PART – A

I. Multiple Choice Questions

1. **Number of significant figures in 0.523**
   Answer: (a) 3

2. **Molarity is defined as number of moles of solute present in**
   Answer: (a) One litre of solution

3. **Rutherford's α-scattering experiment is related to size of**
   Answer: (a) Nucleus

4. **Mendeleev's periodic table is based on**
   Answer: (a) Atomic weight

5. **In XeF_4, number of lone pairs and shared pairs respectively**
   Answer: (b) 2, 4

6. **True for ionic compounds**
   Answer: (c) Soluble in polar solvents

7. **Second law of thermodynamics introduced concept of**
   Answer: (c) Entropy

8. **ΔH (constant pressure) − ΔE (constant volume) > RT when**
   Answer: (b) $\Delta n_{(g)} > 1$

9. **If catalyst is used in reversible reaction**
   Answer: (d) Equilibrium is attained more quickly

10. **Lewis acids are**
    Answer: (b) Electron pair acceptors

11. **During reduction, oxidation number**
    Answer: (b) Decreases

12. **Catenation property is more marked in**
    Answer: (b) Carbon

13. **Methoxy methane and ethyl alcohol are**
    Answer: (c) Functional isomers

14. **According to Markovnikov's rule, hydrogen attaches to**
    Answer: (d) Carbon with maximum number of hydrogen atoms

15. **Hydrocarbon that damages DNA and causes cancer**
    Answer: (d) All of these

II. Fill in the blanks

16. Outer shell of most stable atoms contains **8 electrons**

17. Standard enthalpy of formation of element is taken as **zero**

18. When pH decreases, hydrogen ion concentration **increases**

19. Successive members differ by **$-CH_2-$** group

20. Process of converting n-hexane to benzene is called **aromatisation**

PART – B

III. Answer any five (2 Marks Each)

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21. Define Entropy. What happens when liquid changes into vapour?

Entropy (S) is the measure of randomness or disorder of a system.

When a liquid changes into vapour, randomness increases.
Therefore, **entropy increases**.

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22. Define Heterogeneous Equilibrium. Give Example.

A heterogeneous equilibrium is one in which reactants and products are present in different physical states.

Example:
$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

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23. Isoelectronic Species

Isoelectronic species are atoms or ions having the same number of electrons.

Check electrons:

- Na^+ → 10 electrons
- Cl^- → 18 electrons
- F^- → 10 electrons
- Li^+ → 2 electrons

Isoelectronic pair: Na^+ and F^-

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24. Limitations of Octet Rule (Any Two)

1. It fails to explain molecules with incomplete octet (e.g., BF_3).
2. It fails for expanded octet molecules (e.g., PCl_5).

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25. Define Hydrogen Bond. Example of Intramolecular H-bond.

A hydrogen bond is the attractive force between hydrogen attached to a highly electronegative atom (F, O, N) and another electronegative atom.

Example (Intramolecular H-bond): o-Nitrophenol

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26. Why is Helium Monoatomic? (MOT Explanation)

Helium has electronic configuration:

$1s^2$

Bond order formula:

Bond order = $(N_b - N_a) / 2$

For He_2:
Bond order = $(2 - 2)/2 = 0$

Since bond order is zero, He_2 molecule does not exist.
Therefore, helium is monoatomic.

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27. Stock Notation

i) $HAuCl_4$ → Gold(III) chloride
ii) Fe_2O_3 → Iron(III) oxide

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28. $CH \equiv C–CH=CH–CH_3$

i) Bond-line formula:
Five carbon chain with triple bond at C1 and double bond at C3.

ii) The compound is **unsaturated** (contains double and triple bonds).

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29. Geometrical Isomers of But-2-ene

1. **Cis-but-2-ene**
2. **Trans-but-2-ene**

PART – C

IV. Answer any three (3 Marks Each)

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30. Electronegativity

Definition:
Electronegativity is the tendency of an atom in a molecule to attract the shared pair of electrons towards itself.

Trend in Periodic Table:

- **Across a period (left → right):**
  Electronegativity increases
  (Due to increase in nuclear charge and decrease in atomic size)

- **Down a group (top → bottom):**
  Electronegativity decreases
  (Due to increase in atomic size and shielding effect)

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31. Shape of Ammonia (NH_3) – VSEPR Theory

In NH_3:

- **Nitrogen has 5 valence electrons**
- **It forms 3 bond pairs with hydrogen**
- **1 lone pair remains**

Total electron pairs = 4

According to VSEPR theory:

- **Electron pair geometry = Tetrahedral**
- **Due to one lone pair, the molecular shape becomes Trigonal Pyramidal**
- **Bond angle ≈ 107°**

Lone pair–bond pair repulsion > bond pair–bond pair repulsion
Hence bond angle decreases from 109.5° to 107°.

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32. (a) Dipole Moment

Definition:
Dipole moment is the product of magnitude of charge and distance between the charges.

Dipole moment (μ) = q × d

Unit: Debye (D)

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32. (b) Why Sigma Bond is Stronger than Pi Bond?

- **Sigma bond is formed by head-on overlap of orbitals.**
- **Pi bond is formed by sidewise overlap.**

Head-on overlap is more effective than sidewise overlap.
Therefore, sigma bond is stronger than pi bond.

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33. Lewis Structure of NO_2^-

Total valence electrons:

N = 5
O = 6 × 2 = 12
Extra electron (negative charge) = 1

Total = 18 electrons

Structure:

One N atom in centre bonded to two O atoms.
One double bond and one single bond (resonance structure exists).

Formal charge calculation:

Formal charge = Valence electrons − (Lone pair electrons + ½ bonding electrons)

For double bonded oxygen:
Formal charge = 6 − (4 + 2) = 0

For single bonded oxygen:
Formal charge = 6 − (6 + 1) = −1

Hence, negative charge is on one oxygen (delocalised due to resonance).

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34. Balance the Reaction (Oxidation Number Method)

Reaction:

$Fe^{2+} + MnO_4^- \rightarrow Fe^{3+} + Mn^{2+}$

In acidic medium:

Oxidation:
$Fe^{2+} \rightarrow Fe^{3+} + 1e^-$

Reduction:
$MnO_4^- \rightarrow Mn^{2+}$

Balanced equation:

$5Fe^{2+} + MnO_4^- + 8H^+$
$\rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$

V. Answer any three (3 Marks Each)

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35. Dalton's Atomic Theory (Any Three)

1. **Matter is made up of indivisible atoms.**
2. **Atoms of same element have identical mass and properties.**
3. **Atoms combine in simple whole number ratios to form compounds.**

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36. Limitations of Bohr's Model

1. **Applicable only to hydrogen-like atoms.**

2. **Cannot explain Zeeman effect.**
3. **Cannot explain Stark effect.**

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37. Element with Atomic Number 24

Atomic number 24 → Chromium (Cr)

i) Electronic configuration:
[Ar] $3d^5$ $4s^1$

ii) Number of unpaired electrons:
6 unpaired electrons

iii) Belongs to:
d-block element

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38. Relationship Between Cp and Cv

For ideal gas:

$C_p - C_v = R$

Derivation (final result):

$C_p = C_v + R$

Where:
Cp = Heat capacity at constant pressure
Cv = Heat capacity at constant volume
R = Gas constant

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39. Le Chatelier's Principle

If a system at equilibrium is disturbed by change in temperature, pressure or concentration, the system shifts in such a way as to oppose the disturbance.

For exothermic reaction:

Heat is released.

If temperature increases → equilibrium shifts in backward direction.
If temperature decreases → equilibrium shifts in forward direction.

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40. Relation Between Kp and Kc

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

Where
Δn = moles of gaseous products − moles of gaseous reactants

(i) Kp > Kc

Example:
$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$
Here $\Delta n = 1$
So Kp > Kc

(ii) Kp = Kc

Example:
$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$
Here $\Delta n = 0$
So Kp = Kc

PART – D

VI. Answer any two (5 Marks Each)

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41. (a) Hybridisation of Marked Carbon Atoms

Given compound:

$H_3C^a – HC^c = C^b = CH_2$

- **Carbon a (H_3C–) → single bonds only → sp^3 hybridised**
- **Carbon c (–HC=) → involved in one double bond → sp^2 hybridised**
- **Carbon b (=C=) → involved in two double bonds → sp hybridised**

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41. (b) Carbocation

A carbocation is a positively charged carbon species having six electrons in its valence shell.

Example: CH_3^+

Stability order:
3° > 2° > 1°

(Tertiary > Secondary > Primary)

Reason: Greater alkyl groups give +I effect and hyperconjugation.

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42. (a) Estimation of Carbon (Principle & Calculation)

Principle:
Organic compound is heated with excess CuO in presence of oxygen.
Carbon gets oxidised to CO_2.

CO_2 is absorbed in KOH solution.
Increase in weight gives amount of CO_2.

Calculation:

Mass of carbon =
(12 / 44) × Mass of CO_2 formed

Percentage of carbon =

(Mass of carbon / Mass of compound) × 100

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42. (b) Electrophiles

Electrophiles are electron-deficient species which accept electron pairs.

Examples:
H^+, NO_2^+, BF_3

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43. (a) Ozonolysis Problem

Alkene gives:
Ethanal (CH_3CHO)
Propan-2-one (CH_3COCH_3)

Required alkene:

$CH_3–CH=C(CH_3)_2$

IUPAC Name: 2-methylbut-2-ene

Reaction:

Alkene + O_3 → Ozonide →
On hydrolysis → Ethanal + Propan-2-one

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43. (b) Wurtz Reaction

When alkyl halides are treated with sodium in dry ether, higher alkane is formed.

General reaction:

$2R–X + 2Na \rightarrow R–R + 2NaX$

Example:

$2CH_3Cl + 2Na \rightarrow C_2H_6 + 2NaCl$

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44. (a) Nitration of Benzene

Benzene reacts with concentrated HNO_3 in presence of concentrated H_2SO_4.

Step 1: Formation of electrophile
$HNO_3 + H_2SO_4 \rightarrow NO_2^+ + HSO_4^- + H_2O$

Step 2: Electrophilic attack on benzene ring

Step 3: Deprotonation to restore aromaticity

Product formed: Nitrobenzene

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44. (b) Staggered Conformation of Ethane

In staggered conformation:

- **Hydrogen atoms are maximum apart**
- **Torsional strain is minimum**
- **It is most stable conformation**

(Draw Newman projection in exam)

PART – E

VII. Answer any four (3 Marks Each)

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45. Molecular Formula Determination

Given:

C = 26.66%
H = 2.22%
O = 71.12%

Step 1: Convert to moles

C = 26.66 / 12 = 2.22
H = 2.22 / 1 = 2.22
O = 71.12 / 16 = 4.445

Step 2: Divide by smallest value (2.22)

C = 1
H = 1
O = 2

Empirical formula = CHO_2

Empirical mass = 45

Given molecular mass = 90

Multiplier = 90 / 45 = 2

Molecular formula = $C_2H_2O_4$

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46. Mass of Ammonia Formed

Reaction:

$N_2 + 3H_2 \rightarrow 2NH_3$

Given:

Mass N_2 = 2000 g
Mass H_2 = 1000 g

Moles:

N_2 = 2000 / 28 = 71.4 mol
H_2 = 1000 / 2 = 500 mol

Required ratio = 1 : 3

For 71.4 mol N_2 →
H_2 required = 214.2 mol

Available H_2 = 500 mol

So N_2 is limiting reagent

From reaction:

1 mol N_2 → 2 mol NH_3

71.4 mol N_2 → 142.8 mol NH_3

Mass NH_3 = 142.8 × 17

= 2427.6 g

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47. First Line of Balmer Series

Formula:

$1/\lambda = R [(1/2^2) - (1/3^2)]$

$1/\lambda = 1.09677 \times 10^7 \times (1/4 - 1/9)$

$= 1.09677 \times 10^7 \times (5/36)$

$= 1.52 \times 10^6$ m^{-1}

Wave number = 1.52×10^6 m^{-1}

Wavelength = $1 / (1.52 \times 10^6)$

$= 6.58 \times 10^{-7}$ m

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48. Energy of One Mole of Photons

Given:

$v = 5 \times 10^{14}$ Hz

Formula:

E = hv

$h = 6.626 \times 10^{-34}$ J s

Energy of one photon:

$= 6.626 \times 10^{-34} \times 5 \times 10^{14}$
$= 3.313 \times 10^{-19}$ J

For one mole:

$E = 3.313 \times 10^{-19} \times 6.022 \times 10^{23}$

$= 1.996 \times 10^5$ J

= 199.6 kJ

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49. Enthalpy of Formation of Benzene

Given:

$\Delta Hc(C) = -393.5$ kJ/mol
$\Delta Hc(H_2) = -285.83$ kJ/mol
$\Delta Hc(C_6H_6) = -3267$ kJ/mol

Using Hess law:

$\Delta Hf(C_6H_6) = [6(-393.5) + 3(-285.83)] - (-3267)$

$= -2361 - 857.49 + 3267$

$= +48.51$ kJ/mol

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50. Work Done in Isothermal Expansion

Formula:

$W = -nRT \ln(V_2/V_1)$

Given:

$n = 1$ mol
$R = 8.314$ J/mol K
$T = 298$ K
$V_2/V_1 = 20/10 = 2$

$W = -(1)(8.314)(298) \ln 2$

$= -(2477.5)(0.693)$

$= -1717$ J

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51. ΔG° of Hydrolysis

Formula:

$\Delta G° = -RT \ln K$

Given:

$K = 2 \times 10^{-3}$
$T = 300$ K

$\Delta G° = -(8.314)(300) \ln(2 \times 10^{-3})$

$\ln(2 \times 10^{-3}) = -6.214$

$\Delta G° = 15500$ J

= 15.5 kJ

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52. Equilibrium Composition

Reaction:

$2N_2 + O_2 \rightleftharpoons 2N_2O$

Given:

$N_2 = 0.482$ mol
$O_2 = 0.933$ mol
Volume = 10 L

$K_c = 2 \times 10^{-37}$ (very small)

Since Kc is extremely small → reaction hardly proceeds.

Therefore:

Equilibrium amounts approximately:

$N_2 \approx 0.482$ mol
$O_2 \approx 0.933$ mol
$N_2O \approx 0$