2. Chemistry

New Syllabus

Course Contents

General & Physical Chemistry (Section A)

Unit 1: Language of Chemistry (Review Lecturers) **- 3 teaching hours**

1. Chemical equations, their significances and limitations
2. Balancing chemical equations by :
   i. hit and trail method ii. Partial equation method
3. Types of chemical reaction

Unit 2: Chemical Arithmetic **- 17 teaching hours**

2.1 Dalton's atomic theory and Laws of Stoichiometry:

1. Postulates of Dalton's atomic theory
2. Law of conservation of mass
3. Law of constant proportions
4. Law of multiple proportions
5. Law of reciprocal proportions
6. Law of gaseous volumes
7. Chemical calculations based on stoichiometry

2.2. Atomic Mass and Molecular Mass:

Definition of atomic mass and molecular mass

1. Mole concept
2. Mole in term of mass, volume, number and ions
3. Calculation based on mole concept

2.3. Empirical, Molecular Formula and Limiting Reactants:

1. Percentage compositions
2. Derivation of empirical and molecular formula from percentage composition
3. Chemical calculation based on following chemical equation
   - Limiting reactants
   - Mass-mass relationship
   - Volume – volume relationship
   - Mass volume relationship

   (Solving related numerical problems))

2.4. Avogadro's Hypothesis and Its Applications:

1. Development of Avogadro's hypothesis
2. Definition of Avogadro's hypothesis
3. Application of Avogadro's hypothesis
   i. Deduction of atomicity of elementary gas
   ii. Deduction of relationship between molecular mass and vapour density
   iii. Deduction of molar volume of gases
   iv. Deduction of molecular formula from its volumetric composition (Solving related numerical problems)

2.5. Equivalent Mass:

1. Concept of equivalent mass
2. Equivalent weight of elements, and compounds (Salt, acid , base, oxidising agents, reducing agents)
3. Gram equivalent weight (GEW)
4. Relation between equivalent weight, valency and atomic weight
5. Determination of equivalent weight of metal by
   i. Hydrogen displacement method ii. Oxide formation method (Solving related numerical problems)

Unit 3: State of Matter **- 14 teaching hours**

3.1. Gaseous State:

1. Boyle's law
2. Charle's law and Kelvin scale of temperature
3. Application of Charle's law and Boyle's law
4. Combined gas law, ideal gas equation and universal gas constant
5. Dalton's law of partial pressure
6. Mathematical derivation of Dalton's law and their applications
7. Graham's law of diffusion and its applications
8. Kinetic model of gas and its postulates
9. Ideal and real gases
10. Deviation of gas from ideal behaviour (Solving related numerical problems)

3.2 **Liquid State:**

1. Physical properties of liquid
   i. Evaporation and condensation
   ii. Vapour pressure of liquid and boiling
   iii. Surface tension
   iv. Viscosity
2. Solution and solubility:
   i. Equilibrium in saturated solution
   ii. Solubility and solubility curve and its applications.
   (Solving related numerical problems)

3.3. **Solid Statem:**

1. Crystalline and amorphous solids
2. Water of crystallization
3. Efflorescences
4. Deliquesces
5. Hygroscopic
6. Seven types of crystal system
7. Simple cubic, face centered and body centered

Unit 4: Atomic Structure **- 10 teaching hours**

1. Discovery of fundamental particles of atom (electron, proton and neutron)
2. Concept of atomic number, mass number, fractional atomic mass, isotopes, isobars
3. Rutherford's α ray scattering experiment and nuclear model of atom; limitation
4. Bohr's model of atom and explanation of hydrogen spectra
5. Limitation of Bohr's model of atom
6. Elementary idea of quantum mechanical model
   i. Dual nature of electron (de-Broglie equation)
   ii. Heisenberg's uncertainty principle
   iii. Probability concept
7. Shape of atomic orbital (s and p orbitals only)
8. Quantum numbers
9. Pauli's exclusion principle
10. Hund's rule of maximum multiplicity
11. Aufbau principle and Bohr Bury rule
12. Electronic configuration of the atoms and ions(Z = 1 to 30)

Unit 5: Nuclear Chemistry **- 3 Teaching hours**

1. Concept of radioactivity
2. Radioactive rays (alpha ray, beta ray & gamma ray)
3. Meaning of natural and artificial radioactivity
4. Nuclear reactions, Nuclear energy (fission and fusion)
5. Nuclear isotopes and uses

Unit 6: Electronic Theory of Valency and Bonding **- 8 teaching hours**

1. Basic assumption of electronic theory of valency
2. Octet rule
3. Ionic bonds, ionic compounds and characteristics of ionic compounds. Lewis symbol to represent the formation of ionic compounds
4. Covalent bonds, covalent compounds and characteristics of covalent compounds – Lewis structure of some typical covalent compounds
5. Co-ordinate covalent bonds. Lewis structures of some typical co-ordinate covalent compounds
6. Exception of the octet rule
7. Partial ionic characters of covalent compounds. Non-polar and polar covalent molecules
8. Dipole moments and its application
9. Some special types of bonds: hydrogen bond and its types, metallic bond, vander Waal's bond, Resonance and resonance hybrid structures of O_3, SO_3, SO_4, $CO_3{}^{2-}$, SO_4^{2-}, $PO_4{}^{2-}$, NO_3^-
10. Classification of crystalline solids
    i. Ionic solid
    ii. Covalent solid
    iii. Molecular solid
    iv. Metallic solid

Unit 7: Periodic Classification of Elements **- 6 teaching hours**

1. Introduction
2. Mendeleev's periodic law and periodic table
3. Anamolies of Mendeleev's periodic table
4. Modern periodic law, and modern periodic table
5. Advantages of modern Periodic table

6. Division of elements into s,p, d and f blocks
7. Periodicity of physical properties: valency , atomic radii, ionic radii, ionisation energy, electron affinity and electronegativity (general trends only)

Unit 8: Oxidation and Reduction - 6 teaching hours

1 Classical concept of oxidation and reduction
2 Electronic interpretation of oxidation and reduction
3 Oxidation number and rules for the assignment of oxidation number
4 Differentiate between oxidation number and valency
5 Oxidising and reducing agent
6 Redox reaction
7 Balancing redox reactions by
i. oxidation number method ii. ion-electron method

Unit 9: Equilibria - 5 teaching hours

1. Introduction
2. Equilibrium involving in physical change
3. Chemical equilibrium
- Reversible and irreversible reactions
- Dynamic nature of chemical equilibrium and its characteristics
- Law of mass action
- Equilibrium constant (Kc) and its characteristics
- Homogenous and heterogeneous equilibrium
- Relation between Kp and Kc (derivation)
- Le-chatelier's principle and its application (No numerical is required)

Inorganic Chemistry
Section B
Unit 10: Non – Metals I - 12 teaching hours

10.1 Hydrogen:
1 Position in periodic table
2 Atomic hydrogen , Nascent hydrogen
3 Isotopes of hydrogen
4 Ortho and Para hydrogen
5 Applications

10.2. Oxygen:
1 Position in periodic table
2 Types of oxides
3 Uses of oxygen

10.3. Ozone:
1 Occurrence
2 Preparation from oxygen
3 Structure of ozone
4 Important properties of ozone
5 Ozone layer and ozone hole
6 Uses of ozone

10.4. Water:
1 Structure
2 Solvent property of water
3 Heavy water and uses
4 Uses

10.5 Nitrogen and Its Compounds:
1 Position of nitrogen in Periodic table
2 Uses of nitrogen
3 Types of nitrogen oxides (name and Lewis structure)
4 Ammonia
- manufacture by Haber's synthesis method
- Physical properties, chemical properties and uses
5 Oxyacids of nitrogen (type)
6 Technical production of nitric acid by Ostwald method
- Properties of nitric acid and uses.
- Test of nitrate ion

Unit 11: Non-Metals II - 23 teaching hours

11.1 Halogens: (Chlorine, Bromine and Iodine)
1 Position in periodic table
2 Comparative study on: preparation, properties and uses
3 Manufacture of bromine from carnallite process and manufacture of iodine form
i. sea weeds (principle only) ii. caliche (Principle only)
4 Uses of halogens
5 Comparative study on ; preparation, properties and uses of haloacids (HCl, HBr and HI)

11.2. Carbon:

1 Position in periodic table
2 Allotropes of carbon including fullerenes
3 Laboratory preparation, properties and uses of carbon monoxides

11.3. Phosphorous:

1 Occurrence, position in periodic table
2 Allotropes of phosphorous and uses of phosphorus
3 Preparation, properties and uses of phosphine
4 Oxides and oxyacids of phosphorous (structure and uses)
5 Preparation, properties and uses of orthophospheric acid

11.4. Sulphur:

Position in periodic table and allotropes
1 *Hydrogen Sulphide:* (Laboratory methods and Kipp's apparatus), properties and uses of
2 *Sulphurdioxide :* Laboratory preparation, properties and uses
3 *Sulphuric acid:* Manufacture by contact process, properties and uses
4 *Sodiumthiosulphate (hypo):* formula and uses

11.5. Boron and Silicon:

1 Occurrences, position in periodic table
2 Properties and uses
3 Formula and uses of borax, boric acid, Silicate and Silica

11.6. Noble gas: Position in periodic table, occurrence and uses

11.7. Environmental Pollution:

- Air pollution, photochemical smog
- Acid rain, water pollution
- Green house effect

Unit 12: Metal and Metallurgical Principles **- 6 teaching hours**

1 Characteristics of metals, non-metals and metalloids
2 Minerals and ores
3 Important minerals deposit in Nepal
4 Different process involved in metalurgical process
5 Concentration
6 Calcination and roasting
7 Smelting
8 Carbon reduction process
9 Thermite process
10 Electrochemical reduction
11 Refining of metals: poling, electro-refinement etc

Unit 13: Alkali and Alkaline Earth Metals **- 10 teaching hours**

1 Periodic discussion and general characteristics.
2 Sodium: Occurrence, Extraction from Downs process; properties and uses.
3 Sodium hydroxide: Manufacture, properties and uses.
4 Sodium carbonate: Manufacture, properties and uses.

13.1 Alkaline Earth Metals:

1 Periodic discussion and general characteristics
2 Preparation, properties and uses of i. quick lime, ii. plaster of Paris iii. bleaching powder, iv. magnesia v. Epsom salt.

Organic Chemistry

Section C

Unit 14: Introduction to Organic Chemistry

14.1 Fundamental Principles: **- 6 teaching hours**

1 Definition of organic chemistry and organic compounds
2 Origin of organic compounds (vital force theory)
3 Reasons for the separate study of organic compounds
4 Tetra covalency and catenation property of carbon
5 Classification of organic compounds
6 Functional groups and homologous series
7 Meaning of empirical formula, molecular formula, structural formula and contracted formula
8 Qualitative analysis of organic compounds. (detection of N,S and halogens by Lassaigne's test)

14.2. Nomenclature of Organic Compounds: **- 6 teaching hours**

1 Common names
2 IUPAC system and IUPAC rules of naming hydrocarbons, alcohols, ethers, aldehydes, Ketones, carboxylic acid, amines, ester, acid derivative halogen derivatives, nitriles etc.)

14.3. Structure Isomerism in Organic Compounds: **- 2 teaching hours**

1 Definition of structure isomerism

2 Types of structure isomerism: chain isomerism, position isomerism, functional isomerism and metamerism

14.4 Preliminary Idea of Reaction Mechanism **- 2 teaching hours**

1 Concept of homolytic and herterolytic fission
2 Electrophile , nucleophiles and free- radicals
3 Inductive effect, +I and -I effect

Unit 15: Hydrocarbons

15.1 Sources: **- 4 teaching hours**

Origin of coal and petroleum , hydrocarbon from petroleum cracking and reforming , aliphatic and aromatic hydrocarbon form coal, quality of gasoline, octane number and gasoline additive.

15.2 Alkanes (Saturated Hydrocarbons):

1 General methods of preparations :
- Decarboxylation
- Catalytic hydrogenation
- Reduction of haloalkane
- Kolbe's electrolysis method
- Using Grigrand's reagent
- Wurtz reaction
- From aldelydes and ketones

2 Physical properties
3 Chemical properties: Substitutions reaction, oxidation, pyrolysis or cracking aromatization

15.3. Alkenes : **- 4 teaching hours**

1 General methods of preparation
- Dehydration of alcohol
- Dehydrohalogenation
- Catalytic hydrogenation of alkyne
- Kolbe's electrolysis

2 Laboratory preparation of alkene
3 Chemical properties of alkene: Addition reaction (H_2, X_2, HX, H_2O, O_3, H_2SO_4)
4 Oxidation with alkaline $KMnO_4$ (Baeyer's reaction)
5 Polymerisation
6 Test of ethene and uses

15.4. Alkynes : **- 3 teaching hours**

Ethyne

1 Preparation form i. carbon and hydrogen ii. Kolbes electrolysis iii. 1,2 dibromoethane
2 Lab preparation of ethyne
3 Physical properties
4 Chemical properties: Addition (H_2, X_2, HX, H_2O, O_3), Acidic nature (action with ammonical $AgNO_3$ and ammonical Cu_2Cl_2) , Oxidation with alkaline $KMnO_4$, Polymerization uses of ethyne

Practical

Full Marks: 25 ***Pass Marks: 10***

Students are required to secure the pass marks in the practical paper separately. The following is the list of experiments. The students are required to perform in the practical classes in Grade XI.

A. Experiments based on laboratory techniques:

1. To separate the insoluble component in pure and dry state from the given mixture of soluble and insoluble solids. (NaCl and sand)
2. To separate volatile component form the given mixture of volatile and non volatile (demonstration of sublimation process)
3. To separate a mixture of two soluble solids by fractional crystallization (KNO_3 + NaCl)
4. To prepare a saturated solution of impure salt and obtain the pure crystal of the same salt by crystallization
5. To separate the component of a mixture of two insoluble solids (The being soluble in dil acids)
6. To obtain pure water from given sample of water (Distillation).

B. Experiment to study the different reactions (Neutralization, Precipitation, Redox reaction, electrolysis):

7. To perform precipitation reaction of $BaCl_2$ and H_2SO_4 and obtain solid $BaSO_4$;
8. To neutralize sodium hydroxide with hydrochloric acid solution and recover the crystal of sodium chloride
9. To test the ferrous ions in the given aqueous solution and oxidise it to ferric ion (Ferrous $\rightarrow$ Ferric system) Redox Reaction

10. To study the process of electrolysis and electroplating.

C. Experiments on quantitative analysis:

11. To determine the equivalent weight or weight of metal by hydrogen displacement method;
12. To determine the solubility of the given soluble solid at laboratory temperature;
13. To determine the relative surface tension of unknown liquid by drop count method;and
14. To study the rate of flow of liquid through Ostwald's viscometer and determine the relative viscosity of unknown liquid.

D. Experiments on preparation of gas and study of properties:

15. To prepare and collect hydrogen gas and study the following properties;
    a. Solubility with water, colour, odour;
    b. Litmus test;
    c. Burning match stick test; and
    d. Reducing properties of nascent hydrogen.
16. To prepare and collect ammonia gas and investigate the following properties:
    a. Solubility with water / colour / odour;
    b. Litmus test;
    c. Action with copper sulphate solution; and
    d. Action with mercurous nitrate paper.
17. To prepare carbondioxide gas and investigate the following properties:
    a. Solubility,colour, odour;
    b. Litmus paper test;
    c. Lime water test; and
    d. Action with burning magnesium ribbon.
18. To study the properties of hydrogen sulphide (Physical, analytical and reducing);
19. To study the following properties of sulphuric acid:
    a. Solubility with water;
    b. Litmus paper test;
    c. Precipitating reaction;and
    d. Dehydrating reaction.

E. Experiments on qualitative analysis:

20. To detect the basic radical of the given salt by dry way and the acid radical by dry and wet ways. Basic radicals: Zn^{++}, Al^{+++}, NH_4^+, Ca^{++}, Na^+

    Acid radicals: CO_3^{--}, SO_4^{--} NO_3^-, Br^-, I^-, Cl^-

Note: Experiment from no 1 to 19 requires one practical period of each experiment and the experiment no 20 requires four practical periods. (Two theory periods will be equivalent to one practical period)

Evaluation Scheme

The chemistry theory paper (XI) will consist of three types of questions:

(a) Very short-answer questions (weightage of 2 marks of each);

(b) Short-answer questions (weightage of 5 marks of each); and

(c) Long- answer questions (weightage of 10 marks of each).

According to manner of questions groups are divided into group 'A', group 'B' and group 'C'.

1 Group 'A' will consist of twenty two (22) very short questions, out of which, examinees are required to answer only fifteen (15) questions.

2 Group 'B' will consist of seven (7) short questions, out of which examinees are required to answer five (5) questions.

3 Group 'C' will consist of four (4) questions, out of which examinees are required to answer 2 questions.

The weightage of content distribution for the three types of questions form different sections of the curriculum will be as follows:

	Units	Teaching hours	V.S.O.	S.Q.	L.Q.
	1	3	x		
	2	17	2		
	3	14	2		
	4	10	2		
	5	3	1		
	6	8	2		
	7	6	1		
	8	6	1		
	9	5	1		
	10	12	2		
	11	23	2		
	12	6	1		
	13	10	1		
	14	16	2		
	15	11	2		
Total	15	150	22	7	4

Model Question

Time: 3 hrs.

Full Marks:- 75
Pass Marks:- 27

Group 'A'

Attempt any **fifteen** questions. $15 \times 2 = 30$

Q.1. State Law of reciprocal proportion. **(From Unit 2)**

Q.2. The oxide of an element contains 67.67% of oxygen and V.D of its volatile chloride is 79. Calculate the atomic weight of the element. **(From Unit 2)**
[Ans: 15.28]

Q.3. Define surface tension of liquid. **(From Unit 3)**

Q.4. Write down important differences between crystalline and amorphous solid. **(From Unit 3)**

Q.5. What is Pauli's exclusion principle? **(From Unit 4)**

Q.6. Give the values of all four quantum numbers of 11th electron of magnesium. (At. no. 12). **(From Unit 4)**

Q.7. What is radioactivity? **(From Unit 5)**

Q.8. Write the Lewis structure of (a) H_2O_2 and (b) HNO_3. **(From Unit 6)**

Q.9. Explain why HCl has polar character though it has covalent bond. **(From Unit 6)**

Q.10. State Modern periodic law. **(From Unit 7)**

Q.11. How would you show the following reaction is a redox reaction?

$$Mg + Cl_2 \longrightarrow MgCl_2$$ **(From Unit 8)**

Q.12. State Le-Chateliers principles. **(From Unit 9)**

Q.13. What are the differences between nascent and molecular hydrogen? **(From Unit 10)**

Q.14. Name any two oxides of each of the following:
(i) amphoteric
(ii) neutral. **(From Unit 10)**

Q.15. What is allotropy? Name the latest discovered allotropic form of carbon. **(From Unit 11)**

Q.16. What is meant by acid rain ? Give one major effect of acid rain. **(From Unit 11)**

Q.17. Distinguish between flux and slag with one example of each. **(From Unit 12)**

Q.18. Can sodium be extracted by the electrolysis of aqueous solution of sodium chloride? If not why? **(From Unit 13)**

Q.19. Define electrophile and nucleophile with an example of each. **(From Unit 14)**

Q.20. Give IUPAC name of the following compound. **(From Unit 14)**

(a) $CH_3-CH(Br)-CH(NH_2)-COOH$

(b) $CH_3-C(CH_3)_2-C(OH)H-CH_3$

Q.21. What is meant by thermal cracking and catalytic cracking? **(From Unit 15)**

Q.22. Identify A and B in the following reaction and give their IUPAC name **(From Unit15)**

$$C_2H_6 \xrightarrow[h\nu]{Cl_2} (A) \xrightarrow[\Delta]{alc.\ KOH} (B)$$

Group 'B'

Attempt any five questions. **5×5= 25**

Q.23. Urea [$(NH_2)_2CO$] is prepared by reacting ammonia with carbondioxide:

$2NH_3(g) + CO_2(g)$Æ $(NH_2)_2CO(aq.) + H_2O(l)$

In one process, 637.2g NH_3 is treated with 1142g of CO_2.

(a) Which of the two reactants is the limiting reactant?

(b) Calculate the mass of urea formed.

(c) How much excess reagent (in gram) is left at the end of the reaction? **(From Unit 2.2)**

[Ans: (a) NH_3 (b) 1123.2 g (c) 318.56 g]

Q.24. State Avogadro's law. Using the law to deduce relationship between Molecular mass and vapour density. **(From Unit 2.4)**

Q.25. State and explain Hund's rule of maximum multiplicity. **(From Unit 4)**

Q.26. Specify oxidation half, reduction half, oxidizing agent and reducing agent. Balance the following equation by ion-electron or oxidation number method:

$Fe^{2+} + H^+ + NO^-_3 \rightarrow Fe^{3+} + NO + H_2O$ **(From Unit 8)**

Q.27. Describe the manufacture of nitric acid by Ostwald's process. **(From Unit 10.5)**

Q.28. Explain the laboratory preparation of carbon monoxide in laboratory. What happens when CO is passed through finely divided nickel? (From Unit 11.2)

Q.29. Describe the detection of foreign elements (N, S, X) in organic compounds. **(From Unit 14.1)**

Group 'C'

Attempt any two questions. **2×10=20**

Q.30. State Boyle's law and Charle's law. Derive PV = nRT. What is the density (in gram per litre) of ammonia at STP if the gas in a 1.0L bulb weights 0.672g at 25°C and 733.4 mm Hg pressure. **(From Unit 3.1)**

[Ans: 0.76 gL^{-1}]

Q.31. Describe the manufacture of sodium carbonate by ammonia soda process. Also mention the function of lime stone in the manufacture process. **(From Unit 13)**

Q.32. Write down the process involved in the manufacture of caustic soda by Solvay-Kellner's process. **(From Unit 13)**

Q.33. Write short note on any two:

(a) Relationship between Kp and Kc **(From Unit 9)**

(b) Manufacture of H_2SO_4 along with a flow-sheet diagram. **From Unit 11.4)**

(c) Functional group of organic compound. **(From Unit 14.1)**

(d) Laboratory preparation of ethene gas **(From Unit 15.3)**

HSEB Questions

Section A: General & Physical Chemistry

Unit 1 – Language of Chemistry

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. Define a chemical change and point out its two important characters. **[Q.N.2, 2056]**
2. What information can you obtain from the symbol $^{39}_{19}K$. **[Q.N.6, 2056]**
3. How many electrons and neutrons are present in the symbol $^{27}_{13}Al^{3+}$? **[Q.N.10, 2059]**
4. Two elements A and B have outermost shell electronic configuration $3s^1$ and $2s^2 2p^4$ respectively, then name the chemical formed between them. **[Q.N. 9, 2064]**

Unit 2: Chemical Arithmetic

2.1 Dalton's Atomic Theory and Laws of Stoichiometry

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. State Law of Multiple proportion. **[Q.N.7, 2055]**
2. State the law of Constant proportions. **[Q.N.1, 2056]**
3. State the law of Reciprocal proportions. **[Q.N.1, 2057]**
4. State the law of Reciprocal proportions. **[Q.N.3, 2058]**
5. H and O react separately to give H_2O_2 and H_2O respectively. What law of stoichiometry is illustrated ? State the law. **[Q.N.5, 2062]**
6. State the law of Reciprocal proportion. **[Q.N. 2. 2063]**
7. State the Law of Reciprocal proportion. **[Q.N. 13, 2065]**
8. Phosphorous reacts with Oxygen to Produce P_2O_3 and P_2O_5 respectively. Which chemical law does the data illustrate? State the Law. **[Q.N.4, 2066]**
9. State Law of conservation of mass. Why is this law known as Law of indestructibility of matter ? 1+1 **[Q.N. 5, 2067]**
10. 12 g of Carbon react with 32 g of Oxygen of produce 44 g of Carbondioxide. Which Chemical Law do these data illustrate ? State the law. **[Q.N.2,2068]** 1+1=2
11. State law of Reciprocal proportions. 2**[Q.N. 2, Supp. 2068]**
12. State the law of Multiple proportions. 2**[Q.N. 1, Set 'A' 2069]**
13. State law of equivalent proportions. **[Q.N. 7(Or), Set 'B' 2069]**
14. 16 g of Methane can be produced by combining 12 g of carbon with 4g hydrogen. Which chemical law do this data illustrate? State the Law. **[Q.N. 2, Supp. 2069]**

Short Questions

(All questions are of equal value, 5 marks each.)

1. A metal X, forms two oxides A and B, 3.000 g of A and B contain 0.720 g and 1.160 g of oxygen respectively. Calculate the masses of metal in gram which combine with one gramme of oxygen in each case. What chemical law do these masses of metal illustrate ? State the chemical law. **[Group C, Q.N.1, 2052]**
2. State and explain law of multiple proportion. A certain element X_1 forms three different binary compounds with chlorine, containing 50.68%, 68.95% and 74.75% chlorine, respectively. Show how these data illustrate the law of multiple proportions. **[Q.N.28, 2054]**

3. State the law of Reciprocal proportions.
   (a) 0.46 g of a metal produced 0.77 g of metal oxide.
   (b) 0.805 g of the same metal displaced 760 cc of H_2 gas at NTP from HCl.
   (c) 1.26 g of water was formed by the union of 1.12 g of oxygen with hydrogen.
   Show that these data illustrate the law of Reciprocal proportions.
   [Q.N.25, 2060]

Long Questions

(All questions are of equal value, 10 marks each.)

No any questions have been asked on this section up to now.

2.2 Atomic Mass and Molecular Mass

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. Explain why atomic weights of elements are not whole numbers ? **[Group A, Q.N.7, 2051]**
2. Explain why atomic weights of elements are not whole numbers. **[Q.N.2, 2055]**
3. Explain why atomic weight of the elements are not whole number. **[Q.N.4, 2058]**
4. Which of the followings has larger number of molecules and how? 7 gram of nitrogen or 1 gram of hydrogen. **[Q.N.1, 2066]**
5. Which of the following gases has greater number of hydrogen molecule? 9g of CH_4 or 10g of NH_3. **[Q.N. 1, 2070 'D']**
6. What mass of H_2 gas will react with 22.4 litres of O_2 at STP to produce 36 gram of water. **[Q.N. 2, 2070 'D']**

Long Questions

(All questions are of equal value, 10 marks each.)

1. Write short notes on:
   (a) Victor Meyer's method of determination of molecular weight of volatile substances. **[Q.N.31 (b), 2055]**

Numerical Problems

1. How many molecules are contained in 0.35 mole of N_2? **[Group A, Q.N.6, 2051]**
   [Ans: 2.10×10^{23} molecules]
2. a) How many grams of H_2S are contained in 0.400 mole of H_2S ?
   [Ans: 13.6 g]
   b) How many gram–atoms of H and of S are contained in 0.400 mole of H_2S ?
   [Ans: 0.8 gram atom of H and 0.4 gram atom of S]
   c) How many grams of H and of S are contained in 0.400 mole of H_2S ?
   [Ans: 0.89 g of H and 12.8 g of S]
   d) How many molecules of H_2S are contained in 0.400 mole of H_2S ?
   [Ans: 2.4092×10^{23} molecules]
   e) How many atoms of H and of S of are contained in 0.400 mole of H_2S ?
   5 **[Group C, Q.N.1, 2051]**
   [Ans: 4.818×10^{23} atoms of H, 2.4092×10^{23} atoms of S]
3. Copper has a density of 8.92 g/mL If 1 mole of copper were shaped into a cube, what would be the length of the side of the cube (at. wt. of copper = 63.5) ? 2 **[Group A, Q.N.4, 2052]**
   [Ans: 1.92 cm]
4. How many moles of atoms is contained in 15 g of Zn.
   [Ans: 0.229 moles] 2**[Group A, Q.N.10, 2052]**
5. What is the weight (in g) of 0.5 atom of oxygen ? 2 **[Group A, Q.N.20, 2052]**
   [Ans: 1.328×10^{23} grams]
6. Define atomic mass of an element. Chlorine naturally is made up of 75%, Cl-35 and 25% Cl-37. Calculate the element atomic mass of chlorine.
   [Ans: 35.5] 5 **[Group B, Q.N.3, 2052]**

7. A jar containing 0.400 mol. of H_2S. Calculate the following:
   a) How many grams of H_2S?
   [Ans: 13.6 g]
   b) How many mole of H and S?
   [Ans: 0.8 mol of H and 0.4 mol of S]
   c) How many grams of H and S?
   [Ans: 0.89 g of H and 12.8 g of S]
   d) How many molecules of H_2S?
   [Ans: 2.4092×10^{23} molecules]
   e) How many atoms of H and S?
   [Ans: 4.818×10^{23} atoms of H, 2.4092×10^{23} atoms of S]
   5 [Group C, Q.N.1, 2053]

8. The cost of table sugar ($C_{12}H_{22}O_{11}$) is Rs. 24 per kg. Calculate its cost per mole.
   [Ans: Rs. 8.20 per mole] 2 [Q.N.2, 2054]

9. If 32 g of O_2 contains 6.022×10^{23} molecules at NTP, how many molecules under the same condition 32 g of S will contain? 2 [Q.N.5, 2055]
   [Ans: 6.023×10^{23} molecules]

10. 1 mole of a compound contains 1 mole of C and 1 mole of O. What is the molecular weight of the compound? 2 [Q.N.6, 2055]
    [Ans: Mo. wt = 28]

11. What will be molecular weight of a gas, 11.2 litre of which at NTP weighs 14 g ? 2 [Q.N.13, 2055]
    [Ans: Mol. wt = 28]

12. How many moles of oxygen molecules are present in 112 mL of O_2 gas at NTP.
    [Ans: 0.005 moles] 2 [Q.N.3, 2056]

13. Calculate the number of atoms of carbon present in 25g $CaCO_3$
    (Ca = 40, C = 12, O = 16) 2 [Q.N.2, 2057]
    [Ans: 1.505×10^{23} atoms]

14. Make the following conversions
    (a) 2.62 moles of C to grams of C
    [Ans: 31.44 g]
    (b) 28.0 grams of N_2 to moles of N_2
    [Ans: 1 mole] 2 [Q.N.1, 2058]

15. Find the number of moles of molecules present in 50 mL of an ideal gas exerting a pressure of 770 mm at 25°C (R = 0.082 L atm $mol^{-1}K^{-1}$)
    [Ans: 2.07×10^{-13} moles] 2 [Q.N.1, 2060]

16. How many molecules are contained in 0.35 mole of N_2 ?
    [Ans: 2.10×10^{23} molecules] 2 [Q.N.1, 2061]

17. One atom of an element 'X' weighs 6.644×10^{-23}g. Calculate the number of gram atoms in 80 kg of it. 2 [Q.N.2, 2061]
    [Ans: 2000 gram atoms]

18. How many moles of hydrogen are left when 3×10^{21} molecules of hydrogen are removed from a vessel containing 40mg of hydrogen ?
    [Ans: 0.015 mole of H left] 2 [Q.N.2, 2062]

19. How heavy is one atom of hydrogen ?
    [Ans: 1.67×10^{-24} g] 2 [Q.N.1(a), 2063]

20. How may moles of CO_2 are there in 4.4g of carbon dioxide ?
    [Ans: 0.1 mole] [2 Q.N. 1(b), 2063]

21. Calculate the mass of:
    (a) two atom of nitrogen
    [Ans: 4.64×10^{-23} g]
    (b) one molecule of hydrogen 2 [Q.N. 2, 2064]
    [Ans: 3.35×10^{-24} g]

22. An oxide of trivalent metal contains 32% of oxygen. Calculate the atomic mass of the metal. 2 [Q.N. 3, 2064]
    [Ans: 51]

23. One atom of an element 'A' weighs 6.644×10^{-23}g. Calculate the number of gram atom in 80 kg of it. 2 [Q.N. 1, 2065]
[Ans: 2000 gram atom]

24. Calculate the mass of 120 cc of nitrogen at NTP. How many number of molecules are present in it ? 2 [Q.N. 2, 2065]
[Ans: 0.15g, 3.23×10^{21} molecules]

25. What weight of Na will contain the same number of atoms as are present in 1.2 g of Carbon (C^{12}) ? 2 [Q.N. 1, 2067]
[Ans: 2.3 g]

26. 4 g of a divalent metal reacts with chlorine to produce 11.1 g of its metal chloride. Calculate the atomic mass of metal. 2 [Q.N. 3, 2067]
[Ans: 40 g]

27. 34.2 gram of Sucrose $C_{12}H_{22}O_{11}$ are dissolved in 180 gram of water. Calculate the number of oxygen atoms in the solution. 2 [Q.N.1,2068]
[Ans: 6.68242×10^{24} atoms]

28. What is the mass in gram of a molecule of carbon dioxide (CO_2)?
(Ans: 7.3×10^{-23}g) 1[Q.N. 1(a), Supp. 2068]

29. Calculate the mass in gram of 1×10^{22} molecules of $CuSO_4 . 5H_2O$
(at. wt. of Cu = 63)
(Ans: 4.13 g) 1[Q.N. 1(b), Supp. 2068]

30. Calculate the mass of: [1+1] [Q.N. 2, Set 'A' 2069]
i) one molecule of Nitrogen in gram
[Ans: 3.82×10^{23}g]
ii) a mole of carbondioxide
[Ans: 44g]

32. 73 g of conc. HCl was diluted by adding 144 g of water. How many gram atom of hydrogen are present in the dilute acid?
(Ans: 18 g atom) 2[Q.N. 1, Set 'B' 2069]

33. Calculate the amount of lime (CaO) that can be prepared by heating 200 kg of lime stone ($CaCO_3$) that is 95% pure. 2 [Q.N. 2, Set 'B' 2069]
(Ans: 106.4 kg)

34. Convert the followings.
a. 3 atom of nitrogen into gms.
Ans: 6.97×10^{-23} g
b. 4 g atom of carbon into number. [Q.N. 1, Supp. 2069]
Ans: 2.4088×10^{24}

35. 6 g of an element x combine with 16 g of another element y to give 0.5 mole of a compound xy. What is the molecular mass of xy?
[Ans: 44] [Q.N. 2, 2070 'C']

2.3 Empirical, Molecular Formula and Limiting Reactants

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. What is a limiting reactant ? Why is it essential in stoichiometric calculations ? [Q.N.2, 2060]

Short Questions

(All questions are of equal value, 5 marks each.)

(No any questions have been asked in this section upto now.)

Long Questions

(All questions are of equal value, 10 marks each.)

(No any questions have been asked in this section upto now.)

Numerical Problems

1. Give the equation: $4NH_3(g) + 5O_2 (g) \rightarrow 4NO (g) + 6H_2O (g)$
a) How many moles of NH_3 must react to produce 5.0 moles of NO?
[Ans: 5 moles]

b) How many moles of O_2 must react to produce 5.0 moles of NO?
[Ans: 6.25 moles]
c) How many litres of NH_3 and O_2 must react to produce 100 litres of NO?
[Ans: 100 litres of NH_3 and 125 litres of O_2]
d) How many litres of O_2 will react with 100 gms of NH_3?
[Ans: 164.70 litres]
e) How many litres of NO are formed by reacting 10 moles of NH_3 with 10 moles of O_2? 5 [Q.N.29, 2055]
[Ans: 179.2 litres]

2. Given, $CaCO_3(s) + 2HCl\ (aq) \rightarrow CaCl_2\ (aq) + H_2O\ (l) + CO_2\ (g)$.
If 10 gram of pure $CaCO_3$ are added in a solution containing 7.665 gram of HCl,
a) Find the limiting reactant.
[Ans: $CaCO_3$ is limiting reactant]
b) Calculate the number of moles excess reactant left over unreacted.
[Ans: 0.01 mol]
c) Calculate the volume of CO_2 gas produced at NTP.
[Ans: 2.24 litres]
d) Calculate the number of grams of NaOH required to absorb whole of the CO_2 gas as Na_2CO_3.
[Ans: 8 gram]
(at. mass of Ca=40, C=12, O=16, Cl=35.5, Na=23 and H=1) 5 [Q.N.33, 2056]

3. 5g of pure $CaCO_3$ if treated with 5g of HCl to produce $CaCl_2$, H_2O and CO_2:
(a) Find which one is limiting reactant and why ?
[Ans: $CaCO_3$ is limiting reactant]
(b) Calculate mass of $CaCl_2$ formed.
[Ans: 5.5 g]
(c) How many number of water molecules are produced ?
[Ans: 3.01×10^{22} molecules]
(d) Calculate the volume of CO_2 produced at NTP.
[Ans: 1.12 litres] 5 [Q.N.23, 2062]

4. 10 gram of impure zinc reacts with excess of dilute sulphuric acid to yield zinc sulphate and hydrogen. (Zn = 65, S = 32, O = 16)
a. Calculate the number of moles of H_2SO_4 consumed.
[Ans: 0.154 mol]
b. Calculate the mass of $ZnSO_4$ formed.
[Ans: 24.76 g]
c. What volume of hydrogen is evolved at NTP ?
[Ans: 3449.6 cc] 5 [Q.N. 24. 2063]

5. 10.6 g of pure Na_2CO_3 if treated with 7.9 g of HCl to produce NaCl, H_2O and CO_2.
(a) Find the limiting reagent and calculate mole of unreacted regent left over.
[Ans: Na_2CO_3 is the limiting reagent, 1.64×10^{-02} mole of unreacted reagent left over]
(b) What volume of CO_2 gas is produced at NTP ?
[Ans: 2.24 litres]
(c) Calculate mass of NaCl formed.
[Ans: 11.7 g] 5 [Q.N. 25, 2064]

6. What weight of 60% pure sulphuric acid is required to decompose 25 gram of chalk ($CaCO_3$) ? 2 [Q.N. 1, 2064]
[Ans: 40.83 g]

7. 200 g of 90% pure $CaCO_3$ is completely reacted with excess HCl to produce $CaCl_2$, H_2O and CO_2.
(i) Which one is limiting reagent ?
(ii) Calculate the mass of $CaCl_2$ formed.
(iii) How many moles of water are produced ?

(iv) What volumes of CO_2 are produced if the reaction is carried out at 27°C temperature and 760 mmHg pressure ? 5 [Q.N. 23, 2065]

[Ans: (i) $CaCO_3$ (ii) 199.8 g (iii) 1.8 mol ((iv) 44.33 litre]

8. What volume of CO_2 gas is produced when 20 g of 20% pure $CaCO_3$ is completely heated ? 2 [Q.N.13, 2066]

[Ans: 896 mL]

9. 17 g of ammonia is completely reacted with 45 g of oxygen to produce NO and H_2O.
   i) Which is limiting reagent?
   ii) Calculate the number of moles of unreacted reactent left over.
   iii) What volume of NO are produced at NTP?
   iv) Calculate the mass of water produced. 5 [Q.N.23, 2066]

[Ans: i) NH_3 ii) 0.156 mol iii) 22.4 L iv) 279 g]

10. How many gram atoms of sulphur and how many gram of oxygen are needed to prepare 6.023×10^{24}, molecules of SO_2 ? 2 [Q.N. 2, 2067]

[Ans: 10 gram atoms S and 320 gram of O_2]

11. (i) How much sulphuric acid containing 90% H_2SO_4 by weight is needed for the production of 1000kg of hydrochloric acid containing 42% HCl by weight in the following reaction ?

$$2\,NaCl\,(aq) + H_2SO_4\,(aq) \rightarrow Na_2SO_4(aq) + 2HCl\,(aq)$$

[Ans: 626.48 kg]

(ii) If the above reaction is carried out by mixing 11.7 g of pure NaCl and 10 g of pure H_2SO_4, find the limiting reactant. 5 [Q.N. 23, 2067]

[Ans: NaCl]

12. A chemical reaction was carried out by mixing 25 g of pure Calcium Carbonate and 0.75 mole of pure hydrochloric acid to give $CaCl_2$, H_2O and CO_2.
   i) Which one is limiting reactant and why ?
   [Ans: $CaCO_3$]
   ii) Calculate the mass of $CaCl_2$ produced.
   [Ans: 27.75 g]
   iii) How many number of water molecules are formed ?
   [Ans: 1.5055×10^{23}]
   iv) What mass of NaOH is required to absorb the whole CO_2 produced in the reaction ? 5 [Q.N.25,2068]
   [Ans: 20g]

13. 5 g of pure $CaCO_3$ if treated with 5 g of HCl to produce $CaCl_2$, H_2O and CO_2.
   a) Find which one is limiting reagent and why?
   (Ans: $CaCO_3$: it finishes first)
   b) Calculate mass of $CaCl_2$ formed.
   (Ans: 5.55 g)
   c) How many numbers of water molecules are produced?
   (Ans: 3.011×10^{22})
   d) Calculate the volume of CO_2 produced at NTP.
   [Ans: 1.12 L] 2+1+1+1 [Q.N. 25, Supp. 2068]

14. What is meant by limiting reactant? A chemical reaction was carried out by mixing 22 g of pure NaOH with 24.5 g of pure H_2SO_4 to produce Na_2SO_4 and water.
   a. Which one is limiting reactant?
   Ans: H_2SO_4
   b. Calculate the mass of sodium sulphate produced.
   Ans: 35.5 g
   c. How many moles of water are formed?
   Ans: 0.5 mol
   d. Find the no. of molecules of unreacted reactant left over.
   Ans: 3.011×10^{22} 1+1+1+1+1[Q.N. 25, Supp. 2069]

15. How many gram atoms of sulphur and how many grams of oxygen are needed to prepare 6.023×10^{24} molecules of SO_2?
[Ans: 10 gram atoms of sulphur, 320g of Oxygen] **[Q.N. 1, 2070 'C']**

16. For a reaction,
$Ca(OH)_2 (aq.) + 2NH_4Cl (aq.) \rightarrow CaCl_2 (aq.) + 2NH_3 (g) + 2H_2O (l)$
The reaction is carried out by mixing 7g of pure $Ca(OH)_2$ and 7g of pure NH_4Cl.
a) Find the limiting reactant.
b) Calculate the mole of unreacted reactant left over.
c) How many gram of $CaCl_2$ are formed?
d) What volumes of NH_3 gas are produced at 27°C and 1.5 atmospheric pressure? **[Q.N. 23, 2070 'D']**
[Ans: (a) NH_4Cl (b) 0.03 mol (c) 7.26 g (d) 2.13 L]

17. 20g of 40% pure $CaCO_3$ if reacted with 5g of HCl to produce $CaCl_2$, H_2O and CO_2. 1+1+1+2**[Q.N. 28, 2070 'C']**
a. Find which one is limiting reactant and why?
[Ans: HCl]
b. Calculate mass of $CaCl_2$ formed.
[Ans: 7.6 g]
c. How many number of water molecules are produced?
[Ans: 4.12×10^{22}]
d. Calculate the volume of CO_2 produced at 27°C and 0.5 atms pressure.
[Ans: 3.3723L]

2.4 Avogadro's Hypothesis and Its Applications

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. State Avogadro's Hypothesis. **[Q.N.10, 2054]**
2. Define vapour density. How is it related to molecular mass? **[Q.N.12, 2059]**
3. How did the law of multiple proportions encourage Dalton to introduce an atomic theory ? **[Q.N.3, 2060]**
4. One-volume of hydrogen reacts with one-volume of chlorine to give two-volumes of hydrogen chloride gas. Which law of stoichiometry is illustrated ? State the law. **[Q.N. 12, 2064]**
5. State Avogadro's hypothesis. **[Q.N. 9, 2065]**
6. In what way has Avogadro's hypothesis given support to Dalton's atomic theory ? **[Q.N. 4, 2067]**

Short Questions

(All questions are of equal value, 5 marks each.)

1. State Avogadro's hypothesis. Show that molecular wt = 2 × vapour density. **[Q.N.23, 2057]**
2. State Avogadro's Law. Apply the law to deduce the relationship between the molecular mass and the vapour density. **[Q.N.23, 2061]**

Long Questions

(All questions are of equal value, 10 marks each.)

1. State and explain Avogadro's hypothesis. How this theory can be used to determine the molecular weight of a gas? **[Q.N.31 (a), 2054]**
2. Avogadro's hypothesis and its relation with molecular mass of volatile substance. **[Q.N.33(b), Set 'A' 2069]**

Numerical Problems

1. Calculate the weight of 11.2 litre of CO_2 gas at STP. (Mol. Wt. of CO_2 = 44).
[Ans: 22 g] 2 **[Group A, Q.N.7, 2053]**
2. 1 litre of hydrogen at STP weighs 0.09 g. If 2 litres of a gas at STP weighs 2.880 g, calculate the vapour density and the molecular weight of the gas.
[Ans: VD = 16, mol. wt = 32] 5 **[Q.N.24, 2055]**

3. 16 g of a gas at STP occupies 5.6 L What is the molecular mass of the gas ?
[Ans: Mol mass = 64] 2 [Q.N.5, 2057]
4. Calculate the weight, in gram, of 5.60 litre of chlorine gas (Cl_2) at NTP?
[Ans: 17.75 g] 2 [Q.N.2, 2058]
5. What volume would 5.5 g CO_2 occupy at STP? 2 [Q.N.11, 2059]
[Ans: 2.8 litre]
6. Calculate the volume of 11 g of CO_2 at NTP. 2 [Q.N. 3. 2063]
[Ans: 5.6 litre]
7. An oxide of nitrogen contains its half-volume of nitrogen and its vapour density is 15. Determine its molecular formula. 2 [Q.N. 11, 2064]
[Ans: Molecular Formula : NO]
8. What volume of CO_2 will be delivered at NTP to extinguish fire from a Cylinder of 10 liter Capacity containing 5kg of CO_2 gas.
[Ans: 25354] 2 [Q.N.2, 2066]
9. Calculate the wt. of 11.2 liter of CO_2 gas at STP (Mol.wt. of CO_2 = 44).
[Ans: 22] 2[Q.N. 4, Set 'A' 2069]
10. Define Avogadro's hypothesis. How is this hypothesis applied to show that molecular mass of volatile substance is twice of its vapour density?
A oxide of nitrogen contains same of its own volume of nitrogen and its vapour density is 54. Determine its molecular formula
(Ans: N_2O_5) 1+2+2 [Q.N. 23, Set 'B' 2069]

2.5 Equivalent Mass

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. Equivalent weight of an element is 32.5. What does it mean ? [Q.N.4, 2057]
2. Why is hydrogen displacement method not applicable to determine the equivalent mass of copper? [Q.N.5, 2059]

Numerical Problems

1. Define equivalent weight of an element. A divalent metal has atomic weight 24. What is its equivalent weight ? 2 [Q.N. 3, 2065]
[Ans: 12]
2. Calculate the equivalent wt of following underlined elements :
(i) <u>C</u>Cl_4 (ii) <u>Fe</u>$_2O_3$
[Atomic weight of Carbon = 12, Atomic weight of Oxygen = 55.8]
[Ans: (i) 3, (ii) 18.6] 2 [Q.N.3, 2066]
3. Calculate the equivalent weight of underlined element. 2[Q.N. 9, Set 'A' 2069]
i) <u>C</u>Cl_4 ii) <u>Mg</u>O iii) <u>Fe</u>$_2O_3$ iv) <u>Al</u>Cl_3
[Ans: (i) 3, (ii) 12, (iii) 18.6, (iv) 9]

Unit 3 – State of Matter

3.1 Gaseous State

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. At what condition the value of P × V is always constant ?
[Group A, Q.N.1, 2051]
2. Draw volume (V) and temperature (t°C) relationship of gases at constant pressure. Indicate the temperature at which the volume occupied by the gas becomes zero. Name that temperature. **[Group A, Q.N.7, 2052]**

3. Sketch the diagram for the variation of volume of a given mass of ideal gas with temperatures at constant pressure. Indicate the absolute zero in the diagram. **[Q.N.9, 2054]**
4. What do you mean by ideal gas and real gas? **[Q.N.15, 2054]**
5. What is an ideal gas? Under what conditions will a gas behave nearly like an ideal gas? **[Q.N.7, 2058]**
6. State Dalton's law of partial pressure. **[Q.N.3, 2059]**
7. What is universal gas constant? **[Q.N.8, 2059]**

Short Questions

(All questions are of equal value, 5 marks each.)

1. What are the main points of kinetic theory of gases ? **[Group B, Q.N.3, 2051]**
2. Give the postulates of Kinetic Molecular Model of gas. **[Q.N.23, 2055]**

Long Questions

(All questions are of equal value, 10 marks each.)

1. State and explain Boyle's law. **[Group C, Q.N.3 (a), 2052]**
2. State and explain Dalton's law of partial pressure. **[Group C, Q.N.2 (a), 2053]**
3. Write short notes on:
   (a) Postulates of kinetic theory of gas **[Q.N.33 (d), 2060]**
4. Write the postulates of kinetic theory of gas. **[Q.N. 30,(a) 2063]**
5. State and explain Charle's law. How is Charle's law explained qualitatively in the light of kinetic theory of gas. **[Q.N. 30,(b) 2063]**
6. State and explain Graham's law of diffusion of gases. **[Q.N. 30(a), 2064]**
7. Derive the relation PV = nRT. What is meant by Absolute scale of temperature and Absolute zero? **[Q.N. 30(a), Supp. 2068]**
8. Explain, how Charle's law gave the concept of absolute scale of temperature. Derive the relation PV=nRT. A hydrocarbon CxHy has mass ratio between hydrogen and carbon 1:10.5. One litre of the hydrocarbon at 127°C and 1 atm pressure weights 2.8 g, find the molecular formula of the hydrocarbon. **[Q.N. 32, 2070 'D']**

Numerical Problems

1. A carbon dioxide fire-extinguisher of 3 litre capacity contains 4.4 kg of carbon dioxide. What volume of gas could this extinguisher deliver at NTP ?
   [Ans: 2237 litre of CO_2] 5 **[Group C, Q.N.3 (b), 2052]**
2. A 1.00 L sample of dry gas at 25°C has the following compositions:
   0.8940 g of N_2, 0.2741 g of O_2, 0.0152 g of Ar;
   0.00107 g of CO_2 Given R = 0.0821 1 atm $mol^{-1}K^{-1}$
   What are the partial pressures for each components gas in the mixture? What is the total pressure? 5 **[Group C, Q.N.2 (b), 2053]**
   [Ans: partial pressure of (i) N_2 = 0.78 atm, (ii) O_2 = 0.2079 atm, (iii) Ar = 0.0092 atm, (iv) CO_2 = 0.00058 atm, total pressure = 0.99768 atm]
3. The volume of carbon monoxide gas collected over water at 25°C is 680 cc with a total pressure of 752 mmHg. The vapour pressure of water at 25°C is 23.8 mm Hg. Determine the partial pressure of CO in container.
   [Ans: 728.2 mmHg] 2 **[Group A, Q.N.8, 2053]**
4. A balloon can hold 1000 cc of air before bursting. The balloon can hold 975 cc of air at 5°C. Will it burst when it is taken into a house at 25°C? Assume that the pressure of the gas in the balloon remains constant. 5 **[Q.N.31 (b), 2054]**
   [Ans: balloon will burst]
5. Two grams of hydrogen diffuses from a container in 10 minutes. How many grams of oxygen would diffuse through the same container in the same time under similar conditions? 2 **[Q.N.9, 2055]**
   [Ans: 8 g]

6. One mole of a gas occupies a volume of 1 litre at 27°C. What will be the pressure of the gas? 2 [Q.N.10, 2055]
[Ans: 24.63 atm]

7. State Graham's law of diffusion. How long will it take 600 mL of H_2 gas to diffuse through a porous partition, if 300 mL of O_2 diffuse through it in 10 minute under identical conditions?
[Ans: 5 minute] 5 [Q.N.25, 2056]

8. A gas x diffuses five times as rapidly as another gas y. Calculate the ratio of molecular mass of x and y. 2 [Q.N.3, 2057]
[Ans: x : y :: 1.25]

9. State Boyle's law and Charle's law. Derive the relation PV = nRT. A gas cylinder containing cooking gas can withstand up to pressure 14.9 atm. The pressure gauge of cylinder indicates 12 atm at 27°C. Due to sudden fire in the building its temperature starts rising. At what temperature will the cylinder explode ?
[Ans: above 99.5°C] 10 [Q.N.30, 2057]

10. One litre of a gas at 0°C is heated to 100°C keeping pressure constant. What will be the new volume at 100°C? 2 [Q.N.6, 2058]
[Ans: 1.366 litre]

11. Outline the basic assumptions of kinetic model of gas. What are relative diffusion rates of methane (CH_4) and sulphurdeoxide (SO_2)?

$$\left[\textbf{Ans: } \frac{\text{rate }(CH_4)}{\text{rate }(SO_2)} = 2:1\right]$$

If these two gases are simultaneously introduced into opposite ends of 100 cm tube and allowed to diffuse toward each other, at what distance from the SO_2 end will the molecules of two gases meet? 10 [Q.N.30, 2058]
[Ans: At a distance of 33.33 cm from SO_2 end]

12. State Graham's law of diffusion. A vessel of volume 100 mL contains 10% O_2 and 90% unknown gas. The gases diffuse in 86 s through a small hole of the vessel. If pure oxygen under the same condition diffuses in 75 s, find the molecular mass of the unknown gas.
[Ans: Molecular mass : 43.16] 5 [Q.N.24, 2059]

13. The rate of diffusion of a saturated hydrocarbon (C_nH_{2n+2}) gas is 1.206 times that of SO_2 gas under identical conditions. Find the molecular mass and the value of 'n' for the gas. (Mol. mass of SO_2 = 64) 2 [Q.N.4, 2060]
[Ans: Molecular mass = 44, n = 3]

14. State Grahm's Law of diffusion. How long will it take 500 mL of hydrogen gas to diffuse through a partition if 250 mL of oxygen diffuse in 50 minutes under similar conditions ?
[Ans: 25 minute] 5 [Q.N.24, 2061]

15. Calculate the mass of oxygen gas whose volume is 320mL at 17°C and 2 atmospheric pressure. 2 [Q.N.1, 2062]
[Ans: 0.857 g]

16. State Boyle's Law and Charls Law. Derive PV = nRT. 0.50g of a volatile liquid was introduced into a globe of 1000mL capacity. The globe was heated to 91°C, so that all the liquid vapourised exerted a pressure of 190 mmHg. Calculate the molecular mass of the liquid. (R = 0.082 L atm $mol^{-1}K^{-1}$)
[Ans: Molecular mass = 59.69] 10 [Q.N.30, 2062]

17. A mixture of ozone and oxygen containing 20% by volume of ozone diffused through a porous plug in 172 second, while the same volume of pure oxygen took 164 second to diffuse through the same plug. Calculate the relative density of ozone.
[Ans: relative density of O_3 = 23.95] 5 [Q.N. 23. 2063]

18. 5 gram of hydrogen diffused through a porous membrane in 30 minute. Find the time required to diffuse the same amount of SO_2 gas at identical conditions. 5 **[Q.N. 30(b), 2064]**
[Ans: Time required = 5.29 minute]

19. What are the basic postulates of kinetic theory of gas ? Why do gases not show ideal behaviour at low temperature and high pressure ? An evacuated glass vessel weighs 50 g when empty, 148 g when filled with a liquid of density of 0.98 g/cc, and 50.5 g when filled with an ideal gas at 760 mmHg and at 300K. Determine the molecular mass of the gas.
[Ans: 123.15] 10 **[Q.N. 30, 2065]**

20. State and explain the Graham's Law of diffusion. What is the main application of Graham's Law? A Flask of 0.3 liter capacity was weighed after it had been evacuated. It was then filled with a gas of unknown molecular mass at 1.0 atm pressure and temperature of 300K. The increase in mass of the flask was 0.977g. Calculate the molecular mass of the gas.
[Ans: 80.1] 10 **[Q.N.30, 2066]**

21. State and explain Dalton's Law of partial pressure. What is the main application of this Law ? A vessel Contains 12 g of an ideal gas at t°C temperature and 1 atm pressure. When the temerature is increased by 10°C at the same volume, the pressure increases by 10%. Calculate the volume and initial temperature. 10 **[Q.N. 30, 2067]**
[Ans: 283 K]

22. Derive the relation 'PV=nRT'. Under what condition does a gas follow the above relation ? How would you define universal gas constant 'R' ? How much increase in temperature is necessary to increase volume of half litre of the gas by 40% at 25°C, keeping the pressure constant ? 10 **[Q.N.30,2068]**
[Ans: 119.2°]

23. A baloon can hold 1000 cc of air before bursting. The baloon can hold 975cc of air at 5°C, will it burst when it is taken into a house at 25°C? Assume that pressure of the gas in the balloon remains constant.
[Ans: 1045.15 L, the ballon will brust) 4+2+4 **[Q.N. 30(b), Supp. 2068]**

24. State Boyle's law and Charle's law. Derive the relation PV = nRT. Two moles of ammonia are enclosed in a five litre flask at 27°C. Calculate the pressure exerted by the gas assuming that the gas behaves like an ideal gas.
[Ans: 9.85 atom] 8+2 **[Q.N. 32, Set 'A' 2069]**

25. State Boyle's law and draw sketch graphs of:
i) P against V ii) P against I/V iii) PV against P iv) PV against V.
For a perfect gas at constant temperature. The mass of 525cc of a gaseous compound at 28°C and 730 mm Hg pressure was found to be 0.9g. What will be the volume of 2g of the gas at 30°C and 760 mmHg pressure?
[Given; R = 0.0821 L atm $K^{-1}mol^{-1}$] 2+4+4 **[Q.N. 31, Set 'B' 2069]**
(Ans: 1.13 L)

26. State Boyle's Law and derive the relation $P_1V_1=P_2V_2$, where P is pressure, V is volume of an ideal gas. Draw sketch graphs of
a. P against 1/V
b. P against V
c. PV against P.
10 gm of sample of a gas is introduced into 90 litre vessel at NTP. If the pressure is kept constant and the temperature of the gas is raised to 200°C, how many gram of the gas will escape out the vessel?
[Ans: 7.3 g] 1+2+1+1+1+4**[Q.N. 30, Supp. 2069]**

27. Derive PV=nRT. How did Charle's law give the concept of absolute scale of temperature? Two vessel of capacity 1.5 litre and 2 litres contain hydrogen gas and oxygen gas respectively under a pressure of 750mm and 100mm. The gases are mixed together in a 5 litre vessel. What will be the final pressure of mixture?
[Ans: 265 mmHg] 4+2+4 **[Q.N. 31, 2070 'C']**

3.2 Liquid State

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. Name the physical property behind rise of liquids in capillary tube. **[Group A, Q.N.1, 2052]**
2. Define aqueous tension. Why is it subtracted from the total pressure to determine the pressure of a dry gas ? **[Group A, Q.N.8, 2052]**
3. What is evaporation ? How does it differ from boiling ? **[Group A, Q.N.18, 2052]**
4. Define coefficient of viscosity. How coefficient of viscosity is related with viscous force ? **[Group A, Q.N.22, 2052]**
5. Why is Glycerine more viscous than water? **[Group A, Q.N.4, 2053]**
6. Why is a mercury droplet spherical? **[Group A, Q.N.5, 2053]**
7. In terms of vapour pressure, what do you mean by a boiling point of a liquid? **[Q.N.1, 2054]**
8. The meniscus for mercury in a glass tube is concave downward. Explain. **[Q.N.4, 2054]**
9. What do you understand by viscosity? **[Q.N.13, 2054]**
10. What is the effect of temperature on: **[Q.N.18, 2054]**
    a) Surface tension b) Viscosity c) Vapour pressure of liquid.
11. What do you understand by the term 'Surface tension'? **[Q.N.11, 2055]**
12. How is surface tension of a liquid originated? **[Q.N.5, 2056]**
13. What is meant by viscosity? **[Q.N.10, 2058]**
14. Define surface tension. **[Q.N.2, 2059]**
15. Define Vant Hoff's factor. What for is it used ? **[Q.N.5, 2060]**
16. What do you mean by boiling point and evaporation ? **[Q.N.3, 2061]**
17. What is meant by viscosity ? **[Q.N.5, 2061]**
18. Give reason:
    (a) Falling liquid drops are spherical
    (b) Evaporation takes place from the surface of liquid **[Q.N.3, 2062]**
19. What happens to the vapour pressure of a solvent, when non volatile solute particles are dissolved in it ? **[Q.N. 4. 2063]**
20. Define the term coefficient of viscosity. **[Q.N. 9. 2063]**
21. Why does boiling point of liquid rise on increasing pressure ? **[Q.N. 4, 2064]**
22. What is Surface tension ? Mention any one physical properties of liquid due to surface tension. **[Q.N. 4, 2065]**
23. How is boiling of Liquid different from Evaporation? **[Q.N. 5, 2066]**
24. Give reason :
    (i) It is more efficient to wash clothes in hot water than cold water. 1
    (ii) Evaporation takes place from the surface of the liquid. **[Q.N. 6 , 2067]**
25. Define Surface tension. Write its unit. **[Q.N.3,2068]**
26. What is the effect of pressure on boiling point and viscosity of a liquid? **[Q.N. 3, Supp. 2068]**
27. State the physical principle behind the following phenomenon:
    (i) Rain drops are Spherical.
    (ii) A drop of ether on your skin disappears fast and the skin feels cool. **[Q.N. 3, Set 'B' 2069]**
28. Why does the boiling point of a liquid rise when the pressure is increased? **[Q.N. 3, Supp. 2069]**
29. Define surface tension. Write its unit. **[Q.N. 3, 2070 'C']**
30. Write any two physical properties of liquid caused by surface tension. **[Q.N. 3, 2070 'D']**

Long Questions

(All questions are of equal value, 10 marks each.)

1. What are solubility curves ? What information is obtained from solubility curve ? **[Group C, Q.No. 4(a), 2052]**
2. Write short notes on:
   (a) Solubility curve and its applications **[Q.N.32 (c), 2056]**

Numerical Problems

1. The solubility of salt at 0°C is 12. How much salt will 50 g of its saturated solution contain at that temperature ? 5 [Group C, Q.N.4(b), 2052]
[Ans: 5.35 g]
2. What do you mean by solubility ? Represent different types of solubility curves and give its applications. Calculate the weight of crystal formed on cooling 80g of saturated solution from 60°C to 30°C. Solubility of salt at 60°C and 30°C are 132 and 95 respectively. 10 [Q.N.30, 2054]
[Ans: 12.759 g]
3. Calculate the molality of 4.9% H_2SO_4. (molecular mass of H_2SO_4 = 98)
[Ans: Molality = 0.525] 2 [Q.N.6, 2057]
4. 2.65 g anhydrous sodium carbonate is dissolved in water and prepared exactly 100 cc solution. Calculate the molarity of this solution. (Molecular mass of anhydrous sodium carbonate = 106). 2 [Q.N.1, 2059]
[Ans: Molarity = 0.25 M]
5. Define solubility of a salt. The solubility of salt in water at 75°C is 155. When 80 g. of its saturated solution at 75°C was cooled to 15°C, 40 g of the salt was precipitated. Find the solubility of the salt at 15°C.
[Ans: solubility at 15°C = 27.46%] 5 [Q.N.24, 2060]

3.3 Solid State

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. Explain why sodium chloride does not conduct electricity in solid state but a good conductor when molten. [Q.N. 10, 2065]
2. Distinguish between crystalline and amorphous solid. [Q.N.10, 2066]
3. Distinguish between Crystal Lattice and unit cell. [Q.N.4,2068]
4. Differentiate between crystalline and Amorphous solid. [Q.N. 4, Supp. 2068]
5. Differentiate between Crystalline and Amorphous solid. [Q.N. 3, Set 'A' 2069]
6. Write an example of each of the followings:
   (i) Crystalline solid
   (ii) Hygroscopic substance
   (iii) Water of crystallization
   (iv) Isotropic substance. [Q.N. 4, Set 'B' 2069]
7. Mention one important character and an example of each of deliquescence and efflorescence. [Q.N. 5, Supp. 2069]
8. Give an example of each of the following:
   a. Efflorescent substance
   b. Isotropic substance
   c. Anisotropic substance
   d. Hygroscopic substance. [Q.N. 4, 2070 'C']
9. Distinguish isotropic and anisotropic substance with one example of each. [Q.N. 4, 2070 'D']

Unit 4 – Atomic Structure

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. Give the electronic configuration of copper (At no. 29) in terms of s, p, d, f orbitals. [Group A, Q.N.5, 2051]
2. Write the shapes of s and p orbitals. [Group A, Q.N.2, 2052]
3. Give the electronic configuration of silver (At no. 47) in terms of s, p and d orbitals. [Group A, Q.N.3, 2052]

4. A scientist investigating the electron structure of the element concluded that the K, L and M shells were all full and that the N shell contained four electrons. What is the atomic number of that element ? **[Group A, Q.N.9, 2052]**
5. For n= 4, write all possible values of l and m. **[Group A, Q.N.19, 2052]**
6. Discuss how Bohrs was able to predict the line spectra of a hydrogen atom. **[Group B, Q.N.2, 2052]**
7. Write the electron configuration of copper (atomic number 29) in terms of s.p.d. orbitals. **[Group A, Q.N.9, 2053]**
8. Write the electronic configuration of chromium (At. no. 24) in terms of s.p.d. orbitals. **[Q.N.14, 2054]**
9. An atomic orbital has n = 3, what are the possible values of l and m? **[Q.N.19, 2054]**
10. What is an atomic orbital? What are shapes of a s orbital and p orbital? **[Q.N.1, 2055]**
11. Write the electronic configuration of chromium (At no 24) and copper (At. no 29) **[Q.N.7, 2057]**
12. What are values for n, l and m for $2P_x$ orbital ? **[Q.N.11, 2057]**
13. An electron of an atom possesses the quantum numbers n = 2, ℓ = 0 and m = 0. What do they mean ? **[Q.N.6, 2060]**
14. Write the electronic configuration of elements with the atomic number 19 and 24. Give the name of these elements. **[Q.N.12, 2061]**
15. Write the atomic number of elements whose outermost electronic configuration are represented by (a) $3s^1$ (b) $3p^6$ **[Q.N.9, 2062]**
16. Write the ground state electronic configuration of Cu (Z = 29) and Cr (Z = 24) in terms of s, p and d orbitals. **[Q.N. 5. 2063]**
17. What observation did Rutherford led to conclude that the nucleus of atom is very small but heavy mass ? **[Q.N. 5, 2064]**
18. Why is it that electron does not jump into the nucleus ? **[Q.N. 5, 2065]**
19. What is meant by atomic spectrum ? **[Q.N. 6, 2065]**
20. What are the values of Principal quantum number (n) and azimuthal quantum number (l) for the following orbitals :
    (i) 3S (ii) 4p **[Q.N.6, 2066]**
21. An atom of an element has 24 electrons, what is the total number of s electrons? **[Q.N.7, 2066]**
22. Write one important property of the compound formed by the two atoms x and y whose valence shell electronie configurations are $3s^1$ and $3s^2 3p^5$ respectively. **[Q.N.9, 2066]**
23. Write down all four quantum number for outermost electron of sodium atom. (Z =11) **[Q.N. 7, 2067]**
24. Give the values of all four quantum number of 11th electron of Magnesium (At.no. = 12) **2 [Q.N.5,2068]**
    [Ans: $n = 3, \ell = 0, m = 0, s = +\frac{1}{2}$]
25. What observations did Rutherford make the following conclusions ?
    (i) The atomic center is positively charged.
    (ii) Most of the space inside the atom is hollow. **[Q.N.6,2068]**
26. How many maximum number of electrons that may be present in principle quantum number 3 and azimuthal quantum number 2? **[Q.N. 5, Supp. 2068]**
27. What is Pauli exclusion principle? **[Q.N. 6, Supp. 2068]**
28. An element has 2 electrons in 'K' shell, 8 electrons in 'L' shell and 9 electrons in 'M' shell. Write its electronic configuration and calculate the total numbers of p–electrons. **2 [Q.N. 5, Set 'B' 2069]**
    (Ans: 12 electrons)
29. Write the electronic configuration of the element with atomic number 18 and 26. **[Q.N. 5, Set 'A' 2069]**

30. Name the spectral series which appears visual part of the electromagnetic spectrum. How is such series originated? [Q.N. 6, Set 'A' 2069]
31. How are Balmer Series and Paschen Series orginated in hydrogen spectra? [Q.N. 6, Set 'B' 2069]
32. Write the electric configuration Cr^{++} and O^{--}.
(Atomic number of Cr = 24 and O= 8) [Q.N. 6, Supp. 2069]
33. What is the maximum number of electrons that may be present in all the atomic orbitals with principal quantum number (n = 4) and azimuthal quantum number (1 = 3) 2[Q.N. 7, Supp. 2069]
Ans: 14 electrons
34. What is Hund's rule? 2 [Q.N. 8, 2070 'C']
35. An atom 'A' has atomic number (Z = 29). Calculate the total number of s-electrons of A^{++}. 2[Q.N. 5, 2070 'D']
36. State Pauli-exclusion principle. 2[Q.N. 6, 2070 'D']

Short Questions

(All questions are of equal value, 5 marks each.)

1. Describe Bohr's model of the atom. Draw a picture lebelling pertinent parts. [Group B, Q.N.1, 2051]
2. Define Aufbau Principle. An atom has 2 electrons in first (K) shell, 8 electrons in second (L) shell and 2 electrons in third (M) shell. If so, find out the following :
(i) Electronic configuration of the atom.
(ii) Total number of principle quantum numbers
(iii) Total number of sub-shells
(iv) Total number of s-electrons [Q.N. 24, 2065]
3. How does Bohr's theory explain the Origin of hydrogen spectra? Name the different spectral lines with a labelled diagram. [Q.N.24, 2066]
4. Write down the essential postulates of Bohr's atomic model. How did it overcome the Limitation of Rutherford's atomic model ? [Q.N. 25, 2067]
5. X, Y and Z represent elements of atomic number 1, 6 and 17 respectively.
a) Write the electron structure of X, Y and Z.
b) Place the elements in the appropriate group of the periodic table.
c) Write the formula and the Lewis structures of the covalent compounds formed between: i) X and Y ii) X and Z [Group B, Q.N.3, 2053]
6. What experimental evidence led Rutherford to conclude that (a) the nucleus of the atom contains most of the atomic mass? (b) the nucleus of the atom is positively charged? (c) the atom consists of mostly empty space. [Q.N.21, 2055]
7. State Pauli exclusion principle and Hund's rule. Write the ground state electronic configuration of an atom having atomic mass number 37 and number of neutrons 20. [Q.N.23, 2056]
8. How does Bohr's theory predict the origin of line spectra of hydrogen atom? [Q.N.25, 2059]
9. Write down main postulates of Bohr's atomic model. [Q.N.24, 2062]
10. How does Bohr's atomic theory explain the origin of hydrogen spectra ? [Q.N. 25. 2063]
11. Explain hydrogen spectra in light on Bohr's theory. Why does hydrogen gas show large number of line spectra though H–atom contains one electron ? [Q.N. 23, 2064]
12. Write down the essential postulates of Bohr's atomic model. How did it overcome the Limitation of Rutherford's atomic model ? [Q.N. 25, 2067]
13. How does Bohr's theory explain the origin of hydrogen spectra ? Name the various Spectral Series observed in the atomic spectrum of hydrogen with a well lebelled diagram. [Q.N.23,2068]
14. What are the essential components of Bohrs' atomic model? How does Bohrs' theory predict the origin of line spectra of hydrogen atom? [Q.N. 26, Supp. 2068]
15. State and explain Hund's rule of maximum multiplicity. [Q.N. 25, Set 'A' 2069]

16. What are the conclusions made by Rutherford's α-ray scattering experiment about the structure of atom. Point out its drawbacks. [Q.N. 24, Set 'B' 2069]
17. What are the conclusions of Rutherford's α–particles scattering experiment about the structure of the atom? What is the major drawback of this model? [Q.N. 23, Supp. 2069]
18. How does Bohr's theory explain the origin of hydrogen spectra? Name the different spectral lines with labelled diagram. [Q.N. 23, 2070 'C']
19. What are the conclusions made by Rutherford from his α–ray scattering experiment about the structure of atom? Mention its limitation. [Q.N. 25, 2070 'D']

Long Questions

(All questions are of equal value, 10 marks each.)

1. Write short notes on:
   (a) Quantum numbers [Q.N.31 (a), 2055]
   (b) Bohr's model and explanation of hydrogen spectrum [Q.N.33 (b), 2058]
   (c) Bohr's model of atom and explanation of hydrogen spectrum. [Q.N.33 (a), 2061]
2. (i) Discuss how Rutherfords' nuclear model of atom is introduced on the basis of alpha particle scattering experiment. Point out the limitation of the model.
   (ii) How is the nuclear model of atom improved by Bohr ?
   (iii) Why is Bohr's atomic model appeared to be defective in the light of Heisenberg's uncertainty principle ? [Q.N.31, 2060]

Unit 5 – Nuclear Chemistry

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. What are radioisotopes ? State one use of such isotopes. [Q.N.7,2068]
2. Write two applications of radio-isotopes. [Q.N. 9, supp. 2068]
3. Write any two applications of each of the following isotope: ^{60}Co and ^{131}I. [Q.N. 7(Or), Set 'A' 2069]
4. Define Nuclear fusion and give an example of it. [Q.N. 7, Set 'B' 2069]
5. Give any two differences between Nuclear reactions and chemical reactions. 1 [Q.N. 9, Supp. 2069]
6. Define Nuclear fission reaction and write an example of it. [Q.N. 11, 2070 'C']
7. What is meant by nuclear reaction? Give an example of it. 1+1[Q.N. 7, 2070 'D']

Unit 6 – Electronic Theory of Valency and Bonding

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. What is Lewis base ? Give one example. [Group A, Q.N.2, 2051]
2. Explain why CO_2 got linear structure while H_2O got angular structure. [Group A, Q.N.6, 2052]
3. Distinguish between a covalent and a coordinate covalent bond. [Group A, Q.N.3, 2051]
4. Write the lewis structure of CCl_4. [Group A, Q.N.13, 2052]
5. Define Lewis base. Give an example of Lewis base. [Group A, Q.N.16, 2052]
6. How does a covalent bond differ from an ionic bond ? [Group A, Q.N.25, 2052]
7. Write Lewis electron dot formula for carbon dioxide. [Group A, Q.N.3, 2053]
8. What types of bonds are involved in oxygen and calcium fluoride molecules? [Group A, Q.N.11, 2053]
9. Write the Lewis structure of SO_4^{2-}. [Group A, Q.N.12, 2053]

10. What is octet rule? [Q.N.5, 2054]
11. Write the Lewis structure of orthophosphoric acid and phosphate ions. [Q.N.11, 2054]
12. Write the Lewis structure of SO_2 molecule. [Q.N.17, 2054]
13. Distinguish between a covalent and coordinate covalent bond. [Q.N.3, 2055]
14. Draw Lewis structure of N_2O_5. [Q.N.16, 2055]
15. Write the Lewis structure of NH_4Cl molecule. [Q.N.4, 2056]
16. Give an example of intermolecular hydrogen bond. How is it originated? [Q.N.9, 2056]
17. Give Lewis structure of potassium carbonate. [Q.N.8, 2057]
18. Write the Lewis dot structure of BF_3 molecule and justify the formation of coordinate covalent compounds by BF_3. [Q.N.8, 2058]
19. Why is solid sodium chloride a non conductor of electricity? [Q.N.4, 2059]
20. Why are metals malleable and ductile ? [Q.N.7, 2060]
21. Why are solid sodium chloride and diamond non-conductor of electricity ? [Q.N.6, 2061]
22. Write Lewi's structure of NO_2 and N_2O_3. [Q.N.16, 2061]
23. Write the Lewis Structure of:
    (a) H_2O_2 (b) HNO_3 [Q.N.7, 2062]
24. What is meant by metallic bond ? [Q.N.11, 2062]
25. Write the Lewis structure for NH_4^+ and H_2SO_4. [Q.N. 8, 2063]
26. Write the Lewis structure of the compound of formed by two elements A and B whose atomic numbers are 12 and 17 respectively. [Q.N. 8, 2065]
27. What is dipole moment ? Mention its one important application. [Q.N.11, 2066]
28. How would you explain metallic bond in light of electron-sea model ? [Q.N. 10, 2067]
29. Write down the Lewis structure of :
    (i) H_3BO_3 (ii) NO_2^- [Q.N. 13, 2067]
30. Define dipole moment. What is its unit ? [Q.N.10, 2060]
31. Define hydrogen bond. Give an example of intermolecular hydrogen bond. [Q.N.7, 2061]
32. Write an example of intermolecular and intramolecular hydrogen bondings. [Q.N.4, 2062]
33. Write resonance structures of ozone. [Q.N.12, 2062]
34. What is meant by hydrogen bonding ? Give an example of intermolecular hydrogen bonding. [Q.N. 11. 2063]
35. Each carbon-oxygen bond in CO_2 is polar but CO_2 molecule is non-polar. Give reason. [Q.N. 7, 2064]
36. What is hydrogen bond ? Write an example. [Q.N. 11, 2065]
37. What is Bohr's - Bury rule ? [Q.N. 8, 2067]
38. Explain why :
    (i) HCl has polar character though it has covalent bond.
    (ii) CO_2 is a linear molecule but H_2O is not. [Q.N. 11, 2067]
39. Define Octet rule. Name the two compounds in which Octet rule is not obeyed. [Q.N.7(Or),2068]
40. Write down the Lewis structure of
    i) NH_4NO_3 ii) H_2O_2 [Q.N.9,2068]
41. Give reason:
    i) Ammonia has higher boiling point than Phosphine.
    ii) CO_2 molecule gets linear structure. [Q.N.10,2068]
42. Write Lewis structure of N_2O_5 and N^+H4 1+1[Q.N. 8, Supp. 2068]
43. Give two important properties of electrovalent compound. [Q.N. 9(Or), Supp. 2068]
44. Define dipole moment. Mention one important application of dipole moment. 1+1 [Q.N. 10, Supp. 2068]

45. What type of hydrogen bond is found in ammonia? Give any one important use of hydrogen bond. [Q.N. 7, Set 'A' 2069]

46. Define dipole moment. What is its unit? [Q.N. 8, Set 'A' 2069]

47. Write down the Lewis structure of:

    (i) KNO_3 (ii) HCO_3^- [Q.N. 8, Set 'B' 2069]

48. What is hydrogen bond? Given an example of intermolecular hydrogen bond. [Q.N. 9, Set 'B' 2069]

49. Two elements have the following Lewi's symbols.

    $\overset{\circ}{\mathrm{X}}$ $\overset{\circ}{\underset{\circ}{{}^{\circ}_{\circ}\mathrm{Y}^{\circ}}}$

    Write the Lewis structure of the Covalent compound formed by x and y, and give its one important property. [Q.N. 4, Supp. 2069]

50. Define hydrogen bond and give an example of intra–molecular hydrogen bond. [Q.N. 9(Or), Supp. 2069]

51. What is meant by polar Covalent bond? Give an example of a polar and a non-polar covalent bond. [Q.N. 11, Supp. 2069]

52. What is meant by Octet rule? Give the Lewis structure of $(NH_4)_2SO_4$. [Q.N. 11 (Or), 2070 'C']

53. Write any two applications of dipole moment. [Q.N. 13, 2070 'C']

54. Mention the requirements for a molecule to fulfil for the formation of hydrogen bond. [Q.N. 16, 2070 'C']

55. Write a Covalent Compound formed by nitrogen and oxygen. What is its Lewis structure? [Q.N. 7(Or), 2070 'D']

56. Define Polar Covalent bond and give an example of it. [Q.N. 8, 2070 'D']

57. How does hydrogen bond affect the physical properties of compound? [Q.N. 9, 2070 'D']

Long Questions

(All questions are of equal value, 10 marks each.)

1. The elements X and Y have the atomic numbers 11 and 17 respectively.
   (i) Write the electronic configuration of the elements.
   (ii) State the type of bond when they combine to form a compound.
   (iii) In which group of periodic table do they belong? [Q.N.23, 2059]

2. Discuss the formation of potassium chloride and carbon tetrachloride molecules on the basis of electronic theory of valency. Give any two characteristics of ionic and covalent compounds. [Q.N.23, 2060]

3. Write short notes on :
   (a) Resonance [Q.N.33 (b), 2060]
   (b) Resonance [Q.N.33 (iii), 2062]
   (c) Covalent bonding [Q.N. 33,(b) 2063]

Unit 7 – Periodic Classification of Elements

Very Short Questions

(All questions are of equal value, 2 marks each.)

1. Ar comes before K in the periodic table, but Ar has a larger relative atomic mass than K. Explain. [Group A, Q.N.5, 2052]

2. Is a potassium atom larger, smaller or the same size as a potassium ion ? Explain. [Group A, Q.N.11, 2051]

3. Why does the atomic size increase in going down any family of the periodic table. [Group A, Q.N.11, 2052]

4. Why do metals form positive ions and non-metals from negative ions ? [Group A, Q.N.14, 2052]

5. Why are metals form positive ions and non-metals from negative ions? [Group A, Q.N.1, 2053]

6. Which of the following pairs would have a larger size and why?
a) K or K^+ b) F or F^- **[Group A, Q.N.10, 2053]**
7. Why does the first ionization energy increase from left to right in a given period of the periodic table? **[Group A, Q.N.14, 2053]**
8. Why Ionization energies of alkali metals decreases as the atomic number increases. **[Q.N.6, 2054]**
9. Why do the oxide, O^{2-} and sulphide, S^{2-} have negative charge? **[Q.N.4, 2055]**
10. What is the basis of the classification of elements in Mendeleev's periodic table? **[Q.N.8, 2055]**
11. Why do atomic radii decrease across a period and increase in a group with the increase of atomic number? **[Q.N.7, 2056]**
12. Arrange the elements Na, Li and K in the increasing order of first ionization energy. **[Q.N.12, 2057]**
13. What was the basis of the classification of elements in Mendeleev's periodic table? **[Q.N.9, 2058]**
14. Why is the ionisation energy of lithium greater than sodium? **[Q.N.6, 2059]**
15. Compare the size of $F^{\ominus}$ and $Na^{\oplus}$ with the atomic size of Neon. **[Q.N.8, 2060]**
16. Why is the size of $Cl^{\ominus}$ ion is larger than Cl atom where as size of K^+ ion is smaller than that of K atom ? **[Q.N.6, 2062]**
17. State Modern Periodic law. **[Q.N. 10. 2063]**
18. Why is ionization energy of oxygen less than that of nitrogen ? **[Q.N. 7, 2065]**
19. State Modern Periodic Law. **[Q.N.8, 2066]**
20. State Modern Periodic law. **[Q.N. 9, 2067]**
21. What is the basic difference between modern periodic table and Mendeleev's periodic table ? **[Q.N.8,2068]**
22. State Modern Periodic Law. **[Q.N. 7, Supp. 2068]**
23. Define ionization energy and atomic radius. **[Q.N. 10, Set 'A' 2069]**
24. Write any two demerits of Mendeleevs periodic table. **[Q.N. 10, Set 'B' 2069]**
25. State Mendeleev's Periodic Law and mention a demerit of the Mendeleev's periodic table. **[Q.N. 8, Supp. 2069]**
26. State the Mendeleev's periodic law. **[Q.N. 5, 2070 'C']**
27. On what basis does Mendeleev's periodic Law differ from Modern periodic Law? **[Q.N. 10, 2070 'D']**

Short Questions

(All questions are of equal value, 5 marks each.)

1. "Periodicity in chemical properties is based on atomic structure." Explain. **[Group B, Q.N.2, 2051]**
2. Show your acquaintance with modern Periodic Table. **[Q.N.22, 2055]**
3. How does atomic size vary within a horizontal row of the periodic table? Explain how this variation arises? **[Q.N. 25 (a), 2058]**
4. The Mg^{2+} and Na^+ have same number of electrons (ten). Which ion would you expect to have the smaller radius? Explain. **[Q.N.25 (b), 2058]**
5. Define ionization energy. How do 'Nuclear charge' and 'Size of the atom' influence the magnitude of the ionization energy ? Ionization energy of 'N' is higher than that of 'O'. Give reason. **[Q.N.24,2068]**
6. What is a group in periodic table? How do atomic size and electron affinity vary in a group? 1+2+2 **[Q.N. 23, Supp. 2068]**
7. State Modern Periodic Law. What are the advantages of modern periodic table? 1+4 **[Q.N. 24, Set 'A' 2069]**

Long Questions

(All questions are of equal value, 10 marks each.)

1. Write short notes on:
(a) Modern Periodic Table. **[Group C, Q.N.4 (a), 2053]**
(b) Modern Periodic Table **[Q.N.33 (i), 2062]**
(c) Modern Periodic Table. **[Q.N. 33(a), 2064]**