

CHEMISTRY NOTES FOR FIRST YEAR IMPROVEMENT EXAM

1. SOME BASIC CONCEPTS OF CHEMISTRY



1. State Law of Multiple proportion. Who proposed this law?

It states that when two elements combine to form more than one compound, the different masses of one of the elements that combine with a fixed mass of the other, is in a simple whole number ratio. This law was proposed by **John Dalton**.

2. State Law of Definite (Constant) proportion. Who proposed this law?

It states that a given compound always contains exactly the same proportion of elements by weight. Or, the same compound always contains the same elements combined in a fixed ratio by mass. This law was proposed by **Joseph Proust**.

3. State Law of conservation of mass (matter). Who proposed this law?

It states that matter can neither be created nor destroyed. Or, in a chemical reaction, the total mass of reactants is equal to the total mass of products. This law was proposed by **Antoine Lavoisier**.

4. Define a.m.u

$1/12^{\text{th}}$ the mass of a C^{12} atom is called atomic mass unit (amu).

5. Define molar mass

The mass of one mole of a substance in gram is called its molar mass.

6. Define 1 mole.

1 mole is the amount of substance that contains as many particles as there are atoms in exactly 12 g C^{12} isotope. **1 mole of any substance contains 6.022×10^{23} atoms.** This number is known as **Avogadro number or Avogadro constant** (N_A or N_0).

7. Define mass percentage.

It is the percentage of each elements present in 100g of a substance.

i.e. percentage composition (mass percent) of an element =
$$\frac{\text{Mass of that element in the compound} \times 100}{\text{Molar mass of the compound}}$$

Or, Mass percent =
$$\frac{\text{Mass of solute} \times 100}{\text{Mass of solution}}$$

8. What are empirical and molecular formulae?

Empirical formula is the simplest formula which gives only the ratio of different elements present in the compound. But molecular formula is the actual formula that gives the exact number of different elements present in the sample.

9. Find the number of moles and molecules present in 90 g water.

No. of moles(n) =
$$\frac{\text{Given mass in gram (w)}}{\text{Molar mass (M)}} = \frac{90}{18} = 5 \text{ moles}$$

No. of molecules = no. of moles $\times 6.022 \times 10^{23} = 5 \times 6.022 \times 10^{23}$

10. What is limiting reagent?

The reagent which limits a reaction or the reagent which is completely consumed in a chemical reaction is called limiting reagent or limiting reactant.

11. Define molarity and molality.

Molarity is defined as the number of moles of solute dissolved per litre of solution.

i.e. Molarity (M) =
$$\frac{\text{Number of moles of solute (n)}}{\text{Volume of solution in litre (V)}}$$

Molality is defined as the number of moles of the solute present per kilogram (kg) of the solvent.

$$\text{i.e. Molality (m)} = \frac{\text{Number of moles of solute}}{\text{Mass of solvent in kg}}$$



12. Find the molarity of a solution prepared by dissolving 4g of NaOH dissolves in 250 mL solution.

$$\text{Molarity (M)} = \frac{\text{Number of moles of solute (n)}}{\text{Volume of solution in litre (V)}}$$

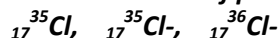
No. of moles of NaOH = Given mass/ Molar mass = $4/40 = 0.1$ mol

Volume of solution in L = $250/1000 = 0.25$ L

Molarity = $0.1/0.25 = 0.4$ M

2. STRUCTURE OF ATOM

1. Find the number of protons, electrons and neutrons in the following.



Ans: No. of Protons & electrons = Atomic number

No. of electrons in an ion = Atomic number + charge of the ion (for –ve ions) or, Atomic number – charge on the ion (for +ve ions)

No. of neutrons = Mass number – Atomic number

Species	No. of Protons	No. of electrons	No. of neutrons
${}_{17}^{35}\text{Cl}$	17	17	$35-17 = 18$
${}_{17}^{35}\text{Cl}^-$	17	18	$35-17 = 18$
${}_{17}^{36}\text{Cl}^-$	17	18	$36-17 = 19$

2. Give the scattering experiment conducted by Rutherford. What are the important observations and conclusions made by Rutherford?

Rutherford proposed an atom model based on his α -particle scattering experiment. He bombarded a very thin gold foil with α -particles. The gold foil had a circular fluorescent zinc sulphide screen around it.

Observations:

- Most of the α - particles passed through the gold foil without any deviation.
- A small fraction of the α -particles was deflected by small angles.
- A very few α - particles (~ 1 in 20,000) bounced back, that is, were deflected by nearly 180° .

Conclusions: From the above observations, Rutherford made the following conclusions:

- Since most of the α -particles passed through the foil without any deviation, most space in the atom is empty.
- A few positively charged α - particles were deflected. This is because the positive charge of the atom is concentrated in a very small volume at the centre called nucleus.
- The volume occupied by the nucleus is negligibly small as compared to the total volume of the atom.

3. What are the postulates of Rutherford atom model?

Rutherford proposed **the nuclear model (Planetary model) of atom**. According to this model:

- All the positive charge and most of the mass of the atom were concentrated in the centre called nucleus.
- Electrons are revolving round the nucleus with a very high speed in circular paths called orbits.
- Electrons and the nucleus are held together by electrostatic forces of attraction.

4. What are the drawbacks of Rutherford atom model?

- He could not explain the stability of the atom.
- He could not explain the electronic structure of atom.

5. What is photoelectric effect? What are its characteristics?

It is the phenomenon of ejection of electrons by certain metals (like potassium, rubidium, caesium etc.) when light of suitable frequency incident on them. The electrons ejected are called photoelectrons.

The important characteristics of photoelectric effect are:

- The electrons are ejected from the metal surface as soon as the beam of light strikes the surface.
- The number of electrons ejected is proportional to the intensity or brightness of light.
- For each metal, there is a minimum frequency (known as threshold frequency [ν_0]) below which photoelectric effect is not observed.

- iv. The kinetic energy of the ejected electrons is directly proportional to the frequency of the incident light.



6. Give the different lines present in the line spectrum of hydrogen atom.

The hydrogen spectrum consists of mainly five series of lines which are **Lyman, Balmer, Paschen, Brackett and Pfund series**. Among these lines, the Balmer series is the only series that we can be visible.

Series	Spectral region
Lyman	Ultra violet
Balmer	Visible
Paschen	Infra red
Brackett	Infra red
Pfund	Infra red

7. Give the important postulates of Bohr Atom model.

- The electron in the hydrogen atom can move around the nucleus in a circular path of fixed radius and energy. These paths are called orbits or stationary states or allowed energy states.
- The energy of an electron in the orbit does not change with time.
- The frequency of radiation absorbed or emitted when transition occurs between two stationary states that differ in energy by ΔE , is given by:

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

- The angular momentum of an electron is an integral multiple of $h/2\pi$.

8. What are the limitations of Bohr atom model?

- It could not explain the fine spectrum of hydrogen atom.
- It could not explain the spectrum of atoms other than hydrogen.
- It was unable to explain **Stark effect** and **Zeeman effect**.
- It could not explain the ability of atoms to form molecules by chemical bonds.

9. What do you mean by the dual nature of matter?

Matter has two types of nature – particle nature and wave nature. This is known as the dual nature of matter.

10. Give the de Broglies equation and explain the terms?

An equation relating wavelength (λ) and momentum (p) of a material particle is known as de Broglies equation. The equation is:

$$\lambda = h/mv \quad \text{Where } m \text{ is the mass of the particle, } v \text{ is the velocity and } p \text{ is the momentum.}$$

11. State Heisenberg's uncertainty principle. Give its mathematical form.

It states that "it is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of a moving microscopic particle like electron".

Mathematically, it can be given as in equation:

$$\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$$

Where Δx is the uncertainty in position and Δp (or, Δv) is the uncertainty in momentum (or velocity) of the particle.

12. What are Quantum numbers? Explain the different quantum numbers.

These are certain numbers used to explain the size, shape and orientation of orbitals. Or, Quantum numbers are the address of an electron. There are four quantum numbers which describe the electron in an atom. They are Principal Quantum number (n), Azimuthal Quantum number (l), Magnetic Quantum number (m or m_l) and Spin Quantum number (s)

1. Principal Quantum Number (n)

The following informations are obtained from n .

- It gives the size the orbit.
- It gives the energy of electron in an orbit.
- It gives the shell in which the electron is found.
- It also gives the average distance between the electron and the nucleus.

The possible values of n are 1, 2, 3, 4, 5 etc.

If $n = 1$ the electron is in K shell

$n = 2$ the electron is in L shell

$n = 3$ the electron is in M shell and so on.

2. Azimuthal Quantum Number [Subsidiary or orbital angular momentum Quantum number] (l)

The following informations are obtained from l .

1. It gives the shape of the orbital.
2. It gives the sub shell or sub level in which the electron is located.
3. It also gives the orbital angular momentum of the electron.

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

For example, when $n = 1$, value of l is only 0. For $n = 2$, the possible value of l can be 0 and 1. For $n = 3$, the possible l values are 0, 1 and 2.

$l = 0$ represents s orbital, $l = 1$ represents p orbital, $l = 2$ represents d orbital and $l = 3$ represents f orbital

3. Magnetic Quantum Number (m or m_l)

It gives information about the orientation of orbitals in space. For a given ' l ' value, there are $2l+1$ possible values for m and these values are given by $m = -l$ to 0 to $+l$

Thus for $l = 0$, $m_l = 0$. i.e. s sub shell contains only one orbital called s orbital.

For $l = 1$, $m_l = -1, 0$ and $+1$. i.e. p subshell contains three orbitals called p orbitals (p_x , p_y and p_z).

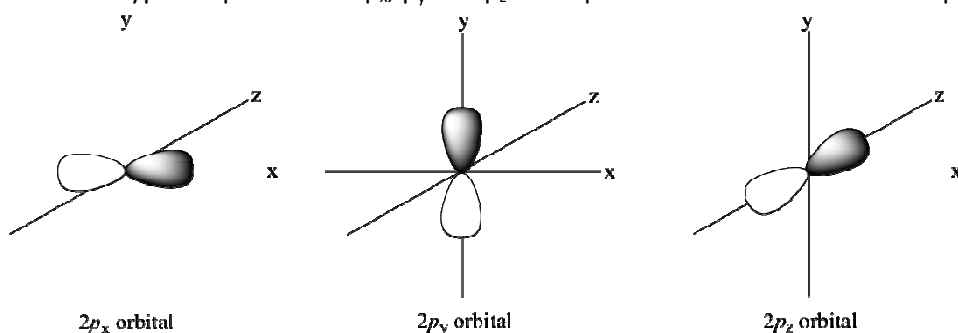
For $l = 2$, $m_l = -2, -1, 0, +1$ and $+2$. i.e. d subshell contains five orbitals called d orbitals (d_{xy} , d_{xz} , d_{yz} , $d_{x^2-y^2}$ and d_{z^2})

4. Spin Quantum Number (s or m_s)

It is the only experimental Quantum number and it gives the spin of electrons. The values for s may be $+\frac{1}{2}$ or $-\frac{1}{2}$. $+\frac{1}{2}$ represents clock-wise spin and $-\frac{1}{2}$ represents anticlock-wise spin.

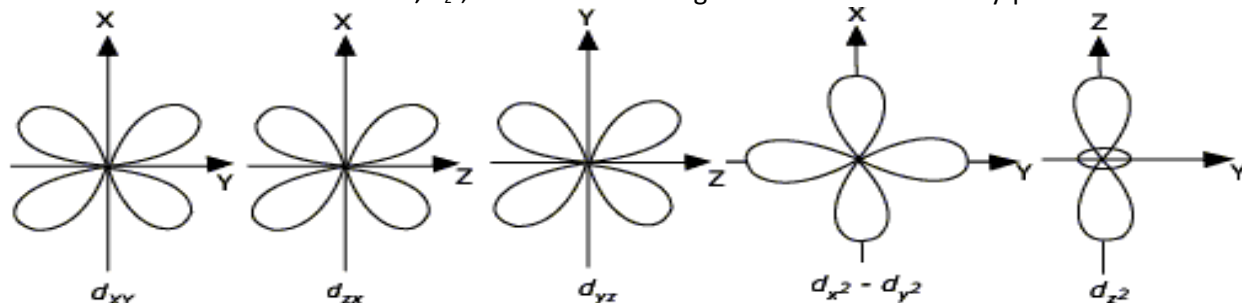
13. Draw the shapes of different p-orbitals.

There are 3 types of p-orbitals – p_x , p_y and p_z . Each p-orbitals has a dumb-bell shape.



14. Which are the different types of d-orbitals? Draw their shapes.

There are 5 types of d-orbitals. They are d_{xy} , d_{xz} , d_{yz} , $d_{x^2-y^2}$ and d_{z^2} . The shapes of the first four d-orbitals are double dumb-bell and that of the fifth one, d_{z^2} , is dumb-bell having a circular collar in the xy-plane.



15. Which are the different rules for filling electrons in orbitals? State them.

The filling of electrons into the orbitals of different atoms takes place according to the 3 rules - aufbau principle, Pauli's exclusion principle and the Hund's rule of maximum multiplicity.

1. Aufbau principle:

It states that **the orbitals are filled in order of their increasing energies.**



This rule has two sub rules:

- The various orbitals are filled in the increasing order of their $(n+l)$ value.
- If two orbitals have the same $(n+l)$ values, the orbital with the lower n value is filled first.

2. Pauli's Exclusion Principle

It states that ***no two electrons in an atom can have the same set of four quantum numbers***. i.e. an orbital can accommodate a maximum of only 2 electrons with opposite spin.

3. Hund's rule of maximum multiplicity

It states that ***electron pairing takes place only after partially filling all the degenerate orbitals***.

16. Which of the following set of quantum numbers are not allowed?

- $n = 3, l = 3, m = -3, s = +\frac{1}{2}$
- $n = 2, l = 1, m = 0, s = -\frac{1}{2}$
- $n = 1, l = 0, m = 0, s = +\frac{1}{2}$
- $n = 0, l = 0, m = 0, s = +\frac{1}{2}$

Ans: Here i and iv are not allowed. In i $n=3$ and $l=3$. The value of l is always less than n . i.e. if $n=3$, the possible l values are 0, 1 and 2.

In iv, $n=0$. It cannot be possible, because the values of n are 1, 2, 3, 4,

3. CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

1. Write a note on Mendeleev's periodic table.

Dimitri Mendeleev classified the elements in the increasing order of their atomic weights. He proposed a **periodic law** which states that **"The properties of elements are the periodic functions of their atomic weights."** That is, when elements are arranged in the increasing order of their atomic weights, their properties repeat after a regular interval. Mendeleev arranged elements in horizontal rows (periods) and vertical columns (groups). The elements with similar properties are arranged in the same group. He arranged elements mainly based on the similarities in the empirical formulae and the properties of the compounds formed by the elements.

2. What are the Merits of Mendeleev's periodic table

- It was the first comprehensive classification of elements.
- He corrected the wrong atomic weights of some elements and placed them in correct position in the periodic table.
- He left vacant places for undiscovered elements and predicted some of their properties.
- Elements with similar properties are placed in the same group.

3. What are the drawbacks of Mendeleev's periodic table?

- Elements with dissimilar properties are found in same group.
- He could not give an exact position for hydrogen.
- He could not give exact position for Lanthanoids and Actinoids and also for isotopes.
- Mendeleev's periodic table did not strictly obey the increasing order of atomic weights.

4. State Modern Periodic law. Who proposed this law?

Modern periodic law states that **"the physical and chemical properties of elements are the periodic functions of their atomic numbers"**. This law was proposed by Moseley.

5. Explain the different blocks in Modern periodic table.

There are 4 blocks in modern periodic table. They are s block, p block, d block and f block.

i. The s block elements

These are elements in which the last electron enters in the outer most s sub shell. They include elements of the groups 1 and 2. Their general outer electronic configuration is ns^1 or ns^2 . They are all reactive metals with low ionization enthalpies.

ii. The p block elements

These are elements in which the last electron enters in the outer most p sub shell. They include elements of the groups 13 to 18. They are also called **Representative elements**. Their general outer electronic configuration is $ns^2 np^{1 \text{ to } 6}$.

iii. The d block elements

These are elements in which the last electron enters in the penultimate d sub shell. They include elements of the groups 3 to 12. They are also called **Transition elements**, since they show a transition (change) from the most electropositive s block elements to the least electropositive p block elements. Their general outer electronic configuration is $(n-1)d^{1 \text{ to } 10} ns^{0 \text{ to } 2}$.



iv. The f block elements

These are elements in which the last electron enters in the anti-penultimate f sub shell. They include *lanthanides of 6th period and actinides of 7th period*. They are also called **Inner transition elements**. Their general outer electronic configuration is $(n-2)f^{1 \text{ to } 14}(n-1)d^{0 \text{ to } 1} ns^2$. They are all metals.

6. What are the characteristics of transition elements?

- They are all metals
- They form coloured compounds or ions
- They show variable oxidation states and valencies,
- They show paramagnetism and catalytic properties.

7. How does atomic radius vary along a group and period and why?

The atomic size decreases from left to right in a period. This is because in a period, the electrons are added to the same shell. Thus the number of shells remains same, but the effective nuclear charge increases. So the atomic radius decreases.

In a group, the atomic radius increases from top to bottom. This is because of the increase in no. of shells and shielding effect.

8. What are isoelectronic species? Give examples.

Atoms and ions having the same number of electrons are called isoelectronic species. E.g. O^{2-} , F^- , Ne, Na^+ , Mg^{2+} etc. (All these contain 10 electrons). Among isoelectronic species, *the cation with greater positive charge will have the smaller radius. The anion with greater negative charge will have the larger radius.*

9. Define ionisation enthalpy.

It is defined as the energy required to remove an electron from the outer most shell of an isolated gaseous atom in its ground state. Its unit is kJ/mol or J/mol.

The important factors which affect ionisation enthalpy are atomic size, nuclear charge, shielding effect etc.

10. How does ionisation enthalpy vary along a period and group? Justify your answer.

Along a *period*, ionisation enthalpy *increases* from left to right. This is because of the *decrease in atomic radius and increase in nuclear charge*. Thus alkali metals have the least $\Delta_i H$ and noble gases have the most.

Down a *group*, $\Delta_i H$ *decreases* due to *increase in atomic radius and shielding effect*. Thus among alkali metals, lithium has the least $\Delta_i H$ and francium has the most.

11. The first ionisation enthalpy of Boron is slightly less than that of Beryllium. Why?

This is because of the completely filled orbitals in Be ($1s^2 2s^2$).

12. The first ionisation enthalpy of N is greater than that of Oxygen. Why?

This is because N has half filled electronic configuration ($1s^2 2s^2 2p^3$), which is more stable and so more energy is required to remove an electron.

13. Define electron gain enthalpy ($\Delta_{eg} H$).

It is the heat change (enthalpy change) when an electron is added to the outer most shell of an isolated gaseous atom. Its unit is kJ/mol. It may be positive or negative depending on the nature of the element.

14. How does electron gain enthalpy vary along a period and group? Justify.

From left to right across a *period*, electron gain enthalpy become *more negative*. This is *because of decrease in atomic radius and increase in nuclear charge*. Down a *group*, electron gain enthalpy becomes less negative. This is *due to increase in atomic radius and shielding effect*.

15. Arrange halogens in the decreasing order of –ve electron gain enthalpy?

Among halogens, the negative $\Delta_{eg} H$ decreases as follows. $Cl > F > Br > I$

16. Chlorine has higher negative electron gain enthalpy than fluorine. Why?

Or, Electron gain enthalpy of fluorine is less negative than chlorine. Why?

This is because, when an electron is added to F, it enters into the smaller 2^{nd} shell. Due to the smaller size, the electron suffers more repulsion from the other electrons. But for Cl, the incoming electron goes to the larger 3^{rd} shell. So the electronic repulsion is low and hence Cl adds electron more easily than F.

17. Electron gain enthalpy values of noble gases are zero. Why?

This is because of their completely filled (stable) electronic configuration.

**18. Define electronegativity. How does it vary along a group and period?**

Electronegativity of an atom in a compound is the ability of the atom to attract shared pair of electron of electrons. It increases across a period and decreases along a group. So in modern periodic table, F has the maximum electronegativity and Fr has the minimum electronegativity.

4. CHEMICAL BONDING AND MOLECULAR STRUCTURE

1. State octet rule. What are its limitations?

This rule was proposed by Lewis and Kossel. According to this rule, *atoms undergo chemical reaction in order to attain octet of electrons in the valence shell. Or, atoms containing 8 electrons in their valence shell are stable.*

Limitations of Octet rule

- 1) It could not explain the stability of compounds containing less than 8 electrons around the central atom. E.g. LiCl, BeH₂, BCl₃ etc.
- 2) It could not explain the stability of molecules containing odd number of electrons like NO, NO₂ etc.
- 3) It could not explain the stability of molecules containing more than 8 electrons around the central atom (i.e. expanded octet). E.g. PF₅, SF₆, H₂SO₄, IF₇ etc.
- 4) Octet rule is based upon the chemical inertness of noble gases. But some noble gases like xenon and krypton form compounds with F and O.
- 5) This theory does not account for the shape of molecules.
- 6) It does not explain the relative stability of the molecules.

2. Define Dipole moment. Give its unit.

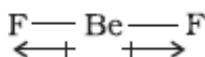
The polarity of a molecule is expressed in terms of *dipole moment* (μ). It is defined as *the product of the magnitude of charge at one end (Q) and the distance between the charges (r).*

Mathematically, $\mu = Q \times r$.

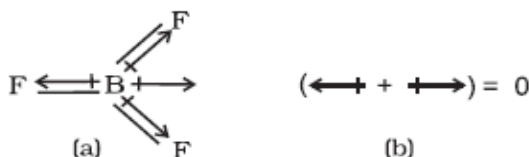
The unit of dipole moment is Coulomb metre (Cm). But it is usually expressed in the unit Debye (D).

3. The dipole moment of BeF₂ is zero. Why?

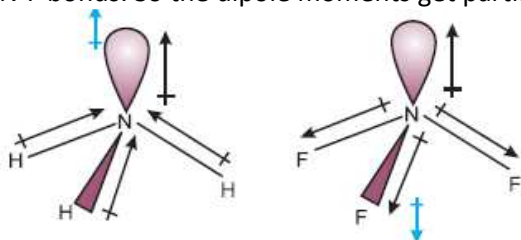
BeF₂, the net dipole moment is zero, since the two equal bond dipoles are in opposite directions and cancel each other.

**4. The dipole moment of BF₃ is zero. Why?**

In BF₃, the net dipole moment is zero. Here the resultant of any 2 bond dipoles is equal and opposite to the third.

**5. Ammonia (NH₃) has higher dipole moment than NF₃, even though F is more electronegative than H. why?**

This is because in the case of NH₃, the orbital dipole due to lone pair is in the same direction as the resultant dipole moment of the three N – H bonds. But in NF₃, the orbital dipole is in the opposite direction to the resultant dipole moment of the three N-F bonds. So the dipole moments get partially cancelled.



6. State Fajan's rule regarding the partial covalent character of an ionic bond.

The partial covalent character of ionic bonds was explained by Fajans in terms of the following rules:

- 1) The smaller the size of the cation and the larger the size of the anion, the greater the covalent character of an ionic bond.
- 2) The greater the charge on the cation, the greater the covalent character of the ionic bond.
- 3) For cations of the same size and charge, the ion with electronic configuration $(n-1)d^n ns^0$ is more polarising than the ion with a noble gas configuration $(ns^2 np^6)$.

7. What are the main postulates of valence shell electron pair repulsion (VSEPR) theory?

This theory was proposed by Sidgwick and Powell and later modified by Nyholm and Gillespie. The important postulates of this theory are:

- 1) The shape of the molecule depends on the number of valence shell electron pairs (VSEPRs) around the central atom.
- 2) The valence shell electron pairs repel each other.
- 3) In order to reduce the repulsion, the electron pairs stay at maximum distance.
- 4) Presence of lone pairs of electron causes distortion in the expected geometry of the molecule.
- 5) The repulsion between two lone pairs of electrons is different from those between two bond pairs or between a lone pair and bond pair. The repulsion decreases in the order lone pair - lone pair > lone pair - bond pair > bond pair - bond pair.
- 6) As the angle between the electron pairs increases, the repulsion decreases.

8. Explain the shape ammonia and water molecules on the basis of VSEPR theory.

In ammonia, the central atom N has 5 valence electrons ($\gamma N - 2, 5$). Among these electrons, three are used for the formation of bonds with hydrogen atoms and the remaining 2 electrons stay as lone pairs. So there are 4 VSEPs. Hence the expected shape of the molecule is tetrahedral. But due to the presence of lone pairs, the shape is distorted to triangular pyramid and the bond angle changes to 107° .

In water, the central atom O has 6 valence electrons ($\gamma O - 2, 6$). Two of them are used for the formation of bonds with hydrogen atoms and the remaining 4 electrons stay as lone pairs. So there are 4 VSEPs. Hence the expected shape of the molecule is tetrahedral. But due to the presence of 2 lone pairs, the shape is distorted bent or angular or inverted 'v' shape and the bond angle changes to 104.5° .

9. Define Hybridisation. What are the important characteristics of Hybridisation?

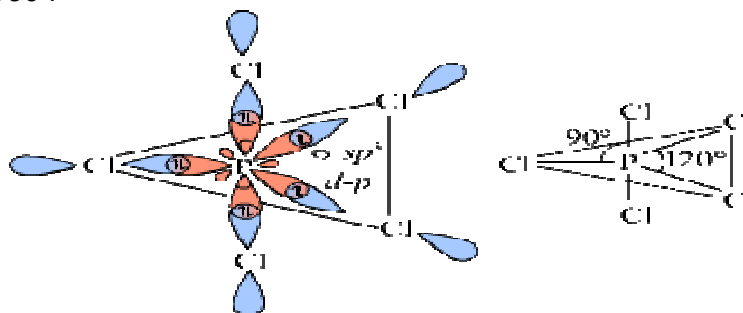
It is the process of inter mixing of atomic orbitals having slightly different energies to form new orbitals having equivalent energy and identical shape. The new orbitals formed are called hybrid orbitals.

The important characteristics of hybridisation are:

- 1) The number of hybrid orbitals formed is equal to the number of atomic orbitals undergo hybridization.
- 2) The hybrid orbitals are always equivalent in energy and identical in shape.
- 3) The hybrid orbitals are more effective in forming stable bonds than the pure atomic orbitals.
- 4) The hybrid orbitals are directed to some fixed positions in space. So the type of hybridization gives the shape of the molecule.

10. Explain the geometry of PCl_5 molecule and account for its high reactivity.

In PCl_5 , phosphorus atom is in sp^3d hybridisation. So the molecule has trigonal bipyramidal shape with bond angles 120° and 90° .

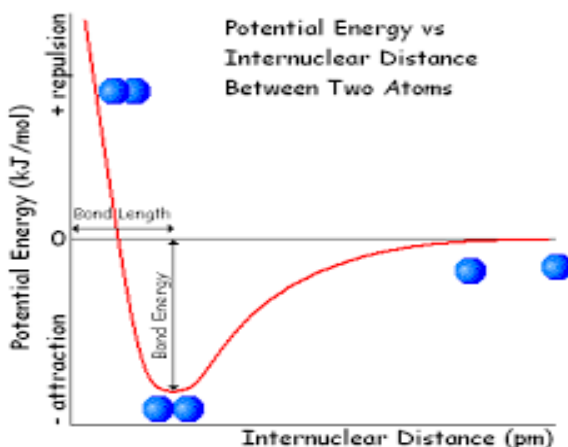


Trigonal bipyramidal structure of PCl_5 .

In PCl_5 , three P-Cl bonds lie in one plane, at an angle of 120° . These three bonds are called equatorial bonds. The other two P-Cl bonds lie one above and one below this plane. They are called axial bonds. The axial bond pairs suffer more

repulsion from the equatorial bond pairs. So the axial bond length is greater than the equatorial bond length. So PCl_5 is highly unstable and is very reactive.

11. **Draw the potential energy curve for the formation of a hydrogen molecule on the basis of inter nuclear distance of the hydrogen atoms.**



12. **Molecular orbitals are formed by the linear combination of atomic orbitals (LCAO). Give the salient features of molecular orbital theory.**

- In molecules, the electrons are present in some special type of orbitals called molecular orbitals (M.Os).
- The atomic orbitals (A.Os) of comparable energy and proper symmetry combine to form molecular orbitals.
- Atomic orbitals are monocentric, while molecular orbitals are polycentric.
- The number of molecular orbitals formed = the number of atomic orbitals combined. i.e. if 2 atomic orbitals combined, 2 molecular orbitals are formed. One is called bonding molecular orbital (BMO) and the other is called anti-bonding molecular orbitals (ABMO).
- The BMO has lower energy and greater stability than the corresponding ABMO.
- The molecular orbitals give the electron probability distribution around a group of nuclei.
- The molecular orbitals are filled according to 3 rules – Aufbau principle, Pauli's exclusion principle and Hund's rule.

13. **Define bond order. How is bond order related to bond length?**

It is defined as the half of the difference between the number of bonding electrons (N_b) and the number of anti-bonding electrons (N_a). i.e. Bond order (B.O) = $\frac{1}{2} [N_b - N_a]$

As the bond order increases, bond length decreases.

14. **He_2 does not exist. Why?**

For He_2 molecule, bond order = 0, the molecule cannot exist.

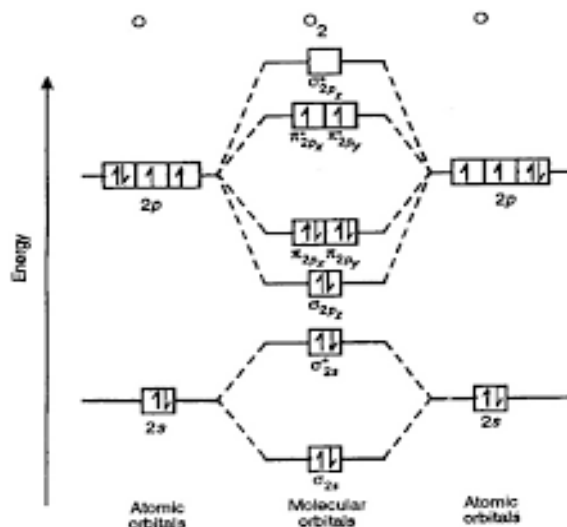
15. **Ne_2 does not exist. Why?**

For Ne_2 molecule bond order = 0, the molecule cannot exist.

16. **Draw the M.O diagram for oxygen molecule (O_2). Give its magnetic character.**

In O_2 there are 16 electrons.

The M.O configuration for O_2 is $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^* 2p_x^1 = \pi^* 2p_y^1$

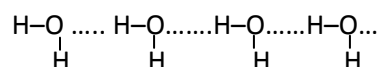


Here there are 2 unpaired electrons. So it is paramagnetic in nature.

17. Define hydrogen bonding with example. Explain the different types of H-bonding.

The weak attractive force between Hydrogen atom of one molecule and electronegative atom (like F, O or N) of the same or different molecule is termed as Hydrogen bond.

e.g. Hydrogen bonding in water



There are two types of H bonds- inter molecular H-bonding and intra molecular H-bonding.

- a) Inter molecular H bonding:** It is the H bond formed by H atom of one molecule and the electronegative atom of another molecule.
- b) Intra molecular H bonding:** It is the H bond formed between H atom and the electronegative atom of the same molecule.

18. Based on bond order compare the relative stability of O_2 and O_2^{2-} .

The M.O configuration of O_2 is $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^* 2p_x^1 = \pi^* 2p_y^1$

$$\text{B.O} = \frac{1}{2} [N_b - N_a] = \frac{1}{2} [10 - 6] = \frac{1}{2} \times 4 = 2$$

Since bond order is positive, it is stable.

The M.O configuration of O_2^{2-} is $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2 \pi^* 2p_x^2 = \pi^* 2p_y^2$

$$\text{B.O} = \frac{1}{2} [N_b - N_a] = \frac{1}{2} [10 - 7] = \frac{1}{2} \times 3 = 1.5$$

Since bond order is positive, it is stable.

5. STATES OF MATTER

1. State Boyle's law. Give its mathematical relationship. Draw the graphical representation.

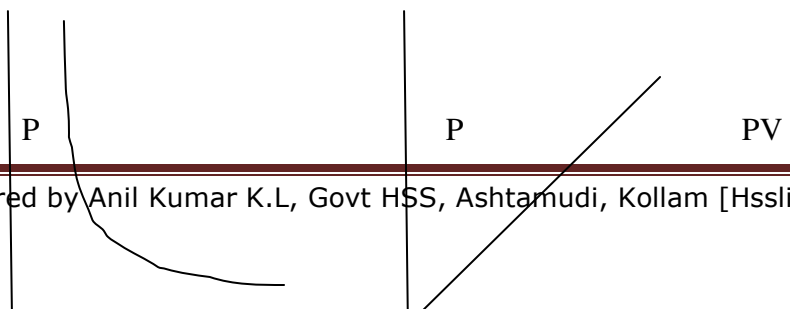
It states that at constant temperature, the volume of a fixed mass of gas is inversely proportional to its pressure.

Mathematically,

$$P \propto 1/V$$

$P = k \times 1/V$, where k is the proportionality constant.

Or, $PV = k$, a constant



$$PV = a \text{ constant}$$



V

1/V

P

2. Name the different types of van der Waal's forces.

Ans: London force, dipole-dipole interaction, dipole induced dipole interaction

3. State Charles's law. Give its mathematical form.

It states that at constant pressure, volume of a fixed mass of gas is directly proportional to its temperature.

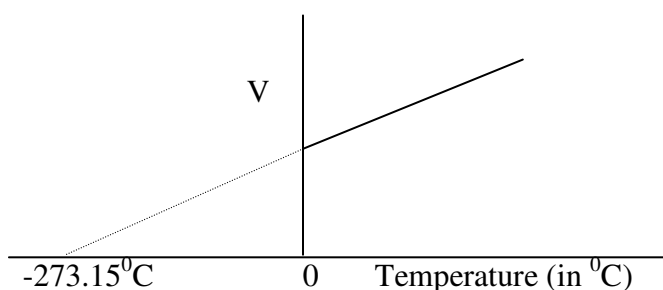
Mathematically, $V \propto T$

Or, $V = k \times T$

Or, $V/T = k$, a constant

4. What are isobars? Draw the isobar related to charle's law.

The graph obtained at constant pressure is called isobar.



5. What is absolute zero of temperature?

The lowest hypothetical or imaginary temperature at which gases are supposed to occupy zero volume is called **Absolute zero of temperature.**

6. State Avogadro's law.

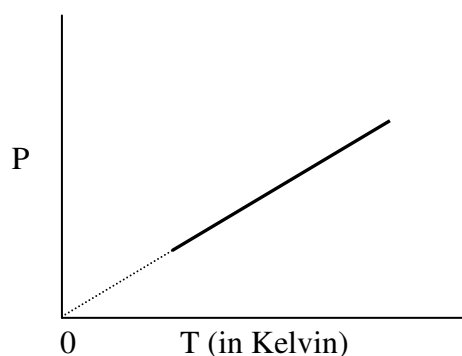
It states that equal volumes of all gases under the same conditions of temperature and pressure contain equal number of moles or molecules.

7. What are isotherm and isochore?

A graph obtained at constant temperature is called isotherm and a graph obtained at constant volume is called isochore.

8. Name the gas law which gives the relationship between the pressure and temperature of a fixed amount of gas at constant volume. Draw the graph to illustrate the above gas law.

Ans: Gay-Lussac's law.



9. Derive ideal gas equation.

The combination of the three gas laws (Boyle's law, Charles' and Avogadro law) give a single equation which is known as **ideal gas equation.**

According to Boyle's law: $V \propto 1/P$ (At constant T and n)

According to Charles' Law: $V \propto T$ (At constant p and n)

According to Avogadro Law: $V \propto n$ (At constant p and T)

On combining these three laws we get:

$$V \propto n \times T \times 1/P$$

Or, $V = R \times n \times T \times 1/P$ (where R is a constant called **universal gas constant**)

$$\text{Or, } \mathbf{PV = nRT} \dots\dots\dots (1)$$

This equation is known as **ideal gas equation**.

10. Write down the values of R in two different units.

1. L atm/K/mol	0.0821
2. L bar/K/mol	0.083
3. Pa m ³ /K/mol	8.314
4. J/K/mol	8.314

11. State Dalton's law of Partial pressures. Give its application.

This law states that at constant temperature, the total pressure exerted by a mixture of non-reacting gases is equal to the sum of the partial pressures of the component gases. Dalton's law can be used to calculate the pressure of a gas collected over water surface.

12. What are the postulates of Kinetic Molecular theory of gases?

The important postulates of this theory are:

- Every gas contains a large number of minute and elastic particles (atoms or molecules). The actual volume of the molecules is negligible compared to the volume of the gas.
- There is no force of attraction between the gas particles.
- The particles of a gas are in constant and random motion in straight line. During this motion they collide with each other and also with the walls of the container.
- The pressure of a gas is due to the wall collisions of the particles.
- All collisions are perfectly elastic. i.e. the total energy of particles before and after collisions remains the same.
- At any particular time, different particles of a gas have different speed and hence different kinetic energy.
- The average kinetic energy of gas molecules is directly proportional to absolute temperature.

13. Why do real gases deviate from ideal gas equation?

The deviation from real gases from ideal behaviour is due to two wrong assumptions of kinetic theory of gases. They are:

- The actual volume of the molecules is negligible compared to the volume of the gas.
- There is no force of attraction between the gas particles.

These two assumptions become wrong at *high pressure and low temperature*.

14. Write the expression for compressibility factor. What is its value for an ideal gas?

The compressibility factor (Z) = PV/nRT

For ideal gas $Z = 1$ at all temperature and pressures.

15. Write the van der Waal's equation of states. Explain the terms.

$$\text{van der Waal's equation is } (P + n^2a/V^2)(V - nb) = nRT$$

Where P is the pressure of the gas, V is the volume, R is the universal gas constant, T is the absolute temperature, n is the no. of moles, 'a' and 'b' are called van der Wall's constants.

16. Define Boyle point.

The temperature at which real gases obey ideal gas equation over an appreciable range of pressure is called Boyle temperature or Boyle point.

17. What is mean by normal boiling point and standard boiling point?

At 1 atm pressure the boiling point is called **normal boiling point**. If pressure is 1 bar then the boiling point is called **standard boiling point**. The normal boiling point of water is 100 °C (373 K) and its standard boiling point is 99.6 °C (372.6 K).

**18. At high altitudes a pressure cooker is used for cooking food. Why?**

At high altitudes (heights) atmospheric pressure is low. Therefore liquids at high altitudes boil at lower temperatures than at sea level. Since water boils at low temperature on hills, the pressure cooker is used for cooking food.

19. Liquid drops assume spherical in shape. Why?

Every liquid tries to reduce their energy by decreasing the surface area. For a given volume sphere has the minimum surface area. So liquid drops assume spherical shape.

20. Define surface tension. What is its SI unit?

It is defined as the force acting per unit length perpendicular to the line drawn on the surface of liquid. Its SI unit is N m

21. What are the factors affecting viscosity of a liquid?

Viscosity depends on temperature and molar mass of the substance. It decreases with temperature and increases with molar mass.

6. THERMODYNAMICS

1. What are extensive and intensive properties? Give examples.

Extensive properties are properties which depend on the amount of matter present in the system.

E.g.: Volume (v), length (l), breadth (b), height (h), internal energy (U), enthalpy (H), entropy (S), Gibb's energy (G), heat capacity etc.

Intensive properties are properties which are independent of the amount of matter present in the system.

E.g. : Temperature (T), Pressure (P), Volume (V), density, refractive index, molar heat capacity, viscosity, surface tension etc.

2. Define state function and path function.

A function or a property that depends only on the initial and final state of a system and not on the path followed is called a state function. E.g.: T, P, V, U, H, S, G etc.

Path functions are properties which depend on the path followed also. E.g. heat (q) and work (w)

3. State the first law of thermodynamics. Give its mathematical form.

It states that energy can neither be created nor be destroyed. Or, the total energy in the universe is always a constant. Mathematically $\Delta U = q + w$

4. Give the relationship between ΔH and ΔU .

ΔH is related to ΔU as $\Delta H = \Delta U + P\Delta V$

5. Define std. enthalpy of formation. Give the std. enthalpy of formation of O_2 molecule.

The enthalpy of formation in the standard state (298K T, 1 bar P and at stable state of aggregation) is called std. enthalpy of formation. By convention, standard enthalpy for formation, $\Delta_f H^\circ$, of an element in its reference state is taken as zero. So the std. enthalpy of formation of O_2 is zero.

6. State Hess's law of constant summation. Give any 2 applications.

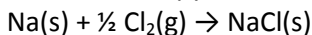
The law states that the total enthalpy change for a physical or chemical process is the same whether the reaction taking place in a single step or in several steps. Or, the total enthalpy change for a process is independent of the path followed. It is used for determining std. enthalpy of combustion and bond enthalpy.

**7. Define lattice enthalpy.**

The lattice enthalpy of an ionic compound is the enthalpy change when one mole of an ionic compound dissociates into gaseous ions.

8. Construct an enthalpy diagram for the determination of lattice enthalpy of sodium chloride.

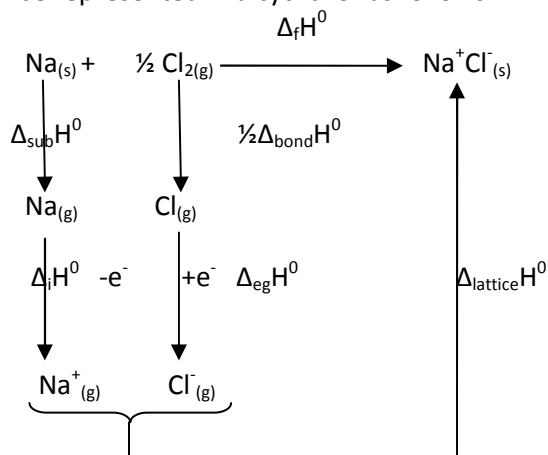
Lattice enthalpy of NaCl is determined by Born-Haber cycle.



This involves the following steps:

- Conversion of solid sodium atom to gaseous sodium atom. The energy change involved in this process is called sublimation energy.
 $\text{Na(s)} \rightarrow \text{Na(g)}; \Delta_{\text{sub}}H^\circ$
- Conversion of gaseous sodium atom to gaseous sodium ion. The energy change in this process is called ionisation enthalpy.
 $\text{Na(g)} \rightarrow \text{Na}^+(\text{g}); \Delta_iH^\circ$
- Conversion of gaseous chlorine molecule to gaseous chlorine atom. The energy change during this process is called bond dissociation enthalpy.
 $\text{Cl}_{2(\text{g})} \rightarrow 2\text{Cl}_{(\text{g})}; \Delta_{\text{bond}}H^\circ$
 $\frac{1}{2} \text{Cl}_{2(\text{g})} \rightarrow \text{Cl}_{(\text{g})}; \frac{1}{2} \Delta_{\text{bond}}H^\circ$
- Conversion of chlorine atom to chloride ion. The energy change involved in this process is called electron gain enthalpy ($\Delta_{\text{eg}}H^\circ$).
 $\text{Cl}_{(\text{g})} \rightarrow \text{Cl}^-(\text{g}); \Delta_{\text{eg}}H^\circ$
- Packing of $\text{Na}^+(\text{g})$ and $\text{Cl}^-(\text{g})$ to form NaCl(s). The energy change in this process is called lattice enthalpy ($\Delta_{\text{lattice}}H^\circ$)
 $\text{Na}^+(\text{g}) + \text{Cl}^-(\text{g}) \rightarrow \text{NaCl(s)}; \Delta_{\text{lattice}}H^\circ$

The different steps can be represented in a cyclic form as follows:



By applying Hess's law we can write: $\Delta_fH^\circ = \Delta_{\text{sub}}H^\circ + \Delta_iH^\circ + \frac{1}{2} \Delta_{\text{bond}}H^\circ + \Delta_{\text{eg}}H^\circ + \Delta_{\text{lattice}}H^\circ$

From this lattice enthalpy can be determined as:

$$\Delta_{\text{lattice}}H^\circ = \Delta_fH^\circ - [\Delta_{\text{sub}}H^\circ + \Delta_iH^\circ + \frac{1}{2} \Delta_{\text{bond}}H^\circ + \Delta_{\text{eg}}H^\circ]$$

Lattice enthalpy of NaCl is determined using Born-Haber cycle as follows:

$$\Delta_{\text{lattice}}H^\circ = \Delta_fH^\circ - [\Delta_{\text{sub}}H^\circ + \Delta_iH^\circ + \frac{1}{2} \Delta_{\text{bond}}H^\circ + \Delta_{\text{eg}}H^\circ]$$

9. What are spontaneous processes? Give examples.

It is a process that takes place without the help of any external agency. All natural processes are spontaneous. E.g. flow of water from high level to low level, flow of heat from hot body to cold body, inter mixing of gases, burning of fuels, melting of ice, evaporation of water etc.

10. What are the criteria for spontaneous process?

Decrease in energy and increase in disorderness.

11. Define entropy. What is its unit?

It is **a measure of degree of disorderness or randomness of a system**. The unit of entropy and entropy change is J/K/mol.

12. Define Gibb's energy. Give the criteria for spontaneity of a process in terms of free energy change (ΔG).

It is defined as **the maximum amount of available energy that can be converted to useful work**. For a process to be spontaneous, its ΔG should be negative.

13. What are the conditions for ΔG to be negative?

- i) If ΔH is negative and ΔS is +ve, ΔG is always -ve and the process is always spontaneous.
- ii) If both ΔH and ΔS are positive, ΔG will be -ve when $T\Delta S > \Delta H$. This is possible at high temperature.
- iii) If both ΔH and ΔS are negative, ΔG will be -ve when $T\Delta S < \Delta H$. This is possible at low temperature.

**7. EQUILIBRIUM**

1. Write the expression for equilibrium constant for the reaction i) $H_2 + I_2 \rightleftharpoons 2HI$ and ii) $N_2 + 3H_2 \rightleftharpoons 2NH_3$.

$$\begin{aligned} \text{i)} \quad K_c &= \frac{[HI]^2}{[H_2][I_2]} \\ \text{ii)} \quad K_c &= \frac{[NH_3]^2}{[N_2][H_2]^3} \end{aligned}$$

2. What is the relation between K_c and K_p ?

$$K_p = K_c (RT)^{\Delta n}$$

3. Give any 3 characteristics of equilibrium constant.

- i. Equilibrium constant is applicable only when the concentrations of the reactants and products have attained their equilibrium state.
- ii. The value of equilibrium constant is independent of the initial concentrations of reactants and products.
- iii. The value of equilibrium constant depends on temperature.

4. State Le-Chatlier Principle.

It states that whenever there is a change in concentration, pressure or temperature of a system at equilibrium, the system will try to readjust in such a way so as to cancel the effect of that change.

5. How do the effect of change in concentration, temperature and pressure affect the rate of chemical reaction?

An increase in concentration of reactants increases the rate of forward reaction (i.e. the equilibrium is shifted to the forward direction) and an increase in concentration of products increase the rate of backward reaction.

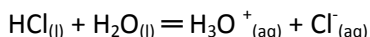
According to Le Chatlier's principle, increase in temperature favours endothermic process and decrease in temperature favours exothermic process.

When the pressure increases, the equilibrium is shifted to that direction in which no. of moles decreases.

6. Write the Arrhenius concept of acids and bases.

According to this concept acids are substances which give hydrogen ion (H^+) or hydronium ion (H_3O^+) in aqueous solution and bases are substances which give hydroxyl ion (OH^-) in aqueous solution.

e.g. HCl is an acid since it produces H_3O^+ in aqueous solution.

**7. Write Bronsted concept of acids and bases?**

According to this concept acids are proton (H^+) donors and bases are (H^+) acceptors.

For example in the reaction $NH_{3(l)} + H_2O_{(l)} = NH_4^+_{(aq)} + OH^-_{(aq)}$

Here NH_3 is a base since it accepts an H^+ ion to form NH_4^+ and H_2O is an acid since it donates an H^+ ion to form OH^- . In the reverse reaction, NH_4^+ is an acid and OH^- is a base.

8. What are conjugate acid – base pairs?

Acid- base pair that differs by only one proton is called a conjugate acid – base pair.

An acid formed from a base is called conjugate acid and a base formed from an acid is called conjugate base.

In general, Acid – $H^+ \rightarrow$ Conjugate base

Base + $H^+ \rightarrow$ Conjugate acid

9. State Lewis concept of acids and bases.

According to this concept acids are electron pair acceptors and bases are electron pair donors.

Substances which donate electron pair are called Lewis bases and substances which accept electron pair are called Lewis acids.

Example for Lewis acids are BF_3 , AlCl_3 , H^+ , Co^{3+} , Mg^{2+} etc.

Example for Lewis bases are NH_3 , H_2O , OH^- , Cl^- , Br^- etc.

10. Define ionisation constant of water or ionic product of water. What is its value at 298K?

It is defined as the product of the molar concentration of hydrogen ion (hydronium ion) and hydroxyl ion in water or in any aqueous solution. At 298K, $K_w = 10^{-14}$

11. Define p^{H} .

It is defined as the negative logarithm of the hydrogen ion or hydronium ion concentration in moles per litre (i.e. molarity).

i.e. $\text{p}^{\text{H}} = -\log[\text{H}^+]$

or $\text{p}^{\text{H}} = -\log[\text{H}_3\text{O}^+]$

12. Define common ion effect.

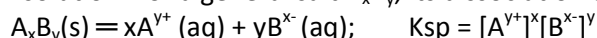
It is the suppression of the dissociation of a weak electrolyte by the addition of a strong electrolyte containing a common ion.

13. What is Buffer solution?

Solution which resists the change in pH on dilution or with the addition of small amount of acid or alkali is called Buffer solution. There are two types of buffer solutions – acidic buffer and basic buffer.

14. Define solubility product.

It is defined as the product of the molar concentration of ions of a sparingly soluble salt or in a saturated solution. For a general salt A_xB_y , its dissociation can be denoted as:



15. If the concentration of hydrogen ion in a soft drink is $3 \times 10^{-3} \text{ M}$, calculate its pH

$$[\text{H}_3\text{O}^+] = 3 \times 10^{-3} \text{ M}$$

$$\text{p}^{\text{H}} = -\log[\text{H}_3\text{O}^+]$$

$$= -\log(3 \times 10^{-3}) = 2.5229$$

16. Consider the equilibrium, $\text{AgCl}(\text{s}) \rightleftharpoons \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq})$. The solubility of AgCl is $1.06 \times 10^{-5} \text{ mol/L}$ at 298K. Find out its K_{sp} at this temperature.

Ans: $\text{AgCl}(\text{s}) \rightleftharpoons \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq})$. Let S be the solubility of AgCl.



$$K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-]$$

$$= \text{S} \times \text{S} = \text{S}^2$$

$$\text{Since } \text{S} = 1.06 \times 10^{-5}, K_{\text{sp}} = (1.06 \times 10^{-5})^2 = 1.1236 \times 10^{-10} \text{ M}^2$$

8. REDOX REACTIONS

1. Define the electronic concept of oxidation and reduction.

According to this concept *oxidation is the process of removal (losing) of electron and reduction is the process of addition (gaining) of electron.*

2. Define the oxidation number concept of oxidation and reduction.

According to this concept, *oxidation is the process of increase in the oxidation number of an element and reduction is the process of decrease in the oxidation number of an element.*

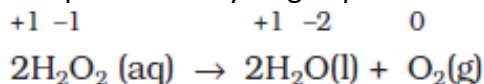
3. What are Stock notations? Using Stock notation, represent the following compounds: i) HAuCl_4 ii) MnO_2

Stock notations are some notations used to represent the oxidation number of a metal in a compound. According to this, the oxidation number is represented in Roman numeral in brackets after the symbol of the metal in the molecular formula. Stock notations for HAuCl_4 is HAu(III)Cl_4 and that for MnO_2 is Mn(IV)O_2 .

4. What is a disproportionation reaction? Give an example.

Disproportionation reaction is a special type of redox reaction in which an element in one oxidation state is simultaneously oxidised and reduced.

E.g. The decomposition of hydrogen peroxide.

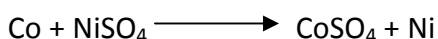


Here the oxygen of peroxide is in -1 state and it is converted to zero oxidation state in O_2 and -2 oxidation state in H_2O .

5. **Fluorine reacts with ice as given below: $\text{H}_2\text{O}(\text{s}) + \text{F}_2(\text{g}) \longrightarrow \text{HF}(\text{g}) + \text{HOF}(\text{g})$. Justify that this is a redox reaction.**

Ans: Here the oxidation number of Oxygen increases from -2 to $+1$ and that of F is decreased from zero to -1 . So this is a redox reaction.

6. **Write the redox reaction involved when metallic cobalt is placed in a nickel sulphate solution.**



7. **In the reaction $\text{Pb}(\text{s}) + \text{PbO}_2(\text{s}) + 2\text{H}_2\text{SO}_4(\text{aq}) \longrightarrow 2\text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$ Identify the following: i) Substance oxidised ii) Substance reduced iii) Oxidising agent iv) Reducing agent.**

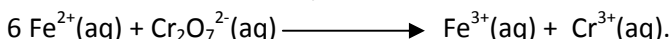
Ans: Here the Oxidation number of Pb increases from zero to $+2$. So Pb is oxidised and it is the reducing agent. Also the oxidation number of PbO_2 decreases from $+4$ to $+2$. So PbO_2 is reduced and hence it is the oxidising agent.

8. **Chemical reactions which involve oxidation and reduction are called redox reactions. The unbalanced equation in the ionic form of a redox reaction is shown below.**

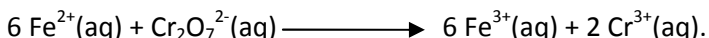
$\text{Fe}^{2+}(\text{aq}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) \xrightarrow{\text{acidic medium}} \text{Fe}^{3+}(\text{aq}) + \text{Cr}^{3+}(\text{aq})$. Balance this equation by oxidation number method.

Ans: Step 1: Assign the oxidation number of each element and identify the atoms which undergo change in oxidation number.

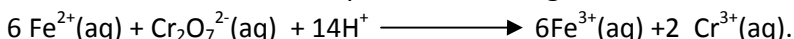
Step 2: Calculate the change in oxidation number per atom and equate them by multiplying with suitable coefficients. Here the oxidation number of iron is increased by 1 and that of Cr is decreased by 3. (since there are two Cr atoms, oxidation number decreased by 6). So multiply Fe^{2+} by 6 and $\text{Cr}_2\text{O}_7^{2-}$ by 1.



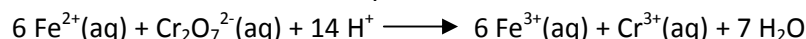
Step 3: Balance all the atoms except oxygen and hydrogen.



Step 4: Now equate the ionic charges on both sides of the equation. Since the reaction takes place in acidic medium, use H^+ ions to balance charges. Here the net ionic charge on LHS is $+10$ and that at RHS is $+24$. So add 14H^+ on LHS to equate ionic charge.



Step 5: Now make the numbers of hydrogen atoms in the expression on the two sides equal by adding water (H_2O) molecules to the reactants or products. Now, also check the number of oxygen atoms.



9. **What are the rules for assigning oxidation number of an element?**

- The oxidation number of all elements in the free or the uncombined state is zero. For e.g. oxidation number of H_2 , O_2 , Cl_2 , O_3 , P_4 , S_8 , Na, Mg, Al etc. is zero.
- For simple ions, the oxidation number is equal to the charge on the ion. Thus Na^+ ion has an oxidation number of $+1$, Mg^{2+} ion $+2$, Fe^{3+} ion $+3$, Cl^- ion -1 , O^{2-} ion -2 and so on.

- c) All alkali metals have oxidation number of $+1$ and all alkaline earth metals have an oxidation number of $+2$. Aluminium shows an oxidation number of $+3$ in all of its compounds.
- d) The common oxidation number of oxygen is -2 . But in peroxides (e.g., H_2O_2 , Na_2O_2), oxidation number of oxygen is -1 and in superoxides (e.g., KO_2 , RbO_2), it is $-\frac{1}{2}$. In oxygen difluoride (OF_2) and dioxygen difluoride (O_2F_2), the oxygen is assigned an oxidation number of $+2$ and $+1$ respectively.
- e) The common oxidation number of hydrogen is $+1$. But in hydrides, H shows an oxidation number of -1 .
- f) The common oxidation number of halogens is -1 . Fluorine shows only -1 oxidation number in all of its compounds. But other halogens show positive oxidation numbers also in their oxides and oxoacids.
- g) The algebraic sum of the oxidation number of all the atoms in a compound is zero.
- h) In polyatomic ion, the sum of the oxidation numbers all the atoms is equal to the charge on the ion.



9. HYDROGEN

1. Justify the position of hydrogen in the periodic table.

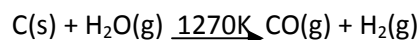
Hydrogen shows resemblance with both Alkali metals of the first group and halogens of the 17th group. Like alkali metals it has one electron in the outer most shell and forms unipositive ions. Like halogens, it requires only one electron to complete the valence shell configuration. So it gains one electron to form uninegative ion. At the same time it shows some differences from alkali metals and halogens. So it is placed separately in the periodic table.

2. What is water gas or syn gas?

A mixture of CO (carbon monoxide) and H_2 .

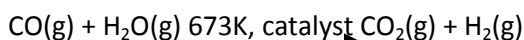
3. What is 'coal gasification'?

The process of producing 'syngas' from coal is called '**coal gasification**'.



4. How is dihydrogen produced by 'water gas shift reaction'?

The production of dihydrogen can be increased by reacting carbon monoxide of syngas mixtures with steam in the presence of iron chromate as catalyst. This is called **water-gas shift reaction**.

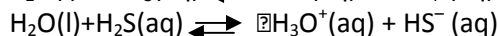
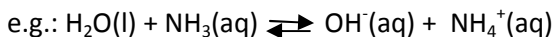


5. Write any two uses of dihydrogen.

- a) It is used as a rocket fuel in space research.
- b) It is used in fuel cells for generating electrical
- c) It is used in the manufacture of vanaspati fat by the hydrogenation of vegetable oils.

6. Water is an amphoteric substance. Justify.

Water can act both as acid and base. So it is an amphoteric substance.



In the first e.g. water acts as an acid and in the second it acts as a base.

7. Which are the different types of hydrides? Give one example for each.

There are three types of hydrides – ionic hydride (e.g. NaH), covalent hydride (e.g. H_2O) and metallic hydride (e.g. $\text{VH}_{0.56}$).

8. What is mean by hardness of water? What is the reason for hardness of water?

Water which does not easily form lather with soap is called hard water. It is due to the presence of calcium and magnesium salts in the form chlorides, sulphates and bicarbonates.

9. Hard water is harmful for boilers. Why?

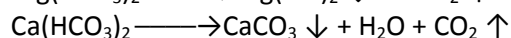
On boiling hard water, the soluble $\text{Mg}(\text{HCO}_3)_2$ is converted into insoluble $\text{Mg}(\text{OH})_2$ and $\text{Ca}(\text{HCO}_3)_2$ is changed to insoluble CaCO_3 . These salts are deposited in the boilers in the form of scales. This reduces the efficiency of boilers.



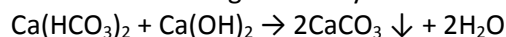
10. What is mean by temporary hardness? Explain the different methods used for its removal?

Hardness which can be removed by simple boiling is called temporary hardness. It is due to the presence of bicarbonate of calcium and magnesium. The following methods are used to remove temporary hardness.

1. **Boiling:** During boiling, the soluble $\text{Mg}(\text{HCO}_3)_2$ is converted into insoluble $\text{Mg}(\text{OH})_2$ and $\text{Ca}(\text{HCO}_3)_2$ is changed to insoluble CaCO_3 , which can be removed by filtration.



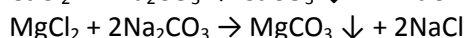
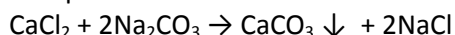
2. **Clark's method:** In this method calculated amount of lime is added to hard water. It precipitates out calcium carbonate and magnesium hydroxide which can be filtered off.



11. What is mean by permanent hardness? Explain the different methods used for its removal?

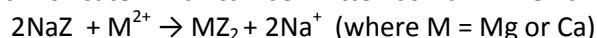
Hardness which cannot be removed by boiling is called Permanent hardness. It is due to the presence of soluble chlorides and sulphates of calcium and magnesium in water. It can be removed by the following methods:

1. **Treatment with washing soda** (Sodium carbonate): Washing soda reacts with soluble calcium and magnesium chlorides and sulphates in hard water to form insoluble carbonates.

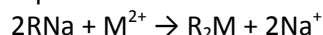


2. **Calgon's method:** Sodium hexametaphosphate ($\text{Na}_6\text{P}_6\text{O}_{18}$) is commercially called 'calgon'. When it is added to hard water, the Ca and Mg ions in hard water are replaced by Na^+ ions.

3. **Ion-exchange method:** This method is also called zeolite/permutit process. Zeolite /permutit is hydrated sodium aluminium silicate which can be written as NaZ . When this is added to hard water, exchange reactions take place.



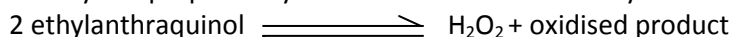
4. **Synthetic resins method:** Cation exchange resins contain large organic molecule with $-\text{SO}_3\text{H}$ group and are water insoluble. Ion exchange resin (RSO_3H) is changed to RNa by treating it with NaCl . The resin exchanges Na^+ ions with Ca^{2+} and Mg^{2+} ions present in hard water and make the water soft.



This method is more suitable to get pure demineralised water.

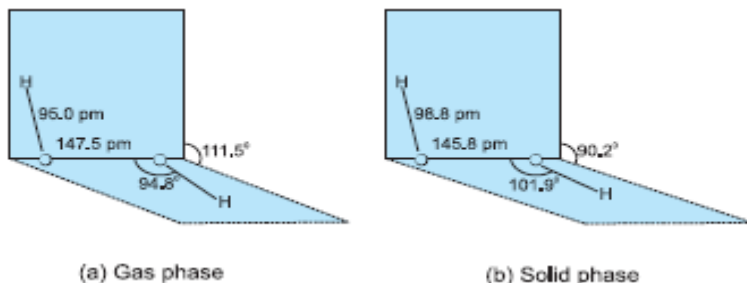
12. Give the industrial preparation of Hydrogen peroxide.

Industrially it is prepared by the auto-oxidation of 2-alkylanthraquinols.



13. Give the structure of Hydrogen Peroxide (H_2O_2).

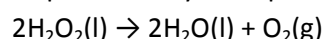
Hydrogen peroxide has a non-planar structure (open book-like structure).



14. Explain why hydrogen peroxide is not stored in glass vessels.

OR, Hydrogen peroxide is stored in plastic vessels in dark. Why?

Ans: H_2O_2 decomposes slowly on exposure to light.



In the presence of traces of alkali (present in glass containers), the above reaction is catalysed. So it is stored in wax-lined glass or plastic vessels in dark.



15. What is heavy water? Give one of its uses.

Deuterium oxide (D_2O) is called heavy water. It is used as a moderator in nuclear power plant.

10. THE S-BLOCK ELEMENTS

1. Alkali metals and their salts give characteristic colour to non-luminous flame. Why?

This is because the heat from the flame excites the outer most orbital electron to a higher energy level. When this electron comes back to the ground level, they emit the radiation in the visible region.

2. How do alkali metals react with air?

Alkali metals react with air to form oxides, peroxides and super oxides. Li forms only monoxide, sodium forms monoxide and peroxide and other alkali metals form monoxide, peroxide and super oxide.

3. Solutions of alkali metals in liquid ammonia are blue in colour. Why?

The alkali metals dissolve in liquid ammonia to give deep blue solutions. The blue colour of the solution is due to the formation of ammoniated electron.

4. Give any two anomalous properties of Lithium.

- a) Li is the least reactive but the strongest reducing agent among all the alkali metals.
- b) It forms only monoxide with oxygen.

5. What is mean by diagonal relationship? Give any two similarities in properties shown by Lithium and Magnesium.

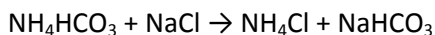
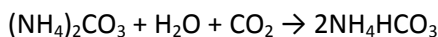
The similarity in properties shown by diagonally placed elements of second and third periods in modern periodic table is called diagonal relationship.

Li shows the following similarities in properties with Mg of the second group.

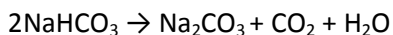
- a) Both react slowly with water.
- b) They do not form superoxides.

6. Name the process used for the industrial preparation of sodium carbonate. Explain the process.

Sodium carbonate is manufactured by Solvay process (Ammonia-Soda Process). In this process, CO_2 is passed through a concentrated solution of NaCl saturated with ammonia. Ammonium carbonate first formed then converted to ammonium bicarbonate and finally reacts with NaCl to form $NaHCO_3$.



Sodium bicarbonate crystals are separated and heated to get sodium carbonate.



7. Give any two uses of sodium carbonate.

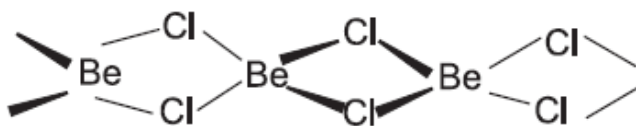
- a) It is used in water softening, laundering and cleaning.
- b) It is used as a laboratory reagent.

8. What are the uses of sodium hydroxide (NaOH)?

- a) in petroleum refining.
- b) in the purification of bauxite
- c) as a laboratory reagent.

9. Give the structure of beryllium chloride ($BeCl_2$).

It has a chain structure in the solid state.



10. Give any two similarities in properties (diagonal relationship) shown by Beryllium and aluminium.

- Like Al, Beryllium is not readily attacked by acids.
- The chlorides of both the elements have bridged structure in vapour phase.

11. What is quick lime and slaked lime? Give their uses.

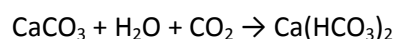
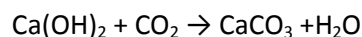
Quick lime is calcium oxide (CaO) and slaked lime is calcium hydroxide [Ca(OH)₂].

Quick lime is used in the manufacture of Na₂CO₃ from caustic soda and for the purification of sugar.

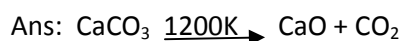
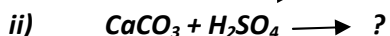
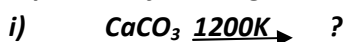
Slaked lime is used in white washing, in glass making, in tanning, for the preparation of bleaching powder and for purification of sugar.

12. What happens when CO₂ is passed through lime water?

When CO₂ is passed through lime water, it turns milky due to the formation of CaCO₃. On passing CO₂ continuously, the solution becomes clear due to the formation of soluble calcium bicarbonate [Ca(HCO₃)₂]



13. Complete the following reactions:



14. What is plaster of paris? Give its preparation and uses?

It is a hemihydrate of calcium sulphate. It is obtained when gypsum [CaSO₄·2H₂O] is heated to 393 K.



It is used in the building industry, in hospitals for setting fracture of bones and for making statues.

15. What is dead burnt plaster?

When plaster of paris (or gypsum) is heated above 393 K, it becomes anhydrous calcium sulphate (CaSO₄).

This is known as 'dead burnt plaster'.

16. Plaster of paris is used for moulding statues. Why?

On mixing with water plaster of paris forms a plastic mass that gets into a hard solid in 5 to 15 minutes.

During this process its volume increases. So it is used for the preparation of statues.

17. What are the major roles of these Na⁺ and K⁺ ions in our body?

Na⁺ ions participate in the transmission of nerve signals, in regulating the flow of water across the cell membranes and in the transport of sugars and aminoacids.

K⁺ ions activate many enzymes, participate in the oxidation of glucose to ATP and helps in the transmission of nerve signals along with sodium ions.

18. Give the name of the metal present in chlorophyll.

Ans: Magnesium

19. What is the use of milk of magnesia?

Ans: It is used as an antacid.

20. What is Baking soda? Give its uses.

Sodium bicarbonate (NaHCO₃) is called baking soda. It is used as a mild antiseptic and in fire extinguisher.

11. The p-Block Elements

1. In group 13 elements, atomic radius increases down the group. But atomic radius of gallium is less than that of aluminium. Why?

This is due to the presence of completely filled d-orbitals in Ga. Due to this the shielding effect is low. Thus atomic radius of gallium is less than that of aluminium.

2. How can you explain the higher stability of BCl_3 as compared to TlCl_3 ?

In BCl_3 , B is in $+3$ oxidation state. Since boron is a small atom and it contains 3 electrons in the outer most shell, it can easily show $+3$ oxidation state. Down the group $+1$ oxidation state becomes more common due to inert pair effect. So Tl mainly shows $+1$ oxidation state.

3. What is inert pair effect? What is its consequence?

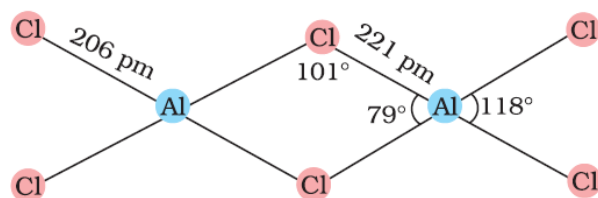
It is the reluctance of s-electrons to participate in chemical binding. It is commonly seen in the elements of groups 13, 14 & 15. In these groups as the atomic size increases (i.e. down the group), the maximum oxidation state is two less than the group oxidation states.

4. PbCl_2 is more stable than PbCl_4 . Why?

This is due to inert pair effect.

5. Give the structure of aluminium chloride (AlCl_3)?

In AlCl_3 , there are only six electrons around the central atom Al. In order to attain stability, it exists as dimer $[\text{AlCl}_3]_2$.



6. What is Borax? What is the action of heat on it?

It is Sodium tetra boratedecahydrate with formula $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$. 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.

7. What is Borax bead test?

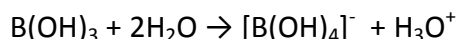
On heating Borax with transition metals, coloured beads are formed. So borax bead test can be used to identify transition metals.

8. A aqueous solution of borax is basic in nature. Justify.

Borax dissolves in water to give NaOH and orthoboric acid. Since NaOH is a strong alkali and orthoboric acid is weak acid, the solution is basic in nature.

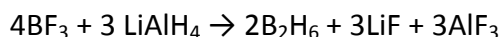
9. Ortho boric acid is monobasic even though it contains three hydrogen atoms. Why?

Orthoboric acid is a weak monobasic non-protic acid. It acts as a Lewis acid by accepting electrons from a hydroxyl ion.

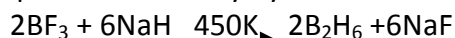


10. Give the preparation of diborane.

Diborane is the simplest boron hydride. It is prepared by treating BF_3 with Lithium aluminium hydride (LiAlH_4) in ether.

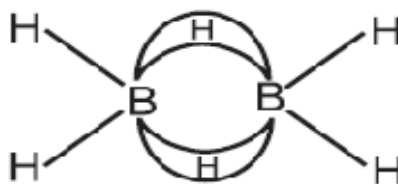
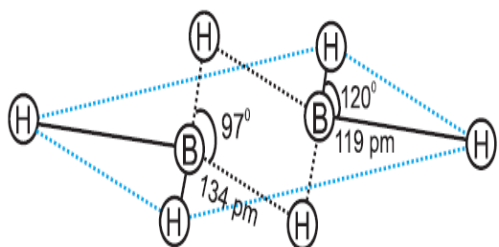


Diborane is prepared industrially by the reaction of BF_3 with sodium hydride.



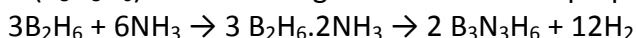
11. Give the structure of Diborane.

In diborane, each boron atom is in sp^3 hybridisation. The two boron atoms and 4 hydrogen atoms lie in one plane. These four H atoms are called terminal hydrogen. The other two hydrogen atoms lie one above and one below this plane. These H atoms are called bridging hydrogen atoms. The four terminal B-H bonds are regular two centre-two electron bonds while the two bridge (B-H-B) bonds are three centre- two electron ($3c-2e$) bonds. Thus diborane is an electron deficient compound.



12. What is inorganic benzene? Give its preparation?

Borazine ($B_3N_3H_6$) is called inorganic benzene. It is prepared by the reaction of ammonia with diborane



Its structure is similar to benzene with alternate BH and NH groups. So it is called inorganic benzene.

13. CCl_4 cannot be hydrolysed. Why?

In CCl_4 , there is no vacant d-orbitals in Carbon. Thus it cannot accommodate the lone pairs of electrons donated by oxygen atoms. So it cannot be hydrolysed.

14. Name the allotropes of carbon?

C exists in crystalline and amorphous allotropes. The important crystalline allotropes of C are diamond, graphite and fullerene. Amorphous forms of carbon are charcoal, coke and carbon black.

15. Diamond is hard and a non-conductor of electricity but graphite is soft and good conductor. Explain.

In diamond, each carbon atom undergoes sp^3 hybridisation and linked to four other carbon atoms in a tetrahedral manner. So it has a rigid three dimensional network of carbon atoms. Therefore, diamond is hard. Due to the absence of free electrons, it is a non-conductor of electricity.

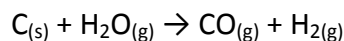
In Graphite each carbon atom is in sp^2 hybridisation and it has a hexagonal layered structure. Different layers are held by weak van der Waals forces. So each layer can slide over one another. Therefore graphite is soft. Due to the presence of free electrons, it is a good conductor of electricity.

16. Write a short note on fullerene.

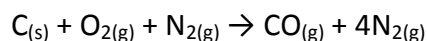
Fullerenes are one of the allotropic forms of carbon. The most commonly known fullerene is C_{60} , which is known as Buckminster fullerene. It contains twenty six- membered rings and twelve five membered rings. A six membered ring is fused with six or five membered rings but a five membered ring can only fuse with six membered rings. All the carbon atoms are equal and they undergo sp^2 hybridisation. Each carbon atom forms three sigma bonds with other three carbon atoms. The remaining electron at each carbon is delocalised in molecular orbitals. This gives an aromatic character to the molecule.

17. What are water gas and producer gas. Give any one of their use?

A mixture of carbon monoxide and hydrogen is called water gas or synthesis gas. It is prepared by passing steam over red hot coke at about 1273K.



A mixture of carbon monoxide and nitrogen is called producer gas. It is prepared by passing air over red hot coke at about 1273K.



Both these gases are used as industrial fuels.

18. Suggest a reason for the toxicity of CO?

In CO, there are one sigma and two pi bonds between carbon and oxygen ($:C \equiv O:$). Because of the presence of lone pair on carbon, CO molecule acts as an electron donor and reacts with metals. So it can easily

form a complex with haemoglobin (carboxy haemoglobin). It is about 300 times stabler than oxyhaemoglobin complex. This prevents haemoglobin in RBC from carrying oxygen round the body.

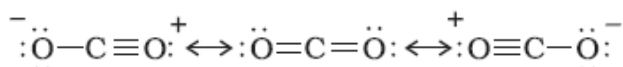


19. What is dry ice? Give its use?

Dry ice is solid carbon dioxide. It is obtained by cooling CO₂ under pressure. It sublimes directly to the vapour phase without liquefying. It does not wet the surface on which it sublimes. So it is called dry ice. It is used as a refrigerant for ice-cream and frozen food.

20. CO₂ is a gas, while SiO₂ is a solid. Account for this?

In CO₂ molecule, C atom undergoes sp hybridisation. So it has a linear shape. It exists as discrete (separate) molecules and there is only a weak attractive force between different CO₂ molecules. The resonance structures of CO₂ are:



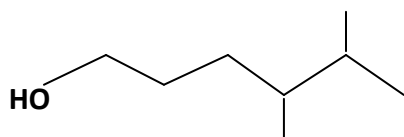
But in silica (SiO₂), each silicon atom undergoes sp³ hybridisation. Here each Si atom is tetrahedrally surrounded by 4 oxygen atoms. So it has a three dimensional network structure and hence it is a solid.

21. What are silicones? Give their uses?

They are a group of organosilicon polymers, which have (-R₂SiO-) as a repeating unit. They are used as sealant, greases, electrical insulators and for water proofing of fabrics. Being biocompatible they are also used in surgical and cosmetic plants.

12. ORGANIC CHEMISTRY – SOME BASIC PRINCIPLES AND TECHNIQUES

1. The bond-line formula of a compound is given below.

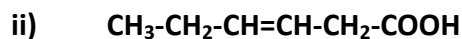
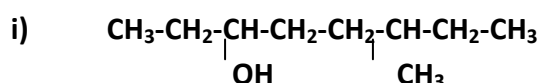


Write its condensed formula and give the IUPAC name.

Ans: Condensed formula is HO-CH₂-CH₂-CH₂-CH(CH₃)-CH(CH₃)₂

IUPAC Name is 4,5-Dimethylhexan-1-ol

2. Give the IUPAC names of the following compounds:



Ans: i) 6-Methyloctan-3-ol ii) Hex-3-en-1-oic acid

3. Draw the structures of the following compounds.

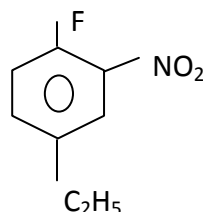
i) 2,3-Dibromo-1-phenylpentane

ii) 4-Ethyl-1-fluoro-2-nitrobenzene

Ans: i) $\text{CH}_2\text{--CH(Br)--CH(Br)--CH}_2\text{--CH}_3$



ii)



4. Write the complete, condensed and bond line structural formulae of 2-bromobutane.

Ans: Complete structural formula: $\begin{array}{ccccccc} & \text{H} & \text{H} & \text{H} & \text{H} \\ & | & | & | & | \\ \text{H} & \text{--C--} & \text{C--} & \text{C--} & \text{C--H} \\ & | & | & | & | \\ & \text{H} & \text{H} & \text{H} & \text{H} \end{array}$



Condensed structural formula: $\text{CH}_3\text{-CHBr-CH}_2\text{-CH}_3$



5. Write the structures of the following organic compounds.

a) 2,5,6 – Trimethyloctane

b) Hexane-2,4-dione

c) 5-oxohexanoic acid

Ans:

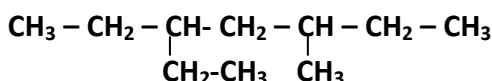
a) $\text{CH}_3\text{-CH(CH}_3\text{)-CH}_2\text{-CH}_2\text{-CH(CH}_3\text{)-CH(CH}_3\text{)-CH}_2\text{-CH}_3$

b) $\text{CH}_3\text{-CO-CH}_2\text{-CO-CH}_2\text{-CH}_3$

c) $\text{CH}_3\text{-CO-CH}_2\text{-CH}_2\text{-CH}_2\text{-COOH}$

6. The IUPAC names of alkanes are based on their chain structure.

a) Give the IUPAC name of



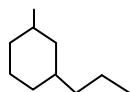
b) Represent 1-Methyl-3-propylcyclohexane using bond line notation.

c) What is the type of hybridization of C in CH_3^+ ? Also predict its shape.

d) Name the type of bond fission resulting in the formation of free radicals?

Ans: a) 3-Ethyl-5-methylhexane

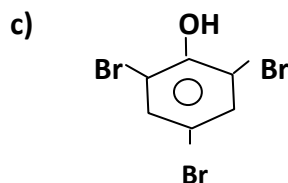
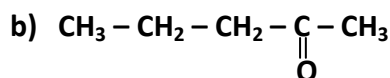
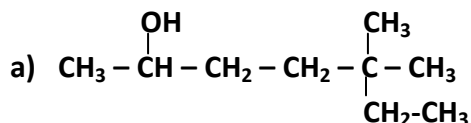
b)



c) Hybridisation: sp^2 shape: planar triangular

d) Homolysis or homolytic bond fission

7. Give the IUPAC names of the following:




Ans: a) 5,5-Dimethyl-2-heptanol

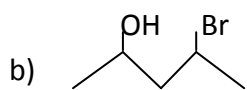
b) 2-pentanone

c) 2,4,6-trinitrophenol

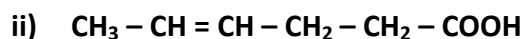
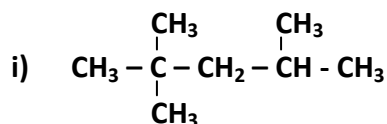
8. The bond line representation of cyclopropane  . Write the bond line structures of

- a) Cyclohexane
 b) $\text{CH}_3 - \text{CH}(\text{OH}) - \text{CH}_2 - \text{CHBr} - \text{CH}_3$

Ans: a) 



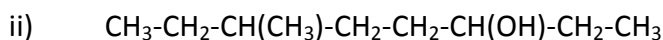
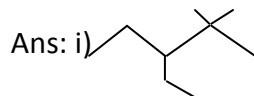
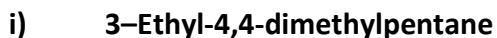
9. Write the IUPAC name of the following:



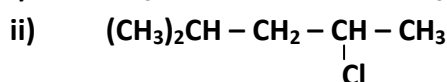
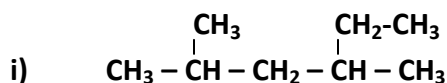
Ans: i) 2,2,4-trimethylpentane

ii) Hex-4-en-1-oic acid

10. Give the structures of the following compounds:



11. Write the IUPAC name of the following compounds :



Ans: i) 2,4-dimethylhexane

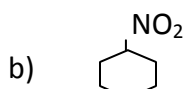
ii) 2-chloro-4-methylpentane

iii) Cyclohexenol

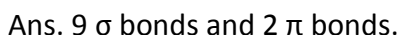
12. Draw the structure of the molecules represented by the IUPAC names :



Ans: a) $\text{CH}_2 = \text{CH} - \text{CH} - \text{CH}(\text{OH}) - \text{CH}_3$



13. How many 'σ' and 'π' bonds are present in the following compounds?


$$\text{CH}_3 - \text{CH}_2 - \underset{\text{OH}}{\underset{|}{\text{CH}}} - \text{CH}_2 - \text{CHO} \text{ and } \text{CH}_3 - \text{CH} = \text{CH} - \text{CH}_2 - \text{C} \equiv \text{C} - \text{CH}_3$$

Ans: 4-oxohexanoic acid

Ans:

Inductive effect (I effect): It is the shifting of sigma electrons through a carbon chain in presence of an atom or group of atom attached to a carbon chain. It is a permanent effect.



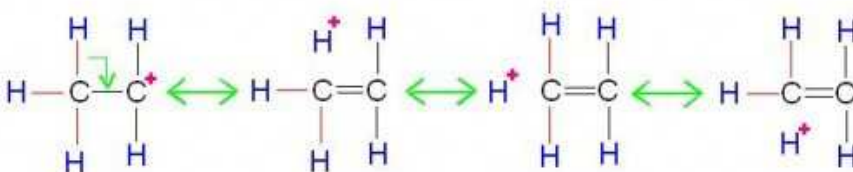
Electromeric effect (E effect): It is the shifting of π -electrons of a multiple bond in presence of an attacking reagent. It is a temporary effect.

Resonance Effect (R effect): It is defined as 'the polarity (charge) produced in the molecule by the interaction of two π -bonds or between a π -bond and lone pair of electrons present on an adjacent atom'.

Hyper conjugation: It is a permanent effect. Here the σ electrons of C—H bond of the alkyl group enter into partial conjugation with the unsaturated system or with the unshared p orbital.

24. Hyper conjugation is a general stabilizing interaction. Write the hyper-conjugative structures of $\text{CH}_3\text{-CH}_2^+$ (ethyl cation)

Ans:



25. Arrange the following carbocations in the increasing order of stability: $(\text{CH}_3)_2\text{CH}^+$, CH_3^+ , $(\text{CH}_3)_3\text{C}^+$, $\text{CH}_3\text{-CH}_2^+$

Ans: $\text{CH}_3^+ < \text{CH}_3\text{-CH}_2^+ < (\text{CH}_3)_2\text{CH}^+ < (\text{CH}_3)_3\text{C}^+$

26. Categorize the following functional groups into those having +R effect and -R effect: $-\text{NH}_2$, $-\text{NO}_2$, $-\text{COOH}$, $-\text{OH}$

Ans:

+R effect: $-\text{NH}_2$, $-\text{OH}$

-R effect: $-\text{NO}_2$, $-\text{COOH}$

27. Various methods for the purification of organic compounds are based on the nature of the compound and the impurity present in it. Explain the principle involved in the following methods of purification:

- a) Distillation b) Steam distillation

Ans:

a) **Distillation:** This method is used to separate (i) volatile liquids from non-volatile impurities and (ii) the liquids having sufficient difference in their boiling points. The principle of this method is that liquids having different boiling points vaporise at different temperatures. The vapours are cooled and the liquids so formed are collected separately.

b) **Steam distillation:** This technique is applied to separate substances which are steam volatile and are immiscible with water. In steam distillation, steam from a steam generator is passed through a heated flask containing the liquid to be distilled. The mixture of steam and the volatile organic compound is condensed and collected. The compound is later separated from water using a separating funnel.

Aniline – water mixture is separated by this method.

28. How is sodium fusion extract prepared? Using this, how will you detect the presence of Nitrogen, Sulphur and Halogen in an organic compound?

Ans: Nitrogen, sulphur and halogens present in an organic compound are detected by "**Lassaigne's test**". Here the organic compound is fused with metallic sodium in a fusion tube. It is then plunged into distilled water taken in a china dish. The solution is boiled and filtered. The filtrate is known as sodium fusion extract. By using sodium fusion extract, elements like N, S and halogens can be detected as follows:

No.	Experiment	Observation	Inference
1.	To one part of sodium fusion extract add freshly prepared ferrous sulphate (FeSO_4) solution. Heated to boiling, cooled and acidified with dil. H_2SO_4 .	Blue or green coloration or precipitate (ppt)	Presence of nitrogen
2.	A little of the sodium fusion extract is acidified with dil. HNO_3 and then silver nitrate (AgNO_3) is added.	White ppt soluble in ammonium hydroxide (NH_4OH)	Presence of Chlorine
		Pale yellow ppt slightly soluble in NH_4OH	Presence of Bromine
		Yellow ppt insoluble in NH_4OH	Presence of Iodine
3.	To a little of the sodium fusion extract, add sodium nitroprusside solution	Violet colouration	Presence of sulphur

29. Write the name of the test used to detect nitrogen, sulphur, halogens and phosphorous present in an organic compound.

Ans: Lessaigne's test



30. Organic compounds have to be purified before analysis.

- Which type of liquids can be purified using distillation under reduced pressure? Suggest an example.
- Name the two main types of chromatographic techniques based on the principle of differential adsorption.
- In the Lassaigne's test for halogens, they are precipitated as
- In what form is nitrogen estimated in the Dumas method?

Ans: a) Liquids having very high boiling points and those, which decompose at or below their boiling points. Glycerol can be separated from spent-lye in soap industry by using this technique.

- i) Column chromatography, and
ii) Thin layer chromatography.
- Silver halides (AgX)
- Molecular nitrogen (N_2)

31. Name the method for estimation of Halogen.

Ans: Carius method

32. Explain any one method for the estimation of nitrogen present in an organic compound.

Ans: **Dumas method:** Here the organic compound is heated with copper oxide in an atmosphere of carbon dioxide so that free nitrogen, carbon dioxide and water are produced. This mixture of gases is collected over an aqueous solution of potassium hydroxide which absorbs carbon dioxide. Nitrogen is collected in the upper part of the graduated tube. From the volume of nitrogen collected, we can determine the % of Nitrogen.

33. In the Carius method of estimation of halogen, 0.15g of an organic compound gave 0.12g of AgBr . Find the percentage of Br in the compound.



Ans: From Carius method,

$$\begin{aligned}\text{Percentage of halogen} &= \frac{\text{Atomic mass of halogen} \times m_1 \times 100}{\text{Molecular mass of AgX} \times m} \\ &= \frac{80 \times 0.12 \times 100}{188 \times 0.15} = 34\%\end{aligned}$$

34. Different methods are used to purify organic compounds. Name any three methods of purification.

Ans: Sublimation, Crystallisation and Distillation.

35. On complete combustion, 0.246g of an organic compound gave 0.198g of CO₂ and 0.1014g of H₂O.

Determine the percentage composition of carbon and hydrogen in the compound.

$$\text{Percentage of hydrogen} = \frac{2 \times m_1 \times 100}{18 \times m} = \frac{2 \times 0.1014 \times 100}{18 \times 0.246} = 4.6\%$$

$$\text{Percentage of carbon} = \frac{12 \times m_2 \times 100}{44 \times m} = \frac{12 \times 0.198 \times 100}{44 \times 0.246} = 21.95\%$$

36. Different techniques are used for the purification of organic compounds based on their nature.

- Suggest a suitable method for the separation of a mixture of aniline and water.**
- Give the chemical name of the compound responsible for the blue colour in the Lessaigne's test for nitrogen.**
- Briefly explain the principle involved in Kjeldahl's method for the estimation of nitrogen.**

Ans: a) steam distillation

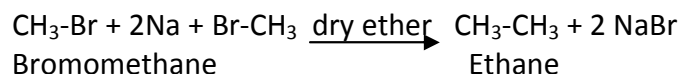
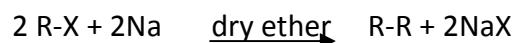
b) Ferriferrocyanide

c) Kjeldahl's method: Here the organic compound containing nitrogen is heated with concentrated sulphuric acid. Nitrogen in the compound gets converted to ammonium sulphate. The resulting acid mixture is then heated with excess of sodium hydroxide. The liberated ammonia gas is absorbed in an excess of standard solution of sulphuric acid. The amount of ammonia produced is determined by estimating the amount of sulphuric acid consumed in the reaction. It is done by estimating unreacted sulphuric acid left after the absorption of ammonia by titrating it with standard alkali solution.

13. HYDROCARBONS

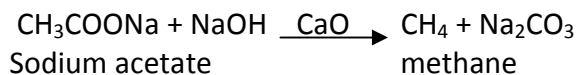
1. What is Wurtz reaction?

Alkyl halides react with metallic sodium in dry ether to form alkanes. This reaction is known as Wurtz reaction.



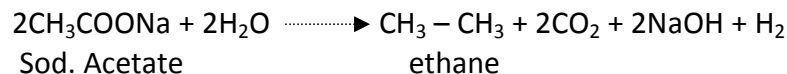
2. How will you convert sodium acetate to methane?

Sodium acetate on heating with soda lime (a mixture of NaOH and CaO), it undergoes decarboxylation to form methane.



3. How will you convert sodium acetate to ethane?

An aqueous solution of Sodium acetate on Kolbe's electrolysis, we get ethane



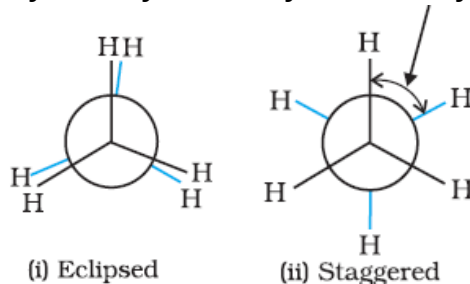
4. What are conformations?

The different spatial arrangements of atoms arising due to free rotation around a C-C single bond are called conformations or conformers.

5. Which are the different types of conformations of ethane? Among these which is most stable

Ethane has mainly two types of conformations – Eclipsed conformation and staggered conformation. Among these, staggered conformation is more stable. This is because here the electron clouds of carbon-hydrogen bonds are very far apart. So there is minimum repulsive forces and maximum stability.

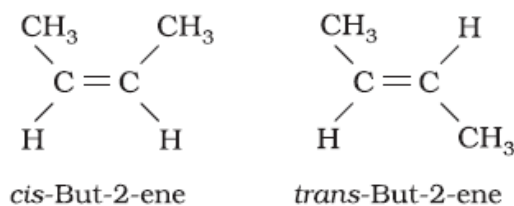
6. Draw the Newman projection formula for the conformations of ethane?



7. What is mean by Geometrical isomerism? Draw the geometrical isomers for 2-butene.

The isomerism arising due to the difference in the spatial arrangement of atoms around carbon-carbon double bond is called geometrical isomerism. There are two types of geometrical isomers – **cis isomer and trans isomer**.

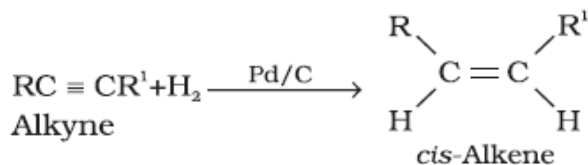
e.g. 2-butene



8. What is Lindlar's catalyst? Give its use.

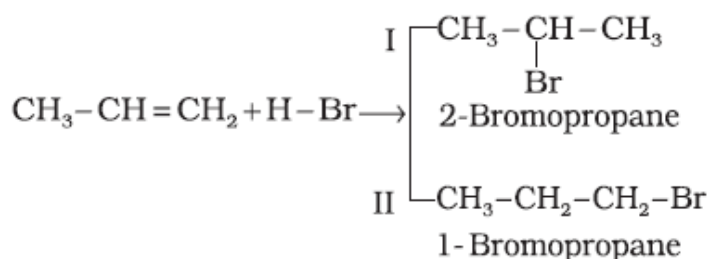
Partially deactivated palladised charcoal is known as **Lindlar's catalyst**.

Alkynes on reduction with Lindlar's catalyst, we get cis-alkene.



9. What is the major product formed on the addition of HBr to 1-propene? Name the rule you used to choose the major product. State it.

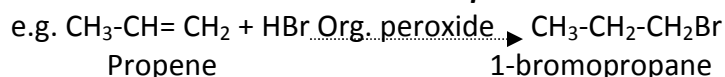
When HBr is added to propene, we get 2 products – 1-bromopropane and 2-bromopropane.



Here the major product is determined by a rule called **Markownikoff's rule**. The rule states that "**when an unsymmetrical reagent is added to an unsymmetrical alkene, the negative part of the addendum (adding molecule) gets attached to the carbon containing lesser number of hydrogen atoms**"

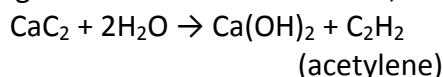
10. What is peroxide or Kharasch effect? Give an example.

In the presence of peroxide, addition of HBr to unsymmetrical alkenes takes place against Markovnikov rule. This is known as **peroxide or Kharash effect or anti-Markovnikov addition reaction**.



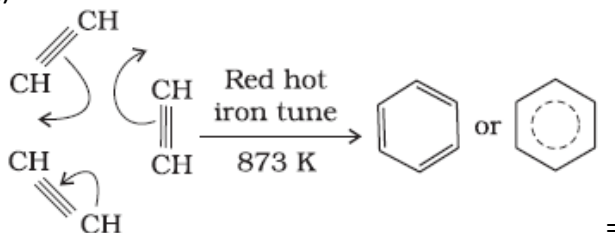
11. Give the preparation of acetone (C_2H_2) from calcium carbide?

On treating calcium carbide with water, we get acetylene gas.



12. How will you convert acetylene to benzene?

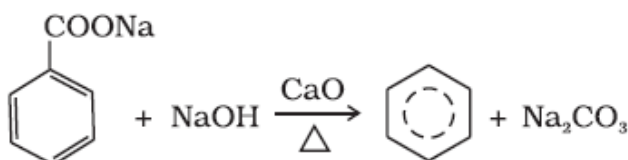
Ethyne (acetylene) on passing through red hot iron tube at 873K, undergoes cyclic polymerisation to form benzene (C_6H_6).



13. State Hickel's ($4n + 2$) π electron rule.

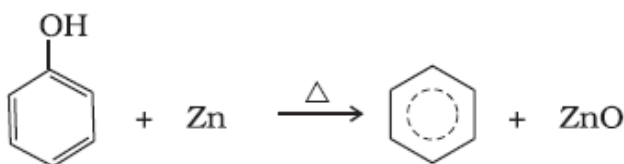
According to this rule, "cyclic, planar, conjugated systems containing $(4n+2)$ π electrons are aromatic". Where n is the number of rings. n may be 1,2,3,....

14. How will you convert sodium benzoate to benzene ?



15. How will you convert phenol to benzene?

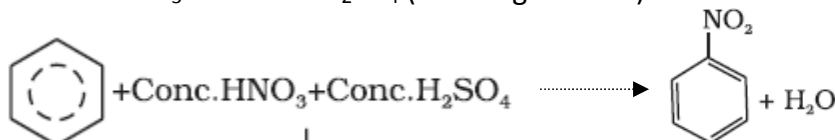
Phenol is reduced to benzene by passing its vapours over heated zinc dust.



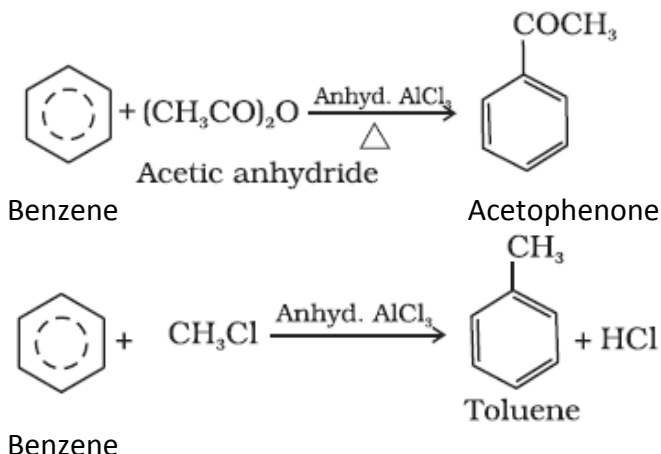
16. What are electrophilic substitution reactions? Give examples.

These are reactions in which a weak electrophile is replaced by a strong electrophile. The important electrophilic substitution reactions are Nitration, Sulphonation, Halogenation and Friedel-Crafts alkylation and acylation.

Nitration: It is the introduction of nitro ($-\text{NO}_2$) group to a benzene ring. For this benzene is heated with a mixture of conc. HNO_3 and conc. H_2SO_4 (nitrating mixture).

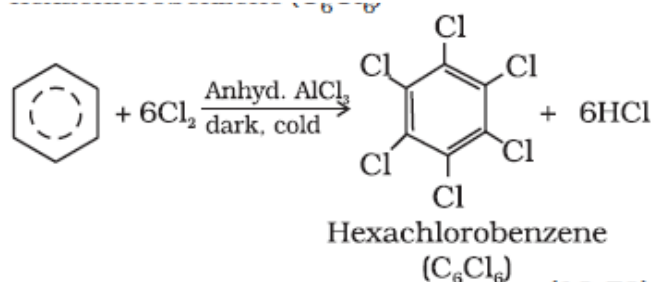


17. Convert benzene to acetophenone and toluene.



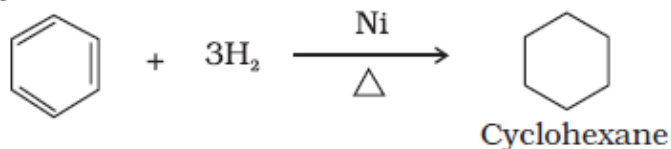
18. What is gammexane or BHC? Give its preparation.

Benzene adds chlorine in presence of uv light to form benzene hexachloride (BHC). It is also known as Gammexane or Lindane or 666.



19. How will you convert benzene to cyclohexane?

Benzene add hydrogen in presence of nickel catalyst at high temperature and pressure to form cyclohexane.



14. ENVIRONMENTAL CHEMISTRY

1. Carbon monoxide is one of the most serious air pollutants. Why?

Carbon monoxide binds to haemoglobin in blood to form carboxyhaemoglobin. It reduces the oxygen carrying capacity of blood. This oxygen deficiency, results into headache, weak eyesight etc.

2. What are the major gases which contribute towards global warming? What can we do to reduce global warming?

Gases like carbon dioxide, methane, ozone, chlorofluorocarbon compounds (CFCs) and water vapour are responsible for global warming. We can reduce Global warming by following method

- to minimise the use of automobiles.
- plant more trees to increase the green cover



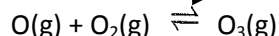
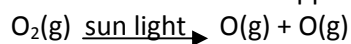
3. What is mean by the green house effect?

On increasing the concentration of CO₂ in the atmosphere, it traps the earth radiation and hence the temperature of the earth surface increases. This phenomenon is known as Green House effect. This results in melting of polar ice caps and flooding of low lying areas.

4. Ozone in the stratosphere is produced by the help of ultraviolet radiations. It protects us from harmful ultraviolet radiations.

- Write equation for the formation of ozone in stratosphere.
- Explain with chemical equation, the destruction of ozone by chlorofluoro carbons causing ozone hole.

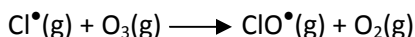
Ans: a) Ozone is formed in the upper stratosphere by the interaction of uv radiation on dioxygen.



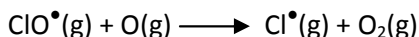
b) In stratosphere, CFC's get broken down by UV radiations, releasing chlorine free radical.



The chlorine radical then react with ozone to form chlorine monoxide radicals and molecular oxygen.



Reaction of chlorine monoxide radical with atomic oxygen produces more chlorine radicals.



The chlorine radicals are continuously regenerated and cause the breakdown of ozone.

5. Suppose that your teacher asks you to conduct a seminar on ozone depletion. Give any three harmful effects of ozone depletion that you would present in the seminar.

- With the depletion of ozone layer, more UV radiation filters into troposphere. UV radiations lead to ageing of skin, cataract, sunburn, skin cancer, killing of many phytoplanktons, damage to fish productivity etc.
- It has also been reported that plant proteins get easily affected by UV radiations which leads to the harmful mutation of cells.
- It also increases evaporation of surface water through the stomata of the leaves and decreases the moisture content of the soil. Increase in UV radiations damage paints and fibres, causing them to fade faster.

6. What is mean by 'acid rain'? Explain the causes and harmful effects of acid rain.

When the pH of the rain water drops below 5.6, it is called acid rain. Oxides of nitrogen and sulphur (e.g. SO₂ and NO₂) are mainly responsible for acid rain. The harmful effects of acid rain are:

- It is harmful for agriculture, trees and plants.
- It causes respiratory ailments and skin cancer in human beings and animals.
- It affects plants and animal life in aquatic ecosystem.
- It corrodes water pipes resulting in the dissolution of heavy metals into the drinking water.
- Acid rain damages buildings and other structures made of stone or metal (e.g. The Taj Mahal).

7. What are the different measures to reduce acid rain?

Acid rain can be reduced by the following methods:

- Reduce the use of automobiles.
- Encourage the use of natural gas instead of coal.
- Catalytic converters must be used in motor vehicles to reduce the emission of CO and oxides of Nitrogen to the atmosphere.
- Acidity of the soil can be reduced by adding powdered limestone.

e) Aware the public about acid rain.

8. There are international standards regarding drinking water. Write any three among them.

The International Standards for drinking water are given below:

Fluoride: For drinking purposes, water should contain fluoride upto 1 ppm (parts per million).

Lead: The upper limit concentration of lead in drinking water is about 50 ppb (parts per billion).

Sulphate: Less than 500 ppm

9. Pollution of water originates mainly from human activities.

a) What do you mean by the term PCB?

b) How do chemical pollutants cause eutrophication?

c) Mention the adverse effects of high fluoride concentration in drinking water.

Ans: a) PCB – Polychlorinated biphenyls (Pollution Control Board)

b) Chemical pollutants cause Eutrophication. Fertilizers contain phosphates and the presence of phosphates in water increases the growth of algae. The increased amount of algae reduces the oxygen concentration in water.

c) Excess fluoride is harmful to teeth and bones.

10. Write any two differences between classical smog and photochemical smog.

Classical smog occurs in cool humid climate. Chemically it is a reducing mixture. So it is called reducing smog.

Photochemical smog occurs in warm, dry and sunny climate. Photochemical smog has high concentration of oxidising agents and so it is also called oxidising smog.

11. Write any two adverse effects of photochemical smog.

It leads to cracking of rubber and extensive damage to plant life.

It also causes corrosion of metals, stones, building materials, rubber and painted surfaces.

12. What is mean by viable and non-viable particulate pollutants? Give examples.

The viable particulates are minute living organisms that are dispersed in the atmosphere. E.g., bacteria, fungi, moulds, algae etc. The non-viable particulates are non-living like smoke, fumes, mist, dust etc.

13. How do organic pollutants affect aquatic life?

As the amount of organic matter in water increases, more oxygen is required to decompose them by bacteria. So the amount of dissolved oxygen in water decreases. This causes oxygen dependent aquatic life to die.

14. What is Biological Oxygen Demand (BOD)?

The amount of oxygen required by bacteria to break down the organic matter present in a certain volume of a sample of water is called Biochemical Oxygen Demand (BOD). The amount of BOD in the water is a measure of the amount of organic material in the water.

15. Give any two applications of Green Chemistry in day-to-day life.

a) **Dry Cleaning of Clothes:** Liquefied carbondioxide, with a suitable detergent is used for dry cleaning clothes.

b) **Bleaching of Paper:** Hydrogen peroxide (H_2O_2) with suitable catalyst is used for bleaching paper.

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