\text{KO}_2$                       d)  $\text{K}_2\text{O}_2$
20. Which of the following act as oxidising agent as well as reducing agent?
- a)  $\text{Na}_2\text{O}$                       b)  $\text{Na}_2\text{O}_2$   
c)  $\text{KNO}_3$                       d)  $\text{NaNO}_3$
21. The active constituent of bleaching powder is -
- a)  $\text{Ca}(\text{OCl})_2$                       b)  $\text{Ca}(\text{OCl})\text{Cl}$   
c)  $\text{Ca}(\text{Cl}_2\text{O}_2)$                       d)  $\text{Ca}(\text{ClO}_2)\text{Cl}$
22. Slaked lime reacts with chlorine to give
- a)  $\text{CaO}$                       b)  $\text{Ca}(\text{OCl})_2$   
c)  $\text{CaOCl}_2$                       d)  $\text{CaCO}_3$
23. Which of the following is most stable -
- a)  $\text{LiCl}$                       b)  $\text{LiBr}$   
c)  $\text{LiI}$                       d)  $\text{LiF}$
24. By adding gypsum to cement -
- a) setting time of cement becomes less                      b) colour of cement become light

- c) shining surface is obtained
- d) setting time of cement increases

25. Match the elements of column - I with the colour they impart to the flame given in column - II

Column - I	Column - II
i) Cs	P - apple green
ii) Sr	Q - Brick red
iii) K	R - Blue
iv) Ca	S - Crimson Red
v) Ba	T - Violet

- a) i) - P, ii) - Q, iii) - S, iv) - R, v) - T
- b) i) - R, ii) - S, iii) - T, iv) - Q, v) - P
- c) i) - Q, ii) - P, iii) - R, iv) - S, v) - T
- d) i) - S, ii) - R, iii) - Q, iv) - P, v) - T

**B. Assertion and Reason :**

**(Each question 1 mark)**

- a) Both assertion and reason are correct and reason is the correct explanation of the assertion.
  - b) Both assertion and reason are correct, but reason is not the correct explanation of the assertion.
  - c) Assertion is correct but reason is incorrect.
  - d) Assertion is incorrect but reason is correct.
1. Assertion : s-Block elements do not occur free in nature.  
Reason : s-Block elements are highly electropositive in nature.
  2. Assertion : Potassium and Caesium are useful as electrodes in photoelectric cells.  
Reason : K & Cs when irradiated with light, the light energy absorbed is sufficient to eject out the electron from an atom.
  3. Assertion : The ionisation enthalpies of alkaline earth metals are fairly low.  
Reason : Generally, the low ionisation enthalpies of alkaline earth metals are due to large size of their atoms.
  4. Assertion : LiF is almost insoluble in water  
Reason : LiF is soluble not only in water but also in acetone.
  5. Assertion : The carbonate of lithium decomposes easily on heating to form LiO and CO<sub>2</sub>.  
Reason : Lithium being very small in size, polarizes large carbonate ion leading to the formation of more stable Li<sub>2</sub>O and CO<sub>2</sub>.
  6. Assertion : BeCO<sub>3</sub> is kept in the atmosphere of CO<sub>2</sub>.  
Reason : BeCO<sub>3</sub> is unstable and decomposes to give BeO and CO<sub>2</sub>.

7. Assertion : NaCl when exposed in air it becomes wet.  
Reason : NaCl contains hygroscopic impurities like  $\text{CaCl}_2$ ,  $\text{MgCl}_2$  etc.
8. Assertion :  $\text{Li}_2\text{SO}_4$  do not form double salt like alum.  
Reason : Atomic size of Li is too small.
9. Assertion : Be has complex formation tendency.  
Reason : Be is s-block element.
10. Assertion : Chlorides of Li, Be and Mg are covalent in nature.  
Reason : Li, Be, Mg have large cationic size in the s-block elements.
11. Assertion : In the solution of K in liquid  $\text{NH}_3$ , blue colour appears.  
Reason : K reacts with liquid  $\text{NH}_3$  to form  $\text{KNH}_2$ .
12. Assertion :  $\text{KHCO}_3$  can not be obtained by solvay process.  
Reason :  $\text{KHCO}_3$  is less soluble in water than  $\text{NaHCO}_3$ .

**C. Very Short answer type questions :**

**(Each question one mark)**

1. Write the electronic configuration of s-block elements ?
2. Which groups of periodic table contains s-block elements ?
3. Which alkali metal is the most radioactive element?
4. Which alkali metals are used as electrode in photo electric cell ?
5. What is the colour of potassium superoxide?
6. Why  $\text{KHCO}_3$  is not prepared by Solvay process?
7. Why do alkali metals not occur in free state?
8. Name the metal which floats on water without any apparent reaction with water.
9. Why is  $\text{BeCl}_2$  soluble in organic solvents?
10. Why is  $\text{CaCl}_2$  added to NaCl in extraction of Na by Down cell ?

11. Which member of the alkaline earth metals family has i) least reactivity, ii) lowest density, iii) highest boiling point, iv) maximum reduction potential ?
12. Why is beryllium not attacked by an acid easily?
13. Lithium is the only alkali metal to form a nitride directly - Why?
14. Why are peroxides and superoxides stable in comparison to other oxides?
15. State one reason to explain why alkaline earth metals have more tendency to form complexes?

**D. Short Answer Type Question :**

**(Each question 2 marks)**

1. Why is first ionization enthalpy of alkali metals lower than those of alkaline earth metals?
2. Alkali metals are strong reducing agents, why?
3. What is the cause of diagonal relationship?
4. Why caesium can be used in photoelectric cell while lithium cannot be?
5. Write three general characteristics of s-block elements which distinguish them from the elements of other block.
6. The stability of peroxides and superoxides of alkali metals increases as we go down the group. Explain.
7. Draw the structure of -
  - i)  $\text{BeCl}_2$  (vapour)
  - ii)  $\text{BeCl}_2$  (solid)
8. Alkali and alkaline earth metals can't be obtained by chemical reduction method - Why?
9. Why do Be and Mg not impart colour to the flame in the flame test?
10. Complete the following -
  - i)  $\text{O}_2^{2-} + \text{H}_2\text{O} \rightarrow$
  - ii)  $\text{O}_2^- + \text{H}_2\text{O} \rightarrow$
11. Why does common salt become wet in the rainy season?
12. Why do the solubility of carbonates and sulphates of alkali metals decrease down the group?
13. Why are lithium salts commonly hydrated and those of other alkali ions usually anhydrous?
14. How would you explain the following?
  - i) BeO is almost insoluble but  $\text{BeSO}_4$  is soluble in nature.
  - ii) BaO is soluble but  $\text{BaSO}_4$  is insoluble in water.

15. Comment on the trend of the following -

- i) Thermal stability of carbonates of Group - 2 elements.
- ii) The solubility and the nature of oxides of Group - 2 elements.

16.  $E^{\circ}$  for  $M_{(aq)}^{2+} + 2e \rightarrow M(s)$  where  $M = Ca, Sr$  or  $Ba$  is nearly constant. How can you account for it?

**E. Short Answer type Question :**

**(Each question 3 marks)**

1. State as to why?

- i) a solution of  $Na_2CO_3$  is alkaline?
- ii) alkali metals are prepared by electrolysis of their fused chlorides?
- iii) Sodium is found to be more useful than Potassium.

2. Arrange the following in the decreasing order of the property mentioned:-

- i)  $Li^+, Na^+, K^+, Rb^+$  (ionic mobility)
- ii)  $Be, Mg, Ca, Sr$  (melting point)
- iii)  $BeO, MgO, CaO$  (enthalpy of formation)

3. Explain the significance of  $Na, K, Mg$  &  $Ca$  in biological fluids.

4. Why are  $BeSO_4$  and  $MgSO_4$  readily soluble in water while  $CaSO_4, SrSO_4$  &  $BaSO_4$  are insoluble?

5. An element of group-2 forms covalent oxide which is amphoteric in nature and dissolves in water to give an amphoteric hydroxide. Identify the element and write chemical reactions of the hydroxide of the element with an alkali and an acid.

6. When water is added to a compound (A) of  $Ca$ , solution of compound (B) is formed, when  $CO_2$  gas is passed into the solution, it turns milky due to formation of compound (C). If excess of  $CO_2$  is passed into the solution milky solution disappears due to formation of compound (D). Identify the compounds A, B, C, D. Explain why the milky solution disappears in the last step.

7. Ions of an element of group - I participate in the transmission of nerve signals and transportation of sugars and aminoacids in cells. This element imparts yellow colour on flame test and form an oxide and a peroxide on reaction with oxygen. Identify the element, write chemical equation to show the reaction of its peroxide with water.

8. Element A of group-2 burns to nitrogen to give an ionic compound B. B reacts with water to give C & D. Solution of C becomes milky on bubbling  $CO_2$ . Identify A, B, C, D.

9. 'The Chemistry of Beryllium is not essentially ionic'. Justify the statement by making a reference to the nature of oxide, chloride and fluoride of Beryllium.

**F. Long Answer Type-Question :**

**(Each question 5 marks )**

1. Present a comparative account of the alkali and alkaline earth metals with respect to the following characteristics -
  - i) Tendency to form ionic / covalent compounds
  - ii) Nature of oxides and their solubility
  - iii) formation of Oxosalts
  - iv) solubility of Oxosalts
  - v) Thermal stability of Oxosalts.
2. The stability of peroxide and superoxides of alkali metals increase as we go down the group. Explain.
3. Lithium Hydride can be used to prepare other useful hydrides.  $\text{BeH}_2$  is one of them. Suggest a route for the preparation of  $\text{BeH}_2$  from  $\text{LiH}$ . Write chemical equations involved in the process.

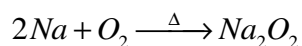
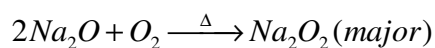
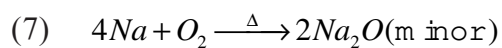
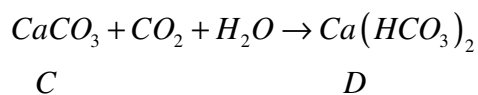
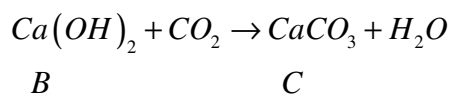
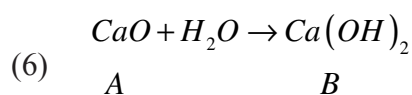
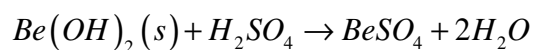
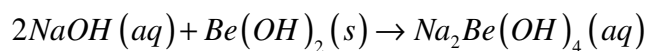
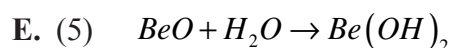


## Solution

- A.**
- |         |         |         |        |
|---------|---------|---------|--------|
| 1 - b,  | 2 - b,  | 3 - d,  | 4 - a  |
| 5 - a,  | 6 - d,  | 7 - b,  | 8 - a  |
| 9 - d,  | 10 - a, | 11 - c, | 12 - d |
| 13 - c, | 14 - d, | 15 - a, | 16 - a |
| 17 - b, | 18 - c, | 19 - b, | 20 - b |
| 21 - b, | 22 - c, | 23 - d, | 24 - d |
| 25 - b. |         |         |        |

**B. Assertion and Reason :**

- |       |        |        |        |
|-------|--------|--------|--------|
| 1 - b | 2 - a  | 3 - a  | 4 - c  |
| 5 - a | 6 - a  | 7 - a  | 8 - a  |
| 9 - b | 10 - c | 11 - b | 12 - c |



Chapter at a glance:

- ◆ The elements in which last electron enters the outermost shell are called p-block elements.
- ◆ The p-block of the periodic table comprise of the gr-13 to gr-18 elements. These elements together with the s-block elements are called Representative elements.
- ◆ p-block of the periodic table is unique in terms of having all types of elements - metal, non-metal and metalloids.
- ◆ Their valence shell electronic configuration is  $ns^2np^{1-6}$  (except -He). Differences in their inner core of the electronic configuration greatly influence their physical and chemical properties. As a consequence of this, a lot of variation in properties among these elements is observed. In addition to the group oxidation state, these elements show other oxidation states as shown in the table.

**Table 11.1 General Electronic Configuration and Oxidation States of p-Block Elements**

Group	13	14	15	16	17	18
General electronic configuration	$ns^2np^1$	$ns^2np^2$	$ns^2np^3$	$ns^2np^4$	$ns^2np^5$	$ns^2np^6$ ( $1s^2$ for He)
First member of the group	B	C	N	O	F	He
Group oxidation state	+3	+4	+5	+6	+7	+8
Other oxidation states	+1	+2, -4	+3, -3	+4, +2, -2	+5, +3, +1, -1	+6, +4, +2

- ◆ In case of a particular group for p-block elements, the highest oxidation state is, (Group Number - 10). For example - for group - 13 elements the highest oxidation state is (13 - 10 = +3). The stability of highest oxidation state of a particular group decreases down a group due to inert pair effect.
- ◆ The compounds formed by highly reactive non-metals with highly reactive metals are generally ionic because of large differences in their electronegativities.
- ◆ Compounds formed by the elements with medium electronegativity difference form polar covalent bonds and have polymeric structures.
- ◆ Compounds formed between non-metals themselves are largely covalent because of small differences in their electronegativities.

- ◆ The combined effect of size and availability of d-orbitals considerably influences the ability of these elements to form  $\pi$  bonds. The lighter elements form  $p\pi - p\pi$  bonds, the heavier ones form  $d\pi - p\pi$  or  $d\pi - d\pi$  bonds.
- ◆ Absence of d-orbitals in second period elements limits their covalency to 4 ( using 2s and three 2p orbitals). In contrast, heavier elements can expand their covalence above four, by using vacant d-orbitals lying between 3p and 4s orbitals.

### Gr - 13 elements:

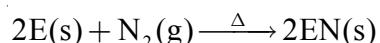
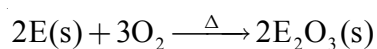
**Table 11.2 Atomic and Physical Properties of Group 13 Elements**

Property	Element					
	Boron B	Aluminium Al	Gallium Ga	Indium In	Thallium Tl	
Atomic number	5	13	31	49	81	
Atomic mass(g mol <sup>-1</sup> )	10.81	26.98	69.72	114.82	204.38	
Electronic Configuration	[He]2s <sup>2</sup> 2p <sup>1</sup>	[Ne]3s <sup>2</sup> 3p <sup>1</sup>	[Ar]3d <sup>10</sup> 4s <sup>2</sup> 4p <sup>1</sup>	[Kr]4d <sup>10</sup> 5s <sup>2</sup> 5p <sup>1</sup>	[Xe]4f <sup>14</sup> 5d <sup>10</sup> 6s <sup>2</sup> 6p <sup>1</sup>	
Atomic radius/pm <sup>a</sup>	(88)	143	135	167	170	
Ionic radius M <sup>3+</sup> /pm <sup>b</sup>	(27)	53.5	62.0	80.0	88.5	
Ionic radius M <sup>+</sup> /pm	-	-	120	140	150	
Ionization enthalpy (kJ mol <sup>-1</sup> )	$\Delta_i H_1$	801	577	579	558	589
	$\Delta_i H_2$	2427	1816	1979	1820	1971
	$\Delta_i H_3$	3659	2744	2962	2704	2877
Electronegativity <sup>c</sup>	2.0	1.5	1.6	1.7	1.8	
Density /g cm <sup>-3</sup> at 298 K	2.35	2.70	5.90	7.31	11.85	
Melting point / K	2453	933	303	430	576	
Boiling point / K	3923	2740	2676	2353	1730	
E <sup>⊖</sup> / V for (M <sup>3+</sup> /M)	-	-1.66	-0.56	-0.34	+1.26	
E <sup>⊖</sup> / V for (M <sup>+</sup> /M)	-	+0.55	-0.79(acid)	-0.18	-0.34	

<sup>a</sup>Metallic radius, <sup>b</sup> 6-coordination, <sup>c</sup> Pauling scale, -1.39(alkali)

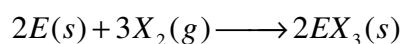
- ◆ Group - 13 elements are Boron (B), Aluminium (Al), Gallium (Ga), Indium (In) and Thallium (Tl).
- ◆ On moving down the group, for each successive member one extra shell of electrons is added so atomic radius is expected to increase. However a deviation can be seen. Atomic radius of Ga(135 pm) is less than that of Al (143 pm). The presence of additional 10 d-electrons offer only poor screening effect for the outer electrons from the increased nuclear charge in gallium.
- ◆ B and Al show oxidation state of +3 only, while Ga, In, Tl show oxidation state of both +1 and +3. Due to inert pair effect, as we move down the group, the stability of +3 oxidation state decreases, while that of +1 oxidation state increases. The compounds in +1 oxidation state are more ionic than those in +3 oxidation state.

- ◆ Due to small size of Boron, the sum of its first three ionization enthalpies is very high. This prevents it to form +3 ions and forces it to form only covalent compounds.
- ◆ Boron and aluminium react with oxygen and nitrogen at high temperature to form oxides and nitrides respectively. Oxide of Boron is acidic, oxides of Al and Ga are amphoteric, oxides of In And Tl are basic in nature.

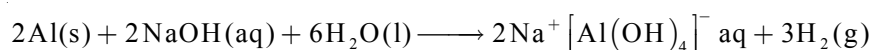


here E = element

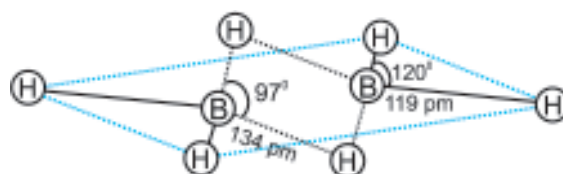
- ◆ Gr - 13 elements react with halogen to form trihalides (Except  $TlI_3$ ).  
(here, X = F, Cl, Br, I)



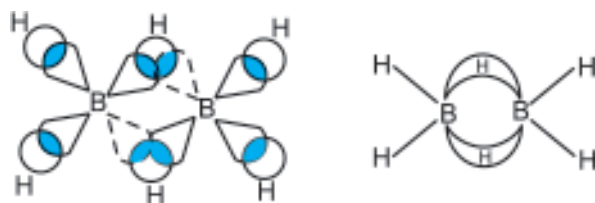
- ◆ Boron is non-reactive at moderate temperature to acids and alkalies. Al reacts with both acids and bases, which shows its amphoteric nature.



- ◆ Boron, the first member of group- 13, differs in many respects from the other members of the group due to small size, high electronegativity and absence of d-orbital. The first member of each group of p-block differs from rest members of their group.
- ◆ Boron is a typical non-metal and the other members are metals. The availability of three valence electrons ( $2s^2 2p^1$ ) for covalent bond formation using four orbitals ( $2s, 2p_x, 2p_y, 2p_z$ ) leads to the electron deficiency in Boron compounds. This deficiency makes them good electron acceptor. So they behave as Lewis acid.
- ◆ Boron forms covalent molecular hydride with dihydrogen as boranes, the simplest form of which is Diborane,  $B_2H_6$ . In addition to four terminal two centre two electron B - H bonds. Diborane contains two bridging hydrogen atoms between two boron atoms. These bridge bonds (B - H - B) are three centred - two electron bonds, which are called bent or banana bond.

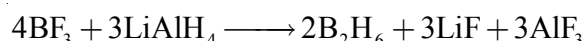


**Fig.** The structure of diborane,  $B_2H_6$

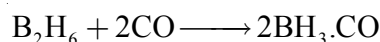
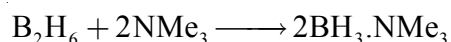


**Fig.** Bonding in diborane. Each B atom ..... banana bonds.

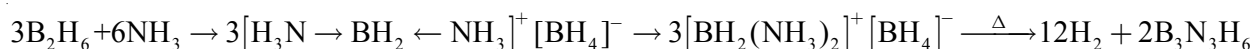
- ◆ Diborane is prepared by treating boron trifluoride with  $\text{LiAlH}_4$  in diethyl ether.



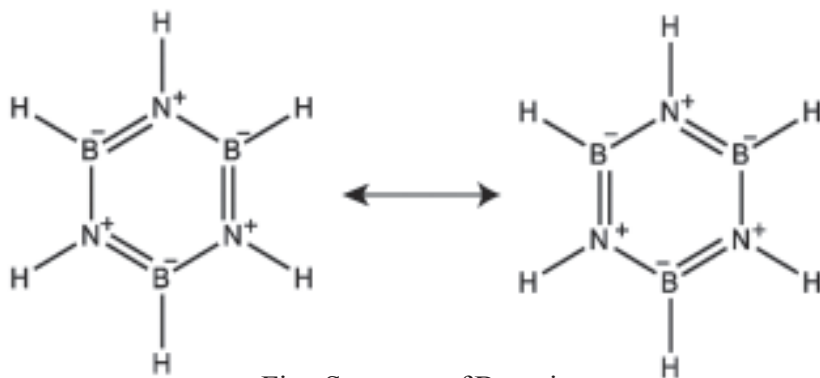
- ◆ Diborane undergoes cleavage reactions with Lewis bases (L) to give borane adducts.



- ◆ Reaction of  $\text{NH}_3$  with  $\text{B}_2\text{H}_6$  gives initially  $\text{B}_2\text{H}_6 \cdot 2\text{NH}_3$  which is formulated as  $[\text{BH}_2(\text{NH}_3)_2]^+ [\text{BH}_4]^-$ ; further heating gives borazine,  $\text{B}_3\text{N}_3\text{H}_6$ .
- ◆  $\text{B}_3\text{N}_3\text{H}_6$  has ring structure with alternate BH and NH groups. So it is known as Inorganic Benzene.

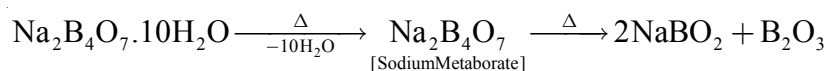


$\text{B}_3\text{N}_3\text{H}_6$  has ring structure with alternate BH and NH groups. So it is known as Inorganic Benzene.



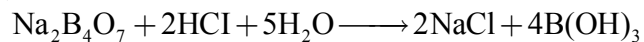
**Fig.:** Structure of Borazine

- ◆ The most important compound of Boron is Borax ( $\text{Na}_2\text{B}_4\text{O}_6 \cdot 10\text{H}_2\text{O}$ ). It is a white crystalline solid. On heating borax first loses water molecules and swells up. On further heating, it turns into a transparent liquid, which solidifies into glass like material known as borax bead.



The metaborates of many transition metals have characteristic colours and so borax bead test can be used to identify them in the laboratory.

- ◆ Orthoboric acid ( $\text{H}_3\text{BO}_3$ ) is a white crystalline solid with soapy touch. It can be prepared by acidifying an aqueous solution of borax.



It has a layer structure in which planar  $\text{B}(\text{OH})_3$  units are joined by hydrogen bonds.

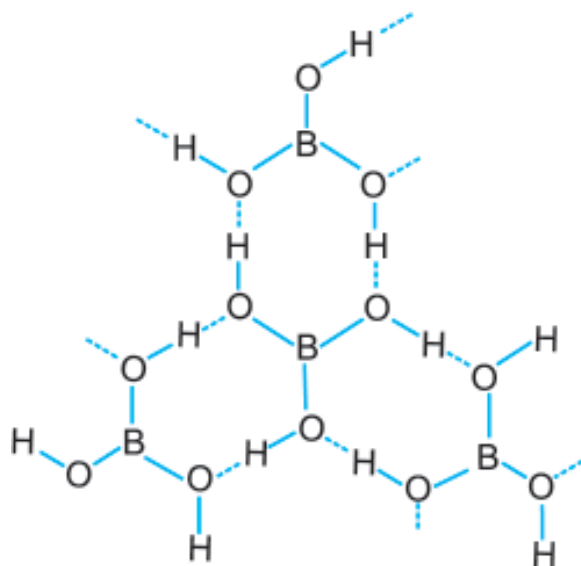
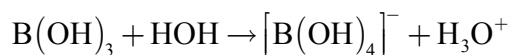


Fig. - Structure of Boric acid, the dotted lines represent hydrogen bonds.

- ◆ Boric acid is a weak monobasic acid. It is not a protonic acid but acts as a lewis acid by accepting electrons from a hydroxyl ion.



#### Group - 14 Elements:

- ◆ Carbon(C), Silicon(Si), Germanium (Ge), Tin (Sn), Lead (Pb) are the members of Group - 14. Carbon is a non metal, Si and Ge are metalloids, Sn and Pb are metals.

**Table 11.3 Atomic and Physical Properties of Group 14 Elements**

Property	Element					
	Carbon C	Silicon Si	Germanium Ge	Tin Sn	Lead Pb	
Atomic Number	6	14	32	50	82	
Atomic mass (g mol <sup>-1</sup> )	12.01	28.09	72.60	118.71	207.2	
Electronic configuration	[He]2s <sup>2</sup> 2p <sup>2</sup>	[Ne]3s <sup>2</sup> 3p <sup>2</sup>	[Ar]3d <sup>10</sup> 4s <sup>2</sup> 4p <sup>2</sup>	[Kr]4d <sup>10</sup> 5s <sup>2</sup> 5p <sup>2</sup>	[Xe]4f <sup>14</sup> 5d <sup>10</sup> 6s <sup>2</sup> 6p <sup>2</sup>	
Covalent radius/pm <sup>a</sup>	77	118	122	140	146	
Ionic radius M <sup>4+</sup> /pm <sup>b</sup>	–	40	53	69	78	
Ionic radius M <sup>2+</sup> /pm <sup>b</sup>	–	–	73	118	119	
Ionization enthalpy/ kJ mol <sup>-1</sup>	$\Delta_i H_1$	1086	786	761	708	715
	$\Delta_i H_2$	2352	1577	1537	1411	1450
	$\Delta_i H_3$	4620	3228	3300	2942	3081
	$\Delta_i H_4$	6220	4354	4409	3929	4082
Electronegativity <sup>c</sup>	2.5	1.8	1.8	1.8	1.9	
Density <sup>d</sup> /g cm <sup>-3</sup>	3.51 <sup>e</sup>	2.34	5.32	7.26 <sup>f</sup>	11.34	
Melting point/K	4373	1693	1218	505	600	
Boiling point/K	–	3550	3123	2896	2024	
Electrical resistivity/ ohm cm (293 K)	10 <sup>14</sup> –10 <sup>16</sup>	50	50	10 <sup>-5</sup>	2 × 10 <sup>-5</sup>	

<sup>a</sup>for M<sup>IV</sup> oxidation state; <sup>b</sup>6-coordination; <sup>c</sup>Pauling scale; <sup>d</sup>293 K; <sup>e</sup>for diamond; for graphite, density is 2.22; <sup>f</sup> $\beta$ -form (stable at room temperature)

- ◆ There is a considerable increase in covalent radius from carbon to silicon, but a small increase from silicon to lead. This is due to the poor shielding effect of completely filled d and f-orbitals.
- ◆ The common oxidation states exhibited by these elements are +4 and +2. Tendency to show +2 oxidation state increases from Ge to Pb. Stability of +4 oxidation state decreases down the group. Pb is stable in +2 oxidation state and acts as an oxidising agent in +4 oxidation state.
- ◆ Presence of d-orbital in elements apart from carbon increases their tendency to form complexes by accepting electron pairs from donors and their halides undergo hydrolysis.
- ◆ Like first member of other groups, Carbon differs from rest of the members of the group. It is due to its smaller size, high electronegativity, higher ionisation enthalpy and unavailability of d-orbitals.
- ◆ Carbon has the unique property of forming p $\pi$ -p $\pi$  multiple bonds with itself and with other atoms of small size and high electronegativity.  
For example: C = C, C $\equiv$ C, C = O, C = S, C $\equiv$ N.  
Heavier elements do not form p $\pi$ -p $\pi$  bonds because their atomic orbitals are too large and diffused to have effective overlapping.

- ◆ Carbon atoms have the tendency to link with one another through covalent bonds to form chains and rings. This property is called catenation.
- ◆ The tendency of catenation decreases as  $C \gg Si > Ge \sim Sn$  due to increase in atomic size and decrease in electronegativity of the elements.
- ◆ Due to catenation and formation of  $p\pi-p\pi$  multiple bonds, allotropic forms of carbon (diamond, graphite, fullerene) exist.

<b>Diamond</b>	<b>Graphite</b>	<b>Fullerence</b>
1. It has crystalline lattice	1. It has a layered structures	1. It has a spherical shape like football
2. It is made up of tetrahedral units	2. It has planar hexagonal rings	2. It has a cage-like structure which looks like ultimately spherical.
3. In diamond each C-atom is $sp^3$ hybridised and is bonded to four other C-atom through a sigma bond.	3. In graphite each C-atom is $sp^2$ - hybridised & is bounded to three adjacent C-atoms through sigma bond. The fourth electron forms a pi-bond.	3. In fullerene each C-atom is $sp^2$ -hybridised and is bonded to three adjacent C-atom through sigma bond. The fourth electron is delocalised in molecular orbitals, which in turn give aromatic character to the molecule.
4. The C - C bond length is 154 pm	4. The C - C bond length is 141.5 pm	4. It has both single and double bonds with C-C bond length 143.5 pm and C = C bond length 138.3 pm
5. It has rigid covalent bond network which is difficult to break.	5. It is quite soft and its layers can be separated easily.	5. It has both six-membered and five-membered rings. A six-membered ring can fuse with a 6-membered or a 5 -membered ring. But a 5-membered ring can only fuse with a 5-membered ring.
6. It acts as an electrical insulator	6. It is a good conductor of electricity	6. It is a good conductor of electricity, but not as good as graphite.





**A. Select the correct answer (MCQ) :**

**(Each question 1 mark)**

- An aqueous solution of borax is  
a) Neutral      b) amphoteric      c) Basic      d) Acidic
- Boric acid is polymeric due to -  
a) Its acidic nature      b) the presence of hydrogen bonds  
c) Its monobasic nature      d) its geometry
- The type of hybridisation of boron in diborane is -  
a) sp      b) sp<sup>2</sup>      c) sp<sup>3</sup>      d) dsp<sup>2</sup>
- Thermodynamically the most stable form of carbon is -  
a) Diamond      b) Graphite      c) Fullerenes      d) Coal
- In the following metals which one has highest abundance in earth's crust?  
a) Al      b) Ca      c) Mg      d) Na
- Which of the following has least Lewis base character?  
a) BF<sub>3</sub>      b) PF<sub>3</sub>      c) CO      d) F<sup>-</sup>
- The hybridisation of C in graphite and diamond are -  
a) sp<sup>3</sup>, sp<sup>3</sup>      b) sp<sup>3</sup>, sp<sup>2</sup>      c) sp<sup>2</sup>, sp<sup>2</sup>      d) sp<sup>2</sup>, sp<sup>3</sup>
- Which does not show inert pair effect?  
a) Al      b) Sn      c) Pb      d) Tl
- Which pairs of compounds do not have similar structure?  
a) BN, Graphite      b) BF<sub>3</sub>, BCl<sub>3</sub>      c) B<sub>2</sub>H<sub>6</sub>, C<sub>2</sub>H<sub>6</sub>      d) B<sub>3</sub>N<sub>3</sub>H<sub>6</sub>, C<sub>6</sub>H<sub>6</sub>
- Carborundum is -  
a) SiC      b) AlCl<sub>3</sub>      c) Al<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub>      d) Al<sub>2</sub>O<sub>3</sub>
- Arrange the following in order of their increasing boiling point.  
CH<sub>4</sub>, SiH<sub>4</sub>, GeH<sub>4</sub>, SnH<sub>4</sub>

- a)  $\text{CH}_4 < \text{SnH}_4 < \text{GeH}_4 < \text{SiH}_4$   
 c)  $\text{CH}_4 < \text{SnH}_4 < \text{GeH}_4 < \text{SnH}_4$
- b)  $\text{SiH}_4 < \text{SnH}_4 < \text{GeH}_4 < \text{CH}_4$   
 d)  $\text{CH}_4 < \text{SiH}_4 < \text{SnH}_4 < \text{GeH}_4$
12. Boric acid acts as an acid because its molecules:  
 a) contain replaceable H-atom  
 b) can accept  $\text{OH}^-$  from water releasing  $\text{H}^+$  ion  
 c) proton producer  
 d) combines with proton from water
13. The tendency of catenation in group - 14 elements follows the order -  
 a)  $\text{C} > \text{Si} > \text{Ge} > \text{Sn}$   
 c)  $\text{C} \gg \text{Si} > \text{Ge} \approx \text{Sn}$
- b)  $\text{Si} > \text{C} > \text{Sn} > \text{Ge}$   
 d)  $\text{Ge} > \text{Si} > \text{Sn} > \text{C}$
14. Silicon has a strong tendency to form silicones. The chain length of silicone polymer can be controlled by  
 a)  $\text{CH}_3\text{SiCl}$       b)  $\text{Me}_3\text{SiCl}$       c)  $\text{Me}_4\text{Si}$       d)  $\text{Me}_2\text{SiCl}_2$
15. Which of the following oxide is acidic in nature?  
 a)  $\text{B}_2\text{O}_3$       b)  $\text{Al}_2\text{O}_3$       c)  $\text{Ge}_2\text{O}_3$       d)  $\text{Sn}_2\text{O}_3$
16. Which is a Lewis acid?  
 a)  $\text{AlCl}_3$       b)  $\text{MgCl}_2$       c)  $\text{CaCl}_2$       d)  $\text{BaCl}_2$
17. Among the following which is the most commonly used reducing agent?  
 a)  $\text{SnCl}_2$       b)  $\text{AlCl}_3$       c)  $\text{PbCl}_2$       d)  $\text{SnCl}_4$
18. Dry ice is  
 a) Solid  $\text{NH}_3$       b) Solid  $\text{SO}_2$       c) Solid  $\text{N}_2$       d) Solid  $\text{CO}_2$
19. An example of a double salt is  
 a) Bleaching powder      b) Hypo  
 c) Potassium Ferrocyanide      d) Potash Alum
20. The substance used as a smoke screen in warfare is -  
 a)  $\text{SiCl}_4$       b)  $\text{PH}_3$       c)  $\text{PCl}_5$       d)  $\text{C}_2\text{H}_2$

21. Which type of bond binds the carbon atoms in graphite?  
 a) Covalent      b) Vander Walls      c) Dipolar      d) Ionic
22. Al-metal is extracted by electrolysis of molten mixture of -  
 a)  $\text{Al}_2\text{O}_3 + \text{HF} + \text{NaAlF}_4$       b)  $\text{Al}_2\text{O}_3 + \text{CaF}_2 + \text{NaAlF}_4$   
 c)  $\text{Al}_2\text{O}_3 + \text{Na}_3\text{AlF}_6 + \text{CaF}_2$       d)  $\text{Al}_2\text{O}_3 + \text{KF} + \text{Na}_3\text{AlF}_6$
23. Which of the following structure is similar to graphite -  
 a)  $\text{B}_3\text{N}_3\text{H}_6$       b)  $\text{B}_4\text{C}$       c)  $\text{B}_2\text{H}_6$       d) BN
24. In Borax bead test which compound is formed -  
 a) Ortho borate      b) Meta borate      c) Per borate      d) Tetra borate
25. Natural source of Boron is  
 a) Bauxite      b) Borax      c) Borazine      d) Borazole

**B. Assertion and Reason:**

**(Each question 1 mark)**

**Direction:**

In the following questions a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below in each question.

- a) Both A and R are correct and R is the correct explanation of A.  
 b) Both A and R are correct but R is not the correct explanation of A.  
 c) A is correct and R not correct.  
 d) A is not correct but R is correct.
- Assertion (A): Silicones are water repelling in nature.  
 Reason (R): Silicones are organosilicon polymers, which have  $(-\text{R}_2\text{SiO}-)$  as repeating unit.
  - Assertion (A): If Al atoms replace a few Si atoms in three dimensional network of  $\text{SiO}_2$ , the overall structure acquires a negative charge.  
 Reason (R): Al is trivalent while Si is tetravalent.
  - Assertion (A):  $\text{SF}_6$  can't be hydrolysed but  $\text{SF}_4$  can be.  
 Reason (R): Six F atoms in  $\text{SF}_6$  prevent the attack of  $\text{H}_2\text{O}$  on the sulphur atom of  $\text{SF}_6$ .
  - Assertion (A): Diborane forms three-centered two electron bonds.  
 Reason (R):  $\text{B}_2\text{H}_6$  does not have sufficient number of electrons to form normal covalent bonds.
  - Assertion (A): Diamond is a bad conductor of electricity.  
 Reason (R): All C - C bond lengths in diamond are of 154 pm.

6. Assertion (A): Boron does not form  $\text{BF}_6^{3-}$  ion while  $\text{AlF}_6^{3-}$  is known.  
Reason (R): B does not have d-orbitals in valence shell while Al has d-orbitals in valence shell.
7. Assertion (A): Glass is not an example of silicates.  
Reason (R): All silicates have tetrahedral  $\text{SiO}_4^{4-}$  units.
8. Assertion (A):  $\text{CO}_2$  has linear geometry involving  $\text{sp}^2$  hybridisation of carbon.  
Reason (R): Solid  $\text{CO}_2$  is known as Dry Ice.
9. Assertion (A): Boron has low electrical conductivity.  
Reason (R): At ordinary temperature boron behaves as metal.
10. Assertion (A): Boric acid behaves as weak monobasic acid.  
Reason (R): Boric acid has only one ionisable hydrogen atom.

**C. Very short answer type questions:**

**(Each question 1 mark)**

- Write general electronic configuration of p-block elements.
- Why do  $\text{BCl}_3$ ,  $\text{AlCl}_3$  behave as Lewis acid?
- Name the simplest hydride of Boron.
- What is inorganic graphite?
- What is catenation?
- Which is the purest allotrope of carbon ?
- Name the man made silicon compounds.
- Between  $\text{AlF}_3$  and  $\text{AlCl}_3$  which one has high melting point?
- Why does boron not form  $\text{B}^{3+}$  ion?
- Why boron and silicon have similar properties?
- What is carbogen?
- What are the isotopes of boron?
- Why is Boron unable to form  $\text{BF}_6^{3-}$  ion?
- Write the chemical formula of 'Inorganic Benzene'?
- Write two examples of electron deficient molecules.
- Why do Boron have unusual high melting point?
- Why do third-period elements expand their covalence above four?
- Why do the element of group -13 are called p-block elements?
- How does metallic and non-metallic character vary in a group?
- Why  $\text{CCl}_4$  behaves as an electron precise molecule?
- Why do lead remain unaffected by water ?
- Name the metalloid found in Group - 14.
- What is silica-gel used as?

24. Write the resonance structure of  $\text{CO}_2$ .
25. What is the common name of recently developed allotrope of carbon (i.e.  $\text{C}_{60}$ ) molecule?

**D. Short answer type questions: (each carry 2 marks)**

1. Is boric acid a protic acid? Explain.
2. Assign the hybridisation of B in  $\text{BF}_3$  and  $\text{BH}_4^-$ .
3. Write the resonance structure of i)  $\text{CO}_3^{2-}$  ii)  $\text{HCO}_3^-$
4. If B-Cl bond has a dipole moment, explain why  $\text{BCl}_3$  molecule has zero dipole moment ?
5. Why the B-F bond lengths in  $\text{BF}_3$  (130pm) and  $\text{BF}_4^-$  (143 pm) differ?
6. What do you mean by inert pair-effect?
7. Why there is a decrease in ionization enthalpy from carbon to silicon ?
8. What is producer gas? How is it prepared commercially ?
9. What are silicones?
10. What is hybridisation of carbon atoms in fullerene structure? Write about its structure.
11. Write two uses for each of Quartz and silica gel.
12. How can carbon monoxide form metal carbonyls ?
13. What are the uses of Zeolites?
14. Write fumes appear around the bottle of anhydrous  $\text{AlCl}_3$  - Why?
15. What happens when boric acid is heated?
16. Why does carbon show anomalous behaviour from other group -14 elements?
17. Diamond is covalent, yet it has high melting point. Why ?
18. How can you explain higher stability of  $\text{BCl}_3$  as compared to  $\text{TiCl}_3$  ?
19. Why is carbon monoxide poisonous ?
20. Explain why -
  - a) Si can form  $\text{SiF}_6^{2-}$ , but corresponding fluoro compound of carbon is not known.

**E. Long answer type questions: (3 marks each)**

1. What are electron deficient compounds? Are  $\text{BCl}_3$  and  $\text{SiCl}_4$  electron deficient species? Explain.
2. Classify the following oxides as neutral, acidic, basic or amphoteric ?  
 $\text{CO}$ ,  $\text{B}_2\text{O}_3$ ,  $\text{SiO}_2$ ,  $\text{Al}_2\text{O}_3$ ,  $\text{PbO}_2$ ,  $\text{Ti}_2\text{O}_3$
3. Boron fluoride exists as  $\text{BF}_3$  but boron hydride does not exist as  $\text{BH}_3$ . In which form does it exist. Give reason.
4. A compound (A) of boron reacts with  $\text{NMe}_3$  to give an adduct (B), which on hydrolysis give a compound (C),  $\text{NMe}$  and  $\text{H}_2$  gas. Compound C is an acid. Identify A, B, C. Give reactions involved.
5. What happens when -

- i. Al is treated with dilute NaOH
  - ii. CO is heated with ZnO
  - iii. SiO<sub>2</sub> is treated with HF.
6. Identify the compounds A, X, Z in the following reactions -
- i.  $A + 2HCl + 5H_2O \longrightarrow 2NaCl + X$
  - ii.  $X \xrightarrow[370K]{\Delta} HBO_2 \xrightarrow[>370K]{\Delta} Z$
7. Complete the following chemical reactions -
- i.  $Z + 3LiAlH_4 \longrightarrow X + 3LiF + 3AlF_3$
  - ii.  $X + 6H_2O \longrightarrow Y + 6H_2$
  - iii.  $3X + 3O_2 \xrightarrow{\Delta} B_2O_3 + 3H_2O$
8. When aqueous solution of borax is acidified with hydrochlorine acid, a white crystalline solid is formed, which is soapy to touch. Is this solid acidic or basic in nature? Explain, write equation involved.
9. A compound X, of boron reacts with NH<sub>3</sub> on heating to give another compound Y, which is called inorganic benzene. X can be prepared by treating BF<sub>3</sub> with LiAlH<sub>4</sub>. Identify X and Y write equations for reactions involved.

#### F. Very Long answer type questions: (5 marks each)

1. Explain -
  - i. Gallium has higher ionisation enthalpy than aluminium.
  - ii. Boron does not exist as B<sup>3+</sup>.
  - iii. Tl(NO<sub>3</sub>)<sub>3</sub> acts as an oxidising agent.
2. Give reasons -
  - i. Silicon does not form a graphite like structure whereas carbon does.
  - ii. PbX<sub>2</sub> is more stable than PbX<sub>4</sub>.
  - iii. Sn<sup>2+</sup> acts as a reducing agent.
3. Account for the following -
  - i. PbO<sub>2</sub> is stronger oxidising agent than SnO<sub>2</sub>
  - ii. The +1 oxidation state of thallium is more stable than its +3 state.
  - iii. AlCl<sub>3</sub> exists as dimer.
4. Explain why?
  - i. Acidity of BF<sub>3</sub> is less than that of BCl<sub>3</sub>.
  - ii. SiCl<sub>4</sub> can be hydrolysed but not CCl<sub>4</sub>.

- iii. Atomic radius of Ga is less than that of Al.
  - iv.  $\text{AlCl}_3$  behave as a Lewis acid.
  - v.  $\text{CO}_2$  is a gas, but  $\text{SiO}_2$  is a solid.
5. Give reasons -
- i. Conc.  $\text{HNO}_3$  can be transported in Aluminium containers.
  - ii. Graphite is used as lubricant.
  - iii. Alloys of Aluminium are used to make aircrafts.
  - iv. pH value fo human blood is between 7.26 to 7.42.
  - v. A mixture of dil. NaOH and Al pieces are used to open drains.

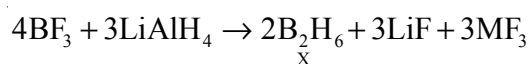
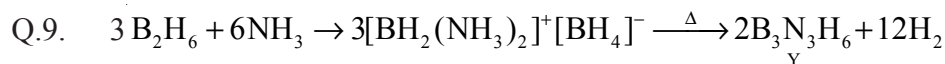
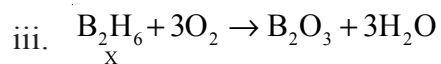
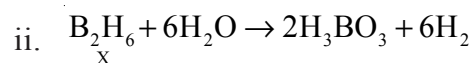
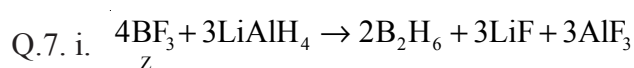
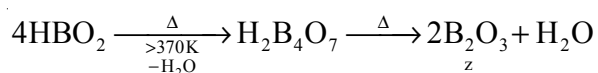
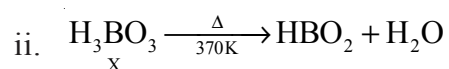
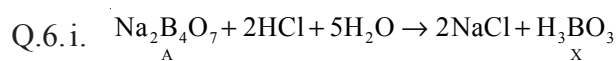
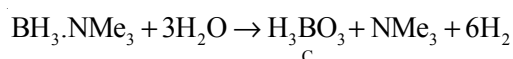


## Solution

### MCQ:

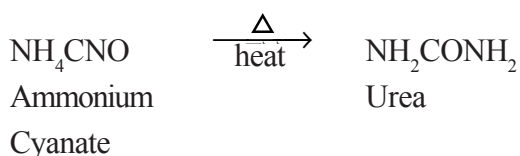
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       09. c    10. a    11. c    12. b    13. c    14. b    15. a    16. a  
       17. a    18. d    19. d    20. a    21. a    22. c    23. d    24. b  
       25. b
- B.    01. b    02. a    03. a    04. a    05. b    6. a    07. d    08. d  
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### E. Long answer type questions:

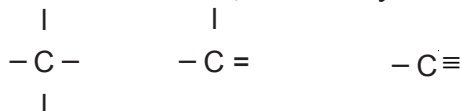


**Chapter at a glance :**

Science of Organic Chemistry is about two hundred years old. Around the year 1780, chemists began to distinguish between organic compounds obtained from plants and animals and inorganic compounds prepared from mineral sources. Berzilius, a Swedish chemist proposed that a "Vital Force" was responsible for the formation of organic compounds. However, this notion was rejected in 1828 when F. Wohler synthesised an organic compound, urea from an inorganic compound, ammonium cyanate.

**1. Tetravalence of Carbon:**

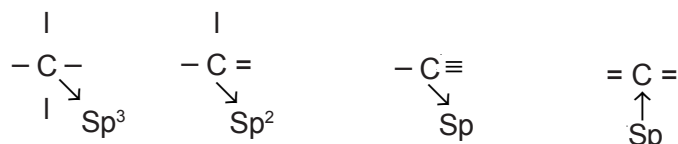
The ground state electronic configuration of carbon atom is  $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^0$ . But in excited state one electron of 2s orbital jump from 2s orbital to 2p<sub>z</sub> orbital and hence in excited state the electronic configuration of carbon is  $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$ . Therefore, due to presence of 4 unpaired electrons in outermost shell, the valency of carbon is 4.

**2. Sigma bond and pi-bond in organic compounds:**

Sigma ( $\sigma$ ) bonds are formed by the end to end overlapping of atomic orbitals whereas pi ( $\pi$ ) bonds are formed by the sideways overlapping of atomic orbitals. All single bonds are sigma bonds and in case of double bond, one sigma and one pi bond whereas in case of triple bond, one sigma and two pi bonds.

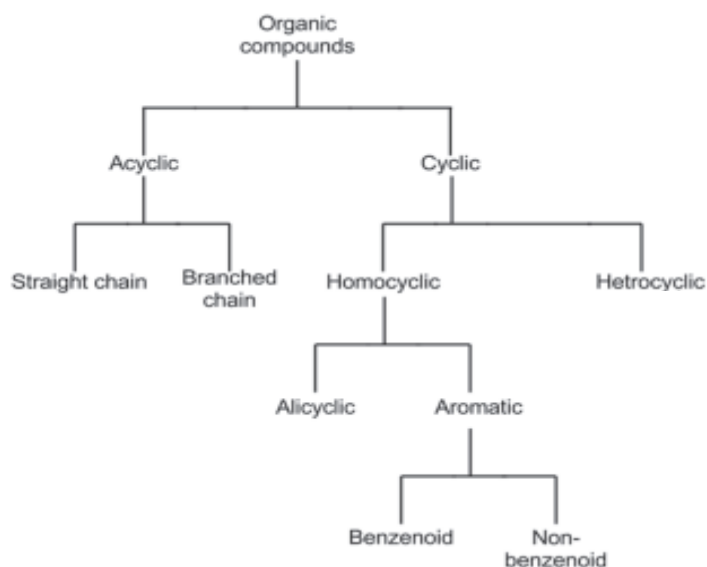
**3. Hybridisation of carbon in organic compounds:**

Carbon atom undergoes in three types of hybridisation. There are  $sp^3$ ,  $sp^2$  and  $sp$ .



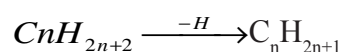
#### 4. Classification of organic compounds:

Organic compounds are broadly classified as follows:

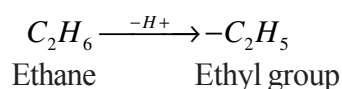
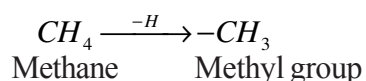


#### 5. Alkyl Group:

The groups which are formed by the elimination of one hydrogen atom from alkanes is known as alkyl group.



Example:



#### 6. Functional Group:

The functional group may be defined as an atom or group of atoms joined in a specific manner which is responsible for the characteristic chemical properties of the organic compounds. The examples are hydroxyl group (-OH), aldehyde group (-CHO) and carboxylic acid group (-COOH) etc.

#### 7. Homologous series:

A group or a series of organic compounds each containing a characteristic functional group forms a homologous series and the members of the series are called homologues. The members of a homologous series can be represented by general molecular formula and the successive members differ from each other in molecular formula by a  $-CH_2$  unit.

## 8. IUPAC Nomenclature of organic compound:

Organic chemistry deals with millions of compounds. In order to clearly identify them, a systematic method of naming has been developed and is known as the IUPAC System of nomenclature. The rules for naming of organic compounds are given below-

- First of all, the longest carbon chain (parent chain) in the molecule is identified.
- The carbon atoms of the parent chain are numbered to identify the parent alkane and to locate the position of the carbon atoms at which branching takes place due to the substitution of alkyl group in place of hydrogen atoms. The numbering is done in such a way that the branched carbon atoms get the lowest possible numbers.
- The names of alkyl groups attached as a branch are then prefixed to the name of the parent alkane and position of the substituents is indicated by the appropriate numbers. If different alkyl groups are present, they are listed in alphabetical order.
- If two or more identical substituent groups are present then the numbers are separated by commas. The names of identical substituents are not repeated, instead prefixes such as di (for 2), tri (for 3), tetra (for 4), penta (for 5), hexa (for 6) etc. are used. While writing the name of the substituents in alphabetical order these prefixes, however, are not considered.
- If the two substituents are found in equivalent positions, lower number is given to the one coming first in the alphabetical listing.
- The branched alkyl groups can be named by following the above mentioned procedures. However, the carbon atom of the branch that attaches to the root alkane is numbered as 1.
- A saturated monocyclic compound is named by prefixing 'cyclo' to the corresponding straight chain alkane. If side chains are present, then the rules given above are applied.
- The longest chain of carbon atoms containing the functional group is numbered in such a way that the functional group is attached at the carbon atom possessing lowest possible number in the chain.
- In case of polyfunctional compounds, the choice of principal functional group is made on the basis of order of preference. The order of decreasing priority for some functional group is -


-COOH > -SO<sub>3</sub>H > -COOR > -COCl > -CONH<sub>2</sub> > -CN > -CHO > C=O >

-OH > -NH<sub>2</sub> > C=C > C≡C

- If more than one functional group of the same type are present, their number is indicated by adding di, tri, etc. before the class suffix.

Example of some functional groups with their prefixes and suffixes are given as follows.

### Some Functional Groups and Classes of Organic Compounds

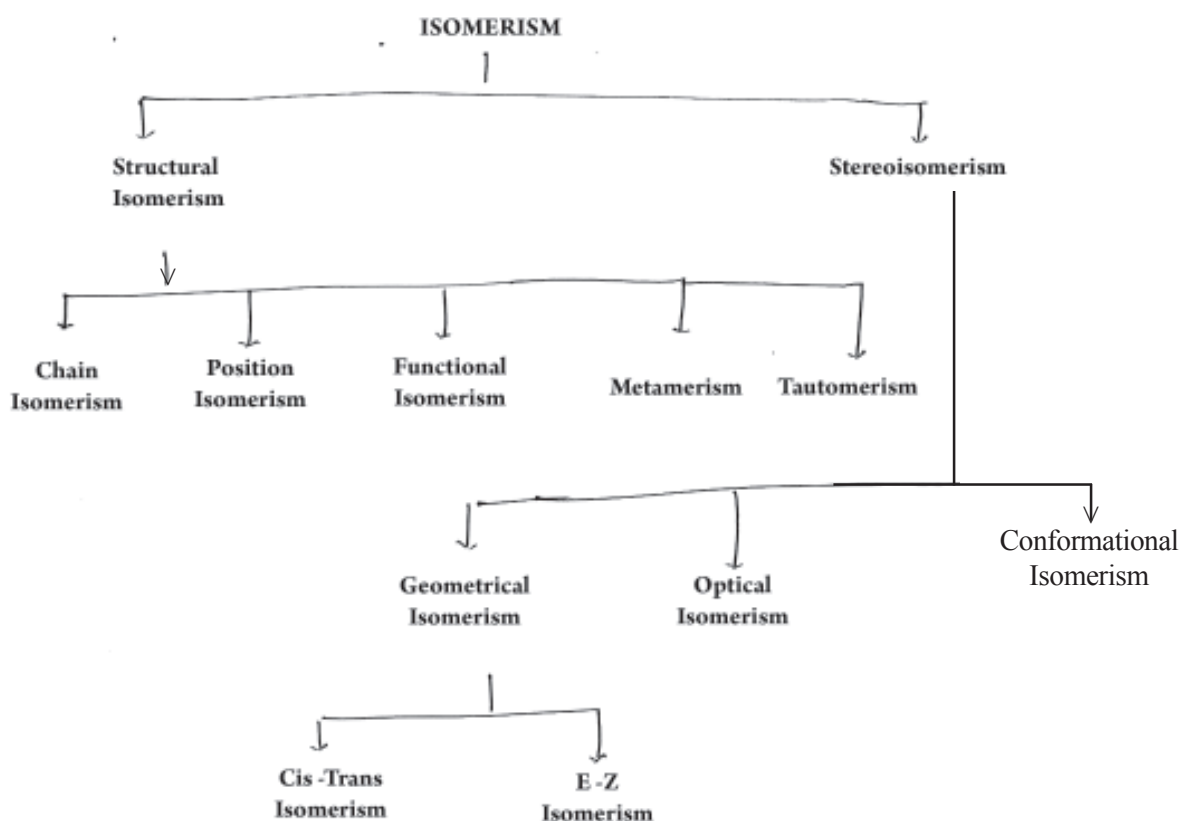
Class of compounds	Functional group structure	IUPAC group prefix	IUPAC group suffix	Example
Alkanes	-	-	-ane	Butane, $\text{CH}_3(\text{CH}_2)_2\text{CH}_3$
Alkenes	$>\text{C}=\text{C}<$	-	-ene	But-1-ene, $\text{CH}_2=\text{CHCH}_2\text{CH}_3$
Alkynes	$-\text{C}\equiv\text{C}-$	-	-yne	But-1-yne, $\text{CH}\equiv\text{CCH}_2\text{CH}_3$
Arenes	-	-	-	Benzene, 
Halides	-X (X=F,Cl,Br,I)	halo-	-	1-Bromobutane, $\text{CH}_3(\text{CH}_2)_2\text{CH}_2\text{Br}$
Alcohols	-OH	hydroxy-	-ol	Butan-2-ol, $\text{CH}_3\text{CH}_2\text{CHOHCH}_3$
Aldehydes	-CHO	formyl, or oxo	-al	Butanal, $\text{CH}_3(\text{CH}_2)_2\text{CHO}$
Ketones	$>\text{C}=\text{O}$	oxo-	-one	Butan-2-one, $\text{CH}_3\text{CH}_2\text{COCH}_3$
Nitriles	$-\text{C}\equiv\text{N}$	cyano	nitrile	Pentanenitrile, $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CN}$
Ethers	-R-O-R-	alkoxy-	-	Ethoxyethane, $\text{CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3$
Carboxylic acids	-COOH	carboxy	-oic acid	Butanoic acid, $\text{CH}_3(\text{CH}_2)_2\text{CO}_2\text{H}$
Carboxylate ions	$-\text{COO}^-$	-	-oate	Sodium butanoate, $\text{CH}_3(\text{CH}_2)_2\text{CO}_2^- \text{Na}^+$
Esters	-COOR	alkoxycarbonyl	-oate	Methyl propanoate, $\text{CH}_3\text{CH}_2\text{COOCH}_3$
Acyl halides	-COX (X=F,Cl,Br,I)	halocarbonyl	-oyl halide	Butanoyl chloride, $\text{CH}_3(\text{CH}_2)_2\text{COCl}$
Amines	$-\text{NH}_2$ , $>\text{NH}$ , $>\text{N}-$	amino-	-amine	Butan-2-amine, $\text{CH}_3\text{CHNH}_2\text{CH}_2\text{CH}_3$
Amides	$-\text{CONH}_2$ , $-\text{CONHR}$ , $-\text{CONR}_2$	-carbamoyl	-amide	Butanamide, $\text{CH}_3(\text{CH}_2)_2\text{CONH}_2$
Nitro compounds	$-\text{NO}_2$	nitro	-	1-Nitrobutane, $\text{CH}_3(\text{CH}_2)_3\text{NO}_2$
Sulphonic acids	$-\text{SO}_3\text{H}$	sulpho	sulphonic acid	Methylsulphonic acid $\text{CH}_3\text{SO}_3\text{H}$

## 9. Nomenclature of substituted benzene compounds:

- For IUPAC nomenclature of substituted benzene compounds, the substituent is placed as prefix to the word benzene.
- If benzene ring is disubstituted, the position of substituents is defined by numbering the carbon atoms of the ring such that the substituents are located at the lowest possible number.
- In the trivial system of nomenclature the terms ortho(o), meta (m) and para (p) are used as prefixes to indicate the relative positions 1, 2-; 1, 3- and 1, 4- respectively.
- Substituent of the base compound is assigned number 1 and then the direction of numbering is chosen such that the next substituents get the lowest number. The substituents appear in the name in alphabetical order.
- When a benzene ring is attached to an alkane with a functional group, it is considered a substituent instead of a parent.

**10. Isomerism:** The compounds having same molecular formula but exist in different structural formula and different properties, then the compounds are called Isomers and the phenomenon is called Isomerism.

Isomerism can be classified into different classes and these are as follows:



### 11. Fundamental concepts in organic reaction mechanism:

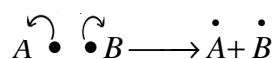
In an organic reaction, the organic molecule reacts with an appropriate attacking reagent and leads to the formation of one or more intermediates and finally products.

A sequential account of each step, describing details of electron movement, energetics during bond cleavage and bond formation, and the rates of transformation of reactants into products is referred to as reaction mechanism.

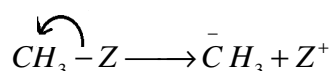
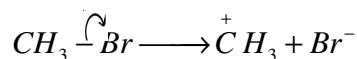
### 12. Fission of co-valent bond:

The process of cleavage of co-valent bond is known as bond fission. A covalent bond can get cleaved either by homolytic fission or heterolytic fission.

**a) Homolytic Fission:** If the shared pair electrons of covalent bond are equally distributed between the bonded atom the fission is called homolytic fission. During the homolytic bond fission, the neutral species obtained is called free radical.



**b) Heterolytic Fission:** If the shared pair electrons of co-valent bond are unequally distributed between the bonded atoms, the fission is called heterolytic fission. During the heterolytic bond fission the positively charged carbon species obtained or called carbocation or carbonium ion and negatively charged carbon species obtained is called carbanion.

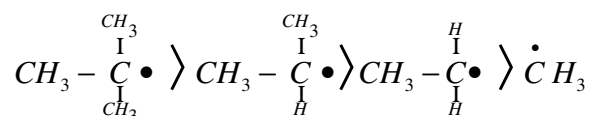


### 13. Reaction Intermediates:

The species obtained during bond fission are called reaction intermediates. During a chemical reaction we get the following reactive intermediates:

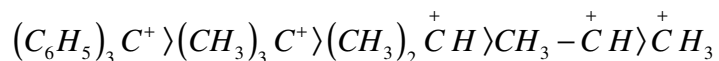
**a) Free radical:** The neutral species obtained by the homolytic bond fission is called free radical. Free radicals are paramagnetic in nature due to presence of unpaired electron. The carbon atom in alkyl free radical is  $sp^2$ . Therefore, it has planar shape.

Stability of free radicals:



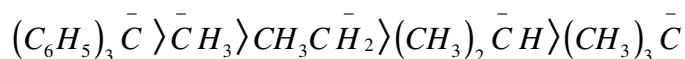
**b) Carbocation:** The positively charged carbon species obtained during heterolytic bond fission is called carbocation. Carbocation has only six electrons in its outermost shell. It is diamagnetic in nature and the carbon atom in alkyl carbonium ion is  $sp^2$ . So, it has a planar shape.

Stability of carbocation:



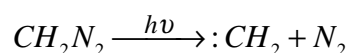
**c) Carbanion:** The negatively charged carbon species obtained during heterolytic bond fission is called carbanion. Carbanion has eight electrons in its outermost shell. It is also diamagnetic in nature and the carbon atom in alkyl carbanion is  $sp^3$ . It has tetrahedral structure. But due to presence of one unpaired electron in one  $sp^3$  hybridised orbital, it has pyramidal shape.

Stability of carbanion:



**d) Carbenes:** The bi-valent neutral carbon species in which carbon atom has six electrons in the outer most shell is called carbene.

- 14. Reagent:** More reactive reactants are called reagents. The attacking reagents can be classified into the following classes.



**a) Electrophile:** A reagent that takes away an electron pair is called electrophile. i.e. electron seeking.

Example of electrophile:  $H^+$ ,  $H_3O^+$ ,  $Cl^+$ ,  $NO^+$  etc (Positively charged)

$AlCl_3$ ,  $BF_3$ ,  $SO_3$  etc (neutral)

**b) Nucleophile:** A reagent that brings an electron pair is called a nucleophile i.e. nucleus seeking.

Example of nucleophile:  $Br^-$ ,  $OH^-$ ,  $CN^-$  etc (negatively charged)

$NH_3$ ,  $PH_3$ ,  $H_2O$  etc (neutral)

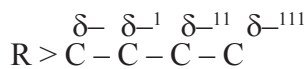
- 15. Electron displacement effects in covalent bonds:**

**a) Inductive effect:** When a covalent bond is formed between atoms of different electronegativity, the electron density is more towards the more electronegative atom of the bond. Such a shift of electron density result in a polar covalent bond. Bond polarity leads to various electronic effects in organic compounds.

The polarisation of  $\sigma$  bond caused by the polarisation of adjacent  $\sigma$ -bond is referred to as the inductive effect. It can be classified into two classes -

**i) (+) I-effect:** If electron donating group is attached to the end of the carbon chain, then the electron density of carbon chain increases, this effect is called (+) I-effect.

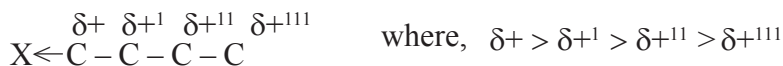




Electron donating group: where,  $\delta- > \delta-1 > \delta-11 > \delta-111$

R is alkyl group as for example - Methyl(-CH<sub>3</sub>), Ethyl(-C<sub>2</sub>H<sub>5</sub>) etc.

**ii) (-) I-effect:** If electron withdrawing group is attached to the end of the carbon chain, then the electron density of carbon chain decreases, this effect is called (-) I-effect.



### Electron withdrawing group:

Halogen (F, Cl, Br, I), Nitro (-NO<sub>2</sub>), Cyano(-CN), Carboxy (-COOH), ester(-COOR), aryloxy (-OAr) etc.

### b) Resonance:

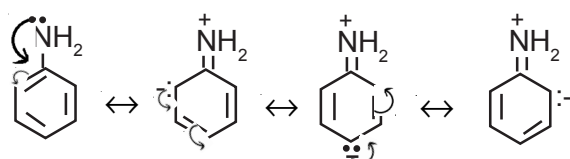
If a molecule can exist more than one structural form but none of which is capable to explain all the known properties of the compound, then the phenomenon is called resonance and the different structural forms are called resonating structure as canonical structure.

The difference in energy between the actual structure and the lowest energy resonance structure is called resonance energy. The more the number of resonating structure, the more is the resonance energy and its more stable.

### Resonance effect:

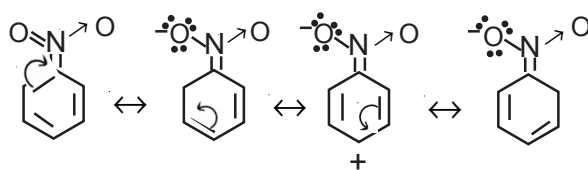
The resonance effect is defined as the polarity produced in the molecule by the interaction of two  $\pi$ -bonds or between  $\pi$ -bond and lone pair of electrons present on an adjacent atom. It is also known as mesomeric effect, and it can be classified into two classes-

**a) (+) R- or (+) M-effect:** In this type, the transfer of electrons is away from an atom or substituent group attached to the conjugated system. This effect in **aniline is shown as:**



+R effect is shown by halogen, -OH, -OR, -OCOR, -NH<sub>2</sub>, -NHR, -NR<sub>2</sub>, -NHCOR groups.

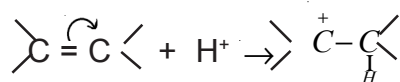
**b) (-) R or (-) M effect:** This effect is observed when the transfer of electron is towards the atom or substituent group attached to the conjugated system. This effect in nitro benzene is shown as:



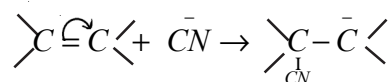
-R effect is shown by halogen, -OCOR, -CHO, >C=O, -CN, -NO<sub>2</sub> groups.

**c) Electromeric effect:** The complete transfer of a shared pair of  $\pi$  electrons to one of the atom joined by a multiple bond on the demand of an attacking reagent is called electromeric effect. It is a temporary effect. There are two types of electromeric effect and these are as follows -

**i) (+) E-effect:** In this effect the  $\pi$  electrons of the multiple bond are transferred to that atom to which the reagent gets attached.



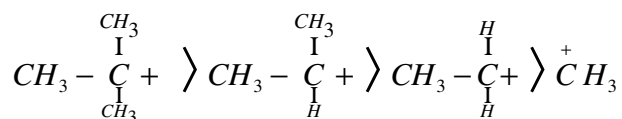
**ii) (-) E-effect:** In this effect the  $\pi$  electrons of the multiple bond are transferred to that atom to which the attacking reagent does not get attached.



**d) Hyperconjugation:**

Hyperconjugation is a general stabilising interaction. It involves delocalisation of  $\sigma$  electron of C-H bond of an alkyl group directly attached to an atom of unsaturated system or to an atom with an unshared p-orbital. The  $\sigma$  electrons of C-H bond of the alkyl group enter into partial conjugation with the attached unsaturated system or with the unshared p-orbital. Hyperconjugation is a permanent effect.

In general, greater the number of alkyl group attached to a positively charged carbon atom, the greater is the hyperconjugation interaction and stabilisation of the cation. Thus, we have the following relative stability of carbocations:



**16. Types of organic reactions:**

Organic reactions can be classified into the following classes -

- a) Substitution reaction.
- b) Addition reaction.
- c) Elimination reaction.
- d) Re arrangement reaction.

**17. Methods of purification of organic compounds:**

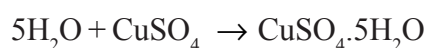
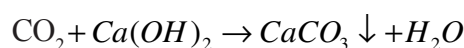
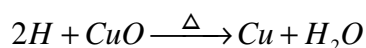
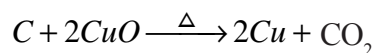
Once an organic compound is extracted from a natural source as synthesised in the laboratory, it is essential to purify it. Various methods used for the purification of organic compounds are based on the nature of the compound and the impurity present in it.

The common techniques used for purification are as follows -

- a) Sublimation
- b) Crystallisation
- c) Distillation
- d) Differential extraction
- e) Chromatography

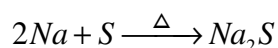
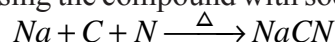
## 18. Qualitative analysis of organic compounds:

**I. Detection of carbon and hydrogen:** Carbon and hydrogen are detected by heating the compound with copper (II) oxide. Carbon present in the compound is oxidised to carbon dioxide (tested with lime-water, which develops turbidity) and hydrogen to water (tested with anhydrous copper sulphate, which turns blue).



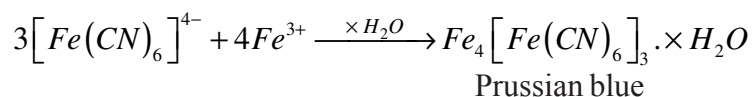
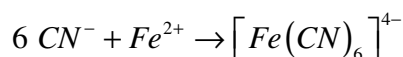
## II. Detection of Nitrogen, Sulphur, Halogens and Phosphorus:

Nitrogen, Sulphur, halogens and phosphorus present in an organic compound are detected by Lassaignes test. The elements present in the compound are converted from covalent form into the ionic form by fusing the compound with sodium metal.

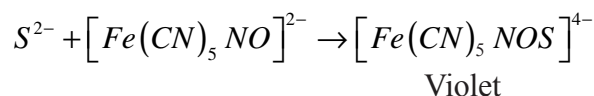


C, N, S and X come from organic compound.

a) The sodium fusion extract is boiled with iron (II) sulphate and then acidified with concentrated sulphuric acid. The formation of prussian blue colour confirms the presence of nitrogen.



b) On treating sodium fusion extract with sodium nitroprusside, appearance of a violet colour further indicates the presence of sulphur.



c) The sodium fusion extract is acidified with nitric acid and then treated with silver nitrate. A white precipitate, soluble in ammonium hydroxide shows the presence of chloride, a yellowish precipitate, sparingly soluble in ammonium hydroxide shows the presence of bromine and a yellow precipitate, insoluble in ammonium hydroxide shows the presence of Iodine.



### 19. Quantitative analysis of organic compounds:

The percentage composition of elements present in an organic compound is determined by the methods based on the following principles:

#### a) Estimation of carbon and hydrogen (Liebig's method):

$$\text{Percentage of Carbon} = \frac{12 \times m_2 \times 100}{44 \times m}$$

$$\text{Percentage of Hydrogen} = \frac{2 \times m_1 \times 100}{18 \times m}$$

Where, m = mass of organic compound.

$m_1$  = mass of water

$m_2$  = mass of carbon dioxide

#### b) Estimation of Nitrogen :

##### i) Dumas Method :

$$\text{percentage of nitrogen} = \frac{28 \times V \times 100}{22400 \times m}$$

where, m = mass of organic compound

v = volume of nitrogen at STP

$$V = \frac{P_1 V_1 \times 273}{760 \times T_1} \text{ mL.}$$

$V_1$  = Volume of nitrogen gas collected

$P_1$  = Pressure at which nitrogen gas is collected

$T_1$  = room temperature (in K)

**b) Kjeldahl's method:**

$$\text{Percentage of Nitrogen} = \frac{14 \times M \times 2 \left( V - \frac{V_1}{2} \right) \times 100}{1000 \times m}$$

where,

M = concentration of  $\text{H}_2\text{SO}_4$  in molarity

m = mass of organic compound

V = volume of M molar  $\text{H}_2\text{SO}_4$ .

$V_1$  = volume of M molar NaOH.

**c) Estimation of Halogens: (Carius Method)**

$$\text{Percentage of Halogen} = \frac{\text{atomic mass of X} \times m_1 \times 100}{\text{molecular mas of AgX} \times m}$$

where, m = mass of organic compound.

$m_1$  = mass of AgX formed.

**d) Estimation of Sulphur: (carius method)**

$$\text{Percentage of Sulphur} = \frac{32 \times m_1 \times 100}{233 \times m}$$

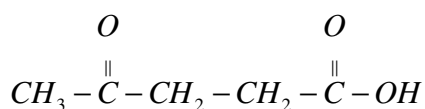
where, m = mass of organic compound.

$m_1$  = mass of  $\text{BaSO}_4$  formed.

**A. Choose the correct answer (MCQ):**

**(Each question of mark 1)**

1. The IUPAC name of the compound



- a) 1-hydroxypentane-1, 4-dione  
b) 1,4-dioxypentanol.  
c) 1-carboxybutan -3-One  
d) 4 - Oxopentanoic acid
2. Which of the following compound contain all the carbon atoms in the same hybridisation state?
- a)  $\text{CH}_3 - \text{C} \equiv \text{C} - \text{CH}_3$   
b)  $\text{CH}_2 = \text{CH} - \text{CH} = \text{CH}_2$   
c)  $\text{CH}_3 - \text{CH} = \text{CH}_2$   
d)  $\text{CH}_3 - \text{CH}_2 - \text{C} \equiv \text{CH}$
3. In which of the following, functional group isomerism is not possible?  
a) Alcohols      b) Aldehydes      c) Cyanides      d) Alkyl halides
4. Which of the following carbocation is most stable?  
a) Methyl      b) Ethyl      c) Isopropyl      d) Tert-butyl
5. The IUPAC name of the compound  
 $\text{CH}_3 - \text{CH} = \text{CH} - \text{CH}_2 - \text{C} \equiv \text{CH}$  is  
a) Hex -2-en-5-yne      b) Hex-4-ene-1-yne  
c) Hex-1-en-4-yne      d) Hex-2-en-4-yne
6. Which of the following is an electrophile?  
a)  $\text{NH}_3$       b) ROH      c)  $\text{AlCl}_3$       d)  $\text{CN}^-$
7. Homolytic fission of a covalent bond leads to the formation of -  
a) Nucleophile      b) Electrophile  
c) Carbocation      d) Free Radical
8. An electrophilic reagent is -  
a) electron rich species      b) a Lewis base  
c) Electron deficient species      d) negatively charged species

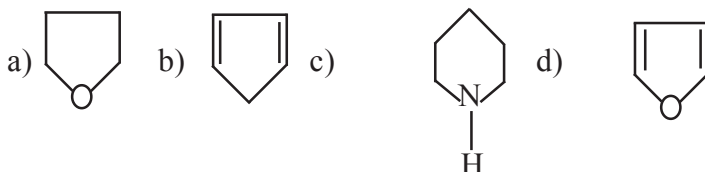






27. What is the type of hybridisation of carbon atom in the compound  $\text{HCONH}_2$  ?  
 a)  $\text{sp}^3$       b)  $\text{sp}^2$       c)  $\text{sp}$       d)  $\text{sp}^3\text{d}$
28. Kjeldahl's method cannot be used for the estimation of nitrogen in -  
 a) Pyridine                      b) Nitrocompounds  
 c) Azo compounds              d) All of these
29. Two members of a homologous series have -  
 a) Different general formula  
 b) Different molecular weights  
 c) Different chemical properties  
 d) Different methods of preparation
30. During Lassaigne's test, N and S present in an organic compound changes into -  
 a)  $\text{NaSCN}$                       b)  $\text{Na}_2\text{SO}_4$  and  $\text{NaCN}$   
 c)  $\text{NaCN}$  and  $\text{Na}_2\text{S}$               d)  $\text{Na}_2\text{S}$  and  $\text{NaCNO}$

31. Which of the following is heterocyclic aromatic compound ?



32. Which of the following has +R (Resonance) effect?

a)  $-\text{CN}$ ,      b)  $-\text{NH}_2$ ,      c)  $-\text{CHO}$ ,      d)  $-\text{NO}_2$

33. The structure of isobutyl group is -

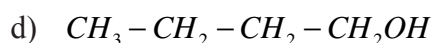
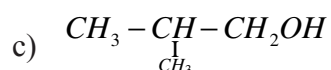
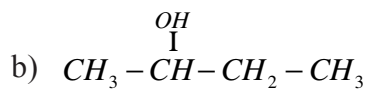
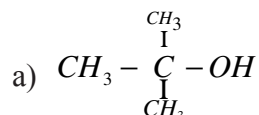
a)  $\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 -$                       b)  $\text{CH}_3 - \underset{\text{I}}{\text{C}} - \text{CH}_2 - \text{CH}_3$

c)  $\text{CH}_3 - \underset{\text{CH}_3}{\overset{\text{CH}_3}{\text{C}}} -$                       d)  $\text{CH}_3 \text{ > } \underset{\text{CH}_3}{\text{C}} - \text{CH}_2 -$

34. The correct IUPAC name of  is

a) iso-propyl benzene                      b) cumene  
 c) 2-phenyl propane                      d) Phenyl iso propane

35. The compound which gives the most stable carbocation on dehydration is -



**B. Assertion Reason type questions:**

**(Each question mark-1)**

- Both Assertion (A) and Reason (R) are correct and reason is the correct explanation of the Assertion.
  - Both Assertion (A) and Reason (R) are correct and Reason is not the correct explanation of Assertion.
  - Assertion (A) is correct but Reason (R) is incorrect.
  - Assertion (A) is incorrect but Reason (R) is correct.
- Assertion (A): Alkanes having more than three carbon atoms exhibit chair isomerism.  
Reason (R): Alkanes have general formula  $\text{C}_n\text{H}_{2n+2}$ .
  - Assertion (A): A mixture of o-nitrophenol and p-nitrophenol can be separated by steam distillation.  
Reason (R): O-nitrophenol is steam volatile while p-nitrophenol is not steam volatile.
  - Assertion (A): All the carbon atoms in  $\text{CH}_2=\text{C}=\text{CH}_2$  are  $\text{sp}^2$  hybridised.  
Reason(R): In this molecule, all carbon atoms are linked to each other by double bonds.
  - Assertion (A): But-1-ene and But-2-ene are position isomers.  
Reason(R): Position isomers differ in the position of functional group or a substituent.
  - Assertion (A): Allyl free radical is more stable than methyl free radical.  
Reason(R): The allyl free radical is stabilized by the resonance.
  - Assertion (A):  $\text{NH}_3$  can behave as a nucleophile.  
Reason (R):  $\text{NH}_3$  has one lone pair of electrons.
  - Assertion (A): Electromeric effect is a temporary effect.  
Reason (R): It operates only in the presence of an outside reagent.
  - Assertion (A): Electrophiles behave as Lewis acids.  
Reason (R): They have empty orbital due to which they can accept electrons.
  - Assertion (A): Homolytic fission of covalent bonds leads to the formation of free radicals.  
Reason (R): The order of stability of alkyl free radicals is :  $\text{CH}_3 < 1^\circ < 2^\circ < 3^\circ$

10. Assertion (A): The positively charged atom in carbocation is  $sp^2$  hybridised.  
Reason (R): It's structure is linear.

**C. Very short answer type question:**

**(Each question 1 mark)**

- Write the IUPAC name of the compound  $CH_3 - O - CH(CH_3)_2$ .  
Ans: \_\_\_\_\_
- Write the structural formula of pent-4-en-2-ol.  
Ans: \_\_\_\_\_
- How many  $\sigma$  and  $\pi$  bonds are present in  $CH_2 = C = CHCH_3$ .  
Ans: \_\_\_\_\_
- What is the type of hybridisation of '\*' marked carbon atom in  $CH_3C^*HO$ ?  
Ans: \_\_\_\_\_
- What is the general formula of alcohols?  
Ans: \_\_\_\_\_
- What is inductive effect?  
Ans: \_\_\_\_\_
- Give two examples of functional groups which show '+R' effect.  
Ans: \_\_\_\_\_
- What is the type of hybridisation of positively charged carbon atom in carbocation?  
Ans: \_\_\_\_\_
- Arrange the following carbocations in order of increasing stability:  
 $CH_3^+CH_2$ ,  $(CH_3)_3C^+$ ,  $CH_3^+$ ,  $(CH_3)_2CH^+$   
Ans: \_\_\_\_\_
- Which type of bond fission produces free radicals?  
Ans: \_\_\_\_\_
- Write bond line formula of heptan-4-one?  
Ans: \_\_\_\_\_
- Which bond is more polar?  $H_3C - H$  or  $H_3C - Br$ ?  
Ans: \_\_\_\_\_
- Which method will you use for separating a mixture of o-nitrophenol and p-nitrophenol?  
Ans: \_\_\_\_\_
- Give one example of electrophile.  
Ans: \_\_\_\_\_
- Which is the best and latest technique for isolation, purification and separation of organic compound?  
Ans: \_\_\_\_\_
- Give one example of elimination reaction.  
Ans: \_\_\_\_\_

17. What is the geometrical shape of the molecule  $H - C \equiv C - C \equiv C - Cl$  ?

Ans : \_\_\_\_\_

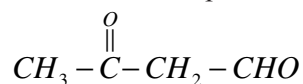
18. Name one hydrocarbon whose all carbon atoms are  $2^{\circ}$ .

Ans : \_\_\_\_\_

19. Give one example of ambident nucleophile.

Ans : \_\_\_\_\_

20. Write the IUPAC name of the compound:

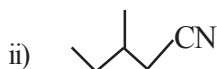
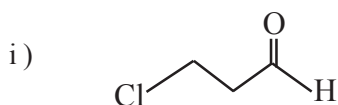


Ans : \_\_\_\_\_

**D. Short answer type questions:**

**( Each questions marks 2)**

1. What are metamers? Give example.
2. What is homolytic fission? Give example.
3. Draw the resonance structures for  $C_6H_5OH$ .
4. Explain the following correct order of acidity of the carboxylic acids:  
 $Cl_3COOH > Cl_2CHCOOH > ClCH_2COOH$
5. Will  $CCl_4$  give white precipitate of  $AgCl$  on heating it with silver nitrate? Give Reason.
6. Explain, how is the electronegativity of carbon atoms related to their state of hybridisation in an organic compound ?
7. Write structural formulae for the followings:  
(i) 2-Chlorobutanoic acid  
(ii) 2-Methyl propan -2-ol
8. Write two points of differences between inductive effect and electronics effect.
9. Why does  $SO_3$  acts as an electrophile?
10. Define homologous series. Give example.
11. Write resonating structures of  $CH_3COO^-$  ion.
12. Why is  $(CH_3)_3C^+$  more stable than  $CHC^+H_2$  ?
13. What is the principle of chromatography ?
14. How is nitrogen detected in an organic compound?
15. What are functional isomers ? Give one example.
16. What are substitution reactions? Give one example.
17. Why is nitric acid added to sodium extract before adding silver nitrate for testing halogens?
18. Which of the two:  $O_2NCH_2CH_2O^-$  or  $CH_3CH_2O^-$  is expected to be more stable and why?
19. Give the IUPAC names of the following compounds:

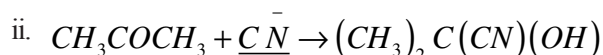
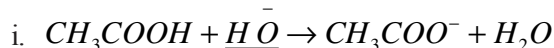


20. Write the structural formulae for structural isomers of pentane ( $C_5H_{12}$ ).

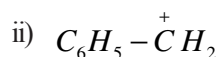
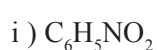
**E. Long answer type questions:**

**Marks - 3**

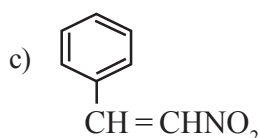
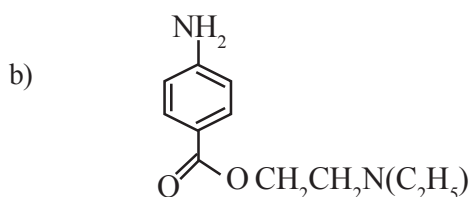
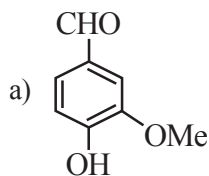
1. What are electrophiles and nucleophiles? Explain with examples.
2. Identify the reagents shown underlined in the following equations as nucleophiles or electrophiles:



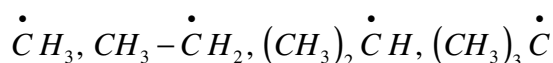
3. What do you mean by +I effect and -I effect? Give example.
4. Draw resonance structures for the followings:



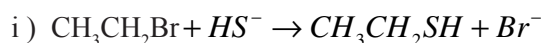
5. What is hyperconjugation? What are its important consequences?
6. What are addition and substitution reactions? Explain with examples.
7. What are different types of reaction intermediates formed by homolytic and heterolytic fission of a covalent bond? Explain with example.
8. Identify the functional groups in the following compounds:

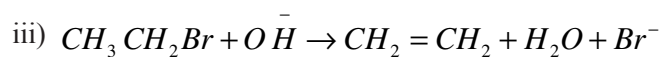
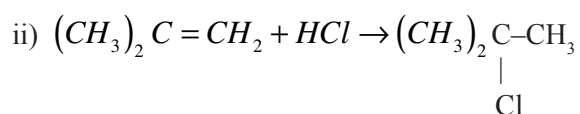


9. An organic compound contains 69% carbon and 4.8% hydrogen, the remainder being oxygen. Calculate the masses of carbon dioxide and water produced when 0.20g of this substance is subjected to complete combustion.
10. i) Explain, why an organic liquid vaporises at a temperature below its boiling point in its steam distillation?  
ii) A liquid has non-volatile impurities. Which method will you use for its purification?
11. Arrange the following free radicals in order of increasing stability and explain the reason.



12. Classify the following reactions in one of the reaction type studied in this unit.



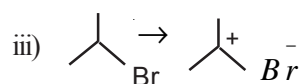
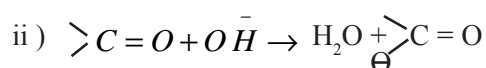
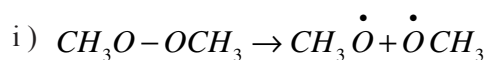


13. 0.3780g of an organic chloro compound gave 0.5740g of silver chloride in carius estimation. Calculate the percentage of chlorine present in the compound.
14. Differentiate between the principle of estimation of nitrogen in an organic compound by -  
 i) Dumas method and (ii) Kjeldahl's method
15. i) Explain why alkyl groups act as electron donors when attached to a  $\pi$  system.  
 ii) Give one example of positional isomerism.

**F. Very long answer type questions:**

**( Each question marks - 5)**

1. For the following bond clearages, use curved arrows to show the electron flow and classify each as homolysis or heterolysis. Identify reactive intermediate produced as free radical, carbocation and carbanion.



2. Write structural formulae of the following compounds:

a) 2 - Methoxypropane

b) 4-phenyl-3-hydroxybutanal

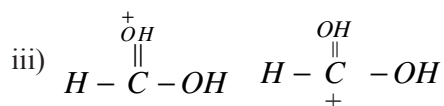
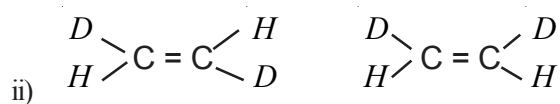
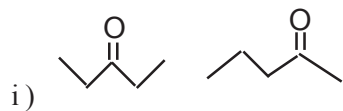
c) 3 - Ethylcyclohexene

d) Hex-2-en-1-oic acid

e) 3 - Chloropropanal.

3. a) Write the principle of estimation of halogen, sulphur and phosphorous present in an organic compound.  
 b) What are free radicals? Give examples.

4. a) What is the relationship between the members of following pairs of structures ? Are they structural or geometrical isomers or resonance contributors?



- b) Write Structural formulae for all the functional isomers of the molecular formula  $C_4H_8O$ .
5. Write formulae for the first five members of each homologous series beginning with the following compounds:
- a)  $H - CHO$       b)  $H - COOH$       c)  $CH_3COOH_3$

## Solution

- A. 1. d      2. b      3. d      4. d      5. b  
6. c      7. d      8. c      9. d      10. b  
11. b      12. a      13. b      14. b      15. d  
16. b      17. b      18. a      19. d      20. c  
21. b      22. d      23. c      24. b      25. b  
26. c      27. b      28. d      29. b      30. c  
31. d      32. b      33. d      34. c      35. a

E. 9)  $\text{CO}_2 = 0.506\text{g}$ ,  $\text{H}_2\text{O} = 0.0864\text{g}$

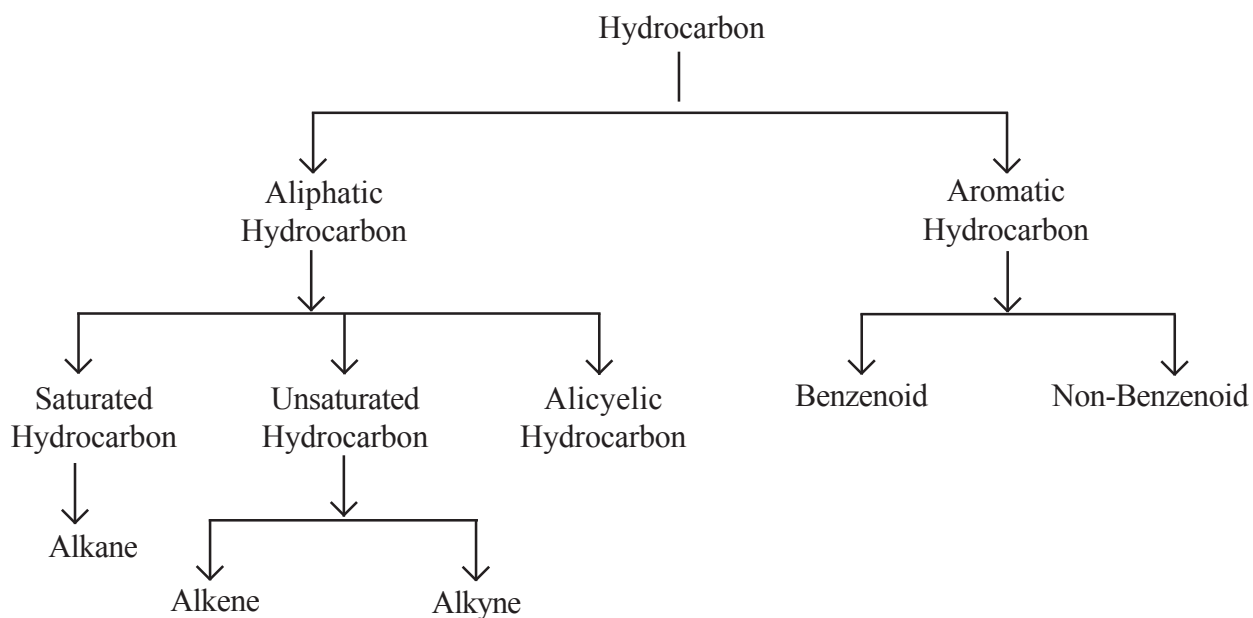
13) % chlorine = 37.57%



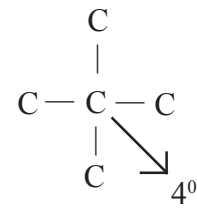
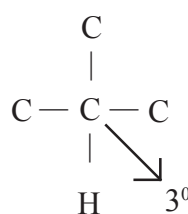
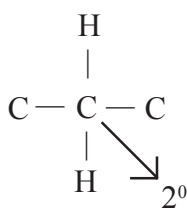
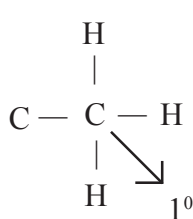
**Chapter at a glance :**

Hydrocarbon is a compound of carbon and hydrogen only. Hydrocarbons play a key role in our daily life. We are familiar with the terms LPG, CNG, Kerosene, Petrol, diesel etc. All these fuels contain mixture of hydrocarbons which are sources of energy. Hydrocarbons are used for the preparation of polymers like polythene, polystyrene, polypropene etc. Higher hydrocarbons are also used as a solvent for paints.

- Hydrocarbons can be classified into different classes. The classification of hydrocarbons are as follows.

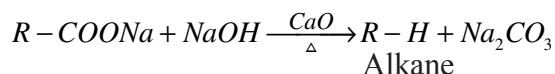


- Alkanes are the saturated hydrocarbons having carbon-carbon single bond. The general formula of alkane is  $C_nH_{2n+2}$ , where  $n = 1, 2, 3 \dots$ . Methane is the first member of alkane series.
- Alkanes can show chain isomerism. On the basis of number of carbon atoms attached to a carbon atom, the carbon atom is termed as Primary ( $1^\circ$ ), Secondary ( $2^\circ$ ), tertiary ( $3^\circ$ ) and quaternary ( $4^\circ$ ).

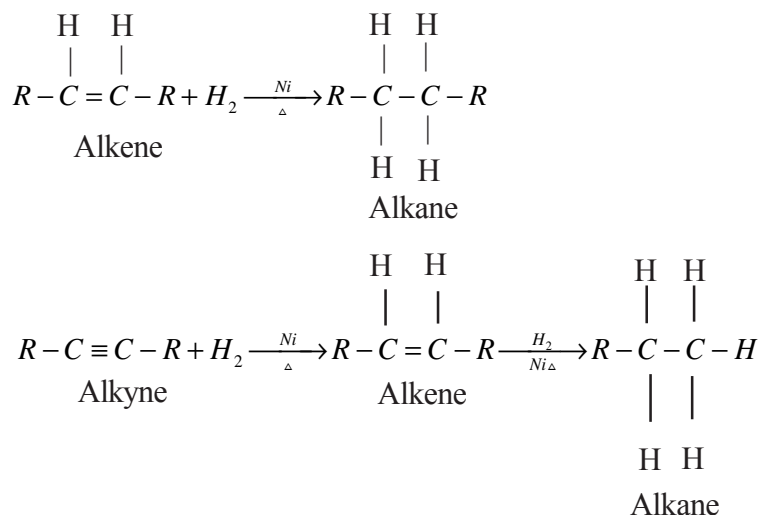


4. Petroleum and natural gas are the main sources of alkanes. However, alkanes can be prepared by following methods.

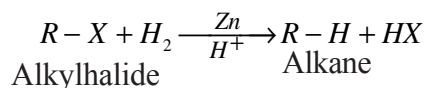
**a) From sodium salt of carboxylic acid :**



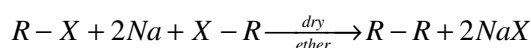
**b) From unsaturated hydrocarbon (Sabatier - Senderens Reaction) :**



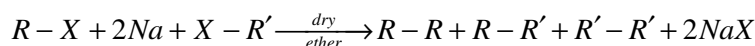
**c) From Alkyl halides (dehydrohalogenation reaction) :**



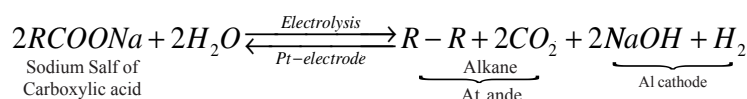
**d) By Wurtz's Reaction :**



Wurtz's reaction is not suitable for the preparation of odd number of carbon containing alkane. Because when two different types of alkylhalides take part in this reaction then a mixture of alkanes are formed.



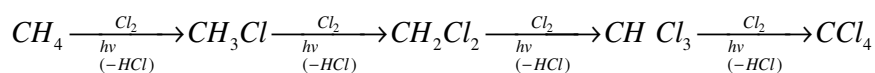
**e) By Kolbe's electrolysis methods :**



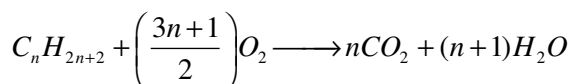
5. Alkanes are almost non-polar in nature. They possess weak vander waal forces. Boiling point of alkanes increases with increase in their molecular mass. This is due to the fact that the intermolecular forces increase with increase of the molecular size. Branched chain alkane have lower boiling point compared to same number of carbon containing straight chain alkane due its small size and weak intermolecular forces.

6. Alkanes are generally inert towards acids, bases, oxidising and reducing agents. However, they undergo the following reactions under certain conditions.

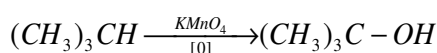
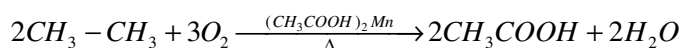
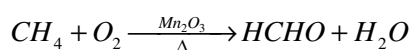
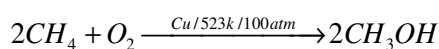
**a) Substitution Reaction :**



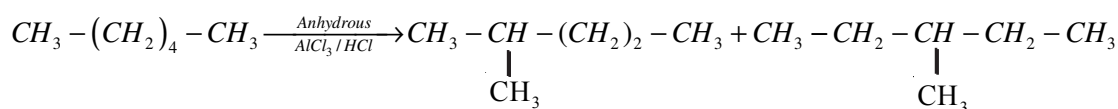
**b) Combustion :**



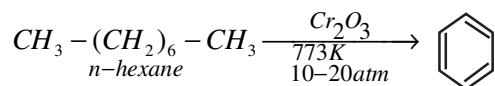
**c) Controlled oxidation :**



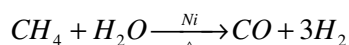
**d) Isomerisation :**



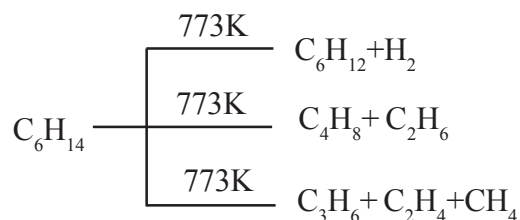
**e) Aromatisation :**



**f) Reaction with steam :**

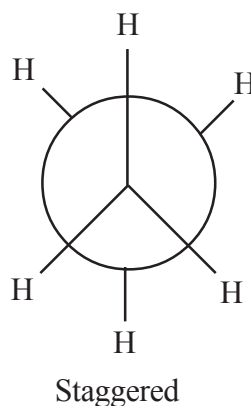
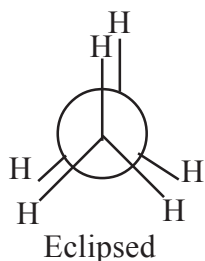


**g) Pyrolysis :**



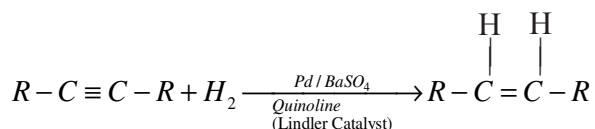
7. Conformations are different arrangements of a molecule which can be obtained by rotation of atoms around c-c single bond. The conformation in which hydrogen (atoms attached to two carbons are as closed) together as possible is called eclipsed conformation and the conformation in which hydrogens are as far apart as possible is known as staggered conformation.

conformation. The conformations can be represented by sawhorse and Newman projections.

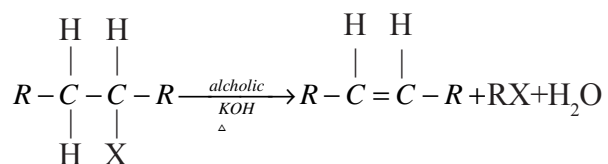


8. Alkenes are unsaturated hydrocarbons containing at least one carbon-carbon (C = C) double bond. The general formula of alkenes is  $C_nH_{2n}$ , where  $n = 2, 3, 4, \dots$ . Alkenes are also known as Olefins (Oil forming)
9. Alkenes show structural as well as geometrical isomerism. In case of geometrical isomerism, trans-isomer is more stable than cis-isomer. The dipole moment of cis-isomer is higher than that of trans-isomer. Hence cis-isomer is more polar compared to trans-isomer.
10. Alkenes can be prepared by using different methods. Some of these methods are as follows:

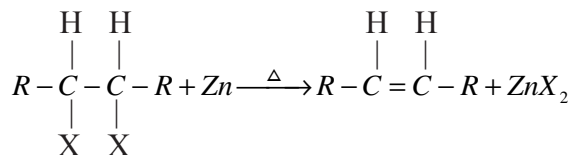
a) From Alkyne :



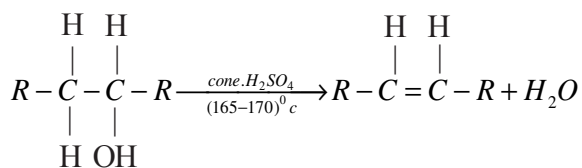
b) From Alkylhalides : (dehydrohalogenation/  $\beta$ -Elimination)



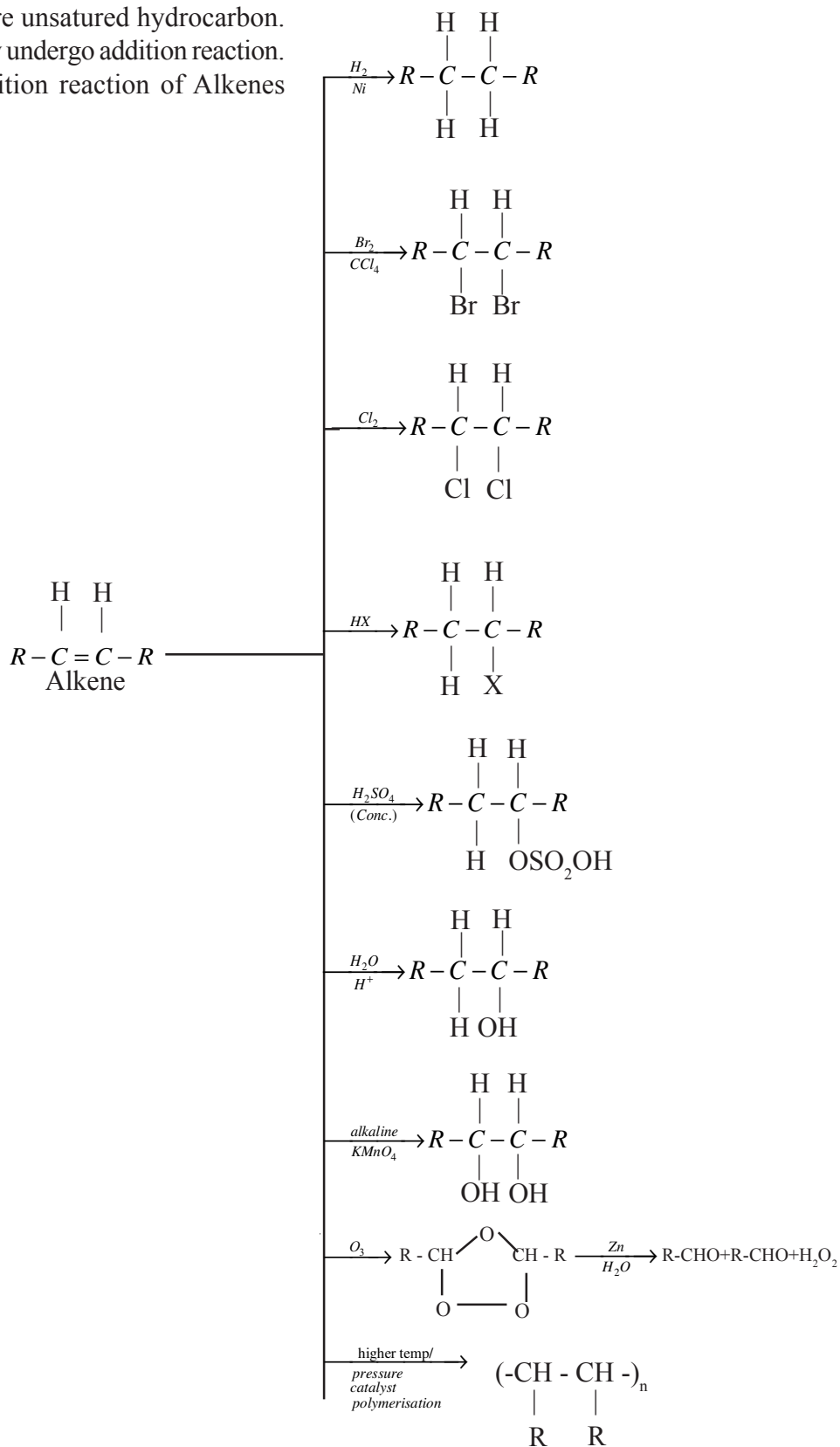
c) From Vicinal dihalides :



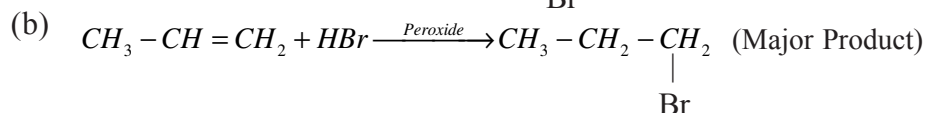
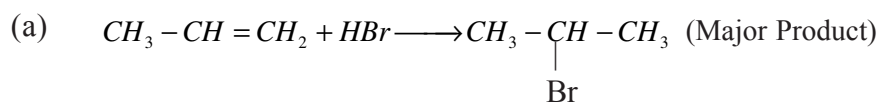
d) From Alcohol: (Acidic dehydration) :



11. Alkenes are unsaturated hydrocarbon. Hence they undergo addition reaction. Some addition reaction of Alkenes are as :



12. Unsymmetrical alkenes when take part in addition reaction with unsymmetrical reagent, then the reaction obey Markovnikov's rule but if the reaction occur in presence of peroxide, then the reaction follows anti Markovnikov's rule. This happens only for HBr but not with HCl and HI. Example for :

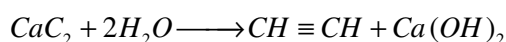


13. Alkynes are also unsaturated hydrocarbon having carbon-carbon ( $C \equiv C$ ) triple bond. The general formula of alkynes is  $C_n H_{2n-2}$ , where  $n = 2, 3, 4, \dots$

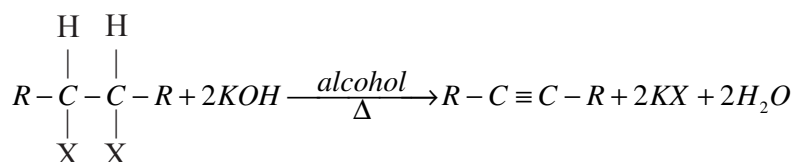
14. In Alkyne the carbon atoms are  $sp$  Hybridised. So the carbon atoms have 50%  $s$  character and more electronegative. Hence the hydrogen attached with the carbon atom of triple bond can be easily replaced by active metals and thus alkyne is acidic in nature.

15. Alkynes can be prepared by using the following methods:

a) From Calcium Carbide :

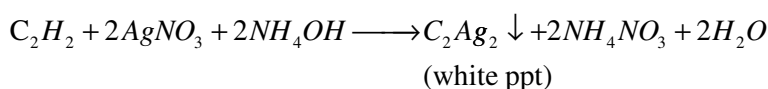
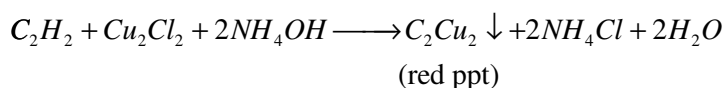
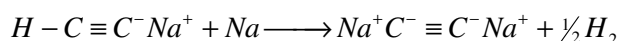
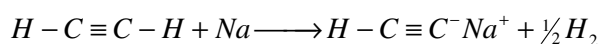


b) From Vicinal dihalides :

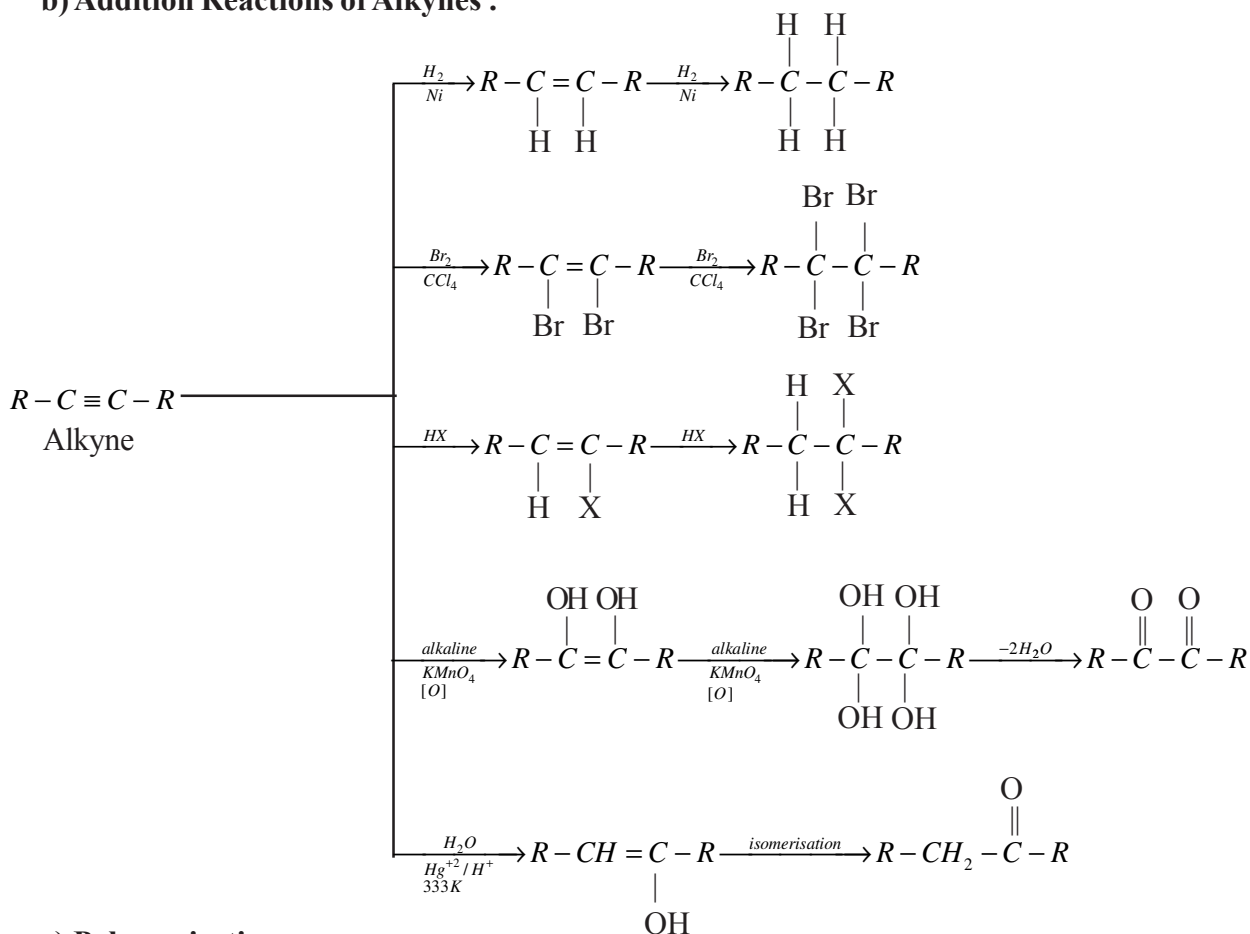


16. Alkynes show acidic nature, addition reactions and polymerisation reactions as follows:

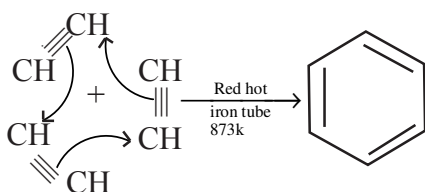
a) Acidic character of Alkyne :



## b) Addition Reactions of Alkynes :



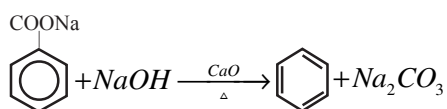
## c) Polymerisation :



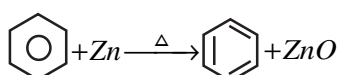
17. Aromatic hydrocarbons are known as 'arenes'. Since most of them possess pleasant odour the class of compounds was named as 'Aromatic Compounds'. Aromatic Compounds containing benzene ring are known as benzenoids and those not containing a benzene ring are known as non-benzenoids.
18. Benzene was considered as parent 'aromatic compound'. Now, the name is applied to all the ring systems whether or not having benzene ring, possessing following characteristics known as aromaticity.
- Planarity.
  - Complete delocalisation of the  $\pi$  electrons in the ring.
  - Presence of  $(4n+2)\pi$  electrons in the ring where  $n$  is an integer ( $n=0, 1, 2, \dots$ ). This is often referred to as Huckel Rule.

19. Benzene is commercially isolated from coaltar. However, it may be prepared in the laboratory by the following methods.

**a) Decarboxylation of Sodium Salt of aromatic acids :**



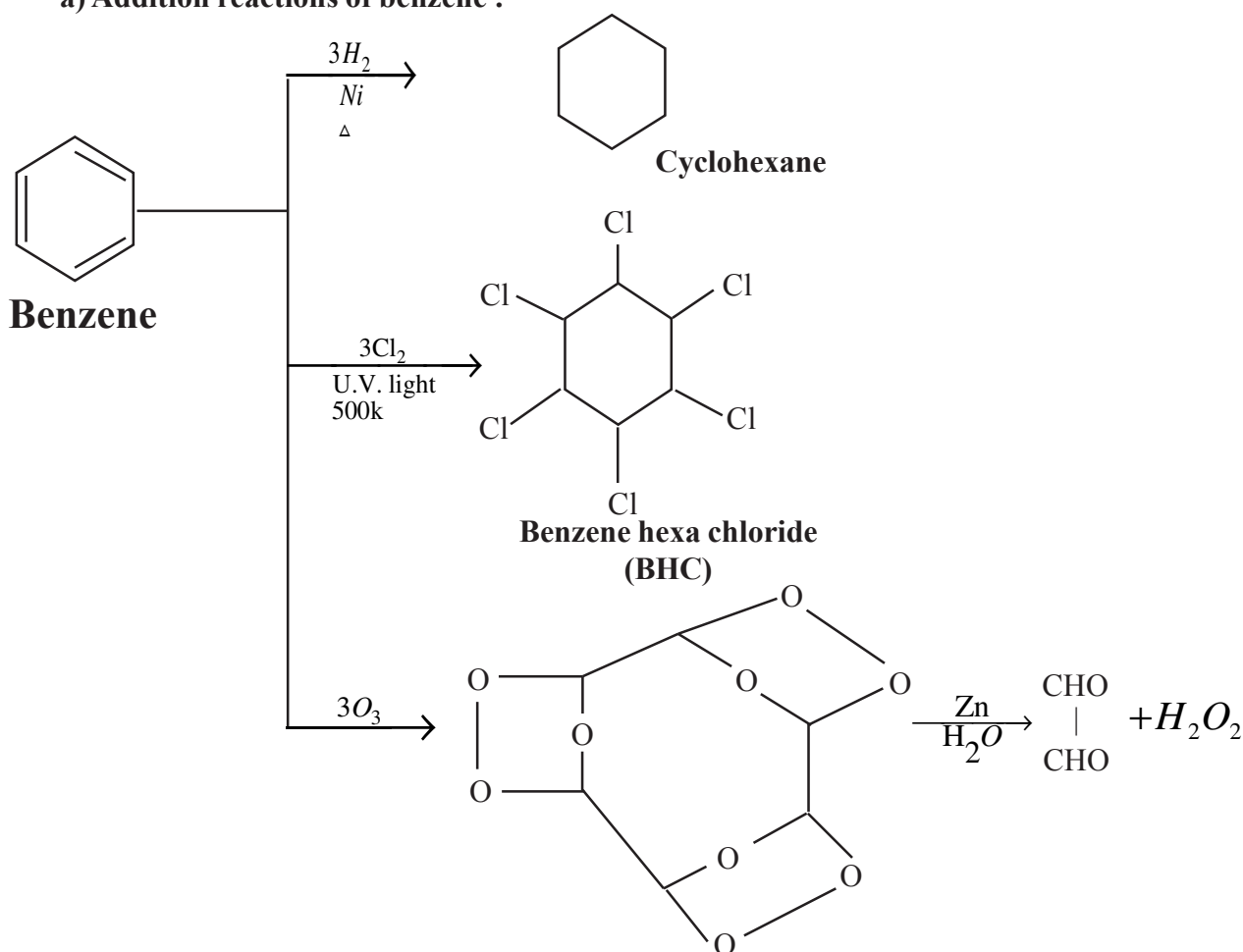
**b) Reduction of Phenol :**



20. Aromatic hydrocarbons are non-polar, colourless liquids or solids with a characteristic aroma. Aromatic hydrocarbons are immiscible with water but are readily miscible with organic solvent. They burn with sooty flame.

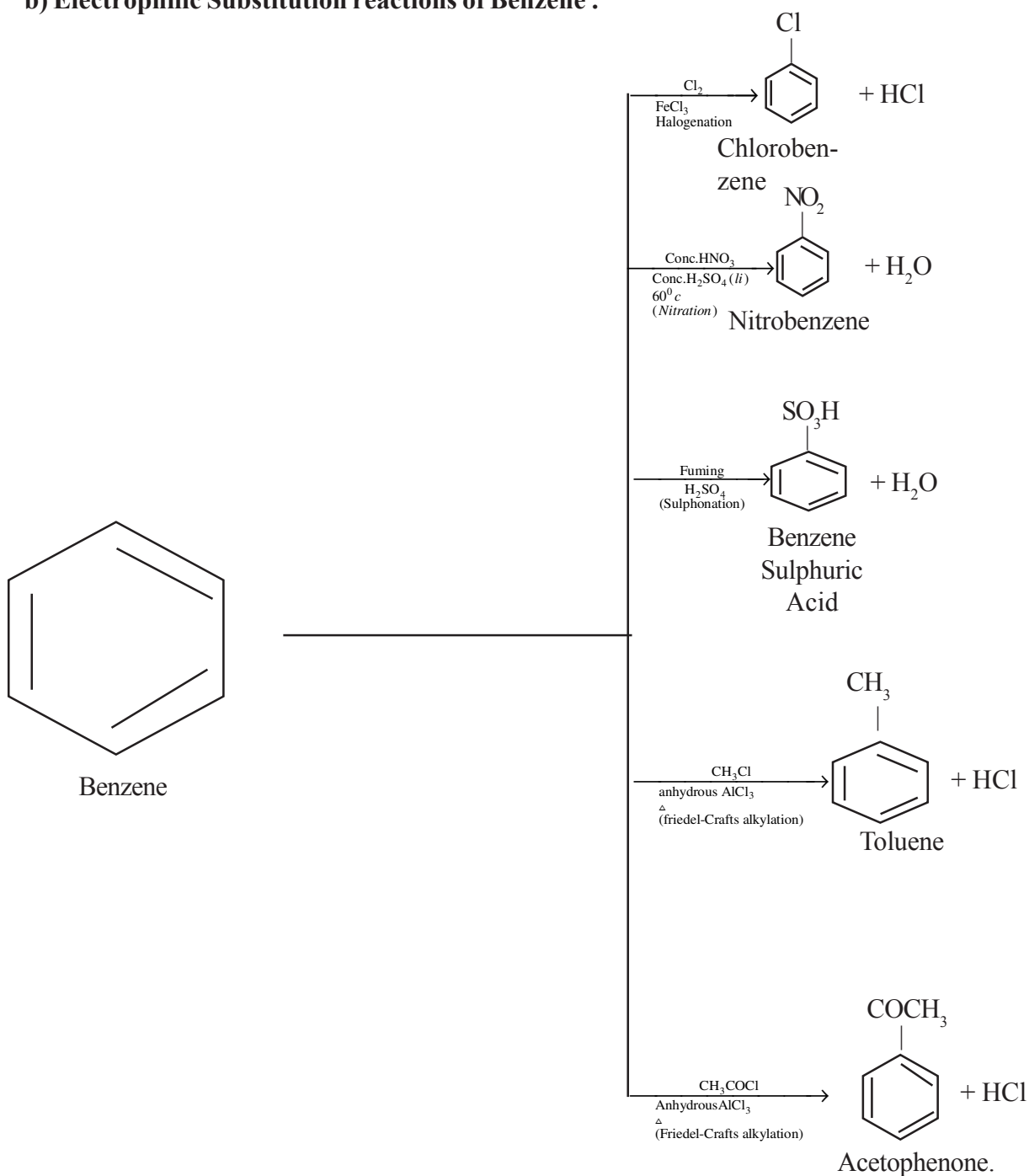
21. Arenes are characterised by electrophilic substitution reaction. However, under special conditions they can also undergo addition and oxidation reactions.

**a) Addition reactions of benzene :**

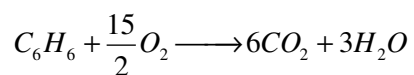
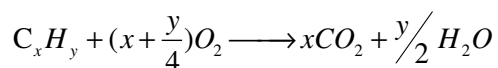




**b) Electrophilic Substitution reactions of Benzene :**



**c) Oxidation (Combustion) :**



22. During electrophilic substitution reaction the groups which direct the incoming group to meta position are called meta directing groups. Some examples of these groups are  $-\text{NO}_2$ ,  $-\text{CN}$ ,  $-\text{CHO}$ ,  $-\text{COOH}$ ,  $-\text{COR}$ ,  $-\text{SO}_3\text{H}$ ,  $-\text{COOR}$  etc.

The groups which direct the incoming group to ortho and para position are called orth-para directing groups. Some examples of these groups are  $-\text{NH}_2$ ,  $-\text{OH}$ ,  $-\text{OR}$ ,  $-\text{R}$ ,  $-\text{X}$  etc.

23. Benzene and Polynuclear hydrocarbon containing more than two benzene rings fused together are toxic and said to passes carcinogenic property. They enter into human body and undergo various biochemical reactions and finally damage DNA and cause cancer.

**A. Multiple Choice Question :**

**( 1 Mark for each question ) :**

1. The General Formula of Alkane is -

- a)  $\text{C}_n\text{H}_{2n+3}$       b)  $\text{C}_n\text{H}_{2n+1}$       c)  $\text{C}_n\text{H}_{2n+2}$       d)  $\text{C}_n\text{H}_{2n}$

2. Which of the following is known as marsh gas?

- a)  $\text{CH}_4$       b)  $\text{C}_2\text{H}_6$       c)  $\text{C}_2\text{H}_2$       d)  $\text{C}_2\text{H}_4$

3. Which of the following method is not used for the preparation of  $\text{CH}_4$ ?

- a) Wurtz's Reaction      b) Kolbe's electrolysis  
c) Sabatier and Senderen's Reaction      d) All of these.

4. The Compound with highest boiling point is -

- a) n - pentane      b) n-hexane      c) 2-methyl butane      d) 2, 2-dimethyl propane

5. Which of the following metal is used for the preparation of alkane in Wurtz's Reaction?

- a) K      b) Na      c) Mg      d) Al

6. Of the five isomeric hexanes, the isomer which can give two monochlorinated compounds is -

- a) n-hexane      b) 2-methylpentane      c) 2, 2-dimethylbutane      d) 2- methylbutane

7. The correct order of chemical reactivity with alkane is -

- a)  $\text{Br}_2 < \text{I}_2 < \text{Cl}_2 < \text{F}_2$   
b)  $\text{I}_2 < \text{Br}_2 < \text{Cl}_2 < \text{F}_2$   
c)  $\text{F}_2 < \text{Cl}_2 < \text{Br}_2 < \text{I}_2$   
d)  $\text{Cl}_2 < \text{F}_2 < \text{Br}_2 < \text{I}_2$

8. Which of the following reagent is used for the preparation of ethene from ethyl bromide?  
 a) alcoholic KOH    b) aqueous KOH  
 c) aqueous NaOH    d) Na-Hg and H<sub>2</sub>O
9. What is the role of Zn metal in Ozonolysis?  
 a) Catalyse the reaction.  
 b) Prevention of Oxidation of Ozonised Products.  
 c) Propagation of Reaction  
 d) All of these.
10. Which of the following take part in Anti Markovnikov addition reaction?  
 a) HCl    b) HI    c) HF    d) HBr
11. Which of the following is more acidic?  
 a) C<sub>2</sub>H<sub>2</sub>    b) C<sub>2</sub>H<sub>4</sub>    c) C<sub>2</sub>H<sub>6</sub>    d) CH<sub>4</sub>
12. The reagent used to distinguish between ethene and ethyne is -  
 a) Br<sub>2</sub> water    b) Ammonical Cu<sub>2</sub>Cl<sub>2</sub>  
 c) H<sub>2</sub> and Ni    d) Cl<sub>2</sub> Water.
13. 1-pentyne  $\xrightarrow[CH_3I, THF]{NaNH_2}$  X  $\xrightarrow[Pd/BaSO_4]{H_2}$  Y, the compound Y is -  
 a) Cis-2-hexene    b) trans-2-hexene  
 c) Hexene    d) 2-Pentene
14. C-C bond length is the longest is -  
 a) C<sub>2</sub>H<sub>4</sub>    b) C<sub>2</sub>H<sub>6</sub>    c) C<sub>2</sub>H<sub>2</sub>    d) C<sub>6</sub>H<sub>6</sub>
15. Which of the following alkene on ozonolysis gives acetone and Acetaldehyde?  
 a) 2 - methyl - 2 - butene  
 b) 2, 3 - dimethyl - 2 - butene  
 c) 2, 4 - dimethyl - 2 - pentene  
 d) 2 - butene
16. CH<sub>3</sub> - C  $\equiv$  CH  $\xrightarrow[333k]{H_2O, Hg^{+2} / H^+}$  X, the product X is -  
 a) Propane    b) Propanol    c) Propanal    d) Propanone

17. Propene when reacts with HBr in presence of peroxide which of the following compound is obtained?

- a) Isopropyl bromide      b) n - Propylbromide  
c) 3 - Bromopropene      d) Allyl bromide

18. Which of the following is an aromatic compound?



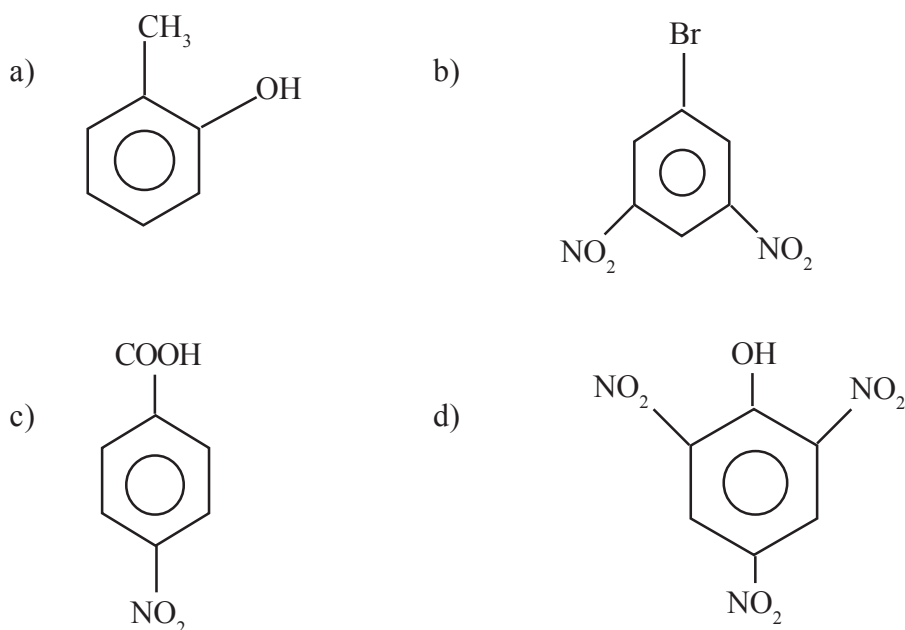
19. The product formed by passing ethyne through red hot iron tube is -

- a) Toluene      b) Benzene      c) ethene      d) cyclohexane

20. Which of the following is not ortho-para directing group in electrophilic aromatic substitution reaction?

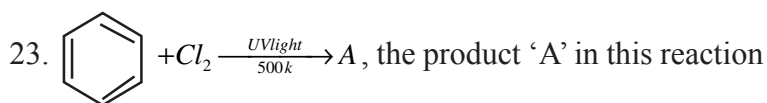
- a) -COOH      b) -OH      c) -NH<sub>2</sub>      d) -CH<sub>3</sub>

21. The structural formula of picric acid is -



22. Which of the following is less reactive towards electrophilic substitution Reaction?

- a) Benzene      b) Toluene      c) Chlorobenzene      d) Nitrobenzene



- a) chlorobenzene      b) benzenehexachloride  
c) chloroform          d) di chlorobenzene

24. When benzene is treated with H<sub>2</sub> in presence of Ni at higher temperature, then the hybridisation of carbon atom changes from-

- a) SP<sup>3</sup> to SP<sup>2</sup>      b) SP<sup>2</sup> to SP          c) SP to SP<sup>2</sup>          d) SP<sup>2</sup> to SP<sup>3</sup>

25. Ethylene reacts with Baeyers reagent to give -

- a) glycol      b) acetaldehyde    c) oxalic acid      d) ethyl alcohol.

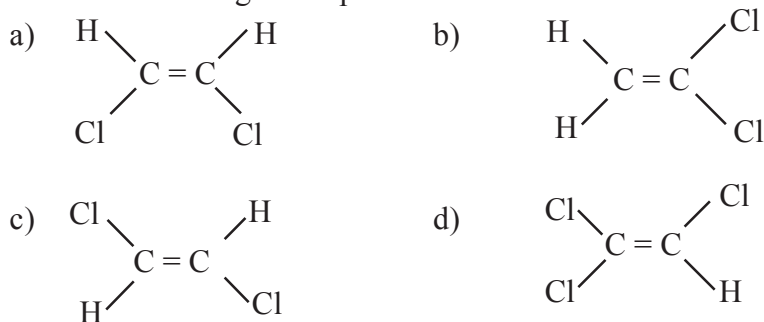
26. Which of the following is more stable?

- a) ethene      b) ethyne    c) ethane    d) benzene

27. Benzene reacts with acetyl chloride in presence of anhydrous AlCl<sub>3</sub> to give -

- a) Toluene      b) Acetophenone      c) Benzophenone      d) Ethyl Benzene

28. Which of the following is non polar?



**B. Assertion and Reason type questions :**

**Each question marks 1**

- a) Both assertion and reason are correct and reason is the correct explanation.  
b) Both assertion and reason are correct but reason is not the correct explanation.  
c) Assertion is correct but reason is incorrect.  
d) Assertion is incorrect but reason is correct.

1. Assertion : The branched chain alkanes are more volatile.  
Reason : Branched chain alkanes are smaller in size.

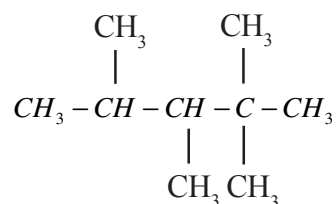
2. Assertion : Addition of bromine to trans-but-2-ene yields meso -2, 3 - di bromo butane.  
Reason : Addition bromine to an alkene is an electrophilic addition.
3. Assertion : Ethyne is more acidic compared to ethane.  
Reason : Carbon atoms in ethyne is sp hybridised
4. Assertion : Cyclo propenyl cation is an aromatic compound.  
Reason : Cyclopropenyl cation obey all the rules of aromaticity.
5. Assertion : Nitration of benzene occur more readily compared to toluene.  
Reason : Toluene is more electron rich compared to benzene.
6. Assertion : Toluene form ortho - Chlorotoluene when reacts with chlorine in presence of  $\text{FeCl}_3$ .  
Reason : Methyl group is ortho/para directing group.
7. Assertion : Alkanes are known as paraffins.  
Reason : Alkanes take part in substitution reaction.
8. Assertion : 2-butane is more stable than ethene.  
Reason : 2-butane have large no. of  $\alpha-H$  compared to ethene.
9. Assertion : Boiling point of n-butane is higher than that of propane.  
Reason : Boiling point increases with increase in molecular mass.
10. Assertion : Cis-2-butene is more polar than trans-2-butene.  
Reason : Dipole moment of Cis-2-butene is less compared to trans-2-butene.

**C. Very short answer question :**

**( Each question 1 mark)**

1. What is the main source of energy for domestic fuel?
2. What type of isomerism arises in alkane?

3. Identify the 4<sup>o</sup> carbon atom in the following compound :



4. How many sigma ( $\delta$ ) bond and Pi ( $\pi$ ) bond are present in but-1-en-3-yne?

5. Sodium salt of which acid will be needed for the preparation of propane?

6. Write the equation for general combustion equation for any alkane.

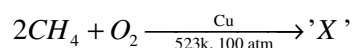
7. Name the catalyst used for hydrogenation of alkene.

8. What is Sodalime?

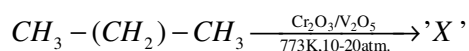
9. When methylchloride is reduced with Zinc and dilute hydrochloric acid, which compound is formed?

10. How boiling point change with increase in molecular mass of alkane?

11. Identify 'X' from the following reaction?

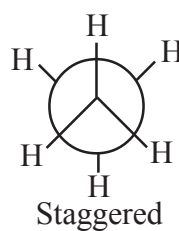
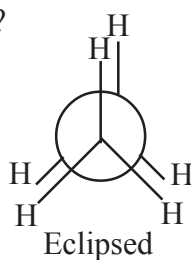


12. Identify 'X' from the following reaction?



13. What is aromatization?

14. Which one is more stable?



16. What is torsional strain?

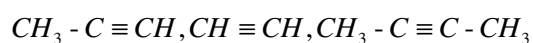
17. Write the structure of cis-2-butene and trans-2-butene.

18. Give an example of  $\beta$ -elimination reaction?

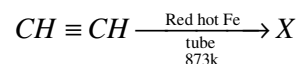
19. Write the name of ozonised products of 2-methyl-2-butene.

20. What is Baeyer's Reagent?

21. Arrange the following in increasing order of acidic nature?



22. Identify 'X' from the following reaction :



23. What is arene?

24. According to Huckel's rule how many  $\pi$  electrons are present in aromatic compounds?

25. What is nitrating mixture?

26. What is anti aromatic compound?

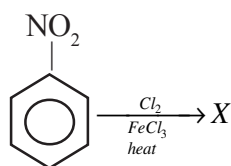
27. Write down the structural formula of Anthracene and phenanthrene.

28. What is the role of anhydrous  $AlCl_3$  in Friedel-Crafts reaction?

29. What is gamaxne?

30. Identify the compound 'X' from the following Reaction?





31. What type of substitution reaction occur in benzene?
32. Arrange the following in increasing order of in their stability -  
2 - butene, ethene, 2-methyl-2-butene, 2, 3-dimethyl-2-butene.

**D. Short Answer type questions :**

**(Each question marks - 2)**

1. Write structure of different chain isomers of alkanes corresponding to the molecular formula  $\text{C}_5\text{H}_{12}$ .
2. What is Wurtz's Reaction? Give example.
3. Write a short note on 'Sabatier and Senderens' reaction.
4. What happened and write down the balanced equation when ethane burns in presence of excess oxygen?
5. What effect does branching of an alkane chain has on its boiling point ?
6. How will you carry out the following conversion?
  - i)  $\text{CH}_4 \longrightarrow \text{HCHO}$
  - ii)  $(\text{CH}_3)_3\text{CH} \longrightarrow (\text{CH}_3)_3\text{COH}$
7. What is eclipsed and staggered conformation?

8. Draw the eclipsed and staggered form of ethane.
9. How can you prepare ethene from the following?  
a) Ethyne            b) Ethyl Bromide
10. State Markovnikov's rule with example.
11. What is anti Markovnikov's addition? Give example.
12. Explain why ethyne is acidic in nature.
13. What is polymerisation reaction? Give example.
14. Name the monomer unit of polyethene and P. V. C.
15. Complete the following reactions:
- a)  $CH \equiv CH + H_2O \xrightarrow[333K]{Hg^{+2}/H^+}$
- b)  $CH \equiv CH \xrightarrow[NH_4OH]{CuCl_2}$
16. Explain why benzene is an aromatic compound.
17. How can you prepare benzene from the following compounds?  
a) Sodium Benzoate   b) Phenol

18. Write a short 'Nitration of benzene'.
19. Why is benzene extraordinary stable though it contains three double bonds?
20. Out of benzene, toluene and m-dinitrobenzene which will undergo nitration reaction most easily and why?
21. Write a short note on 'Toxicity of Polynuclear Hydrocarbons.'
22. Write the resonating structure of benzene.
23. How will you convert the following:
  - a) Benzene  $\longrightarrow$  Toluene
  - b) Benzene  $\longrightarrow$  Benzene Sulphonic acid.
24. What do you mean by ortho and para directing groups?

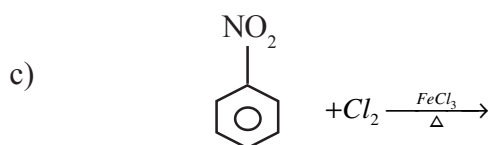
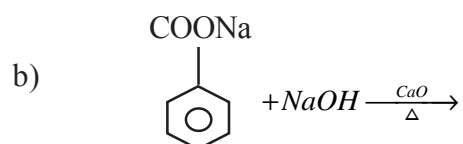
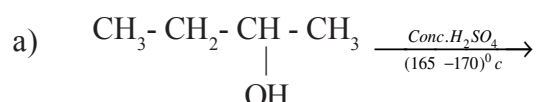
**E. Short Answer type Questions:****( Each question marks - 3)**

- Write IUPAC name of the following compounds:
  - $(\text{CH}_3)_3\text{CCH}_2\text{C}(\text{CH}_3)_3$
  - $(\text{CH}_3)_2\text{C}(\text{C}_2\text{H}_5)_2$
  - $(\text{CH}_3)\text{C}(\text{CH}_3)_2\text{C}(\text{CH}_3)_2\text{CH}_3$
- Write a short note on -
  - Kolbe's electrolysis Reaction.
  - Friedel-Crafts alkylation Reaction.
- What is Grignard Reagent? How can you prepare ethane by using Grignard's Reagent?
- Write chemical equations for combustion reaction of the following hydrocarbons :
  - Butane
  - Benzene
  - Ethyne
- Write the different steps involved in the mechanism of chlorination of methane in presence of sunlight.
- Identify A, B and C from the following reaction and also write their IUPAC name:
  - $\text{CH}_3 - \text{CH} = \text{CH}_2 + \text{HI} \longrightarrow \text{A}$
  - $\text{CH}_3 - \text{C} = \text{CH}_2 + \text{HBr} \xrightarrow{\text{Peroxide}} \text{B}$
  - $\text{CH}_3 - \text{CH} = \text{CH}_2 + \text{H}_2\text{O} \xrightarrow{\text{H}^+} \text{C}$
- What is polymerisation reaction? How can you prepare the following polymers? (chemical reactions only)
  - Polypropene
  - Polythene
- How will you convert the following compounds?
  - Ethene  $\longrightarrow$  ethanol
  - 2, 3, - di bromo butane  $\longrightarrow$  2 butene.
  - ethyne  $\longrightarrow$  ethanal
- Why is Wurtz reaction not preferred for the preparation of alkanes containing odd number of carbon atoms? Illustrate your answer by taking one example.

10. How will you convert benzene into -

- a) P- nitrotoluene    b) Benzenesulphonic acid  
c) Acetophenone

11. Complete the following reaction :-



12. a) What happens when calcium carbide reacts with cold water? Write the balance chemical equation.

b) What is Paraffin?

13. Write the structures for the following:

a) An alkyne has molecular formula  $\text{C}_4\text{H}_6$  do not react with ammoniacal Cuprous chloride solution.

b) An alkene on ozonolysis gives two molecules of Acetaldehyde.

c) An alkane has molecular formula  $\text{C}_5\text{H}_{12}$  which gives only one monochlorinated compound.

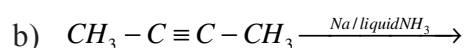
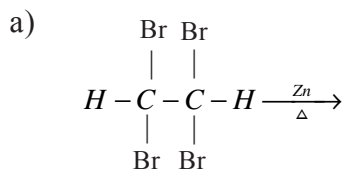
14. An alkyl halide  $\text{C}_5\text{H}_{11}\text{Br}$  (A) reacts with alcoholic KOH to give an alkene 'B' which reacts with  $\text{Br}_2$  water to give compound 'C' which on dehydrobromination gives an alkyne 'D'. Compound when treated with Na metal form a sodium salt. Identify the compound A, B, C and D. Give the reactions involved.

15. An unsaturated hydrocarbon 'A' adds two molecules of  $\text{H}_2$  and on ozonolysis gives ethanal, butane-1, dial and propanone. Give the structure of 'A'. Write its IUPAC name and explain the reactions involved.

**F. Long Answer type question:****(Each questions marks - 5)**

1. i) Addition of HBr to propene yields 2-bromopropane, while in the presence of benzyl peroxide, the same reaction yields 1-bromopropane. Explain and give mechanism.

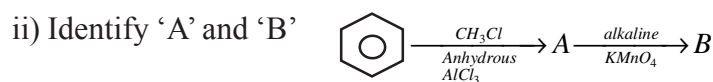
ii) Complete the following reaction:



2. i) What are the necessary conditions for any system to be aromatic?

ii) What is pyrolysis of alkane? Give example.

3. i) What is Sulphonation of Benzene? Explain with Mechanism.



4. Conversion:

a) Benzene  $\longrightarrow$  m-di Nitrobenzene

b) Benzoic Acid  $\longrightarrow$  Benzene

c) Benzene  $\longrightarrow$  Benzaldehyde

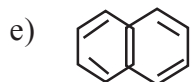
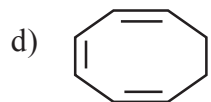
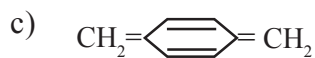
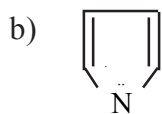
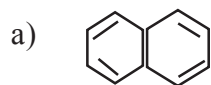
d) Benzene Sulphonic acid  $\longrightarrow$  Benzene

e) Benzene diazonium Chloride  $\longrightarrow$  Benzene

5. i) Draw Newmann and Sawhorse projection for the eclipsed and staggered conformers of ethane. Justify their stability.

ii) "Rotation around Carbon-Carbon (C - C) Single bond of ethane is not completely free" - Explain.

6. Which of the following compounds are aromatic and why according to Huckel's rule?



3+2

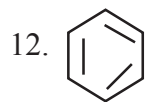
7. i) 896 ml of a hydrocarbon 'A' having carbon 87.80% and hydrogen 12.19% weighs 3.28g at STP. Hydrogenation of 'A' gives 2-methylpentane. Also 'A' on hydration in the presence of  $\text{H}_2\text{SO}_4$  and  $\text{HgSO}_4$  gives a Ketone 'B' having molecular formula  $\text{C}_6\text{H}_{12}\text{O}$ . Identify A and B and also write the reactions involved.

ii) Write the resonance structures of Nitrobenzene.

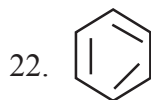
## Solution

- A.
- |       |       |       |
|-------|-------|-------|
| 1. c  | 11. a | 21. d |
| 2. a  | 12. b | 22. d |
| 3. d  | 13.   | 23. b |
| 4. b  | 14. b | 24. d |
| 5. b  | 15. a | 25. a |
| 6. c  | 16. d | 26. d |
| 7. c  | 17. a | 27. b |
| 8. a  | 18. c | 28. a |
| 9. b  | 19. b |       |
| 10. d | 20. a |       |

- C. 11.  $\text{CH}_3\text{OH}$

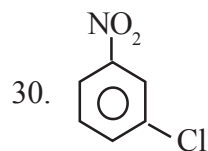


21.  $\text{CH}_3 - \text{C} \equiv \text{C} - \text{CH}_3 < \text{CH}_3 - \text{C} \equiv \text{CH} < \text{CH} \equiv \text{CH}$

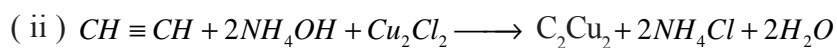




24.  $(4n+2)\pi$  electrons where,  $n = 0, 1, 2, 3, 4, \dots$



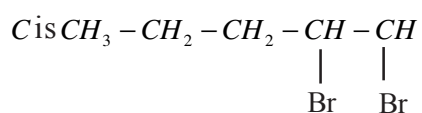
D. 15. (i)  $\text{CH}_3\text{CHO}$



E.

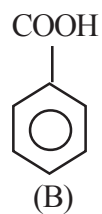
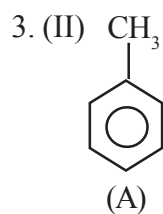
14. A is  $\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{Br}$

B is  $\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH} = \text{CH}_2$

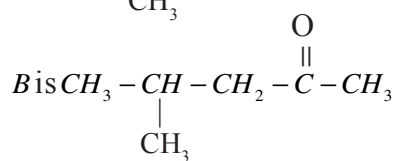


D is  $\text{CH}_3 - \text{CH}_2 - \text{CH}_3 - \text{CH} \equiv \text{CH}$

F.



7. A is  $\text{CH}_3 - \underset{\text{CH}_3}{\text{CH}} - \text{C} \equiv \text{C} - \text{CH}_3 (\text{C}_6\text{H}_{10})$



**Chapter at a glance**

Environmental chemistry deals with the study of the origin, transport, reactions, effects and fates of chemical species in the environment. Environmental pollution is the effect of undesirable changes in our surroundings that have harmful effects on plants, animal and human beings. Environmental pollution happens due to natural phenomenon or due to human activity. We can divide environmental pollution in three part -

- i) Atmospheric pollution                      ii) Water pollution                      iii) Soil pollution

**Pollutants:**

A substance, which causes pollution, is known as pollutant. Pollutants can be solid, liquids or gaseous and their concentration is higher than natural abundance. A small amount of pollutant in the air become more harmful compared to similar levels present in food. Some pollutants may remain in the environment in an unchanged form for many decades and can be more harmful for environment. This pollutants cannot be degraded by natural processes, example: DDT, Plastic, heavy metal, nuclear waste etc. These are also called as non-biodegradable pollutants.

**Atmospheric pollution:**

Among the four layers of atmosphere, tropospheric (upto 10Km from Sea level) and Stratospheric (from 10Km to 50 Km) pollution is generally studied as atmospheric pollution. Troposphere contains dust, turbulent air, water vapour and clouds. Stratosphere contains di-nitrogen, di-oxygen, Ozone and water vapour.

**Tropospheric pollution occurs due to following gaseous and particulate pollutants:**

1. Gaseous air pollutant: Oxides of sulphur, nitrogen, carbon, hydrocarbons etc.
2. Particulate Pollutant: These are further classified in two parts.
  - i. Viable Particulate like - Bacteria, fungi, moulds, algae etc.
  - ii. Non-Viable particulate -  
Cigarette smoke, burning of fossil fuel, oil smoke, dust, saw dust from wood work, cement and fly ash, pulverized coal, sulphuric acid mist, mist of herbicides and pesticides, organic solvent, metals and metallic oxides, or fume particles, photochemical smog etc.

**Harmful effects of some gaseous air pollutants:**

1. Oxides of Sulphur: Causes asthma, bronchitis, emphysema, irritation to the eyes resulting in tears and redness, lung diseases, lung cancer, stiffness of flower buds cough etc.
2. Oxides of Nitrogen: Causes leaf spotting and breakdown of plant tissues, decreases the rate of photosynthesis, bronchitis and acute respiratory problem for children, toxic to living tissues harmful to various type of fibres and metals.
3. Hydro Carbon: Hydrocarbons are carcinogenic, they harm plants by causing aging, breakdown of tissues and shedding of leaves, flowers and twigs.
4. Oxides of Carbon: Carbon mono oxide combines with haemoglobin to form carboxy haemoglobin and oxygen transportation capability of haemoglobin as a result headache, weak eye sight,

nervousness, Cardiovascular disorder, premature birth of baby, spontaneous abortion, deformed babies etc may occur.  $\text{CO}_2$  is responsible global warming.

### Global warming and Green house effect:

Some portion of the reflected IR radiations are absorbed by mainly  $\text{CO}_2$  and other gases like  $\text{CH}_4$ ,  $\text{O}_3$ ,  $\text{N}_2\text{O}$ ,  $\text{CFCl}_3$ ,  $\text{CF}_2\text{Cl}_2$  etc. This radiations are re-emitted back towards earth's surface and heat up the atmosphere. This is known as Green house gas because these gases act as natural green house.

### Sources of green house gas:

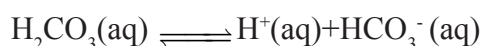
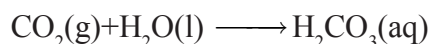
- Methane is produced naturally when vegetation is burnt, digested or rotted in the absence of oxygen, large quantity of methane are released in paddy fields.
- Coal mines, from lakes, fossil, fuels etc.
- Water vapour, ozone, nitrous oxide are produced naturally
- Chlorofluorocarbons are human made industrial waste.

### Acid Rain:

The pH value of rain water is generally 5.6. If the value falls below 5.6 then it is known as acid rain. Burning of petrol or diesel in motor engine, fossil fuels and coals in nuclear reactors thermal power plants produces large quantity of  $\text{SO}_2$  and oxides of nitrogen. This gases are oxidised by soot particles present in atmosphere, mix up with rain water to produce acid rain.

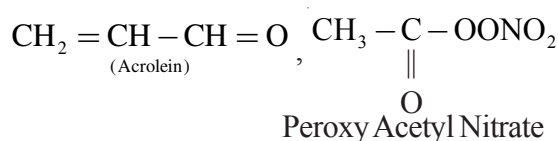
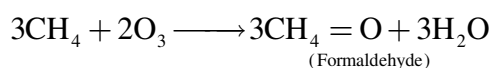
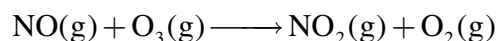
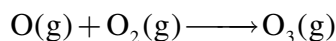
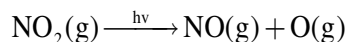


Gaseous  $\text{CO}_2$  dissolves in rain water and produce weak carbonic acid -



### Photochemical Smog:

Generally, Photo chemical smog are composed of ozone, nitric oxide, acrolein, formal dehyde and peroxy acetyl nitrate (PAN). The reactions involved in the making of photo chemical smog are as follows:



### Harmful effects of photo chemical smog:

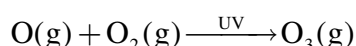
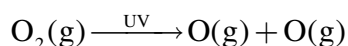
Ozone, PAN, aldehydes cause irritation of eyes and burning sensation in throat and respiratory tract, headache, chest pain, dryness of throat, cough and difficulty in breathing. Photo chemical smog causes corrosion of metals, stones, building materials, rubber and painted surfaces.

### Control of Photo chemical smog:

Use of catalytic converter in motor vehicles can transform nitrogeneous oxide and hydrocarbon to other less harmful gas or compounds, certain plant like pinus, Juniparus, vitis etc can metabolise nitrogen dioxide and reduce its concentration in atmosphere.

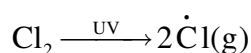
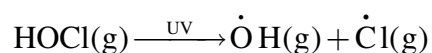
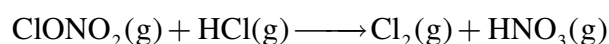
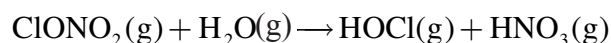
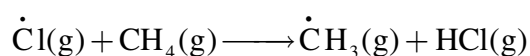
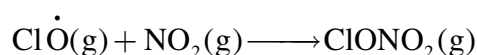
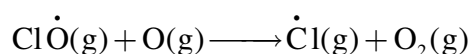
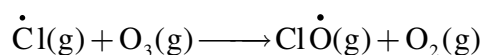
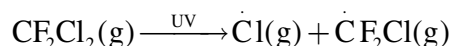
### Ozone and Ozone hole:

The UV radiation split apart molecular oxygen(O) atoms. These oxygen atoms combine with the molecular oxygen to form Ozone.



CFC's (Chloro Fluro Carbons) broke down by powerful UV radiations, releasing chlorine free radical. The Chlorine free radical react with ozone to form oxygen molecule and chlorine monoxide radicals. Chlorine monoxide radical further react with atomic oxygen and produces chlorine free radical and the chain reaction propagates and depletion of ozone layer takes place. Depletion of ozone layer in south pole is specially known as ozone hole.

Some reactions involved in depletion of ozone layer are as follows -



### Effects of Ozone layer depletion:

Due to depletion of Ozone layer UV radiation can penetrate through troposphere to reach earth. UV radiation lead to ageing of skin, cataract, sunburn, skin cancer, killing of many phytoplankton and dame

to fish productivity, increased evaporation of surface water through the stomata of the leaves, decrease moisture content of the soil, damage of paints, fibres etc can take place.

### **Water Pollution:**

Surface water or ground water gets polluted by different activity of humans. Municipal and industrial discharge pipes etc where pollutants enter the water source and can be identified easily are known as point source. On the other hand, agricultural run-off, acid rain, storm water draining etc where the source of pollution cannot be easily identified are known as non-point source of pollution.

### **Reason of Water pollution:**

- i. Pathogens: bacteria like *Escherichia coli*, *Streptococcus fecalis* etc from human excreta.
- ii. Organic wastes: leaves, grass, trash, excessive growth of phytoplanktons etc which are biodegradable.
- iii. Chemical pollutants and their harmful effects:
  - a) *Cadmium, mercury, nickel*: damage kidney, central nervous system, liver etc.
  - b) *Pesticides, polychlorinated diphenyls*: they are carcinogenic.
  - c) *Detergents*: increase algae in water, reduction of dissolved oxygen, death of fish and other animals, eutrophication.
  - d) *Fluoride*: very low concentration can cause tooth decay, over concentration (>2ppm) can lead to brown mottling of teeth excessive fluoride (over 10ppm) causes decay of bones and teeth.
  - e) *Lead*: above 50 ppm in drinking water, causes disease in kidney, liver, reproductive system etc.
  - f) *Sulphate*: excess sulphate in drinking water (>500ppm) causes laxative effect.
  - g) *Nitrate*: above the level of 50 ppm can cause 'blueberry syndrome' or metheno globinemia.

### **Dissolved oxygen, Biochemical Oxygen demand (BOD):**

In cold water dissolved oxygen can reach a concentration upto 10ppm, it is generally known as dissolved oxygen, on the other hand the amount of oxygen required by bacteria to breakdown the organic matter present in a certain volume of a sample of water is called Bio chemical Oxygen Demand (BOD). In pure water its value is less than 5 ppm.

### **Soil Pollution:**

The main sources of soil pollution are dumping of industrial wastes, urban wastes, volatile organic compounds, radio-active wastes, chemical used in agricultural activities, fertilizers, pesticides, herbicides and agents to increase fertility of soil etc.

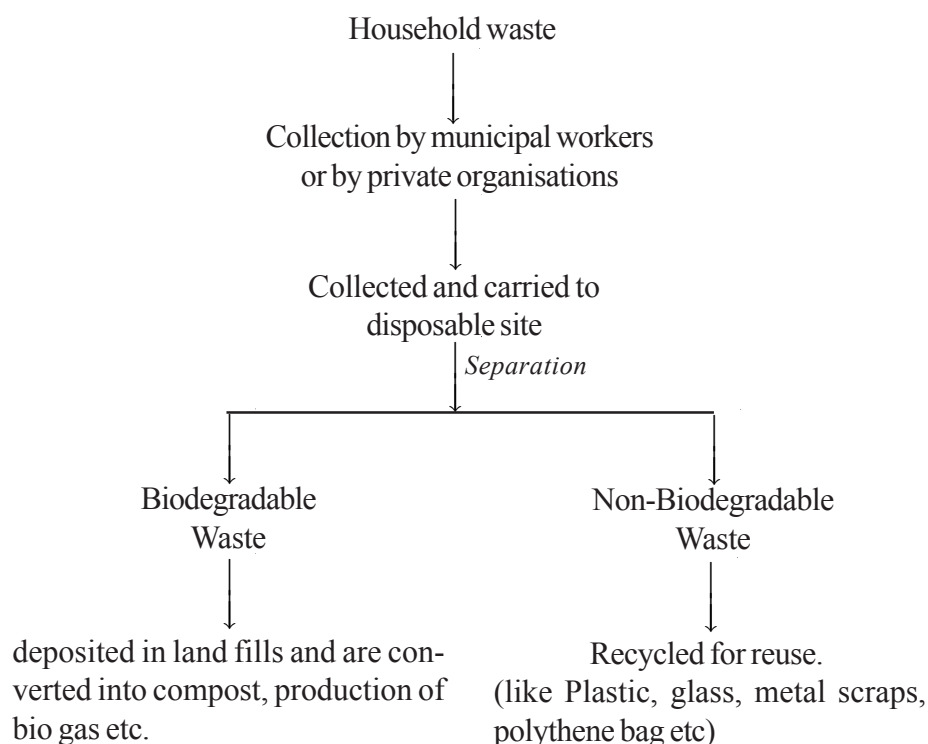
### **Different types of Pesticides and their harmful effects:**

1. Insecticides: DDT, malathion, BHC etc may cause physiological and metabolic disorders. Organic phosphates and carbamates are, very harmful for our nervous system and overall health.
2. Herbicides: Sodium Chlorate ( $\text{NaClO}_3$ ), Sodium Arsenate ( $\text{NaAsO}_3$ ), 2,4 di chloro phenoxy acid, triazine etc may cause birth defect. It has been found that excessive use of herbicides may increase insect growth in fields in the long run.
3. Fungicides: Phenyl mercury acetate, bordeaux mixture, methyl mercury etc are poisonous for human health. Methyl mercury can even cause deaths.

### Industrial Waste:

Biodegradable industrial waste are generated by cotton mills, food processing units, paper mills and textile factories. Non Biodegradable industrial waste like flyash, slag, mud, tailings, gypsum, inflammable hazardous chemicals, composite explosives, highly reactive substances etc. If not properly disposed can cause serious environmental pollution.

### Waste Management Pollution :

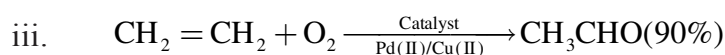


### Green Chemistry:

The idea of utilising the existing knowledge and principles of chemistry and other sciences to reduce the harmful effects of development on environment and at the same time maintaining the speed of sustainable development is known as green chemistry.

### Examples of application of Green Chemistry in day to day life:

- i. Using liquified carbondioxide with suitable detergent in place of carcinogenic tetra chloro ethene ( $\text{Cl}_2\text{C}=\text{CCl}_2$ ) for dry wash, use of  $\text{H}_2\text{O}_2$  in laundry for bleaching of clothes.
- ii. Use of  $\text{H}_2\text{O}_2$  with suitable catalyst for bleaching of paper in place of chlorine gas



**1. Find out the correct answer (MCQ) :**

**(Each question 1 Mark)**

1. The main contributor of acid rains is :  
a)  $\text{H}_2\text{SO}_4$     b)  $\text{CaCO}_3$     c)  $\text{CH}_3\text{COOH}$     d) PAN
2. Acid rain has pH in the range:  
a) 0-3                      b) 5- 6                      c) 6-7                      d) 7 - 10
3. Ozone in the stratosphere is depleted by -  
a)  $\text{CF}_2\text{Cl}_2$     b)  $\text{C}_7\text{F}_{16}$     c)  $\text{C}_6\text{H}_6\text{Cl}_6$     d)  $\text{C}_6\text{F}_6$
4. Besides  $\text{CO}_2$ , other greenhouse gas is -  
a)  $\text{CH}_4$                       b)  $\text{N}_2$                       c) Ar                      d)  $\text{O}_2$
5. Increased concentration of  $\text{CO}_2$  in atmosphere is responsible for:  
a) Green house effect                      b) Acid Rain  
c) Lack of photosynthesis                      d) Death of aquatic life
6. Burning of fossil fuels is the main source of -  
a) Nitrogen dioxide                      b) Nitric Oxide  
c) Nitrous Oxide                      d) Sulphur DiOxide
7. Depletion of Ozone layer causes -  
a) Blood Cancer                      b) Breast Cancer  
c) Lung Cancer                      d) Skin Cancer
8. Bhopal Gas tragedy was caused by:  
a) Methyl amine                      b) Methyl isocyanate  
c) Phosgene                      d) Chloroform
9. COD stands for:  
a) Chemical Oxygen Demand                      b) Controlled Oxygen Demand  
c) Clouds Causing Ozone Depletion                      d) Chlorinated Oxygen Demand
10. In upper stratosphere, Ozone protects us from -  
a) Infrared Radiation                      b) Ultra-Violet Radiations  
c) Carbon Dioxide                      d) Peroxy Acetyl Nitrate

11. DDT is:
- a) an antibiotic
  - b) Biodegradable Pollutant
  - c) Non-Degradable Pollutant
  - d) Non Pollutant
12. The clean water should have BOD value -
- a) less than 5 ppm
  - b) 10 ppm
  - c) 20 ppm
  - d) 100 ppm
13. Which of the following statement is incorrect?
- a) Taj Mahal is affected by hydrocarbon
  - b) Building are adversely affected by acid rain
  - c) Due to acid rain, micro organism are affected
  - d) Large amount of acid rain decreases soil fertility
14. Which of the following metal will pollute water?
- a) Cd
  - b) Na
  - c) K
  - d) None of these
15. The Major cause of air pollution in big cities is -
- a) Burning of Coal
  - b) Domestic exhaust
  - c) Burning of Cooking gas
  - d) Vehicular exhaust
16. Green Chemistry means such reaction which -
- a) Produce colour during reaction
  - b) Reduce the use and production of hazardous chemicals
  - c) Are related to the depletion of Ozone layer
  - d) Study the reactions in plant
17. The major air pollutant is -
- a) CO
  - b) Oxides of nitrogen
  - c) Soot
  - d) Oxides of Sulphur
18. Which one of the following particulate is most toxic?
- a) Fly ash
  - b) Soot
  - c) Inorganic Compound
  - d) Smog



19. The consequences of global warming may be -  
a) Increase in average temperature of the earth  
b) Melting of Himalayan Glaciers  
c) Increased biochemical oxygen demand  
d) All the above
20. Which of the following belongs to secondary air pollutants?  
a) CO  
b) Hydrocarbon  
c) Peroxy Acetyl Nitrate  
d) NO
21. Which of the following is a sink for CO?  
a) Micro Organism Present in the soil.  
b) Ocean  
c) Plant  
d) Haemoglobin
22. Among the following, which one is not a green house gas?  
a) Nitrous Oxide  
b) Methane  
c) Ozone  
d) Sulphur Dioxide
23. Which one of the following is not a common component of photo chemical smog?  
a) Ozone  
b) Acrolein  
c) Peroxy acetyl Nitrate  
d) Chloro Fluoro Carbon
24. Which is not a renewable source -  
a) Sunlight  
b) Coal  
c) Water  
d) Wind Energy
25. PAN stands for-  
a) Peroxy Acetyl Nitrate  
b) Pyruvic Acid Nitrate  
c) Phospho Acetyl Nitrate  
d) None of the above

**B. Assertion and reason type questions -**

**(Each question 1 mark)**

- a. Both assertion and reason are correct and reason is the correct explanation of the assertion.  
b. Both assertion and reason are correct but reason is not the correct explanation of the assertion.  
c. Assertion is correct but reason is incorrect.  
d. Assertion is incorrect but reason is correct.

1. Assertion(A): Green house effect was observed in houses used to grow plant and these are made of glass.  
Reason (R): Green house name has been given because glass houses are made of green glass.
2. Assertion(A): The pH of acid rain is less than 5.6.  
Reason (R): Carbon di-oxide present in the atmosphere dissolves in rain water and forms carbonic acid.
3. Assertion(A): Photochemical smog is oxidising in nature.  
Reason(R): Photochemical smog contains  $\text{NO}_2$  and  $\text{O}_3$ , which are formed during the sequence of reaction.
4. Assertion(A): Carbon di-oxide is one of the important green house gases.  
Reason(R): It is largely produced by respiratory function of animals and plants.
5. Assertion(A): Ozone is destroyed by Solar radiation in upper stratosphere.  
Reason(R): Thinning of the Ozone layer allows excessive UV radiations to reach the surface of earth.
6. Assertion(R): Excessive use of chlorinated synthetic pesticides causes soil and water pollution.  
Reason(R): Such pesticides are non-biodegradable.
7. Assertion(A): If BOD level of water in a reservoir is less than 5 ppm it is highly polluted.  
Reason(R): High biological oxygen demand means low activity of bacteria in water.
8. Assertion(A): Use of DDT in agricultural field is causing serious problem for our environment.  
Reason(R): DDT is a bio-degradable pollutant.
9. Assertion(A):  $\text{H}_2$  is a green house gas.  
Reason (R):  $\text{H}_2$  does not take part in trapping of reflected solar energy in the atmosphere.
10. Assertion (A): Liquified  $\text{CO}_2$  is used with detergent for dry cleaning now a days instead of tetrachloro ethene.  
Reason(R): Liquified  $\text{CO}_2$  with detergent will pollute water more compared to tetrachloro ethene.

**C. Very short answer type questions:**

**(Each question mark -1)**

1. Write the full form of DDT.
2. Which region of atmosphere contains Ozone?
3. Name two main green house gases.
4. What is PAN?
5. Name the full form of PCB.
6. Name major air pollutant.
7. Give example of photo chemical smog.
8. Write the full form of BOD.
9. Write the main component of photochemical smog.
10. Name two major water pollutant.

11. What happens when CO reacts with haemoglobin?
12. Give example of bio degradable pesticide?
13. Give example of non-biodegradable pesticide?
14. Give example of herbicide.
15. Name two secondary components of photochemical smog.
16. Name the acids present in acid rain?
17. Name two sources of dissolved oxygen in water.
18. Name a disease caused by excess amount of oxides of sulphur present in air.
19. Is there any effect of high concentration of  $\text{NO}_2$  on photosynthesis?
20. Name the green house gas used in Air Conditioner.
21.  $\text{CaCO}_3 + \text{H}_2\text{SO}_4 \rightarrow$   
Complete the above reaction.
22. Write down the disease caused by lead in human body.
23. Name two diseases caused by mercury pollution.
24. Name some diseases caused by photo chemical smog.
25. Why catalytic converter is used in motor vehicles?
26. What is blue baby syndrome?
27. Mention the bad effect of Tetrachloro ethylene used in dry wash.
28. Name one disease caused by arsenic contamination.
29. Pesticides are harmful- Give one example.
30. Name some tree which can help in metabolism of oxides of nitrogen.

**D. Short answer type question:**

**(Each question 2 marks)**

1. Define green house effect.
2. Define smog.
3. What is photo chemical smog? Give example?
4. What do you mean by Ozone hole?
5. Write two effects of depletion of Ozone layer.
6. Define eutrophication.
7. What do you mean by bio-degradable pollutant?
8. What do you mean by non-biodegradable pollutant?
9. What is global warming?
10. How can we control the global warming?
11. Name the sources of dissolved oxygen in water?
12. How ozone is produced in stratosphere?
13. Give chemical reactions for the conversion of  $\text{SO}_2$  into  $\text{SO}_3$ .
14. Name some pathogens which create water pollution.

15. How easily available detergents are causing water pollution?
16. Mention some negative effect of excess amount  $\text{NO}_2$  in air.
17. How excess  $\text{SO}_2$  in air creates health problems for human?
18. Write two steps to control soil pollution.
19. What is the importance of measuring BOD of a water body?
20. How do chlorofluorocarbons cause thinning of Ozone layer in stratosphere?
21. Write about some harmful effects of acid rain.
22. What is PCBs? How do they contaminate water?
23. How lead act as particulate pollutant?
24. What is MIC? Name tragedy associated with this gas in Bhopal?
25. Give example of one photochemical smog and classical smog.

**E. Short answer type question:**

**( Each question 3 marks)**

1. Write short note on formation of Ozone hole.
2. Write three possible consequences of global warming.
3. Write about harmful effects of oxides of sulphur.
4. Write down some harmful effects of oxides of nitrogen.
5. How acid rain is being created?
6. What is smog? Write about different types of smog.
7. Write down about causes of water pollution.
8. Mention some steps to reduce the rate of global warming.
9. How excess amount of lead, sulphate, nitrate, create health related problems.
10. Write the reactions involved in the formation process of photo chemical smog.
11. What is green chemistry? How it can help us reducing environmental pollution?
12. What will happen if all the green house gases disappear from atmosphere?
13. Write down the reactions involved in decaying of Ozone layer.
14. How house hold waste can be transformed into organic fertilizer?
15. What do you mean by point source and non-point sources of water pollution. Explain.
16. Carbon monoxide gas is more dangerous than carbon dioxide gas. Why?
17. Statues and monuments in India are affected by acid rain. How?
18. A large number of fish are suddenly found floating dead on a lake. There is no evidence of toxic dumping but you find an abundance of phytoplankton suggest a reason for the fish kill.
19. What do you mean by viable and non-viable particulates? Give example.
20. Give three examples where green chemistry has been applied.

**F. Long answer type questions:**

**( Each question 5 Marks)**

1. a. Why smoking is prohibited for pregnant women?  
b. What is polar stratopheric cloud? How it helps in Ozone depletion. Write the reaction involved in the above process.
2. Briefly describe the causes of water pollution.
3. What is green chemistry? Briefly describe how green chemists can help to protect our environment and lead to sustainable development.
4. What is green house effect? How global warming is creating negative environmental consequences? How can we reduce the speed of global warming?

## Solution

### A. Find out the correct answer : MCQ

- |       |       |       |
|-------|-------|-------|
| 1. a  | 11. c | 21. d |
| 2. b  | 12. a | 22. d |
| 3. a  | 13. a | 23. b |
| 4. a  | 14. a | 24. b |
| 5. a  | 15. d | 25. a |
| 6. d  | 16. b |       |
| 7. d  | 17. d |       |
| 8. b  | 18. d |       |
| 9. a  | 19. d |       |
| 10. b | 20. c |       |

### B. Assertion & Reason

- |      |      |      |      |       |      |
|------|------|------|------|-------|------|
| 1. a | 2.a  | 3.a  | 4. b | 5. b  | 6. d |
| 6. a | 7. c | 8. c | 9. d | 10. c |      |

## NOTE

## NOTE