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Model questions on Atomic Structure

Carrying 1 mark each

1. What is the ratio of e/m for a cathode ray?
2. The ionization potential of hydrogen atom is 13.6 eV. What is the energy required to remove an electron in the n = 2 state of hydrogen atom?
3. What is the ratio of the radii of the first three Bohr orbit?
4. Which of the following sets of quantum numbers is possible for an electron in a 4f orbital?

   (a) $n = 4, l = 2, m = 2, s = +\frac{1}{2}$ (b) $n = 4, l = 4, m = +4, s = +\frac{1}{2}$

   (c) $n = 4, l = 3, m = +1, s = -\frac{1}{2}$ (d) $n = 4, l = 3, m = 4, s = +\frac{1}{2}$
5. The number of orbitals in n = 3 are _____
6. What is the orbital angular momentum of an electron in 2s orbital?
7. Which of the following relates to photons both as wave motion and as a stream of particles?

   (a) Interference (b) $E = mc^2$

   (c) Diffraction (d) E = h□
8. The number of nodal planes in a p_x orbital is _____
9. For a hydrogen atom, the energies that an electron can have are given by the expression, $E = -13.58/n^2$ eV, where n is an integer. The smallest amount of energy that a hydrogen atom in the ground state can absorb is _____

10. The radial probability distribution curve obtained for an orbital wave function (ψ) has 3 peaks and 2 radial nodes. The valence electron of which one of the following metals does this wave function (ψ) correspond to?

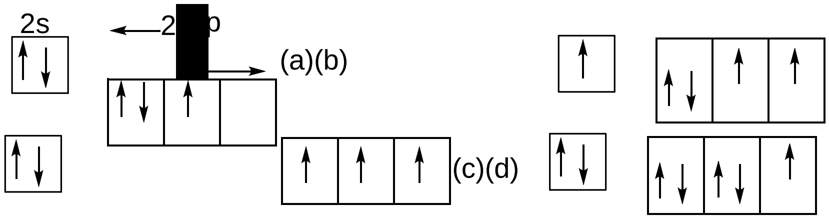
(a) Co (b) Li (c) K (d) Na

11. A photon with initial frequency 10^{11} Hz scatters off an electron at rest. Its final frequency is 0.9×10^{11} Hz. The speed of scattered electron is close to
($h = 6.63 \times 10^{-34}$ Js, $m_e = 9.1 \times 10^{-31}$ kg)

(a) 4×10^3 ms^{-1} (b) 3×10^2 ms^{-1} (c) 2×10^6 ms^{-1} (d) 30 ms^{-1}

Carrying 1.5 marks each

1. What is the ratio of 2nd, 4th and 6th orbits of hydrogen atom?
2. The frequency of a green light is 6×10^{14} Hz. Calculate its wavelength?
3. The values of charge on the oil droplets experimentally observed were -1.6×10^{-19}, -2.4×10^{-19} and -4×10^{-19} coulomb. What is the value of the electronic charge, indicated by these results?
4. What is the ratio of the difference in energy between the first and second Bohr orbits to that between the second and third Bohr orbit?
5. The electrons, identified by quantum numbers n and l, (i) n = 4, l = 1 (ii) n = 4, l = 0, (iii) n = 3, l = 2 and (iv) n = 3, l = 1; place in order of increasing energy, from the lowest to highest.
6. What is the order of wavelength associated with a golf ball weighing 200g and moving at a speed of 5 m/h?
7. A gas absorbs a photon of 355 nm and emits two wavelengths. If one of the emission is at 680 nm, the other is at?
8. In which of the following orbital diagram 'Aufbau principle' is violated? Explain.



9. Binding energy per nucleon of three nuclei A, B and C are 5.5, 8.5 and 7.5 respectively. Which nucleus is most stable? Explain.
10. The mass of ${}_3^7Li$ is 0.042 less than the mass of 3 protons and 4 neutrons. What is the binding energy per nucleon in ${}_3^7Li$?

11. The highest energy in Balmer series, in the emission spectra of hydrogen is represented by ____ cm^{-1}

Carrying 2 marks each

1. Determine the de Broglie wavelength of
   a. an electron moving at 1/10 the speed of light.
   b. a 400 g Frisbee moving at 10 km/h.
   c. an 8.0-pound bowling ball rolling down the lane with a velocity of 2.0 meters per second.
   d. a 13.7 g hummingbird flying at a speed of 30.0 miles per hour.

2. The transition from the $n = 7$ to the $n = 2$ level of the hydrogen atom is accompanied by the emission of radiation slightly beyond the range of human perception, in the ultraviolet region. Determine the energy and wavelength.

3. Emissions are observed at wavelengths of 383.65 and 379.90 nm for transitions from excited states of the hydrogen atom to the $n = 2$ state. Determine the quantum numbers n_h for these emissions.

4. State the following principles
   a. Pauli exclusion principle
   b. Hund's rule of maximum multiplicity

5. For the $3p_z$ and $4d_{xz}$ hydrogen-like atomic orbitals, sketch the following:
   a. The radial function R
   b. The radial probability function $a_0 r^2 R^2$

6. What is the quantum number l for g orbitals? How many orbitals are in a g subshell?
7. Define and state the conditions of normalized and orthogonal wave functions.
8. Write down the limitations of Aufbau principle.
9. Write down the limitations of Bohr's theory.
10. Write down the limitations of Sommerfeld's extension to Bohr's theory.
11. State Heisenberg's uncertainty principle and its significances.
12. State significances of ψ and ψ^2.

13. What is Hamiltonian operator (H)? Express Schrödinger wave equation in terms of H.
14. State significances of four quantum numbers.
15. Describe the angular nodal surfaces for a d_z2 orbital, whose angular wave function is

$$Y = \frac{1}{4}\sqrt{\frac{5}{\pi}}\,\frac{(2z^2 - x^2 - y^2)}{r^2}$$

16. Write down the equation for the solutions of Schrödinger wave equation in terms of polar coordinates.
17. Does an orbital exist if there's no electron in it?

Carrying 6 marks each

1. Briefly describe various postulates and limitations of Bohr's theory.

2. Derive Schrödinger equation for the wave mechanical model of an atom. Discuss significances of ψ and ψ^2. Express Schrödinger wave equation in terms of Hamiltonian operator (H).

3. Write down Schrödinger equation for the wave mechanical model of an atom. State the names and significances of quantum numbers yielded by the solution of this equation.

4. a) What do understand by dual character of matter? Derive de Broglie equation. How was it verified experimentally?
   b) Differentiate matter waves and electromagnetic waves. Why wave nature of matter is not apparent in our daily observations.
   c) Explain that concept of probability follows Heisenberg's uncertainty principle.

The orbital $4f_{z(x^2-y^2)}$ has the angular function 5.

$$Y = (\text{constant})\ z(x^2 - y^2)/r^3$$

a. How many radial nodes does this orbital have?

b. How many angular nodes does it have?

c. Write equations to define the angular nodal surfaces. What shapes are these surfaces?

d. Sketch the shape of the orbital, and show all radial and angular nodes.

6. Derive expressions for complete wave functions of s/p/d-orbital. Explain how these expressions account for various shapes of s/p/d-orbital.

7. Clearly distinguish the following aspects of the structure of an atom and sketch the appropriate function for 1s,2s,2p, 3s and 3p orbitals. a. Radial wave function b) Radial probability function
   c) Angular wave function
   d) Angular probability function

Model questions on **Periodic Properties of elements**

Carrying 1 mark each

1. An element belongs to period 2 and group 2, then number of valence electrons in the atoms of this element is?

2. If the aufbau principle had not been followed, Ca (Z = 20) would have been placed in which block?

3. What is the atomic number of the element with the maximum number of unpaired 4p electron?

4. If period number and group number of any representative element(s) are same then which of the following statement is incorrect regarding such type element(s) in their ground state:
   (Period number and group number are according to modern form of periodic table)
   A. The possible value of principal quantum number is 2
   B. The possible value of azimuthal quantum number is zero
   C. The possible value of magnetic quantum number is 1
   D. The species could be paramagnetic

5. How does the energy gap between successive energy levels in an atom vary from low to high values?

6. Write down the increasing order of acidity of the oxides of Mn (MnO, MnO_2, Mn_2O_7).

7. The correct order of ionic size of N3–, Na+, F–, Mg2+ and O2– is _____

8. Incorrect order of ionic size is:
   A. $La_{3+} > Gd_{3+} > Eu_{3+} > Lu_{3+}$
   B. $V_{2+} > V_{3+} > V_{4+} > V_{5+}$ C. $Tl^{+} > In^{+} > Sn^{2+} > Sb^{3+}$
   D. $K_{+} > Sc_{3+} > V_{5+} > Mn_{7+}$

9. Among NaF , NaCl , NaBr and NaI , the NaF has highest melting point because _____

10. The elements with the lowest atomic number that has a ground state electronic configuration of $(n - 1)d^6 ns^2$ is located in the period _____

11. The period number and group number of “Tantalum” (Z = 73) are respectively ___ and ___

12. [Xe] $4f^{14} 5d^1 6s^2$ belongs to ___ period and ___ group.

Carrying 1.5 marks each

1. Explain the correct order of second I. E. of C, N, O and F.

2. Consider the following changes $A \rightarrow A^+ + e^-$: E1 and $A^+ \rightarrow A^{2+} + e^-$: E2

   The energy required to pull out the two electrons are E and E respectively. The correct relationship between two energies would be E1_____E2.

3. The correct order of increasing electron affinity of O, S, F and Cl is _____

4. If the ionization and electron gain enthalpy of an element are 275 and 86 Kcal.mol^{-1} respectively, then the electro negativity of the element on the Mulliken scale is ____

5. The atomic numbers of the metallic and non-metallic elements which are liquid at room temperature respectively are ___ and ___

6. What is the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar?

7. Sodium generally does not show oxidation state of +2. Why?

8. In the fourth period of the periodic table, how many elements have one or more 4d electrons?

9. Assuming that elements are formed to complete the seventh period, what would be the atomic number of the alkaline earth metal of the eighth period?

10. Why Zn and Cd metals do not show variable valency?

11. A compound contains three elements A, B and C, if the oxidation number of A = +2, B = +5 and C = –2, the possible formula of the compound is _____

12. Consider the following information about element P and Q:

Element	Period	Group
P	2	15
Q	3	2

Then formula of the compound formed by P and Q element is _____

13. Which is the most stable cation of element 113? Explain.

14. Using Slater's rules, determine Z * for a 2 p electron in O^{2-}.

15. Why Sn often forms an ion having a charge of 2+?

Carrying 2 marks each

1. What states are possible for a d^3 configuration? Determine the Coulombic and exchange energies for each, and rank the states in terms of relative energy.
2. The electron configuration of Ti is [Ar] $4s^2\ 3d^2$, but that of Cr^{2+} is [Ar] $3d^4$. Explain.
3. Give electron configurations for the following:
   a. V b. Br c. Ru^{3+} d. Hg^{2+} e. Sb
4. The electron configuration of the molybdenum atom is [Kr] $5s^1\ 4d^5$ rather than [Kr] $5s^2\ 4d^4$. Explain.
5. The most common ion formed by zinc has a 2+ charge. Explain.
6. Cm has the outer electron configuration $s^2\ d^1\ f^7$ rather than $s^2\ f^8$. Explain.
7. The second ionization of carbon ($C^+ \rightarrow C^{2+} + e^-$) and the first ionization of boron ($B \rightarrow B^+ + e^+$) both fit the reaction $1s^2\ 2s^2\ 2p^1 \rightarrow 1s^2\ 2s^2$ + e-. Compare the two ionization energies (24.383 eV and 8.298 eV, respectively) and the effective nuclear charge Z*. Is this an adequate explanation of the difference in ionization energies? If not, suggest other factors.
8. Ionization energies should depend on the effective nuclear charge that holds the electrons in the atom. Calculate Z* (Slater's rules) for N, P, and As. Do their ionization energies seem to match these effective nuclear charges? If not, what other factors influence the ionization energies?
9. In each of the following pairs, pick the element with the higher ionization energy and explain your choice. a) Fe, Ru b) P, S c) K, Br d) C, N e) Cd, In f) Cl, F
10. On the basis of electron configurations, explain why
    a. sulfur has a lower electron affinity than chlorine.
    b. boron has a lower ionization energy than beryllium.
11. Calculate the effective nuclear charge on a 5s, 5p, and 4d electron in a tin atom.
12. Which of the halogens, X_2, would you expect to be most likely to form a cation, X^+? Discrete X^+ ions are not known in chemical compounds but X_2^+, X_3^+ and X_5^+ are known. Why should the latter be more stable than X^+?
13. Discuss the formula for calculating Sanderson's electron density ratio with suitable examples.

Carrying 6 marks each

1. What is screening effect? Discuss the trend of screening effect along the periods and groups. Describe Slater's rules. Discuss its role in determining the effective nuclear charge for any electron in an atom and its limitations.

2. Using Slater's rules, determine Z* for
   a) a 3p electron in P, S, Cl, and Ar. Is the calculated value of Z* consistent with the relative sizes of these atoms?

   b) a 2p electron in O^{2-}, F^-, Na^+ and Mg^{2+}. Is the calculated value of Z* consistent with the relative sizes of these ions?

   c) a 4s and a 3d electron of Cu. Which type of electron is more likely to be lost when copper forms a positive ion?

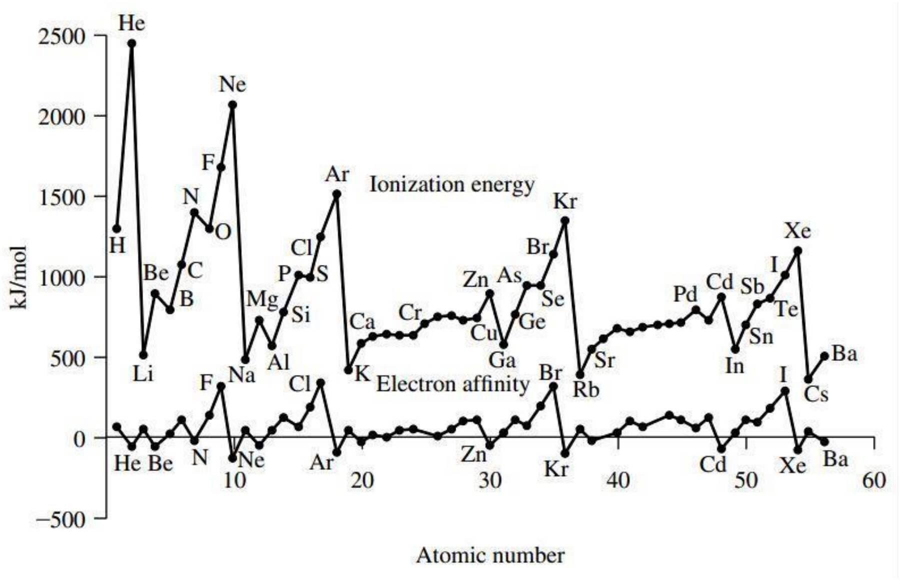
   d) a 4f electron in Ce, Pr, and Nd. There is a decrease in size, commonly known as the lanthanide contraction, with increasing atomic number in the lanthanides. Are your values of Z* consistent with this trend?

3. Select the best choice, and briefly indicate the reason for each choice:
   a) Largest radius: Na^+ Ne F^-
   b) Greatest volume: S^{2-} Se^{2-} Te^{2-}
   c) Highest ionization energy: Na Mg Al
   d) Most energy necessary
   to remove an electron: Fe Fe^{2+} Fe^{3+} e.
   Highest electron affinity: O F Ne
   f. Smallest radius: Y^{3+} Zr^{4+} Nb^{5+}

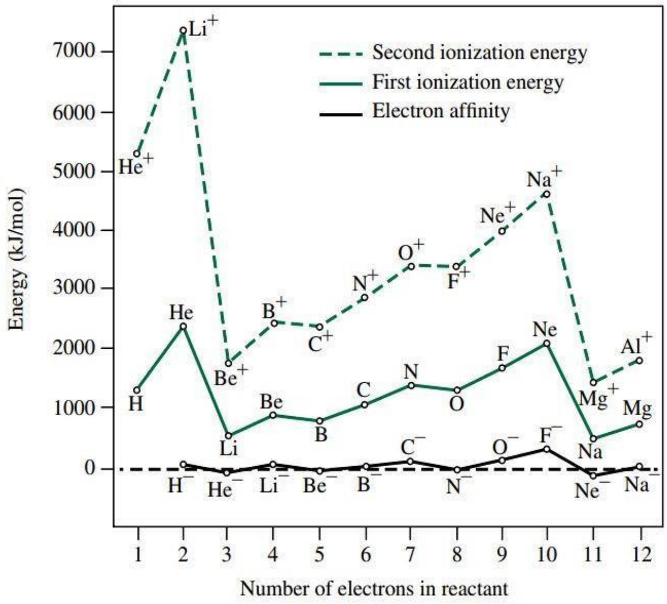
4. **a)** The graph of ionization energy versus atomic number for the elements Na through Ar (Figure) shows maxima at Mg and P and minima at Al and S. Explain these maxima and minima.
   b) The graph of electron affinity versus atomic number for the elements Na through Ar (Figure) also shows maxima and minima, but shifted by one element in comparison with the ionization energy graph. Why are the maxima and minima shifted in this way?

   c) For a graph of third ionization energy against atomic number, predict the positions of peaks and valleys for elements through atomic number 12. Compare the positions of these peaks and valleys with those for first ionization energies shown in Figure.

   d) How would a graph of second ionization energy vs. atomic number for the elements helium through neon compare with the graph of first ionization energy in the Figure? Be specific in comparing the positions of peaks and valleys.



5. **a)** Explain why all three graphs in the Figure have maxima at 4 electrons and minima at 5 electrons.
b) How would a graph of third ionization energies against the number of electrons in reactant compare with the other graphs shown in the Figure? Explain briefly.



6. The size of the transition-metal atoms decreases slightly from left to right in the periodic table. What factors must be considered in explaining this decrease? In particular, why does the size decrease at all, and why is the decrease so gradual? Also discuss trend of electron gain enthalpy along the periods and groups.

7. What is Ionization enthalpy? Discuss successive ionization enthalpies and its trend with suitable examples and factors affecting ionization energy. Elaborate applications of ionization enthalpy.

8. What is electronegativity? Write down the formula for calculation using Pauling's and Mulliken's electronegativity scales. Discuss the variation of electronegativity with bond order, partial charge and hybridization.

Model questions on **Chemical Bonding**

Carrying 1 mark each

1. The Geometry of $TeCl_4$ is _____

2. The bond angle of Cl_2O is

   (a) Smaller than that of F_2O (b) Smaller than that of H_2O

   (c) Greater than that of H_2O (d) Same as that of F_2O

3. The bond order in O_2^+ is_____

4. According to VSEPR model, the shape of $[XeOF_5]^-$ is_____
5. The C–H bond length is greatest in

   (a) CH_4 (b) CH_3F (c) CH_2F_2 (d) CHF_3

6. Among the following, the number of molecules which do not exist are______ IF_7, SH_6, XeH_4, PH_5, ClF_3, ClH_3, BrF_5

7. The number of molecular orbital nodel plane (s) in BMO formed from p-p sideways overlapping is (are) ______

8. The F–B–F bond angle in BF_3 is _______ degree.

9. The number of molecules with bond angle near to 90º are__________ CH_4, SiF_4, BrF_5, XeF_4, PH_3, H_2S

10. The number of molecules having bond angle greater than H_2O are______ CH_4, NH_3, H_2S, PH_3, Cl_2O, F_2O

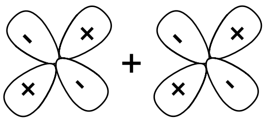
11. The percentage s-character in lone pair of NH_3 is _______%

12. If p_z and d_{zx} orbital overlapped, then the bond formed will be _____ bond, if the internuclear axis is ‘x’ axis.

13. Which is incorrect representation of mixing:

    (a) □□□ □□□

    (c) 

    (d) 

    (b)

14. The hybridization state in triangle bipyramidal CH_5^+ is _____

15. Which is the strongest bond among:

    (a) $2p$□-$2p$□ (b) $3p$ -$3p$ □ (c) $3d$ -$3d$ □ (d) $3d$ -$2p$□ □

16. Hybridization on cubic $[PaF_8]$□□□ is ______

17. Give an example where d□ □p bond is possible.

18. In octahedral structure the pair of ‘d’ orbitals involved is ____ and ____

19. The ONO angle is maximum in (NO_3^- /NO_2^- / NO_2 / NO_2^+)

20. Why P_2 coverts into P_4?

Carrying 1.5 marks each

1. Write down the correct bond angle order. CH_4, NH_3, PH_3, H_2O, H_2S

2. What is the correct order of p-p bond length in P_2F_4, $P_2(CH_3)_4$, $P_2(CF_3)_4$ and P_2H_4?

3. In a rectangular octahedral molecule MX_6, the number of X-M-X bonds at 180^0C is_____.

4. Justify the correct order of bond angle

   (a) $CCl_4 > BF_3 > NO_2^+$ (b) $NH_3 > NCl_3 > NBr_3$ (c) $Br_2O > Cl_2O > OF_2$ (d) $PCl_3 > PBr_3 > PI_3$

5. What is the ratio of sigma and pi-bonds in naphthalene?

6. If CH_4 is assumed to give two disubstituted products, then what is the shape of CH_4?

7. State Fajan’s rules.

8. Which molecule has the smallest bond angle in each series?
   (a) $OSeF_2$ $OSeCl_2$ $OSeBr_2$ (halogen–Se–halogen angle)
   (b) $SbCl_3$ $SbBr_3$ SbI_3
   (c) PI_3 AsI_3 SbI_3

9. Give Lewis dot structures and sketch the shapes of the following:

   (a) $SeCl_4$ b. I_3^- c. $PSCl_3$ (P is central)

10. Compare the structures of the azide ion, N_3^-, and the ozone molecule, O_3.
11. How would you expect the structure of the ozonide ion, O_3^-, to differ from that of ozone?

12. PCl_5 is a stable molecule, but NCl_5 is not. Explain.
13. Why SF_4 and SF_6 are known, but OF_4 and OF_6 are not?

14. Are the CF_3 groups in $PCl_3(CF_3)_2$ more likely axial or equatorial? Explain briefly.

15. Are the axial or equatorial bonds likely to be longer in $SbCl_5$? Explain briefly.

16. Select the molecule having the smallest bond angle and briefly explain: NH_3, PH_3, or AsH_3
17. $[S_2CN(CH_3)_2]^-$ Draw the resonating structures of this ion.
18. Draw the resonating structures of dimethylthiocarbamate ion.

19. $NSCl_2^-$ and $OSCl_2$: Which of these species has the smaller Cl---S---Cl angle? Explain briefly.

20. Methanol, CH_3OH, has a much higher boiling point than methyl mercaptan, CH_3SH. Explain.

Carrying 2 marks each

1. In PF_4^+ the F---F and P—F distances are 238 pm and 145.7 pm, respectively. Predict the P—F distance in POF_3, which has an F—P—F angle of 101.1°.

Briefly account for the following observations:

2. The bond angle in NCl_3 is nearly 5 degrees larger than in NF_3.
3. The S–F axial distance in SOF_3 is longer than the S–F equatorial distance.
4. In $Te(CH_3)_2I_2$ the methyl groups are in equatorial, rather than axial, positions.
5. The O—S—O bond angle in $FSO_2(OCH_3)$ is larger than in $FSO_2(CH_3)$.
6. Carbon monoxide has slightly higher melting and boiling points than N_2.
7. The ortho isomer of hydroxybenzoic acid [$C_6H_4(OH)(CO_2H)$] has a much lower melting point than the meta and para isomers.
8. The boiling points of the noble gases increase with atomic number.
9. Acetic acid in the gas phase has a significantly lower pressure (approaching a limit of one half) than predicted by the ideal gas law.
10. Mixtures of acetone and chloroform exhibit significant negative deviations from Raoult's law, which states that the vapor pressure of a volatile liquid is proportional to its mole fraction. For example, an equimolar mixture of acetone and chloroform has a lower vapor pressure than either of the pure liquids.
11. Carbon monoxide has a greater bond-dissociation energy (1072 kJ/mol) than molecular nitrogen (945 kJ/mol).

Discuss the following.

12. General characteristics of an ionic bond
13. Limitations of radius ratio rule
14. Madelung constant
15. Applications of Born-Haber cycle
16. Solvation energy
17. Postulates of VBT
18. Heitler-London approach to VBT
19. Limitations of VBT
20. Resonance energy
21. M.O. Diagram of N2 / O2 / C2 / B2 / F2 / CO / NO / CO^+ / NO^+ / NO^-
22. Covalent character in ionic compounds
23. Postulates of VSEPR theory
24. Limitations of VSEPR theory
25. Polarizing power and polarizability
26. Ionic character in covalent compounds
27. Percentage ionic character from dipole moment and EN difference
28. Metallic Bond
29. Van der Waals forces
30. Ion-dipole forces
31. Dipole-dipole interactions
32. Induced dipole interactions
33. Instantaneous dipole-induced dipole interactions
34. Hydrogen bonding

35. Which has the smaller F—P—F angle, PF_4^+ or PF_3O? Which has the longer fluorine–fluorine distance? Explain briefly.

Carrying 6 marks each

1. (*a*) What is an ionic bond ?

   (*b*) Discuss carefully the formation of an ionic bond in terms of decrease of energy. Derive from it the conditions for the formation of ionic bonds.

   (*c*) Explain giving suitable examples why ions with three positive or three negative charges are rare ?

2. (*a*) Discuss the formation of a covalent bond on the basis of the orbital overlap concept. What is meant by a sigma bond ?

   (*b*) Differentiate between a sigma bond and a pi bond. Explain why a sigma bond is stronger than a pi bond ?

   (*c*) Explain the term covalency. How would you account for variable covalency of an element ? What is the maximum covalency which an element can exhibit ?

   (*d*) Construct a plot of potential energy versus internuclear distance between two hydrogen atoms.

3. (*a*) What do you understand by polar and non-polar molecules ? Is it possible for the non-polar molecules to have polar bonds ? Justify you answer with examples.

   (*b*) Explain why CO_2 and CCl_4 molecules are non-polar while $CHCl_3$ molecule is polar ?

   (*c*) What is meant by dipole moment ? Can its magnitude give an idea about the structure of a molecule ? Explain giving examples.

   (*d*) Which is more polar, H_2S or H_2O and why ?

4. (*a*) What is the difference between coordinate bond and covalent bond ?

   (*b*) Under what circumstances can a fully filled orbital participate in bond formation? Discuss with an example.

   (*c*) PCl_5 exists while NCl_5 does not. Explain.

5. What is Valence Bond theory ? Explain.

   In the potential energy diagram for H_2 molecule, the combined wave function ψ for two hydrogen atoms $\psi = \psi_A(1)\psi_A(2)$ (where $\psi_A(1)$ and $\psi_B(2)$ represent wave functions for two hydrogen atoms H_A and H_B associated with electrons 1 and 2, respectively), results in decrease in energy of the system by 24 kJ mol^{-1}. But the experimental value is 458 kJ mol^{-1}. What improvements in the wave functions of the system are made to account for this difference?

6. (*a*) What is resonance ? Discuss by taking example of CO_3^{2-} ion.

   (*b*) What do you understand by the term resonance energy ? What are the essential rules for writing resonating structures?

7. (*a*) What is Molecular Orbital theory ?

   (*b*) Give the electron charge density diagrams for bonding and antibonding molecular orbitals.

   (*c*) Predict the bond order in C_2 molecule with the help of MO diagram.

8. (*a*) What do you understand by the terms bonding and antibonding molecular orbitals ? Why are they so called ? Illustrate your answer by taking into consideration the structure of hydrogen ion H_2^+.

   (*b*) Is B_2 molecule paramagnetic or diamagnetic ? Discuss.

9. Explain how the atomic orbitals combine to form bonding and antibonding molecular orbitals. What are the limitations to such combinations ? Explain why in a homonuclear molecule, an *s* orbital of one atom cannot combine with a p_x or a p_y orbital of another atom of the same element.

10. (*a*) Draw molecular orbital diagram for HF molecule.

11. Explain on the basis of MO theory as to why:

10. Derive Born-Lande equation. Define Madelung constant and upon which factors does it depen

   (*a*) Oxygen molecule is paramagnetic while nitrogen molecule is diamagnestic.

   (*b*) Hydrogen forms diatomic molecule while helium remains monoatomic.

   (*c*) The bond order in O_2^- is less than that in O_2 which, in turn, is less than that in O_2^+.

12. (*a*) ~~The internuclear distances in some O – O~~ bonds are as follows :

O_2^+	O_2	O_2^-	O_2^{2-}
1·12 Å	1·21 Å	1·30 Å	1·49 Å

   (*i*) Account for the gradation of the internuclear distances in the above series.

(*ii*) Which of these species are paramagnetic ?

(*b*) Is it correct to say that bond energy in a diatomic molecule always increases when an electron is lost ?

(*c*) "While in the case of F_2 molecule, all the orbitals are molecular, in case of HF molecule there is only one orbital which is molecular." Comment on this statement.

13. (*a*) Discuss with the help of the MO theory the formaiton of N_2, O_2 and F_2 molecules. How does the theory explain the difference in the reactivity of N_2, O_2 and F_2 ?

(*b*) Discuss carefully the electronic configurations of NO, CO, HF and HCl molecules.

14. Draw the MO diagrams for the following species :

Be_2, HHe (hypothetical), O_2^-, CO, BN

Explain the magnetic character of these species.

15. Write the electronic configuration of NO molecule.

(*a*) What is the bond order ?

(*b*) Will the bond length be shorter or larger than in NO^+ ?

(*c*) How many unpaired electrons will be present ?

(*d*) What will be the bond order in NO^+ ?

16. With the help of molecular orbital diagrams explain why

(*a*) the bond order in N_2^+ ion is less than that in N_2 molecule whereas the bond order in O_2^+ is greater than that in O_2 molecule.

(*b*) the bond energy of NO^+ is higher than that of NO.

(*c*) the bond length of CO^+ is larger than that of CO.

17. (*a*) Write the electronic configuration of a diatomic molecule having bond order three.

(*b*) Compare briefly valence bond and molecular orbital theories.

18. What is meant by metallic bond ? Illustrate the nature of metallic bond on the basis of (*i*) Electron sea model. (*ii*) Valence bond model and (*iii*) The band model (MO approach).

19. Discuss in details the band model of metallic bond. What type of energy bands are obtained in, (*i*) sodium metal (*ii*) an insulator (*iii*) semi-conductor

20. What is meant by hydrogen bonding ? Discuss the nature and consequences of hydrogen bonding. Discuss the importance of hydrogen bonding in sustaining life. Why do H_2O and HF have abnormally high boiling points ?

21. (*a*) What are intermolecular forces ? How do these forces originate ?

(*b*) What are London forces ? What is their origin ?

Carrying 6 marks each

1. (*a*) What do you understand by hybridisation ? Explain giving examples.
   (*b*) What type of geometry is possible for molecules undergoing the following types of hybridisation :
   (*i*) sp (*ii*) sp^2 (*iii*) sp^3 (*iv*) sp^3d (*v*) sp^3d^2 (*vi*) sp^3d^3
2. (*a*) Why do molecules have definite geometry ? In which of the following types of geometry, all the bond angles are not the same :
   (*i*) Octahedral (*ii*) trigonal planar (*iii*) trigonal bipyramidal
   (*iv*) tetrahedral (*v*) pentagonal bipyramidal. **[Ans. (*iii*), (*v*)]**
   (*b*) Beryllium (Z = 4) has no unpaired electron in its ground state. However, in all its compounds beryllium shows divalency. Explain.
3. (*a*) What is hybridisation ? What are the important characteristics of hybridisation ?
   (*b*) The central atom in the molecules of CH_4, NH_3 and H_2O involves sp^3 hybridisation of its valence shell orbitals. But the bond angles in these molecules are 109·5, 107° and 104·5°, respectively. Account for this.
   (*c*) What shapes are associated with the molecules involving sp^3d^2 and sp^3d^3 hybridisation ?
4. Discuss the stereochemistry of the following molecules on the basis of hybridisation : PF_5, $SnCl_2$, NH_3, ClF_3
5. The ground state electronic configuration of chlorine is $1s^22s^2p^63s^2p^5$. Account for the geometry of the following molecules/ions of chlorine :
   (*a*) ClF_3 (*b*) ClO_4^- (*c*) ClO_3^-
6. On the basis of hybridisation, discuss the geometry of the following molecules :
   SF_6, IF_7, $SnCl_2$, XeF_4, SF_4, H_2O, PCl_5 and NH_3.
7. (*a*) According to electronic configuration, water should have bond angle H – O – H of 90° and ammonia should have three mutually perpendicular N – H bonds but actually both the molecules have unsymmetrical tetrahedral structure. How would you explain this ?
   (*c*) Discuss the geometry of molecules having
   (*i*) two lone pairs and two bond pairs (*ii*) one lone pair and three bond pairs
   (*iii*) one lone pair and four bond pairs (*iv*) one lone pair and five bond pairs
8. Discuss the geometry of the following ions :
   (*a*) NO_3^- (*b*) CO_3^{2-} (*c*) SO_4^{2-} (*d*) ClO_3^- (*e*) ClO_4^-
9. (*a*) Why is it necessary to involve hybrid orbitals while trying to account for the geometry of methane molecule ?
   (*b*) Which of PF_5 or SF_6 does not have symmetrical geometry and why ?
   (*c*) Account for the geometry of PH_3, OF_2 and IF_5 on the basis of concept of hybridisation.
   (*d*) Discuss sp^3d^3 hybridisation with one example.
10. Construct wave functions for sp, sp^2 and sp^3 hybrid orbitals.
11. Involving the wave mechanical treatment, discuss the orientation of hybrid orbitals in space.
12. Discuss the VSEPR theory. Illustrate carefully that VSEPR theory has to be combined with the concept of hybridisation to account for the geometry of covalent molecules.

Model questions on **Redox Reactions**

Carrying 1 mark each

1. Which of the following is not an example of redox reaction?
   (i) $CuO + H_2 \longrightarrow Cu + H_2O$
   (ii) $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$

(iii) $2K + F_2 \longrightarrow 2KF$

(iv) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$

2. The more positive the value of E⊖, the greater is the tendency of the species to get reduced. Using the standard electrode potential of redox couples given below find out which of the following is the strongest oxidising agent.

$$\textbf{E}^{\ominus}\textbf{ values: } Fe^{3+}/Fe^{2+} = +0.77;\ I_2(s)/I^- = +0.54;$$
$$Cu^{2+}/Cu = +0.34;\ Ag^+/Ag = +0.80V$$

3+

(i) Fe (ii) I_2 (s)

(iii) Cu^{2+}

(iv) Ag^+

3. E⊖ values of some redox couples are given below. On the basis of these values choose the correct option.

$$\textbf{E}^{\ominus}\textbf{ values: } Br_2/Br^- = +1.90;\ Ag^+/Ag(s) = +0.80$$
$$Cu^{2+}/Cu(s) = +0.34;\ I_2(s)/I^- = +0.54$$

–

(i) Cu will reduce Br

(ii) Cu will reduce Ag

(iii) Cu will reduce I^- (iv) Cu will reduce Br_2

4. Using the standard electrode potential, find out the pair between which redox reaction is not feasible.

$$\textbf{E}^{\ominus}\textbf{ values: } Fe^{3+}/Fe^{2+} = +0.77;\ I_2/I^- = +0.54;$$
$$Cu^{2+}/Cu = +0.34;\ Ag^+/Ag = +0.80\ V$$

(i) Fe^{3+} and I^- (ii) Ag^+ and Cu

(iii) Fe^{3+}

(iv) Ag and Fe^{3+}

5. Thiosulphate reacts differently with iodine and bromine in the reactions given below:

$$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$$
$$S_2O_3^{2-} + 2Br_2 + 5H_2O \rightarrow 2SO_4^{2-} + 2Br^- + 10H^+$$

Which of the following statements justifies the above dual behaviour of thiosulphate? (i) Bromine is a stronger oxidant than iodine.

(ii) Bromine is a weaker oxidant than iodine.

(iii) Thiosulphate undergoes oxidation by bromine and reduction by iodine in these reactions.

(iv) Bromine undergoes oxidation and iodine undergoes reduction in these reactions.

6. The oxidation number of an element in a compound is evaluated on the basis of certain rules. Which of the following rules is not correct in this respect?

(i) The oxidation number of hydrogen is always +1.

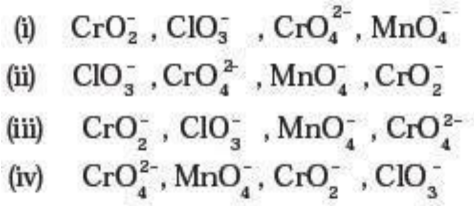
(ii) The algebraic sum of all the oxidation numbers in a compound is zero.
(iii) An element in the free or the uncombined state bears oxidation number zero.

(iv) In all its compounds, the oxidation number of fluorine is – 1.

7. In which of the following compounds, an element exhibits two different oxidation states.

(i) NH_2OH
(ii) NH_4NO_3
(iii) N_2H_4
(iv) N_3H

8. Which of the following arrangements represent increasing oxidation number of the central atom?



(i) CrO_2^-, ClO_3^-, CrO_4^{2-}, MnO_4^-
(ii) ClO_3^-, CrO_4^{2-}, MnO_4^-, CrO_2^-
(iii) CrO_2^-, ClO_3^-, MnO_4^-, CrO_4^{2-}
(iv) CrO_4^{2-}, MnO_4^-, CrO_2^-, ClO_3^-

9. The largest oxidation number exhibited by an element depends on its outer electronic configuration. With which of the following outer electronic configurations the element will exhibit largest oxidation number?

(i) $3d^1 4s^2$
(ii) $3d^3 4s^2$
(iii) $3d^5 4s^1$
(iv) $3d^5 4s^2$

10. Identify disproportionation reaction

(i) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
(ii) $CH_4 + 4Cl_2 \rightarrow CCl_4 + 4HCl$
(iii) $2F_2 + 2OH^- \rightarrow 2F - + OF_2 + H_2O$
(iv) $2NO_2 + 2OH^- \longrightarrow NO_2^- + NO_3^- + H_2O$

11. Which of the following elements does not show disproportionation tendency?

(i) Cl (ii) Br (iii) F (iv) I

Carrying 1.5 or 2 marks each

1. The reaction
$Cl_2 (g) + 2OH^- (aq) \longrightarrow ClO^- (aq) + Cl^- (aq) + H_2O(l)$ represents the process of bleaching. Identify and name the species that bleaches the substances due to its oxidising action.

2. MnO_4^{2-} undergoes disproportionation reaction in acidic medium but MnO_4^- does not. Give reason.

3. PbO and PbO_2 react with HCl according to following chemical equations:
   $2PbO + 4HCl \longrightarrow 2PbCl_2 + 2H_2O$
   $PbO_2 + 4HCl \longrightarrow PbCl_2 + Cl_2 + 2H_2O$
   Why do these compounds differ in their reactivity?

4. Nitric acid is an oxidising agent and reacts with PbO but it does not react with PbO_2. Explain why?

5. Write balanced chemical equation for the following reactions:

   (i) Permanganate ion (MnO_4^-) reacts with sulphur dioxide gas in acidic medium to produce Mn^{2+} and hydrogensulphate ion. (Balance by ion electron method)

   (ii) Reaction of liquid hydrazine (N_2H_4) with chlorate ion (ClO_3^-) in basic medium produces nitric oxide gas and chloride ion in gaseous state. (Balance by oxidation number method)

   (iii) Dichlorine heptaoxide (Cl_2O_7) in gaseous state combines with an aqueous solution of hydrogen peroxide in acidic medium to give chlorite ion (ClO_2^-) and oxygen gas.

   (Balance by ion electron method)

6. Calculate the oxidation number of phosphorus in the following species.

   (a) HPO_3^{2-} and (b) PO_4^{3-}

7. Calculate the oxidation number of each sulphur atom in the following compounds:

   (a) $Na_2S_2O_3$ (b) $Na_2S_4O_6$ (c) Na_2SO_3 (d) Na_2SO_4

8. Balance the following equations by the oxidation number method.
   (i) $Fe^{2+} + H^+ + Cr_2O_7^{2-} \longrightarrow Cr^{3+} + Fe^{3+} + H_2O$
   (ii) $I_2 + NO_3^- \longrightarrow NO_2 + IO_3^-$
   (iii) $I_2 + S_2O_3^{2-} \longrightarrow I^- + S_4O_6^{2-}$
   (iv) $MnO_2 + C_2O_4^{2-} \longrightarrow Mn^{2+} + CO_2$

9. Identify the redox reactions out of the following reactions and identify the oxidising and reducing agents in them.
   (i) $3HCl(aq) + HNO_3(aq) \longrightarrow Cl_2(g) + NOCl(g) + 2H_2O(l)$
   (ii) $HgCl_2(aq) + 2KI(aq) \longrightarrow HgI_2(s) + 2KCl(aq)$
   (iii) $Fe_2O_3(s) + 3CO(g) \xrightarrow{\Delta} 2Fe(s) + 3CO_2(g)$
   (iv) $PCl_3(l) + 3H_2O(l) \longrightarrow 3HCl(aq) + H_3PO_3(aq)$
   (v) $4NH_3 + 3O_2(g) \longrightarrow 2N_2(g) + 6H_2O(g)$

10. Balance the following ionic equations

(i) $Cr_2O_7^{2-} + H^+ + I^- \longrightarrow Cr^{3+} + I_2 + H_2O$

(ii) $Cr_2O_7^{2-} + Fe^{2+} + H^+ \longrightarrow Cr^{3+} + Fe^{3+} + H_2O$

(iii) $MnO_4^- + SO_3^{2-} + H^+ \longrightarrow Mn^{2+} + SO_4^{2-} + H_2O$

(iv) $MnO_4^- + H^+ + Br^- \longrightarrow Mn^{2+} + Br_2 + H_2O$

Carrying 6 marks each

1. Explain redox reactions on the basis of electron transfer. Give suitable examples.
2. On the basis of standard electrode potential values, suggest which of the following reactions would take place? (Consult the book for $E^\ominus$ value).
   (i) $Cu + Zn^{2+} \longrightarrow Cu^{2+} + Zn$
   (ii) $Mg + Fe^{2+} \longrightarrow Mg^{2+} + Fe$
   (iii) $Br_2 + 2Cl^- \longrightarrow Cl_2 + 2Br^-$
   (iv) $Fe + Cd^{2+} \longrightarrow Cd + Fe^{2+}$
3. Why does fluorine not show disporportionation reaction?
4. Write redox couples involved in the reactions (i) to (iv) given in question 34.
5. Find out the oxidation number of chlorine in the following compounds and arrange them in increasing order of oxidation number of chlorine.
   $NaClO_4$, $NaClO_3$, $NaClO$, $KClO_2$, Cl_2O_7, ClO_3, Cl_2O, $NaCl$, Cl_2, ClO_2.
   Which oxidation state is not present in any of the above compounds?
6. Which method can be used to find out strength of reductant/oxidant in a solution? Explain with an example.