VIVEKANANDA MATRIC HIGHER SEC. SCHOOL, SIRKALI

2 AND 3 MARK QUESTIONS - INORGANIC XI CHEMISTRY

1. CHEMICAL CALCULATIONS

Textual Questions*:

1. By applying the knowledge of chemical classification, classify each of the following into elements, compounds or mixtures.
   a) Sugar b) Sea water c) Distilled water
   d) Carbon dioxide e) Copper wire f) Table salt
   g) Silver plate h) Naphthalene balls
2. Calculate average atomic mass from the following data: **(TB)**

Isotope	Isotopic Atomic mass	Abundance (%)
Mg^{24}	23.99	78.99
Mg^{25}	24.99	10.00
Mg^{26}	25.98	11.01

3. Define: Relative atomic masses, A_r. **(TB)**
4. Calculate the relative atomic mass of H and Glucose ($C_6H_{12}O_6$).
5. Calculate the molar mass of the following.
   a) Ethanol (C_2H_5OH) d) Sucrose ($C_{12}H_{22}O_{11}$)
   b) Potassium permanganate ($KMnO_4$)
   c) Potassium dichromate ($K_2Cr_2O_7$)
6. Which contains the greatest no. of moles of oxygen atoms? **(TB)**
   a) 1 mole of ethanol b) 1 mole of formic acid c) 1 mole of H_2O
7. Calculate the molar mass of the following compounds. **(TB)**
   a) Urea [$CO(NH_2)_2$] c) Boric acid [H_3BO_3]
   b) Acetone [CH_3COCH_3] d) Sulphuric acid [H_2SO_4]
8. What do you meant by 'mole'? Write Avogadro number. **(TB)**
9. Define: Equivalent mass. **(TB)**
10. Mention the difference between molecular mass and molar mass. Find molecular mass and molar mass of carbon monoxide. **(TB)**
11. Distinguish between oxidation and reduction. **(TB)**
12. What is oxidation number (Or) Oxidation state? **(TB)**

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Problems related to Mole concept:

1. How many moles of H_2 is required to produce 10 moles of NH_3?
2. Calculate the amount of H_2O produced by combustion of 32 g CH_4.
3. How much volume of carbon dioxide is produced when 50 g of solid calcium carbonate is heated under standard conditions?
4. How many moles of ethane is needed to produce 44 g of $CO_{2\,(g)}$ after combustion?
5. How much volume of chlorine is required to form 11.2 L of HCl at 273 K and 1 atm. pressure?
6. Calculate the percentage composition of the elements present in magnesium carbonate. How many kilogram of CO_2 can be obtained by heating 1 kg of 90 % pure magnesium carbonate?
7. Density of CO_2 is equal to 1.965 $kg.m^{-3}$ at 273 K and 1 atm. pressure. Calculate the molar mass of CO_2.

Problems related to Equivalent Mass:

1. Calculate equivalent mass: a) Zn in $ZnSO_4$ b) H_2SO_4 c) KOH.
2. Calculate the equivalent mass of Mn in $KMnO_4$ (in acid medium).
3. The equivalent mass of a trivalent metal element is 9 $g.eq^{-1}$. Find the molar mass of its anhydrous oxide.
4. If 0.456 g of a metal gives 0.606 g of its chloride. Calculate equivalent mass of the metal.
5. Calculate the equivalent mass of potassium dichromate. The reduction half-reaction in acid medium is,

$$Cr_2O_7^{2-} + 14\ H^+ + 6\ e^- \rightarrow 2\ Cr^{3+} + 7\ H_2O$$

Problems related to Limiting and Excess:

1. What is limiting and excess agent? **(TB)**
2. The balanced equation for a reaction is given below

$$2\ x + 3\ y \rightarrow 4\ \ell + m$$

When 8 moles of x react with 15 moles of y, then

a) Which is the limiting reagent?

b) Calculate the amount of products formed.

c) Calculate the excess reactant left at the end of reaction.

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3. In a process, 646 g of ammonia is allowed to react with 1.144 kg of CO_2 to form urea. Find the limiting and excess reagent in the reaction. The balanced equation: $2\ NH_3 + CO_2 \rightarrow H_2NCONH_2 + H_2O$
4. In a reaction $X + Y + Z_2 \rightarrow XYZ_2$. Identify the Limiting reagent if any, in the following reaction mixtures.
   a) 200 atoms of X + 200 atoms of Y + 50 molecules of Z_2
   b) 50 atoms of X + 25 atoms of y + 50 molecules of Z_2
   c) 2.5 mole of X +5 mole of Y + 5 mole of Z_2
5. Experimental analysis of a compound containing the elements x, y and z on analysis gave following data: x = 32 %, y = 24 %, z = 44 %. The relative no. of atoms of x, y and z are 2, 1 and 0.5, respectively. (Molecular mass of compound is 400 g). Find out
   a) The atomic masses of the element x, y and z.
   b) Empirical and Molecular formula of the compound.

Problems related to Molecular and Empirical formula:

1. What is the empirical formula of the following?
   a) Fructose ($C_6H_{12}O_6$) found in honey.
   b) Caffeine ($C_8H_{10}N_4O_2$) a substance found in tea and coffee.
2. An acid found in tamarinds on analysis shows following percentage composition: 32% Carbon; 4% Hydrogen; 64% Oxygen. Find the empirical formula of the compound.
3. An organic compound present in vinegar has 40 % carbon, 6.6 % hydrogen and 53.4 % oxygen. Find its Empirical formula.
4. A Compound on analysis gave the following percentage composition C=54.55%, H=9.09%, O=36.36%. Determine the empirical formula and molecular formula of the compound.
5. Calculate the empirical and molecular formula of a compound containing 76.6% carbon, 6.38 % hydrogen and rest oxygen its vapour density is 47.
6. A Compound on analysis gave Na = 14. 31% S = 9. 97% H = 6. 22% and O = 69.5%. Calculate molecular formula of the compound if all the hydrogen in the compound is present as water. (Molecular mass of the compound is 322).

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Problems related to Oxidation state:

1. Find the oxidation state of underlined element:
   a) $\underline{C}O_2$ b) $H_2\underline{S}O_4$ c) $\underline{C}H_2F_2$ d) $\underline{S}O_2$ e) $\underline{Cr_2}O_7^{2-}$
   f) $\underline{Mn}O_4^-$ g) $\underline{O}F_2$ h) $K\underline{O}_2$ i) $H_2\underline{O}_2$ j) $Na_2\underline{S}_2O_3$
   k) $K\underline{Cl}O_3$ l) Cl_2 m) HNO_3 n) Cr_2O_3
2. Arrange the following based on their reactivity – Justify.

   $ZnSO_4$, $AgNO_3$ and $CuSO_4$

Creative Questions:

1. Define: a) Matter b) Compound c) Element d) mixture*
2. Define: Molar mass and Molar volume.
3. How will you find the equivalent mass of an acid? Give example.
4. How will you find the equivalent mass of a base? Give example.
5. How to find equivalent mass of an oxidizing agent? Give example.
6. How to find equivalent mass of a reducing agent? Give example.
7. Define Stoichiometry.
8. Give day today example for redox reaction.
9. Write note on Combination reaction.
10. Write note on decomposition reaction.
11. Write note on displacement reaction.
12. Write note on Disproportionation reaction. Give example.**
13. Write note on competitive electron transfer reaction.
14. Predict the type of Following reactions:*
    a) $2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$
    b) $C + O_2 \rightarrow CO_2$
    c) $H_2O_2 \rightarrow 2\ H_2O + O_2$

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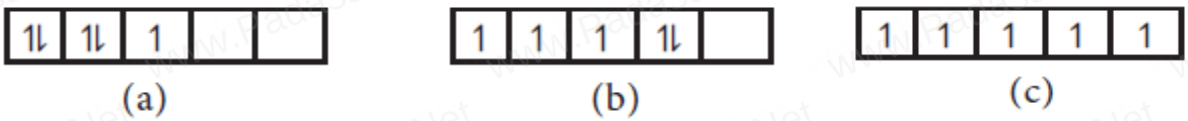
2. QUANTUM MECHANICAL MODEL OF AN ATOM

Textual Questions*:

1. An atom of an element contains 35 electrons and 45 neutrons. Deduce **(TB)**
   a) The number of protons
   b) The electronic configuration for the element
   c) All the four quantum numbers for the last electron.
2. Write the electronic configuration of the following and count the no. of unpaired electrons, also comment on their stability: **(TB)**
   a) Ni^{2+} b) Fe^{3+} c) Cr^{3+} d) Mn^{2+}
3. Show that the circumference of the Bohr orbit for H atom is an integral multiple of the de Broglie wavelength. **(TB)**
4. State and explain the following: **(TB)**
   a) Aufbau principle. b) Pauli's exclusion principle. c) Hund's rule
5. For the following, give sub level designation, allowable 'm' values and the number of orbitals. **(TB)**
   a) n = 4, ℓ =2 bi) n =5, ℓ = 3 c) n=7, ℓ =0
6. Define the terms: orbit and Orbital. **(TB)**
7. Which quantum number reveals the information about the shape, orientation and size of orbitals? **(TB)**
8. How many orbitals are possible for n =4? **(TB) (Practice)**
9. Calculate the number of electrons per shell, n= 1, 2, 3 and 4.
10. Account for the following: **(TB)**
    a) The stabilization of a half filled d - orbital is more pronounced than that of p-orbital.
    b) Consider following electron arrangements for d^5 configuration.



(a) (b) (c)

i) Which of these represents the Ground state?
ii) Which configuration has maximum exchange energy?

11. How many radial and Angular nodes possible for 2s, 4p, 3d, 4d and 4f orbitals? **(TB)**

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12. Find the 'n' and 'ℓ' values for the following orbitals: **(TB)**
    a) $3p_x$ and $4d_{x^2-y^2}$ electron. b) 4d, 5p and 3s orbitals.
13. Discuss time independent Schrodinger wave equation? **(TB)**
14. Explain the meaning of the symbol $4f^2$. Write all four quantum numbers for these electrons.
15. Determine the values of all the four quantum numbers of
    a) 8th electron in O atom.
    b) 15th electron in Cl atom and
    c) The last electron in chromium (Z = 24).

Problems related to Bohr's Atom Model:

1. The quantum mechanical treatment of the hydrogen atom gives the energy value:

$$E_n = -13.6/n^2 \text{ eV.atom}^{-1}$$

    i) Using this expression, find ΔE between n = 3 and n = 4.
    ii) Calculate the wavelength corresponding to the above transition.
2. Energy of an electron in H atom in ground state is -13.6 eV. What is the energy of the electron in the second excited state?
3. Li^{2+} ion is H like ion. Calculate the Bohr radius for the 3^{rd} orbit and the energy of an electron in 4^{th} orbit.
4. Calculate the energy required for the process.

$$He^{+}_{(g)} \rightarrow He^{2+}_{(g)} + e^-$$

    IE of H atom in its ground state is - 13.6 eV. $atom^{-1}$.

Problems related to de-Broglie Relation:

1. Calculate de Broglie wavelength in the following two cases:
    i) A 6.626 kg iron ball moving with 10 ms^{-1}
    ii) An electron moving at 72.73 ms^{-1}
2. Calculate the de-Broglie wavelength of an electron that has been accelerated from rest through a potential difference of 1 keV.
3. Find de Broglie wave length of an electron, which is accelerated from the rest, through a potential difference of 100V?
4. What is the de Broglie wavelength (in cm) of a 160g cricket ball travelling at 140 $Km.hr^{-1}$.

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5. How fast must a 54g tennis ball travel in order to have a de Broglie wavelength that is equal to that of a photon of green light 5400Å?
6. Protons can be accelerated in particle accelerators. Calculate the wavelength (in Å) of such accelerated proton moving at 2.85×10^{8} ms^{-1}, (Mass of proton is 1.673×10^{-27} Kg).

Problems related to Heisenberg's Principle:

1. Calculate uncertainty in the velocity of the electron in hydrogen atom. Assuming that the position of an electron in this orbit is determined with the accuracy of 0.5 % of the radius.
2. Calculate the uncertainty in the position of an electron, if the uncertainty in its velocity is 5.7×10^{5} ms^{-1}.
3. Suppose the uncertainty in determining the position of an electron in an orbit is 0.6 Å. What is its uncertainty in momentum?
4. Show that if the measurement of the uncertainty in the location of the particle is equal to its de Broglie wavelength, the minimum uncertainty in its velocity is equal to its velocity / 4π.

Creative Questions:

1. What did Rutherford's alpha ray scattering experiment revel?
2. Mention defects of Rutherford's model?
3. State: Bohr's Quantum conditions.*
4. Write Bohr's expression of radius of an orbit
5. Write note on limitations of Bohr's model of an atom*.
6. What are Zeeman and Stark Effect?**
7. What is Probability density?
8. Write the significance of de-Broglie equation.
9. Calculate angular momentum of electron in s, p, d and f orbitals.
10. State: Heisenberg's uncertainty principle.*
11. Arrange the orbitals in the increasing order of their energies:

    4p, 3s, 6s, 5d, 2s, 3p, 3d, 4f
12. Explain '$n+\ell$' rule with suitable example.
13. Define: Exchange energy. What is nodal surface or radial node?
14. Write the electronic configuration of Cr (z = 24) and Cu (z = 29).

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3. PERIODIC PROPERTIES OF ELEMENTS

Textual Questions*:

1. State: Modern Periodic law. **(TB)**
2. What is the basic difference in approach between Mendeleev's periodic table and modern periodic table?
3. The elements with atomic number 120, 123 and 126 have not been discovered so far. What would be their IUPAC name and symbol for these elements? Predict the possible electronic configuration.
4. Predict the position of an element in periodic table with electronic configuration: $(n-1)\ d^2\ ns^2$, where n = 5.
5. In what period and group will an element with atomic number, Z = 118 will be present? **(TB)**
6. Mention any two anomalous properties of 2^{nd} period elements. **(TB)**
7. Give the electronic configuration of lanthanides and actinides? **(TB)**
8. Using Slater's rule, calculate the effective nuclear charge on a 3p electron in Aluminium and chlorine. Explain how these results relate to the atomic radii of the two atoms.
9. What are iso-electronic ions? Give examples. **(TB)**
10. A student reported the ionic radii of isoelectronic species X^{3+}, Y^{2+} and Z^- as 136 pm, 64 pm and 49 pm respectively. Is that order correct?
11. What is screening effect? **(TB)**
12. What is Effective nuclear charge? **(TB)**
13. Explain periodic trend of ionization potential. **(TB)**
14. How will you explain the fact that the second ionization potential is always higher than first ionization potential?
15. Magnesium loses electrons successively to form Mg^+, Mg^{2+} and Mg^{3+} ions. Which step will have the highest ionization energy? Why? **(TB)**
16. Explain the following, give appropriate reasons. **(TB)**
    a) Ionization potential of N is greater than O.
    b) First ionization potential of C is greater B atom, whereas the reverse is true is for the second ionization potential.
    c) Electron affinity values of Be, Mg and noble gases are zero and those of N (0.02 eV) and P (0.80 eV) are very low.

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17. Account for the following: **(TB)**
    a) The electronic configuration of atom is one of the important factors which affect the value of ionization potential and electron gain enthalpy. Explain.
    b) Fifth period of the periodic table should have 18 elements on the basis of quantum numbers.
    c) Is the following definition correct?
    "Ionization enthalpy is the energy required to remove the most loosely bound electron from the valence shell of an atom".
18. Why is a halogens act as an Oxidizing agent? **(TB)**
19. Why is alkali metals act as a Reducing agent?
20. Briefly give basis for Pauling's scale of electro-negativity. **(TB)**
21. Give appropriate reason for high electron affinity of Noble gases and halogens in the periodic table.
22. Formation of $F^-_{(g)}$ from $F_{(g)}$ is exothermic while that of $O^{2-}_{(g)}$ from $O_{(g)}$ is endothermic – Reason out. **(TB)**
23. Define: Electronegativity. **(TB)**
24. Explain periodic trend of electronegativity in group and periods. **(TB)**
25. Elements a, b, c and d have following configurations: **(TB)**
    a: $1s^2, 2s^2, 2p^6$ b: $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$
    c: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ d: $1s^2, 2s^2, 2p^1$
    Which among these will belong to same group of periodic table?

Creative Questions:

1. What is Triad and Octaves?
2. Write note on Chancourtois report?
3. What is atomic and ionic radius?
4. What is covalent radius? Why it is shorter than actual atomic radius?
5. What is metallic radius?
6. Why first ionization enthalpy of Na is lower than that of Mg while its second ionization enthalpy is higher than that of Mg?
7. Define the following:
    a) Ionization energy b) Electron Gain enthalpy/Electron affinity.

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4. HYDROGEN

Textual Questions*:

1. Why H is not placed with halogen in the periodic table? Explain. **(TB)**
2. Justify the position of hydrogen in the periodic table. **(TB)**
3. What are isotopes? Write isotopes of hydrogen. **(TB)**
4. How do you convert Para hydrogen into ortho hydrogen? **(TB)**
5. What is water-gas shift reaction? **(TB)**
6. Explain preparation of hydrogen by electrolysis. **(TB)**
7. Give the uses of heavy water. **(TB)**
8. Do you think that heavy water is suitable for drinking? Give reason.
9. An ice cube at 0 ^{0}C is placed in some liquid water at 0 ^{0}C, ice cube sinks - Why? **(TB)**
10. What are the three types of covalent hydrides? Give example. **(TB)**
11. Predict the hydrides are a gas or a solid a) HCl b) NaH. **(TB)**
12. Write the expected formulas of hydrides of 4th period elements. In what way the first 2 members of the series different from others. **(TB)**
13. a) Give uses of Deuterium. **(TB)** b) Write exchange reactions of D.
14. Give reason for the following: **(TB)**
    a) Interstitial hydrides have a lower density than the parent metal.
    b) Metallic hydrides are useful for the hydrogen storage.
15. Account for the following: **(TB)**
    a) Arrange NH_3, H_2O and HF in the increasing order of H-bonding.
    b) NH_3 has very high melting and boiling point as compared to those of the hydrides of the remaining element of group 15.
16. Discuss the types of Hydrogen Bonding? (or) Differentiate between Inter and intra molecular H-Bonding.
17. Compare the properties of H_2O and H_2O_2.

Problems:

1. A group I metal (A) which is present in common salt reacts with (B) to give compound (C) in which hydrogen is present in –1 oxidation state. (B) on reaction with O_2 to give universal solvent (D). Compound (D) on reacts with (A) gives (E), a strong base. Identify A, B, C, D and E and explain the reactions.

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2. An isotope of hydrogen (A) reacts with diatomic molecule of element which occupies group number 16 and period number 2 to give compound (B) is used as a moderator in nuclear reaction. (A) adds on to a compound (C), which has the molecular formula C_3H_6 to give (D). Identify A, B, C and D.
3. Write chemical equation for following: **(TB)**
   a) Hydrogen gas and chlorine gas.
   b) How Na metal reacts with Acetylene gas?
   c) i) $Li + H_2 \rightarrow ?$ ii) $Ca + H_2 \rightarrow$
   d) In what way water reacts with Na, Ca and Fe? Mention the required conditions?
   e) Reaction of hydrogen with the following:
      i) Tungsten (VI) oxide ii) CuO on heating. iii) Acetylene.
4. Complete the Following: **(TB)**
   a) $KMnO_4 + H_2O_2 \rightarrow$
   b) $CrCl_3 + H_2O \rightarrow$
   c) $CaO + H_2O \rightarrow$

Creative Questions:

1. How is water gas or syngas prepared?*
2. The electron affinity of H is positive while that of Br is negative. Why?
3. Compare the isotopes of Hydrogen.
4. How is Tritium prepared? Mention its properties.
5. Write any three uses of Hydrogen*.
6. Compare Para and Ortho hydrogen.
7. What are soft water and hard water? Which is easily removable one?
8. How will you remove temporary hardness of water?*
9. How is temporary hardness removed by Clark's method?
10. Suggest a method to remove harness of water?
11. Hard water produces less foam with detergents. Why?*
12. How Hydrogen peroxide prepared industrially?
13. Mention the uses of hydrogen peroxide.*
14. Give the importance of Hydrogen bonding.

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5. ALKALI AND ALKALINE EARTH METALS

Textual Questions*:

1. Why NaOH is much more water soluble than NaCl? **(TB)**
2. What is meant by efflorescence? **(TB)**
3. Write the important common features of Group 2 elements. **(TB)**
4. Write the chemical equations involved in Solvay process of preparation of sodium carbonate. **(TB)**
5. An alkali metal (X) forms a hydrated sulphate, X_2SO_4. 10 H_2O. Is metal more likely to be sodium or potassium? Why? **(TB)**
6. Substantiate with reason: **(TB)**
   a) Lithium fluoride has the lowest solubility among Group 1 fluorides.
   b) Beryllium halides are Covalent while magnesium halides are ionic.
   c) Alkaline earth metals are harder than alkali.
7. Which would you expect to have a higher melting point, magnesium oxide or magnesium fluoride? Explain your reasoning
8. Give the systematic names for the following: **(TB)**
   a) Milk of magnesia. b) Lye c) Lime d) Trona
   e) Washing Soda f) Soda ash f) Caustic Potash
9. How do you prepare Plaster of Paris? Give its uses. **(TB)**
10. Write any six uses of Gypsum. **(TB)**

Problem:

1. Alkaline earth metal (A) belongs to 3rd period reacts with oxygen and nitrogen to form compound (B) and (C) respectively. It undergoes metal displacement reaction with $AgNO_3$ to form compound (D). Find A, B, C and D. Explain the reaction.

Creative Questions:

1. List down the uses of washing soda.
2. Give the uses of sodium bicarbonate.
3. What is quick and slaking of lime? Write their formula.
4. List out bleaching powder prepared?
5. How is Lithium different from its rest of its group members?
6. How is Beryllium different from its rest of its group members?

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7. Discuss briefly the similarities between lithium and magnesium
8. Mention the uses of Be and Mg.
9. Mention the uses of Ca, r and Ba.
10. Why is the alkali and alkaline metals (their salts) show colour upon heating?
11. Match the flame colours of the alkali and alkaline earth metal salts in the bunsen burner

a) Sodium	(1) Brick red
b) Potassium	(2) Blue
c) Calcium	(3) Yellow
d) Barium	(4) Violet
e) Strontium	(5) Apple green
f) Caesium	(6) Crimson red

12. What helps the cake to be so soft and fluffy? Give reason.
13. What is dead burnt plaster?

3 MARK QUESTIONS - PHYSICAL CHEMISTRY

6. GASEOUS STATE

Textual Questions*:

1. What are the gases exist as gases at room temperature?
2. State: a) Boyle's and Charles law. b) Avogadro's hypothesis. c) Gay Lussac's law. **(TB)**
3. Write notes on Dalton's law of partial pressure.
4. Distinguish between diffusion and effusion. **(TB)**
5. A balloon filled with air at room temperature and cooled to a much lower temperature can be used as a model for Charles law? **(TB)**
6. Name two items that serve as a model for the Gay Lussac's law and explain. **(TB)**
7. Give the mathematical expression that relates gas volume and moles. Describe in words. (or) Avogadro's law. **(TB)**
8. What are ideal gases? In what way real gases differ from it. **(TB)**
9. Can a Van der Waals gas with a = 0 be liquefied? Explain. **(TB)**

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10. Would it be easier to drink water with a straw on the top of Mount Everest? **(TB)**
11. Aerosol cans carry clear warning of heating of the can. Why? **(TB)**
12. Suppose there is a tiny sticky area on the wall of a container of gas. Molecules hitting this area stick there permanently. Is pressure greater or lesser than on the ordinary area of walls? **(TB)**
13. Explain the following observations: **(TB)**
    a) Aerated water bottles are kept under water during summer.
    b) Liquid ammonia bottle is cooled before opening the seal.
    c) Tyre of an automobile is inflated to slightly lesser pressure in summer than in winter
    d) Size of a weather balloon becomes larger and larger as it ascends up into the larger altitude.
14. Give suitable explanation for the following facts about gases: **(TB)**
    a) Gases don't settle at the bottom of a container.
    b) Gases diffuse through all the space available to them.
15. Suggest why there is no hydrogen (H_2) in our atmosphere. Why does the moon have no atmosphere?
16. Explain whether a gas approaches ideal behavior or deviates from ideal behaviour if **(TB)**
    a) It is compressed to a smaller volume at constant temperature.
    b) The temperature is raised at while keeping the volume constant
    c) More gas is introduced into the same volume and at the same temperature
17. Which of the following gases would you expect to deviate from ideal behaviour under conditions of low temperature F_2, Cl_2 or Br_2? Explain.
18. When the driver of an automobile applies brake, the passengers are pushed toward the front of car but a helium balloon is pushed toward back of the car. Upon forward acceleration the passengers are pushed toward the front of the car. Why? **(TB)**
19. Why Astronauts wear protective suits when they are on moon? **(TB)**
20. When ammonia combines with HCl, NH_4Cl is formed as white dense fumes. Why do more fumes appear near HCl? **(TB)**

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Problems:

1. A sample of gas at 15 °C at 1 atm. has a volume of 2.58 dm^3. When the temperature is raised to 38 °C at 1 atm. does the volume of the gas increase? If so, calculate the final volume.
2. A sample of gas has a volume of 8.5 dm^3 at an unknown temperature. When the sample is submerged in ice water at 0 °C, its volume gets reduced to 6.37 dm^3. What is its initial temperature?
3. Of two samples of nitrogen gas, sample A contains 1.5 moles of nitrogen in a vessel of volume of 37.6 dm^3 at 298 K, and the sample B is in a vessel of volume 16.5 dm^3 at 298 K. Calculate the number of moles in sample B.
4. Sulphur hexafluoride is a colourless, odourless gas; calculate the pressure exerted by moles of the gas in a vessel of volume 5.43 dm^3 at 69.5 °C, assuming ideal gas behaviour.
5. Argon is an inert gas used in light bulbs to retard the vaporization of the tungsten filament. A certain light bulb containing argon at 1.2 atm. and 18 °C is heated to 85 °C at constant volume. Calculate its final pressure in atm.
6. A small bubble rises from the bottom of a lake where the temperature and pressure are 6 °C and 4 atm. to the water surface, where the temperature is 25 °C and pressure is 1 atm. Calculate the final volume in (mL) of the bubble, if its initial volume is 1.5 mL.
7. Hydrochloric acid is treated with a metal to produce hydrogen gas. Suppose a student carries out this reaction and collects a volume of 154. 4 × 10^{-3} dm^3 of a gas at a pressure of 742 mm of Hg at a temperature of 298 K. What mass of H gas (in mg) did the student collect?
8. It takes 192 sec for an unknown gas to diffuse through a porous wall and 84 sec for N_2 gas to effuse at the same temperature and pressure. What is the molar mass of the unknown gas?
9. A tank contains a mixture of 52.5 g of oxygen and 65.1 g of CO_2 at 300 K the total pressure in the tanks is 9.21 atm. Calculate partial pressure (in atm.) of each gas in the mixture.
10. A combustible gas is stored in a metal tank at a pressure of 2.98 atm. at 25 °C. The tank can withstand a maximum pressure of 12 atm. after

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which it will explode. The building in which the tank has been stored catches fire. Now predict whether the tank will blow up first or start melting? (Melting point of the metal = 1100 K).

Creative Questions:

1. Mention the differences between gas and vapour.
2. What is Joule Thomson effect?*
3. What is inversion temperature?*
4. What are the different methods used for liquefaction of gases?
5. What is Boyle temperature or Boyle point?
6. Define absolute zero or Kelvin scale.*
7. Mention the application of Dalton's law.
8. Write note on Graham's law of diffusion.
9. Write Vander Waal's equation and explain the terms.
10. Show that Excluded volume of single molecule is four times of volume of single molecule.
11. What is compressibility factor, Z?
12. Write the value of critical temperature (T_c), critical pressure (P_c) and critical volume (V_c).*

7. THERMODYNAMICS

Textual Questions*:

1. State: First law of thermodynamics. **(TB)**
2. State: Third law of thermodynamics. **(TB)**
3. Define Hess's law of constant heat summation. **(TB)**
4. Explain intensive properties with two examples. **(TB)**
5. What is state and path functions? Give two examples. **(TB)**
6. State Kelvin statement of second law of thermodynamics. **(TB)**
7. What is lattice energy?
8. Write down the Born-Haber cycle for the formation of $CaCl_2$. **(TB)**
9. Identify the state and path functions out of the following: **(TB)**

a) Enthalpy b) Entropy c) Heat
d) Temperature e) Work f) Free energy.

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10. Define the following terms: **(TB)**
    a) Isothermal process b) Adiabatic process
    c) Isobaric process d) Isochoric process
11. What is the usual definition of entropy? What is the unit of entropy?
12. Predict the feasibility of a reaction when **(TB)**
    i) Both ΔH and ΔS positive
    ii) Both ΔH and ΔS negative
    iii) ΔH decreases but ΔS increases
13. Define: Entropy. Give its unit. **(TB)**
14. Define: Gibb's free energy. List out it's Characteristics. **(TB)**
15. Define enthalpy of combustion. **(TB)**
16. Define: Molar heat capacity. Give its unit. **(TB)**
17. Define the calorific value of food. What is its unit? **(TB)**
18. Define enthalpy of neutralization. **(TB)**
19. What are spontaneous reactions? What are the essential conditions for spontaneity of a process? **(TB)**
20. The equilibrium constant of a reaction is 10, what will be the sign of ΔG? Will this reaction be spontaneous?
21. Enthalpy of neutralization is always a constant when a strong acid is neutralized by a strong base: account for the statement.
22. List the characteristics of internal energy. **(TB)**

Problems:

1. Calculate the work done when 2 moles of an ideal gas expands reversibly and isothermally from a volume of 500 ml to a volume of 2 L at 25°C and normal pressure.
2. In a constant volume calorimeter, 3.5 g of a gas with molecular weight 28 was burnt in excess oxygen at 298 K. The temperature of the calorimeter was found to increase from 298 K to 298.45 K due to the combustion process. Given that the calorimeter constant is 2.5 kJ K . Calculate the enthalpy of combustion of the gas in kJ mol^{-1}.
3. Calculate the entropy change in the system, and surroundings, and the total entropy change in the universe during a process in which 245 J of

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heat flow out of the system at 77°C to the surrounding at 33°C.

4. 1 mole of an ideal gas, maintained at 4.1 atm. and at a certain temperature, absorbs heat 3710 J and expands to 2 litres. Calculate the entropy change in expansion process.
5. 30.4 kJ is required to melt one mole of sodium chloride. The entropy change during melting is 28.4 JK^{-1} mol^{-1}. Calculate the melting point of sodium chloride.
6. Calculate the standard heat of formation of propane, if its heat of combustion is −2220.2 kJ mol^{-1}. The heats of formation of $CO_{2(g)}$ and $H_2O_{(l)}$ are −393.5 and −285.8 kJ mol^{-1} respectively.
7. You are given normal boiling points and standard enthalpies of vapourisation. Calculate Entropy of vapourisation for listed liquids:

S. No	Liquid	Boiling points (^{0}C)	ΔH (kJ mol^{-1})
1.	Ethanol	78.4	+ 42.4
2.	Toluene	110.6	+ 35.2

8. For the reaction $Ag_2O_{(s)} \rightarrow 2\ Ag_{(s)} + 1/2\ O_{2\ (g)}$: $\Delta H = 30.56$ kJ mol^{-1} and $\Delta S = 6.66 JK^{-1}mol^{-1}$ (at 1 atm.). Calculate the temperature at which ΔG is equal to zero. Also predict the direction of the reaction
(i) at this temperature and (ii) below this temperature.

9. What is the equilibrium constant Keq for the reaction at 400K.

$$2\ NOCl_{(g)} \rightarrow 2\ NO_{(g)} + Cl_{2\ (g)},$$

Given that $\Delta H0 = 77.2$kJ mol^{-1}; and $\Delta S^0 = 122 JK^{-1}mol^{-1}$.

10. Cyanamide (NH_2CN) is completely burnt in excess oxygen in a bomb calorimeter, ΔU was found to be −742.4 kJ mol^{-1}, calculate enthalpy change of the reaction at 298K.

$$NH_2CN_{(s)} + 3/2\ O_{2(g)} \rightarrow N_{2\ (g)} + CO_{2\ (g)} + H_2O_{(l)} \qquad \Delta H = ?$$

11. Calculate the enthalpy of hydrogenation of ethylene from the following data. Bond energies of C - H, C – C , C = C and H – H are 414, 347, 618 and 435 kJ mol^{-1}.

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12. Calculate the lattice energy of CaCl2 from the given data

$$Ca_{(s)} + Cl_{2\,(g)} \rightarrow CaCl_{2\,(s)} \quad \Delta H_f^0 = -795 \text{ kJ mol}^{-1}$$

Atomization: $Ca_{(s)} \rightarrow Ca_{(g)}$ $\Delta H_1^0 = +121$ kJ mol−1

Ionisation: $Ca_{(g)} \rightarrow Ca^{2+}_{(g)} + 2e^-$ $\Delta H2^0 = +2422$ kJ mol−1

Dissociation: $Cl_{2\,(g)} \rightarrow 2Cl_{(g)}$ $\Delta H3^0 = +242.8$ kJ mol−1

Electron affinity: $Cl_{(g)} + e^- \rightarrow Cl^-_{(g)}$ $\Delta H4^0 = -355$ kJ mol−1

13. Calculate enthalpy change for reaction $Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$ from the following data.

a) 2 Fe + 3/2 O2 → Fe2O3; ΔH = −741 kJ

b) C + 1/2 O2 → CO; ΔH = −137 kJ

c) C + O2 → CO2; ΔH = − 394.5 kJ

14. When 1-pentyne (A) is treated with 4N alcoholic KOH at 175°C, it is converted slowly into an equilibrium mixture of 1.3% 1-pentyne (A) , 95.2% 2-pentyne (B) and 3.5% of 1, 2 pentadiene (C) the equilibrium was maintained at 175 °C, calculate ΔG^0 for the following equilibria.

15. At 33 K, N_2O_4 is fifty percent dissociated. Calculate the standard free energy change at this temperature and at one atmosphere.

16. The standard enthalpies of formation of SO_2 and SO_3 are −297 kJ mol^{-1} and −396 kJ mol^{-1} respectively. Calculate the standard enthalpy of reaction for the reaction:

$$SO_2 + 1/2\ O_2 \rightarrow SO_3$$

17. For the reaction at 298 K: 2A +B → C.

ΔH=400 J mol^{-1} ; ΔS = 0.2 JK^{-1} mol^{-1} Determine the temperature at which the reaction would be spontaneous.

18. Find out the value of equilibrium constant for the following reaction at 298K, $2\ NH_{3\,(g)} + CO_{2\,(g)} \rightarrow NH_2CONH_{2\,(aq)} + H_2O_{(l)}$. Standard Gibbs energy change, ΔG_r^0 at the given temperature is −13.6 kJ mol^{-1}.

19. A gas mixture of 3.67 lit of ethylene and methane on complete combustion at 25 °C and at 1 atm. pressure produces 6.11 lit of carbondioxide. Find out the amount of heat evolved in kJ, during this combustion. ($\Delta H_C(CH_4)$= − 890 kJ mol^{-1} and ($\Delta HC(C_2H_4)$= −1423 kJ mol^{-1}

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Creative Questions:

1. What are system, surrounding and boundary?
2. What are the types of systems depending on the nature of boundary?
3. What is reversible and irreversible process?
4. What are the applications of Hess law.*
5. State Zeroth law of thermodynamics.*
6. What are the applications of bomb calorimeter?
7. What is standard entropy of formation?*
8. Derive the relationship between standard free energy change and the equilibrium constant
9. State Claussius statement of second law of thermodynamics.
10. Define the terms: a) Entropy of fusion. b) Entropy of vapourization. c) Entropy of transition
11. What are the various statements of first law of thermodynamics?**

8. PHYSICAL AND CHEMICAL EQUILIBRIUM

Textual Questions*:

1. If there is no change in the concentration, why is chemical equilibrium state considered as dynamic? **(TB)**
2. State: Law of mass action. **(TB)**
3. Define: Equilibrium constant, K. **(TB)**
4. Write equilibrium constant for the following: **(TB)**
   a) $H_2O_{2\,(g)} \rightleftarrows H_2O_{(g)} + \frac{1}{2} O_{2\,(g)}$
   b) $2\,SO_{2\,(g)} + O_{2\,(g)} \rightleftarrows 2\,SO_{3\,(g)}$
   c) $2\,CO_{(g)} \rightleftharpoons CO_{2\,(g)} + C_{(s)}$
   d) $CO_{(g)} + H_2O_{(g)} \rightleftarrows CO_{2\,(g)} + H_{2\,(g)}$
5. What is the relation between K_P and K_C? Give an example for which K_P is equal to K_C. **(TB)**
6. For a gaseous reaction at equilibrium, number of moles of the products is greater than that of reactants. Is K_C is larger or smaller than K_P? **(TB)**
7. Write a balanced equation for an equilibrium reaction for which the equilibrium constant is given by expression: **(TB)**

$$K_c = \frac{[NH_3]^4}{[NO]^4} \frac{[O_2]^5}{[H_2O]^6}$$

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8. Define: Reaction Quotient, Q. **(TB)**
9. How will you predict the direction of equilibrium reaction with the help of Reaction Quotient, Q? **(TB)**
10. When the value of reaction quotient is greater than equilibrium constant, in which direction does reaction proceed to reach equilibrium? **(TB)**
11. State: Le Chatelier's principle. **(TB)**
12. What is the effect of adding inert gas (Say Argon) on the chemical reaction at equilibrium? **(TB)**
13. Consider the following reactions, **(TB)**

a) $H_{2\,(g)} + I_{2\,(g)} \rightleftharpoons 2\,HI$

b) $CaCO_{3\,(s)} \rightleftharpoons CaO_{\,(s)} + CO_{2\,(g)}$ c) $S_{\,(s)} + 3\,F_{2\,(g)} \rightleftharpoons SF_{6\,(g)}$

In each of the above reaction find out whether you have to increase or decrease the volume to increase the yield of the product.

Problems:

1. One mole of PCl_5 is heated in one litre closed container. If 0.6 mole of chlorine is found at equilibrium, calculate the value of equilibrium constant.
2. For the reaction: $SrCO_{3\,(s)} \rightleftharpoons SrO_{\,(s)} + CO_{2\,(g)}$, the value of equilibrium constant $K_P = 2.2 \times 10{-}4$ at 1002 K. Calculate K_C for the reaction.
3. To study the decomposition of hydrogen iodide, a student fills an evacuated 3 litre flask with 0.3 mol of HI gas and allows the reaction to proceed at 500 0C. At equilibrium he found the concentration of HI which is equal to 0.05 M. Calculate K_C and K_P.
4. Oxidation of nitrogen monoxide was studied at 200 ^{0}C with initial pressures of 1 atm. NO and 1 atm. of O_2. At equilibrium, the partial pressure of oxygen is found to be 0.52 atm. calculate K_P value.
5. 1 mol of CH_4, 1 mole of CS_2 and 2 mols of H_2S are 2 mol of H_2 are mixed in a 500 ml flask. Equilibrium constant for the reaction $K_C = 4 \times 10^{-2}$ mol^2 lit^{-2}. In which direction will the reaction proceed to reach equilibrium?
6. At particular T, $K_C = 4 \times 10^{-2}$ for the reaction: $H_2S_{\,(g)} \rightleftharpoons H_{2\,(g)} + \frac{1}{2}\,S_{2\,(g)}$ Calculate K_C for each of the following reaction:

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a) $2\ H_2S_{(g)} \rightleftharpoons 2\ H_{2\ (g)} + S_{2\ (g)}$

b) $3\ H_2S_{(g)} \rightleftharpoons 3H_{2\ (g)} + 3/2\ S_{2\ (g)}$

7. 28 g of Nitrogen and 6 g of hydrogen were mixed in a 1 litre closed container. At equilibrium 17 g NH_3 was produced. Calculate the weight of nitrogen, hydrogen at equilibrium.
8. The equilibrium for the dissociation of XY2 is given as,

$$2\ XY_{2\ (g)} \rightleftharpoons 2\ XY_{(g)} + Y_{2\ (g)}$$

If the degree of dissociation x is so small compared to one. Show that $2\ K_P = PX_3$ where P is the total pressure and K_P is the dissociation equilibrium constant of XY_2.

9. A sealed container was filled with 1 mol of $A_{2\ (g)}$, 1 mol $B_{2\ (g)}$ at 800 K and total pressure 1.00 bar. Calculate the amounts of components in the mixture at equilibrium given that K = 1 for the reaction $A_{2\ (g)} + B_{2\ (g)} \rightleftharpoons 2\ AB_{(g)}$

Creative Questions:

1. Give an example and explain the Solid –liquid physical equilibrium.
2. Distinguish between Physical and Chemical equilibrium.
3. What happens when Δng = 0, Δng = +ve, Δng = -ve in a gaseous reaction?
4. Derive relationship between formation and dissociation equilibrium constant? Give one example.
5. Dissociation of PCl_5 decreases in the presence of increase in Cl_2 why?

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3 MARK QUESTIONS - ORGANIC CHEMISTRY

11. FUNDAMENTALS OF ORGANIC CHEMISTRY

Textual Questions*:

1. Write the properties of Organic compounds. **(TB)**
2. What is homologous series? Give example. **(TB)**
3. Write molecular formula of the first six members of homologous series of nitro alkanes. **(TB)**

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4. What are functional groups? Identify functional group present in the following compounds: **(TB)**
   a) Acetaldehyde c) Oxalic acid
   b) Dimethyl ether d) Methyl amine
5. Write structural formula of first four members of homologous series of carboxylic acids. **(TB)**
6. Give general formula for the following classes of organic compounds:
   a) Aliphatic monohydric alcohols.
   b) Aliphatic ketones. c) Aliphatic amines. **(TB)**
7. Describe reaction involved in the detection of Nitrogen by Lassaigne's method. **(TB)**

Problems related to Estimation of Carbon and Hydrogen:

1. 0.26 g of an organic compound gave 0.039 g of water and 0.245 g of carbon di oxide on combustion. Calculate percentage of Carbon and Hydrogen.
2. 0.2346 g of organic compound containing C, H and O, on combustion gives 0.2754 g of water and 0.4488 g carbon di oxide. Calculate percentage of C, H and O.
3. 0.30 g of a substance gives 0.88 g of carbon di-oxide and 0.54 g of water. Calculate percentage of carbon and hydrogen in it. **(TB)**

Problems related to Estimation Sulphur:

1. 0.32 g of an organic compound, after heating with fuming nitric acid and barium nitrate crystals is a sealed tube game 0.466 g of barium sulphate. Determine the percentage of S in the compound. **(TB)**
2. In an estimation of sulphur by Carius method, 0.2175 g of the substance gave 0.5825 g of $BaSO_4$ Calculate the percentage of Sulphur in the compound.
3. 0.16 g of an organic compound was heated in a Carius tube and H_2SO_4 acid formed was precipitated with $BaCl_2$. The mass of $BaSO_4$ was 0. 35 g. Find the percentage of S.

Problems related to Estimation of Halogens:

1. 0.24g of an organic compound gave 0.287 g of silver chloride in the Carius method. Calculate the percentage of chlorine. **(TB)**

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2. 0.284 g of an organic substance gave 0.287 g AgCl in a Carius method for the estimation of halogen. Find the Percentage of Cl in it.
3. 0.185 g of an organic compound when treated with Conc. HNO3 and silver nitrate gave 0.320 g of silver bromide. Calculate percentage of bromine in the compound.
4. 0.40 g of an iodo-substituted organic compound gave 0.235 g of AgI by Carius method. Calculate the percentage of iodine in the compound.

Problems related to Estimation of Halogens

1. 0.24 g of organic compound containing Phos-phorous gave 0.66 g of $Mg_2P_2O_7$ by the usual analysis. Calculate the percentage of P in the compound
2. 0.33 g of organic compound containing Phosphorous gave 0.397 g of $Mg_2P_2O_7$ by the usual analysis. Calculate the percentage of P in the compound.

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5 MARK QUESTIONS - INORGANIC XI CHEMISTRY

1. CHEMICAL CALCULATIONS

Textual Questions*:

1. Balance by oxidation number method: **(TB)**
   a) $K_2Cr_2O_7 + KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + I_2 + H_2O$
   b) $KMnO_4 + Na_2SO_3 \rightarrow MnO_2 + Na_2SO_4 + KOH$
   c) $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$
   d) $KMnO_4 + H_2C_2O4 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + CO_2 + H_2O$
   e) $FeSO_4 + KMnO_4 + H_2SO_4 \rightarrow Fe_2(SO_4)_3 + MnSO_4 + K_2SO_4 + H_2O$
   f) $As_2S_3 + HNO_3 + H_2O \rightarrow H_3AsO_4 + H_2SO_4 + NO$
2. Balance by ion-electron method: **(TB)**
   a) $KMnO_4 + SnCl_2 + HCl \rightarrow MnCl_2 + SnCl_4 + H_2O + KCl$
   b) $C_2O_4^{2-} + Cr_2O_7^{2-} \rightarrow Cr^{3+} + CO_2$ (acid medium)
   c) $Na_2S_2O_3 + I_2 \rightarrow Na_2S_4O_6 + NaI$ (acid medium)
   d) $Zn + NO_3^- \rightarrow Zn^{2+} + NO$
   e) $FeSO_4 + KMnO_4 + H_2SO_4 \rightarrow Fe_2(SO_4)_3 + MnSO_4 + K_2SO_4 + H_2O$

Created Questions:

1. What is Matter? Explain how it is classified in detail.
2. Explain the different types of Redox reactions with an example.
3. How will you find the equivalent mass of
   a) An acid? Give example.
   b) A base? Give example.
   c) An oxidizing agent? Give example.
   d) A reducing agent? Give example.

2. QUANTUM MECHANICAL MODEL OF AN ATOM

Textual Questions*:

1. Account for the following: **(TB)**
   Electronic configuration of chromium is [Ar] $3d^5$ $4s^1$ but not [Ar] $3d^4$ $4s^2$ – Explain based on the Concept of Exchange energy.
2. Explain briefly the time independent Schrodinger wave equation. **(TB)**
3. Discuss the four Quantum numbers in detail.
4. Derive de-Broglie Relation. Give its significance. **

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5. Identify the missing quantum numbers and the sub energy level **(TB)**

n	ℓ	m	Sub energy level
?	?	0	4d
3	1	0	?
?	?	?	5p
?	?	-2	3d

Created Questions:

1. Write the Postulates of Bohr's atom Model. **
2. Explain the shapes of s, p and d orbitals.
3. Discuss Davisson and Germer experiment. (or) Prove wave nature of an electron with an experiment.

3. PERIODIC PROPERTIES OF ELEMENTS

Textual Questions*:

1. Explain Pauling method for the determination of ionic radius.
2. By using Pauling's method calculate the ionic radii of K^+ and Cl^- ions in KCl. Given that $d_{K^+-Cl^-} = 3.14\ A^0$.
3. a) Explain the periodic trend of Atomic or ionic radius.
   b) Calculate the covalent radius of H using the experimental d_{H-Cl} value is 1.28 Å and the covalent radius of chlorine is 0.99 Å.
   c) Calculate Effective nuclear charge of 4s and 3d electron of Scandium.

Created Questions:

1. Explain the factors affecting ionization energy.
2. Explain the factors affecting Electron gain enthalpy.
3. Differentiate Electron gain enthalpy and Electronegativity.
4. Explain Slater rule to calculate screening constant with an example.

4. HYDROGEN

Textual Questions*:

1. Hydrogen peroxide can function as an oxidizing agent and reducing agent. Substantiate this statement with suitable examples. **(TB)**
2. Explain the different types of hydrides with suitable example.
3. Compare the isotopes of Hydrogen with their Physical properties.
4. What are Ortho and Para Hydrogen? Differentiate them accordingly.

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Created Questions:

1. List out the importance of Hydrogen.
2. Explain the way to remove Temporary and Permanent hardness of the water.
3. List out the importance of Hydrogen bonding in day today life.
4. Explain the types of H-bonding in detail. Comment on the substance's solubility and boiling point.

5. ALKALI AND ALKALINE EARTH METALS

Textual Questions*:

1. Discuss briefly the similarities between Beryllium and Aluminum.
2. Describe briefly biological importance of Calcium and magnesium.
3. Write the balanced chemical equation for the following processes:

   I) Heating

   a) Calcium in oxygen.
   b) Calcium hydrogen carbonate.
   c) Sodium bi carbonate.
   d) Calcium Carbonate.
   e) Calcium Oxide with Carbon.

   II) Reaction of Alkali's

   a) Lithium with nitrogen gas.
   b) Rubidium with oxygen gas.
   c) Solid KOH with CO2.

Created Questions:

1. In what way Lithium is different from its rest of members.
2. Discuss the similarities between Lithium and magnesium.
3. Explain Sodium potassium pump in detail.
4. Write the uses of alkaline earth metals in different fields.
5. How is sodium hydroxide prepared commercially from brine solution?

5 MARK QUESTIONS - PHYSICAL CHEMISTRY

6. GASEOUS STATE

Textual Questions*:

1. Derive the values of T_c, P_c and V_c from van der Waals constants.
2. Write the Van der Waals equation for a real gas. Explain the correction term for pressure and volume

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Created Questions:

1. Explain Andrews Isotherm with neat diagram.
2. Derive the value of Gas constant R, with different units.
3. a) State and derive Dalton's law of Partial pressure.
   b) Derive equation of state.

7. THERMODYNAMICS

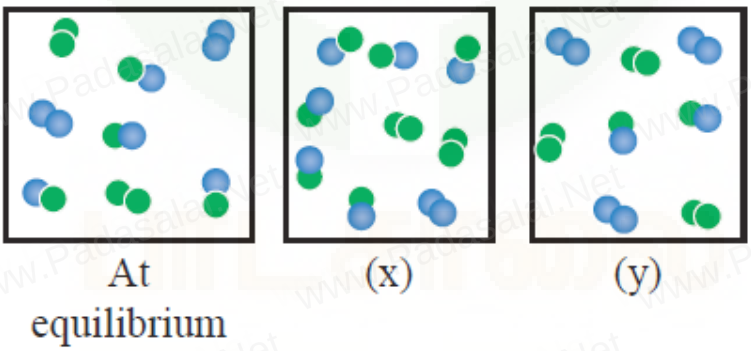
Textual Questions*:

1. State various statements of second law of thermodynamics.
2. Write down the Born-Haber cycle for the formation of $CaCl_2$.
3. List out the Characteristics of Gibbs free energy, G.
4. List out the Characteristics of Entropy, S.

8. PHYSICAL AND CHEMICAL EQUILIBRIUM

Textual Questions*:

1. Derive the general relation between Kp and Kc.
2. Derive the general expression of Kp and Kc for the formation of NH_3.
3. For the reaction: $A_{2\,(g)} + B_{2\,(g)} \rightleftharpoons 2\,AB_{(g)}$ ΔH -ve.
   Following molecular scenes represent different reaction mixture (A – Green, B – Blue)



   a) Write equilibrium constant K_P and K_C.
   b) For the above reaction mixture represented by scene (x) and (y), the reaction proceeds in which directions?
   b) What is the effect of increase in pressure for the mixture at equilibrium?
4. Deduce the Van't Hoff equation.

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5. Define: Reaction Quotient. Explain how you will predict the direction of equilibrium with the help of Q.

Created Questions:

1. Derive the general expression of Kp and Kc for the formation of HI.
2. Derive the general expression of Kp and Kc for the dissociation of PCl_5.
3. For the ammonia formation reaction, Apply Le-Chatelier's principle.
4. Discuss the effect of P, C and T on the reaction:

$$N_{2g)} + O_{2(g)} \rightleftarrows 2NO_{(g)}$$

5 MARK QUESTIONS - ORGANIC CHEMISTRY

11. FUNDAMENTALS OF ORGANIC CHEMISTRY

Textual Questions*:

1. Explain the classification of Organic compounds. **(TB)**
2. Explain the types of constitutional (or structural) isomerism in organic compounds. **(TB)**
3. a) What is Optical isomerism? Give suitable example. **(TB)**
   b) Mention the condition for a compound to be optically active.
4. Write the principle involved in the estimation of halogens by Carius method. **(TB)**
5. Briefly explain geometrical isomerism in alkene by considering 2-butene as an example. (Or) Explain Cis-Trans isomerism in detail. **(TB)**
6. Explain paper chromatography.
7. Give a brief description of the principles of
   i) Fractional distillation.
   ii) Column Chromatography.

Created Questions:

1. Explain tautomerism in detail.
2. Discuss Carius method to estimate Sulphur.
3. Discuss Carius method to estimate Phosphorous.
4. Discuss estimation of Carbon and Hydrogen by Liebig's method.
5. Explain Kjeldhal method to estimate Nitrogen.

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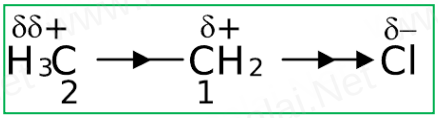
12. BASIC CONCEPTS OF ORGANIC REACTION

(3 Marks)

1. Explain Inductive effect with suitable Example. (TB)

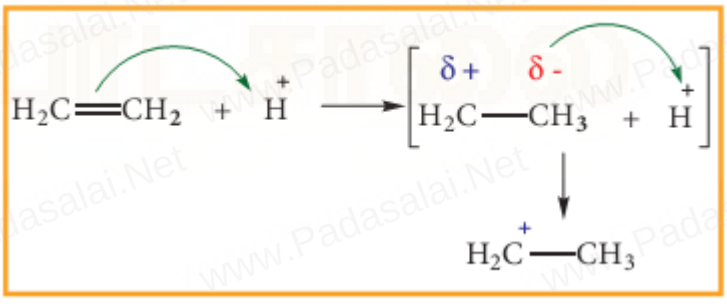
- Inductive effect is defined as the change in the polarization of a covalent bond due to presence of adjacent bonds, atoms or groups in a molecule.
- This is a permanent phenomenon.
- This involves sigma bonds.



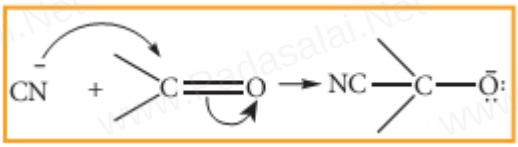
- Positive Inductive effect:
  - Atoms or Groups attached to Carbon that release electron through σ covalent bond are said to have +I effect.
  - Eg: $-CH_3$, $-CH_2CH_3$, $-CH(CH_3)_2$, $-C(CH_3)_3$
- Negative Inductive effect:
  - Atoms or Groups attached to Carbon that withdraw electron through σ covalent bond are said to have -I effect.
  - Eg: $-NO_2$, - OH, -F, -Br.

2. Explain Electromeric effect with an example. (TB)

- Electromeric effect is defined as temporary polarity produced in unsaturated (multiple bonded) compounds in the presence of attacking reagents.
- This is a Temporary phenomenon.
- This involves Pi bonds.
- Positive Electromeric effect:
  - When the π electron is transferred towards the attacking reagent, it is called + E effect.



- Negative Electromeric effect:
  - When the π electron is transferred away from the attacking reagent, it is called, -E effect.



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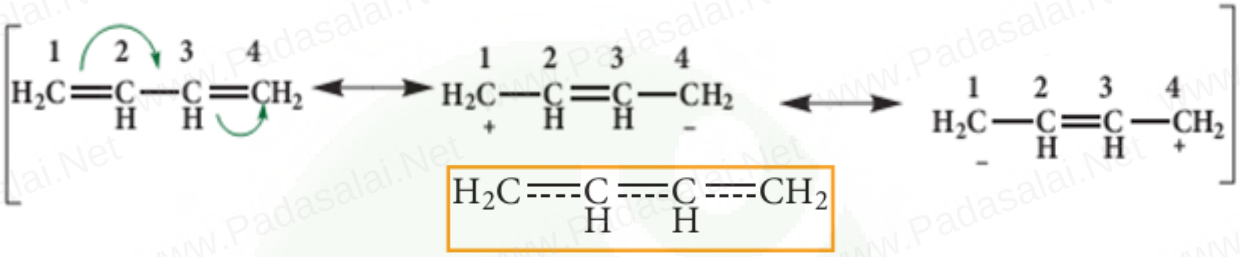
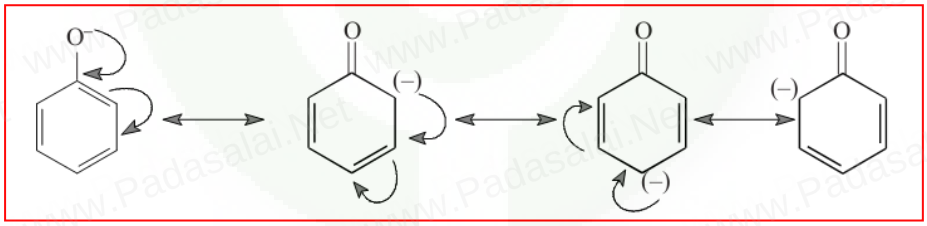
3. Discuss: Resonance Effect with an Example. (TB)

- Certain organic compounds can be represented by more than one structure and they differ only in the position of bonding and lone pair of electrons. These structures are called the resonance structures and this is called Resonance (R) or Mesomeric effect (M).
- Positive Resonance effect:
  - * It occurs when the electrons move away from substituent attached to the conjugated system.
  - * Groups: $-CH_3$, $-OCH_3$, -OH, -SH, $-NH_2$.
  - * Eg: 1, 3-butadiene.

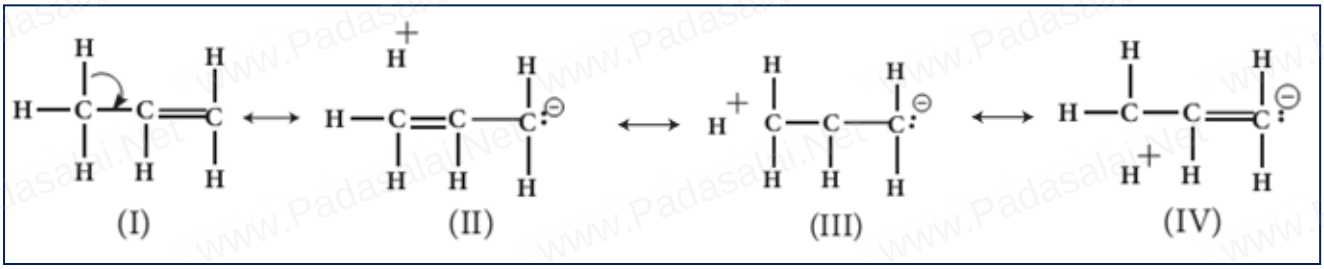


- Negative Resonance effect:
  - * It occurs when the electrons move towards the substituent attached to the conjugated system.
  - * Groups: $-NO_2$, -COOH, -CHO, >C=O-, -CN.
  - * Eg: Phenoxide ion.



4. Explain Hyperconjucative effect in detail. (TB)

- The delocalization of σ electrons is called Hyperconjucation.
- This is a Permanent phenomenon.
- Conditions:
  - * It requires α- H or atom with lone pair (like N, O) is adjacent to sp^2 hybridized carbon (π bond).
  - * Eg: All Carbon in Propene has same bondlength.



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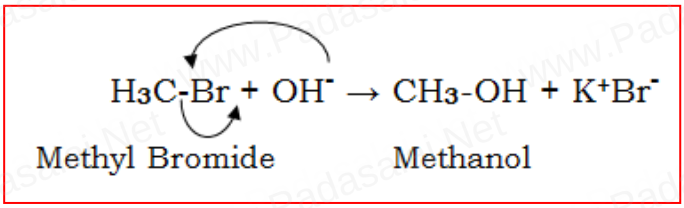
5. Show that Heterolysis of a covalent bond and complete the following equations. Identify the nucleophile. (TB)

a) CH_3-Br + KOH →

CH_3-Br + K^+OH^- → CH_3-OH + K^+Br^-

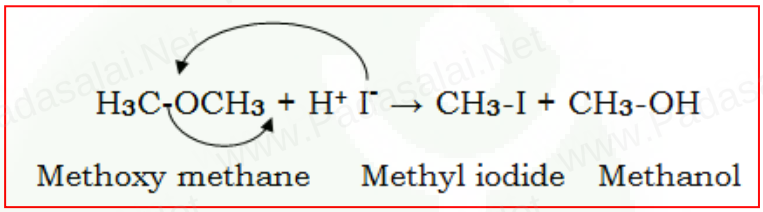
- Nucleophile: OH^-



b) CH_3-O-CH_3 + HI →

CH_3-O-CH_3 + HI → CH_3I + HO-CH_3

- Nucleophile: I^-

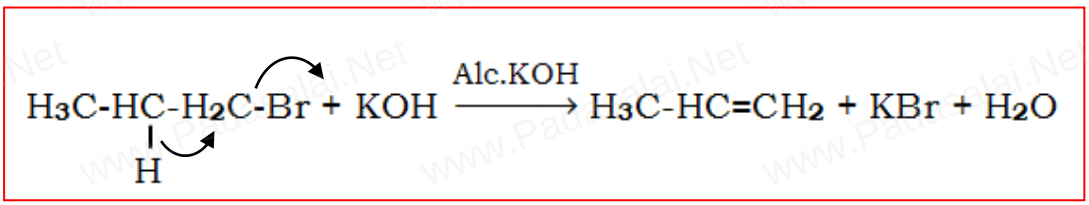


6. What are Electrophiles and Nucleophiles? Explain. (TB)

	Electrophiles (Electron + Loving)		**Nucleophiles** (Nucleus + Loving)
1	Electrophiles are attracted towards the negative charge or the electron rich centre.	1	Nucleophiles are reagent that has high affinity for electro positive centers.
2	They are either positively charged ions or the electron deficient neutral molecules.	2	They are either negatively charged ions or electron rich neutral molecules.
3	They have empty orbitals to accept electrons.	3	They have filled orbitals to donate the electrons.
4	They are Lewis Acids.	4	They are Lewis bases.
5	Positive: NO_2^+, Cl^+, H^+, O^+ Neutral: BF_3, $AlCl_3$, CO_2	5	Negative: Cl^-, CN^-, OH^-, RO^-. Neutral: ROR, H_2O, NH_3.

7. Give example for the following reactions: (TB)

a) β Elimination.



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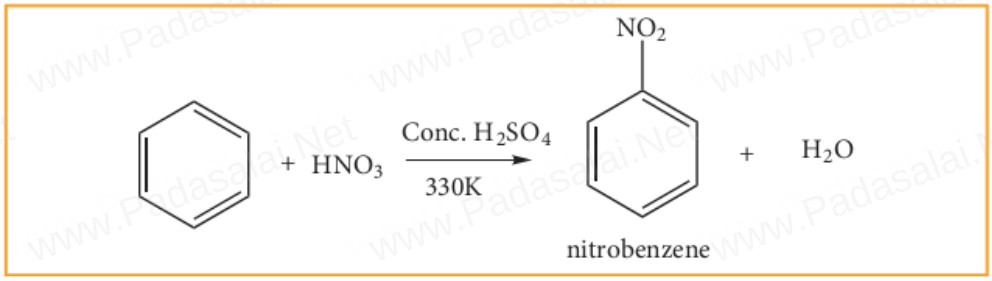
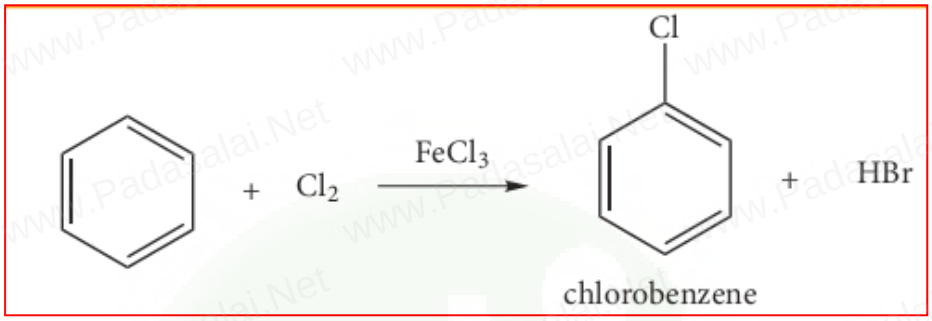
b) Electrophilic substitution.

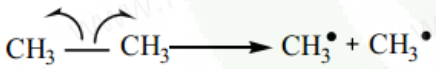
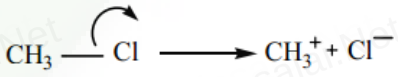
i) Nitration of benzene.



ii) Chlorination of benzene.



8. Differentiate between Homolysis and Heterolysis.

	Heterolytic Fission		Homolytic Fission
1	The process where a covalent bond breaks symmetrically in such way that each of bonded atoms retains one electron. 	1	The process where a covalent bond breaks unsymmetrically in such way that one of the bonded atoms retains the bond pair of electrons. 
2	It requires special conditions: • High Temperature/Light • Peroxide medium.	2	It occurs in normal condition.
3	Electronegative difference is zero or negligible.	3	Electronegative difference is Large.
4	It yields Free radicals.	4	It yields ions i.e. Cation and Anion.
5	Single headed Fish hook arrow is used.	5	Double headed curved arrow is used.

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