Atomic Structure

Chapter 2 — Complete Chapter Notes

1st Year Chemistry

Concepts, worked examples, Quick Checks, MCQs & short questions

Table of Contents

SECTION 1: ATOMIC NUMBER, PROTON NUMBER AND NUCLEON NUMBER

Moseley's Law

Statement: The square root of the frequency of the X-rays is directly proportional to the atomic number of an element (Z).

Mathematically:

$$\sqrt{\nu} \propto Z$$

Moseley's Experiment (1913)

- Performed X-ray experiments with different elements
- When different elements were bombarded with cathode rays, X-rays of some characteristic frequencies were produced
- Concluded that atomic number (Z) was a fundamental property of an element
- Atomic number is also called **proton number**

Mass Number or Nucleon Number

The sum of protons and neutrons in the nucleus of an atom is called its **nucleon number (A)**. It is also called **mass number**.

$$A = Z + N$$

The number of neutrons (N) in an atom can be calculated as:

$$N = A - Z$$

An atom of an element X having atomic number Z and mass number A is described as:

$$^{A}_{Z}X \text{e.g.,} ^{35}_{17}Cl$$

Example: $^{27}_{13}\mathbf{Al}$

- Atomic number / proton number (Z) = 13
- Mass number / nucleon number (A) = 27
- Neutron number (N) = 27 - 13 = 14

For Ions:

$^{27}_{13}Al$ atom loses three electrons to form Al^{3+}:

- No. of protons = 13
- No. of neutrons = 14

- No. of electrons = 13 - 3 = 10

$^{35}_{17}Cl$ gains an electron to form Cl^{-} ion:

- No. of protons = 17
- No. of neutrons = 18
- No. of electrons = 17 + 1 = 18

Number of Protons, Electrons and Neutrons in Different Ions:

Species	Neutrons	Protons	Electrons
O^{2-}	8	8	10
S^{2-}	16	16	18
P^{3-}	16	15	18

SECTION 2: EFFECT OF ELECTRIC FIELD ON FUNDAMENTAL PARTICLES

The behaviour of particles in an electric field depends upon their mass and charge. When beams of electrons, protons and neutrons are passed through an electric field at the same speed:

Particle	Behaviour in Electric Field
Neutron	Neutral, not deflected; travels in straight path perpendicular to field
Proton	Positively charged; deflected towards negative plate
Electron	Negatively charged; deflected towards positive plate

Note: Electrons are deflected to a greater extent than protons because they are 1/1836 times lighter than protons.

Factors Affecting Deflection:

- **Angle of deflection** ∝ charge / mass
- **Radius of deflection** ∝ mass / charge

After deflection, the particle moves in a circular path.

Properties of Three Fundamental Particles:

Particle	Charge (coulomb)	Relative Charge	Mass (kg)	Mass (amu)
Proton	$+1.6022 \times 10^{-19}$	+1	1.6726×10^{-27}	1.0073
Neutron	0	0	1.6750×10^{-27}	1.0087
Electron	-1.6022×10^{-19}	-1	9.1095×10^{-31}	5.4858×10^{-4}

🔍 Quick Check 2.1

a) Calculate the number of neutrons in the following elements:

$^{39}_{19}K$:

- Z = 19
- A = 39
- N = 39 - 19 = 20 neutrons

$^{35}_{17}Cl$:

- Z = 17
- A = 35
- N = 35 - 17 = 18 neutrons

$^{40}_{18}Ar$:

- Z = 18
- A = 40
- N = 40 - 18 = 22 neutrons

b) Which of electron, proton and neutron is deflected the most in the magnetic field?

In a magnetic field, angle of deflection ∝ charge/mass. Since electron has the least mass, its angle of deflection will be highest. Thus, **electron** is deflected the most.

SECTION 3: EXPERIMENTAL EVIDENCES FOR ELECTRONIC CONFIGURATION

The modern theory of electronic structure originates from the Bohr Model of atom. Evidences are derived from:

1. Atomic spectra
2. Ionization energies

(i) ATOMIC SPECTRA (LINE SPECTRA)

A visual display or dispersion of the components of an electromagnetic radiation is called a **spectrum**.

An absorption or emission spectrum which consists of lines is called a **discontinuous or line spectrum**. Each line shows a specific wavelength absorbed or emitted from the atoms of an element.

Types of Atomic Spectra: **(a) Atomic Emission Spectrum**

Definition: The spectrum formed by emission of radiations by an atom.

Formation: When an element in gaseous state is heated to high temperatures or subjected to electrical discharge, radiation of certain wavelengths is emitted. The spectrum of this radiation contains coloured lines.

(b) Atomic Absorption Spectrum

Definition: The spectrum formed by absorption of radiations by an atom.

Formation: When a beam of white light is passed through a gaseous sample of an element in cold state, certain wavelengths are absorbed by atoms. The absorbed wavelengths show up as dark lines on the spectrum.

Note: The wavelengths of the dark lines in the absorption spectrum are exactly the same as in the emission spectrum.

Why Atomic Spectra are Called "Finger Prints" of Elements:

- Radiations are emitted or absorbed when electrons jump from one energy level to another
- Each element has a unique arrangement of electrons and unique fixed energy levels
- Every element absorbs or emits unique wavelengths of radiation
- Therefore, every element has its own characteristic spectrum

(ii) IONIZATION ENERGY

Definition: The minimum amount of energy required to remove an electron from its gaseous atom is called ionization energy.

Relationship Between Ionization Energy and Energy Levels:

Different electrons in different shells have different energies. The electronic configuration can be investigated by measuring ionization energies in two ways:

1. Successive ionization energies of the same atom
2. First ionization energies for different atoms

Successive Ionization Energies:

The energy required to remove each electron, one by one, from an atom:

$$X \rightarrow X^{+} + e^{-} \Delta H_1 X^{+} \rightarrow X^{2+} + e^{-} \Delta H_2 X^{2+} \rightarrow X^{3+} + e^{-} \Delta H_3$$

Example: Magnesium Atom

Graph Analysis:

Electrons Removed	Shell	Energy Requirement
First two	Outermost shell	Less energy required
Third	Inner shell	Large increase (closer to nucleus)
Next seven	Second shell	Gradual increase
11th and 12th	First shell (innermost)	Enormous jump

🔍 Quick Check 2.2

a) Write equations that describe: **i. The 1st ionisation energy of calcium:**

$$Ca_{(g)} \rightarrow Ca^{+}_{(g)} + e^{-} \Delta H_1$$

ii. The 3rd ionisation energy of potassium:

$$K^{2+}_{(g)} \rightarrow K^{3+}_{(g)} + e^{-} \Delta H_3$$

iii. The 2nd ionisation energy of lithium:

$$Li^{+}_{(g)} \rightarrow Li^{2+}_{(g)} + e^{-} \Delta H_2$$

iv. The 5th ionisation energy of sulfur:

$$S^{4+}_{(g)} \rightarrow S^{5+}_{(g)} + e^{-} \Delta H_5$$

b) For element aluminium (Z=13), draw a sketch graph between log10 of successive ionisation energies against number of electrons removed.

Answer: The graph shows three distinct jumps:

- Small jumps for first 3 electrons (valence shell)
- Large jump for 4th electron (inner shell)
- Very large jumps for remaining electrons (inner shells)

c) First and second ionisation energies (kJ/mol) of a few elements:

Element	ΔH_1	ΔH_2
I	237	2525
II	520	7300
III	900	1760
IV	1680	3380

Analysis:

i. A reactive metal: Element II (least first ionization energy)

ii. A reactive non-metal: Element IV (element I and IV are non-metals; IV has relatively lower ionization energy than I)

iii. A noble gas: Element I (highest first ionization energy)

iv. A metal that forms a stable binary halide of formula AX_2: Element II (has two valence electrons; ΔH_2 is not exceptionally high)

SECTION 4: QUANTUM NUMBERS

Bohr's Model vs. Schrodinger's Model:

Bohr's Model	Schrodinger's Model
One-dimensional model	Three-dimensional model
One quantum number (n)	Three quantum numbers
Described size and energy of orbit	Describes orbitals of electrons

The Four Quantum Numbers:

1. Principal Quantum Number (n)

Symbol: n

Values: 1, 2, 3, 4... (designated by K, L, M, N...)

Significance:

- Describes the size of the orbital
- Larger n = greater average distance from nucleus
- Larger n = higher energy (less tightly bound)

Calculations for one electron system:

- Radius: $r_n = 0.529(n^2)\text{Å}$
- Energy: $E_n = -2.18 \times 10^{-18}/n^2 \text{ J}$

Maximum number of electrons in a shell: $2n^2$

n	Shell	Max Electrons
1	K	2
2	L	8
3	M	18
4	N	32

2. Azimuthal Quantum Number (l) / Angular Quantum Number

Values: 0 to (n-1)

Significance: Describes the shape of the orbital

l value	Subshell	Shape
0	s	Spherical
1	p	Dumbbell
2	d	Complex
3	f	Very complex

3. Magnetic Quantum Number (m)

Values: -l to +l (including 0)

Significance: Describes the orientation of the orbital in space

Number of orbitals in a subshell: (2l + 1)

Subshell	l value	m values	Number of orbitals
s	0	0	1
p	1	-1, 0, +1	3
d	2	-2, -1, 0, +1, +2	5
f	3	-3, -2, -1, 0, +1, +2, +3	7

4. Spin Quantum Number (s)

Values: +½ or -½

Significance: Describes the spin of the electron (clockwise or anticlockwise)

Shapes of Orbitals:

s-Orbitals:

- l = 0, m = 0 (only one orbital)
- Shape: Spherical
- Electron density uniformly distributed in all directions
- Size increases with increase in n

p-Orbitals:

- l = 1, m = -1, 0, +1 (three orbitals)
- Shape: Two lobes on any axis
- Named: px, py, pz (mutually perpendicular)

d-Orbitals:

- l = 2, m = -2, -1, 0, +1, +2 (five orbitals)
- Shapes and orientations: d_{xy}, d_{xz}, d_{yz}, $d_{x^2-y^2}$, d_{z^2}

- dxy, dxz, dyz: lobes between axes
- dx^2-y^2: lobes along x and y axes
- dz^2: two lobes along z-axis with "doughnut" in xy plane

f-Orbitals:

- l = 3, m = -3, -2, -1, 0, +1, +2, +3 (seven orbitals)
- Shapes: Very complicated

SECTION 5: ELECTRONIC CONFIGURATION

Electronic Configuration is the distribution of electrons among available shells, subshells, or orbitals of an atom or ion.

Distribution of Electrons in Shells:

Shell	n value	Max Electrons ($2n^2$)
K	1	2
L	2	8
M	3	18
N	4	32

Example Distribution:

Sodium (Na): 2, 8, 1

Chlorine (Cl): 2, 8, 7

Distribution of Electrons in Subshells:

Notation: Subshell symbol with superscript indicating number of electrons

Example: Two electrons in 1s orbital = $1s^2$

Aufbau Principle

The subshells in an atom are filled with electrons in increasing order of their energy values.

Also known as: Building up principle

The energy of a subshell depends upon:

- Principal quantum number (n)
- Azimuthal quantum number (l)

The order of filling is obtained from the **(n + l)** value.

(n + l) Rule:

- Lower the (n + l) value, lower the energy of subshell → filled first
- If two subshells have same (n + l) value, lower n value → filled first

(n + l) Values of Various Subshells:

n	l	Subshell	(n+l) value
1	0	1s	1
2	0	2s	2
2	1	2p	3
3	0	3s	3
3	1	3p	4
3	2	3d	5
4	0	4s	4
4	1	4p	5
4	2	4d	6
4	3	4f	7
5	0	5s	5
5	1	5p	6
5	2	5d	7
5	3	5f	8
6	0	6s	6
6	1	6p	7

Order of Filling Orbitals:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s$$

Why 4s has lower energy than 3d:

Subshell	(n+l) value	n value	Remarks
4s	4 + 0 = 4	4	Lower energy
3d	3 + 2 = 5	3	Higher energy

Thus, 4s orbital is filled before 3d.

Pauli's Exclusion Principle

Statement: No two electrons in an atom can have the same values for all four quantum numbers.

Alternatively: Two electrons in an orbital will always have opposite spins.

Example: Helium (He) Atom

Electron	n	l	m	s
Electron 1	1	0	0	+½
Electron 2	1	0	0	-½

Hund's Rule

Statement: When degenerate orbitals are available and more than two electrons are to be placed in them, they should be placed in separate orbitals with the same spin rather than in the same orbital with opposite spins.

Example: Filling of p-orbitals in Carbon, Nitrogen and Oxygen:

Element	Configuration	Unpaired Electrons
Carbon	$1s^2\ 2s^2\ 2p^2$	2
Nitrogen	$1s^2\ 2s^2\ 2p^3$	3
Oxygen	$1s^2\ 2s^2\ 2p^4$	2

Electronic Configurations of Elements (1-36):

Z	Element	Electronic Configuration
1	H	$1s^1$
2	He	$1s^2$
3	Li	$1s^2\ 2s^1$
4	Be	$1s^2\ 2s^2$
5	B	$1s^2\ 2s^2\ 2p^1$
6	C	$1s^2\ 2s^2\ 2p^2$
7	N	$1s^2\ 2s^2\ 2p^3$
8	O	$1s^2\ 2s^2\ 2p^4$
9	F	$1s^2\ 2s^2\ 2p^5$
10	Ne	$1s^2\ 2s^2\ 2p^6$
11	Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$

Z	Element	Electronic Configuration
12	Mg	$1s^2\ 2s^2\ 2p^6\ 3s^2$
13	Al	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
14	Si	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
15	P	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
16	S	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
17	Cl	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
18	Ar	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
19	K	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
20	Ca	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
21	Sc	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$
22	Ti	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$
23	V	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3\ 4s^2$
24	Cr	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$
25	Mn	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$
26	Fe	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$
27	Co	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$
28	Ni	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$
29	Cu	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$
30	Zn	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$
31	Ga	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^1$
32	Ge	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^2$
33	As	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^3$
34	Se	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^4$
35	Br	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^5$
36	Kr	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$

SECTION 6: ELECTRONIC CONFIGURATION AND THE PERIODIC TABLE

Blocks in the Periodic Table:

Block	Columns	Subshell Being Filled	Example Groups
s-block	2 columns	s orbitals	Group 1 (ns^1), Group 2 (ns^2)

Block	Columns	Subshell Being Filled	Example Groups
p-block	6 columns	p orbitals	Group 13-18
d-block	10 columns	d orbitals	Transition metals
f-block	14 columns	f orbitals	Lanthanides, Actinides

Valence Electrons:

An electron in an atom outside the noble-gas or pseudo-noble-gas core is called a **valence electron**. These are involved in chemical reactions.

Examples:

Element	Electronic Configuration	Valence Electrons
Na	[Ne] $3s^1$	1
Mg	[Ne] $3s^2$	2
Al	[Ne] $3s^2\ 3p^1$	3

🔍 Quick Check 2.5

a) Write electron configurations using noble-gas inner core: **i. 48Cd:**

$$[\mathrm{Kr}]5s^2 4d^{10}$$

ii. 57La:

$$[\mathrm{Xe}]6s^2 5d^1$$

b) Complete electron configuration for antimony (Sb, Z=51):

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^3$$

c) How many unpaired electrons are there in each atom of 51Sb?

Since 5p orbital has 3 electrons, all three are unpaired.

d) Valence shell electronic configuration:

i. 13Al: [Ne] $3s^2\ 3p^1$

ii. 14Si: [Ne] $3s^2\ 3p^2$

iii. 28Ni: [Ar] $4s^2\ 3d^8$

🔍 **Quick Check 2.6**

i. An element has electronic configuration:

$$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^5$$

- **Block:** p-block (last electron in 5p)
- **Group:** 17 (valence shell: $5s^2 5p^5$, 7 valence electrons)
- **Period:** 5 (n = 5)
- **Element:** Iodine (Z = 53)

ii. Element with configuration:

$$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$$

- **Block:** d-block (d-orbital being filled)
- **Element:** Chromium (Cr)

SECTION 7: ELECTRONIC CONFIGURATION OF IONS AND FREE RADICALS

IONS: **Positive Ions (Cations):**

Formed when electrons are removed from atoms.

Example: Na^+ (Z=11, 10 electrons)

$$\text{Na}^+ : 1s^2 2s^2 2p^6 (\text{same~as~Ne})$$

Negative Ions (Anions):

Formed when atoms gain electrons.

Example: S^{2-} (Z=16, 18 electrons)

$$\text{S}^{2-} : 1s^2 2s^2 2p^6 3s^2 3p^6 (\text{same~as~Ar})$$

d-Block Elements:

When d-block elements form ions, 4s electrons are lost first.

Examples:

- Ti atom: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$

- Ti^{2+} ion: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2$
- Cr atom: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$
- Cr^{3+} ion: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3$

FREE RADICALS:

A **free radical** is a species that has one or more unpaired electrons. Free radicals are very reactive.

Examples:

Free Radical	Electronic Configuration
Cl•	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
OH•	Groups of atoms with unpaired electron
CH_3•	Groups of atoms with unpaired electron

SECTION 8: ELECTRONIC CONFIGURATION AND SEMICONDUCTORS

Semiconductors are materials that can conduct electricity under some conditions. Examples: Silicon, Germanium, Arsenic.

Silicon (Si, Z=14):

$$\text{Electronic~Configuration:~}1s^2 2s^2 2p^6 3s^2 3p^2 (2, 8, 4)$$

- Has 4 valence electrons
- In pure crystalline form, each Si atom bonds to four other Si atoms
- No possibility of electronic conduction

P-type Semiconductors:

- Formed by **doping** with trivalent impurities (Boron, Aluminum)
- Trivalent impurity has 3 valence electrons
- Creates "holes" (positive charge carriers)

N-type Semiconductors:

- Formed by **doping** with pentavalent impurities (Phosphorus, Arsenic)
- Pentavalent impurity has 5 valence electrons
- Adds extra electrons (negative charge carriers)

SECTION 9: MULTIPLE CHOICE QUESTIONS

Topic-wise MCQs: **Atomic Number and Mass Number:**

#	Question	Answer
1	The property atomic number of an element was discovered by:	**Moseley**
2	The mass of an electron is how many times less than that of a proton?	**1836**
3	The atomic number and mass number of S are 16 and 32 respectively. The total number of fundamental particles in S^{2-} is:	**50**
4	The number of neutrons in $^{23}_{11}Na$ is:	**12**

Atomic Spectra:

#	Question	Answer
5	Atomic spectrum is of how many types?	**Two**
6	Atoms form which type of spectrum?	**Line absorption spectrum and line emission spectrum**
7	Which is considered as the fingerprint of the elements?	**Atomic spectra**
8	The evidences for Bohr's theory were obtained from:	**Both ionization energy and atomic spectra**

Ionization Energy:

#	Question	Answer
9	Minimum energy required to remove an electron from its gaseous atom is called:	**Ionization energy**
10	The graph between log10 of successive ionization energies for Mg shows two large jumps. This shows Mg has:	**Three shells**
11	Which element has highest ionization potential?	**Be**
12	Elements which have highest ionization energy in each period are:	**Noble gases**

Quantum Numbers:

#	Question	Answer
13	Which quantum number is not directly derived from Schrodinger's theory?	**Spin**
14	Quantum number values for 4f orbital are:	**n=4, l=3**
15	If n=2 and l=2, how many orientations in space are possible?	**5**
16	The number of electrons in a sub-shell can be obtained by:	**2(2l+1)**
17	An orbital which is spherical and symmetrical is:	**s-orbital**

Electronic Configuration:

#	Question	Answer
18	Maximum number of electrons in an orbital is:	**2**
19	(n+l) value of 6d orbital is:	**8**
20	The electron in a subshell is filled according to formula:	**2(2l+1)**
21	Maximum number of electrons in f-subshell is:	**14**
22	The element which has maximum number of unpaired electrons is:	**Fe (Z=26)**
23	Electronic configuration $1s^2\ 2s^2\ 2p^4$ has how many unpaired electrons?	**2**
24	After filling of 4f, the entering electron goes into:	**5d**
25	When 4s orbital is complete, the electron goes into:	**3d**
26	When 5d orbital is completed then entering electron goes into:	**6p**

Entrance Test MCQs:

#	Question	Answer
1	In atomic particles:	**Charge of proton is almost equal to charge of electron**
2	Proton number of an element with four unpaired electrons in ground state:	**26**
3	Which quantum number helps to study orientation of orbital in space?	**Magnetic Quantum Number**
4	Quantum number values for 2p orbitals are:	**n=2, l=1**
5	Which pair shows abnormal electronic configuration?	**Cu and Cr**
6	Correct order of energy in given subshells:	**5s > 3d > 4s > 3p**
7	Number of electrons in outermost shell of chloride ion (Cl^-):	**8**
8	Relative energies of 4s, 4p and 3d orbitals:	**4s < 3d < 4p**
9	Electronic configuration of Manganese (Mn):	**[Ar] $4s^2\ 3d^5$**
10	Which pair has same electronic configuration as Neon (Ne-10)?	**Na^+, F^-**
11	Correct order with respect to increasing energy of orbitals:	**4s < 4p < 4d < 4f**
12	Element with ground state configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$:	**Ar**
13	Correct electronic configuration of nitrogen:	**$1s^2$, $2s^2$, $2px^1$, $2py^1$, $2pz^1$**
14	Configuration representing element that forms ion with charge -3:	**$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$**
15	Correct electronic configuration of ^{24}Cr:	**$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$**
16	Which two elements are isotopes?	**$^{12}_{6}X$ and $^{13}_{6}Y$**

#	Question	Answer
17	According to which scientist, probability of finding electron at a certain position is possible?	**Schrodinger**
18	When 6d orbital is filled, the entering electron goes into:	**7p**
19	Element with ground state configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$:	**Ar**
20	The p-orbital has:	**2 lobes**
21	Correct electronic configuration for carbon:	**$1s^2\ 2s^2\ 2p^2$**

Additional MCQs:

#	Question	Answer
1	How many electrons in a sub-shell with n=3, l=1?	**6**
2	Electronic configuration $1s^2\ 2s^2\ 2p^4$ has how many unpaired electrons?	**2**
3	When atoms are volatilized, they form:	**Line spectrum**
4	Splitting of spectral lines under electric and magnetic fields shows presence of:	**Sub-shells**
5	Quantum number values for 2p orbitals:	**n=2, l=1**
6	Electron in a shell filled according to:	**$2n^2$**
7	Which element can be doped with Si to prepare n-type semiconductor?	**As**
8	Free radical among following:	**Cl**

Answer Keys:

Topic-wise MCQs (1-12):

#	Ans	#	Ans	#	Ans	#	Ans
1	A	4	D	7	A	10	**B**
2	A	5	A	8	C	11	**B**
3	C	6	B	9	A	12	**B**

Additional MCQs (13-26):

#	Ans	#	Ans	#	Ans	#	Ans
13	D	17	A	21	D	25	**B**
14	C	18	D	22	C	26	**B**
15	C	19	A	23	B		
16	D	20	B	24	A		

Entrance Test MCQs:

#	Ans	#	Ans	#	Ans	#	Ans
1	D	6	B	11	A	16	C
2	D	7	D	12	A	17	D
3	C	8	B	13	A	18	C
4	A	9	A	14	B	19	A
5	D	10	D	15	A	20	A

SECTION 10: SHORT QUESTIONS

Q2. (a) There are three orientations of p-orbital due to three values of magnetic quantum number. Justify it.

Answer:

The magnetic quantum number 'm' describes the orientation of an orbital in space. Within a subshell, the value of 'm' depends on the value of ℓ.

For p-subshell, ℓ = 1, then there are three values of m: -1, 0, and +1. Thus, there are three orientations of p-orbitals. The p orbitals are named as px, py, and pz, which are perpendicular to each other.

(b) 'I_2' of Mg is much bigger than 'I_1'. Justify. **Answer:**

In Mg, there are two electrons in the valence shell. When the first electron is removed, the second electron experiences greater attraction from the nucleus. Thus, when two electrons of the outer shell have been removed, the third has to be removed from the inner shell that is closer to the nucleus. So, more energy is required. Hence, I_2 is much bigger than I_1.

(c) Among the elements Li, K, Ca, S and Kr which one has the lowest first ionisation energy? Which has the highest first ionization?

Answer:

Generally, larger atomic size means lesser electrostatic attraction between electrons and nucleus, so lower ionization energy and vice versa.

Atomic size order: K > Ca > Li > S > Kr

Ionization energy order: K < Ca < Li < S < Kr

- **Lowest first ionization energy:** K (Potassium)
- **Highest first ionization energy:** Kr (Krypton)

(d) Consider the electronic configuration of potassium atom (atomic number 19):

(i) Write the full electronic configuration:

$$_{19}K : 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$$

(ii) Explain why the 4s subshell is filled before the 3d subshell: **Answer:**

According to Aufbau principle, subshells are filled in increasing order of their energy values. The order of filling is obtained by (n + l) values. Lower the (n + l) value, lower the energy:

Subshell	(n+l) value	n value	Remarks
4s	4 + 0 = 4	4	Lower energy
3d	3 + 2 = 5	3	Higher energy

Thus, 4s orbital has lower energy than 3d and is filled first.

(e) An atom of element X has atomic number 17 and mass number 35:

(i) Determine number of protons, neutrons, and electrons:

- Z = 17, A = 35
- Protons = 17
- Neutrons = 35 - 17 = 18
- Electrons = 17

(ii) If element forms ion with charge -1:

- Protons = 17
- Neutrons = 18
- Electrons = 17 + 1 = 18

(f) In the ground state of mercury ($_{80}Hg$):

Electronic configuration:

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10}$$

i. How many electrons occupy atomic orbitals with n = 3?

- 3rd shell has 18 electrons ($3s^2$ $3p^6$ $3d^{10}$)

ii. How many electrons occupy 4d atomic orbitals?

- 10 electrons ($4d^{10}$)

iii. How many electrons occupy 4p atomic orbital?

- 6 electrons ($4p^6$)

iv. How many electrons have spin "up" (s = +½)?

- All orbitals completely filled
- Out of 80 electrons, half (40 electrons) will have spin up

(g) Successive ionization energies for an unknown element:

$$I_1 = 896\text{~kJ/mol}, I_2 = 1752\text{~kJ/mol}, I_3 = 14,807\text{~kJ/mol}, I_4 = 17,948\text{~kJ/mol}$$

Which family does it belong to?

The data shows that when two electrons of the outer shell have been removed, the third electron requires much higher energy. This shows there are two valence electrons. So, the element belongs to **Group 2** of the periodic table.

(h) Consider ionization energies for aluminum:

$$I_1 = 580\text{~kJ/mol}, I_2 = 1815\text{~kJ/mol}, I_3 = 2740\text{~kJ/mol}, I_4 = 11600\text{~kJ/mol}$$

(i) Account for the trend in the values:

When electrons are removed successively, the hold of the nucleus on remaining electrons increases. So, it becomes difficult to remove the next electron. Thus:

$$I_1 < I_2 < I_3 < I_4$$

(ii) Explain the large increase between I_3 and I_4:

This shows there are three electrons in the valence shell of Al. The fourth electron has to be removed from the inner shell (closer to nucleus). So more energy is required.

(iii) List four aluminum ions in order of increasing size:

When an atom loses electrons, it becomes a positive ion. Higher positive charge means smaller size.

Order of increasing size:

$$Al^{4+} < Al^{3+} < Al^{2+} < Al^{+}$$

(i) State the general order of filling orbitals up to the 4p subshell:

$$1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p$$

(j) Draw the orbital box diagram for valence electrons of phosphorus (Z=15):

Electronic configuration:

$${}_{15}P : 1s^2 2s^2 2p^6 3s^2 3p^3$$

P has 5 valence electrons ($3s^2 3p^3$).

Orbital box diagram for valence electrons:

3s		3p		
↑↓		↑	↑	↑

SECTION 11: SUBJECTIVE QUESTIONS

Q2. Answer the following short questions:

(i) Differentiate between line spectrum and continuous spectrum. **Answer:**

- **Line spectrum:** Consists of discrete lines at specific wavelengths; produced by atoms in gaseous state.
- **Continuous spectrum:** Contains all wavelengths in a continuous range; produced by hot solids or liquids.

(ii) Calculate the number of protons, neutrons and electrons in $^{56}_{26}Fe$.

Answer:

- Z = 26, A = 56
- Protons = 26
- Neutrons = 56 - 26 = 30
- Electrons = 26 (neutral atom)

(iii) How an emission spectrum is obtained? **Answer:**

When an element in gaseous state is heated to high temperatures or subjected to electrical discharge, radiation of certain wavelengths is emitted. The spectrum of this radiation contains coloured lines and is called atomic emission spectrum.

(iv) Define Moseley's law. Give its mathematical expression. **Answer:**

Moseley's law states that the square root of the frequency of the X-rays is directly proportional to the atomic number (Z) of an element.

$$\sqrt{\nu} \propto Z$$

(v) Distribute the electrons in ^{29}Cu and ^{35}Br. **Answer:**

^{29}Cu: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

^{35}Br: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^5$

(vi) Define spectrum. Name its two types. **Answer:**

A visual display or dispersion of the components of an electromagnetic radiation is called a spectrum.

Two types:

1. Continuous spectrum
2. Line spectrum (Atomic absorption spectrum and Atomic emission spectrum)

(vii) Differentiate between orbit and orbital. **Answer:**

Orbit	Orbital
Well-defined circular/elliptical path of electron	Region of space with maximum probability of finding electron
Bohr model concept	Quantum mechanical concept
One-dimensional	Three-dimensional
Has fixed energy and radius	Has variable probability distribution

(viii) State Pauli's Exclusion Principle with an example. **Answer:**

No two electrons in an atom can have the same values for all four quantum numbers.

Example: In helium, two electrons in 1s orbital have:

- Electron 1: n=1, l=0, m=0, s=+½
- Electron 2: n=1, l=0, m=0, s=-½

(ix) In an electric field, electron is deflected to a greater extent than proton. Explain.

Answer:

Angle of deflection ∝ charge/mass. Electrons have the same charge magnitude as protons but are 1/1836 times lighter. Therefore, electrons are deflected to a greater extent than protons.

(x) How p-type semiconductors are designed? **Answer:**

P-type semiconductors are formed by "doping" a pure semiconductor material (like Si) with trivalent impurities (like Boron or Aluminum). The trivalent impurity has three valence electrons, which creates "holes" (positive charge carriers) in the lattice.

(xi) What are free radicals? Give examples. **Answer:**

Free radicals are species that have one or more unpaired electrons. They are very reactive.

Examples: Cl• (chlorine atom), OH•, CH_3•

(xii) Write equations that describe: **(a) The 2nd ionization energy of Be:**

$$Be^{+}_{(g)} \rightarrow Be^{2+}_{(g)} + e^{-} \Delta H_2$$

(b) The 5th ionization energy of sulfur:

$$S^{4+}_{(g)} \rightarrow S^{5+}_{(g)} + e^{-} \Delta H_5$$

Q3. (a) What do you mean by successive ionization energies? How the electronic shell structure of aluminium (Al) is derived from the successive ionization energies?

Answer:

Successive ionization energies are the energies required to remove each electron one by one from an atom.

For Aluminum (Z=13):

- Shell structure: 2, 8, 3 (K=2, L=8, M=3)
- First three electrons removed from valence shell (M shell) require relatively low energy
- Fourth electron requires much higher energy (removed from inner L shell)
- Subsequent electrons show gradual increase from L shell
- 11th and 12th electrons show enormous jump (removed from K shell, closest to nucleus)

This pattern clearly shows three distinct electron shells in aluminum.

(b) What are quantum numbers? Explain the significance of azimuthal quantum number.

Answer:

Quantum numbers are sets of numbers that describe the properties of electrons in atoms.

Azimuthal Quantum Number (l):

- Also called angular momentum quantum number
- Values: 0 to (n-1)
- Describes the shape of the orbital
- Determines the subshell:

l value	Subshell	Shape
0	s	Spherical
1	p	Dumbbell
2	d	Complex
3	f	Very complex

KEY TERMS GLOSSARY

Term	Definition
Atomic Number (Z)	Number of protons in an atom

Term	Definition
Mass Number (A)	Sum of protons and neutrons in an atom
Nucleon	A proton or neutron in the nucleus
Isotopes	Atoms of same element with same Z but different A
Spectrum	Visual display of electromagnetic radiation components
Emission Spectrum	Spectrum formed by emission of radiation
Absorption Spectrum	Spectrum formed by absorption of radiation
Ionization Energy	Energy required to remove an electron from a gaseous atom
Principal Quantum Number (n)	Describes size and energy of orbital
Azimuthal Quantum Number (l)	Describes shape of orbital
Magnetic Quantum Number (m)	Describes orientation of orbital in space
Spin Quantum Number (s)	Describes spin of electron
Aufbau Principle	Electrons fill orbitals in order of increasing energy
Pauli's Exclusion Principle	No two electrons can have same four quantum numbers
Hund's Rule	Electrons fill degenerate orbitals singly first
Valence Electrons	Electrons in outermost shell involved in bonding
Free Radical	Species with unpaired electrons
Semiconductor	Material with intermediate conductivity
Doping	Adding impurities to semiconductor

SUMMARY

This chapter covers fundamental concepts of atomic structure including:

1. **Atomic Number and Mass Number** - Identity and composition of atoms
2. **Moseley's Law** - Relationship between atomic number and X-ray frequency
3. **Fundamental Particles** - Properties of protons, neutrons, and electrons
4. **Atomic Spectra** - Emission and absorption spectra as fingerprints of elements
5. **Ionization Energy** - Successive ionization energies revealing shell structure
6. **Quantum Numbers** - Four quantum numbers describing electron properties
7. **Electronic Configuration** - Distribution of electrons in shells, subshells, and orbitals
8. **Aufbau Principle** - Order of filling orbitals (n+l rule)
9. **Pauli's Exclusion Principle** - Limitation on electron quantum states
10. **Hund's Rule** - Filling of degenerate orbitals
11. **Periodic Table Relationship** - Blocks (s, p, d, f) based on electronic configuration
12. **Semiconductors** - Application of electronic configuration in technology

End of Chapter 2 - Atomic Structure

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