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CLASS XI

CHEMISTRY

Chemistry +1

- 1 **Some basic concepts of Chemistry**
- 2 **Structure of Atom**
- 3 **Classification of elements and periodicity in properties**
- 4 **Chemical Bonding and Molecular Structure**
- 5 **Thermodynamics**
- 6 **Chemical Equilibrium**
- 7 **Redox Reactions**
- 8 **Organic Chemistry - Some Basic Principles and
Techniques**
- 9 **Hydrocarbons**

ആമുഖം

ഹയർസെക്കൻഡറി തലത്തിൽ ഒന്നാംവർഷ രസതന്ത്രപഠനം ലക്ഷ്യകരിക്കുന്നതിന് വേണ്ടിയുള്ള സമഗ്രവും ലളിതവുമായുള്ള ഒരു പഠനസഹായിയാണ് ഇത്. ഇതിൽ ഓരോ യൂണിറ്റിലേയും ഏറ്റവും പ്രധാന ആശയങ്ങളും അതുമായി ബന്ധപ്പെട്ട ചോദ്യങ്ങളും ഉത്തരങ്ങളും ഈ ബുക്ക്ലെറ്റിൽ ഉൾപ്പെടുത്തിയിട്ടുണ്ട്.

സ്റ്റേറ്റ് പ്രോജക്ട് ഡയറക്ടർ

Unit 1

Some basic concepts of Chemistry

1. Complete the table

Measurements in gram	Scientific notation	No of significant figures	Round upto three significant figure
34.216	3.4216×10^1	5	34.2
10.4107			
0.04597			
0.06200			

2. Molar mass of $C_6H_{12}O_6$ is 180 g/mol ($6 \times 12 + 12 \times 1 + 6 \times 16 = 180$)

Calculate molar mass of the following

Atomic mass of C = 12u, O = 16u, S = 32u, Na = 23u, Cl = 35.5u

- (a) Na_2CO_3
- (b) H_2SO_4
- (c) HCl
- (d) $NaOH$
- (e) $C_2H_4O_2$
- (f) H_2O

3. Complete the table

Compound	Formula mass	No. of moles	mass of compound	No. of molecules	No. of atoms	Volume of gas in Litre at 273K and 1 atm
$CH_4(g)$		5	5×16	$5 \times 6.022 \times 10^{23}$	$5 \times 6.022 \times 10^{23} \times 5$	5×22.4
$NH_3(g)$			170			
$H_2O(l)$				12.044×10^{23}		

4. The following data are obtained when N_2 and O_2 react together to form different compounds.

SI No.	Mass of N_2 in gram	Mass of O_2 in gram
1	14	16
2	14	32
3	28	32
4	28	80

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

5. H_2 and O_2 react to form H_2O according to the equation $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ complete the table.

Number of moles of H_2	Number of moles of O_2	Limiting reagent	Exces of re-agent in moles	Number of moles of H_2O formed
4	1	O_2	$4-2=2$	$1 \times 2=2$
6	2			
2	4			
4	3			
2	1			

6. Calculate the mass percentage of carbon, hydrogen & oxygen in the following compounds.

Compound	Molecular mass	Percentage composition of carbon	Percentage composition of Hydrogen	Percentage composition of Oxygen
		$= \frac{\text{total mass of carbon} \times 100}{\text{molecular mass}}$	$= \frac{\text{total mass of Hydrogen} \times 100}{\text{molecular mass}}$	$= \frac{\text{total mass of Oxygen} \times 100}{\text{molecular mass}}$
CH_4O	32	$= \frac{(12 \times 1) \times 100}{32}$ $= 37.5\%$	$= \frac{(4 \times 1) \times 100}{32}$ $= 12.5\%$	$= \frac{(16 \times 1) \times 100}{32}$ $= 50.00\%$
$C_2H_4O_2$				
$C_6H_{12}O_6$				

7. Determine the empirical formula and molecular formula of organic compound (molecular mass = 78u), which has 92.30% carbon and 7.69% dioxygen

Component elements	atomic mass (u)	Percentage composition	Simple ratio	whole no ratio
		atomic mass		
C				
H				

empirical formula =

molecular formula = n (empirical formula)

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

8. Determine the empirical formula and molecular formula of organic compound with molar mass 62g/mol, which has 38.70% carbon, 6.45% hydrogen and 51.61% oxygen by mass.

9. Match the following

A	B
1. Mass per cent	A. It is defined as the number of moles of solute in one litre of solution
2. Mole fraction	B. It is defined as the number of moles of solute present in one Kg of solvent.
3. Molality	C. Mass of solute present in 100g of solution.
4. Molarity	D. It is the ratio of number of moles of a particular component to the total number of moles of all the components.

10. Calculate mass percent of solute, mole fraction of solute and molality of solution.

(a) 80 g of NaOH in 500 g of water

(b) 146 g of HCl in 600 g of water

(c) 212 g of Na_2CO_3 in 900 g of water

Qn	mass of solute in gram	mass of solvent in gram	mass Percent of solute $= \frac{\text{mass of solute} \times 100}{\text{mass of solution}}$	number of moles of solvent $n_A = \frac{\text{mass of solvent}}{\text{molar mass}}$	number of moles of solute $n_B = \frac{\text{mass of solute}}{\text{molar mass}}$	mole fraction of solute $x_B = \frac{n_B}{n_A + n_B}$	molality of solution (m) $= \frac{n_B \times 1000}{\text{mass of solvent in gram (mole kg}^{-1}\text{)}}$
a	80	500	$= \frac{80 \times 100}{(500 + 80)} = 13.79$	$n_A = \frac{500}{18} = 27.77$	$n_B = \frac{80}{40} = 2$	$x_B = \frac{2}{2 + 27.77} = 0.067$	$= \frac{2 \times 1000}{500} = 4 \text{ m}$
b							
c							

Solutions

1.

Measurements in gram	Scientific notation	No of significant figures	Round upto three significant figures
34.216	3.4216×10^1	5	34.2
10.4107	1.04107×10^1	6	10.4
0.04597	4.597×10^{-2}	4	0.0460
0.06200	6.200×10^{-2}	4	0.0620

2.

(a) 106 g/mol

(b) 98 g/mol

(c) 36.5 g/mol

(d) 40 g/mol

(e) 60 g/mol

(f) 18 g/mol

3.

Compound	Formula mass	No. of moles	mass of compound	No. of molecules	No. of atoms	Volume of gas in Litre at 273K and 1 atm
$CH_4(g)$	16 u	5	5x16	$5 \times 6.022 \times 10^{23}$	$5 \times 6.022 \times 10^{23} \times 5$	5x22.4
$NH_3(g)$	17 u	10	170	$10 \times 6.022 \times 10^{23}$	$10 \times 6.022 \times 10^{23} \times 4$	10x22.4
$H_2O(l)$	18 u	2	36 g	12.044×10^{23}	$12.044 \times 10^{23} \times 3$	-

4. Law of multiple proportion

5.

Number of moles of H ₂	Number of moles of O ₂	Limiting reagent	Exces of re-agent in moles	Number of moles of H ₂ O formed`
4	1	O ₂	4-2=2	1x2=2
6	2	O ₂	6-4=2	2x2=4
2	4	H ₂	4-1=3	2
4	3	H ₂	3-2=1	4
2	1	-	-	2

6.

Compound	Molecular mass	Percentage composition of carbon	Percentage composition of Hydrogen	Percentage composition of Oxygen
		$\frac{\text{total mass of carbon} \times 100}{\text{molecular mass}}$	$\frac{\text{total mass of Hydrogen} \times 100}{\text{molecular mass}}$	$\frac{\text{total mass of Oxygen} \times 100}{\text{molecular mass}}$
CH ₄ O	32	$\frac{(12 \times 1) \times 100}{32} = 37.5\%$	$\frac{(4 \times 1) \times 100}{32} = 12.5\%$	$\frac{(16 \times 1) \times 100}{32} = 50.00\%$
C ₂ H ₄ O ₂	60	$\frac{24}{60} \times 100 = 40\%$	$\frac{4}{60} \times 100 = 6.67\%$	$\frac{32}{60} \times 100 = 53.33\%$
C ₆ H ₁₂ O ₆	180	$\frac{72}{180} \times 100 = 40\%$	$\frac{12}{180} \times 100 = 6.67\%$	$\frac{96}{180} \times 100 = 53.33\%$

7.

Component elements	atomic mass (u)	Percentage composition	Simple ratio	whole no ratio
		atomic mass		
C	12	$\frac{92.30}{12} = 7.69$	$\frac{7.69}{7.69} = 1$	1
H	1	$\frac{7.69}{1} = 7.69$	$\frac{7.69}{7.69} = 1$	1

empirical formula = CH

molecular formula = n (empirical formula)

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

molecular formula = 6 (CH)



8.

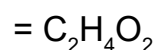
Component elements	atomic mass (u)	Percentage composition	Simple ratio	whole no ratio
		atomic mass		
C	12	$\frac{38.7}{12} = 3.16$	1	1
H	1	$\frac{6.45}{1} = 6.45$	2	2
O	16	$\frac{51.61}{16} = 3.2$	1	1

empirical formula = CH₂O

molecular formula = n (empirical formula)

$$n = \frac{62}{30} = 2.06 \approx 2$$

molecular formula = 2 (CH₂O)



9.

A	B
1. Mass per cent	C. Mass of solute present in 100g of solution.
2. Mole fraction	D. It is the ratio of number of moles of a particular component to the total number of moles of all the components.
3. Molality	B. It is defined as the number of moles of solute present in one Kg of solvent.
4. Molarity	A. It is defined as the number of moles of solute in one litre of solution

10.

Qn	mass of solute in gram	mass of solvent in gram	mass Percent of solute $= \frac{\text{mass of solute} \times 100}{\text{mass of solution}}$	number of moles of solvent $n_A = \frac{\text{mass of solvent}}{\text{molar mass}}$	number of moles of solute $n_B = \frac{\text{mass of solute}}{\text{molar mass}}$	mole fraction of solute $x_B = \frac{n_B}{n_A + n_B}$	molality of solution (m) $= \frac{n_B \times 1000}{\text{mass of solvent in gram (mole Kg}^{-1}\text{)}}$
a	80	500	$= \frac{80 \times 100}{(500 + 80)} = 13.79$	$n_A = \frac{500}{18} = 27.77$	$n_B = \frac{80}{40} = 2$	$x_B = \frac{2}{2 + 27.77} = 0.067$	$= \frac{2 \times 1000}{500} = 4$
b	146	600	$= \frac{146 \times 100}{(146 + 600)} = 19.57$	$n_A = \frac{600}{18} = 33.33$	$n_B = \frac{146}{36.5} = 4$	$x_B = \frac{4}{4 + 33.33} = 0.1072$	$= \frac{4}{600} \times 1000 = 6.67$
c	212	900	$= \frac{212 \times 100}{(212 + 900)} = 19.06$	$n_A = \frac{900}{18} = 50$	$n_B = \frac{212}{106} = 2$	$x_B = \frac{2}{2 + 50} = 0.0385$	$= \frac{2 \times 1000}{900} = 2.22$

Unit 2

Structure of Atom

1. Cathode rays (electrons) were discovered through discharge tube experiments. Which of the following results of this experiment is/are INCORRECT?

- (a) Cathode rays start from anode and move towards cathode.
- (b) These rays are invisible but their behavior can be observed with the help of fluorescent or phosphorescent material.
- (c) In the absence of electrical or magnetic field, these rays travel in straight lines
- (d) The characteristics of cathode rays depend upon the material of the electrodes and the nature of the gas present in the cathode ray tube.

2. Electrical discharge carried out in the modified cathode ray tube led to the discovery of positively charged particles called canal rays. Identify the CORRECT statement (s) about canal rays.

- a) The properties of canal rays depend on the nature of the gas present in the cathode ray tube
- b) The charge to mass ratio of these particles is independent of the nature of the gas present in the cathode ray tube.
- c) These rays are deviated towards the positive plate of electrical field
- d) Canal rays are positively charged gaseous ions

3. Rutherford proposed the nuclear model of atom based on α -particle scattering experiment. The observations obtained from the experiment are given in column A. Write the conclusions of each observation in Column B.

Observations (A)	Conclusions (B)
(i) Most of the α -particles passed through the gold foil undeflected.	
(ii) A small fraction of the α -particles was deflected by small angles.	
(i) A very few α -particles bounced back	

4. What are the important postulates of Rutherford's nuclear model of atom? Also write the demerits of this model.

9. In order to calculate the wave number of spectral lines in the hydrogen spectrum,

Johannes Rydberg proposed the equations $\bar{\nu} = \frac{1}{\lambda} = 109677 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \text{ cm}^{-1}$ calcu-

late the wave length of the light emitted when the electron in hydrogen atom undergoes transition from an energy level with $n = 4$ to an energy level with $n=2$?

10. The radius of the first stationary state of hydrogen atom is called Bohr radius, which is equal to ---- pm.

11. Calculate the energy of the electron in the ground state of hydrogen atom.

12. Write any 3 postulates of Bohr's model of hydrogen atom. Also write any 2 limitations of this model

13 What will be the wavelength of an electron moving with a velocity of 10m/s ? [mass of electron = $9.1 \times 10^{-31} \text{ kg}$]

14. It is impossible to determine simultaneously, the exact position and exact momentum of an electron. Identify the principle related to this statement. Write the mathematical expression of the above statement.

15. What is the maximum possible number of emission lines in the line spectrum of hydrogen when the excited electron in $n = 4$ drop to the ground state?

- a) 10 b) 6 c) 12 d) 15

16. Calculate the kinetic energy of the ejected electron when ultraviolet radiation of frequency $1.6 \times 10^{15} \text{ Hz}$ strikes the surface of sodium metal. [Threshold frequency of sodium metal is $4.4 \times 10^{14} \text{ Hz}$]

17. Which of the following statement does not form a part of Bohr model of hydrogen atom?

- The electron in the hydrogen atom can move around the nucleus in a circular path of fixed radius and energy called orbit or energy level.
- The energy of an electron in the orbit does not change with time.
- The angular momentum of an electron is quantised.
- When an electron jumps from one energy level to another, the energy change takes place in a continuous manner.

18. Which of the following is not a characteristic of Planck's quantum theory of radiation?

- Atoms and molecules could emit or absorb energy not in a continuous manner but in the form of small packet of energy called quantum.
- The energy of a quantum of radiation is proportional to its frequency.
- The energy of a quantum of radiation is proportional to its intensity.

19. Heisenberg's uncertainty principle and the wave nature of matter are significant to which of the following and why?

- A moving electron
- A moving cricket ball

20. Fill in the blanks using the appropriate quantum number given in the list below:

[Principal quantum number, Azimuthal quantum number, Magnetic quantum number, spin quantum number]

- The quantum number which is related to the spin of electron is
- gives the three-dimensional shape of the orbital.
- The quantum number which determines the size and most of the energy of an orbital is
- defines the spatial orientation of orbitals with respect to standard set of co-ordinate axes.

21. Complete the following table:

Orbital	n	l
3s	3	0
2p
.....	4	0
.....	5	3
4d

22. Which of the following sets of quantum numbers are not possible? Give reason.

- $n = 0, l = 0, m_l = 0, m_s = +\frac{1}{2}$
- $n = 1, l = 0, m_l = 0, m_s = -\frac{1}{2}$
- $n = 3, l = 3, m_l = -3, m_s = +\frac{1}{2}$
- $n = 2, l = 1, m_l = 0, m_s = -\frac{1}{2}$

23. Match the following:

Orbital	Shape
s- orbital	Double dumb-bell
p-orbital	Diffused shape
d-orbital	Spherical shape
f-orbital	Dumb-bell shape

24. I am an orbital close to the nucleus. My quantum numbers are $n = 1$, $l = 0$, $m_l = 0$ and $m_s = +\frac{1}{2}$. Identify me. Also draw my boundary surface diagram.

25. Complete the following table:

Orbital	No. of angular nodes	No. of radial nodes
2p	1	0
3s		
4d		
5f		
3p		

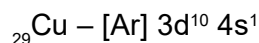
26. Find out the rule or principle relating to the following statements:

- No two electrons in an atom can have the same set of four quantum numbers.
- The pairing of electrons will start in the p, d and f orbitals with the entry of 4th, 6th and 8th electron respectively, in these subshells.
- The valence electron of potassium is found in 4s orbital and not in 3d orbital.

27. Find the odd one based on the electronic configuration. Give reason.

Titanium ($Z = 22$), Chromium ($Z = 24$), Manganese ($Z = 25$), Cobalt ($Z = 27$)

28. Identify the correct electronic configuration of copper ($Z = 29$) from the following:



Justify your answer.

SOLUTIONS

1. The incorrect statements are:

- a) Cathode rays start from anode and move towards cathode. [*Correct statement: Cathode rays start from cathode and move towards anode*]
- d) The characteristics of cathode rays depend upon the material of electrodes and the nature of the gas present in the cathode ray tube. [*Correct statement: The characteristics of cathode rays do not depend upon the material of electrodes and the nature of the gas present in the cathode ray tube*]

2. The correct statements are:

- a) The properties of canal rays depend upon the nature of gas present in the cathode ray tube.
- d) Canal rays are positively charged gaseous ions.

3.

Observations (A)	Conclusions (B)
Most of the α - particles passed through the gold foil undeflected.	Most of the space in the atom is empty.
A small fraction of the α -particles was deflected by small angles.	The positive charge of the atom is concentrated in a very small volume at the centre called nucleus.
A very few α - particles bounced back.	The volume occupied by the nucleus is negligibly small as compared to the total volume of the atom.

4. Postulates of Rutherford's nuclear model of atom are:

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

Demerits of Rutherford atom model are:

- (i) He could not explain the stability of an atom.
- (ii) He says nothing about the electronic structure of atoms. i.e., how the electrons are distributed around the nucleus and what are the energies of these electrons.

5.

SI No.	Species	P	e	n
(i)	${}^{13}_6\text{C}$	6	6	7
(ii)	${}^{32}_{16}\text{S}^{2-}$	16	18	16
(iii)	${}^3_1\text{H}$	1	1	2

6. Diffraction and interference can be explained by the wave nature of electromagnetic radiation.

7. The incorrect statement is:

(b) The number of electrons ejected is proportional to the frequency of light beam used. [Correct statement: The number of electrons ejected is proportional to the intensity or brightness of light]

8.

Series	Spectral Region
Lyman	<u>Ultraviolet</u>
<u>Balmer</u>	Visible
Paschen	<u>Infrared</u>
Brackett	<u>Infrared</u>
Pfund	Infrared

9. Here $n_1 = 2$ and $n_2 = 4$

$$\begin{aligned}\bar{\nu} &= 109677 (1/n_1^2 - 1/n_2^2) \text{ cm}^{-1} \\ &= 109677 (1/2^2 - 1/4^2) \text{ cm}^{-1} \\ &= 109677 \times 3/16 \text{ cm}^{-1} = \mathbf{20564.4 \text{ cm}^{-1}}\end{aligned}$$

10. Bohr radius is 52.9 pm

11. The energy of electron in the n^{th} orbit of Hydrogen atom is given by the expression:

$$E_n = -R_H (1/n^2), \text{ where } R_H = 2.18 \times 10^{-18} \text{ J}$$

For ground state, $n = 1$

$$\text{So, } E_1 = -2.18 \times 10^{-18} \text{ J}$$

12. The important postulates of Bohr model of hydrogen atom are:

- (i) The electron in the hydrogen atom can move around the nucleus in circular paths of fixed radius and energy. These paths are called orbits or stationary states or allowed energy states.
- (ii) The energy of an electron in an orbit does not change with time. However, when an electron absorbs energy, it will move away from the nucleus.
- (iii) The angular momentum of an electron is quantized. i.e. it is an integral multiple of $h/2\pi$

13. According to de Broglie's equation, $\lambda = \frac{h}{mv}$

Here $m = 9.1 \times 10^{-31} \text{ kg}$, $v = 10 \text{ m s}^{-1}$ and $h = 6.626 \times 10^{-34} \text{ J s}$

$$\text{So, } \lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}}{9.1 \times 10^{-31} \times 10} = 7.28 \times 10^{-5} \text{ kg}$$

14. Heisenberg's uncertainty principle.

The mathematical expression is: $\Delta x \cdot \Delta p \geq h/4\pi$

OR, $\Delta x \cdot m \Delta v \geq h/4\pi$

15. Maximum number of emission lines = $\frac{n(n+1)}{2}$

Here $n = 4$. So, the number of emission lines = 10

16. K.E of emitted electron = $h\nu - h\nu_0 = h(\nu - \nu_0)$
 $= 6.626 \times 10^{-34} (1.6 \times 10^{15} - 4.4 \times 10^{14})$
 $= 1031 \times 10^{-20} \text{ J}$

17. d) When an electron jumps from one energy level to another, the energy change takes place in a continuous manner. [Correct statement: When an electron jumps from one energy level to another, the energy change does not take place in a continuous manner].

18. c) The energy of a quantum of radiation is proportional to its intensity.

19. Heisenberg's uncertainty principle and the wave nature of matter are significant to a moving electron. For macroscopic bodies, since the mass is higher, the product of uncertainty in position and velocity is negligible. So, Heisenberg's uncertainty principle is not significant. Also, the wavelengths associated with ordinary objects are so short (because of their large masses) that their wave properties cannot be detected.

20. a) Spin quantum number
 b) Azimuthal quantum number
 c) Principal quantum number
 d) Magnetic quantum number

21.

Orbital	n	l
3s	3	0
2p	2	1
4s	4	0
5f	5	3
4d	4	2

22. The set of Quantum numbers which are not allowed are:

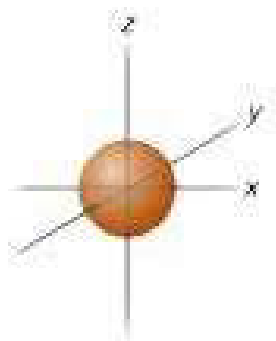
(a) $n = 0, l = 0, m_l = 0, m_s = +\frac{1}{2}$ [n cannot be equal to zero](c) $n = 3, l = 3, m_l = -3, m_s = +\frac{1}{2}$ [the value of l cannot be equal to that of n]

23.

Orbital	Shape
s-orbital	Spherical shape
p-orbital	Dumb-bell shape
d-orbital	Double dumb-bell shape
f-orbital	Diffused shape

24. 1s orbital

Boundary surface diagram of 1s orbital is:



25.

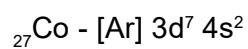
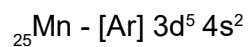
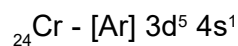
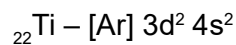
Orbital	No. of angular nodes = l	No. of radial nodes = n - l - 1
2p	1	0
3s	0	2
4d	2	1
5f	3	1
3p	1	1

26. a) Pauli's Exclusion principle

b) Hund's rule

c) Aufbau principle

27. Electronic configuration of:



The electronic configuration of Cr is the odd one. This is because of the extra stability of half-filled electronic configuration.

28. ${}_{29}\text{Cu} - [\text{Ar}] 3d^{10} 4s^1$

This is because of the extra stability of completely filled electronic configuration.

Unit 3

Classification of elements and periodicity in properties

1. Match the following.

	A	B	C
1.	Johann Dobereiner	Mendeleev's periodic law	Properties of elements are periodic functions of their atomic weights
2.	John Alexander Newlands	Modern periodic law	Every eighth element has the properties similar to the first element
3.	Dimitri Mendeleev.	Triads	Physical and Chemical properties of elements are periodic functions of their atomic number.
4.	Henry Moseley	Law of octaves	Properties of middle element were in between those of other two elements

2. Which statement is correct about the merits of Mendeleev's periodictable

- (a) Mendeleev corrected atomic masses of some elements .
- (b) Mendeleev left certain gaps in his periodic table for some of the then undiscovered elements
- (c) Mendeleev gave more importance to similarities in properties than to atomic mass.
- (d) Mendeleev was the first to successfully classify all the known elements at that time.
- (e) all are correct

3. Which statement is correct about the demerits of Mendeleev's periodic table.

- (a) No fixed position could be given to hydrogen
- (b) Lighter elements are placed after heavier elements, contrary to the law.
- (c) Chemically similar elements were placed in different groups and vice versa.
- (d) There were no definite places for isotopes
- (e) All are correct

4. Identify the incorrect statement about the merits of modern periodic table.

- (a) The table is based on atomic numbers of elements.
- (b) It correlates the position of the elements with the electronic configuration clearly.
- (c) The reason for placing isotopes at one place is justified as the classification is on the basis of atomic numbers.

- (d) The Lanthanides and actinides are placed separately at the bottom of the periodic table
- (e) There is nine vertical columns (groups) and seven horizontal columns (Periods)

5. Identify the correct position for the following elements in the periodic table.

- (a) atomic number (Z) = 32
- (b) Z = 53
- (c) Z = 25
- (d) Z = 38
- (e) Z = 17
- (f) Z = 46
- (g) Z = 39 .

Sl No	atomic number (z)	electronic configuration	position in the period (number of outer energy level as 'n' value)	position in the group (Total no. of electron in (n-1) d orbital and nth shell)
(a)	32	$3d^{10} 4s^2 4p^2$	4	$10+4 = 14$
(b)	53			
(c)	25	$3d^5 4s^2$	4	$5 + 2 = 7$
(d)	38			
(e)	17			
(f)	46			
(g)	39 .			

6. Write the names for the following elements according to the IUPAC nomenclature.

- (i) atomic number - 109 (ii) atomic number -119
- (iii) atomic number - 120 (iv) atomic number - 121

Sl No	Atomic Number	IUPAC nomenclature	Symbol
(i)	109	Unnilennium	Uue
(ii)	119		
(iii)	120	Unbinilium	Ubn
(iv)	121		

7. Complete the table.

SI No	General electronic configuration	Block	Name	Position in the periodic table
1	ns^{1-2}			1 st and 2 nd group elements
2			Representative or main block elements	
3		d - Block		
4				Two series of elements each contain 14 elements placed separately at the bottom of the periodic table

8. Complete the table.

SI No	Periodic properties	defenition	Periodic trend down a group	periodic trend across the period from left to right
1	Atomic Radii	Is the distance of the outermost shell of an atom from the centre of the nucleus	moving down the group atomic radii increases	across the period atomic radii decreases
2	Ionization Enthalpy			
3	Electron gain Enthalpy			
4	Electronegativity			

9. Match the following, regarding nature of the oxides

A	B
1. N_2O	(a) Amphoteric
2. BaO	(b) Neutral
3. As_2O_3	(c) Acidic
4. Cl_2O_7	(d) Basic

10. Which is the correct order of the size of the following.

- (a) $Ca^{2+} < Ar < K^+ < Cl^- < S^{2-}$
- (b) $Ca^{2+} < K^+ < Ar < Cl^- < S^{2-}$
- (c) $Ar < Ca^{2+} < K^+ < Cl^- < S^{2-}$
- (d) $Ca^{2+} < K^+ < Ar < S^{2-} < Cl^-$

11. Which of the following shows diagonal relationship.

- (a) Boron and Carbon
- (b) Boron and Gallium
- (c) Boron and Silicon
- (d) Boron and Aluminium

Solutions

1.	A	B	C
1.	Johann Dobereiner	Triads	Properties of middle element were in between those of other two elements
2.	John Alexander Newlands	Law of octaves	Every eighth element has the properties similar to the first element
3.	Dimitri Mendeleev.	Properties of elements are periodic functions of their atomic weights	Mendeleev's periodic law
4.	Henry Moseley	Modern periodic law	Physical and Chemical properties of elements are periodic functions of their atomic number.

2. (e)

3. (e)

4. (e)

5.

Z	group	period
32	14	4
53	17	5
25	7	4
38	2	5
17	17	3
46	10	5
39	3	5

Sl No	atomic number (z)	electronic configuration	position in the period (number of outer energy level as 'n' value)	position in the group (Total no. of electron in (n-1) d orbital and nth shell)
(a)	32	$3d^{10} 4s^2 4p^2$	4	$10+4 = 14$
(b)	53	$4d^{10} 5s^2 5p^2$	5	$10+2+5=17$
(c)	25	$3d^5 4s^2$	4	$5 + 2 = 7$
(d)	38	$5s^2$	5	2
(e)	17	$3s^2 3p^2$	3	17
(f)	46	$4d^8 5s^2 5p^2$	5	$8+2=10$
(g)	39	$4d^1 5s^2$	5	$1+2=3$

6.

SI No	Atomic Number	IUPAC nomenclature	Symbol
(i)	109	Unnilennium	Une
(ii)	119	Ununennium	Uue
(iii)	120	Unbinilium	Ubn
(iv)	121	Unbiunium	Ubu

7.

SI No	General outer electronic configuration	Block	Name	Position in the periodic table
1	ns^{1-2}	s - Block	Alkali metals and alkaline earth metals	1 st and 2 nd group elements
2	$ns^2 np^{1-6}$	p - Block	Representative or main block elements	group 13 to group 18
3	$(n-1)d^{1-10} ns^{0-2}$	d- Block	Transition metals	group 3 to group 12
4	$(n-2)f^{1-14} (n-1)d^{0-1} ns^2$	f - Block	Inner transition elements	Two series of elements each contain 14 elements placed separately at the bottom of the periodic table

8.

SI No	Periodic properties	definition	Periodic trend down a group	periodic trend across the period from left to right
1	Atomic Radii	Is the distance of the outermost shell of an atom from the centre of the nucleus	moving down the group atomic radii increases	across the period atomic radii decreases
2	Ionization Enthalpy	It is the amount of energy required to remove an electron from an isolated gaseous atom in its ground state	decreases	increases
3	Electron gain Enthalpy	It is the amount of energy released when an electron is added to a neutral gaseous atom to convert it into a negative ion.	decreases	increases
4	Electronegativity	It is the quantitative measure of an atom in a chemical compound to attract shared electrons to its self	decreases	increases

9.

A	B
1. N_2O	(b) Neutral
2. BaO	(d) Basic
3. As_2O_3	(a) Amphoteric
4. Cl_2O_7	(c) Acidic

10. (b)

11. (c)

UNIT 4

CHEMICAL BONDING AND MOLECULAR STRUCTURE

1. In Lewis dot representation valence electrons are indicated as dots. Using this idea fill in the blanks.

Atom	Valance electrons	Lewis structure
Be	2	$\begin{array}{c} \cdot \\ \text{Be} \\ \cdot \end{array}$
N	5	----
C	4	---

2. The Lewis dot structure of water (H_2O) molecule is



3. Which among the following is not true with octet theory?

- a. It could not explain the formation of compounds of inert gases like XeF_2 , XeOF_2 , KrF_2 etc
- b. It could explain the shape of the molecule
- c. It could not explain the relative stability of molecule.

4. The molecule which does not obey Octet rule is

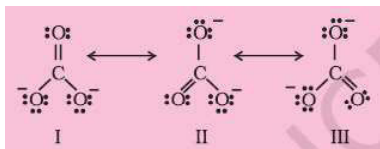
- a. BeCl_2 b. H_2O c. NH_3 d. Cl_2

5. Place the following in to their appropriate columns:



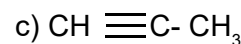
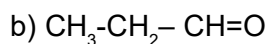
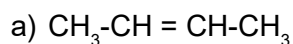
incomplete octet molecules	expanded octet molecules

6. The resonance structure of CO_3^{2-} is



. Draw the resonance structure of CO_2 molecule.

7. Identify the number of σ bonds and π bonds in the following compounds:



8. Which of the following overlapping of atomic orbitals can form a pi bond?

- a. s-s b. s-p c. p-p d. all of these.

9. Which of the following is NOT a characteristic feature of hybridisation?

- (a) The number of hybrid orbitals is equal to the number of the atomic orbitals that get hybridised.
 (b) The hybridised orbitals are always equivalent in energy and shape.
 (c) Atomic orbitals are more effective in forming stable bonds than hybrid orbitals.
 (d) The type of hybridisation indicates the geometry of the molecules.

10. Ammonia molecule contains one lone pair and three bond pairs of electrons. Differentiate between lone pair and bond pair of electrons.

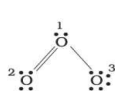

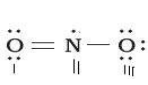
11. Complete the following table:

Molecule	Hybridisation	Shape of the molecule	Bond Angle
CH_4	sp^3	Tetrahedral	$109^\circ 28'$
Ethene (C_2H_4)			
H_2O			
Ethyne (C_2H_2)			
NH_3			

12. Which of the following is not a statement in Fajans rule.

- a. Smaller the size of cation and larger the size of anion, greater the covalent character of an ionic bond
 b. The greater the charge on the cation, greater the covalent character of an ionic bond.
 c. Cations with similar size and charge, the one with electronic configuration of transition metal has greater polarizing power than one with noble gas configuration.
 d. Electron repulsion is in the order $lp-lp > lp-bp > bp-bp$.

13. Complete the following table:

Molecule.	Atom in the molecule.	Total number of valence electrons around the atom.(v)	number of non bonding electrons around the atom.(n)	Number of bonding electrons around the atom.(b)	Formal charge of the atom in the molecule. $F = (v-n) - \frac{1}{2}(b)$
	1	6	2	6	+1
	2	6	4	4	... (a)...
	3	6	... (b)...	2	..(c)...
	N	5	3	4	... (d)..
	O	... (e) (f) ...	4	0
	I	6	... (g)..	4	... (h)...
	II	5	1	... (i)...	... (j)...
	III	... (k)...	6	2	... (l)..

14. Which among the following is a correct statement about bond length?

- Larger the bond length stronger will be the bond.
- More the number of bonds greater the bond length.
- Smaller the bond length stronger will be the bond.
- With increase of bond order bond length increases.

15. Write whether the following statements regarding bond order is true or false. If false, give the correct statement.

- Bond order is the number of bonds between two atoms in a molecule.
- Isoelectronic molecules and ions have identical bond orders.
- With increase of bond order bond length increases.
- With increase of bond order bond enthalpy increases.

16. Differentiate between sigma and pi bond.

17. Write the main postulates of VSEPR theory.

18. Match the following:

Number of bond pairs of electrons	Shape of molecule	Example
2	Tetrahedral	SF ₆
3	Octahedral	BF ₃
4	Linear	PCl ₅
5	Trigonal planar	BeCl ₂
6	Trigonal bipyramidal	CH ₄

19. Considering the X axis as the internuclear axis which out of the following will not form a sigma bond?

- e. 1s and 1s b. 1s and 2px c. 2py and 2py d. 1s and 2s.

20. Give the conditions for the combination of Atomic Orbitals based on MO theory.

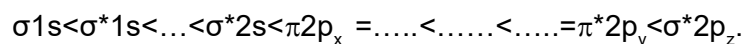
21. Complete the chart

Molecule	Molecular orbital electronic configuration	Number of bonding electrons (Nb)	Number of antibonding electrons (Na)	Bond order = $\frac{1}{2}(Nb - Na)$	Stable or unstable
H ₂	$\sigma 1s^2$	2	0	1	stable
Be ₂	$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2$	4	4	0	unstable
N ₂	$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 \pi 2p_y^2 \sigma 2p_z^2$	10	4
He ₂	-----	2	2

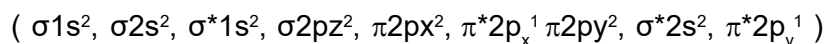
22. Match the Following

Li ₂	$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \pi 2p_y^2 \pi^* 2p_x^2 \pi^* 2p_y^2$	paramagnetic
B ₂	$\sigma 1s^2 \sigma s^* 1s^2 \sigma 2s^2$	diamagnetic
F ₂	$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^1 \pi 2p_y^1$	diamagnetic

23. Fill in the blanks with suitable terms given in brackets to complete the increasing energy level of molecular orbital of C₂. ($\pi 2p_y$, $\sigma 2s$, $\pi^* 2p_x$, $\sigma 2p_z$)



24. Arrange the given molecular orbitals in bracket in the increasing order to get the molecular orbital electronic configuration of O₂ molecule.



25. Complete the table:

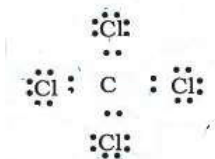
Type of Hydrogenbonding	Nature	Example
Inter molecular	Between molecule
Intramolecular	o-nitro phenol

Solutions

1. In Lewis dot representation valance electrons are indicated as dots. Using this idea fill in the blanks.

Atom	Valance electrons	Lewis structure
Be	2	$\cdot\text{Be}\cdot$
N	5	$\cdot\ddot{\text{N}}\cdot$
C	4	$\cdot\ddot{\text{C}}\cdot$

2.



3.

b It could explain the shape of the molecule

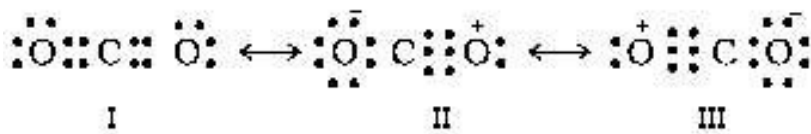
4.

a. BeCl_2

5.

incomplete octet molecules	expanded octet molecules
BeCl_2	SF_6
BF_3	PCl_5
LiCl	H_2SO_4

6.



7.

a) σ bonds – 11, π bonds - 1

b) σ bonds – 9, π bonds - 1

c) σ bonds – 6 π bonds - 2

8.

c. p-p

9.

(c) Atomic orbitals are more effective in forming stable bonds than hybrid orbitals.

10. The valence electron pairs involved in bond formation are called bond pairs and the electron pairs not involved in the bond formation are called lone pairs.

11. Complete the following table:

Molecule	Hybridisation	Shape of the molecule	Bond Angle
CH ₄	sp ³	Tetrahedral	109°28'
Ethene (C ₂ H ₄)	sp ²	Triagonal planar	120°
H ₂ O	sp ³	Bent or V shape	104.5°
Ethyne (C ₂ H ₂)	sp	Linear	180°
NH ₃	sp ³	pyramidal	107°

12.

d. Electron repulsion is in the order lp-lp > lp-bp > bp-bp.

13.

a. 0 b.6 c.-1 d.0 e.6 f.4 g.4 h.0 i.6 j.+1 k.6 l.-1

14.

c. Smaller the bond length stronger will be the bond.

15.

- a. True.
- b. True.
- c. False. With increase of bond order bond length decreases.
- d. True.

16.

sigma bond	pi bond
formed by linear overlapping of atomic orbitals	formed by lateral overlapping of atomic orbitals
Strong bond	Weak bond

17. The important postulates of this theory are:

- 1) The shape of the molecule depends on the number of valence shell electron pairs around the central atom.
- 2) The valence shell electron pairs repel each other.
- 3) In order to minimize repulsion and maximise stability, the electron pairs stay at maximum distance.
- 4) 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.
- 5) As the angle between the electron pairs increases, the repulsion decreases.

18. Match the following:

Number of bond pairs of electrons	Shape of molecule	Example
2	Linear	BeCl ₂
3	Trigonal planar	BF ₃
4	Tetrahedral	CH ₄
5	Trigonal bipyramidal	PCl ₅
6	Octahedral	SF ₆

19. c. 2p_y and 2p_y

20. 1. The combining atomic orbitals must have the same or nearly same energy.
 2. The combining atomic orbitals must have the symmetry about the molecular axis.
 3. The combining atomic orbitals must overlap to the maximum extent.

21.

Molecule	Molecular orbital electronic configuration	Number of bonding electrons (Nb)	Number of antibonding electrons (Na)	Bond order = $\frac{1}{2} (Nb - Na)$	Stable or unstable
H ₂	$\sigma 1s^2$	2	0	1	stable
Be ₂	$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2$	4	4	0	unstable
N ₂	$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 \pi 2p_y^2 \sigma 2p_z^2$	10	4	3	stable.
He ₂	$\sigma 1s^2 \sigma^* 1s^2$	2	2	0	unstable

22.

Li_2	$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2$	diamagnetic
B_2	$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^1 \pi 2p_y^1$	paramagnetic
F_2	$\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2 \pi 2p_x^2 \pi 2p_y^2 \pi^* 2p_x^2 \pi^* 2p_y^2$	diamagnetic

23. $\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < \pi 2p_x = \pi 2p_y < \sigma 2p_z < \pi^* 2p_x = \pi^* 2p_y < \sigma^* 2p_z$.24. $\sigma 1s^2 \quad \sigma^* 1s^2 \quad \sigma 2s^2 \quad \sigma^* 2s^2 \quad \sigma 2p_z^2 \quad \pi 2p_x^2 \quad \pi 2p_y^2 \quad \pi^* 2p_x^1 \quad \pi^* 2p_y^1$

25.

Type of Hydrogenbonding	Nature	Example
Inter molecular	Between molecule	Water, HF
Intramolecular	within the molecule	o-nitro phenol

UNIT 5

THERMODYNAMICS

1. Complete the table:

No.	Type of System	Nature of the movements of matter and energy	Example
i	Open System	Exchange of matter and energy between system and surroundings is possible	Reactants in an open test tube
ii	Closed System	_____	_____
iii	_____	_____	Reactants taken in a thermos flask

2. The values of state variables or state functions depend only on the state of the system and not on how it is reached. Identify state variables from the given list.

Heat (q), pressure (p), temperature (T), volume (V), work (w)

3. Identify and state the law represented by the equation, $\Delta U = q + w$ and fill in the table appended.

No.	Process	q	w	$\Delta U = q + w$
I	Adiabatic, but work is done by the system.	0	$-w_{ad}$	$\Delta U = -w_{ad}$
II	_____	-q	0	$\Delta U = -q$
III	w amount of work is done by the system and q amount of heat is supplied to the system	_____	_____	_____
IV	Compression of a gas from volume, V_i to V_f , achieved in a single step against an external pressure, p_{ex} and at constant temperature, T.	q	_____	$0 = q - p_{ex}(V_f - V_i)$
V	Isothermal (T = constant) expansion of an ideal gas into vacuum	_____	_____	_____
VI	Isothermal reversible change of an ideal gas from volume, V_i to V_f at temperature, T	_____	_____	_____

4. Enthalpy of a system can be defined by the equation, $H = U + pV$.

- (i) If a process occurs at constant pressure, give an expression for enthalpy change, ΔH .
- (ii) If $\Delta U = q_p - p \Delta V$, what will be the relationship between enthalpy change and heat absorbed at constant pressure?
- (iii) If $p \Delta V = \Delta n_g RT$, give the equation for calculating ΔH from ΔU and vice versa.

5. Identify the correct definition and examples of intensive properties from the following:

Definitions:

- (i) Those properties which depend on the quantity or size of matter present in the system.
- (ii) Those properties which do not depend on the quantity or size of matter present

Examples: temperature, density, pressure, mass, volume, internal energy, enthalpy, heat capacity, molar volume, and molar heat capacity.

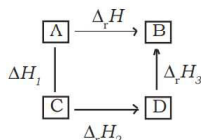
6. Derive the relationship between C_p and C_v for one mole of an ideal gas.

Hint: $q_v = C_v \Delta T = \Delta U$, $q_p = C_p \Delta T = \Delta H$ and $pV = nRT$

7. Match the following:

No.	Type of enthalpy change	Example (Thermochemical equation)
1	Enthalpy of vapourisation	$\text{Na}^+\text{Cl}(\text{s}) \rightarrow \text{Na}^+(\text{g}) + \text{Cl}^-(\text{g})$; $\Delta H^\ominus = +788 \text{ kJ mol}^{-1}$
2	Enthalpy of formation	$\text{H}_2(\text{g}) \rightarrow 2\text{H}(\text{g})$; $\Delta H^\ominus = 435.0 \text{ kJ mol}^{-1}$
3	Enthalpy of combustion	$\text{HCl.25 aq.} + 15 \text{ aq.} \rightarrow \text{HCl.40 aq.}$; $\Delta H = -0.76 \text{ kJ mol}^{-1}$
4	Bond dissociation enthalpy	$2\text{C}(\text{graphite, s}) + 3\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{C}_2\text{H}_5\text{OH}(\text{l})$; $\Delta H^\ominus = -277.7 \text{ kJ mol}^{-1}$
5	Lattice enthalpy	$\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$; $\Delta H^\ominus = +40.79 \text{ kJ mol}^{-1}$
6	Enthalpy of dilution	$\text{C}_2\text{H}_5\text{OH}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$; $\Delta H^\ominus = -1367 \text{ kJ mol}^{-1}$

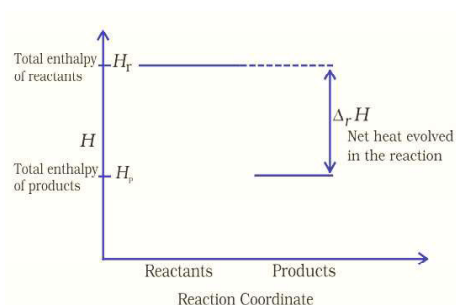
8. State Hess' law and apply it to find $\Delta_r H$ from the following diagram:



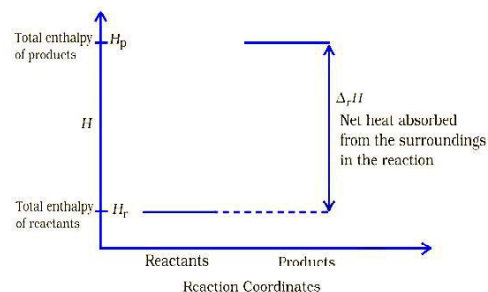
9. Construct Born – Haber cycle for the formation of NaCl(s) and write the equation to find lattice enthalpy of NaCl(s) from it.

10. From the following diagrams identify the one representing endothermic reaction:

(i)



(ii)



11. Identify the correct statements:

- For a given substance, the crystalline solid state is the state of lowest entropy and the gaseous state is state of highest entropy.
- Entropy, like any other thermodynamic property such as internal energy U and enthalpy H is a state function and S is independent of path.
- Heat added to a system at lower temperature causes greater randomness than when the same quantity of heat is added to it at higher temperature.
- Both for reversible and irreversible expansion of an ideal gas, under isothermal conditions, $\Delta U = 0$, but ΔS_{total} is not zero for irreversible process.
- ΔU does not discriminate between reversible and irreversible process, whereas ΔS does.

12. Which of the following is dimensionally different from others?

- (a) ΔU (b) ΔH (c) ΔS (d) ΔG (e) $T \Delta S$

13. Write Gibb's equation and explain why ΔG is called 'free energy of reaction'?

14. What is meant by a spontaneous process? Derive the criterion of spontaneity in terms of free energy of reaction.

15. Identify the law corresponding to each of the following statements:

- The entropy of any pure crystalline substance approaches zero as the temperature approaches absolute zero.
- Increase in entropy in an isolated system is the natural direction of a spontaneous change.
- The energy of an isolated system is constant.

16. Give the relation connecting standard free energy of a reversible reaction (a reaction, which can proceed in either direction simultaneously) and its equilibrium constant.

SOLUTION**1. Complete the table:**

No.	Type of System	Nature of the movements of matter and energy	Example
i	Open System	Exchange of matter and energy between system and surroundings is possible	Reactants in an open test tube
ii	Closed System	Exchange of energy between system and surroundings is possible, but not matter	Reactants in a closed vessel made of Cu
iii	Isolated System	No exchange of matter or energy between system and surroundings is possible	Reactants taken in a thermos flask

2. State variables:

pressure (p), temperature (T), volume (V)

3. First law of thermodynamics. The law states that the energy of an isolated system is constant.

No.	Process	q	w	$\Delta U = q + w$
I	Adiabatic, but work is done by the system.	0	$-w_{ad}$	$\Delta U = -w_{ad}$
II	No work is done but q amount of heat is given to surroundings	-q	0	$\Delta U = -q$
III	w amount of work is done by the system and q amount of heat is supplied to the system	q	-w	$\Delta U = q - w$
IV	Compression of a gas from volume, V_i to V_f , achieved in a single step against an external pressure, p_{ex} and at constant temperature, T.	q	$-p_{ex}(V_f - V_i)$	$0 = q - p_{ex}(V_f - V_i)$
V	Isothermal (T = constant) expansion of an ideal gas into vacuum	0	0	$\Delta U = 0$
VI	Isothermal reversible change of an ideal gas from volume, V_i to V_f at temperature, T	q	$-nRT \ln(V_f/V_i)$	$0 = q - nRT \ln(V_f/V_i)$

4. Enthalpy of a system can be defined by the equation, $H = U + pV$.

$$(i) \Delta H = \Delta U + p\Delta V$$

$$(ii) \Delta H = \Delta U + p\Delta V$$

$$\Delta H = q_p - p\Delta V + p\Delta V, \text{ (since } \Delta U = q_p - p\Delta V)$$

$$\text{Therefore, } \Delta H = q_p$$

$$(iii) \Delta H = \Delta U + p\Delta V$$

$$\text{i.e., } \Delta H = \Delta U + \Delta n_g RT, \text{ (since } p\Delta V = \Delta n_g RT)$$

5. **Correct definition:** (ii) Those properties which do not depend on the quantity or size of matter present are known as intensive properties.

Examples: temperature, density, pressure, molar volume, and molar heat capacity.

6. **Hint:** $q_v = C_v \Delta T = \Delta U$, $q_p = C_p \Delta T = \Delta H$ and $pV = nRT$

The relation between C_p and C_v can be derived for an ideal gas as:

For a mole of an ideal gas, $\Delta H = \Delta U + \Delta(pV) = \Delta U + \Delta(RT) = \Delta U + R\Delta T$

On putting the values of ΔH and ΔU , we have $C_p \Delta T - C_v \Delta T = R\Delta T$ or $C_p - C_v = R$

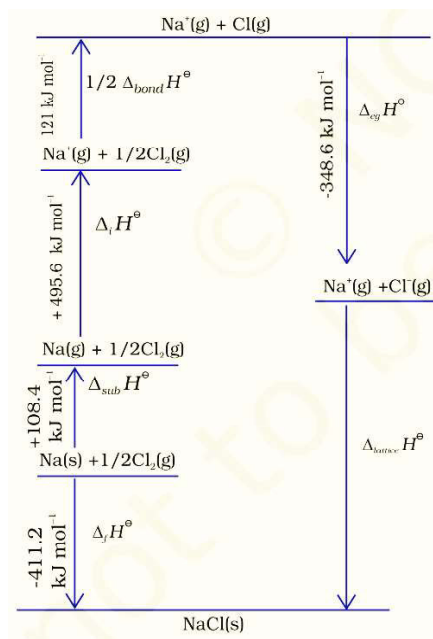
7. Match the following:

No.	Type of enthalpy change	Example (Thermochemical equation)
1	Enthalpy of vapourisation	$\text{H}_2\text{O(l)} \rightarrow \text{H}_2\text{O(g)}; \Delta H^\ominus = + 40.79 \text{ kJ mol}^{-1}$
2	Enthalpy of formation	$2\text{C (graphite, s)} + 3\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{C}_2\text{H}_5\text{OH(l)};$ $\Delta H^\ominus = - 277.7 \text{ kJ mol}^{-1}$
3	Enthalpy of combustion	$\text{C}_2\text{H}_5\text{OH(l)} + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O(l)};$ $\Delta H^\ominus = - 1367 \text{ kJ mol}^{-1}$
4	Bond dissociation enthalpy	$\text{H}_2(\text{g}) \rightarrow 2\text{H(g)}; \Delta H^\ominus = 435.0 \text{ kJ mol}^{-1}$
5	Lattice enthalpy	$\text{Na}^+\text{Cl}^-(\text{s}) \rightarrow \text{Na}^+(\text{g}) + \text{Cl}^-(\text{g}); \Delta H^\ominus = +788 \text{ kJ mol}^{-1}$
6	Enthalpy of dilution	$\text{HCl.25 aq.} + 15 \text{ aq.} \rightarrow \text{HCl.40 aq.}; \Delta H =$ $- 0.76 \text{ kJ mol}^{-1}$

8. If a reaction takes place in several steps then its standard reaction enthalpy is the sum of the standard enthalpies of the intermediate reactions into which the overall reaction may be divided at the same temperature.

$$\Delta_r H = \Delta_r H_1 + \Delta_r H_2 + \Delta_r H_3$$

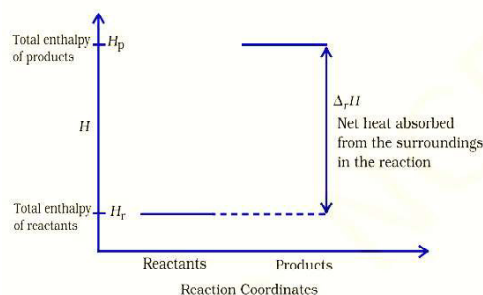
9.



$$\Delta_{\text{lattice}} H^{\ominus} = \Delta_{\text{f}} H^{\ominus} - (\Delta_{\text{sub}} H^{\ominus} + \Delta_{\text{i}} H^{\ominus} + \frac{1}{2} \Delta_{\text{bond}} H^{\ominus} + \Delta_{\text{eg}} H^{\ominus})$$

10. Diagram representing endothermic reaction:

(ii)



11. Identify the correct statements: All the five given statements are correct.

12. Dimensionally different from other given quantities: (c) ΔS (dimension- J K^{-1})(a) ΔU (b) ΔH (d) ΔG (e) $T \Delta S$ (these have dimension- J)13. Gibb's equation: $\Delta G = \Delta H - T \Delta S$

ΔH is the enthalpy change of a reaction, $T \Delta S$ is the energy which is not available to do useful work. So ΔG is the net energy available to do useful work and is thus a measure of the 'free energy'. Therefore it is known as the free energy of the reaction.

14. A process, which is having the potential to proceed without the assistance of external agency is called a spontaneous process or A spontaneous process is an irreversible process and may only be reversed by some external agency.

$$\text{We know that, } \Delta S_{\text{total}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}}$$

If the system is in thermal equilibrium with the surrounding, then the temperature of the surrounding is same as that of the system. Also, increase in enthalpy of the surrounding is equal to decrease in the enthalpy of the system. Therefore, entropy change of surroundings,

$$\Delta S_{\text{surr}} = \frac{\Delta H_{\text{surr}}}{T} = -\frac{\Delta H_{\text{sys}}}{T}$$

$$\Delta S_{\text{total}} = \Delta S_{\text{sys}} + \left(-\frac{\Delta H_{\text{sys}}}{T}\right)$$

Rearranging the above equation:

$$T \Delta S_{\text{total}} = T \Delta S_{\text{sys}} - \Delta H_{\text{sys}}$$

For spontaneous process, $\Delta S_{\text{total}} > 0$, so $T \Delta S_{\text{sys}} - \Delta H_{\text{sys}} > 0$

i.e., $-\Delta G > 0$, (since $\Delta G = \Delta H - T \Delta S$) or $\Delta G < 0$ or ΔG must be negative for a process to be spontaneous at constant temperature and pressure.

15.

- (i) The entropy of any pure crystalline substance approaches zero as the temperature approaches absolute zero: Third law of thermodynamics
- (ii) Increase in entropy in an isolated system is the natural direction of a spontaneous change: Second law of thermodynamics
- (iii) The energy of an isolated system is constant: First law of thermodynamics

$$16. 0 = \Delta_r G^\circ + RT \ln K \text{ or } \Delta_r G^\circ = -RT \ln K \text{ or } \Delta_r G^\circ = -2.303 RT \log K$$

Unit 6
Chemical Equilibrium

1. Which of the following is not a general characteristics of Physical equilibria

- a) All measurable properties of the system remain constant.
- b) All the physical processes stop at equilibrium
- c) Equilibrium is possible only in a closed system.
- d) Both the opposing processes occur at the same rate and there is a dynamic condition.

2. Under what conditions ice-water system is equilibrium at 1 atm.

- a) Below 273K
- b) At 273 K
- c) Above 273 K
- d) At Zero K

3. If the equilibrium constant value for the equilibrium $2\text{HI}(\text{g}) \rightarrow \text{H}_2(\text{g}) + \text{I}_2(\text{g})$ is K , then its value for the equilibrium $\frac{1}{2} \text{H}_2(\text{g}) + \frac{1}{2} \text{I}_2(\text{g}) \rightarrow \text{HI}(\text{g})$ is

- a) $\frac{1}{K}$ b) $\frac{1}{K^2}$ c) $\frac{1}{\sqrt{K}}$ d) K^2

4. Identify the heterogenous equilibria from the following.

- (a) $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
- (b) $\text{Fe}^{3+}(\text{aq}) + \text{SCN}^{-}(\text{aq}) \rightarrow [\text{Fe}(\text{SCN})]^{2+}(\text{aq})$
- (c) $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightarrow 2\text{HI}(\text{g})$
- (d) $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

5. Find the value of Δn in the relation $K_p = K_c (RT)^{\Delta n}$ for the following equilibria .

Equilibria	Δn
$\text{N}_{2(\text{g})} + \text{O}_{2(\text{g})} \rightleftharpoons 2\text{NO}(\text{g})$
$\text{N}_{2(\text{g})} + \text{O}_{4(\text{g})} \rightleftharpoons 2\text{NO}_2(\text{g})$

6. One mole each of H_2O and CO are taken in 10 L vessel and heated to 725K. At equilibrium 70% of H_2O reacts with CO according to the equation.



The equilibrium constant for the reaction is

- a) $\frac{49}{9}$ b) $\frac{7}{9}$ c) $\frac{9}{49}$ d) $\frac{9}{7}$

7. The yield of NH_3 can be increased as per the equilibrium $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$ by

- Decreasing the pressure
- Removal of N_2
- Increasing the pressure
- Removal of H_2

8. Complete the table

Species	Conjugate acid	Conjugate base
H_2O	H_3O^+	OH^-
NH_3	
HSO_4^-	

9. Pick out the Lewis acid and Lewis base from the following.

$\text{NH}_3, \text{CN}^-, \text{CH}_3\text{OH}, \text{BF}_3, \text{Al}^{3+}$

Lewis Acid	Lewis Base

10. pH is the negative logarithm of numerical value of hydronium ion concentration.

Classify the given solution into acidic, basic and neutral based on their pH value.

Solution	pH	Nature of Solution
A	4.7
B	11.3
C	7.0
D	9.1

11. Reaction quotient Q_c is useful in predicting the direction of reaction by comparing it with K_c .

If $Q_c > K_c$

- No Net reaction occurs.
- Net reaction goes from left to right
- Net reaction goes from right to left
- Equilibrium is attained.

12. K_a ionisation constant of acid is a measure of the strength of the acid.

Arrange the following acids in increasing order of their acid strength

Acid	K_a
Benzoic acid	6.5×10^{-5}
Hydrocyanic acid	4.9×10^{-10}
Nitrous acid	4.5×10^{-4}
Niacin	1.5×10^{-5}

13. 0.2 M solution of a weak acid is dissociated to an extent of 1.5%, The dissociation constant of the weak acid is

- a. 4.5×10^{-5} b. 1.44×10^{-5} c. 2.45×10^{-4} d. 1.75×10^{-4}

14. For reactions involving gases the equilibrium constant is expressed in terms of partial pressure

- How it is related to equilibrium constant in terms of concentration.
- Write a chemical reaction in which equilibrium constant in terms of concentration is equal to that in terms of partial pressure

15. The dissociation of a weak electrolyte is suppressed by addition of a strong electrolyte containing a common ion, if so the addition of nitric acid does not suppress the dissociation of

[Acetic acid, Hydrochloric acid, Benzoic acid, Hydrogen sulphide]

16. Buffer solutions resist the change in pH on addition of small amount of acid or alkali
Classify the given buffers into acidic and basic buffer

[HCN + KCN, $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$, $\text{HNO}_2 + \text{KNO}_2$]

17. Solubility product constant relates the equilibrium between solid solute and dissociated ions in solution, the solubility product value for CaF_2 is

- a. S^2 b. $4S^3$ c. $27S^4$ d. S^4

18. Salts are formed by the reaction between acids and bases in definite proportions. Analyse the salt and complete the following table

Salt	Acid	Base
NH_4Cl	HCl	NH_4OH
CH_3COOK
Na_2SO_4
KCl

19. Which of the following salt solution is having pH value greater than 7

- a. Ammonium nitrate
b. Ammonium Chloride
c. Potassium Chloride
d. Sodium acetate

20. KCl is formed by the reaction between KOH and HCl.

- a. Name the salt formed by the reaction between KOH and acetic acid (CH_3COOH)
b. Predict the nature of its aqueous solution

21. Which of the following salt does not undergo hydrolysis

- a. Sodium sulphate
b. Ammonium Sulphate
c. Aluminium sulphate
d. All the above

22. Temperature changes affect the reaction rate and equilibrium constant

- a. How the increase in temperature affects the equilibrium constant of
i) Endothermic reaction ii) Exothermic reaction

23. The conjugate acid-base pair differ only by a proton.

Select the conjugate pair from the following

Conjugate Acid	Conjugate Base
H_2CO_3	NH_2^-
NH_3	OH^-
HF	HCO_3^-
H_2O	F^-

24. Polyprotic acids have more than one ionisation constant values

- What are polyprotic acids. Give an example
- Write the ionisation reaction for the above polyprotic acid

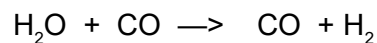
SOLUTIONS

1. b. All the physical processes stop at equilibrium
2. b. At 273 K
3. c. $1/\sqrt{K}$
4. d. $\text{CaCO}_3(\text{s}) \longrightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

5.

Equilibria	Δn
$\text{N}_2 + \text{O}_2 \rightleftharpoons 2\text{NO}$	0
$\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$	1

6.



Initial

Conc (mol/l) 1 1 0 0

% dissociated =

$$1 \times 70/100 = 0.7$$

Equb Conc

1-0.7 1-0.7 0.7 0.7

$$K_c = \frac{[\text{CO}][\text{H}_2]}{[\text{H}_2\text{O}][\text{CO}]}$$

$$= \frac{0.7 \times 0.7}{0.3 \times 0.3} = \frac{0.49}{0.09} = 49/9$$

7. c. Increasing the Pressure

8.

Species	Conjugate acid	Conjugate base
H_2O	H_3O^+	OH^-
NH_3	NH_4^+	NH_2^-
HSO_4^-	H_2SO_4	SO_4^{2-}

9.

Lewis Acid	Lewis Base
$\text{BF}_3, \text{Al}^{3+}$	$\text{NH}_3, \text{CN}^-, \text{CH}_3\text{OH}$

10. If PH = 7, Neutral, PH > 7 Basic PH < 7 Acidic

Solution	PH	Nature of Solution
A	4.7	Acidic
B	11.3	Basic
C	7.0	Neutral
D	9.1	.Basic

11. c. Net reaction goes from right to left

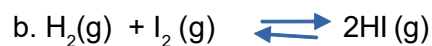
12. Greater the K_a value stronger will be the acid and its strength

The increasing order of acidic strength is

Hydrocyanic acid < Niacin < Benzoic acid < Nitrous acid

$$\begin{aligned}
 13. K_a &= C\alpha^2 & \alpha &= 1.5/100 = 0.015 \\
 &= (0.2) \times (0.015)^2 \\
 &= 4.5 \times 10^{-5}
 \end{aligned}$$

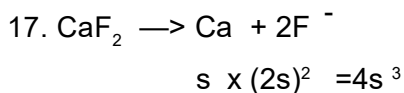
$$14. a. K_p = K_c(RT)^{\Delta n}$$



15. Hydrochloric acid, as it is a strong electrolyte

16. Acidic buffer = HCN + KCN, HNO₂ + KNO₂

Basic buffer = NH₄OH + NH₄Cl



18.

Salt	Acid	Base
NH_4Cl	HCl	NH_4OH
CH_3COOK	CH_3COOH	KOH
Na_2SO_4	H_2SO_4	NaOH
KCl	HCl	KOH

19. d. Sodium acetate, which on hydrolysis give NaOH (strong base) and CH_3COOH (weak acid)

20. a. CH_3COOK

b. Basic in nature

21. a. Sodium sulphate

22. a. i) The equilibrium constant for an endothermic reaction increases as the temperature increases

ii) The equilibrium constant for an exothermic reaction decreases as the temperature increases

23.

Conjugate Acid	Conjugate Base
H_2CO_3	HCO_3^-
NH_3	NH_2^-
HF	F^-
H_2O	OH^-

24. a. The acids which have more than one ionisable proton per molecule of the acid are known as polyprotic acids

eg : H_2SO_4

b. The ionisation reaction for H_2SO_4 can be represented as



UNIT 7

REDOX REACTIONS

1. Complete the table based on different concepts of oxidation and reduction:

No.	Concept	Definition of:		Example
		Oxidation	Reduction	
1	Classical concept			
2	Electronic concept			
3	Oxidation number concept			

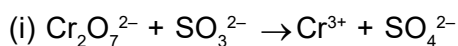
2. Give an example for competitive electron transfer reaction.

3. Fill in the table given below:

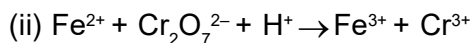
No.	Compound	Identify the metal and its oxidation number in the compound	Stock Notation
1	HAuCl ₄	Au; O.N. = 3	HAu(III)Cl ₄
2	Tl ₂ O		
3	FeO		
4	Fe ₂ O ₃		
5	CuI		
6	CuO		
7	MnO		
8	MnO ₂		

4. Explain different types of redox reactions.

5. Balance the following chemical equations:



(Acidic medium; use oxidation number method)



(Use half reaction method)

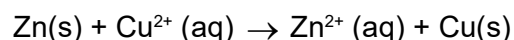
6. What is meant by a redox couple?

SOLUTION**1. Complete the table based on different concepts of oxidation and reduction:**

No.	Concept	Definition of:		Example
		Oxidation	Reduction	
1	Classical concept	The addition of oxygen/ electronegative element to a substance or removal of hydrogen/ electropositive element from a substance.	The removal of oxygen/ electronegative element from a substance or addition of hydrogen/ electropositive element to a substance	Oxidation: $2\text{H}_2\text{S}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{S}(\text{s}) + 2\text{H}_2\text{O}(\text{l})$ Reduction: $2 \text{HgO} (\text{s}) \rightarrow 2 \text{Hg} (\text{l}) + \text{O}_2 (\text{g})$
2	Electronic concept	Loss of electron(s) by any species.	Gain of electron(s) by any species.	$2\text{Na}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{Na}^+\text{Cl}^-(\text{s})$ loss of $2e^-$ gain of $2e^-$
3	Oxidation number concept	An increase in the oxidation number of the element in the given substance.	A decrease in the oxidation number of the element in the given substance.	$\overset{-4}{\text{C}}\overset{+1}{\text{H}}_4(\text{g}) + 4\overset{0}{\text{Cl}}_2(\text{g}) \rightarrow \overset{+4}{\text{C}}\overset{-1}{\text{Cl}}_4(\text{l}) + 4\overset{+1}{\text{H}}\overset{-1}{\text{Cl}}(\text{g})$ Oxidation Reduction

2. Give an example for competitive electron transfer reaction.

The reaction between metallic zinc and the aqueous solution of copper nitrate:

**3. Fill in the table given below:**

No.	Compound	Identify the metal and its oxidation number in the compound	Stock Notation
1	HAuCl_4	Au; O.N. = 3	HAu(III)Cl_4
2	Tl_2O	Tl; O.N. = 1	$\text{Tl}_2(\text{I})\text{O}$
3	FeO	Fe; O.N. = 2	Fe(II)O
4	Fe_2O_3	Fe; O.N. = 3	$\text{Fe}_2(\text{III})\text{O}_3$
5	CuI	Cu; O.N. = 1	Cu(I)I
6	CuO	Cu; O.N. = 2	Cu(II)O
7	MnO	Mn; O.N. = 2	Mn(II)O
8	MnO_2	Mn; O.N. = 4	Mn(IV)O_2

4. Explain different types of redox reactions.

(i) Combination reactions: A combination reaction may be denoted as

$A + B \rightarrow C$ Here either A or B or both A and B must be in the elemental form.

e.g., $C(s) + O_2(g) \rightarrow CO_2(g)$

(ii) 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: $C \rightarrow A + B$.

e.g., $2NH_3(g) \rightarrow N_2(g) + 3H_2(g)$

(iii) 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: $X + YZ \rightarrow XZ + Y$

e.g., $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

(iv) Disproportionation reactions: In a disproportionation reaction an element in one oxidation state is simultaneously oxidised and reduced.

e.g., $+1 \quad -1 \qquad \qquad +1 \quad -2 \qquad \qquad 0$

$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

Here the oxygen of peroxide, which is present in -1 state, is converted to zero oxidation state in O_2 and decreases to -2 oxidation state in H_2O .

5. Balance the following chemical equations:

(i) $Cr_2O_7^{2-} + SO_3^{2-} \rightarrow Cr^{3+} + SO_4^{2-}$

(Acidic medium; use oxidation number method)

Step 1: The skeletal equation is: $Cr_2O_7^{2-} + SO_3^{2-} \rightarrow Cr^{3+} + SO_4^{2-}$

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

$+6 \quad -2 \quad +4 \quad -2 \qquad \qquad +6 \quad -2$

$Cr_2O_7^{2-} + SO_3^{2-} \rightarrow Cr^{3+} + SO_4^{2-}$

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).

$Cr_2O_7^{2-} + 3SO_3^{2-} \rightarrow 2Cr^{3+} + 3SO_4^{2-}$

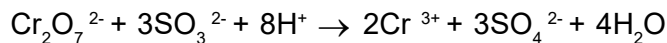
Step 4: Now balance all the atoms except Oxygen and Hydrogen

$Cr_2O_7^{2-} + 3SO_3^{2-} \rightarrow 2Cr^{3+} + 3SO_4^{2-}$

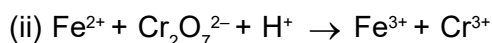
Step 5: Now balance the ionic charges on both sides. Here the net ionic charge on LHS is -8 and on RHS is 0 . To equate them add $8H^+$ on LHS, since the reaction takes place in acidic medium.

$Cr_2O_7^{2-} + 3SO_3^{2-} + 8H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-}$

Step 6: Now balance hydrogen atoms by adding sufficient number of H₂O molecules. Here add 4H₂O molecules on RHS.

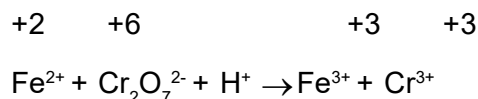


Now the equation is balanced



(Use half reaction method)

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

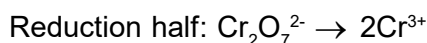


Here Fe²⁺ is oxidized and Cr(VI) is reduced.

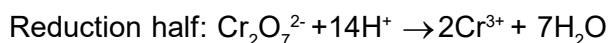
Step-2: Separate the equation into 2 half reactions -oxidation half reaction and reduction half reaction.



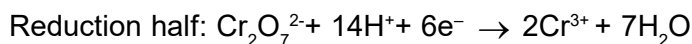
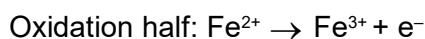
Step-3: Balance the atoms other than O and H in each half reaction individually. Oxidation half: $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+}$



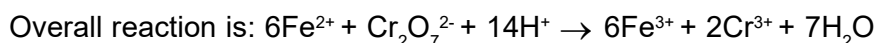
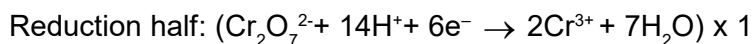
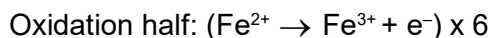
Step-4: Now balance O and H atoms. Add H₂O to balance O atoms and H⁺ to balance H atoms since the reaction occurs in acidic medium.



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.

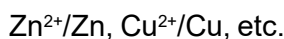


Now the equation is balanced.

6. What is meant by a redox couple?

A redox couple is defined as having together the oxidised and reduced forms of a substance taking part in an oxidation or reduction half reaction.

Example:



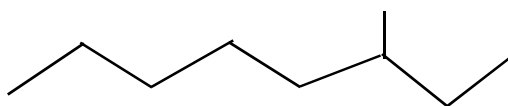
Unit 8

Organic Chemistry - Some Basic Principles and Techniques

1. Complete the following table.

Compound	No. of σ -bonds	No. of π -bonds	Type of Hybridisation
$CH_2 = CH_2$	5	1	SP^2
$CH_3 - CH_3$
$CH \equiv CH$
C_6H_6

2. Bond line formula of an organic compound is given below:

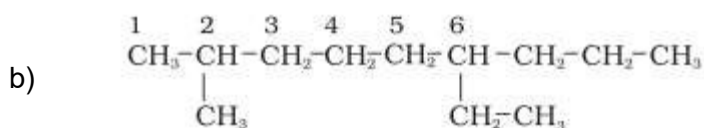
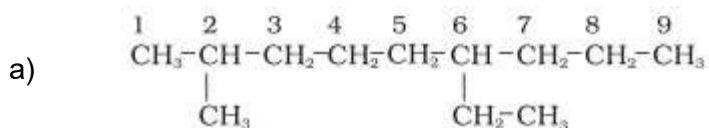


- a) Write its condensed formula
- b) Give IUPAC name of the compound.

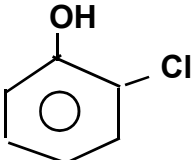
3. Match the type of compounds in column A with their functional groups in column B.

A	B
(i) Alcohols	(a) -CHO
(ii) Amines	(b) -C \equiv N
(iii) Carboxylic acids	(c) -OH
(iv) Aldehydes	(d) -NH ₂
	(e) -COOH

4. Which of the following is to be chosen while the numbering the carbon chain?



5. Complete the following table.

Compound	IUPAC name
(i) $\begin{array}{c} \text{CH}_3 \quad \quad \text{CH}_3 \\ \quad \quad \\ \text{CH}_3 - \text{C} - \text{CH}_2 - \text{CH} - \text{CH}_3 \\ \\ \text{CH}_3 \end{array}$
(ii) $\begin{array}{c} \text{O} \\ \\ \text{CH}_3 - \text{C} - \text{CH}_2 - \text{CH}_2 - \text{COOH} \end{array}$
(iii) 

6. Complete the following table regarding structural isomerism.

SI No.	Type of Structural Isomerism	Description	example
1	Chain isomerism	Two or more Compounds have similar molecular formula but different carbon skeleton.
2	$\begin{array}{c} \text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{OH} \\ \text{CH}_3 - \text{CH} - \text{CH}_3 \\ \\ \text{OH} \end{array}$
3	Functional group Isomerism
4	Different alkyl chains on either sides of functional groups in the molecule

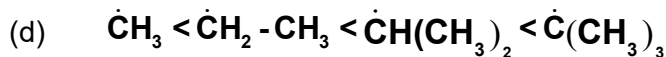
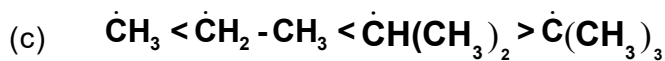
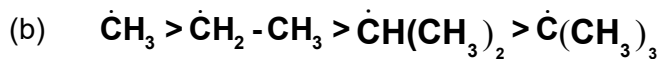
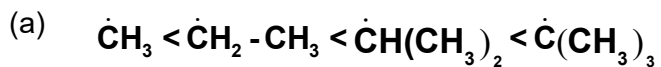
7. Classify the following species into Nucleophiles and Electrophiles.



8. Choose the correct answer about Carbocations.

- Formed as result of hetrolytic bond fission
- Carbon atom possess a positive charge
- Carbon atom possess sextet of electron
- All of the above

9. Choose the correct order of stability of free radicals from the following



10. which of the following electron displacement effect is temporary?

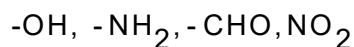
- (a) Inductive effect.
- (b) Resonance effect.
- (c) Hyperconjugation.
- (d) Electromeric effect.

11. Fill the blanks in the following statements by choosing the correct terms given in the following table.

**Hyper conjugation, Inductive effect, Electromeric effect,
Resonance effect, Temporary effect, Permanent effect.**

- (i) Polarisation of σ -bond caused by the polarisation of adjacent bond is called _____
- (ii) 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 is called _____
- (iii) Hyperconjugation is a _____ effect
- (iv) Complete transfer of a shared pair of π -electrons to one of the atoms joined by a multiple bond in the presence of attacking reagent is called _____

12. Classify the following into groups showing +R effect and -R effect



13. Identify the technique used for the separation/purification of the given pair of organic compounds.

- (i) Chloroform (b.p 334K) and aniline (bp. 457K)
- (ii) Crude oil in petroleum industry
- (iii) Glycerol from spent lye in soap industry.
- (iv) Aniline from Aniline - water mixture.

14. Choose the correct reagent, used to detect nitrogen, halogens, sulphur and phosphorous present in an organic Compound in Lassaigne's test. from the given box.

$\text{FeSO}_4 + \text{Con. H}_2\text{SO}_4$, Sodium nitroprusside,
 $\text{HNO}_3 + \text{AgNO}_3$, $\text{HNO}_3 + \text{Ammonium phosphomolybdate}$

15. Match the following.

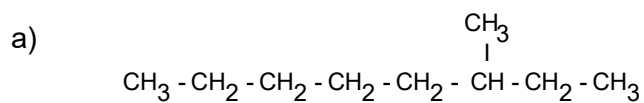
A	B
(i) Dumas Method	(a) Ammonium Sulphate
(ii) Kjeldahl's method	(b) Silver halides
(i) Carius method	(c) Molecular nitrogen

Solutions

1.

Compound	No. of σ -bonds	No. of π -bonds	Type of Hybridisation
$CH_2 = CH_2$	5	1	sp^2
$CH_3 - CH_3$	7	0	sp^3
$CH - CH$	3	2	sp
C_6H_6	12	3	sp^2

2.



b) 3 - Methyloctane

3.

(i) \rightarrow (c)(ii) \rightarrow (d)(iii) \rightarrow (e)(iv) \rightarrow (a)

4.

a)

5.

(i) 2, 2, 4 - Trimethylpentane

(ii) 4 - Oxopentanoic acid

(iii) 2 - Chlorophenol

6.

SI No.	Type of Structural Isomerism	Description	example
1	Chain isomerism	Two or more Compounds have similar molecular formula but different carbon skeleton.	$\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{CH}_3$ & $\text{CH}_3 - \underset{\text{CH}_3}{\text{CH}} - \text{CH}_2 - \text{CH}_3$
2	Position isomerism	Two or more compounds differ in the position of substituent atom or functional group on the carbon skeleton	$\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{OH}$ & $\text{CH}_3 - \underset{\text{OH}}{\text{CH}} - \text{CH}_3$
3	Functional group Isomerism	Two or more compounds have same molecular formula but different functional group	$\text{CH}_3 - \text{CO} - \text{CH}_3$ $\text{CH}_3 - \text{CH}_2 - \text{CHO}$
4	Metamerism	Different alkyl chains on either sides of functional group in the molecule	$\text{CH}_3\text{O} \text{ C}_3\text{H}_7$ & $\text{C}_2\text{H}_5\text{O} \text{ C}_2\text{H}_5$

7.

Nucleophiles :- OH^- , H_2O , R_3N , CN^-

Electrophiles :- $^+\text{CH}_3$, $\text{R} - \overset{\text{O}}{\parallel}{\text{C}} - \text{H}$

8. (d)

9. (a)

10. (d)

11.

- (i) Inductive effect
- (ii) Resonance effect
- (iii) Permanent effect
- (iv) Electromeric effect

12. +R effect :- $-\text{OH}$, $-\text{NH}_2$ -R effect :- $-\text{CHO}$, $-\text{NO}_2$

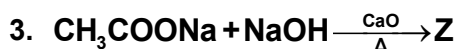
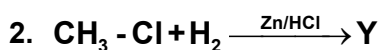
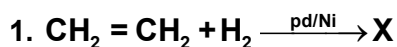
13. (i) Distillation
(ii) Fractional distillation
(iii) Distillation under reduced pressure
(iv) Steam distillation
14. (i) nitrogen \rightarrow $\text{FeSO}_4 + \text{Con. H}_2\text{SO}_4$
(ii) sulphur \rightarrow Sodium nitroprusside
(iii) halogens \rightarrow $\text{HNO}_3 + \text{AgNO}_3$
(iv) phosphorous \rightarrow $\text{HNO}_3 + \text{Ammonium phosphomolybdate}$
15. (i) \rightarrow (c)
(ii) \rightarrow (a)
(iii) \rightarrow (b)

Unit 9 Hydrocarbons

1. Complete the following table

Compound	IUPAC name
1. $\text{CH}_3 - \underset{\text{C}_2\text{H}_5}{\underset{ }{\text{C}}} - \text{CH}_2 - \text{CH}_2 - \text{CH}_3$
2. $\text{CH}_2 = \underset{\text{CH}_3}{\underset{ }{\text{C}}} - \text{CH}_2 - \text{CH}_3$
3. $\text{HC} \equiv \text{C} - \text{CH}_2 - \text{CH}_3$

2. Identify and name the products X, Y & Z in the following reactions.

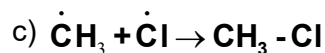
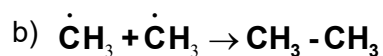
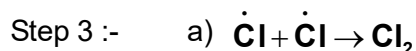
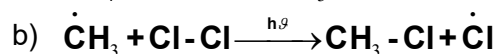
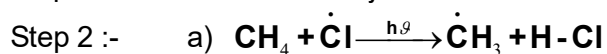
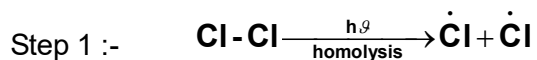


3. "Alkylhalides on treatment with sodium metal in dry ethereal solutions give higher alkanes".

a) Name the reaction stated above.

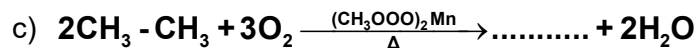
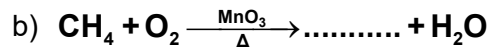
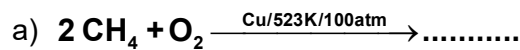
b) Write the chemical reaction equation for the Preparation of ethane by the above reaction

4. Mechanism for the chlorination of methane is given below :



Write the names of step 1, step 2 and step 3 of the above mechanism

5. Identify the products obtained on the controlled oxidation of alkanes given below:



6. "Infinite number of conformations are possible for alkanes".

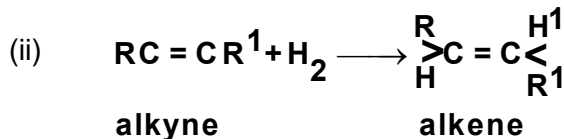
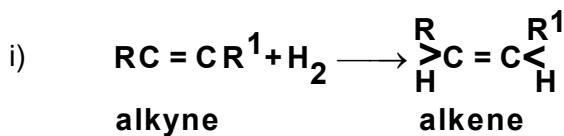
- What are conformations?
- Draw the Newman projections for eclipsed and staggered conformations of ethane
- Which among staggered and eclipsed conformer has least torsional strain?

7. Which of the following compounds will show cis-trans isomerism?

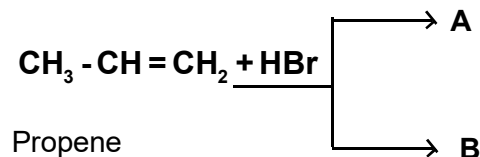
- $(\text{CH}_3)_2\text{C} = \text{CH} - \text{C}_2\text{H}_5 + \text{HBr}$
- $\text{CH}_2 = \text{CBr}_2$
- $\text{C}_6\text{H}_5\text{CH} = \text{CH} - \text{CH}_3$
- $\text{CH}_3\text{CH} = \text{CClCH}_3$

8. Alkynes on partial reduction give alkenes.

Identify the reagents used for effecting following conversions.



9. Analyse the following addition reaction.



- Identify and name the products A & B
- which among A & B is major product?
- Name and State the rule behind the formation of major product.

10. What are the products obtained during the addition reactions of HBr to hex-1-ene;

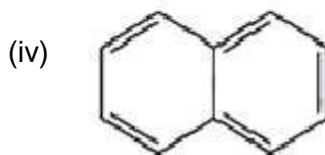
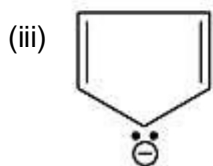
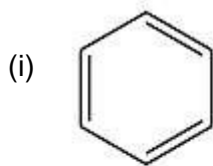
- in the absence of peroxide.
- in the presence of peroxide.

11. An alkene 'A' on Ozonolysis gives a mixture of ethanal and pentan-3-one. Write the structure and IUPAC name of 'A'.

12. Which of the following is/are required for a ring system to show aromaticity?

- (a) Planarity
- (b) Complete delocalisation of π electrons in the ring
- (c) Presence of $(4n+2)$ π electrons in the ring where $n = 0, 1, 2, \dots$
- (d) All of the above

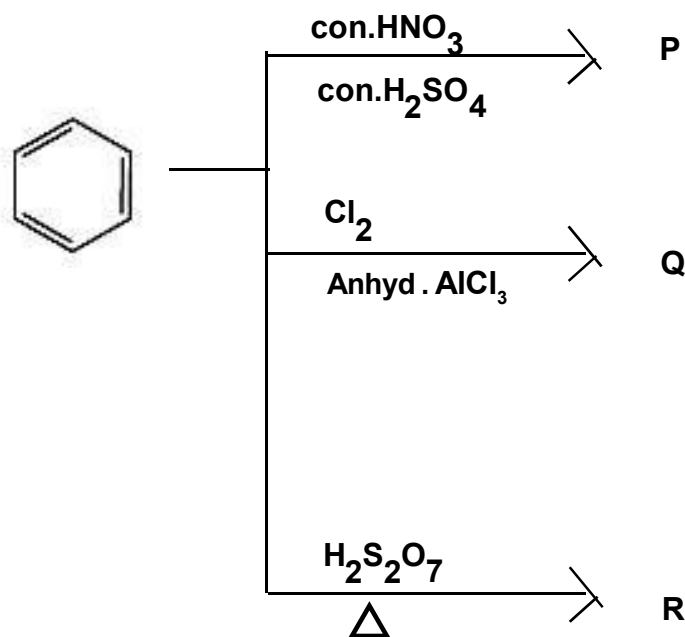
13. Which among the following is not aromatic?



14. How will you effect the following conversions?

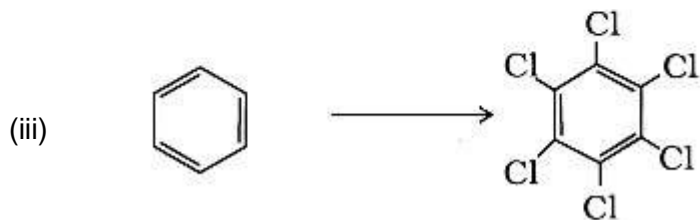
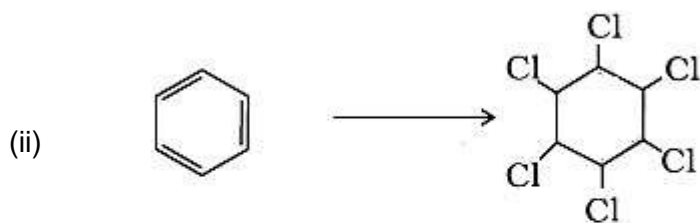
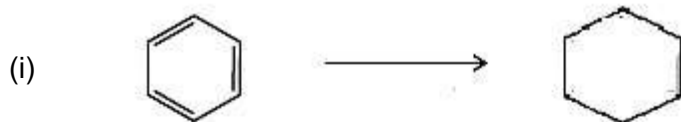
- (a) Ethyne to Benzene
- (b) Phenol to Benzene

15. Identify the products P, Q, R



16. The box given below contains some reagents choose the correct one's for effecting the following conversion.

H_2 / Ni , $Cl_2 / UV / 500 K$, $Cl_2 / Anhyd AlCl_3$ in dark cold condition



Solutions

1.

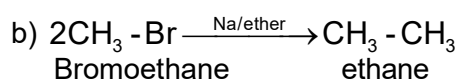
1. 3 - methylhexane
2. 2- methylbut-1-ene
3. But-1-yne

2.

1. $X = \text{CH}_3 - \text{CH}_3$
2. $Y = \text{CH}_4$
3. $Z = \text{CH}_4$

3.

a) Wurtz reaction



4.

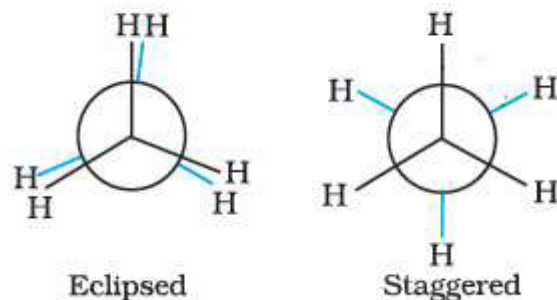
- Step 1 :- Initiation
 Step 2 :- Propagation
 Step 3 :- Termination

5.

- a) $2 \text{CH}_3 - \text{OH}$
- b) HCHO
- c) $2 \text{CH}_3\text{COOH}$

6. (a) spatial arrangements of atoms which can be converted into one another by rotation around a C-C single bond are called conformations.

(b)



(c) Staggered conformation.

7.

(iii) & (iv)

8.

(i) Pd/C (*Lindlar's catalyst*)(ii) Na / liquidNH₃

9.



2- Bromopropane

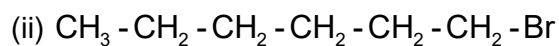
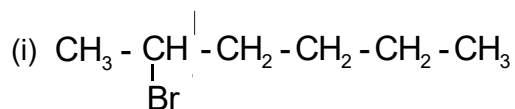
1- Bromopropane

(ii) Major product 2- Bromopropane (A)

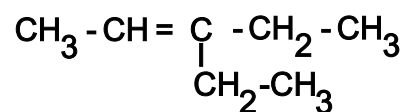
(iii) Markovnikov Rule

Rule states that negative part of the addendum gets attached to that carbon atom which possesses lesser number of hydrogen atoms.

10.



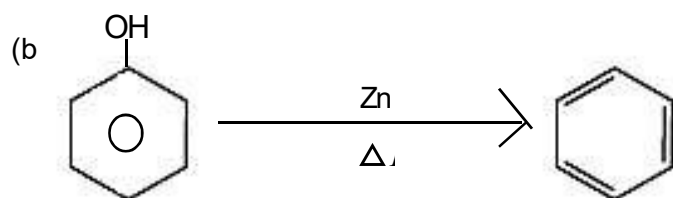
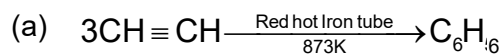
11. 3- Ethylpent-2-ene



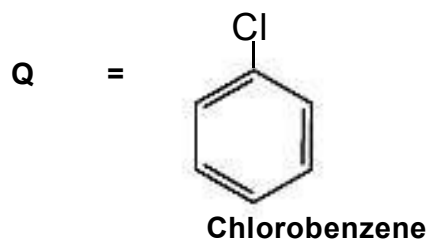
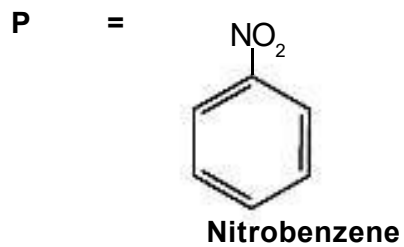
12. (d)

13. (ii)

14.



15.



16.

- (i) H_2 / Ni
- (ii) $\text{Cl}_2 / \text{UV} / 500 \text{ K}$
- (iii) $\text{Cl}_2 / \text{Anhyd AlCl}_3$ in dark cold condition