ESE 2027

UPSC ENGINEERING SERVICES EXAMINATION

PRELIMINARY EXAMINATION

General Studies & Engineering Aptitude

Basics of Material Science and Engineering

Comprehensive Theory with **Practice Questions** and **ESE Prelims Solved Questions**

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Preface

B. Singh (Ex. IES)

The book **Basics of Material Science and Engineering** has been compiled with the objective of providing a concise yet comprehensive resource for aspirants of the Engineering Services Examination (ESE). It is specifically designed to help students develop a clear understanding of the concepts covered under the General Studies and Engineering Aptitude section of ESE.

This textbook offers detailed and descriptive, step-by-step theoretical explanations presented in lucid and easy-to-understand language. It ensures thorough coverage of fundamental concepts along with well-structured objective-type questions. The concise and systematic presentation enables readers to grasp the subject with clarity and apply their knowledge effectively to solve objective questions with speed and accuracy.

The book covers the complete ESE syllabus in a holistic manner and is equally useful for other competitive examinations. Each topic has been given due emphasis so that a careful reading of the text itself is sufficient to build strong conceptual clarity.

Special effort has been made to include solved practice questions and previous years' solved questions from the General Studies and Engineering Aptitude (ESE Prelims) section. These solutions are explained in a clear, step-by-step manner to strengthen problem-solving skills. At the end of every chapter, practice question sets with answer keys and detailed explanations are provided to help readers assess their understanding and enhance their exam readiness.

We have made sincere efforts to ensure accuracy and eliminate errors. However, we would greatly appreciate it if readers bring to our notice any printing or conceptual errors.

Though it is not possible to acknowledge everyone individually, we extend our heartfelt gratitude to all the authors, editors, and reviewers whose valuable contributions made this publication possible.

With Best Wishes
B. Singh (Ex. IES)
CMD, MADE EASY Group

Basics of Material Science and Engineering

Contents

Chapter 1

Introduction 1

1.1 Historical Perspective 1
1.2 Material Science 1
1.3 Classification of Materials 2
1.4 What is Material? 3
1.5 Engineering Needs of Material 3

Chapter 2

Chemical Bonding 5

2.1 Basic Laws of Chemistry 5
2.2 Fundamental Concepts 6
2.3 Electrons in Atoms 7
2.4 The Periodic Table 7
2.5 Comparison of Alpha (a), Beta (b) and Gamma (g) Rays 8
2.6 Quantum Number 9
2.7 Electron Affinity 11
2.8 Electronegativity 11
2.9 Pauli's Exclusion Principle 11
2.10 Auf-bau Principle 12
2.11 Hund's Rule 12
2.12 Heisenberg Uncertainty Principle 12
2.13 Gay Lussac's Law of Gaseous Volumes 13
2.14 Dalton's Atomic Theory 13
2.15 Rutherford Model 13
2.16 Bohr Model 13
2.17 Sommerfeld's Model 14
2.18 De Broglie Wave Equation 14
2.19 Octet Rule 14
2.20 Boyle's Law (Pressure-Volume Relationship) 15
2.21 Charles Law (Temperature-Volume Relationship) 15
2.22 Gay-Lussac's Law (Pressure-Temperature Relationship) 16
2.23 Avogadro Law (Volume-Amount Relationship) 16
2.24 Ideal Gas Equation 17
2.25 Dalton's Law of Partial Pressures 17
2.26 Chemical Bonding 17
2.27 Ionic Bond 18
2.28 Covalent Bond 19
2.29 Metallic Bond 20
2.30 Comparison of Primary Bonds 21
2.31 Van der Waal Bond 21
2.32 Dispersion Bonds 22
2.33 Dipole Bonds 22
2.34 Hydrogen Bonds 23
2.35 Properties of Water (H2O) 23
2.36 Directional Bond 25
2.37 Non-directional Bond 25
2.38 Molecular Orbital Theory 25
2.39 Hybridization 26
Objective Brain Teasers 30

Chapter 3

Crystallography 32

3.1 Introduction 32
3.2 Comparison of Crystalline and Noncrystalline Solids 32
3.3 Lattice Points, Space Lattice and Crystal Structures 33
3.4 Unit Cell and Primitive Unit Cell 34
3.5 Bravais Lattices 34
3.6 Cubic Crystal Structures 36
3.7 Hexagonal Closed Packing 39
3.8 Miller Indices 40
3.9 Interplanar Spacing 41
3.10 Crystal Imperfections 41
3.11 Point Imperfections 42
3.12 Line Defect 44
3.13 Surface Defect 46
3.14 Ionic Crystal Structure 47
3.15 Bragg's Law 48
3.16 Polymorphism and Allotropy 49
Objective Brain Teasers 55

Chapter 4

Electric Properties of Materials 60

4.1 Introduction 60
4.2 Ohm's Law and Electrical Conductivity 61
4.3 Energy Band Structure in Solids.................... 63
4.4 Classification of Materials based upon Energy Band Diagram 65
4.5 Electrical Resistivity of Metals........................ 66
4.6 Thermal Conductivity of Metals–Wiedemann Franz law........................ 67
4.7 Thermoelectric Phenomenon........................ 68
4.8 Insulators........................ 68
4.9 Dielectrics........................ 70
4.10 Electric Dipole Moment and Polarization... 72
4.11 Types of Polarization 73
4.12 Phase Difference and Dielectric Loss 74
4.13 Polar Molecules........................ 75
4.14 Nonpolar Materials 75
4.15 Other Electrical Characteristics of Materials75
4.16 Use of Dielectrics........................ 78
4.17 Semiconductor Materials 78
4.18 Electrons and Holes in an Intrinsic Semiconductor (Pure Semiconductor)........ 79
4.19 Extrinsic Materials 79
4.20 Charge Densities in a Semiconductor 80
4.21 Electrical Properties of Semiconductors..... 81
4.22 Hall Effect 82
4.23 Thermistors........................ 82
4.24 Photoconductors........................ 83
Objective Brain Teasers 85

Chapter 5

Magnetic Properties of Materials.................... 89

5.1 Introduction 89
5.2 Magnetic Parameters........................ 89
5.3 Classification of Magnetic Materials............ 90
5.4 Curie Temperature 95
5.5 Laws of Magnetic Materials 96
5.6 Domain Theory........................ 97
5.7 Magnetisation Curve and Magnetic Hysteresis Loop 98
5.8 Soft Magnetic Materials 99
5.9 Hard Magnetic Materials 101
5.10 Magnetic Storage........................ 103
5.11 Superconductivity........................ 103
Objective Brain Teasers 107

Chapter 6

Ceramics 110

6.1 Introduction 110
6.2 Silicate Ceramics........................ 111
6.3 Ceramics Used in Electrical Applications.. 113
6.4 Stress-Strain Behavior of Ceramics 114
6.5 Properties of Ceramics 115
6.6 Glass 116
6.7 Piezoelectric Ceramics........................ 117
6.8 Cement........................ 117
6.9 Carbon........................ 117
Objective Brain Teasers 121

Chapter 7

Polymers........................ 124

7.1 Introduction 124
7.2 Basic Definitions 124
7.3 General Characteristics of Polymer............ 126
7.4 Molecular Structure of Polymers 126
7.5 Different Types of Polymerizations Reactions 127
7.6 Co-polymer........................ 128
7.7 Classification of Plastics 128
7.8 Thermoplastic Materials 129
7.9 Thermosetting Materials 131
7.10 Mechanical Behaviour of Plastics 134
7.11 Compounding Materials........................ 134
7.12 Comparison of Polymers with Ceramics and Metals 135
7.13 Elastomer 135
7.14 Fibre 135
7.15 Coatings........................ 136
7.16 Vulcanization 136
7.17 Stress-Strain Behavior of Polymers 137
7.18 Advanced Polymeric Materials.................... 138
7.19 Crystallization of Polymers........................ 138
Objective Brain Teasers 142

Chapter 8

Composites 146

8.1 Introduction 146
8.2 General Characteristics of Composites...... 146
8.3 Natural Composites........................ 148

8.4 Particle-Reinforced Composites ... 148
8.5 Fibre-Reinforced Composites ... 150
8.6 Laminar Composites ... 151
8.7 Polymer-Matrix Composites (PMCs) ... 152
8.8 Fibre Phase ... 153
8.9 Matrix Phase ... 154
Objective Brain Teasers ... 155

Chapter 9

Mechanical Properties of Materials ... 156

9.1 Introduction ... 156
9.2 Normal Stress ... 156
9.3 Strain ... 157
9.4 Tension Test for Mild Steel ... 157
9.5 Common Terms of Mechanical Properties ... 159
9.6 Fracture ... 165
9.7 Hooke's Law ... 165
9.8 Elastic Constants ... 166
9.9 Difference between Linearly and Non-linearly Elastic Materials ... 167
9.10 Fatigue ... 167
Objective Brain Teasers ... 172

Chapter 10

Ferrous Metals ... 175

10.1 Introduction ... 175
10.2 Pig Iron ... 175
10.3 Cast Iron ... 176
10.4 Wrought Iron ... 179
10.5 Steel ... 180
10.6 Special Alloys Steels ... 183
Objective Brain Teasers ... 187

Chapter 11

Non-Ferrous Metals and Alloys ... 191

11.1 Introduction ... 191
11.2 Aluminium ... 192
11.3 Aluminium Alloys ... 193
11.4 Copper ... 194
11.5 Lead ... 198
11.6 Tin ... 198
11.7 Nickel ... 199
11.8 Magnesium ... 199
11.9 Titanium ... 199
11.10 Tungsten ... 200
11.11 Some Special Alloys ... 200
11.12 Babbit Metals ... 201
11.13 Solder Material ... 201
11.14 Refractory Materials ... 201
11.15 Super Alloys ... 202
Objective Brain Teasers ... 204

Chapter 12

Introduction to Nanomaterial and Metamaterial ... 208

12.1 Nano Technology ... 208
12.2 Meta Materials ... 210
Objective Brain Teasers ... 211

Chapter 13

Basics of Corrosion, Creep & Heat Treatment ... 213

13.1 Corrosion ... 213
13.2 Self Potential ... 213
13.3 Polarization ... 213
13.4 Oxidizing Power Vs Corrosion Rate ... 214
13.5 Types of Corrosion ... 214
13.6 Corrosion Penetration Rate (CPR) ... 215
13.7 Creep Curve ... 216
13.8 Recovery, Recrystallisation and Grain Growth ... 216
13.9 Difference between hot working and cold working ... 217
13.10 Strengthening Mechanisms of Materials ... 217
13.11 Hot Shortness or Sulphur-Embrittlement ... 218
Objective Brain Teasers ... 220

■■■■

2 Chemical Bonding

2.1 Basic Laws of Chemistry

Atomic mass unit

- It is defined as a mass exactly equal to one-twelfth the mass of one carbon - 12 atom.

$$1 \text{ amu} = 1.66056 \times 10^{-24} \text{g}$$

Mole

- The mole is unit of measurement for mass of substance in the International System of Units (SI).
- "One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the ^{12}C isotope".

$$1 \text{ mole} = 6.023 \times 10^{23} \text{ atoms}$$

- The mass of one mole of a substance in grams is called molar mass.

Avogadro number

- In 1811, Avogadro proposed that equal volumes of gases at the same temperature and pressure should contain equal number of molecules. The number of entities or atoms in a mole is termed as Avogadro number, denoted by N_A.

$$N_A = 6.023 \times 10^{23} \text{ atoms/mol}$$

Law of conservation of mass

Antoine Lavoisier in 1789 stated that:

- "Matter can neither be created nor destroyed by chemical reactions or physical transformations."
- He performed careful experimental studies for combustion reactions for reaching to the above conclusion.
- According to this in a closed system for chemical reaction, the mass of the products must be equal to the mass of the reactants.

Law of Definite Proportions

A French chemist, Joseph Proust stated that:

- "A given compound always consist exactly the fixed and same proportion of elements by mass." It is sometimes also referred as Law of definite composition.
- An example is CO_2. This gas is produced from a variety of reactions, often by the burning of materials, wood or fossil fuels. The structure of the gas always consists of one atom of carbon and two atoms of oxygen.

Law of Multiple Proportions

John Dalton proposed this law in 1803, he states that:

- When two elements combine to form more than one compound, the mass of one element that combines with a fixed mass of the other element, will always be in a ratio of whole numbers.

Example: Hydrogen combines with oxygen to form two compounds, named as water and hydrogen peroxide.

Hydrogen (2g) + Oxygen (16g) $\rightarrow$ Water (18g)

Hydrogen (2g) + Oxygen (32g) $\rightarrow$ Hydrogen Peroxide (34g)

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

2.2 Fundamental Concepts

- A substance which can not be decomposed into other substances is known as element. The smaller particle of an element which takes part in chemical reaction is known as an atom.
- An atom consists of a very small nucleus at its centre which is composed of protons and neutrons. The nucleus is encircled by moving electrons.
- The electron is a negatively charged particle and it has mass of about 1/1836 that of the neutron. Proton has positive charge while a neutron is an uncharged particle having mass equal to the proton.

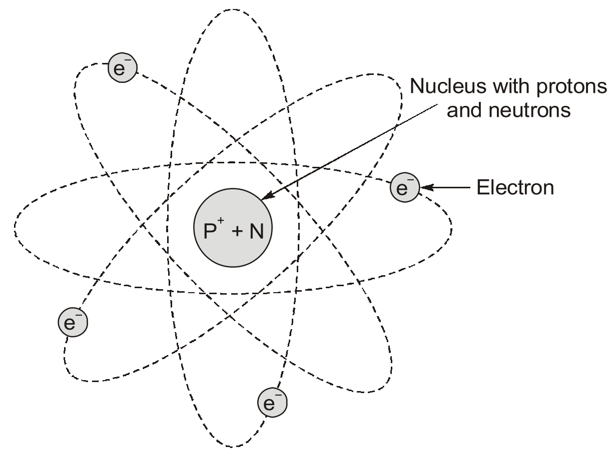

Fig. 2.1 *Concept of an atom*

- Each chemical element is characterized by the number of protons in the nucleus or the **atomic number (Z).** For an electrically neutral or complete atom, the atomic number also equals the number of electrons.
- The **atomic mass (A)** of a specific atom may be expressed as the sum of the masses of protons and neutrons within the nucleus.
- The atomic weight of an element corresponds to the weighted average of the atomic masses of the atom's naturally occurring isotopes. The **atomic mass unit (amu)** may be used for computations of atomic weight.

NOTE

A scale has been established whereby 1 amu is defined as 1/12 of the atomic mass of the most common isotope of carbon, (carbon –12, A = 12.00000). Within this scheme, the masses of protons and neutrons are slightly greater than unity and

$$A \cong Z + N$$

where N = Number of neutrons.

- **Isotopes:** Isotopes have different atomic weights but they have same atomic number. Isotopes are chemically inseparable as they posses identical chemical properties. Isotopes of the same element have the same atomic number and the same charge on the nucleus. $_1H^1$, $_1H^2$ and $_1H^3$ are isotopes of hydrogen while $_{17}Cl^{37}$ and $_{17}Cl^{35}$ are isotopes of chlorine.
- **Isobars:** Atoms which have same mass number (atomic weight or number of protons and neutrons) but they differ in atomic number are called isobars. Isobars are atoms of different chemical elements but they have same atomic mass number. Ar^{40} and Ca^{40} are isobars having same atomic mass number of 40 but they have varying number of protons (atomic number) and neutrons.

- **Isotones:** Atoms whose nuclei have the same number of neutrons but different number of protons. Thus, chlorine-37 and potassium-39 are isotones as their nuclei contain 17 and 19 protons respectively but same 20 neutrons. Isotones have different atomic number and different chemical properties.

2.3 Electrons in Atoms

- The electrons, protons and neutrons in atoms of various elements are identical. Thus, it follows that electrons, protons and neutrons are the fundamental particles of the universe. If it is so, then why do various elements behave differently? This is because of the difference in the number and arrangement of the electrons, protons and neutrons of which each atom is composed.
- All the electrons of an atom do not move in the same orbit. The electrons in an atom are arranged in different orbits or shells.

Remember

In general a shell (or orbit) can contain a maximum of $2n^2$ electrons, where n is the number of shell (or orbit). But according to this rule, there is an exception, the outermost orbit in an atom can not accommodate more than eight electrons. The electrons present in the outermost shell (or orbit) are called valence electrons.

- All the elements have been arranged in a periodic table according to the electronic arrangements in their atoms. The element placed in one vertical column have very similar properties.

2.4 The Periodic Table

- All the elements have been classified according to electron configuration in the **periodic table**. Here, the elements are situated, with increasing atomic number, in seven horizontal rows called periods.
- The arrangement is such that all elements arrayed in a given column or group have similar valence electron structures, as well as chemical and physical properties. These properties change gradually, moving horizontally across each period and vertically down each column.
- The elements positioned in Group 0, the rightmost group, are the inert gases, which have filled electron shells and stable electron configurations. Group VIIA and VIA elements are one and two electrons deficient, respectively, from having stable structures.
- The Group VIIA elements (F, Cl, Br, I and At) are sometimes termed the halogens. The alkali and the alkaline earth metals (Li, Na, K, Be, Mg, Ca etc.) are labeled as Groups IA and IIA, having, respectively, one and two electrons in excess of stable structures.
- The elements in the three long periods, Groups IIIB through IIB, are termed as the transition metals, which have partially filled *d* electron states and in some cases one or two electrons in the next higher energy shell.
- Groups IIIA, IVA and VA (B, Si, Ge, As etc.) display characteristics that are intermediate between the metals and nonmetals by virtue of their valence electron structures.
- As may be noted from the periodic table, most of the elements really come under the metal classification. These are sometimes termed **electropositive** elements, indicating that they are capable of giving up their few valence electrons to become positively charged ions.
- Furthermore, the elements situated on the right-hand side of the table are **electronegative;** that is, they readily accept electrons to form negatively charged ions or sometimes they share electrons with other atoms.
- Atoms are more likely to accept electrons if their outer shells are almost full and if they are less "shielded" from (i.e., closer to) the nucleus.

Key: 29 ← Atomic number; Cu ← Symbol; 63.54 ← Atomic weight

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII	VIII	VIII	IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
1 H 1.0080																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.054	17 Cl 35.453	18 Ar 39.948
19 K 39.099	20 Ca 40.08	21 Sc 44.956	22 Ti 47.87	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.69	29 Cu 63.54	30 Zn 65.41	31 Ga 69.72	32 Ge 72.64	33 As 74.922	34 Se 75.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 105.4	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.30
55 Cs 132.91	56 Ba 137.34	Rare earth series	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.2	76 Os 190.23	77 Ir 192.2	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.19	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (225)	Acti-nide series	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hd (227)	109 Mt (268)	110 Ds (281)								

Rare earth series	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.35	63 Eu 151.96	64 Gd 157.25	65 Tb 158.92	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
Actinide series	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Fig. 2.2 *The periodic table of the elements. The numbers in parentheses are the atomic weights of the most stable or common isotopes.*

2.5 Comparison of Alpha (α), Beta (β) and Gamma (γ) Rays

2.5.1 Alpha Rays

- Alpha rays or alpha particles are the positively charged particles.
- These particles have highly active & energetic helium nucleus which contains two protons and two neutrons.
- Alpha particles have the least penetration power but the greatest ionization power.
- They cannot penetrate the skin but this does not mean that they are not dangerous. Since they have a great ionization power, so if they get into the body they can cause serious damage.

2.5.2 Beta Rays

- Beta particles are extremely energetic electrons that are liberated from the inner nucleus.
- They are negatively charged and have a negligible mass. On the emission of a beta particle, a neutron in the nucleus splits into a proton and an electron.
- Hence, it is the electron that is emitted by the nucleus at a rapid pace.
- Beta particles have a higher penetration power and lower ionization power when compared to alpha particles and can travel through the skin with ease.

2.5.3 Gamma Rays

- The waves arising from the high-frequency end of the electromagnetic spectrum that has no mass are known as gamma rays.
- They hold the highest power of penetration. They are the most penetrating but least ionizing and very difficult to resist them from entering the body.
- The Gamma rays carry a large amount of energy and can also travel via thick concrete and thin lead.

2.5.4 Properties of Alpha, Beta and Gamma rays

Table 2.1 *Characteristics of Alpha, Beta and Gamma Rays*

Characteristic	Alpha (α)	Beta (β)	Gamma (γ)
Emission of	2P + 2N	1 electron – High K.E.	Photon - very high frequency electromagnetic radiation
Changes from	Uranium to Plutonium	Radium to Polonium	No change
Charge (C)	+2	–1	0
Mass (amu)	4	1/1850	0
Speed km/s	15000	300000	300000
% of speed of light	5%	Close to 100%	100%
K.E.	3-6 MeV	5 MeV to 1 MeV	100 keV less than 10 MeV
Penetration Power	Low - Large mass & charge -can be stopped by a thin sheet of paper	Moderate - Medium mass and charge- can be stopped by a few mm thick metal	Very high - no mass, no charge, can be stopped only by a very thick cement or steel block
Ionization power	Very high - Large charge	Moderate - Low charge	Low - No charge

2.6 Quantum Number

Quantum numbers are used to find the electron configuration of an atom and the probable location, energy level, other characteristics like ionization and atomic radius of an electron in the atom. Quantum numbers are also applied to check the movement and orbit of each and every electron within an atom. Quantum numbers are of four types:

2.6.1 The principal quantum number (*n*)

- Principal quantum number of any electron in an atom stands for the main energy level or shell to which an electron belongs. Energy of an electron and its average distance from nucleus depends upon principal quantum number. Increasing the value of 'n' results the distance of electron from its nucleus and its energy also start increasing.
- Shells are specified by a principal quantum number *n*, which may require an integral values beginning with unity; sometimes these shells are designated by the letters *K*, *L*, *M*, *N*, *O* and so on, which correspond, respectively to n = 1, 2, 3, 4, 5,..., shown in Table 2.2.

Table 2.2 *Electron states in some of electron shells & subshells*

Principal quantum Number (n)	Shell Designation	Subshells	No. of States	Number of Electrons	
				Per Subshell	Per Shell
1	*K*	*s*	1	2	2
2	*L*	*s*	1	2	8
		p	3	6	
3	*M*	*s*	1	2	18
		p	3	6	
		d	5	10	
4	*N*	*s*	1	2	32
		p	3	6	
		d	5	10	
		f	7	14	

2.6.2 The orbital angular momentum/ azimuthal quantum number (*l*):

l is the second quantum number which represents the sub-shell, denoted by letters *s*, *p*, *d* or *f*; it is related to the shape of the sub-shell of electron. The number of these sub-shells is defined by the magnitude of *n*. It specifies the number of units of angular momentum connected with an electron in a given orbit and finds the shape of the orbit and the energy of the sublevel.

Note: For any value of n quantum number l can have any integral value from 0 to $n - 1$. Hence we can have 4d, 5f, 2p and 2s electrons whereas 1p, 2d, 3f subshell electrons do not exist.

Example: For

$n = 1$, $l = 0$

$n = 2$, $l = 0, 1$

$n = 3$, $l = 0, 1, 2$

We have observed that n is the principal quantum number that defines principal shell. l provides the possible orbital sub-shells. The sub-shells in the main shell are s, p, d, f, g and h with quantum number $l = 0, 1, 2, 3, 4$ and 5 respectively. We can demonstrate it as follows:

For $n = 1$, $l = 0$, the electron is said to be in 1s sub-shell

$n = 2$, $l = 1$, the electron is said to be in 2p sub-shell

$n = 2$, $l = 0$, the electron is said to be in 2s sub-shell

2.6.3 The magnetic quantum number (m_l)

The third quantum number m_l is used to determine the number of orbitals for each subshell. For an 's' subshell, there is a single energy state, whereas for p, d and f subshells three, five and seven states exist, respectively (Table 2.2).

- The value of m_l varies between $+l$ to $-l$ with zero and as we know that m_l have $(2l + 1)$ values for a given l. For any specific value of l, an electron may have integral values of its inner quantum number m_l from $+l$ through 0 to $-l$. Thus for $l = 2$, m_l can take on the values +2, +1, 0, –1 and –2. Thus we get

For $l = 0$, $m_l = 0$

$l = 1$, $m_l = -1, 0, 1$

$l = 2$, $m_l = -2, -1, 0, +1, +2$

2.6.4 The Magnetic Spin Quantum Number (m_s)

- The electron can spin either in the clockwise or anticlockwise direction and values of spin can be $+\frac{1}{2}$ and $-\frac{1}{2}$, depending upon the direction of spin. m_s is used to represent the spin of an electron.
- We need to remember that the three quantum numbers n, l and m_l can have the same values for two electrons in an atom and that these two electrons will have their spins oriented in opposite directions.

Example 1.

Write the four quantum numbers for each of the electrons in the outermost shell of a boron atom?

Solution:

For boron, $Z = 5$, obviously, it has 5 electrons in it. Out of these 2 are in K shell and remaining 3 in the L-shell of the 3 electrons in the L shell 2 are 2 s electrons and 1 is a p electron. Hence the quantum numbers of the electrons in the L-shell are as follows:

n	l	m_i	m_s
2	0	0	$+\frac{1}{2}$
2	0	0	$-\frac{1}{2}$
2	1	0	$\pm\frac{1}{2}$

Example 2.

When the quantum number l = 4, the quantum number m_l takes the following number of values.

(a) 8 (b) 10

(c) 9 (d) 5

Ans: (c)

$$m_l = (2l + 1) = 2 \times 4 + 1 = 9$$

2.7 Electron Affinity

- This is the amount of energy released, when an electron is added by a neutral atom.
- The energy required to transfer an electron from one atom to another atom is the difference between the ionization energy I_1 and the electronic affinity E_{12} of the respective atoms, $I_1 - E_{12}$.
- Chlorine has the highest electron affinity.

NOTE

1. Electron affinity decreases with increase in atomic radius.
2. Sign given to electron affinity is negative because energy is released.
3. When force of attraction decreases electron affinity decreases.

2.8 Electronegativity

- Electronegativity is a chemical property that defines the tendency of an atom to attract a bonding pair of electrons towards itself. When an element strongly attracts electrons then it means that element has high electronegativity.
- Electronegativity is a measure of the ability of an atom in a chemical compound to attract shared electrons to it.
- The higher the associated electronegativity number, the more an element or compound attracts electrons towards it.
- The electronegativity of any given element varies depending on the element to which it is bound.
- Caesium is the least electronegative element in the periodic table (= 0.79), while fluorine is most electronegative (= 3.98). Electro-positivity is opposite of electronegativity that is a measure of an element's ability to donate electrons.

2.9 Pauli's Exclusion Principle

He stated that, “No two electrons within the same atom can have the same numerical values for their set of four quantum numbers.” This principle states that each electron state can hold no more than two electrons, which must have opposite spins. Thus, *s*, *p*, *d* and *f* sub-shells may each accommodate a total of 2, 6, 10 and 14 electrons respectively.

- Two electrons in an atom cannot be in the same quantum state i.e., their quantum numbers must be different.
- Of the four quantum numbers at least one must be different for the two electrons. For example n, l, m_l may be the same for the two electrons in an atom but the fourth quantum number m_s must be different for the two electrons. If m_s have $+\frac{1}{2}$ for one electron then it must have $-\frac{1}{2}$ for the other.
- Maximum number of electrons for a shell will be = $2n^2$ where n is the principal quantum number.

Previous ESE Prelims Questions

Q.1 In the case of ionic bonding, the molecule is stable as long as the number of bonding electrons is
(a) Equal to the number of antibonding electrons
(b) Less than the number of antibonding electrons
(c) Greater than the number of antibonding electrons
(d) Equal to the number of antibonding neutrons [[ESE Prelims : 2017]

Ans. (c)
For stable molecular structure, number of electrons in bonding molecular orbital must be greater than number of electrons in antibonding molecular orbital.

Q.2 Consider the following characteristics with respect to Alpha particles:
1. They have large specific ionization values.
2. They dissipate their energy rather slowly.
3. They can penetrate the outer layer of human skin.
4. Their emitters are heavy elements.

Which of the above statements are correct?
(a) 1 and 4 only (b) 1 and 3 only
(c) 2 and 4 only (d) 2 and 3 only [ESE Prelims : 2018]

Ans. (a)
Alpha particles can not penetrate human skin. Emitters of α-particles are heavy elements and they have large ionization power.

Q.3 **Statement (I) :** Atoms can neither be created nor destroyed.
Statement (II) : Under similar conditions of temperature and pressure, equal volumes of gases do not contain an equal number of atoms.
(a) Both Statement (I) and Statement (II) are individually true and Statement (II) is the correct explanation of Statement (I)
(b) Both Statement (I) and Statement (II) are individually true, but Statement (II) is not the correct explanation of Statement (I)
(c) Statement (I) is true, but Statement (II) is false
(d) Statement (I) is false, but Statement (II) is true [ESE Prelims : 2018]

Ans. (c)
Under similar conditions of temperature and pressure, equal volumes of gases contain an equal number of atoms.

Q.4 In which one of the following types of bonds, the bond formation is by free moving electrons in an array of positive ions?
(a) Homopolar bond (b) Electrostatic bond
(c) Metallic bond (d) Covalent bond [ESE Prelims : 2020]

Ans. (c)
Metallic bond is formed when atoms give up their electrons and become positive ions. So the bond is formed between free moving electrons and positive ions.

Q.5 Which of the following are the characteristics of covalent compounds?
1. They are mostly gases and liquids.
2. They are usually electric insulators.
3. They are directional in nature.
4. They are insoluble in polar solvents like water but are soluble in non-polar solvents.

(a) 1, 2 and 3 only (b) 1, 2 and 4 only
(c) 1, 3 and 4 only (d) 1, 2, 3 and 4 [ESE Prelims : 2020]

Objective Brain Teasers

Q.1 The bond between two identical non-metal atoms has a pair of electrons
(a) with identical spins
(b) unequally shared between the two
(c) transferred fully from one to another
(d) equally shared between them

Q.2 Which type of bonding would be expected between S and Cl?
(a) ionic (b) non-polar covalent
(c) polyionic (d) polar covalent

Q.3 A strontium atom differs from a strontium ion in that the atom has a greater
(a) number of electrons
(b) number of protons
(c) mass number
(d) atomic number

Q.4 Which combination of atoms can form a polar covalent bond?
(a) H and H (b) Na and Br
(c) H and Br (d) N and N

Q.5 Which kind of bonding can be found in a sample of H_2O?
(a) Hydrogen bonds only
(b) Both polar covalent and hydrogen bonds
(c) Ionic and non-polar hydrogen bonds
(d) Non-polar covalent bonds only

Q.6 Atoms bond primarily to
(a) increase their potential energy and lose stability
(b) reduce their potential energy and gain stability
(c) increase their potential energy and gain stability
(d) make more atoms

Q.7 Which one of the following compounds is most likely to be an ionic compound?
(a) CO_2 (b) CS_2
(c) KF (d) CCl_4

Q.8 Solid material chemical bonds are
(a) Ionic, molecular and fusion
(b) Covalent, fusion and fission
(c) Ionic, covalent and molecular
(d) Fission, molecular and ionic

Q.9 Compounds having Ionic bond have general characteristics
1. High strength
2. Stress–strain curve is a straight line upto fracture point.
3. These solids do not form closed pack structures.

Which of the above statement(s) is/are correct?
(a) 1 and 2 only (b) 2 and 3 only
(c) 1 only (d) 1, 2 and 3

Q.10 Which of the following bonds is directional in nature?
(a) Ionic bond (b) Metallic bond
(c) Dipole bonds (d) None of these

Q.11 Consider the following statements:
1. Secondary bonds are formed due to Vander waals forces of attraction between molecules.
2. Hydrogen bond is a special kind of Dipole Bond.

Which of the above statements is/are correct?
(a) 1 only (b) 2 only
(c) 1 and 2 both (d) None of these

Q.12 **Statement (I):** Covalent solids do not form closed pack structures.
Statement (II): Covalent bonding is directional in nature.
(a) Both Statement (I) and Statement (II) are individually true and Statement (II) is the correct explanation of Statement (I)
(b) Both Statement (I) and Statement (II) are individually true but Statement (II) is not the correct explanation of Statement (I)
(c) Statement (I) is true but Statement (II) is false
(d) Statement (I) is false but Statement (II) is true

Q.13 Consider the following statements:
1. Valence orbit electrons are tightly bound to nucleus due to very strong attraction force.
2. First orbit electrons in an atom posses highest amount of energy

Which of the above statements is/are correct?
(a) 1 only (b) 2 only
(c) both 1 and 2 (d) neither 1 nor 2

Q.14 **Assertion (A):** Water is a diamagnetic material.
Reason (R): Diamagnetism results due to unpaired electrons in molecular orbitals.
(a) Both Assertion (A) and Reason (R) are individually true and Reason (R) is the correct explanation of Assertion (A).
(b) Both Assertion (A) and Reason (R) are individually true but Reason (R) is NOT the correct explanation of Assertion (A).
(c) Assertion (A) is true but Reason (R) is false.
(d) Assertion (A) is false but Reason (R) is true.

Q.15 Which types of bonds are formed when elements having small numbers of valence electrons which are loosely held and can be released to a common pool?
(a) Metallic bonds
(b) Ionic bonds
(c) Covalent bonds
(d) Dipole bonds

Q.16 Which of the following are pair of isobaric elements?
(a) Ar, C (b) Ar, O
(c) Ca, Ar (d) Ca, Mg

Q.17 Match the following:

Type of Bond
A. Ionic Bond
B. Covalent Bond
C. Metallic Bond

Bond formation
1. Sharing the pairs of electrons
2. Electrostatic force of interaction
3. Electrostatic force of attraction.

Codes:

	A	B	C
(a)	1	3	2
(b)	1	2	3
(c)	3	1	2
(d)	3	2	1

Q.18 What is correct order in which bonding energy is increasing for following type of bond in solids?
(a) Ionic Bond > Van der Waals > Covalent
(b) Covalent Bond > Ionic Bond > Van der Waals Bond
(c) Ionic Bond > Covalent Bond > Van der Waal Bond
(d) Covalent Bond > Van der Waal Bond > Ionic Bond

Answers

1. (d)	2. (d)	3. (a)	4. (c)	5. (b)
6. (b)	7. (c)	8. (c)	9. (a)	10. (d)
11. (c)	12. (a)	13. (d)	14. (c)	15. (a)
16. (c)	17. (c)	18. (c)		

Explanations

8. (c)
Solid material chemical bonds are ionic Bond, covalent bond and molecular bond. **Ionic bond** is the strong electrostatic attraction between cation and anion. **Covalent bonds** are one in which there is a sharing of one or more electrons from the adjacent atoms.

9. (a)
- Ionic solids fracture in brittle mode, so they have linear stress-strain curve upto fracture point.
- Ionic solids form closed pack structures while covalent solids do not form closed pack structures.

10. (d)
Covalent bond is directional in nature.

13. (d)
- Valence orbit electrons are the outermost orbit electrons so attraction force of nucleus is minimum.
- First orbit electrons possess lowest amount of energy.

14. (c)
Diamagnetism results due to all paired electrons in molecular orbitals.

16. (c)
Isobars are atoms of different elements having different Atomic Number but same Mass Numbers. eg: ${}_{20}Ca^{40}$ / ${}_{18}Ar^{40}$.

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