

SOME BASIC CONCEPTS OF CHEMISTRY

1. NO and NO₂ are two oxides of nitrogen. Which law of chemical combination is illustrated by these compounds? State the law. Who proposed this law?
Ans: **Law of Multiple proportion.** 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?
Ans: It states that a given compound always contains exactly the same proportion of elements by weight. This law was proposed by **Joseph Proust**.
3. State Law of conservation of mass (matter). Who proposed this law?
Ans: 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
Ans: 1/12th the mass of a C¹² atom is called atomic mass unit (amu).
5. Define 1 mole.
Ans: 1 mole is the amount of substance that contains as many particles as there are atoms in exactly 12 g C¹² isotope.
6. Find the number of oxygen atoms present in 5 moles of glucose (C₆H₁₂O₆).
Ans: 1 mol glucose contains 6 mol Oxygen atoms.
So, 5 mol glucose contain $6 \times 5 = 30$ mol Oxygen atoms
 $= 30 \times 6.022 \times 10^{23}$ Oxygen atoms
7. Write the number of significant figures in the following:
(i) 0.0205 (ii) 3.00×10^{-5} (iii) 1005
Ans: (i) 3 (ii) 3 (iii) 4
8. Define mass percentage.
Ans: It is the percentage of each element present in 100g of a substance.
i.e. percentage composition (mass percent) of an element
$$= \frac{\text{Mass of that element in the compound} \times 100}{\text{Molar mass of the compound}}$$
9. What are empirical and molecular formulae?
Ans: 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.

10. A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% chlorine. Its molecular mass is 98.96. What are the empirical and molecular formulae

Ans:

Element	Percentage	Atomic mass	Percentage/Atomic mass	Simple ratio	Simplest whole no. ratio
C	24.27	12	$24.27/12 = 2.02$	$2.02/2.02 = 1$	1
H	4.07	1	$4.07/1 = 4.07$	$4.07/2.02 = 2$	2
Cl	71.65	35.5	$71.65/35.5 = 2.02$	$2.02/2.02 = 1$	1

Empirical Formula = CH_2Cl

Empirical Formula Mass (EFM) = $12+2+35.5 = 49.5$

Molar mass (MM) = 98.96

$$n = \text{MM}/\text{EFM} = 98.96/49.5 = 2$$

Molecular formula = Empirical formula $\times n$

$$= (\text{CH}_2\text{Cl}) \times 2 = \text{C}_2\text{H}_4\text{Cl}_2$$

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

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

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

12. What is limiting reagent?

Ans: It is the reagent which limits a reaction. OR, it is the reagent which is completely consumed in a chemical reaction.

13. Differentiate between molarity and molality.

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

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

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

$$\text{OR, Molality (m)} = \frac{\text{Number of moles of solute}}{\text{Mass of solvent in kg}}$$

Molality is temperature independent, while molarity changes with temperature.

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

Ans:

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

$$\text{No. of moles of NaOH} = \frac{\text{Given mass}}{\text{Molar mass}} = \frac{4}{40} = 0.1 \text{ mol}$$

$$\text{Volume of solution in L} = 250/1000 = 0.25 \text{ L}$$

$$\text{Molarity} = \frac{0.1}{0.25} = 0.4\text{M}$$

STRUCTURE OF ATOM

1. Give any two characteristic properties of cathode rays.

Ans: (i) They are negatively charged particles.

(ii) They can produce x-rays

2. Write any two characteristic properties of canal rays (anode rays).

Ans: (i) They are positively charged particles.

(ii) Their properties depend on the nature of gas present in the cathode ray tube.

3. What are the important observations and conclusions made by Rutherford in his α -particle scattering experiment ?

Ans: The important observations are:

i. Most of the α - particles passed through the gold foil without any deviation.

ii. A small fraction of the α -particles was deflected by small angles.

iii. A very few α - particles were rebounded (Or, deflected by nearly 180°).

Conclusions:

i. Since most of the α -particles passed through the foil without any deviation, most space in the atom is empty.

ii. In an atom, there is a positively charged centre called nucleus.

iii. The volume of the nucleus is negligibly small as compared to the total volume of the atom.

4. What are the postulates of Rutherford atom model (Planetary model of atom)?

Ans: The important postulates are:

i. All the positive charge and most of the mass of the atom are concentrated at the centre called nucleus.

ii. Electrons are revolving round the nucleus with a very high speed in circular paths called orbits.

iii. Electrons and the nucleus are held together by electrostatic forces of attraction.

5. What are the drawbacks of Rutherford atom model?

Ans: (i) He could not explain the stability of the atom.

(ii) He could not explain the electronic structure of atom.

6. What is photoelectric effect?

Ans: 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.

7. What are the important characteristics of photoelectric effect? [Write the important results observed during photoelectric effect].

The important characteristics of photoelectric effect are:

i. The electrons are ejected from the metal surface as soon as the beam of light strikes the surface.

ii. The number of electrons ejected is proportional to the intensity or brightness of light.

iii. 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.

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

Ans: The hydrogen spectrum consists of mainly five series of lines which are Lyman, Balmer, Paschen, Brackett and Pfund series.

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

9. The series of lines in hydrogen spectrum that is visible to our naked eye is

Ans: Balmer series

10. 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 $h/2\pi$.

11. Write any 3 limitations of Bohr atom model?

Ans: (i) It could not explain the fine spectrum of hydrogen atom.

(ii) It could not explain the spectrum of atoms other than hydrogen.

(iii) It could not explain Stark effect and Zeeman effect.

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

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

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

Ans: de Broglies equation is $\lambda = h/p$

Or, $\lambda = h/mv$ Where λ is the wavelength, m is the mass, v is the velocity and p is the momentum of the particle.

14. What will be the wavelength of a ball of mass 0.1 kg moving with a velocity of 10 m/s ? ($h = 6.626 \times 10^{-34}$ Js)

Ans: Here $m = 0.1$ kg and $v = 10$ m/s

From the de Broglie's equation, $\lambda = \frac{h}{mv}$

$$= \frac{6.626 \times 10^{-34}}{(0.1 \times 10)} = \underline{\underline{6.626 \times 10^{-34} \text{ m}}}$$

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

Ans: It states that "it is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron".

Mathematically,

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

$$\text{OR, } \Delta x. m\Delta v \geq \frac{h}{4\pi}$$

Where Δx is the uncertainty in position, Δp is the uncertainty in momentum, Δv is the uncertainty in velocity and m is the mass of electron.

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

Ans: 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:

1. **Principal Quantum Number (n):** It gives the size the orbit, the energy of electron in an orbit, the shell in which the electron is found and the average distance between the electron and the nucleus.

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

2. **Azimuthal Quantum Number (l):** It gives the shape of the orbital, the sub shell in which the electron is found and the orbital angular momentum of the electron. The possible values of l are: $l = 0, 1, 2, \dots, (n-1)$.

3. **Magnetic Quantum Number (m or m_l):** It gives 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 $+l$

4. **Spin Quantum Number (s or m_s):** 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.

17. Represent the orbital with the following quantum numbers:

(i) $n = 4$ and $l = 0$ (ii) $n = 5, l = 3$ (iii) $n = 3, l = 2$

Ans: (i) 4s (ii) 5f (iii) 3d

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

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

(ii) $n = 2, l = 1, m = 0, s = -\frac{1}{2}$

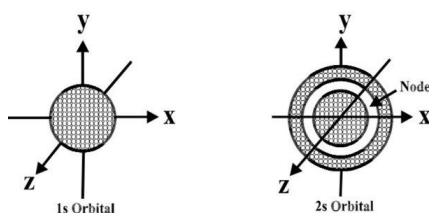
(iii) $n = 1, l = 0, m = 0, s = +\frac{1}{2}$

(iv) $n = 0, l = 0, m = 0, s = +\frac{1}{2}$

Ans: Here (i) and (iv) are not allowed. (i) is not possible, since the value of l cannot be equal to n . (iv) is not possible, because the values of n cannot be equal to zero.

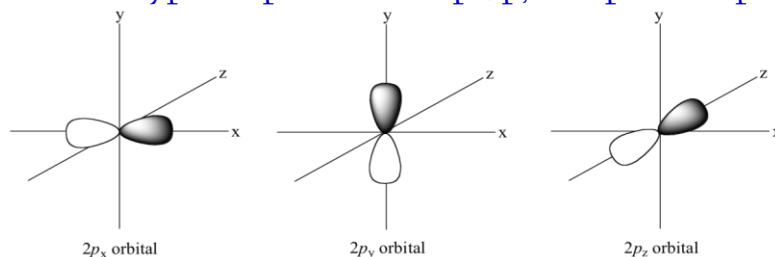
19. Draw the shapes of 1s and 2s orbitals.

Ans:



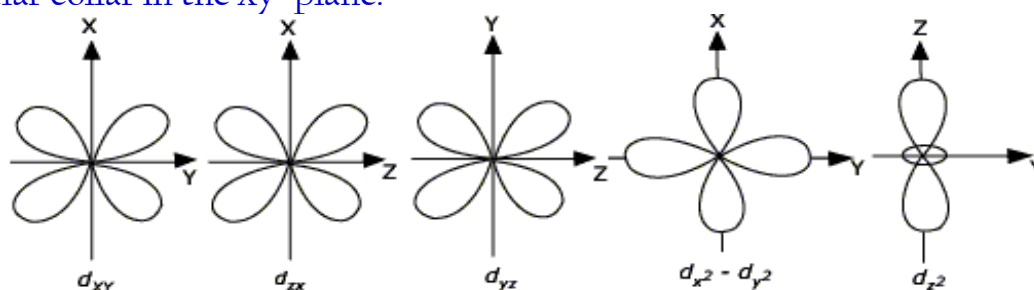
20. Draw the shapes of different p-orbitals.

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



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

Ans: 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.



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

The filling of electrons in different orbitals 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 or $(n + l)$ values.

2. Pauli's Exclusion Principle

It states that no two electrons in an atom can have the same set of 4 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.

23. Write the number of radial nodes and angular nodes for 5f orbital.

Ans: Radial nodes = 1 [Equation for radial nodes = $n - l - 1 = 5 - 3 - 1 = 1$]

Angular nodes = 3 [Equation for angular nodes = l]

24. Write the subshell electronic configurations of Cu ($Z = 29$) and Cr ($Z = 24$). Give reason for their extra stability.

Ans: $_{29}\text{Cu}$: $[\text{Ar}] 3d^{10}4s^1$ OR, $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

$_{24}\text{Cr}$: $[\text{Ar}] 3d^54s^1$ OR, $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$

Half-filled and completely filled electronic configurations have extra stability due to the symmetrical arrangement of electrons and greater exchange energy.

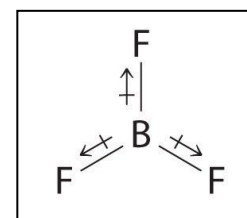
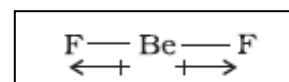
CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

1. Who introduced the periodic law of elements for the first time? State the law.
Ans: Mendeleev. It states that “the properties of elements are the periodic functions of their atomic masses.”
2. State Modern Periodic law. Who proposed this law?
Modern periodic law states that “the properties of elements are the periodic functions of their atomic numbers.” This law was proposed by Henry Moseley.
3. 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 .
 - 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. 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. 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 and actinides. Their general outer electronic configuration is $(n-2)f^{1\text{ to }14}(n-1)d^{0\text{ to }1} ns^2$.
4. Write any 4 properties of transition elements?
Ans: They are all metals, they form coloured compounds or ions, they show variable oxidation states and valencies, they show paramagnetism and catalytic properties.
5. How does atomic radius vary along a group and period and why?
Ans: The atomic size decreases from left to right in a period, due to increase in the effective nuclear charge. Also, the number of shells remain constant in a period.
Down a group, the atomic radius increases from top to bottom, due to the increase in the number of shells and shielding effect.
6. What are isoelectronic species? Give examples.
Ans: 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).
7. Define ionisation enthalpy.
Ans: 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.
8. How does ionisation enthalpy vary along a period and group? Justify your answer.
Ans: Along a period, ionisation enthalpy increases from left to right, due to decrease in atomic radius and increase in nuclear charge.
Down a group, ionization enthalpy decreases due to increase in atomic radius and shielding effect.

9. The first ionisation enthalpy of Boron is slightly less than that of Beryllium. Why?
Ans: Because of the stable fully filled electronic configuration of B^+ ion. ($B^+ - 1s^2 2s^2 2p^1$)
10. The first ionisation enthalpy of Nitrogen is greater than that of Oxygen. Why?
Ans: This is because Nitrogen has half-filled p-orbital configuration ($1s^2 2s^2 2p^3$), which is more stable and so more energy is required to remove an electron.
11. Define electron gain enthalpy ($\Delta_{eg}H$).
Ans: It is the enthalpy change when an electron is added to the outer most shell of an isolated gaseous atom. Its unit is kJ/mol.
12. Chlorine has higher negative electron gain enthalpy than fluorine. Why?
Or, Electron gain enthalpy of fluorine is less negative than chlorine. Why?
Ans: Due to the compactness of the 2p subshell of F, electron-electron repulsion is greater in F and hence it does not easily add electron. [Or, Due to larger size and less electron-electron repulsion in chlorine].
13. Electron gain enthalpy values of noble gases are zero. Why?
Ans: This is because of their stable fully filled electronic configuration.
14. Define electronegativity.
Ans: Electronegativity of an atom in a compound is the ability of the atom to attract shared pair of electrons.
15. Phosphorus forms PCl_5 while nitrogen cannot form NCl_5 . Why?
Ans: Because of the absence of vacant d-orbitals in nitrogen.
16. Write the IUPAC name of the element with the atomic numbers 109 & 117.
Ans: 109 – Unnilennium
117 – Ununseptium

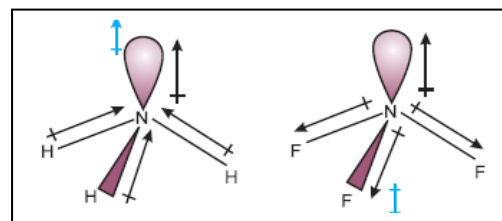
CHEMICAL BONDING AND MOLECULAR STRUCTURE

1. State octet rule. Who proposed this rule?
Ans: Octet rule states that atoms containing 8 electrons in their valence shell are stable. OR, atoms undergo chemical reaction in order to attain 8 electrons in the valence shell. It was proposed by Lewis and Kossel.
2. Give any 2 limitations of Octet rule.
Ans: (i) It could not explain the stability of molecules containing odd number of electrons.
(ii) It could not account for the shape of molecules.
3. Define dipole moment. Give its unit.
Ans: Dipole moment is the product of charge at one end (e) and distance between the charges (r). i.e. $\mu = e \times r$. Its unit is Debye (D)
4. The dipole moment of BeF_2 is zero. Why?
Ans: BeF_2 , the net dipole moment is zero, since the two equal bond dipoles are in opposite directions and cancel each other.
5. The dipole moment of BF_3 is zero. Why?
Ans: In BF_3 , the net dipole moment is zero. Here the resultant of any 2 bond dipoles is equal and opposite to the third.



6. Ammonia (NH_3) has higher dipole moment than NF_3 , even though F is more electronegative than H. why?

Ans: This is because in the case of NH_3 , 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 NF_3 , 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.



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

Ans: (i) The smaller the size of the cation and the larger the size of the anion, the greater the covalent character of an ionic bond.

(ii) The greater the charge on the cation, the greater the covalent character of the ionic bond.

(iii) 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 $(ns^2 np^6)$.

8. What are 3 postulates of valence shell electron pair repulsion (VSEPR) theory?

Ans: (i) The shape of the molecule depends on the number of valence shell electron pairs (VSEPRs) around the central atom.

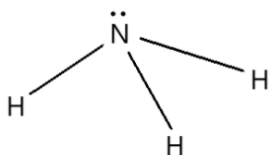
(ii) The valence shell electron pairs repel each other.

(iii) As the angle between the electron pairs increases, the repulsion decreases.

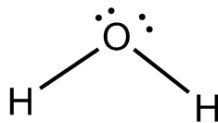
9. Explain the shape of ammonia and water molecules on the basis of VSEPR theory.

Ans: In ammonia, there are 4 VSEPs (3 bond pairs and one lone pair). 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, there are 4 VSEPs (2 bond pairs and 2 lone pairs). 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° .



NH_3 Molecule



H_2O Molecule

10. Write any 3 important characteristics of Hybridisation?

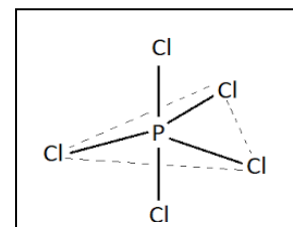
Ans: (i) The number of hybrid orbitals formed is equal to the number of atomic orbitals undergoing hybridization.

(ii) The hybrid orbitals are always equivalent in energy and identical in shape.

(iii) The hybrid orbitals are directed to some fixed positions in space. So the type of hybridization gives the shape of the molecule.

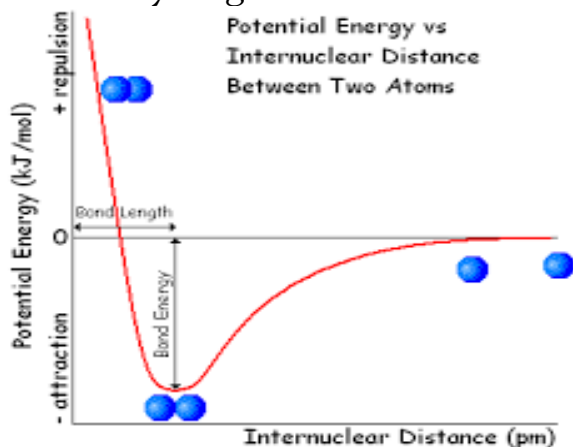
10. Explain the geometry of PCl_5 molecule and account for its high reactivity.

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



In PCl_5 , there are two types of P–Cl bonds – two axial bonds and three equatorial bonds. Due to the greater repulsion, 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. Define bond order.

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) = $\frac{1}{2} [N_b - N_a]$

13. He_2 does not exist. Why?

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

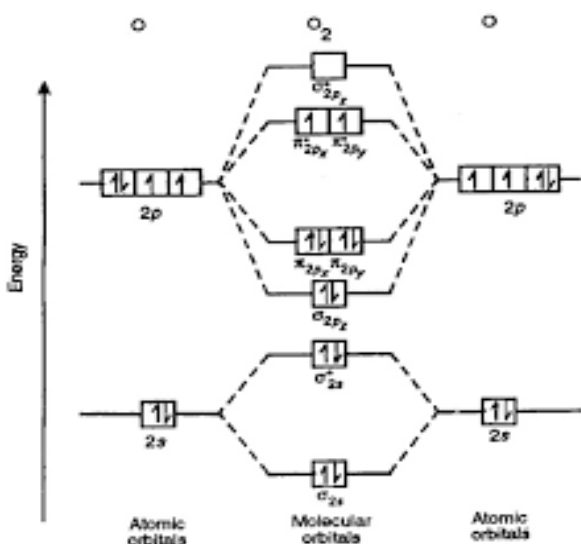
14. Ne_2 does not exist. Why?

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

15. 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.

16. Write the molecular electronic configuration of the N_2 molecule. Predict the stability and magnetic property of N_2 with reasons.

Ans: The M.O configuration of N_2 is $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2$

$$\text{Bond Order} = \frac{1}{2} [N_b - N_a] = \frac{1}{2} [10 - 4] = \frac{1}{2} \times 6 = 3$$

Since bond order is positive N_2 is stable. Also, due to the absence of unpaired electrons, N_2 is diamagnetic.

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

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

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

Since bond order is positive, it is stable.

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

$$B.O = \frac{1}{2} [N_b - N_a] = \frac{1}{2} [10 - 8] = \frac{1}{2} \times 2 = 1$$

Since bond order of O_2 is greater than O_2^{2-} , it is more stable.

18. Define hydrogen bonding with example. Which are the different types of H-bonding?

Ans: 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 H-F H-F.....H-F.....H-F.....H-F.....

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

19. Which has higher boiling point: o-nitrophenol or p-nitrophenol? Give reason.

Ans: p-Nitrophenol. Because of the presence of intermolecular H-bonding in it.

THERMODYNAMICS

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

Ans: Extensive properties are properties which depend on the amount of matter present in the system. E.g.: Volume, length, height, 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, Pressure, molar volume, density, refractive index, molar heat capacity, viscosity, surface tension etc.

2. Define state function and path function.

Ans: 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 and work

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

Ans: 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 .

Ans: $\Delta H = \Delta U + P\Delta V$

5. State and illustrate Hess's law of constant summation.

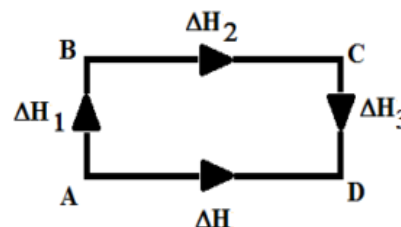
Ans: The law states that the total enthalpy change for a process is the same whether the reaction taking place in a single step or in several steps. Or, the total enthalpy change for a process is independent of the path followed.

Illustration:

Consider a process in which the reactant A is converted to product D in a single step by

involving heat change, ΔH . Let the same reactant A is first converted to B, then to C and finally to D involving heat changes ΔH_1 , ΔH_2 and ΔH_3 respectively.

Then according to Hess's law: $\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$

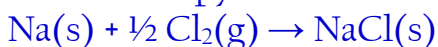


6. Define lattice enthalpy.

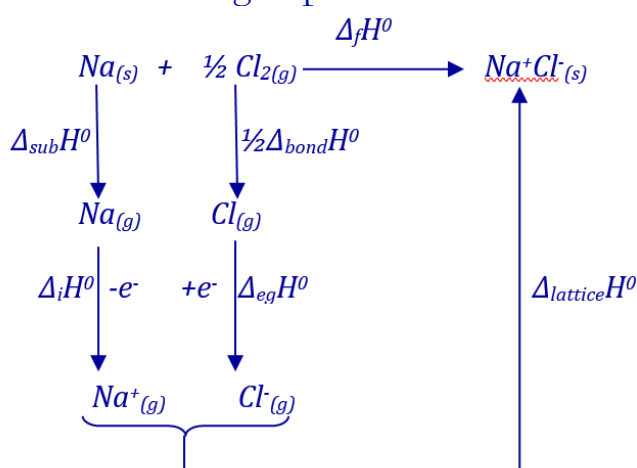
Ans: The lattice enthalpy of an ionic compound is the enthalpy change when one mole of an ionic compound dissociates into gaseous ions.

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

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



This involves the following steps:



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

From this lattice enthalpy can be determined as:

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

8. What are spontaneous processes? Give examples.

Ans: It is a process that takes place without the help of any external agency. E.g. flow of water from high level to low level, burning of fuels, melting of ice, evaporation of water etc.

9. What are the criteria for spontaneous process?

Ans: Decrease in energy and increase in disorderness.

10. Define entropy. What is its unit?

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

11. What happens to the entropy during the following changes?

- a) A gas condenses into liquid b) $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

Ans: a) Entropy decreases.

b) Entropy increases

12. State the second law of thermodynamics.

Ans: It states that the entropy of the universe always increases during every spontaneous process.

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

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

14. Calculate the work done for the reversible isothermal expansion of 1 mole of an ideal gas at 27°C, from a volume of 10 dm³ to a volume of 20 dm³.

Ans: Here $n = 1$ mol, $R = 8.314$ J/K/mol, $V_1 = 10$ dm³, $V_2 = 20$ dm³ and $T = 25 + 273 = 298$ K

$$\begin{aligned}\text{For isothermal reversible expansion, work done, } W_{\text{exp}} &= -2.303nRT \log(V_2/V_1) \\ &= -2.303 \times 1 \times 8.314 \times 298 \times \log(20/10) \\ &= \underline{-1717.46 \text{ J}}\end{aligned}$$

15. Enthalpy and entropy changes of a reaction are 40.63 kJ/mol and 108.8 J/K/mol. Predict the feasibility of the reaction at 27°C.

Ans: Given $\Delta H = 40.63$ kJ/mol = 40630 J/mol, $\Delta S = 108.8$ J/K/mol and $T = 27 + 273 = 300$ K.

From Gibb's equation, $\Delta G = \Delta H - T\Delta S$

$$= 40630 - 300 \times 108.8 = \underline{7990 \text{ J/mol}}$$

Since ΔG^0 is positive, the process is non-spontaneous at this temperature.

16. Calculate the enthalpy of formation of CH₄ from the following data:

(i) $\text{C(s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}); \Delta H = -393.7$ kJ/mol

(ii) $\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O(l)}; \Delta H = -285.8$ kJ/mol

(iii) $\text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O(l)}; \Delta H = -890.4$ kJ/mol

Ans: The required equation is: $\text{C(s)} + 2 \text{H}_2(\text{g}) \rightarrow \text{CH}_4(\text{g})$

Given: $\text{C(s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}); \Delta H = -393.7$ kJ/mol

Multiply equation (ii) $\times 2$; $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O(l)}; \Delta H = -285.8 \times 2 = -571.6$ kJ/mol

Reverse equation (iii) $\text{CO}_2(\text{g}) + 2 \text{H}_2\text{O(l)} \rightarrow \text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}); \Delta H = 890.4$ kJ/mol

Add the above three equations we get $\text{C(s)} + 2 \text{H}_2(\text{g}) \rightarrow \text{CH}_4(\text{g});$

$$\begin{aligned}\Delta H &= -393.7 + -571.6 + 890.4 \\ &= \underline{-74.9 \text{ kJ/mol}}\end{aligned}$$

17. Calculate the standard free energy (ΔG^0) for the conversion of oxygen to ozone:

$\frac{3}{2} \text{O}_2(\text{g}) \rightarrow \text{O}_3(\text{g})$ at 298K, if the equilibrium constant for the conversion is 2.47×10^{-29} . (Given $R = 8.314$ J K⁻¹ mol⁻¹).

Ans: Here $K = 2.47 \times 10^{-29}$, $R = 8.314$ J/K/mol and $T = 298$ K

$$\begin{aligned}\Delta G^0 &= -2.303RT \log K \\ &= -2.303 \times 8.314 \times 298 \times \log(2.47 \times 10^{-29}) \\ &= 163000 \text{ J/mol} = \underline{163 \text{ kJ/mol}}\end{aligned}$$

EQUILIBRIUM

1. Write the expression for equilibrium constant for the reaction i) $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$ and ii) $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$.

Ans: i) $K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$

ii) $K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$

2. What is the relation between K_c and K_p ?

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

3. Give the relation between K_p and K_c , for the reaction $2\text{NOCl(g)} \rightleftharpoons 2\text{NO(g)} + \text{Cl}_2(\text{g})$

Ans: Here $\Delta n = 3 - 2 = 1$. So $K_p = K_c.RT$

4. Give any 3 characteristics of equilibrium constant.

Ans: (i) Equilibrium constant is applicable only when the concentrations of the reactants and products have attained their equilibrium state.

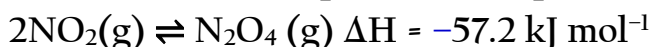
(ii) The value of equilibrium constant is independent of the initial concentrations of reactants and products.

(iii) The value of equilibrium constant depends on temperature.

5. State Le-Chatlier Principle.

Ans: 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.

6. Explain the effects of temperature and pressure on the following equilibrium.



Ans: Here the forward reaction is exothermic. So, when temperature increases, the rate of forward reaction decreases.

Here the forward reaction results in the decrease in no. of moles of gaseous species. So, high pressure favours the forward reaction.

7. Predict the effect of change of pressure in the equilibrium : $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$

Ans: Here the total no. of moles of reactants are products are same. So, pressure has no effect.

8. Write the Arrhenius concept of acids and bases. Give one example each for acid and base.

Ans: According to this concept acids are hydronium ion (H_3O^+) producers and bases hydroxyl ion (OH^-) producers in aqueous solution.

E.g. for acid is HCl and for base is NaOH .

9. Write Bronsted concept of acids and bases? Give example.

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

For example in the reaction $\text{NH}_3(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$

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^- .

10. What are conjugate acid – base pairs? Illustrate using a suitable example.

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



Here NH_4^+ is the conjugate acid of the base NH_3 , since they differ only by a proton.

11. Write the conjugate acid for the following: a) CO_3^{2-} b) H_2O

Ans: a) HCO_3^- b) H_3O^+ [Conjugate acid = given substance + H^+]

12. Write the conjugate base for the following: a) NH_3 b) OH^-

Ans: a) NH_2^- b) H_2O [Conjugate base = Given substance – H^+]

13. State Lewis concept of acids and bases.

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

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.

14. Define ionisation constant or ionic product of water.

Ans: 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. i.e. $K_w = [\text{H}^+][\text{OH}^-]$

15. Define p^{H} .

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

16. Define common ion effect.

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

17. What is Buffer solution?

Ans: 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.

18. Write one example each for acidic buffer and basic buffer.

Ans: Acidic buffer – Equimolar mixture of acetic acid and sodium acetate around a p^{H} 4.75.

Basic buffer – Equimolar mixture of NH_4OH and NH_4Cl around a p^{H} 9.25

19. Define solubility product.

Ans: It is defined as the product of the molar concentration of ions of a sparingly soluble salt or in a saturated solution.

20. Predict the nature of aqueous solutions of the following: (i) Sodium acetate (ii) NaCl (iii) NH_4Cl (iv) K_2CO_3

Ans: (i) Sodium acetate (CH_3COONa) is a salt of weak acid, acetic acid and strong base NaOH . So its aqueous solution is **basic**.

(ii) NaCl is a salt of strong acid HCl and strong base NaOH . So its aqueous solution is **neutral**.

(iii) NH_4Cl is a salt of strong acid HCl and weak base NH_4OH . So its aqueous solution is **acidic**.

(iv) K_2CO_3 is a salt of weak acid carbonic acid and strong base KOH . So its aqueous solution is **basic**.

21. If the concentration of hydrogen ion in a soft drink is $3 \times 10^{-3} \text{ M}$, calculate its pH . Identify whether the solution is acidic or basic.

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

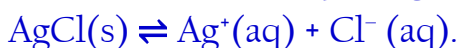
$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

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

Since the $\text{pH} < 7$, the solution is acidic.

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

Ans: Let S be the solubility of AgCl .



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

$$= S \times S = S^2$$

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

REDOX REACTIONS

1. State the electronic concept of oxidation and reduction.

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

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

Ans: 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. Calculate the oxidation number of (i) Cr in $\text{K}_2\text{Cr}_2\text{O}_7$ (ii) P in $\text{H}_2\text{P}_2\text{O}_5$ (iii) S in H_2SO_4 (iv) Mn in KMnO_4

Ans: (i) +6 (ii) +4 (iii) +6 (iv) +7

4. What are Stock notations? Using Stock notation, represent the following compounds:

(i) HAuCl_4 (ii) MnO_2 (iii) FeO (iv) TiO_2

Ans: (i) HAu(III)Cl_4 (ii) Mn(IV)O_2 (iii) Fe(II)O (iv) Ti(IV)O_2

5. Redox reactions are classified into four types. Describe any three of them with suitable examples.

Ans: Combination reactions: A combination reaction may be denoted as $\text{A} + \text{B} \rightarrow \text{C}$

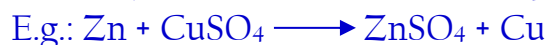
Here either A or B or both A and B must be in the elemental form.

E.g. $\text{C(s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$

Decomposition reactions: Decomposition reactions are the opposite of combination reactions. It involves the breakdown of a compound into two or more components, in which at least one must be in the elemental state. It may be denoted as: $C \rightarrow A + B$.



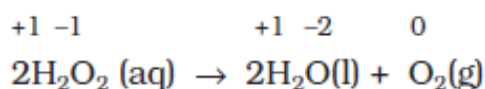
Displacement reactions: Here an ion (or an atom) in a compound is replaced by an ion (or an atom) of another element. It may be denoted as: $X + \text{YZ} \rightarrow \text{XZ} + \text{Y}$



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

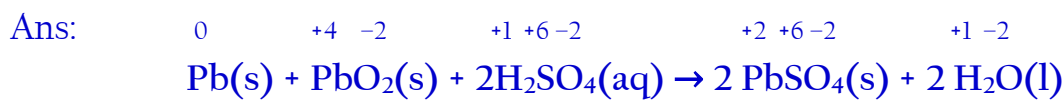
Ans: Disproportionation reaction is a special type of redox reaction in which an element in one oxidation state is simultaneously oxidised and reduced.

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 .

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



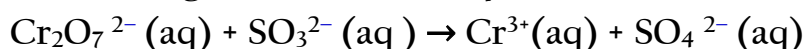
Substance oxidised: Pb

Substance reduced: Pb in PbO_2

Oxidising agent: Pb in PbO_2

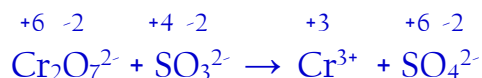
Reducing agent: Pb

8. Balance the following Redox reaction by oxidation number method (Acid medium):



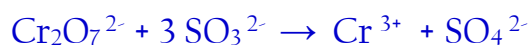
Ans: Step 1: The skeletal equation is: $\text{Cr}_2\text{O}_7^{2-} + \text{SO}_3^{2-} \rightarrow \text{Cr}^{3+} + \text{SO}_4^{2-}$

Step 2: Assign oxidation number each element and identify the elements undergoing change in oxidation number.

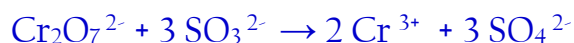


Here the oxidation numbers of Cr and S are changed.

Step 3: Calculate the change in oxidation number and make them equal by multiplying with suitable number. Here the oxidation number of Cr is decreased by 3 and that of S is increased by 2. In order to equate them multiply SO_3^{2-} by 3 (since there are 2 Cr atoms on LHS).



Step 4: Now balance all the atoms except Oxygen and Hydrogen



Step 5: Now balance the ionic charges on both sides. Here the net ionic charge on LHS is -8 and on RHS is 0. To equate them add 8H^+ on LHS, since the reaction takes place in acidic medium.

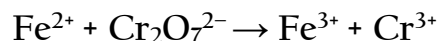


Step 6: Now balance hydrogen atoms by adding sufficient number of H_2O molecules. Here add 4 H_2O molecules on RHS.



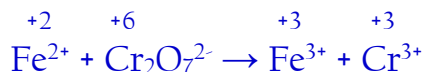
Now the equation is balanced.

9. Balance the following redox equation in acidic medium by half reaction method :



Ans:

Step-1: Assign the oxidation number of each element and find out the substance oxidized and reduced.



Here Fe is oxidized and Cr is reduced.

Step-2: Separate the equation into 2 half reactions -oxidation half reaction and reduction half reaction.



Step-3: Balance the atoms other than O and H in each half reaction individually.



Step-4: Now balance O and H atoms. Add H_2O to balance O atoms and H^+ to balance H atoms since the reaction occurs in acidic medium.



Step-5: Now balance the ionic charges. For this add electrons to one side of the half reaction.



Step-6: Now add the two half reactions after equating the electrons.



Now the equation is balanced.

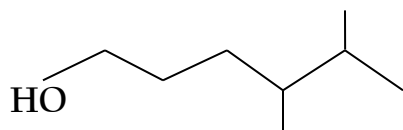
10. When CuSO_4 solution stored in iron vessel, the blue colour changes to pale green. Do you agree with it? Justify.

Ans: Yes. Due to greater electropositivity of iron, it can displace copper from CuSO_4 solution and form FeSO_4 . So, the blue colour changes to pale green.



ORGANIC CHEMISTRY – SOME BASIC PRINCIPLES AND TECHNIQUES

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

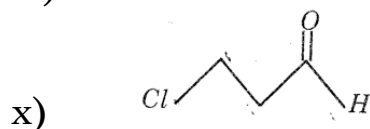
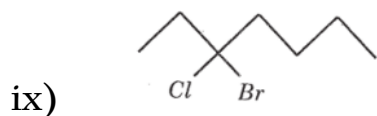
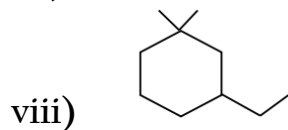
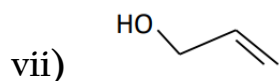
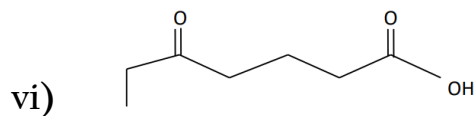
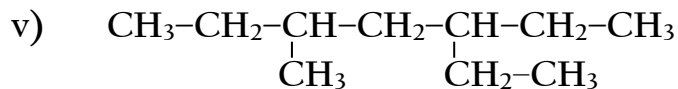
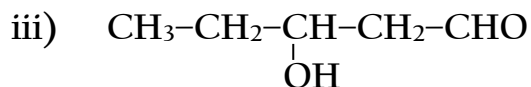


Write its condensed formula and give the IUPAC name.

Ans: Condensed formula is $\text{HO}-\text{CH}_2-\text{CH}_2-\text{CH}_2-\overset{\text{CH}_3}{\underset{\text{CH}_3}{\text{CH}}}-\text{CH}-\text{CH}_3$

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

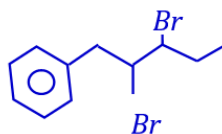
iii) 3-Hydroxypentan-1-al (3-Hydroxypentanal)

- iv) Hept-2-en-5-yne
- v) 3-Ethyl-5-methylheptane
- vi) 5-Oxoheptan-1-oic acid (5-Oxoheptanoic acid)
- vii) Prop-2-en-1-ol (Prop-2-enol)
- viii) 3-Ethyl-1,1-dimethylcyclohexane
- ix) 3-Bromo-3-chloroheptane
- x) 3-Chloropropan-1-al (3-Chloropropanal)

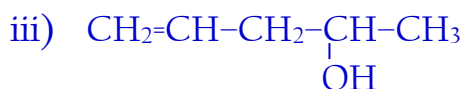
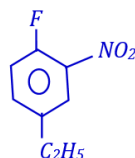
3. Write the structural formula of the following compounds.

- i) 2,3-Dibromo-1-phenylpentane
- ii) 4-Ethyl-1-fluoro-2-nitrobenzene
- iii) Pent-4-en-2-ol
- iv) 6-Hydroxyheptanal
- v) 2,5,6-Trimethyloctane
- vi) Hexane-2,4-dione
- vii) 5-oxohexanoic acid
- viii) 3-Ethyl-2,2-dimethylpentane
- ix) 6-Methyloctan-3-ol
- x) 2-Chloro-4-methylpentane

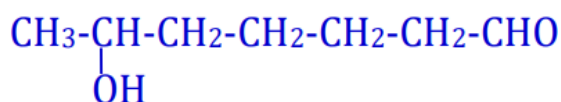
Ans: i)



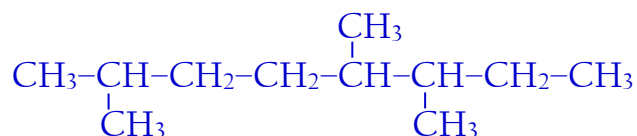
ii)



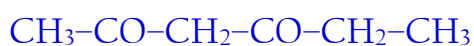
iv)



v)

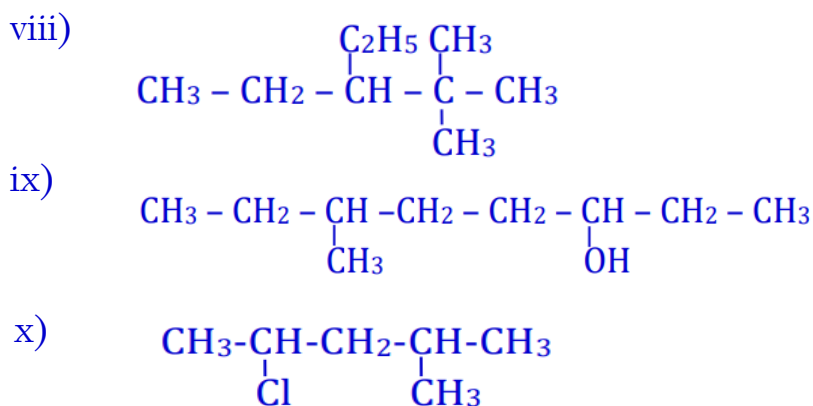


vi)



vii)

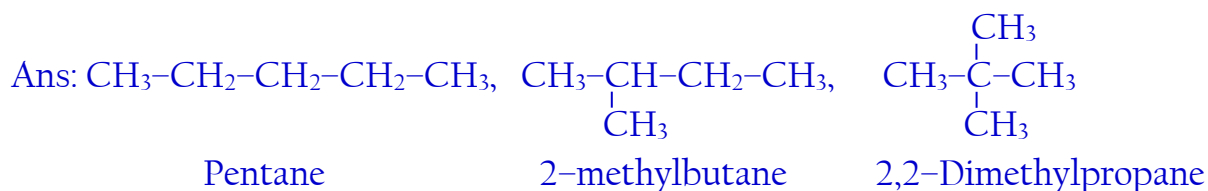




4. What is homologous series?

Ans: It is a series or group of organic compounds in which adjacent members are differed by a $-\text{CH}_2$ group.

5. Write all the possible chain isomers of the compound with molecular formula C_5H_{12} .

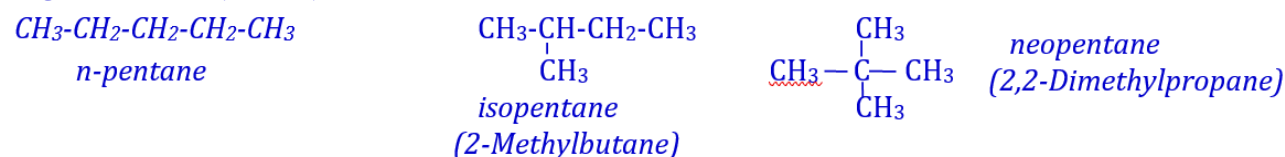


6. Explain with examples, the different types of structural isomerism shown by organic compounds.

Ans: There are mainly four types of structural isomerism:

i) **Chain Isomerism:** Isomers differ in carbon chain or skeleton are called chain isomers and the phenomenon is called chain isomerism.

E.g.: Pentane (C_5H_{12})



ii) **Position isomerism:** Isomers which differ in the position of the substituent or side chain are called position isomers and the phenomenon is called position isomerism.

E.g. : Alcohol with molecular formula $\text{C}_4\text{H}_{10}\text{O}$ may be 1-butanol or 2-butanol



iii) **Functional group isomerism:** Isomers which differ in the functional group are called functional group isomers and the phenomenon is called functional group isomerism. This isomerism is shown by alcohols and ethers and aldehydes and ketones.

E.g. compound with the molecular formula $\text{C}_2\text{H}_6\text{O}$ may be an alcohol ethanol ($\text{CH}_3 - \text{CH}_2\text{OH}$) or an ether methoxymethane ($\text{CH}_3 - \text{O} - \text{CH}_3$).

iv) **Metamerism:** Isomers which differ in the carbon chain (alkyl groups) around the functional group are called metamers and the phenomenon is called metamerism. It is commonly shown by ethers.

E.g.: Ether with formula $C_5H_{12}O$ may be methoxybutane ($CH_3-O-CH_2-CH_2-CH_2-CH_3$) or ethoxypropane ($CH_3-CH_2-O-CH_2-CH_2-CH_3$).

7. Differentiate homolytic cleavage from heterolytic cleavage of covalent bonds.

Ans:

Homolytic bond cleavage (Homolysis)	Heterolytic bond cleavage (Heterolysis)
It is a type of bond fission in which each of the bonded atoms gets one of the electrons of the shared pair.	Here, the bond breaks in such a manner that the shared pair of electrons remains with one of the fragments.
Movement of a single electron occurs.	Movement of a pair of electrons occurs.
The species formed as a result of homolysis is called free radical.	The species formed as a result of heterolysis, may be carbocations or carbanions.

8. Explain the nucleophiles and electrophiles with suitable examples.

Ans: **Nucleophiles:** It is a reagent that brings an electron pair. Or, nucleophiles are electron rich species attack at electron deficient centre.

Example: OH^- , CN^- , NO_2^- , Cl^- , Br^- , I^- , H_2O , NH_3 , $R-NH_2$ etc.

Electrophiles: It is a reagent that takes away an electron pair. Or, electrophiles are electron deficient species attack at electron rich centre.

Example for electrophiles are carbocations (R^+), $-CHO$, $>CO$ etc.

9. Explain the different types electron displacement effects in covalent bonds.

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 it. It is a permanent effect.

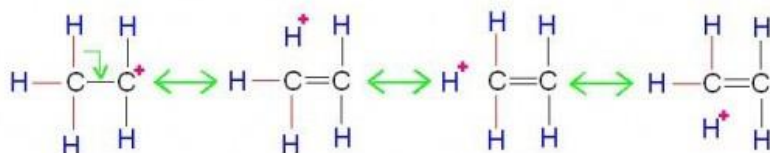
Electromeric effect (E effect): It is the shifting of π -electrons of a multiple bond in presence of an attacking reagent. It is a temporary effect.

Resonance Effect (R effect): It is defined as 'the polarity 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.

10. Hyper conjugation is a general stabilizing interaction. Write the hyper-conjugative structures of $CH_3-CH_2^+$ (ethyl cation)

Ans:



11. Arrange the following carbocations in the increasing order of stability: $(CH_3)_2CH^+$, CH_3^+ , $(CH_3)_3C^+$, $CH_3-CH_2^+$

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

12. 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: $-\text{NO}_2$, $-\text{COOH}$

13. Briefly describe the principles of the following techniques, taking an example in each case. a) Distillation b) Steam distillation c) Distillation under reduced pressure d) Paper chromatography

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. A mixture of chloroform and aniline are separated by distillation.

b) **Steam distillation:** This method is used to separate substances which are steam volatile and are immiscible with water. In steam distillation, steam 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. Aniline is separated from aniline–water mixture by this method.

c) **Distillation under reduced pressure:** This method is used to purify liquids having very high boiling points and those, which decompose at or below their boiling points. These liquids can boil at a temperature below its normal boiling points by reducing the pressure on their surface. Glycerol can be separated from spent–lye in soap industry by using this technique.

d) **Paper chromatography:** It is a type of partition chromatography. Here a strip of chromatography paper, spotted at the base with the solution of the mixture, is suspended in a suitable solvent or a mixture of solvents. The paper selectively retains different components according to their differing partition in the two phases.

14. 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)	Blue or green coloration or precipitate (ppt)	Presence of nitrogen

	solution. Heated to boiling, cooled and acidified with dil. H_2SO_4 .		
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

15. Give the principle of estimation of nitrogen by Dumas method.

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 percentage of Nitrogen.

16. 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,

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

17. On complete combustion, 0.246g of an organic compound gave 0.198g of CO_2 and 0.1014g of H_2O . Determine the percentage composition of carbon and hydrogen in the compound.

Ans: Here, mass of the organic compound (m) = 0.246 g, mass of CO_2 (m_2) = 0.198 g, mass of H_2O (m_1) = 0.1014 g

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

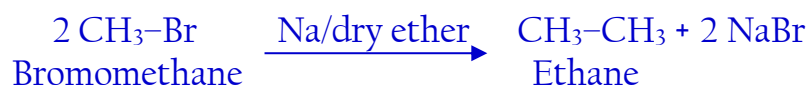
$$\text{Percentage of carbon} = \frac{12 \times m_2 \times 100}{44 \times m} \% = \frac{12 \times 0.198 \times 100}{44 \times 0.246} = 21.95\%$$

HYDROCARBONS

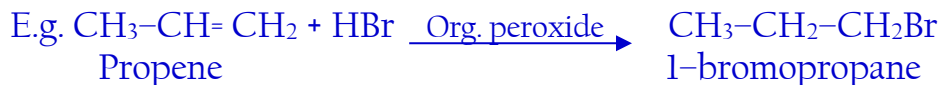
1. Briefly describe the following with suitable chemical equations :

(a) Wurtz reaction (b) Kharash effect

Ans: (a) Wurtz Reaction: Alkyl halides react with metallic sodium in dry ether to form alkanes. This reaction is known as Wurtz reaction.



(b) Kharash effect (Peroxide effect): In the presence of organic peroxide, addition of HBr to unsymmetrical alkenes takes place against Markovnikov rule. This is known as peroxide or Kharash effect.

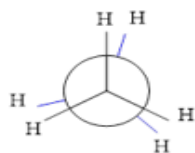


2. What are conformations?

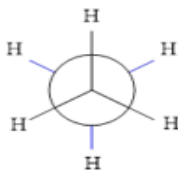
Ans: The different spatial arrangements of atoms arising due to free rotation around a C-C single bond are called conformations or conformers.

3. Draw Newman's projections for the eclipsed and staggered conformations of ethane. Which among these is more stable? Justify.

Ans:



(i) Eclipsed

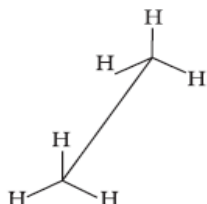


(ii) Staggered

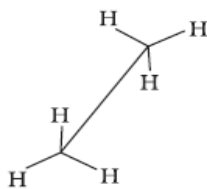
Staggered conformation is stabler than eclipsed form due to minimum repulsive forces between the electron clouds of C-H bonds.

4. Draw the sawhorse projections for eclipsed and staggered forms of an ethane molecule.

Ans:



(i) Eclipsed



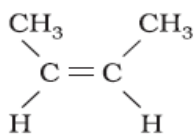
(ii) Staggered

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

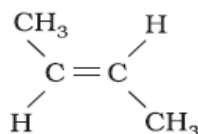
Ans: The isomerism arising due to the difference in the spatial arrangement of atoms around C = C bond is called geometrical isomerism.

There are two types of geometrical isomers – cis isomer and trans isomer.

E.g. 2-butene



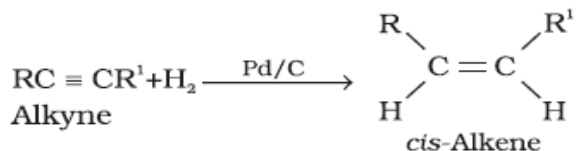
cis-But-2-ene



trans-But-2-ene

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

Ans: Partially deactivated palladised charcoal is known as Lindlar's catalyst. Alkynes on reduction with Lindlar's catalyst, we get cis-alkene.



7. $\text{CH}_3\text{-CH=CH}_2 + \text{HBr} \rightarrow \text{A} + \text{B}$.

(a) Identify A and B.

(b) Which is the major product in the above reaction ?

(c) Name and state the rule which decides the major product in the above reaction.

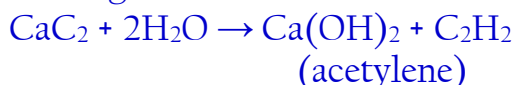
Ans: (a) A is $\text{CH}_3\text{-CHBr-CH}_3$ (2-Bromopropane) and B is $\text{CH}_3\text{-CH}_2\text{-CH}_2\text{Br}$ (1-Bromopropane).

(b) The major product is $\text{CH}_3\text{-CHBr-CH}_3$.

(c) Markownikoff's rule. The rule states that when an unsymmetrical reagent is added to an unsymmetrical alkene, the negative part of the reagent gets attached to the carbon containing lesser number of hydrogen atoms.

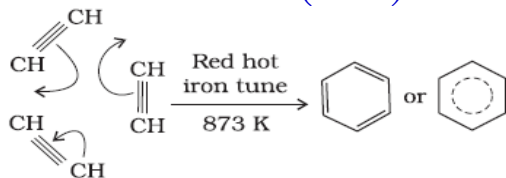
8. Give the preparation of acetylene (C_2H_2) from calcium carbide?

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



9. How will you convert acetylene to benzene?

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



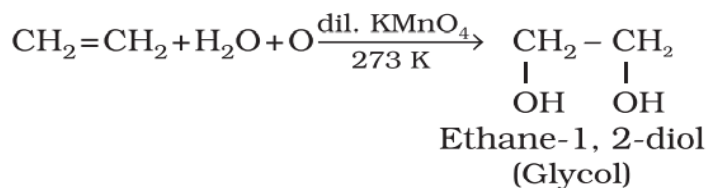
10. State Hickel's $(4n + 2)$ π electron rule.

Ans: According to this rule, "cyclic, planar, conjugated systems containing $(4n+2)$ π electrons are aromatic".

11. Name a polymer that can be used as electrodes in batteries. Ans: Polyacetylene

12. What is Baeyer's reagent? Write the chemical equation of its reaction with ethylene.

Ans: Baeyer's reagent is cold dilute aqueous KMnO_4 .



13. Give the chemical equation for the conversion of hexane to benzene. Write the name of the process.

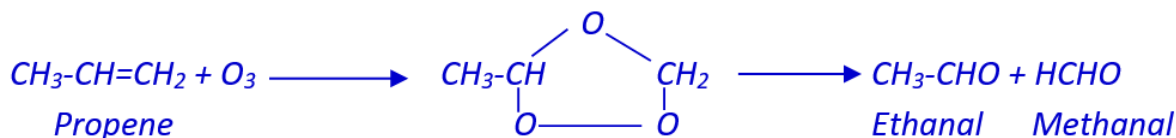


The process is known as Aromatisation.

14. What is ozonolysis? Write the names of the products obtained when propene undergoes ozonolysis?

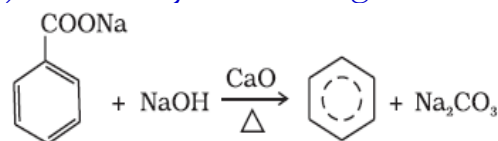
Ans: Alkenes add ozone to form an ozonide which on hydrolysis in presence of Zn to form aldehydes or ketones. This reaction is known as ozonolysis.

Propene on ozonolysis gives ethanal (acetaldehyde) and methanal (formaldehyde).

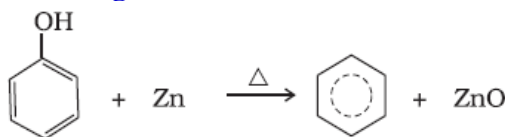


15. How will you convert (i) sodium benzoate to benzene (ii) Phenol to Benzene (iii) Benzene to Toluene

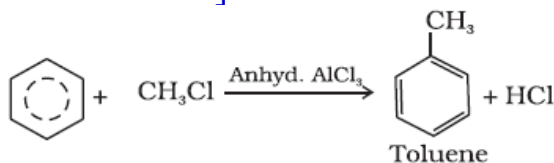
Ans: (i) Decarboxylation using sodalime.



(ii) By heating with zinc dust.



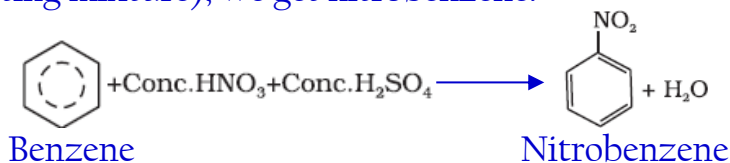
(iii) By treating with methyl chloride in presence of anhydrous AlCl_3 . [Friedel-Craft alkylation reaction]



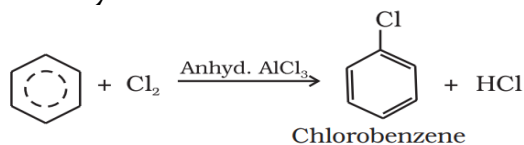
16. What are electrophilic substitution reactions? Give examples.

Ans: These are reactions in which a weak electrophile is replaced by a strong electrophile.

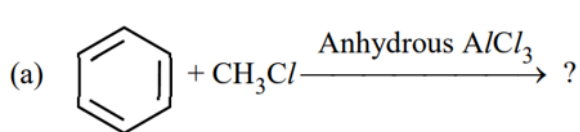
E.g.: Nitration: When benzene is heated with a mixture of conc. HNO_3 and conc. H_2SO_4 (nitrating mixture), we get nitrobenzene.



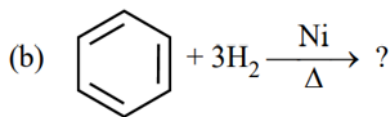
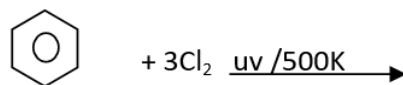
17. How will you convert benzene to chlorobenzene?



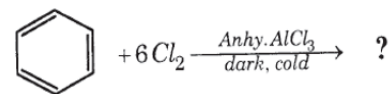
18. Complete the following reactions:



(c)

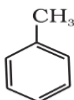


(d)



Ans:

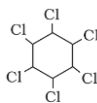
(a) Toluene, OR



(b) Cyclohexane, OR



(c) Benzenehexachloride, OR



(d) Hexachlorobenzene, OR

