XI CHEMISTRY

INDEX

CLASS - XI

Unit 1 Some basic Concepts of Chemistry

Q.No.1-12 : 1 Mark, 13-19 = 2 Marks, and 20-22 = 3 Marks Total = 35 Marks

1. Define law of multiple proportions with example.
2. Calculate the molecular mass of $C_{12}H_{22}O_{11}$
3. Calculate the no. of atoms present in 11.5 litres of H_2 at N.T.P.
4. Calculate the no. of moles of 5.68 gm. of iron.
5. ---------- is the no. of atoms present 2.00 gm. of calcium.
6. Emperical Formula X ----------- = Molecular Formula.
7. The total number of protons in 10 g of calcium carbonate is :

   (1) 3.0115×10^{24} (2) 1.5057×10^{24}

   (3) 2.0478×10^{24} (4) 4.0956×10^{24}
8. The molecular mass of CO_2 is 44 amu and Avogadro's number is 6.02×10^{23}. Therefore, the mass of one molecule of CO_2 is :

   (1) 7.31×10^{-23} (2) 3.65×10^{-23}

   (3) 1.01×10^{-23} (4) 2.01×10^{-23}
9. What is the effect of temperature on molality and molarity?
10. One mole of Water occupies 22.4 L at STP. T/F
11. An atom of an element is 10.1 times heavier than the mass of a carbon atom. What is its mass in a.m.u.?
12. Explain with example, limiting reagent. 12x1=12
13. Differentiate between molarity and molality.
14. 1.82 g. of glucose (molar mass-180) is dissolved in 25g of water. Calculate (a) the molality (b) mole fraction of glucose and water.
15. The molecular mass of an organic compound is 90 and its %age composition is C-26.6%; O=71.1% and H=2.2%. Determine the molecular formula of the compound.
16. How chemical equations are made more informative?
17. How Avogadro's hypothesis used to deduce atomicity of elementary gases?
18. Verify law of Reciprocal proportions or law of equivalent proportions, with example.
19. Define formula mass and how does it differs from molecular mass? 7x2=14
20. Discuss Dalton's Atomic theory and its limitations?
21. Discuss Modern Atomic theory. Why it is better than Dalton's Atomic theory?
22. Commercially available sulphuric acid contains 91% acid by mass and has a density of 1.83g mL^{-1} (i) Calculate the molarity of the solution (ii) volume of concentrated acid required to prepare 3.5L of 0.50 M H_2SO_4

Some More Questions :

23. A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% of chlorine. Its molar mass is 98.96g. What are its empirical and molecular formulas?
24. How much copper can be obtained from 110gm of $CuSO_4$?
25. What is Gay Lussac's law? Explain with two examples.
26. What are empirical and molecular formulae? How are they related to each other?
27. Differentiate between normality and molarity?
28. Why molality is preferred over molarity in expressing the concentration of a solution?
29. Explain with the help of an example law of conservation of mass and energy and also the law of constant proportions.
30. Discuss Avogadro's hypothesis.

Unit 2 Structure of Atom

Q.No.1-12 : 1 Mark, 13-19 = 2 Marks, and 20-22 = 3 Marks Total = 35 Marks

1. From the following nuclei select the isotopes and isobars.

   ${}^{238}_{92}U$, ${}^{234}_{90}Th$, ${}^{234}_{92}U$, ${}^{234}_{91}Pa$, ${}^{238}_{93}Np$

2. What is Zeeman effect and Stark effect?
3. Write electronic configurations, of Cr and Zn^{2+}?
4. Define Aufbau's Principle.
5. Which of the following subshell is possible.

   1 f, 1 d, 3 d, 3 f

6. Explain Hund's rule of maximum multiplicity by taking an example of phosphorous.
7. Write down the quantum numbers 'n', 'l' and 'm' for the following orbitals.

   (i) $3d_{x^2-y^2}$ (ii) $4d_z^2$ (iii) $3d_{xy}$ (iv) $4d_{xz}$

8. Which of the following sets of quantum numbers are not possible?

   (i) $n = 3,\ l = 2,\ m = 0,\ s = -\frac{1}{2}$

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

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

   (iv) $n = 3,\ l = 1,\ m = 0,\ s = +\frac{1}{2}$

9.
10. 4f orbital can have maximum of 14 electrons. T/F
11. Why are Bohr's orbits called Stationary States?
12. What is the difference between atomic mass and mass number? 1X12 =12
13. Explain why the uncertainty principle is significant only for the microscopic particles and not for the macroscopic particles?
14. Why half-filled and fully filled orbitals are extra stable?
15. Why config of 'Cr' is $3d^5\ 4s^1$ and not $3d^4\ 4s^2$ and 'Cu' is $3d^{10}\ 4s^1$ and not $3d^9\ 4s^2$?
16. Give differences between orbit and orbital.
17. What is photoelectric effect? What is the effect of frequency and intensity on photoelectric effect?
18. Why large no. of lines appear in the spectrum of hydrogen although it contains only one electron?
19. Derive de Broglie relationship and give its significance. 2 X7 =14
20. Give important postulates of Bohr's model of an atom.
21. Discuss Planck's Quantum theory of Radiation.

22. Using the s, p, d, f, notations describe the following quantum no.

(a) n=1, l=0 (c) n=4; l=3 (d) n=4; l=2

(b) n=3, l=2 (d) n=5; l=4 (e) n=6; l=4

Some more questions.

23. Discuss important facts about photoelectric effect.
24. Discuss black body radiation. Also explain its reason.
25. What are emission and absorption spectra? Why dark lines appear in the absorption spectra?
26. What is the frequency and wavelength of a photon emitted during a transition from n=5 state to n=2 state in the hydrogen atom.
27. Discuss drawbacks of Rutherford's Model.
28. Explain Heisenberg's uncertainty Principle.
29. What do you understand by an atomic orbital? Briefly describe the shapes of s, p & 'd' orbitals?
30. State and explain Aufbau's principle, Pauli's exclusion principle.
31. Explain the properties of cathode rays.
32. How are anode rays produced?
33. Which of the following orbitals are not possible?

    1p, 2s, 2p, 3f, 3d, 4f, 4d

Unit 3 Classification of Elements

Q.No.1-12 : 1 Mark, 13-19 = 2 Marks, and 20-22 = 3 Marks   Total = 35 Marks

1. What are magic numbers?
2. Define Modern periodic law.
3. What are Dobereiner's triads?
4. Give general electronic configuration of 'd'-block and 'f'-block elements.
5. What are the defects of long form of the periodic table?
6. What is the cause of periodicity?
7. For electron affinity of halogen, which of the following is correct?

   (1) Br > F   (2) F > Cl   (3) Br > Cl   (4) F > I
8. Which one of the following arrangement represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species?

   (1) S < O < Cl < F   (2) Cl < F < S < O   (3) F < Cl < O < S   (4) O < S < F < Cl
9. Aluminium is diagonally related to ---------.
10. An element of atomic number 29 belongs to ------- block.
11. The electronic configuration of transition elements is exhibited by $ns^2 (n-1)d^{10}$. T/F
12. What are successive Ionization enthalpies?   12 X 1 =12
13. Why Ionization enthalpy of 'Be' is more than 'B' and of 'N' is more than 'O'. Explain?
14. Why electron gain enthalpies of Noble gases are positive while those of 'Mg' and 'P' are almost zero?
15. Why electron gain enthalpy of flourine is less negative than that of chlorine?
16. What are iso electronic species? How are their sizes vary in iso electronic series?
17. Which of the following will have the largest and smallest size and why?

    Cl, Cl^{-1}, Al, Al^{3+}
18. Why d- and f-block elements are less electropositive than group 1 and 2 elements?
19. What is diagonal relationship? Explain it with the help of 'Be' and 'Al'.   7 X 2 =14
20. What is ionisation enthalpy? On what factors it depends?
21. What is electron gain enthalpy? On what factors it depends. How it varies in a group and in a period?
22. How will you justify presence of 18 elements in 5th period and presence of 32 elements in 6th period?

Unit 4 Chemical Bonding and Molecular Structure

Q.No.1-12 : 1 Mark, 13-19 = 2 Marks, and 20-22 = 3 Marks Total = 35 Marks

1. Why do atoms combine?
2. What is the significance of Lewis Symbols?
3. Give structure of BrF_5.
4. Which of the following molecules are formed by p-p overlapping?
   (1) Cl_2 (2) HCl (3) H_2O (4) NH_3
5. CO_2 is isostructual with ------------.
6. The compound with the highest boiling point is:
   (1) CH_3OH (2) CH_3Br (3) CH_3Cl (4) CH_4
7. The hydrogen bond is strongest in:
   (1) O – H- - - S (2) S – H - - - O (3) F – H - - - F (4) F – H - - - O
8. The sigma and π-bonds present in benzene ring are --------- and ---------.
9. Why H_2O is liquid and H_2S is a gas?
10. Why NH_3 is liquid and PH_3 is a gas?
11. Boiling point of p-nitrophenol is more than O-nitrophenol why?
12. How is paramagnetic character of a compound is related to the no. of unpaired electrons? 12 X1 =12
13. Describe a co-ordinate bond with an example. How does it differs from a covalent bond?
14. How is MgF_2 and Al_2O_3 formed?
15. What is an Octet rule? What are its limitations?
16. Which out of NH_3 and NF_3 has higher dipole moment and why?
17. Draw molecular orbital diagram for N_2^+ molecule.
18. HCl is a covalent compound but it ionises in the solution?
19. The molecule of CO_2 is linear whereas that of $SnCl_2$ is angular why? 7X2 =14
20. Give molecular orbital energy level diagram of CO. Write its electronic configuration, magnetic behaviour and bond order.
21. How is ionic bond formed? On what factors it depends?
22. Calculate the lattice enthalpy of KCl from the following data by Born-Haber's Cycle.
    Enthalpy of sublimation of K=89 KJ mol^{-1}
    Enthalpy of dissociation of Cl_2 = 244 KJ mol^{-1}
    Ionization enthalpy of potassium = 425 KJ mol^{-1}
    Electron gain enthalpy of chlorine = - 355 KJ mol^{-1}
    Enthalpy of formation of KCl = -438 KJ mol^{-1}

More questions

23. How do atoms combine? Describe briefly.
24. Give characteristics of ionic compounds.
25. How is covalent bond formed discuss with the help of N_2, CH_4, C_2H_2?
26. Give postulates of VSEPR theory.
27. Discuss types of covalent bonds with the help of example. Why pi-bond can't exist independently?
28. Discuss the factors affecting bond enthalpy
29. Discuss the partial ionic character of covalent bond by taking an example.
30. Give applications of dipole moment.
31. Discuss partial covalent character of ionic bonds.
32. What is hybridisation? Discuss facts about hybridisation.
33. Give salient features of Molecular orbital theory.
34. Differentiate between bonding and anti bonding molecular orbitals.
35. Discuss the conditions for the combination of atomic orbitals to form molecular orbitals.
36. What are the consequences of hydrogen bonding?
37. Discuss types and conditions for hydrogen bonding.
38. Why density of water is maximum at 277K? Discuss.
39. Why KHF_2 exists while $KHCl_2$ does not?
40. Which is more polar and why, CO_2 or N_2O?

Unit 5 States of Matter

Q.No.1-12 : 1 Mark, 13-19 = 2 Marks, and 20-22 = 3 Marks Total = 35 Marks

1. How is the pressure of a gas related to its density at a particular temperature?
2. A gas occupies 180 mL at a pressure of 0.740 bar at 20^oC. How much volume will it occupy when it is subjected to external pressure of 1.025 bar at the same temp.?
3. A sample of gas occupies 2.50 L at 25^oC. If the temp is raised to 65^oC, what is the new volume of the gas if pressure remains constant?
4. Give physical significance of Gay Lussac's Law in daily life.
5. How is gas constant 'R' related to work?
6. Why drop of a liquid is spherical in shape?
7. 300 mL of a gas at 27^oC is cooled to -3^oc at constant pressure; the final volume is:
   (1) 540 mL (2) 135 mL (3) 270 mL (4) 350 mL
8. 273 mL of a gas at STP was taken to 27^oC and 600 mm pressure. The final volume of the gas would be:
   (1) 273 mL (2) 300 mL (3) 380 mL (4) 586 mL
9. The density of a gas is equal to P/RT T/F
10. The volume of a gas at 0^oC is 273 mL. It volume at 1^oC and same pressure will be:
    (1) $\frac{274}{273}$ mL (2) 274 mL (3) 272 mL (4) $\frac{273}{274}$ mL
11. If the pressure and absolute temperature of 2 litre of carbon dioxide are doubled, the volume of carbon dioxide would become ---------------------.
12. The V.D. of a gas is 11.2. The volume occupied by 11.2 g of this gas at NTP is:
    (1) 1 litre (2) 2.24 litre (3) 11.2 litre (4) 22.4 litre
13. Derive Ideal gas equation.
14. CO_2 is heavier than N_2 and O_2 gases present in the air but it does not form the lower layer of the atmosphere. Explain?
15. What is an ideal gas? Why do real gases deviate from ideal behaviour?
16. Will water boil at higher temp at Sea level or at the top of mountains and why?
17. Why does the temp. of the boiling liquid remains constant even though heating is continued?
18. Calculate the temp. of 5.5 moles of a gas occupying 6 dm^3 at 3.35 bar (R=0.083 bar dm^3 k^{-1} mol^{-1})
19. The drain cleaner contains small bits of aluminium, which react with caustic soda to produce H_2 gas. What volume of H_2 gas at 20^oC and one bar pressure will be released when 0.15 gm of 'Al' reacts?
20. Give various postulates of kinetic molecular theory of gases and also give its justification.
21. What is liquefacation of a gas? Discuss Andrew's isotherms far CO_2 and important conclusions.

22. According to kinetic theory, the forces of attraction between the gas molecules are negligible. Discuss it.

More questions :

23. Why mercury is used as a liquid in a barometer? Explain.
24. A one litre flask contain helium gas and 1.5 litre flask contains xenon gas at the same temp and pressure. What is the ratio of number of atoms in the two flasks?
25. How will you determine pressure of a dry gas by using Dalton's law of partial pressures?
26. Differentiate between diffusion and effusion. What is the cause of diffusion?
27. Two gases A & B having same volume diffuse through a porous partition in 30 secs. and 20 secs. respectively. The molecular mass of A is 45. Find the molecular mass of B.
28. Calculate the volume of oxygen that will diffuse in the same time as 50 ml of SO_2.
29. Discuss the factors on which vapour pressure depends.
30. What is the effect of temp. and pressure on surface tension and viscosity?
31. Discuss dipole – induced dipole forces with example.
32. Give characteristics of London forces.
33. Why CO_2 and NH_3 can be liquefied easily where as H_2, O_2 and N_2 cannot be liquefied.
34. Which out of the following will have higher vapour pressure at a given temp. and why? (a) Polar liquids like water (b) Non-Polar liquids like ether.
35. Why do gases deviate from ideal behaviour?
36. Compare the properties of solids, liquids and gases.
37. Why do ionic compds have higher m.pt. ?

Unit 6 Thermodynamics

One mark questions.

1. What is meant by extensive and intensive properties?
2. What is meant by State function and path function?
3. What is a perpetual motion machine? Is it possible?
4. Express the change in internal energy of a system when 'W' amount of work is done by the system and 'q' amount of heat is supplied to the system. What type of system would it be?
5. A system absorbs energy equivalent to 415 J and performs work equivalent to 205.15J. Calculate the change in internal energy of the system.
6. Why it is necessary to define the standard state?
7. Why a real crystal does have more entropy than an ideal crystal?
8. Predict the enthalpy change, free energy change and entropy change when ammonium chloride is dissolved in water and the solution becomes colder.
9. Decrease in enthalpy the only criterion for spontaneity. T/F
10. Tendency towards maximum randomness the sole criterion for spontaneity. T/F
11. Define Hess's law of constant heat Summation with suitable example.
12. Absolute value of internal energy cannot be determined. T/F
13. Ethanoic acid and hydrochloric acid react with sodium hydroxide solution. The enthalpy of neutralisation of ethanoic acid is -55.8 KJ mol^{-1} while that of hydrochloric acid is -57.3 KJ mol^{-1}. Can you think of the difference?
14. Discuss the effect of temperature on the spontaneity of an exothermic and endothermic reaction.
15. (i) Absoute value of internal energy cannot be determined. Explain.
    (ii) When ΔG is positive, the process is always non spontaneous. Explain.
16. (i) Explain the meaning of driving force of a chemical reaction. How is ΔG related to ΔH and ΔS in a reaction?
    (ii) How does $T\Delta S$ determine the spontaneity of a process?
17. (i) How will you justify that both 'q' and 'w' are not state functions, yet (q+w) is a state function?
    (ii) ΔH is negative for exothermic reaction and positive for endothermic reaction. Explain.
18. For a reaction both ΔH and ΔS are positive. Under what conditions will the reaction be spontaneous?
19. Determine $\Delta H^o{}_r$ at 298 K for the reaction.

    C(graphite) + $2H_{2(g)} \rightarrow CH_{4(g)}$; $\Delta H^o{}_r = ?$

    you are given

    (i) C(graphite) + $O_{2(g)} \rightarrow CO_{2(g)}$ $\quad \Delta H^o{}_r = -\ 393.5$ KJ mol^{-1}

    (ii) $H_{2(g)} + ½\ O_{2(g)} \rightarrow H_2O_{(l)}$ $\Delta H^o r = -285.8$ KJ mol^{-1}

    (iii) $CO_{2(g)} + 2H_2O_{(l)} \rightarrow CH_{4(g)} + 2O_2(g)$; $\Delta H^o{}_r = +890.3$ KJ mol^{-1}

20. For the reaction $A_{(g)}+3B_{(g)} \rightarrow 2C_{(g)}$, the enthalpy change is -90.2 KJ mol^{-1} and ΔS is -0.1584 KJ K^{-1} mol^{-1}. Predict whether the reaction is feasible or not at 298 K?

21. Enthalpy and entropy changes of a reaction are 49.57 KJ mol^{-1} and 123.2 J K^{-1} mol^{-1}. Calculate the free energy change of the reaction at $27^{o}C$.

22. Give reason why heat of neutralization less than 57.1 KJ mol^{-1} when. 0.1N Solution of acetic acid is neutralized by 0.1 N NaOH solution?

Unit 07 Equilibrium

Total = 30 Marks

One mark questions :

1. What do you mean by homogenous and heterogenous equilibria?
2. Write the expression for the equilibrium constant 'K' for each of the following reaction.
   (i) $2\ NOCl_{(g)} \rightleftharpoons 2\ NO_{(g)} + Cl_{2(g)}$
   (ii) $C_{(S)}+CO_{2(g)} \rightleftharpoons 2CO_{(g)}$
3. (i) $I_2(S)+5F_{2(g)} \rightleftharpoons 2IF_5$ write 'K'
   (ii) $FeO_{(S)} + CO_{(g)} \rightleftharpoons Fe(S) + CO_{2(g)}$ write 'K'.
4. What is the effect of reducing the volume on the system in equilibrium represented below :
   $2C_{(S)}+O_2(g) \rightleftharpoons 2\ CO_{(g)}$
5. What is the effect of increase of temperature on equilibrium constant for the following reaction.
   $I_{2(g)} \rightleftharpoons 2I_{(g)}$
6. The equilibrium constant expression for a gaseous reaction is $K_c = \dfrac{[NH_3]^4\ [O_2]^5}{[NO]^4\ [H_2O]^6}$
   write the balanced chemical equation corresponding to this expression.
7. What are conjugate acid-base pairs? Give an example.

Two marks questions.

8. Give limitations of Arrhenius concept of Acids and bases.
9. Give advantages of Bronsted-Lowry concept over Arrhenius concept.
10. (i) What will be the conjugate bases for the following Bronsted acids?
    HF, H_2SO_4, HCO_3, H_3PO_4
    (ii) What will be the conjugate acids for the following Bronsted bases?
    NH_2^-, NH_3, $HCOO^-$, ClO_4^-
11. Why PO_4^{3-} ion is not amphiprotic?
12. What is a buffer solution? Ammonium acetate is a buffer where as sodium chloride is not. Why?
13. What are acidic buffers? Explain with the help of an example.
14. What are basic buffers? Explain with the help of an example.

Three marks questions.

15. (i) Derive an expression for the calculation of the degree of ionization of a weak electrolyte.
    (ii) Why is ammonia termed as lewis base? Illustrate with two examples.
16. (i) Addition of a drop of HCl to an acidic buffer of acetic acid and sodium acetate does not produce any appreciable change in the pH of the solution. Why?
    (ii) A chemical equilibrium is dynamic in nature. Explain.

17. What do you mean by strength of an acid? How can the strength of the two acids be compared?

More questions.

18. What are the important characteristics of chemical equilibrium?
19. What is the difference between amphoteric and amphiprotic?
20. Using Le-chatelier's principle, predict the effect of (a) decreasing the temperature (b) increasing the pressure on the following system.

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2\ NH_{3(g)} + \text{Heat}$$

21. The dissociation constant of NH_4OH at 298 K is 1.8×10^{-5}. Calculate the degree of dissociation of 0.01 M Sol. of NH_4Cl. K_w at 298 K = 10^{-14}
22. Calculate the hydrolysis constant of the salt containing NO_2^- ions. K_a for HNO_2 = 4.5×10^{-10}
23. Determine the degree of hydrolysis at 0.10 M solution of sodium acetate at 298 K. (K_a for $CH_3COOH = 1.8 \times 10^{-5}$ & $K_w = 1 \times 10^{-14}$). Also calculate hydrolysis constant and pH.
24. The aqueous sol. of all Salts of weak acids and strong bases are alkaline. Justify it with the help of an example.
25. The aqueous sol. of all Salts of weak bases and strong acids are acidic. Justify it with the help of an example.
26. All Arrhenius acids are Bronsted-Lowry acids but all Arrhenius bases are not Bronsted-Lowry bases. Justify this statement with example.

Unit 08 Redox Reactions

Total = 30 Marks

One mark questions :

1. Calculate the oxidation number of Mn in $KMnO_4$ and 'Cr' in $K_2Cr_2O_7$.
2. Identify the oxidant and reductant in the following reactions.
   (i) $2Zn_{(S)} + O_{2(g)} \rightarrow 2\ ZnO_{(S)}$
   (ii) $I_{2(g)} + H_2S_{(g)} \rightarrow 2HI_{(g)} + S_{(s)}$
3. Which elements always have positive oxidation state?
4. What is the function of salt bridge?
5. Give applications of electrochemical series.
6. What are direct and indirect redox reactions?
7. Oxidation and reduction go side by side in a redox reaction. Justify it.

Two marks questions:

8. (i) Why are redox reactions called electron transfer reaction?
   (ii) Can the same element have different oxidation numbers in different compounds? Justify.
9. (i) What happens when a zinc rod is dipped in a copper sulphate solution?
   (ii) What are combination redox reactions and decomposition redox reactions? Give examples.
10. H_2S acts as a reducing agent while SO_2 acts as an oxidising as well as reducing agent. Explain.
11. Give important features of Half-cell reactions.
12. HNO_3 acts as an oxidising agent while HNO_2 can act both as a reducing agent as well as oxidising agent explain.
13. Give differences between oxidation no. and valency.
14. Are all decomposition reactions redox reactions? Comment.

Three marks questions :

15. What do you understand by metal displacement redox reactions? How these differ from non-metal displacement reactions?
16. (i) What would happen if no salt bridge is used in $ZnSO_4 - CuSO_4$ electro chemical cell?
    (ii) What happens when copper rod is dipped in $AgNO_3$ Solution?
17. Mention oxidation, reduction, oxidising agent and reducing agent in the following reactions.
    (i) $FeS_2+O_2 \rightarrow Fe_2O_3+SO_2$
    (ii) $NH_3+O_2 \rightarrow NO+H_2O$
    (iii) $SnO_2+C \rightarrow Sn+CO$

More questions :

18. Balance following equations by oxidation no. method.

    (i) $SnO_2 + C \rightarrow Sn + CO$

    (ii) $Zn + NO_3^- + H^+ \rightarrow Zn^{2+} + N_2O + H_2O$

    (iii) $NH_3 + O_2 \rightarrow NO + H_2O$

    (iv) $H_2S + Fe^{3+} \rightarrow Fe^{2+} + S + H^+$

19. Balance following equations by Ion-Electron method.

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

    (ii) $NO_3^- + Zn \rightarrow Zn^{2+} + NH_4^+$

20. Give differences between Electrochemical cell and Electrolytic cell.
21. What are disproportionation redox reaction? Give example.
22. Give limitations of concept of oxidation number.
23. Give advantage of electron density concept over oxidation no. concept.
24. Discuss the role of redox titrations in volumetric titrations.
25. Chlorine, bromine and iodine disproportionate in alkaline medium but fluorine does not. Why?
26. Give an important application of non-metal displacement redox reactions in qualitative mixture analysis.
27. Assign oxidation no. of the followings :

    (i) 'P' in NaH_2PO_4 (ii) 'S' in $NaHSO_4$

    (iii) 'P' in $H_4P_2O_7$ (iv) 'S' in $KAl(SO_4)_2$

    (v) 'Pt' in $[Pt(C_2H_4)Cl_3]^-$ (vi) 'Cl' in $KClO_4$

Unit 9 Hydrogen

Q.No.1-7 : 1 Mark, 8-14 = 2 Marks, and 15-17 = 3 Marks Total = 30 Marks

1. What are nuclear spin isomers of Hydrogen?
2. Why is dihydrogen not preferred in balloons?
3. Phosphorus forms only PH_3 and not PH_5 why?
4. How is temporary hardness of water removed?
5. Write the names of isotopes of hydrogen. What is the mass ratio of these isotopes?
6. What are electron deficient and electron rich compounds of hydrogen? Give examples.
7. Is distilled water useful for drinking purpose? If not, how can it be made useful?
8. What is autoprotolysis of water? What is its significance?
9. How does H_2O_2 behave as a bleaching agent?
10. What properties of water make it useful as a solvent? What types of compound can it dissolve and hydrolyse?
11. What do you understand by terms hydrolysis and hydration? Give examples.
12. What do you understand by term hydrogen economy?
13. H_2O_2 act both as oxidising and reducing agent, Justify it with the help of examples.
14. (i) How does H_2O_2 reacts with $KMnO_4$ in alkaline medium?
    (ii) How does H_2O_2 reacts with $K_2Cr_2O_7$ in acidic medium?
15. Give uses of heavy water. Can heavy water be used for drinking?
16. (i) How water act both as an oxidising and reducing agent? Give examples.
    (ii) What is coal gasification?
17. How hydrogen does resembles halogens and alkali metals and how it differs from them.

Unit 10 S-Block Elements   Alkali Metals

Q.No.1-7 : 1 Mark, 8-14 = 2 Marks, and 15-17 = 3 Marks   Total = 30 Marks

1. What is the cause of diagonal relationship?
2. Alkali metals have the lowest ionisation enthalpy in each period. Why?
3. The second ionisation enthalpies of alkali metals are very high?
4. All the alkali metals impart characteristic colour to flame. Why?
5. Alkali metals show photoelectric effect. Why?
6. Cesium show photoelectric effect to the maximum extent. Why?
7. Why alkali metals are soft and have low m.pt and b.pt.?
8. Alkali metals are very reactive. Justify with the help of examples.
9. Alkali metals are kept in kerosene oil why?
10. When alkali metals dissolves in liquid ammonia, the solution can acquire different colours. Explain the reason.
11. Why lithium shows anomalous behaviour?
12. Why lithium is the strongest reducing agent where as its ionization enthalpy is highest?
13. What is polarisation discuss it by taking example of lithium?
14. The hydroxides of alkali metals are strongly basic why?
15. How is sodium carbonate prepared by Solvay process?
16. How is sodium hydroxide prepared by Castner Kellner cell?
17. (i) Can we store sodium in water or not? Why.

    (ii) Write balanced equations for the following

    (a) Na_2O_2 and H_2O (b) KO_2 and H_2O
18. LiH is more stable than NaH. Explain.

Unit 10 S-Block Elements   Alkaline Earth Metals

Q.No.1-7 : 1 Mark, 8-14 = 2 Marks, and 15-17 = 3 Marks   Total = 30 Marks

1. Although IE_1 values of alkaline earth metals are higher than those of alkali metals, the IE_2 values of alkaline earth metals are much smaller than those of alkali metals why?
2. Calcium and strontium give characteristic colour to the flame but beryllium and magnesium do not give any characteristic flame colours. Why?
3. Hydration enthalpies of alkaline earth metals are larger than those of the corresponding alkali metals. Why?
4. Why beyllium and magnesium form complexes?
5. The hydroxides of alkaline earth metals are less basic than alkali metals of the corresponding periods. Why?
6. What is cement? What is its composition?
7. What is dead burnt plaster? How is it obtained?
8. (i) $BeCl_2$ can be easily hydrolysed why?

   (ii) What is the difference between quick lime, Slaked lime and lime water?
9. (i) Why are halides of beryllium polymeric?

   (ii) Explain why can alkali and alkaline earth metals not be obtained by chemical reduction methods?
10. What happens when

    (i) 'Mg' is burnt in air.   (ii) Quick lime is heated with silica

    (iii) Chlorine reacts with Slaked lime. (iv) Calcium nitrate is heated.
11. Why does the solubility of alkaline earth metal carbonates and sulphates decrease down the group?
12. Why does solubility of alkaline earth metal hydroxides increase down the group?
13. (i) What is hydrolith?

    (ii) Which out of Mg^{2+}, Ba^{2+}, Ca^{2+} has maximum ionic mobility in water and why?
14. How does quick lime reacts with water, carbon dioxide and phosphorous pentaoxide.
15. How is lime stone manufactured and how it reacts with HCl and H_2SO_4?
16. Discuss chemical properties of Slaked lime.
17. How beryllium behaves differently as compare to magnesium or compare physical and chemical properties of beryllium and calcium.

Unit 11 *p*-Block Elements   Boron. Family

Q.No.1-7 : 1 Mark, 8-14 = 2 Marks, and 15-17 = 3 Marks   Total = 30 Marks

1. Why atomic radii of 'Ga' is smaller than 'Al'?
2. BCl_3 is known but $Tl\ Cl_3$ is not known. Why?
3. The metallic character increases from boron to aluminium and then decreases from aluminium to thallium. Explain.
4. Boron is a non-metal where as 'Al' is a metal. Why?
5. What is inert pair effect?
6. The reducing character of elements of gr. 13 goes on decreasing down the group. Why?
7. BCl_3 acts as a lewis acid. How?
8. Discuss structure of diborane.
9. Why BCl_3 is a stronger lewis acid than BF_3?
10. BCl_3 exists as a monomer where as $AlCl_3$ exists as a dimer why?
11. Borazine is more reactive than benzene. Why?
12. Why alumina cannot be reduced by Carbon?
13. Why anhydrous aluminium chloride has a lower melting point than anhydrous aluminium flouride?
14. Why boron and thallium does not form B^{3+} and Tl^{3+} ions?
15. (i) Why ionisation enthalpy of 'Ga' is higher than that of 'Al'?
    (ii) Thallous compounds (Tl^{+}) are more stable than thallic (Tl^{3+}) compounds. Why?
16. Boron and Silicon are diagonally related to each other. Give chemical reactions to prove this.
17. (i) What is thermite welding?
    (ii) Why B-F bond length in BF_3 is Smaller than the expected value?
    (iii) BF_3 is not hydrolysed where as BCl_3 get easily hydrolysed. Explain.

Unit 11 *p*-Block Elements The Carbon-Family

One mark questions:

1. Tin and lead show '+2' and '+4' oxidation states but for lead compounds +2 oxidation state is more stable. Why?
2. $[SiF_6]^{2-}$ is possible where as $[CF_6]^{-2}$ is not possible. Why?
3. Which allotropes of Carbon acts as an abrasive and which as a lubricant?
4. Why is diamond denser than graphite?
5. What are Silicones?
6. Diamond is covalent, yet it has high melting point. Why?
7. $[SiF_6]^{2-}$ is known but $[SiCl_6]^{2-}$ is not. Give reason.

Two mark questions.

8. Why does carbon not form either C^{4+} or C^{4-} ions?
9. Give the differences in structures of the following pair of compounds: CO_2 and SiO_2.
10. Why $N(CH_3)_3$ is Pyramidal but $N(S_iH_3)_3$ is planar?
11. Why elemental Silicon does not form a graphite like structure, as carbon does?
12. Give Chemical reaction to show that Tin (II) is a reducing agent, whereas, lead (II) is not.
13. Why milkiness disappears when excess of CO_2 gas is passed through lime water?
14. The ionization enthalpy of lead is more than tin. Why?

Three marks questions.

15. (i) Why CO_2 has no net dipole moment? 1½
    (ii) Carbon forms covalent compounds whereas lead forms ionic compounds. Why? 1½
16. (i) Why BCl_3 and CCl_4 behave differently towards water?
    (ii) What are Silicates?
17. Write a short note on fullerenes.

Unit 12 Organic Chemistry : Some basic Principles and Techniques

Q.No.1-7 : 1 Mark, 8-14 = 2 Marks, and 15-17 = 3 Marks Total = 30 Marks

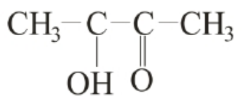
1. What is catenation?
2. What are homocyclic and heterocyclic compounds? Give examples.
3. What is a homologous series?
4. Define structural isomerism. Give structural isomers of butane.
5. Explain metamerism with example.
6. What is tautomerism? Give example.
7. Give all possible isomers of Hexane.
8. What is positive and negative inductive effect? Give examples.
9. What is electromeric effect? Discuss it with the help of an example.
10. Give resonating structures of $C_6H_5NH_2$ molecule.
11. What is homolytic and heterolytic fission?
12. What are free radicals? Which is the most stable free radical and why?
13. What is carbocation? Why tertiary carbocation is most stable?
14. What is carbanion? Which is the most reactive carbanion and why?
15. What are electrophiles and nucleophiles and what are their types? Discuss in detail.
16. What is resonance effect? Discuss positive and negative (+R; -R) effect with example.
17. What is hyperconjugation? Give applications of hyperconjugation.

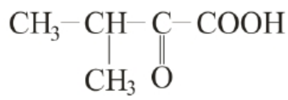
   Or

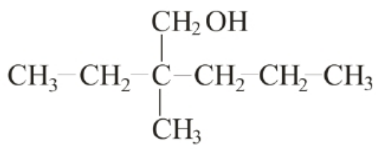
   Discuss Addition reaction and Elimination reaction in detail.

   Or

   Give IUPAC Name of the following:

   (i) 

   (ii) 

   (iii) 

Unit 13 Hydrocarbons   Alkanes & Alkenes

Q.No.1-7 : 1 Mark, 8-14 = 2 Marks, and 15-17 = 3 Marks   Total = 30 Marks

1. What are Saturated and unsaturated hydrocarbons?
2. How are alkanes prepared by Grignard's reagent?
3. Give mechanism of Wurtz reaction.
4. How will you convert acetaldehyde to ethane and acetone to propane?
5. Boiling points of isomeric alkanes goes on decreasing with increased branching. Why?
6. Alkanes with even no. of carbon atoms have high melting point as compare to alkanes with odd no. of carbon atoms why?
7. Give mechanism of sulphonation of alkanes?
8. *n*-pentane has higher boiling point than neo pentane. Explain.
9. Mention primary, secondary and tertiary carbons and hydrogens in the following compound.

$$\begin{array}{c} \qquad\qquad\quad CH_3 \\ \qquad\qquad\quad | \\ CH_3-CH_2-C-CH_3 \\ \qquad\qquad\quad | \\ \qquad\qquad\quad CH_3 \end{array}$$

10. Eclipsed conformation is less stable than staggered conformation of ethane. Explain.
11. What is geometrical isomerism and what is its cause?
12. What are the necessary conditions for the geometrical isomerisation?
13. How are alkenes prepared by Kolbe's Electrolytic process?
14. Why alkenes undergo electrophilic addition and not electrophilic substitution reaction?
15. (i) Explain and Justify Markownikoff's rule.
    (ii) Give mechanism of Kharash effect.
16. (i) Give ozonolysis reaction of ethene.
    (ii) How is structure of alkene elucidated by ozonolysis ?
17. (i) What is lindlar's catalyst? What is its use?
    (ii) Cis alkenes show higher boiling point as compared to trans-isomer. Why?

Unit 13 Hydrocarbons Alkenes and Alkynes

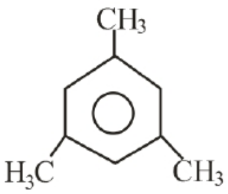
Q.No.1-7 : 1 Mark, 8-14 = 2 Marks, and 15-17 = 3 Marks Total = 30 Marks

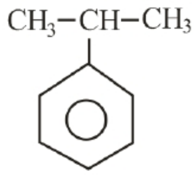
1. How will you prepare acetylene from calcium carbide?
2. Discuss Kolbe's electrolytic method to prepare acetylene.
3. Convert Chloroform into acetylene.
4. Convert methane into acetylene.
5. Alkynes do not exhibit geometrical isomerism while alkenes do so why?
6. Alkynes are less reactive than alkenes towards electrophilic addition reaction why?
7. Convert acetylene into ethanol.
8. Why does acetylene behave like a weak acid?
9. Write two reactions to show acidic nature of acetylene.
10. What is peroxide effect? Why is it applicable only in case of addition of HBr and not in case of HCl and HI?
11. Alkynes undergo both electrophilic and nuclephilic addition reactions. Why?
12. Discuss structure of alkyne.
13. Alkynes are acidic in nature. Explain.
14. Give reaction for the detection of terminal alkynes.
15. (i) Give mechanism of addition of halogens to alkynes.
    (ii) Why alkynes undergo nucleophilic addition reactions while simple alkenes do not?
16. (i) How will you convert acetylene into oxalic acid?
    (ii) How will you convert propyne into ethanoic acid?
    (iii) How will you convert acetylene into acrylic acid?
17. (i) How will you distinguish between Ethane and Ethyne? Give reaction.
    (ii) How will you distinguish between Ethene and Ethyne? Give reaction.
    (iii) How will you distinguish between propane and cyclopropane ? Give reaction.

Unit 13 Hydrocarbons Benzene

Q.No.1-7 : 1 Mark, 8-14 = 2 Marks, and 15-17 = 3 Marks Total = 30 Marks

1. What are aromatic hydrocarbons?
2. Give IUPAC names of the following compounds.

(i) 

(ii) 

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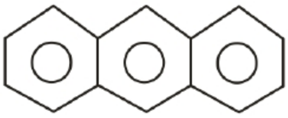
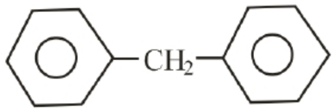
(i) 

(ii) 

4. What is aromaticity?
5. How will you convert n-hexane to benzene?
6. How will you convert benzene to benzoic acid?
7. How will you convert benzene to benzaldehyde?
8. What are activating groups? Explain it with example.
9. Give mechanism of nitration of chlorobenzene.
10. What are electron withdrawing groups? Why are they meta-directing?
11. Give mechanism of chlorination of Nitrobenzene.
12. Give mechanism of Friedal-craft acylation reaction.
13. (i) How will you convert benzene to benzophenone?
    (ii) How will you convert benzene to acetophenone?
14. Give mechanism of Sulphonation of benzene.
15. (i) Give mechanism of nitration of benzene.
    (ii) How will you prepare benzene from diene's?
16. How is structure of benzene deduced? Discuss in detail.
17. Discuss evidences in favour of resonating structure of benzene.

18. Why does benzene undergo electrophilic substitutions reactions easily and nucleophilic substitutions with difficulty?

19. How would you convert following compounds into benzene?

    (i) Ethyne (ii) Ethene (iii) Hexane

20. Arrange benzene, n-hexane and ethyne in decreasing order of acidic behaviour. Also give reasons.

21. Out of benzene, m-dinitrobenzene and toluene which will undergo nitration most easily and why?

22. Although benzene is highly unsaturated yet it does not prefer to undergo addition reactions. Explain.

23. Why is benzene extra ordinarily stable though it contains three double bonds?

24. What are the necessary conditions for any system to be aromatic?

Unit 14 Environmental Chemistry

Total 30 marks

One mark questions

1. List gases which are responsible for green house effect.
2. What is Smog?
3. How is classical smog different from Photochemical Smog?
4. Give one advantage and one disadvantage of ozone in the atmosphere.
5. What is meant by term 'Sink' and target with respect to pollution?
6. How plant nutrients and pesticides act as water pollutants?
7. What are polynuclear aromatic hydrocarbons (PAH)?

Two mark questions.

8. What are the harmful effects of PAH?
9. How NO_x pollution can be controlled? Explain.
10. (i) How lead halides enter into atmosphere as pollutants?
    (ii) How do contaminants differ from pollutants?
11. What are the primary and secondary pollutants of the air?
12. What is chemical oxygen demand? Explain.
13. What is biochemical oxygen demand? Explain.
14. What is meant by inversion temperature in different regions of the atmosphere?

Three mark questions.

15. Chlorine radical plays an important role in the destruction of ozone. Explain.
16. CO_2 is inert and harmless gas, yet it is thought to be a serious pollutant. Explain.
17. What are the reactions involved for ozone layer depletion in the stratosphere?

Some more questions.

18. Write down the reactions involved during the formation of Photochemical Smog.
19. Explain tropospheric pollution.
20. What are the harmful effects of photochemical Smog and how can they be controlled?
21. What do you mean by green chemistry? How will it help in decreasing environmental pollution?
22. How can domestic waste be used as manure?
23. What is acid rain? Give some of its harmful effects?
24. What is incineration? Explain.
25. Name and explain any four methods of waste management.