

PM SHRI KENDRIYA VIDYALAYA KARGIL

HOLIDAYS HOMEWORK

Class-XI

Biology

1. Revise complete Plant & Human Physiology Units.
2. Complete Biology Notebook & Practical Record file.
3. Complete Biology Notebook & Practical Record file.
4. Revision Schedule for PT-2 and Final Exams
5. Weekly modules will be shared in class group for Unit-1, 2 & 3 revision

4 th December 2023- 9 th December 2023	1. Photosynthesis in Higher Plants 2. Respiration in Plants
10 th December 2023- 16 th December 2023	3. Plant Growth and Development 4. Breathing and Exchange of Gases
17 th December 2023- 23 rd December 2023	5. Body Fluids and Circulation 6. Excretory Products and Their Elimination
24 th December 2023- 30 th December 2023	7. Locomotion and Movement 8. Neural Control and Coordination
31 st December 2023- 6 th January 2024	9. Chemical Coordination and Integration 10. The Living World
7 th January 2024 – 13 th January 2024	11. Biological Classification 12. Plant Kingdom 13. Animal Kingdom
14 th January 2024- 20 th January 2024	14. Morphology of Flowering Plants 15. Anatomy of Flowering Plants 16. Structural Organisation in Animals
21 st January 2024- 27 th January 2024	17. Cell: The Unit of Life 18. Biomolecules 19. Cell Cycle and Cell Division
28 th January 2024- 3 rd February 2024	Sample Ques Paper 1 & 2
4 th February 2024- 10 th February 2024	Sample Ques Paper 3 & 4

6.

Mathematics

1. DO ALL REMAINING QUESTIONS OF EXERCISES 12.2 AND MISCELLANEOUS EXERCISES OF LIMIT AND DERIVATIVES.
2. Review and revise the syllabus thoroughly.
3. Be prepared for periodic test 2 and do maximum practice on questions and examples of straight line, probability and conic section.
4. Syllabus for periodic test II are straight lines, probability, conic section.
5. Do maximum practice on each questions, examples and exercises of each units that is unit 1 set function, unit 2 algebra , unit 3 coordinate geometry, unit 4 calculus unit 5 statistics and probability.
6. Special attention must be paid towards algebra and set function as they contains 25 marks and 23 marks respectively in board exams.

ANSWER THE FOLLOWING QUESTIONS

Q-1 A man has 6 friends. In how many ways can be invite one or more of them to a dinner party?

Q-2 A boy has 3 library tickets and 8 books of his interest in the library. Of these 8, he does not want to borrow Mathematics part-II, unless Mathematics Part-1 is also borrowed. In how many ways can he choose the three books to be borrowed?

Q-3 We wish to select 6 persons from 8, but if the person A is chosen, then B must be chosen. In how many ways can the selection be made?

Q-4 In how many ways can a Football Eleven be selected out of 15 players?

- i) In how many of them a particular player is included? ii) In how many of them he is excluded?

Q-5 In a plane there are 27 straight lines, of which 13 pass through the point A and 11 pass through the point B. Besides, no three lines pass through one point, no line passes through both points A and B, and no two are parallel. Find the number of points of intersection of the straight lines.

Q-6 Find the number of all possible arrangements of the letters of the word "MATHEMATICS" taken four at a time. Also find the arrangements when a pair of repeated letters and two distinct letters are used.

PART 11

Complete the following Activities in your Activity Record note book

1. Trigonometric Ratios in different Quadrants
2. Linear Inequalities
3. Permutations and combinations
4. Sample space of a coin

Chemistry

1. REVISION OF ALL THE CHAPTERS.
2. PRACTICE NCERT INTEXT QUESTION/EXERCISE QUESTION.
3. SOLVE MCQ/REASONING/ONE WORD QUESTION.
4. REVISION OF IMPORTANT CONCEPTS AS SHOWN IN CLASS

Physics

Prepare a notebook and write this homework in that notebook. Do at least 10 numericals covering each topic from every chapter

BY KARAN KINDRA

70-day holiday homework covering the full syllabus for Class 11 Physics. I'll provide a general outline, but please adjust the plan according to your specific syllabus and pacing. This plan assumes a five-day study week:

Days 1-5: Units and Measurements

Days 1-2: Read the chapter and understand the basic concepts.

Days 3-4: Solve Example Problems and Conceptual Questions.

Day 5: Work on Exercise Questions and Numericals.

Days 6-15: Motion in a Straight Line

Days 6-7: Read the chapter and understand the fundamentals.

Days 8-10: Solve problems from different exercises.

Days 11-13: Focus on conceptual questions.

Days 14-15: Work on numerical problems and revise the chapter.

Days 16-25: Laws of Motion

Days 16-17: Understand Newton's laws and their applications.

Days 18-20: Solve problems related to friction, tension, and inclined planes.

Days 21-23: Focus on numerical problems.

Days 24-25: Conduct experiments related to laws of motion if possible.

Days 26-35: Work, Energy, and Power

Days 26-27: Study the concepts of work, energy, and power.

Days 28-30: Solve numerical problems and problems from different exercises.

Days 31-32: Explore applications of work, energy, and power.

Days 33-35: Conduct experiments or demonstrations related to work and energy.

Days 36-45: System of Particles and Rotational Motion

Days 36-37: Understand the center of mass and its motion.

Days 38-40: Study rotational motion and torque.

Days 41-42: Solve numerical problems.

Days 43-45: Focus on concept-based questions.

Days 46-55: Gravitation

Days 46-47: Understand the law of gravitation.

Days 48-50: Solve numerical problems and problems from different exercises.

Days 51-52: Explore applications of gravitational force.

Days 53-55: Conduct experiments or demonstrations related to gravitation.

Days 56-65: Mechanical Properties of Solids and Fluids

Days 56-57: Study stress, strain, and elastic constants.

Days 58-60: Focus on fluid mechanics and hydrostatics.

Days 61-62: Solve numerical problems and problems from different exercises.

Days 63-65: Conduct experiments or demonstrations related to the mechanical properties of solids and fluids.

Days 66-70: Waves, Oscillations, and Thermodynamics

Days 66-67: Understand the basics of wave motion and oscillations.

Days 68-69: Study thermodynamics, solve numerical problems, and revise the entire syllabus.

Day 70: Take a comprehensive review of all chapters, focusing on integration and connections between topics.

Remember to maintain a consistent study schedule, review regularly, solve practice papers, and seek clarification from me on doubts by messages or calls. Adjust the plan based on your understanding and pace. Good luck!

हिन्दी

विषयाध्यापक :- डी आर ईणखिया Mr. D R INKHIYA

1. दो पत्र लेखन (औपचारिक पत्र व अनौपचारिक पत्र)
2. आत्मपरिचय और दिन जल्दी जल्दी ढलता है के प्रश्नोत्तर लिखना।
3. अभिव्यक्ति और माध्यम पुस्तक से समस्या इकाईयों से पाँच - पाँच प्रश्न-उत्तर तैयार करना।
4. पठित और अपठित गद्यांश का अध्ययन करना और उसके प्रश्नोत्तर लिखना।

5. अपठित काव्यांश को हल करना , जो प्रतिदर्श प्रश्नपत्र व अध्ययन सामग्री में दिए गए हैं।
6. विद्यालय की तरफ़ से दिए गए अध्ययन सामग्री में जो सारांश दिए गए हैं उसे पढ़ना और लिखना।
7. पाठ्यक्रम की कोई भी एक कविता याद करके तथा उसका वीडियो बनाकर मुझे व्यक्तिगत प्रेषित करें। (कविता डेढ़ मिनट से पाँच मिनट)
8. 'कारगिलविजय दिवस' पर रिपोर्ट (प्रतिवेदन) लिखें।
9. ज़िला कारगिल के दर्शनीय-स्थलों की जानकारी देते हुए दिल्ली में बसे अपने मित्र को पत्र लिखिए।
10. अपने पाठ्यक्रम से जनसंचार के कोई पाँच प्रश्न
12. अभी तक जो भी पढ़ाया गया है, उसको कम-से कम पाँच बार दोहराएँ। आते ही उसी से प्रश्न पूछे जाएँगे।
13. पढ़ाये गए सभी पाठों से प्रत्येक पाठ से 10 -10 बहुविकल्पीय व सारांश पढ़कर प्रश्नोत्तर तैयार करना।

English

Prescribed Books

- Hornbill: English Reader published by National Council Education Research and Training, New Delhi
- The Portrait of a Lady (Prose)
- A Photograph (Poem)
- We're Not Afraid to Die...if We Can be Together
- Discovering Tut: the sage Continues
- The Laburnum Top (Poem)
- The Voice of the Rain (Poem)
- The Adventure
- Silk Road (Prose)
- Father to Son
- Snapshots: Supplementary Reader published by National Council of Education Research and Training , New Delhi
- The Summer of the Beautiful White Horse (Prose)
- The Address (Prose)
- Mother's Day (Play)
- Birth (Prose)
- The Tale of Melon City

Section A – 26 Marks
Reading Skills

I Reading Comprehension through Unseen Passages

1. One unseen passage to assess comprehension, interpretation , analysis, Inference and vocabulary. The passage may be factual, descriptive or literary,
2. One unseen case-based factual Passage with verbal/visual inputs like statistical data,

Note: The combined word limit for both the passages will be 600-750

Multiple Choice Questions / Objective Type Questions will be asked.

3. Note Making and Summarization based on a passage of approximately 200-250 words.

5 Marks

- i. Note Making:

: Title : 1

: Numbering and indenting : 1

: Key/glossary : 1

: Notes: 2

- ii. Summary (up to 50 words) :

3 Marks

Section B – 23 Marks

Grammar and creative writing skills

II. Grammar

7 Marks

4. Questions on Gap filling (Tenses, Clauses)
 5. Questions on re-ordering/transformation of sentences
- (Total seven questions to be done out of the eight given).**

III. Creative Writing Skills

6. Short writing task – Classified Advertisements, up to 50 words. One out of the two given questions to be answered (**3 Marks**: Format :1 / Expression:1)
7. Short writing task – Poster up to 50 words. One out of the two given questions to be answered (3 Marks: Format : 1 / Content :1 / Expression : 1)
8. Long Writing task : Speech in 120-150 words based on verbal / visual cues related to contemporary /age-appropriate topic. One out of the two given questions to be answered.(5 Marks: Format :1/Content:2/ Expression:2)
9. Long Writing Task : Debate based on visual/verbal inputs in 120-150 words, Thematically related to contemporary , topical issues, One out of the two given questions to be answered. (5 Marks: Format:1 /Content;2/Expression:2)

Physical Education

1. In which year, the first Asian Games were organised in India?
2. After how many years, the Asian Games are organised?
3. Jesse Owens Trophy is related to which sports?
4. In which year, the first National Basketball Championships was held
5. Thomas cup is related to which game?
6. Name the historical cricket test series which is played between England and Australia?
7. Aga Khan Cup is related to which game?
8. FIFA Cup is related to which game?
9. In which year, the IPL related to cricket was started?
10. Wimbledon Cup is related to which game?
11. In which year 'Khelo India' programme was launched?

Fill in the blanks.

1. The aim of physical education is harmonious..... of an individual.
2. Coaching career is acareer is the field of physical education.
3. Nowadays, people have become more.....conscious.
- 4.....opportunities are widely available in book writing in the field of physical education and sports.
5. The World Cup of Hockey for men and women was started in..... and 1974, respectively.
- 6.....programme has been introduced to revive the sports culture in India.

State True or False.

1. aim of physical education is harmonious development of an individual
2. Khelo India programme was launched during 2017-2018.
3. Physical fitness of school-going children is not measured under the scheme of Khelo Ind (True/False)
4. Kehlo India programme is not helpful for promoting rural games. (True/False)

Choose the correct answer.

Competitions are helpful for the development of

- (a) Social qualities (c) Sports skills
(b) Selection of players (d) All the above

2. Uber Cup is related to:

- (a) Table Tennis (c) Badminton
(b) Lawn Tennis
(d) None of these

3. Which one of the following components is not included in Khelo India Programme?

- (a) Playfield development
(b) Talent search and development
(c) Sports for women
(d) To organise Cricket Competitions

4. Which one of the following objectives is not the objective of physical education ment by H Clarke?

- (a) Physical fitness (c) Social efficiency
(b) Worthy use of leisure time (d) Culture

Short Answer Questions-I

1. Enlist the objectives Write short notes on any two of the following:

- (a) Career in book writing.
(b) Career in sports photography.
(c) Career in sports industry.

7. Elaborate the 'Khelo India' Programme in detail. physical education and explain any one of them.

2. Discuss about teaching careers in physical education in brief.

3. Enumerate the objectives of physical education in brief.

4. Briefly discuss about any three International Competitions of athletics.

5. Discuss the three national competitions of cricket in brief.
6. Write down any nine tournaments of football.
7. Write down a brief note on 'Khelo India' Programme.

Short Answer Questions

1. Define physical education and explain its aim and objectives in detail.
2. What are the various career options in physical education? Discuss the teaching career in detail.
3. Elucidate the coaching career in detail.
4. Discuss the health-related and administration-related careers in detail.
5. Explain in detail the sports media related careers in the field of physical education.
6. Write short notes on any two of the following:
 - (a) Career in book writing.
 - (b) Career in sports photography.
 - (c) Career in sports industry.

7. Elaborate the 'Khelo India' Programme in detail.

Project Work

1. Physical Fitness Test: SAI Khelo India Test, Brockport Physical Fitness Test (BPFT)*
2. Proficiency in Games and Sports
(Skill of any one IOA recognized Sport/Game of Choice) **
3. Yogic Practices
4. Record File ***

***Record File shall include:

Practical-1: Fitness tests administration. (SAI Khelo India Test)

Practical-2: Procedure for Asanas, Benefits & Contraindication for any two Asanas for each lifestyle disease.

Practical-3: Any one IOA recognized Sport/Game of choice. Labelled diagram of Field & Unit 6. test measurement and Evaluation

1. define test measurement and evaluation
2. importance of test measurement and evaluation in sports.
3. Calculation of BMI Waist -- hip ratio, skin fold.
4. somato type (Endomorphy, Mesomorphy and Ectomorphy)
5. measurement of health- related fitness.

Unit 7. fundamentals of anatomy, physiology in sports

1. definition and importance of anatomy physiology in exercise and sports
2. function of skeletal system, classification of bones, and types of joints.
3. properties and functions muscles.
4. structure and functions of circulatory system and heart .
5. structure and functions of respiratory system.

Unit.8 fundamentals of kinesiology and biomechanics in sports.

Unit 9 psychology and sports.

Unit10. training and doping in Sports

Complete your projects and practical copy.

Computer Science

Revise all topics and if any doubt feel free to ask on whatsapp group.

Unit 1:

Introduction to Computer System Introduction to computer and computing

Components of a computer system and their interconnections

Input/output devices.

Computer Memory: Units of memory, types of memory – primary and secondary

Software: purpose and types – system and application software, generic and specific purpose software.

Unit 2:

Introduction to Python

Python interpreter - interactive and script mode

The structure of a program, indentation, identifiers, keywords, constants, variables

Types of operators, precedence of operators

Data types, mutable and immutable data types, data type conversion,

Statements, expressions, evaluation and comments, input and output statements, debugging. **Control Statements:** if-else, if-elif-else, while loop, for loop

Lists: list operations - creating, initializing, traversing and manipulating lists, list methods

- **Built-in functions:** len(), list(), append(), insert(), count(), index(), remove(), pop(), reverse(), sort(), min(), max(), sum()

Dictionary: Concept of key-value pair, creating, initializing, traversing, updating and deleting elements, dictionary methods

- **Built-in functions:** dict(), len(), keys(), values(), items(), update(), del(), clear()

Unit 3:

Database concepts and the Structured Query Language Database Concepts:

Introduction to database concepts and its need

Database Management System

Relational data model: Concept of domain, tuple, relation

Keys: candidate key, primary key, alternate key

Structured Query Language: Data Definition Language, Data Query Language and Data Manipulation Language, Introduction to MySQL, creating a database using MySQL, Data Types

Data Definition: CREATE DATABASE, CREATE TABLE, DROP, ALTER

Data Query: SELECT, FROM, WHERE with relational operators, BETWEEN, logical operators, IS NULL, IS NOT NULL

Data Manipulation: INSERT, DELETE, UPDATE

Unit 4:

Society, Law and Ethics - Cyber safety

- Cyber safety: safely browsing the web, identity protection, confidentiality, social networks, cyber trolls and bullying
- Appropriate usage of social networks: spread of rumours, and common social networking sites (Twitter, LinkedIn, and Facebook) and specific usage rules.
- Safely accessing web sites: adware, malware, viruses, Trojans
- Safely communicating data: secure connections, eavesdropping, phishing and identity verification.

Find out the Answers of the following Questions and write down on your class notebook.

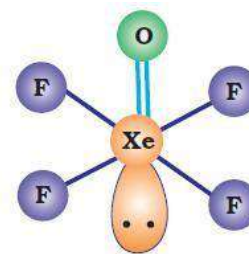
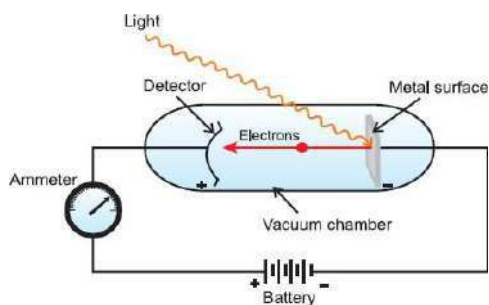
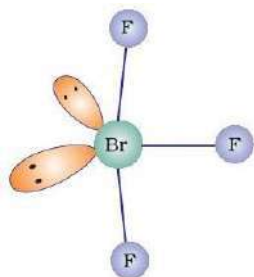
1. After practicals, Atharv left the computer laboratory but forgot to sign off from his email account. Later, his classmate Revaan started using the same computer. He is now logged in as Atharv. He sends inflammatory email messages to few of his classmates using Atharv's email account. Revaan's activity is an example of which of the following cyber crime? Justify your answer.
 - a) Hacking
 - b) Identity theft
 - c) Cyber bullying
 - d) Plagiarism
2. How would you recognise if one of your friends is being cyber bullied?
 - a) Cite the online activities which would help you detect that your friend is being cyber bullied?
 - b) What provisions are in IT Act 2000, (amended in 2008) to combat such situations.
3. Write the differences between the following
 - a) Copyrights and Patents A+B Rationalised 2023-24 SOCIETAL IMPACT 249
 - b) Plagiarism and Copyright infringement
 - c) Non-ethical hacking and Ethical hacking
 - d) Active and Passive footprints
 - e) Free software and Free and open source software
4. If you plan to use a short text from an article on the web, what steps must you take in order to credit the sources used?
5. When you search online for pictures, how will you find pictures that are available in the free public domain. How can those pictures be used in your project without copyright violations?
6. Describe why it is important to secure your wireless router at home. Search the Internet to find the rules to create a reasonably secure password. Create an imaginary password for your home router. Will you share your password for home router with following people. Justify your answer. a) Parents b) Friends c) Neighbours d) Home Tutors
7. List down the steps you need to take in order to ensure a) your computer is in good working condition for a longer time. b) smart and safe Internet surfing.
8. What is data privacy? Websites that you visit collect what type of information about you?
9. Sumit got good marks in all the subjects. His father gifted him a laptop. He would like to make Sumit aware of health hazards associated with inappropriate and excessive use of laptop. Help his father to list the points which he should discuss with Sumit.
10. Cite examples depicting that you were a victim of following cyber crime. Also, cite provisions in IT Act to deal with such a cyber crime. a) Identity theft b) Credit card account theft

Write All program given below:

1. Write a program to read a list of n integers and find their median. Note: The median value of a list of values is the middle one when they are arranged in order. If there are two middle values then take their average. Hint: You can use an built-in function to sort the list
2. Write a program to read a list of elements. Modify this list so that it does not contain any duplicate elements, i.e., all elements occurring multiple times in the list should appear only once.
3. Write a program to read a list of elements. Input an element from the user that has to be inserted in the list. Also input the position at which it is to be inserted. Write a user defined function to insert the element at the desired position in the list.
4. Write a program to read elements of a list. a) The program should ask for the position of the element to be deleted from the list. Write a function to delete the element at the desired position in the list. b) The program should ask for the value of the element to be deleted from the list. Write a function to delete the element of this value from the list.
5. Read a list of n elements. Pass this list to a function which reverses this list in-place without creating a new list.
6. Write a program to read email IDs of n number of students and store them in a tuple. Create two new tuples, one to store only the usernames from the email IDs and second to store domain names from the email IDs. Print all three tuples at the end of the program. [Hint: You may use the function `split()`]
7. Write a program to input names of n students and store them in a tuple. Also, input a name from the user and find if this student is present in the tuple or not.

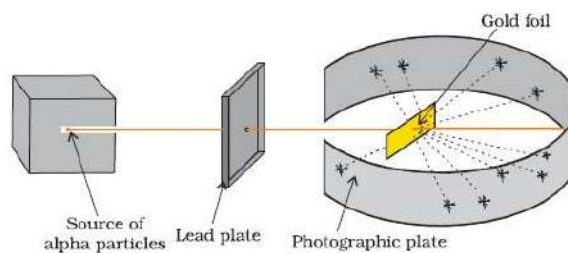
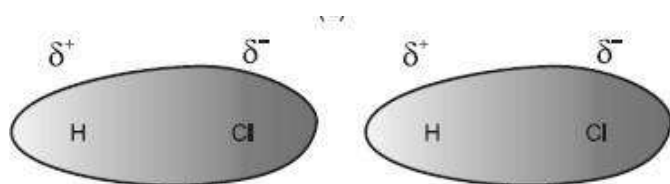
We can accomplish these by: (a) writing a user defined function (b) using the built-in function

8. Write a Python program to find the highest 2 values in a dictionary.
9. Write a program to input your friends' names and their Phone Numbers and store them in the dictionary as the key-value pair. Perform the following operations on the dictionary:
 - a) Display the name and phone number of all your friends
 - b) Add a new key-value pair in this dictionary and display the modified dictionary
 - c) Delete a particular friend from the dictionary
 - d) Modify the phone number of an existing friend
 - e) Check if a friend is present in the dictionary or not
 - f) Display the dictionary in sorted order of names.
10. How is built-in function `pow()` function different from function `math.pow()` ? Explain with an example.
11. Write a program to find the sum of digits of an integer number, input by the user.
12. Write a function that checks whether an input number is a palindrome or not.
13. Write a program to print any 3 different types patterns
14. Write a program to check if the year entered by the user is a leap year or not.
15. Write a program to find the sum of $1 + \frac{1}{8} + \frac{1}{27} + \dots + \frac{1}{n^3}$, where n is the number input by the user.



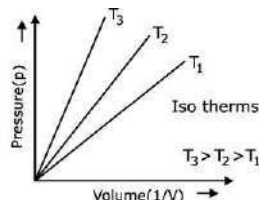
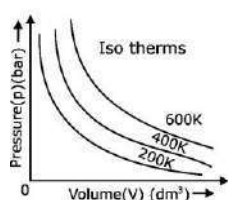
(d) Square pyramidal

HOLIDAY HOME WORK 2023 -24



A. Rutherford's scattering experiment

S.No.	Name of unit
1	Some basic principles of chemistry
2	Structure of atom
3	Classification of Elements and Periodicity in properties
4	Chemical Bonding and Molecular Structure
5	Thermodynamics
6	Equilibrium

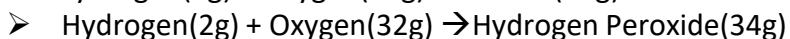
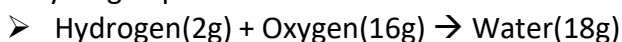


CHEMISTRY

PM SHRI KENDRIYA VIDYALAYA KARGIL

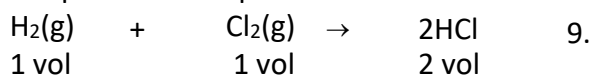
1.SOME BASIC CONCEPTS OF CHEMISTRY

1. In **scientific notation (exponential Notation)** : we can write 232.508 as 2.32508×10^2 in scientific notation. Similarly, 0.00016 can be written as 1.6×10^{-4} .
2. **Precision** refers to the closeness of various measurements for the same quantity.
3. **Accuracy** is the agreement of a particular value to the true value of the result.
4. **Significant figures** are meaningful digits which are known with certainty.
5. **Law of Conservation of Mass** : In a chemical change total mass remains conserved. i.e. mass before reaction is always equal to mass after reaction.
6. **Law of Definite Proportions** : It states that a given compound always contains exactly the same proportion of elements by weight. All chemical compounds are found to have constant composition irrespective of their method of preparation or sources Example : In water (H_2O), Hydrogen and Oxygen combine in 2 : 1 molar ratio, this ratio remains constant whether it is tap water, river water or sea water or produced by any chemical reaction.
7. **Law of Multiple Proportions** states that if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers. For example, hydrogen combines with oxygen to form two compounds, namely, water and hydrogen peroxide.



Here, the masses of oxygen (i.e. 16 g and 32 g) which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e. 16:32 or 1: 2.

8. **Gay Lussac's Law of Gaseous Volumes**: According to him elements combine in a simple ratio of atoms, gases combine in a simple ratio of their volumes provided all measurements should be done in the same temperature and pressure



Avogadro's hypothesis : Equal volume of all gases have equal number of molecules (not atoms) at same temperature and pressure condition. **S.T.P.** (Standard Temperature and Pressure) At S.T.P. condition : temperature = $0^\circ C$ or 273 K Pressure = 1 atm = 760 mm of Hg and volume of one mole of gas at STP is found to be experimentally equal to 22.4 litres which is known as molar volume.

10. One **atomic mass unit** is defined as a mass exactly equal to one twelfth the mass of one carbon - 12 atom.
11. **Molecular mass** is the sum of atomic masses of the elements present in a molecule. It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together.
12. **Mole** : A mole is the amount of substance that contains as many entities (atoms, molecules or other particles) as there are atoms in exactly 0.012 kg (or 12 gm) of the carbon-12 isotope.
13. The number of entities in 1 mol is so important that it is given a separate name and symbol known as **Avogadro constant denoted by N_A** . i.e. on the whole we can say that 1 mole is the collection of 6.02×10^{23} entities..
14. An **empirical formula** represents the simplest whole number ratio of various atoms present in a compound whereas the molecular formula shows the exact number of different types of atoms present in a molecule of a compound.
15. **Limiting reagent** : The reactant which is consumed first and limits the amount of product formed in the reaction, and is there fore, called limiting reagent. .
16. **Mass per cent** = Mass of solute per 100 g of solution

$$\text{Mass per cent} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

17. **Mole Fraction**: It is ratio of number of moles of a particular component to the total number of moles of all the components.

• **Mole-fraction of solute(B)** =
$$\frac{\text{No. of moles of solute}}{\text{No. of moles of solute} + \text{No. of moles of solvent}}$$

$$x^B = \frac{n_B}{n_A + n_B} = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A} + \frac{W_B}{M_B}}$$

- **Mole-fraction of solvent (A)** = $\frac{\text{No. of moles of solvent}}{\text{No. of moles of solute} + \text{No. of moles of solvent}}$

$$x^A = \frac{n_A}{n_A + n_B} = \frac{\frac{W_A}{M_A}}{\frac{W_A}{M_A} + \frac{W_B}{M_B}} \quad \text{Thus } x_B + x_A = 1$$

18. **Molality (m)**. It is defined as number of moles of solute per 1000 g or 1 kg of solvent.

$$(\text{m}) = \frac{\text{No. of moles of solute}}{\text{Kg. of solvent}} = \frac{W_{\text{solute}}}{M_{\text{solute}}} \times \frac{1000}{W_{\text{solvent}} \text{ in grams.}}$$

19. **Molarity (M)**. It is expressed as the number of moles of solute per litre of solution..

$$\text{Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{Litres of solution}} = \frac{W_{\text{solute}}}{M_{\text{solute}}} \times \frac{1000}{\text{Volume of solution (ml)}}$$

$$M = \frac{W_{\text{solute}}}{M_{\text{solute}}} \times \frac{1000 \times \text{density of Solution}}{\text{Mass of solution.}}$$

20. Molarity is an unit that depends upon temperature. varies inversely with temperature. Mathematically :

Molarity decreases as temperature increases.

$$\text{Molarity} \propto \frac{1}{\text{temperature}} \propto \frac{1}{\text{volume}}$$

21. If a particular solution having volume V_1 and molarity = M_1 is diluted to V_2 ml the $M_1V_1 = M_2V_2$

M_2 : Resultant molarity

Answer the following.....

- Express the following in the scientific notation: (i) 0.0048 (ii) 234,000 (iii) 8008 (iv) 500.0 (v) 6.0012
- How many significant figures are present in the following? (i) 0.0025 (ii) 208 (iii) 5005 (iv) 126,000 (v) 500.0 (vi) 2.0034
- What is the difference between 160 cm and 160.0 cm
- Round up the following upto three significant figures: (i) 34.216 (ii) 10.4107 (iii) 0.0459 (iv) 2808
- The following data are obtained when dinitrogen and dioxygen react together to form different compounds :

S.No	Mass of dinitrogen	Mass of dioxygen
1	14 g	16 g
2	14 g	32 g
3	28 g	32 g
4	28 g	80 g

- Which law of chemical combination is obeyed by the above experimental data? Give its statement.
- Convert the following into basic units: (i) 28.7 pm (ii) 15.15 pm (iii) 25365 mg
- State and illustrate the law of constant proportions.
- State and explain the law of multiple proportions.
- Copper oxide obtained by heating copper carbonate or copper nitrate contains copper and oxygen in the same ratio by mass. Which law is illustrated by this observation? State the law.
- What is the percentage of carbon, hydrogen and oxygen in ethanol?
- Calculate the mass per cent of different elements present in sodium sulphate (Na_2SO_4).
- How much copper can be obtained from 100 g of copper sulphate (CuSO_4)?
- In a reaction $\text{A} + \text{B}_2 \rightarrow \text{AB}_2$ Identify the limiting reagent, if any, in the following reaction mixtures. (i) 300 atoms of A + 200 molecules of B (ii) 2 mol A + 3 mol B (iii) 100 atoms of A + 100 molecules of B (iv) 5 mol A + 2.5 mol B (v) 2.5 mol A + 5 mol B
- In the combustion of methane, what is the limiting reactant and why?
- Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation: $\text{N}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$ (i) Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen

reacts with 1.00×10^3 g of dihydrogen. (ii) Will any of the two reactants remain unreacted? (iii) If yes, which one and what would be its mass?

16. Determine the molecular formula of an oxide of iron in which the mass per cent of iron and oxygen are 69.9 and 30.1 respectively.
17. A compound contains 25% hydrogen and 75% carbon by mass. Determine the empirical formula of the compound.
18. In three moles of ethane (C_2H_6), calculate the following: (i) Number of moles of carbon atoms. (ii) Number of moles of hydrogen atoms. (iii) Number of molecules of ethane.
19. If ten volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?
20. A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.
21. Calculate the amount of water (g) produced by the combustion of 16 g of methane.
22. How many moles of methane are required to produce 22 g CO (g) after combustion?
23. What will be the mass of one ^{12}C atom in g?
24. A compound contains 4.07 % hydrogen, 24.27 % carbon and 71.65 % chlorine. Its molar mass is 98.96 g. What are its empirical and molecular formulas?
25. How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?
26. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction, $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$. What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl?
27. Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction $4HCl(aq) + MnO_2(s) \rightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$. How many grams of HCl react with 5.0 g of manganese dioxide?
28. Define the following terms. (Write the mathematical formulas related to terms) (a) Mass percent (b) Molarity (c) Molality (d) Mole-fraction (e) Mass percent
29. Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250 mL of the solution.
30. A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass per cent of the solute.
31. Calculate molality of 2.5g of ethanoic acid (CH_3COOH) in 75g of benzene.
32. How does molality & Molarity depend on temperature? Out of molality & Molarity which one is better way to express concentration?
33. Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. (Molar mass of sodium acetate is $82.0245 \text{ g mol}^{-1}$).
34. Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g mL^{-1} and the mass per cent of nitric acid in it being 69%.
35. What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2L?
36. If the density of methanol is 0.793 kg L^{-1} , what is its volume needed for making 2.5 L of its 0.25 M solution?
37. A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass). (i) Express this in percent by mass. (ii) Determine the molality of chloroform in the water sample.
38. The density of 3 M solution of NaCl is 1.25 g mL^{-1} . Calculate molality of the solution.
39. Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040 (assume the density of water to be one)

2. Structure of Atom

1. Properties of Fundamental Particles Sub-atomic particles (Electron, Protons and Neutrons)

PARTICLE	ELECTRON	PROTON	NEUTRON
Discovery	Sir. J. J. Thomson (1869)	Goldstein (1886)	Chadwick (1932)
Nature of charge	Negative	Positive	Neutral
Amount of charge	1.6×10^{-19} Coloumb	1.6×10^{-19} Coloumb	0
Mass	9.11×10^{-31} kg	1.672614×10^{-27} kg	1.67492×10^{-27} kg

2. **Thomson Model of Atom** : An atom possesses a spherical shape (radius approximately 10^{-10} m) in which the positive charge is uniformly distributed. The electrons are embedded into it in such a manner as to give the most stable electrostatic arrangement.

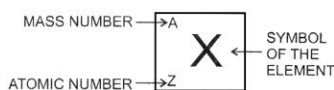
3. Rutherford's Nuclear Model of Atom:

- Most of the space in the atom is empty as most of the α –particles passed through the foil undeflected.
- A few positively charged α – particles were deflected. The deflection must be due to enormous repulsive force showing that the positive charge of the atom is not spread throughout the atom as Thomson had presumed. The positive charge has to be concentrated in a very small volume that repelled and deflected the positively charged α – particles.
- Calculations by Rutherford showed that the volume occupied by the nucleus is negligibly small as compared to the total volume of the atom. The radius of the atom is about 10^{-10} m, while that of nucleus is 10^{-15} m.
- On the basis of above observations and conclusions, Rutherford proposed the nuclear model of atom (after the discovery of protons). According to this model :(i) The positive charge and most of the mass of the atom was densely concentrated in extremely small region. This very small portion of the atom was called nucleus by Rutherford. (ii) The nucleus is surrounded by electrons that move around the nucleus with a very high speed in circular paths called orbits. Thus, Rutherford's model of atom resembles the solar system in which the nucleus plays the role of sun and the electrons that of revolving planets. (iii) Electrons and the nucleus are held together by electrostatic forces of attraction. ${}^1_1\text{H}$

4. **Atomic number (Z)** : The number of protons present in the nucleus is equal to atomic number (Z).the nucleus is equal to atomic number (Z).i.e. **Atomic number (Z) = number of protons in the nucleus of an atom = number of electrons in a neutral atom**

5. **Mass number (A)** Protons and neutrons present in the nucleus are collectively known as nucleons. The total number of nucleons is termed as **mass number (A)** of the atom.mass number (A) = number of protons (Z) + number of neutrons (n)

6. Representation of atoms :



10. **Isotopes**: atoms with identical atomic number but different atomic mass number are known as **Isotopes**. e.g ${}^{12}_6\text{C}$, ${}^{13}_6\text{C}$, ${}^{14}_6\text{C}$ and ${}^1_1\text{H}$ (**protium**) ${}^2_1\text{H}$ (**deuterium**) ${}^3_1\text{H}$ (**tritium**)

11. **Isobars** are the atoms with same mass number but different atomic number for example, ${}^{14}_6\text{C}$ and ${}^{14}_7\text{N}$

12. **Isoelectronic species**: These are those species which have the same number of electrons. For example, O^{2-} , F^- , Na^+ , Mg^{2+} , Al^{3+} , Ne all contain 10 electrons each and hence they are isoelectronic

13. **Drawbacks of Rutherford Model** : According to the electromagnetic theory of Maxwell, charged particles when accelerated should emit electromagnetic radiation .Therefore, an electron in an orbit will emit radiation, the energy carried by radiation comes from electronic motion. The orbit will thus continue to shrink. Calculations show that it should take an electron only 10^{-8} s to spiral into the nucleus. Thus, the Rutherford model cannot explain the stability of an atom.

14. **Relationship between velocity, wavelength and frequency of a wave**. As frequency is the number of waves passing through a point per second and λ is the length of each wave, hence their product will give the velocity of the wave. Thus $v = \nu \times \lambda$

15. Order of wavelength in electromagnetic spectrum

Cosmic rays < γ -rays < X-rays < Ultraviolet rays < Visible < Infrared < Micro waves < Radio waves.

16. **Black body**: An ideal body, which emits and absorbs all frequencies, is called a black body.

17. Planck's Quantum Theory-

- Planck suggested that atoms and molecules could emit (or absorb) energy only in discrete quantities and not in a continuous manner, a belief popular at that time.
- Planck gave the name **quantum** to the smallest quantity of energy that can be emitted or absorbed in the form of electromagnetic radiation.
- The energy (E) of a quantum of radiation is proportional to its frequency (ν) and is expressed by i.e. $E \propto \nu$ or $E = h\nu$ where $h =$ Planck's constant $= 6.626 \times 10^{-27}$ Js
- Energy is always emitted or absorbed as integral multiple of this quantum. $E = nh\nu$ Where $n = 1, 2, 3, 4, \dots$
- With this theory, Planck was able to explain the distribution of intensity in the radiation from black body as a function of frequency or wavelength at different temperatures.

18. **Photoelectric effect:** The phenomenon of ejection of electrons from the surface of metal (for example potassium, rubidium, caesium etc.) when light of suitable frequency strikes it is called photoelectric effect. The ejected electrons are called photoelectrons.

- Photoelectric work function (W_0): The minimum energy required to eject electrons is called photoelectric work function. $W_0 = h\nu_0$

- Energy of the ejected electrons : $h\nu = h\nu_0 + \frac{1}{2} m_e v^2$
- Energy of the ejected electrons : $h(\nu - \nu_0) = \frac{1}{2} m_e v^2$

19. Spectrum is of two types: continuous and line spectrum

- The spectrum which consists of all the wavelengths is called **continuous spectrum**.
- A spectrum in which only specific wavelengths are present is known as a **line spectrum**. It has bright lines with dark spaces between them.

20. **Electromagnetic spectrum** is a continuous spectrum. It consists of a range of electromagnetic radiations arranged in the order of increasing wavelengths or decreasing frequencies. It extends from radio waves to gamma rays.

21. Spectrum is also classified as emission and line spectrum.

- a) **Emission spectrum:** The spectrum of radiation emitted by a substance that has absorbed energy is called an emission spectrum.
- b) **Absorption spectrum** is the spectrum obtained when radiation is passed through a sample of material. The sample absorbs radiation certain wavelengths. The wavelengths which are absorbed are missing and come as dark lines.

22. Spectra lines of Hydrogen Atom:

Series	n_1	$n_2 > n_1$	Region of Spectrum
Lyman	1	2, 3, ..., ∞	Ultraviolet ($2 \rightarrow 1$) $H_{\alpha}, (3 \rightarrow 1)H_{\beta}$ & soon.
Balmer	2	3, 4, ..., ∞	Visible ($3 \rightarrow 2$) $H_{\alpha}, (4 \rightarrow 2)H_{\beta}$ & so on.
Paschen	3	4, 5, ..., ∞	Infrared ($4 \rightarrow 3$) $H_{\alpha}, (5 \rightarrow 3)H_{\beta}$ & so on.
Brackett	4	5, 6, ..., ∞	Infrared ($5 \rightarrow 4$) $H_{\alpha}, (6 \rightarrow 4)H_{\beta}$ & so on.
Pfund	4	6, 7, ..., ∞	Infrared ($6 \rightarrow 5$) $H_{\alpha}, (7 \rightarrow 5)H_{\beta}$ & soon.

- $\frac{n(n-1)}{2} =$ Max. no of s lines possible when an electron returns to ground state from n^{th} orbit.

23. **Rydberg equation** : The Swedish spectroscopist, Johannes Rydberg, noted that all series of lines in the hydrogen spectrum could be described by the following expression

- $\frac{1}{\lambda} = \bar{\nu} = R_H \left[\frac{1}{n_1^2} - \frac{1}{n_2^2} \right] \times Z^2$

$$\bar{\nu} = 109,677 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \text{ cm}^{-1}$$

24. Bohr's Model For Hydrogen Atom

- a) The electron in the hydrogen atom can move around the nucleus in a circular path of fixed radius and energy. These paths are called orbits, stationary states or allowed energy states. These orbits are arranged concentrically around the nucleus.
- b) An electron can move only in those orbits for which its angular momentum is integral multiple of $h/2\pi$ that is why only certain fixed orbits are allowed. The angular momentum of an electron in a

$$m_e v r = n \cdot \frac{h}{2\pi} \quad n = 1, 2, 3, \dots$$

given stationary state can be expressed as in equation

- c) The energy of an electron in the orbit does not change with time. However, the electron will move from a lower stationary state to a higher stationary state when required amount of energy is absorbed by the electron or energy is emitted when electron moves from higher stationary state to lower stationary state. The energy change does not take place in a continuous manner.
- d) The frequency of radiation absorbed or emitted when transition occurs between two stationary states that differ in energy by ΔE , is given by :

$$\nu = \frac{\Delta E}{h} = \frac{E_2 - E_1}{h}$$

25. Bohr's theory can also be applied to the ions containing only one electron, similar to that present in hydrogen atom. For example, He^+ , Li^{2+} , Be^{3+} and so on. The energies of the stationary states associated with these kinds of ions (also known as hydrogen like species) are given by the expression

$$E_n = -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2} \right) \text{J}$$

and radii by the expression

$$r_n = \frac{52.9 (n^2)}{Z} \text{pm}$$

26. **Limitations of Bohr's Model:**, Bohr's theory was also unable to explain the splitting of spectral lines in the presence of magnetic field (**Zeeman effect**) or an electric field (**Stark effect**).
27. **Dual Behaviour of Matter:** The French physicist, de Broglie in 1924 proposed that matter, like radiation, should also exhibit dual behaviour i.e., both particle and wavelike properties.
28. The **de Broglie relation**. :de Broglie relation state that the wavelength associated with a moving object or an electron is inversely proportional to the momentum of the particle.

$$\lambda = \frac{h}{mv} = \frac{h}{p} \quad \text{where } p \text{ is the momentum of particle} = mv.$$

29. **Heisenberg's uncertainty principle:** It states that it is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron. The product of their uncertainties is always equal to or greater than $h/4\pi$. If ' Δx ' is uncertainty in position and ' ΔP ' is uncertainty in momentum then:

$$\Delta x \cdot \Delta P \geq \frac{h}{4\pi}$$

10. **Quantum numbers.** They are used to get complete information about electron, i.e., location, energy, spin, etc. These quantum numbers also help to designate the electron present in an orbital.
11. **Principal quantum number.** It specifies the location and energy of an electron. It is measure of the effective volume of the electron cloud. It is denoted by ' n '. Its possible values are 1, 2, 3, 4,
12. **Angular momentum quantum number.** It is also called 'azimuthal quantum number'. It determines the shape of the orbital. It is denoted by ' l '. The permitted values of ' l ' are 0, 1, 2, etc., upto $n-1$. For a given value of n , $l = 0$ to $n-1$. e.g., if value of n is 4, l can have values 0, 1, 2, 3. It determines angular momentum.

$$mvr = \sqrt{l(l+1)} \frac{h}{2\pi}$$

Subshell notation	s	p	d	f	g
Value of ' l '	0	1	2	3	4
Number of orbitals	1	3	5	7	9

13. **Magnetic quantum number.** It is denoted by ' m ' and its value depends on value of ' l ' since magnetism is due to angular momentum. It determines the magnetic orientation of an orbital, i.e., the direction of orbital

relative to magnetic field in which it is placed. Its permitted values are $-l$ to $+l$ including zero, e.g., when $l = 1$, then $m = -1, 0, +1$. It has total number of values equal to $2l + 1$.

14. **Spin quantum number.** It indicates, the direction in which electron revolves. Spin is magnetic property and is also quantized. It has two permitted values $+\frac{1}{2}$ or $-\frac{1}{2}$. The spin angular momentum of an electron is constant and cannot be changed.
15. **(n+l) rule:** The relative order of energies of various sub-shells in a multi-electron atom can be predicted with the help of (n+l) rule (also called Bohr-Bury rule) According to this rule a sub-shell with lower values of (n+l) has lower energy. In case two sub-shell has equal value of (n+l), the sub-shell with lower value of n has lower energy
16. **Pauli's Exclusion Principle.** No two electrons in an atom can have all the four quantum numbers same. It can also be stated as – An orbital can have maximum two electrons and they must be of opposite spin quantum numbers.
17. **Aufbau principle.** Electrons are filled in the various orbitals in the increasing order of their energies, i.e., orbital having lowest energy will be filled first and the orbital having highest energy will be filled last. **Increasing energy of atomic orbitals for multi-electron atoms**
 $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s$
18. **Exception of Aufbau principle :** Extra stability is associated with the exactly half-filled and fully-filled orbitals. Thus the $p^3, p^6, d^5, d^{10}, f^7, f^{14}$ etc. have extra stability, i.e., lower energy and therefore, more stable.
19. **Hund's rule of maximum multiplicity.** No electron pairing takes place in p, d and f orbitals until each orbital in the given sub-shell contains one electron, e.g., N (7) has electronic configuration $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$ according to Hund's rule and not $1s^2 2s^2 2p_x^2 2p_y^1$.
20. **Degenerate orbitals:** Orbitals having the same energy are called degenerate orbitals.
21. **Shape of p and d-orbitals :** See NCERT

Answer the following:

1. Calculate the number of electrons which will together weigh one gram.
2. Calculate the mass and charge of one mole of electrons.
3. Calculate the total number of electrons present in one mole of methane.
4. Calculate the number of protons, neutrons and electrons in $^{80}_{35}\text{Br}$.
5. The number of electrons, protons and neutrons in a species are equal to 18, 16 and 16 respectively. Assign the proper symbol to the species.
6. Give the number of electrons in the species H_2^+, H_2 and O_2^+ .
7. An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol.
8. An ion with mass number 37 possesses one unit of negative charge. If the ion contains 11.1% more neutrons than the electrons, find the symbol of the ion.
9. An ion with mass number 56 contains 3 units of positive charge and 30.4% more neutrons than electrons. Assign the symbol to this ion.
10. Which of the following are isoelectronic species i.e., those having the same number of electrons? $\text{Na}^+, \text{K}^+, \text{Mg}^{2+}, \text{Ca}^{2+}, \text{S}^{2-}, \text{Ar}$.
11. Compare Electron, Protons and Neutrons.
12. Describe Thomson Model Of Atom.
13. Explain Rutherford's scattering experiment.. What conclusions regarding the structure of atom were drawn by Rutherford on the basis of the observations of experiment? Give the major drawbacks of Rutherford Model.
14. In Rutherford's experiment, generally the thin foil of heavy atoms, like gold, platinum etc. have been used to be bombarded by the α -particles. If the thin foil of light atoms like aluminium etc. is used, what difference would be observed from the above results ?
15. Define the terms Atomic Number, Mass Number, Isobars, Isotopes,
16. The Vividh Bharati station of All India Radio, Delhi, broadcasts on a frequency of 1,368 kHz (kilo hertz). Calculate the wavelength of the electromagnetic radiation emitted by transmitter. Which part of the electromagnetic spectrum does it belong to?

17. The wavelength range of the visible spectrum extends from violet (400 nm) to red (750 nm). Express these wavelengths in frequencies (Hz). ($1\text{nm} = 10^{-9}\text{ m}$)
- 18. Calculate (a) wavenumber and (b) frequency of yellow radiation having wavelength 5800Å.**
- 19. Yellow light emitted from a sodium lamp has a wavelength of 580 nm. Calculate the frequency and wavenumber of the yellow light.**
20. Find energy of each of the photons which (i) correspond to light of frequency $3 \times 10^{15}\text{Hz}$. (ii) have wavelength of 0.50 \AA .
21. Calculate the wavelength, frequency and wave number of a light wave whose period is $2.0 \times 10^{-10}\text{s}$.
22. Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 \AA . Calculate threshold frequency (ν_0) and work function (W_0) of the metal.
23. The work function for caesium atom is 1.9 eV . Calculate (a) the threshold wavelength and (b) the threshold frequency of the radiation. If the caesium element is irradiated with a wavelength 500 nm , calculate the kinetic energy and the velocity of the ejected photoelectron.
24. Write a note on the Spectral Lines for Atomic Hydrogen.
- 25. What are the frequency and wavelength of a photon emitted during a transition from $n = 5$ state to the $n = 2$ state in the hydrogen atom?**
26. Calculate the energy associated with the first orbit of He^+ . What is the radius of this orbit?
27. What is the wavelength of light emitted when the electron in a hydrogen atom undergoes transition from an energy level with $n = 4$ to an energy level with $n = 2$?
28. How much energy is required to ionise a H atom if the electron occupies $n = 5$ orbit? Compare your answer with the ionization enthalpy of H atom (energy required to remove the electron from $n = 1$ orbit).
- 29. What is the maximum number of emission lines when the excited electron of a H atom in $n = 6$ drops to the ground state?**
30. The energy associated with first orbit in the hydrogen atom is $-2.18 \times 10^{-18}\text{J atom}^{-1}$. What is the energy associated with the fifth orbit?
31. Calculate the radius of Bohr's fifth orbit for hydrogen atom.
32. Calculate the wavenumber for the longest wavelength transition in the Balmer series of atomic hydrogen.
33. What is the energy in joules, required to shift the electron of the hydrogen atom from the first Bohr orbit to the fifth Bohr orbit and what is the wavelength of the light emitted when the electron returns to the ground state? The ground state electron energy is $-2.18 \times 10^{-18}\text{ergs}$.
34. The electron energy in hydrogen atom is given by $E_n = (-2.18 \times 10^{-18})/n^2\text{J}$. Calculate the energy required to remove an electron completely from the $n = 2$ orbit. What is the longest wavelength of light in cm that can be used to cause this transition?
35. What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition $n = 4$ to $n = 2$ of He^+ spectrum?
- 36. Explain Dual behaviour of matter.**
- 37. State de Broglie's relation. Give its mathematical expression.**
- 38. What will be the wavelength of a ball of mass 0.1 kg moving with a velocity of 10 m s^{-1} ?**
- 39. The mass of an electron is $9.1 \times 10^{-31}\text{ kg}$. If its K.E. is $3.0 \times 10^{-25}\text{ J}$, calculate its wavelength.**
- 40. Calculate the mass of a photon with wavelength 3.6 \AA .**
41. The velocity associated with a proton moving in a potential difference of 1000 V is $4.37 \times 10^5\text{ms}^{-1}$. If the hockey ball of mass 0.1 kg is moving with this velocity, Calculate the wavelength associated with this velocity.
42. If the velocity of the electron in Bohr's first orbit is $2.19 \times 10^6\text{ms}^{-1}$, calculate the de Broglie wavelength associated with it.
43. Similar to electron diffraction, neutron diffraction microscope is also used for the determination of the structure of molecules. If the wavelength used here is 800 pm , calculate the characteristic velocity associated with the neutron.
44. Calculate the wavelength of an electron moving with a velocity of $2.05 \times 10^7\text{ms}^{-1}$
45. Why de Broglie's relation is not associated with ordinary objects.
- 46. State Heisenberg's Uncertainty Principle. Give its mathematical expression.**

47. A microscope using suitable photons is employed to locate an electron in an atom within a distance of 0.1 \AA . What is the uncertainty involved in the measurement of its velocity?
48. If the position of the electron is measured within an accuracy of $\pm 0.002 \text{ nm}$, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is $h/4\pi \times 0.05 \text{ nm}$, is there any problem in defining this value.
49. Show that the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie wavelength associated with the electron revolving around the orbit.
50. What information is provided by the four quantum numbers?
51. Using s, p, d, f notations, describe the orbital with the following quantum numbers
(a) $n = 2, l = 1$, (b) $n = 4, l = 0$, (c) $n = 5, l = 3$, (d) $n = 3, l = 2$ (e) $n = 1, l = 0$ (f) $n = 3, l = 1$ (g) $n = 4, l = 2$ (h) $n = 4, l = 3$.
52. What is the total number of orbitals associated with the principal quantum number $n = 3$?
53. What is the lowest value of n that allows g orbitals to exist?
54. An electron is in one of the 3d orbitals. Give the possible values of n, l and m_l for this electron.
55. An atomic orbital has $n = 3$. What are the possible values of l and m_l ?
56. List the quantum numbers (m_l and l) of electrons for 3d orbital.
57. Which of the following orbitals are possible? 1p, 2s, 2p, 2d, 4f, 6d and 3f.
58. Explain, giving reasons, which of the following sets of quantum numbers are not possible. (a) $n = 0, l = 0, m_l = 0, m_s = +\frac{1}{2}$ (b) $n = 1, l = 0, m_l = 0, m_s = -\frac{1}{2}$ (c) $n = 1, l = 1, m_l = 0, m_s = +\frac{1}{2}$ (d) $n = 2, l = 1, m_l = 0, m_s = -\frac{1}{2}$ (e) $n = 3, l = 3, m_l = -3, m_s = +\frac{1}{2}$ (f) $n = 3, l = 1, m_l = 0, m_s = +\frac{1}{2}$
59. (i) How many electrons in an atom may have the following quantum numbers?
(a) $n = 4, m_s = -\frac{1}{2}$ (b) $n = 3, l = 0$ (ii) How many sub-shells are associated with $n = 4$? (iii) How many electrons will be present in the sub-shells having m_s value of $-1/2$ for $n = 4$?
60. State (n+l) rule Aufbau rule & Pauli rule.
61. Write electronic configuration of elements with atomic number : 7, 13, 17, 19, 23, 24, 27, 29 & 30 elements.
62. Explain the exceptional configuration of copper and chromium.
63. Give the electronic configurations of the following ions: Cu^{2+} Cr^{3+} Fe^{2+} S^{2-} Fe^{2+} O^{2-} Na^+
64. Explain Hund's rule of maximum multiplicity with an example.
65. Indicate the number of unpaired electrons in : (a) P, (b) Si, (c) Cr, (d) Fe and (e) Kr. (i) Write the electronic configurations of the following ions: (a) H^- (b) Na^+ (c) O^{2-} (d) F^-
66. What are the atomic numbers of elements whose outermost electrons are represented by (a) $3s^1$ (b) $2p^3$ and (c) $3p^5$?
67. Which atoms are indicated by the following configurations? (a) $[\text{He}] 2s^1$ (b) $[\text{Ne}] 3s^2 3p^3$ (c) $[\text{Ar}] 4s^2 3d^1$
68. An atom of an element contains 29 electrons and 35 neutrons. Deduce (i) the number of protons and (ii) the electronic configuration of the element.
69. Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge? (i) 2s and 3s, (ii) 4d and 4f, (iii) 3d and 3p.
70. The unpaired electrons in Al and Si are present in 3p orbital. Which electrons will experience more effective nuclear charge from the nucleus.
71. The bromine atom possesses 35 electrons. It contains 6 electrons in 2p orbital, 6 electrons in 3p orbital and 5 electrons in 4p orbital. Which of these electrons experiences the lowest effective nuclear charge ?
72. Draw the shapes of s, p, d & f orbitals.

3. Classification of elements and periodicity in properties

- Mendeleev's Periodic Law:** The physical and chemical properties of elements are the periodic function of their atomic masses.
- Modern Periodic law (Moseley's Periodic Law):** Physical and chemical properties of elements are the periodic functions of their atomic number. If the elements are arranged in order of their increasing atomic number, after a regular interval, elements with similar properties are repeated.
- Notation for IUPAC Nomenclature of Elements With Z > 100**

Digit	Name	Abbreviation
0	Nil	n
1	Un	u
2	Bi	b
3	Tri	t
4	quad	q
5	pent	p
6	Hex	H

7	sept	S
8	oct	O
9	enn	E

The latin words for various digits of the atomic number are written together in the order of which makes the atomic number and suffix 'ium' is added at the end. In case of bi and tri one 'i' is omitted.

- s-Block Elements** (s-block elements lie on the extreme left of the periodic table.)
 - Group 1 (alkali metals) & 2 (alkaline earth metals) elements constitute the s-block.
 - General electronic configuration is [inert gas] ns^{1-2}
- p-Block Elements** (P-block elements lie on the extreme right of the periodic table)
 - Group 13 to 18 elements constitute the p-block.
 - General electronic configuration is [inert gas] $ns^2 np^{1-6}$
- d-Block Elements**
 - Group 3 to 12 elements constitute the d-block.
 - General electronic configuration is [inert gas] $(n-1) d^{1-10} ns^{0-2}$
- f-Block Elements**
 - General electronic configuration is $(n-2) f^{1-14} (n-1) d^{0-1} ns^2$
 - All f-block elements belong to 3rd group. They are metals
- Prediction of period, group and block :**
 - Period of an element corresponds to the principal quantum number of the valence shell
 - The block of an element corresponds to the type of subshell which receives the last electron
 - The group is predicted from the number of electrons in the valence shell or/and penultimate shell as follows.
 - For **s-block** elements Group number = the no. of valence electrons
 - For **p-block** elements Group number = 10 + no. of valence electrons
 - For **d-Block** elements Group number = no. of electrons in $(n-1)$ d sub shell + no. of electrons in valence shell.
- Van der Waals' radius > Metallic radius > Covalent radius, $r_{\text{covalent}} < r_{\text{crystal}} < r_{\text{vander Walls}}$**
- Variation in Atomic Radius**

Variation in Atomic Radius in Period	Variation in Atomic Radius in Group
Atomic size decreases with increase in atomic number in a period from left to right	Atomic size atom increases from top to bottom in a given group.
Reason: because within the period the outer electrons are in the same valence shell and the effective nuclear charge increases. (Z_{eff} increases) $r_n \propto \frac{1}{Z^*}$	Reason: (a) principal quantum number, n increases (no. of orbits) (b) because the inner energy levels are filled with electrons, which serve to shield the outer electrons from the pull of the nucleus. The effect of increase number of atomic shells overweight the effect of increased screening effect. As a result of this the size of atom increases from top to bottom in a given group.

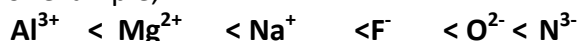
11. The **atomic radii of noble gases** are not considered here. Being monoatomic, their (non-bonded radii) values are very large. In fact radii of noble gases should be compared not with the covalent radii but with the van der Waals radii of other elements.
12. A **cation is smaller than its parent atom** because in a Cation, the number of positive charges on the nucleus is greater than the number of orbital electrons leading to increased inward pull of remaining electrons causing contraction in size of the ion. For example the atomic radius of sodium is 186 pm compared to the ionic radius of 95 pm for Na^+ .

For example :	Na	Na^+
Number of Electrons :	11	10
Number of Protons :	11	11

13. **The size of an anion will be larger than that of the parent atom** because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge. For example, the ionic radius of fluoride ion (F^-) is 136 pm whereas the atomic radius of fluorine is only 64 pm.

	Cl	Cl^-
Number of Electrons :	17	18
Number of Protons :	17	17

14. **Isoelectronic species** : The species containing the same number of electrons but differ in the magnitude of their nuclear charges are called as isoelectronic species. For example, N^{3-} , O^{2-} , F^- , Ne, Na^+ , Mg^{2+} and Al^{3+} are all isoelectronic species with same number of electrons (i.e. 10) but different nuclear charges of +7, +8, +9, +10, +11, +12 and +13 respectively
15. **Variation in ionic radii of isoelectronic species** Within a series of isoelectronic species as the nuclear charge increase, the force of attraction by the nucleus on the electrons also increases. As a result, the ionic radii of isoelectronic species decrease with increase with increases in the magnitude of nuclear charges. For example,



16. Ionization Enthalpy:

- It represents the energy required to remove an electron from an isolated gaseous atom (X) in its ground state. In other words, the first ionization enthalpy for an element X is the enthalpy change ($\Delta_i H$) for the reaction depicted in equation . $\text{X}(\text{g}) \rightarrow \text{X}^+(\text{g}) + \text{e}^-$.
- In general : $(\text{IE})_1 < (\text{IE})_2 < (\text{IE})_3 < \dots$ because, as the number of electrons decreases, the attraction between the nucleus and the remaining electrons increases considerably and hence subsequent I.E. (s) increase.

17. Factors Influencing Ionisation energy

- **Size of the Atom** : Ionisation energy decreases with increase in atomic size.
- **Nuclear Charge** : The ionisation energy increases with increase in the nuclear charge.
- **Shielding effect** : The larger the number of electrons in the inner shells, greater is the screening effect and smaller the force of attraction and thus (IE) decreases.
- **Penetration Effect of the Electron** : The ionisation energy increases as the penetration effect of the electrons increases. Within the same energy level, the penetration effect decreases in the order $S > P > d > f$ Greater the penetration effect of electrons more firmly the electron will be held by the nucleus and thus higher will be the ionisation energy of the atom.
- **Electronic Configuration** : If an atom has exactly half-filled or completely filled orbitals, then such an arrangement has extra stability. As noble gases have completely filled electronic configurations, they have highest ionisation energies in their respective periods.

18. Variation in Ionization Enthalpy

Variation in Ionization Enthalpy in Group	Variation in Ionization Enthalpy in Period
As we descend in a group the first ionization enthalpy generally decreases	The first ionization enthalpy generally increases as we go across a period.
Reason: With increase in atomic size, the distance	Reason: With increase in the effective nuclear

between the outermost electrons and the nucleus increases, the force of attraction between the valence shell electrons and the nucleus decreases. As a result, outer most electrons are held less firmly and lesser amount of energy is required to knock them out. thus ionisation energy decreases in a group from top to bottom.

charge, the electrons of the outer most shell are more firmly held by the nucleus and thus greater amount of energy is required to pull out an electron from the atom. Thus , ionisation energy increases as we move from left to right.

Exceptions :

Be(IE₁)>B(IE₁)

Be (Z = 4) $\underbrace{1s^2, 2s^2}_{\text{completely filled orbital more stable}}$ B (Z = 5) $\underbrace{1s^2, 2s^2, 2p^1}_{\text{Partially filled orbital less stable}}$

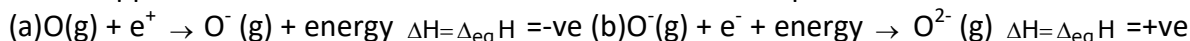
N(IE₁)>O(IE₁)

N (Z = 7) $1s^2 2s^2 2p^3$ O (Z = 8) $1s^2 2s^2 2p^4$
exactly half-filled orbitals more stable

19. **Electron Gain Enthalpy** :When an electron is added to a neutral gaseous atom (X) to convert it into a negative ion, the enthalpy change accompanying the process is defined as the electron gain enthalpy. Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion. $X(g) + e^- \rightarrow X^-(g)$; $\Delta H = \Delta_{eg}H$

20. When an electron is added to the atom and the energy is released, the electron gain enthalpy is negative and when energy is needed to add an electron to the atom, the electron gain enthalpy is positive.

21. The addition of second electron to an anion is opposed by electrostatic repulsion and hence the energy has to be supplied for the addition of second electron. For example :



22. Group 17 (halogens) have very high negative electron gain enthalpies because they can attain stable noble gas electronic configuration by picking up an electron.

23. Noble gases have large positive electron gain enthalpies because the electrons has to enter the next higher energy level leading to a very unstable electronic configuration.

24. **Variation in electron gain enthalpies in Group & period:**

Variation in electron gain enthalpies in Group	Variation in electron gain enthalpies in Period
<p>We should also expect electron gain enthalpy to become less negative as we go down a group because the size of the atom increases and the added electron would be farther from the nucleus. This is generally the case .</p> <p>Exceptions :</p> <p>Electron gain enthalpy of O or F is less than S or Cl. This is due to the fact that when an electron is added to O or F, the added electron goes to the smaller n = 2 energy level and experiences significant repulsion from the other electrons present in this level. In S or Cl, the electron goes to the larger n = 3 energy level and consequently occupies a larger region of space leading to much less electron repulsion.</p>	<p>As a general rule, electron gain enthalpy becomes more negative.</p> <p>Reason: because left to right across a period effective nuclear charge increases and consequently it will be easier to add an electron to a small atom.</p>

- **Factors Influencing electron gain enthalpies** (i) Electron gain enthalpy $\propto \frac{1}{\text{Atomic size}}$ (ii) Electron gain enthalpy \propto Effective nuclear charge (Z_{eff}) (iii) Electron gain enthalpy $\propto \frac{1}{\text{screening effect}}$ (iv) stability of

half filled and completely filled orbitals of a subshell is comparatively more and the addition of an extra electron to such a system is difficult and hence the electron affinity value decreases.

25. **Electronegativity:** Electronegativity is a measure of the tendency of an element to attract electrons towards itself in a covalent bonded molecules.

26. **VALENCE ELECTRONS:** The electrons present in outermost shell are called as valence electron. Because the electrons in the outermost shell determine the valency of an element.

27. ELECTROPOSITIVE OR METALLIC CHARACTER:

- **Periodicity:** In period- The electropositive or metallic characters decreases from left to right in a period.
- In group- The electropositive or metallic characters increases from top to bottom in a group.

28. ELECTRO-NEGATIVE OR NON- METALLIC CHARACTERS:

- In period- The electro-negative or non- metallic characters increases from left to right in a period.
- In group- The electro-negative or non-metallic characters decreases from top to bottom in a group.

29. REACTIVITY OF METALS: Periodicity:

- In period- The tendency of an element to lose electrons decreases in a period. So the reactivity of metals decreases from left to right in a period.
- In group- The tendency of an element to lose electrons increases in a period. So the reactivity of metals increases from top to bottom in a group.

30. REACTIVITY OF NON- METALS:

- (i) In period- The tendency of an element to gain electrons increases in a period. So the reactivity of non-metals increases from left to right in a period.
- (ii) In group- The tendency of an element to gain electrons decreases in a group. So the reactivity of non-metals increases from top to bottom in a group.

31. **Anomalous Properties of Second Period Elements:** The first element of each of the groups 1 (lithium) and 2 (beryllium) and groups 13-17 (boron to fluorine) differs in many respects from the other members of their respective group. For example, lithium unlike other alkali metals, and beryllium unlike other alkaline earth metals, form compounds with pronounced covalent character; the other members of these groups predominantly form ionic compounds. In fact the behaviour of lithium and beryllium is more similar with the second element of the Group 1, 2, 13, 14, 15, 16, 17. following group i.e., magnesium and aluminium, respectively. This sort of similarity is commonly referred to as **diagonal relationship** in the periodic properties.

Answer the following :

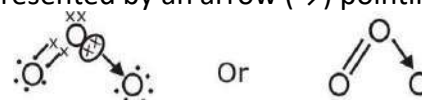
1. What is the basic difference in approach between the Mendeleev's Periodic Law and the Modern Periodic Law?
2. What would be the IUPAC name and symbol for the elements with atomic number 101,107,109,110 &120?
3. The element with atomic number 119 has not been discovered. What would be the IUPAC name and symbol of this element? Also predict the electronic configuration, group and period of this element.
4. How is group, period and block of elements predicted.
5. Give the electronic configuration and in terms of period group and block where would you locate the elements with $Z = 17, 19, 24, 26, 29, 31, 34, 38, 40, 51, 55, 114, 117$ and 120 ?
6. How would you justify the presence of 18 elements in the 5th period of the Periodic Table?
7. On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.
8. Write the general outer electronic configuration s, p, d, & f block elements.
9. Why do elements in the same group have similar physical and chemical properties?
10. How is atomic radius expressed in terms of different types of radius?
11. How does the atomic radius vary in a period and in a group? How do you explain the variation?
12. Why atomic radius of noble gases are not considered in comparison.
13. Explain why cations are smaller and anions larger in radii than their parent atom?
14. Describe the theory associated with the radius of an atom as it (a) gains an electron (b) Loses an electron
15. What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions. (i) F^- (ii) Ar (iii) Mg^{+2} (iv) Rb^+

16. Consider the following species: N^{3-} , O^{2-} , F^- , Na^+ , Mg^{+2} , and Al^{+3} (a) What is common in them? (b) Arrange them in the order of increasing ionic radii.
17. Arrange the following in order of increasing radii: (a) N, O, P (b) F, Cl, Br (c) I^- , I^+
18. Which of the following species will have the largest and the smallest size Mg, Mg^{2+} , Al, Al^{3+} .
19. Pick out among the following, Li^+ , Al^{3+} , K^+ , Mg^{2+} - species having smallest ionic radius
20. A student reported the radii of Cu, Cu^+ and Cu^{+2} as 96, 122 & 72 ppm respectively do you agree with results.
21. Define Ionization enthalpy. Give its unit. Explain the term successive ionization energies.
22. How does the Ionization vary in a period and in a group? How do you explain the variation?
23. Among the second period elements the actual ionization enthalpies are in the order $\text{Li} < \text{B} < \text{Be} < \text{C} < \text{O} < \text{N} < \text{F} < \text{Ne}$ Explain why (i) Be has higher ionization enthalpy than B. (ii) O has lower ionization enthalpy than N and F?
24. How would you explain the fact the first ionization enthalpy of sodium is lower than that of Magnesium but its second ionization enthalpy is higher than that of Magnesium?
25. What are the various factors due to which the ionization enthalpy of the main group elements tends to decrease down a group?
26. Arrange C, N, O and F in the decreasing order of their second ionization enthalpy and explain briefly.
27. Explain why the first ionization enthalpy of Carbon is more than that of Boron but the reverse is true for second ionization enthalpy
28. Would you expect the first ionization enthalpy for two isotopes of the same element to be same or different? Justify your answer.
29. Define the term electron gain enthalpy.
30. How does the electron gain enthalpy in a period and in a group? How do you explain the variation?
31. Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first? Justify your answer?
32. Which of the following pairs of elements would have a more negative electron gain enthalpy? (i) O or F (ii) F or Cl (iii) O or S. Give reason to support your answer
33. Which of the following will have the most negative electron gain enthalpy and which the least negative? P, S, Cl, F. Explain your answer.
34. In each of the following sets, arrange the elements in the increasing order of their negative electron gain enthalpies: (i) C, N, O (ii) O, N, S (iii) Cl, S, Ar (iv) F, Cl, Br
35. Which element do you think would have been named by
(i) Lawrence Berkeley Laboratory (ii) Seaborg's group?
36. What is the difference between the terms electron gain enthalpy and electronegativity?
37. Use the periodic table to answer the following questions.
a) Identify an element with five electrons in the outer subshell. (b) Identify an element that would tend to lose two electrons. (c) Identify an element that would tend to gain two electrons. (d) Identify the group having metal, non-metal, liquid as well as gas at the room temperature.
38. Assign the position of the element having outer electronic configuration (i) ns^2np^4 for $n=3$ (ii) $(n-1)d^2 ns^2$ (iii) $(n-2)f^7 (n-1)d^1 ns^2$ for $n=6$ in the periodic table.
39. Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements.
a) Lithium and oxygen (b) Magnesium and nitrogen (c) Aluminium and iodine (d) Silicon and oxygen (e) Phosphorus and fluorine (f) Element 71 and fluorine (g) silicon and bromine (h) aluminium and sulphur.
40. Considering the elements B, Al, Mg, and K, the correct order of their metallic character is 11. Considering the elements B, C, N, F, and Si, the correct order of their non-metallic character is:
41. Show by a chemical reaction with water that Na_2O is a basic oxide and Cl_2O_7 is an acidic oxide.
42. Considering the atomic number and position in the periodic table, arrange the following elements in the increasing order of metallic character : Si, Be, Mg, Na, P.

4. Chemical bonding and Molecular Structure

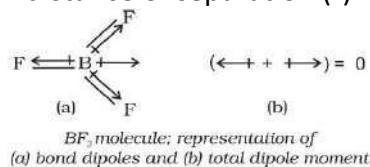
1. **Lewis dot structures** are a shorthand to represent the valence electrons of an atom. The structures are written as the element symbol surrounded by dots that represent the valence electrons.
2. **Electrovalent bond or Ionic Bond**: The chemical bond as result of transfer of electron from one atom (electropositive) to another atom (electronegative). Ionic bonds will be formed more easily between elements with comparatively low ionization enthalpies and elements with comparatively high negative value of electron gain enthalpy. Most ionic compounds have cations derived from metallic elements and anions from non-metallic elements.
3. **Electrovalency**: The number of electrons lost or gain by an atom of an element is called as electrovalency. The element which give up electrons to form positive ions are said to have positive valency, while the elements which accept electrons to form negative ions are said to have negative valency
4. **Formation of Ionic Bond** : It is favoured by, (i) the low ionisation enthalpy of a metallic element which forms the cations, (ii) High electron gain enthalpy of non- metallic element which forms the anions, (iii) Large lattice enthalpy i.e; the smaller size and the higher charge of the atoms
5. **Covalent Bonds**- The bond formed between two atoms by mutual sharing of electrons between them so as to complete their octets or duplets.
6. **COVALENCY**:The number of electrons which an atom contributes towards mutual sharing during the formation of a chemical bond called its covalency in that compound.
7. **SINGLE COVALENT BOND**: A covalent bond formed by the mutual sharing of one pair of electrons is called a single covalent bond, or simply a single bond. A single covalent bond is represented by a small line (–) between the two atoms. H_2
8. **DOUBLE COVALENT BOND**: A covalent bond formed by the mutual sharing of two pair of electrons is called a double covalent bond, or simply a double bond. A double covalent bond is represented by two small horizontal lines (=) between the two atoms. E.g. $O=O$, $O=C=O$ etc.
9. **TRIPLE COVALENT BOND**: A covalent bond formed by the mutual sharing of three pair of electrons is called a triple covalent bond, or simply a triple bond. A triple covalent bond is represented by three small horizontal lines (\equiv) between the two atoms. E.g. $N\equiv N$, $H-C\equiv C-H$ etc.
10. **Octet Rule**- Kossel and Lewis in 1916 developed an important theory of chemical combination between atoms known as electronic theory of chemical bonding. According to this, atoms can combine either by transfer of valence electrons from one atom to another or by sharing of valence electrons in order to attain their octet. This is known as octet rule.
11. **Limitations of octet rule-**
 - a) Incomplete octet of the central atom: In some compounds the number of electrons surrounding the central atom is less than eight. This is especially the case with elements having less than four valence electrons. Examples- $LiCl$, $BeCl_2$, BCl_3
 - b) Odd-electron molecules: In molecules with an odd number of electrons like nitric oxide, NO and nitrogen dioxide, the octet rule is not satisfied for all the atoms.
 - c) The expanded octet : Elements in and beyond the third period of the periodic table have, apart from 3s and 3p orbitals, 3d orbitals also available for bonding. In a number of compounds of these elements there are more than eight valence electrons around the central atom. This is termed as the expanded octet. Some of examples of such compounds are PF_5 , SF_6 .
 - d) This theory does not account for the shape of molecules.
12. **POLAR COVALENT BOND**: The bond between two unlike atoms which differ in their affinities for electrons is said to be polar covalent bond. E.g. $H-Cl$
13. **COORDINATE BOND**: The bond formed when one sided sharing of electrons take place is called a coordinate bond. Such a bond is also known as dative bond. It is represented by an arrow (\rightarrow) pointing

towards the acceptor atom. Example $:O_3$ (ozone)



14. **Bond length** is defined as the equilibrium distance between the nuclei of two bonded atoms in a molecule.
15. **Bond Angle**: It is defined as the angle between the orbital containing bonding electron pairs around the central atom in a molecule/complex ion. It gives some idea regarding the distribution of orbital around the central atom in a molecule/complex ion and hence it helps us in determining its shape
16. **Bond enthalpy**: It is defined as the amount of energy required to break one mole of bonds of a particular type between two atoms in a gaseous state. The unit of bond enthalpy is kJ mol^{-1}
17. **Bond Order** :The Bond Order is given by the number of bonds between the two atoms in a molecule. E.g.: Bond Order Of $\text{O}_2 = 2$. With increase in bond order, bond enthalpy increases and bond length decreases.
18. **Resonance**: According to the concept of resonance, whenever a single Lewis structure cannot describe a molecule accurately, a number of structures with similar energy, positions of nuclei, bonding and the non-bonding pairs of electrons are taken as the canonical structures of the hybrid which describes the molecule accurately.
19. **Dipole moment** : As a result of polarization, the molecule possesses the dipole moment which can be defined as the product of charge and the distance between the centers of positive and negative charge. It is usually designated by a Greek letter ' μ '. Mathematically, it is expressed as follows:

$$\text{Dipole moment } (\mu) = \text{charge } (Q) \times \text{distance of separation } (r)$$



20. **Sigma (σ) bond** : This type of covalent bond is formed by the end to end (hand-on) overlap of bonding orbitals along the internuclear axis. This is called as head on overlap or axial overlap.
21. **pi (π) bond** : In the formation of π bond the atomic orbitals overlap in such a way that their axes remain parallel to each other and perpendicular to the internuclear axis. The orbitals formed due to sidewise overlapping consists of two saucer type charged clouds above and below the plane of the participating atoms.
22. In case of sigma bond, the overlapping of orbitals takes place to a larger extent. Hence, it is stronger as compared to the pi bond where the extent of overlapping occurs to a smaller extent.
23. **VSEPR theory**
- The shape of a molecule depends upon the number of valence shell electron pairs (bonded or nonbonded) around the central atom.
 - The repulsive interaction of electron pairs decrease in the order: Lone pair (lp) – Lone pair (lp) > Lone pair (lp) – Bond pair (bp) > Bond pair (bp) – Bond pair (bp)

24. **Geometry of Molecules** on the basis of VSEPR Theory

Bond pair	Lone pair	Shape	Eg.
2	0	Linear	BeCl_2
3	0	Trigonal Planar	BCl_3
2	1	Bent	SO_2
4	0	Tetrahedral	CH_4
3	1	Pyramidal	NH_3 , PH_3
2	2	V-shape	H_2O

5	0	Trigonal bipyramidal	PCl_5
4	1	See saw	SF_4
3	2	T-shaped	ClF_3 ,
2	3	Linear	XeF_2
6	0	Octahedral	SF_6
5	1	Square pyramidal	ClF_5 , IF_5
4	2	Square planar	XeF_4 ,
6	1	Distorted Octahedral	XeF_6

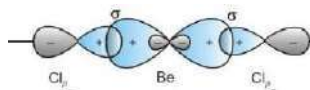
24. **Hybridization:** It can be defined as the process of intermixing of the orbitals of slightly different energies so as to redistribute their energies, resulting in the formations of new set of orbitals of equivalent energies and shape.

• **Salient Features of hybridization :**

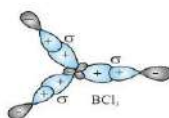
- The number of hybrid orbitals is equal to the number of the atomic orbitals that get hybridised.
- The hybridised orbitals are always equivalent in energy and shape.

25. Types of Hybridisation

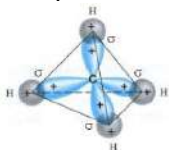
a) **sp hybridisation:** This type of hybridisation involves the mixing of one s and one p orbital resulting in the formation of two equivalent sp hybrid orbitals. Each sp hybrid orbital has 50% s-character and 50% p-character. Such a molecule in which the central atom is sp-hybridised and linked directly to two other central atoms possesses linear geometry. Example BeCl₂ molecule



b) **sp² hybridisation :** In this hybridization there is involvement of one s and two p-orbitals in order to form three equivalent sp² hybridised orbitals. Each sp² hybrid orbital has 33.33% s-character and 66.66 % p-character. Such a molecule in which the central atom is sp² hybridised and linked directly to three other central atoms possesses trigonal planar geometry. Example BCl₃ molecule.

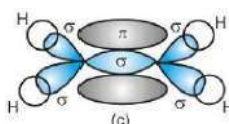


c) **sp³ hybridisation:** In this type of hybridisation there is mixing of one s-orbital and three p-orbitals of the valence shell to form four sp³ hybrid orbital of equivalent energies and shape. There is 25% s-character and 75% pcharacter in each Sp³ hybrid orbital. The four Sp³ hybrid orbitals so formed are directed towards the four corners of the tetrahedron Example CH₄ molecule

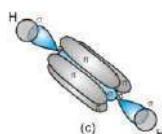


26. sp³ hybridisation in C₂H₆ molecule

27. sp² hybridisation in C₂H₄ molecule

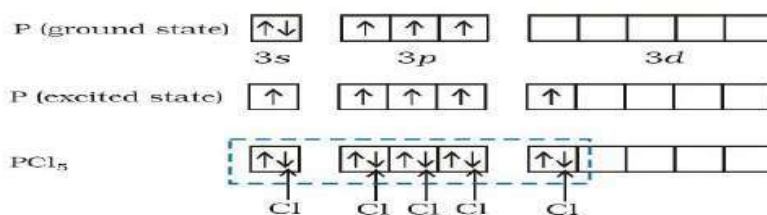


28. sp hybridisation in C₂H₂ molecule



29. Formation of PCl₅

(i) **Formation of PCl₅ (sp³d hybridisation):** The ground state and the excited state outer electronic configurations of phosphorus (Z=15) are represented below.



sp³d hybrid orbitals filled by electron pairs donated by five Cl atoms.

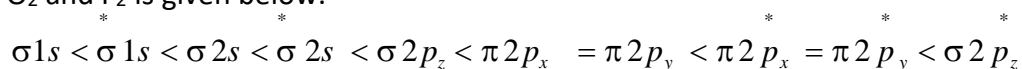
30. Formation of SF₆: Do Your self

31. **Molecular orbital.** It gives electron probability distribution around a group of nuclei in a molecule. They are filled in the same way as atomic orbitals. Molecular orbitals are formed by linear combination of atomic orbitals.
32. **Sigma (σ) molecular orbitals.** A molecular orbital which is formed from the overlap of two s atomic orbitals or head to head overlap of one s and p-atomic orbitals or head to head overlap of two p-atomic orbitals, is known as sigma molecular orbital.
33. **pi (π) molecular orbitals.** A molecular orbital which is formed by lateral overlap of two parallel p-orbitals is known as pi (π) molecular orbital.
34. **Bonding molecular orbital.** A molecular orbital that is formed by addition overlap (i.e., when the lobes of atomic orbitals overlap with the same sign) of two atomic orbitals is known as bonding molecular orbital. It is represented as $\Psi_{MO} = \Psi_A + \Psi_B$ Its energy is lower than the atomic orbitals from which it is formed. It favours bonding.
35. **Anti-bonding molecular orbital.** A molecular orbital that is obtained by the subtraction overlap (i.e., when the lobes of atomic orbitals overlap with the opposite sign) of two atomic orbitals is known as anti-bonding molecular orbital. It is represented as $\Psi_{MO}^* = \Psi_A - \Psi_B$ Its energy is higher than the atomic orbitals from which it is formed. It does not favour bonding.
36. **Bond order.** It is defined as half of the difference between number of electrons in bonding and anti-bonding orbitals, i.e., $B.O. = \frac{1}{2} (N_b - N_a)$ 'where N_b are number of electrons in bonding orbitals' and N_a are number of electrons in anti-bonding orbitals. Bond order helps in estimating stability of atom.

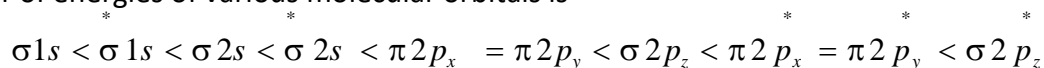
37. Significance of Bond Order

- **It bond order = 0, it means species does not exist.**
- **Bond order of 1, 2 & 3 corresponds to a single bond, double & triple bond respectively**
- **Bond order increases \uparrow stability of molecule increases**
 \uparrow **Bond length decreases \downarrow**
- **If the molecule has one or more unpaired electron, it will be paramagnetic, while if all the electrons are paired it will be diamagnetic.**

38. **Energy level Diagrams for Molecular Orbitals.** The increasing order of energies of various molecular orbitals for O₂ and F₂ is given below.

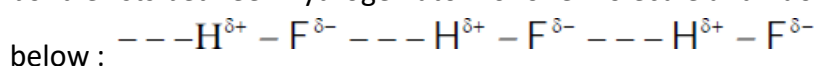


However, this sequence of energy levels of molecular orbitals is not correct for remaining molecules Li₂, Be₂, B₂, C₂, N₂. For instance, it has been observed experimentally that for molecules such as B₂, C₂, N₂ etc., the increasing order of energies of various molecular orbitals is

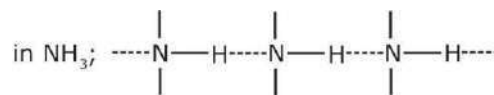
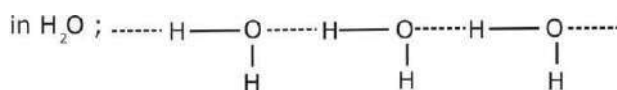


The important characteristic feature of this order is that the energy of $\overset{*}{\sigma}2p_z$ **molecular orbital is higher than that of $\overset{*}{\pi}2p_x$ and $\overset{*}{\pi}2p_y$ molecular orbitals** in these molecules.

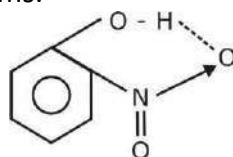
39. **Hydrogen bond** can be defined as the attractive force which binds hydrogen atom of one molecule with the electronegative atom (F, O or N) of another molecule. **For example**, in HF molecule, the hydrogen bond exists between hydrogen atom of one molecule and fluorine atom of another molecule as depicted below:



40. **Intermolecular hydrogen bond** : It is formed between two different molecules of the same or different compounds. For example, H-bond in case of HF molecule, alcohol or water molecules, etc.



41. **Intramolecular hydrogen bond** : It is formed when hydrogen atom is in between the two highly electronegative (F, O, N) atoms present within the same molecule. For example, in o-nitrophenol the hydrogen is in between the two oxygen atoms.



Answer the following :

- Define Covalent Bond. Explain its types with examples.
- Draw the Lewis structures for the following molecules & ions: H_2S , SiCl_4 , BeF_2 , CO_3^{2-} , BeCl_2 , BCl_3 , SiCl_4 , AsF_5 , H_2S , PH_3 , PCl_5 , SF_6 , NH_3 , SF_4 , ClF_3 , BrF_5 , XeF_4 , NO_3^-**
- Explain some compounds which are exceptions to Octet rule.
- What is Formal Charge? Calculate the formal charge on all elements of O_3 & CO_3^{2-} .**
- Define Electrovalent Bond or Ionic Bond. Write the favorable factors for the formation of ionic bond.
- Is CaF_2 linear or bent or neither of the two ? Justify.**
- Use Lewis symbols to show electron transfer between the following atoms to form Cations and anions: (a) K and S (b) Ca and O (c) Al and N (d) Al and O
- Define octet rule. Write its significance and limitations.
- What is Coordinate Bond or Dative Bond.
- Give the main points of VSEPR theory
- Discuss the shape using the VSEPR model: H_2S , SiCl_4 , BeF_2 , CO_3^{2-} , BeCl_2 , BCl_3 , SiCl_4 , AsF_5 , H_2S , PH_3 , PCl_5 , SF_6 , NH_3 , SF_4 , ClF_3 , BrF_5 , XeF_4 , NO_3^-**
- Draw the structure of the following : (i) BrF_3 (ii) XeO_3**
- CH_4 , NH_3 and H_2O contain same number of electrons but their shapes are different ?
- Which of the following has maximum bond angle? H_2O , CO_2 , NH_3 , CH_4 ?
- Draw an orbital diagram showing the formation of bonds in C_2H_4 .**
- Draw and predict the shape of (i) ClF_3 molecule (ii) PCl_5**
- Although geometries of NH_3 and H_2O molecules are distorted tetrahedral, bond angle in water is less than that of ammonia. Discuss.**
- Define Resonance. Explain the structure of CO_3^{2-} ion in terms of resonance.**
- Write the resonance structures for SO_3 , NO_2 and NO_3^- , CO_2
- Explain the Dipole moment. How it is helpful in predicting polar & Non polar nature of compounds.
- Explain why BeH_2 molecule has zero dipole moment although the B-H bonds are polar.**
- Which out of NH_3 and NF_3 has more dipole moment and why?**
- Arrange the bonds in order of increasing ionic character in the molecules: LiF , K_2O , N_2 , SO_2 and ClF_3 .**
- Explain the formation of Hydrogen molecule on basis of Valence Bond Theory.
- Distinguish between a sigma (σ) and a pi (π) bond**
- What is the total number of sigma and pi bonds in the following molecules? (a) C_2H_2 (b) C_2H_4 .**
- Name the type of hybridization of each C atom in a molecule of (i) propene (ii) propyne. How many σ and π -bonds are present in each case?**
- What is meant by hybridization of atomic orbitals?**
- Describe sp , sp^2 , sp^3 hybrid orbitals using suitable examples.**
- Describe the change in hybridization (if any) of the Al atom in the following reaction:**
 $\text{AlCl}_3 + \text{Cl}^- \rightarrow \text{AlCl}_4^-$
- Is there any change in the hybridization of B and N atoms as a result of the following reaction:**
 $\text{BF}_3 + \text{NH}_3 \rightarrow \text{F}_3\text{B.NH}_3$
- Predict the hybrid state of central atom in the following compounds: H_2S , SiCl_4 , BeF_2 , CO_3^{2-} , BeCl_2 , BCl_3 , SiCl_4 , AsF_5 , H_2S , PH_3 , PCl_5 , SF_6 , NH_3 , SF_4 , ClF_3 , BrF_5 , XeF_4 .**

- 33. Explain the concept of hybridization in PCl_5 . Why are axial bonds longer as compared to equatorial bonds in PCl_5 .**
34. Although both CO_2 and H_2O are tri-atomic molecules, the shape of H_2O molecule is bent while that of CO_2 is linear.
35. How will you justify identical nature of all the C-O bonds in CO_3^{2-} ion?
- 36. Explain the concept of hybridization in SF_6 .**
- 37. Which hybrid orbitals are used by carbon atoms in the following molecules? (a) $\text{CH}_3\text{-CH}_3$ (b) $\text{CH}_3\text{-CH=CH}_2$ (c) $\text{CH}_3\text{-CH}_2\text{-OH}$ (d) $\text{CH}_3\text{-CHO}$ (e) $\text{CH}_3\text{-COOH}$ (f) $\text{H}_2\text{C=CH-CH}_2\text{-C}\equiv\text{CH}$**
38. Write the conditions for the combination of atomic orbitals.
39. Assuming Z-axis as molecular axis, label the molecular orbitals formed by the following combination of atomic orbitals :
- (i) $1s + 1s$ (ii) $2p_y - 2p_y$ (iii) $2p_z + 2p_z$ (iv) $2s + 2s$ (v) $2p_x + 2p_x$**
40. Distinguish between bonding molecular orbital & anti bonding molecular orbital.
- 41. Write the molecular orbital configuration and energy diagram for (i) O_2^+ , O_2 , O_2^- , O_2^{2-} (ii) N_2 , N_2^+ , N_2^- (iii) Be_2 , H_2 , C_2**
42. Draw the energy diagram for H_2 , Be_2 , N_2 , & O_2 .
- 43. When a magnet is dipped in a jar of liquid O_2 , some O_2 clings to it. Why ?**
- 44. What is meant by the term bond order. Write the significance of bond order**
- 45. What information does MOT provide for O_2 , O_2^+ and O_2^- molecular species with regards to :**
- Bond dissociation energy
 - Bond length.
- 46. Compare the relative Stability of the following species and indicate their magnetic properties. (i) O_2^+ , O_2 , O_2^- , O_2^{2-} (ii) N_2 , N_2^+ , N_2^-**
- 47. Use the molecular orbital theory to explain why Be_2 molecule does not exist.**
- 48. Why the He_2 molecule does not exist, explain on the basis of molecular orbital theory.**
- 49. Define hydrogen bond .Explain its types with suitable examples.**
50. N and Cl have the same electronegativity, H-bonding is present in NH_3 but not in HCl explain ?
51. Assign suitable reason for HF is liquid while HCl is gas ?
- 52. Assign suitable reasons for the following :**
- H_2O is liquid at room temperature while H_2S is gas.**
 - o-nitrophenol is steam volatile while p-nitrophenol is not.**
53. Is hydrogen bond weaker or stronger than the van der Waals forces?
- . Give reason (i) All P-Cl bonds in PCl_5 molecule are not equal. (ii) N_2 is less reactive than O_2 (iii) BF_3 is non polar while NH_3 is a polar molecule.