CHAPTER 02

# ATOMIC STRUCTURE

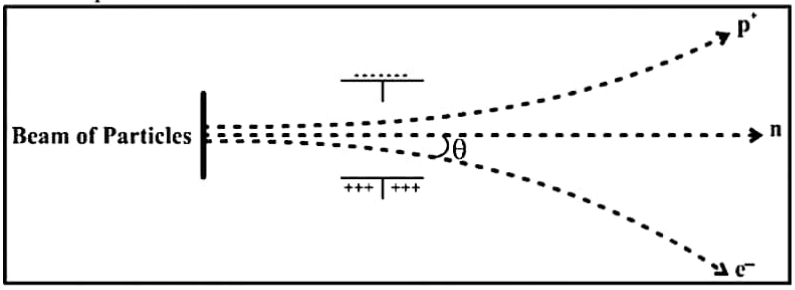
## Introduction

The matter around us is made up of atoms. Earlier, scientists believed that atoms were indivisible. However, modern research has shown that atoms are divisible and consist of more than 100 subatomic particles. Among these, the electron, proton, and neutron are considered the fundamental particles.

## EFFECT OF ELECTRIC FIELD ON FUNDAMENTAL PARTICLES

The behaviour of particles in an electric field depends upon their mass and charge. If we allow the beams of electrons, protons and neutrons to pass one by one at the same speed through an electric field, they show their behaviour as follows.

1. **Neutrons** being neutral are not deflected but travel in a straight path perpendicular to the direction of electric field.
2. **Protons** being positively charged are deflected towards the negative plate.
3. **Electrons** being negatively charged are deflected towards the positive plate, to greater extent since they are 1/1836 times lighter than protons.




### Factors Affecting Deflection of Charged Particles

The amount of deviation of a charged particle from its original path depends on two main factors:

**(i)** Angle of deflection $\propto \dfrac{\text{charge}}{\text{mass}}$ **(ii)** Radius of deflection $\propto \dfrac{\text{mass}}{\text{charge}}$

This occurs because, after deflection, the particle begins to move in a circular path. Therefore, the factors affecting the radius of deflection are the reciprocal of those affecting the angle of deflection. This means that when the angle of bending is large, the circular path is small—and when the angle is small, the path is larger.

**Properties of Fundamental Particles**

| | Electron | Proton | Neutron |
|---|---|---|---|
| **(i) Discovery** | J. J. Thomson | Eugene Goldstein | James Chadwick |
| **(ii) Charge** | $-1.6022 \times 10^{-19}$ C | $+1.6022 \times 10^{-19}$ C | Neutral |
| **(iii) Mass** | | | |
| **(a) kg** | $9.1095 \times 10^{-31}$ kg | $1.6726 \times 10^{-27}$ kg | $1.6750 \times 10^{-27}$ kg |
| **(b) a.m.u** | $5.4858 \times 10^{-4}$ amu | 1.0073 amu | 1.0087 amu |
| **(iv) Heavier than $e^-$** | Equal | 1836 times | 1843 times |

**Atomic Number (Proton Number)**

***"The number of protons in the nucleus of an atom is called atomic number (Z) or proton number."***

- It is the fundamental property of an element.
- In modern periodic table, the elements are arranged in the ascending order of their atomic numbers.
- The concept of atomic number was introduced by Moseley from his research work on X-rays.

**Mass Number** (Nucleon Number)

***"The number of protons and neutrons in the nucleus of an atom is collectively called its mass number (A) or nucleon number."***

- It is always in whole number.
- The mass of electron being very small is not included in the mass number.
- Protons and neutrons are collectively called nucleons.

**Relation between Atomic number (Z) and Mass number (A)**

Atomic number is related to the mass number by the following equation

Mass number = Atomic number + Number of neutrons

$$A = Z + N$$

**Representing atomic number (Z) and mass number (A) of an element.**

$${}^{A}_{Z}X$$

**Example 1:** **Calculate the number of neutrons in an atom of ${}^{27}_{13}Al$.**

**Solution:** Atomic number (Z) of Al = 13, Mass number (A) of Al = 27

**Number of neutrons = Mass number – Atomic number**

= 27 – 13

$\boxed{\textbf{Number of neutrons} = 14}$

**Example 2:** **Calculate the number of electrons, protons and neutrons in an ion of $Al^{3+}$.**

**Solution:** ${}^{27}_{13}Al$ atom loses 3 electrons to form $Al^{3+}$ ion.

Number of protons = 13
Number of neutrons = 14
Number of electrons = 13 – 3 = 10

**Example 3:** **Calculate the number of electrons, protons and neutrons in an ion of $Cl^{-}$.**

**Solution:** ${}^{35}_{17}Cl$ atom loses 1 electron to form $Cl^{-}$ ion.

Number of protons = 17
Number of neutrons = 18
Number of electrons = 17 + 1 = 18

**Number of protons, electrons and neutrons in different ions**

| Atom | Species | Neutrons | Protons | Electrons |
|---|---|---|---|---|
| ${}^{16}_{8}O$ | $O^{2-}$ | 8 | 8 | 10 |
| ${}^{32}_{16}S$ | $S^{2-}$ | 16 | 16 | 18 |
| ${}^{31}_{15}P$ | $P^{3-}$ | 16 | 15 | 18 |

**Quick Check 2.1**

**Calculate the number of neutrons in the following elements. ${}^{39}_{19}K$, ${}^{35}_{17}Cl$, ${}^{40}_{18}Ar$.**

**Ans: (i) ${}^{39}_{19}K$** (Atomic number (Z) of K = 19, Mass number (A) of K = 39)

**Number of neutrons = Mass number – Atomic number**

= 39 – 19

$\boxed{\textbf{Number of neutrons} = 20}$

**(ii) ${}^{35}_{17}Cl$** (Atomic number (Z) of Cl = 35, Mass number (A) of Cl = 17)

**Number of neutrons = Mass number – Atomic number**

= 35 – 17

$\boxed{\textbf{Number of neutrons} = 18}$

**(iii) ${}^{40}_{18}Ar$** (Atomic number (Z) of Ar = 40, Mass number (A) of Ar = 18)

**Number of neutrons = Mass number – Atomic number**

= 40 – 18

$\boxed{\textbf{Number of neutrons} = 22}$

## MOSELEY'S CONTRIBUTION TO ATOMIC STRUCTURE

In 1913, Moseley observed that when different elements were bombarded with cathode rays, the X-rays of some characteristic frequencies were produced. A very simple relationship was found between the frequency ($\upsilon$) of a particular line of X-rays and the atomic number Z of the element emitting it.

**Moseley's Law**

***"The square root of the frequency of the X-rays was directly proportional to the atomic number of an element Z."***

$$\sqrt{\upsilon} \propto Z$$

X-rays of some characteristics frequencies were produced, when different elements were bombarded with cathode rays (electrons).

**Moseley's Law and Idea of Atomic Number**

This law convinces us that characteristic properties of the element (both physical and chemical) are determined by the atomic number and not by the atomic mass. He concluded that this number, i.e. the atomic number Z was a fundamental property of an element.

**Q.** What are X-rays? What is their origin?
**Q.** How was the idea of atomic number derived from X-rays?
**Q.** What is Moseley's law?

### Experimental Evidences for the Electronic Configuration

The modern theory of electronic structure originates from the Bohr Model of atom. Evidence for this and later models of the atoms derives principally from two sources;

**(i)** Atomic spectra **(ii)** ionization energies

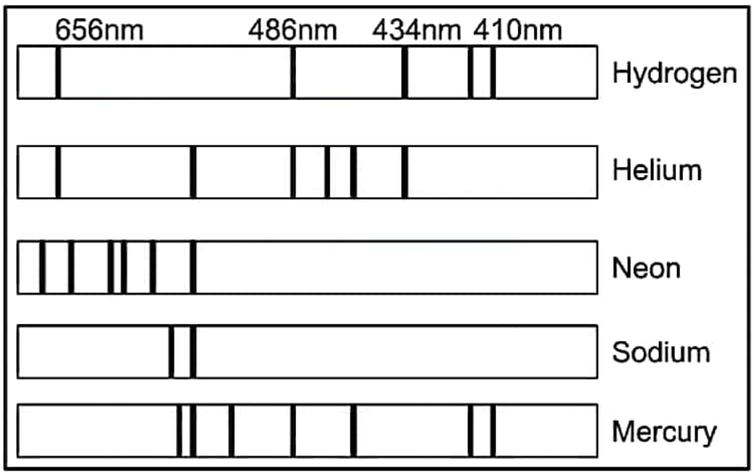
## ATOMIC SPECTRA

***"The spectrum of electromagnetic radiation emitted or absorbed by an electron during transitions between different energy levels within an atom is called atomic spectrum or line spectrum."***

- We see distinct lines separated by dark spaces.
- The boundary line between the colours can be marked.
- It is the characteristic of an atom.
- Each element has its own line spectrum.

### Atomic Spectra: The Fingerprints of Elements

- Each element has a unique arrangement of electrons and thus a unique range of fixed energy levels.
- It follows that the wavelengths and frequencies of the radiation absorbed or emitted when electrons jump from one energy level to another must also be unique.
- This uniqueness convinces us to conclude that every element has its own characteristic spectrum. Therefore, every element is identified by its characteristic spectrum.
- Hence, we can say that atomic spectra are the finger prints of the elements.



### Types of Atomic Spectrum

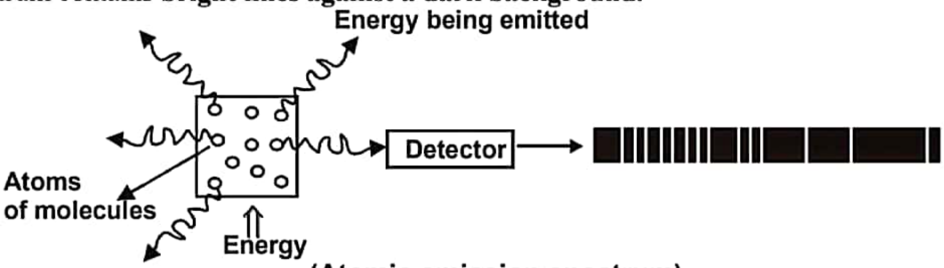
There are two ways in which an atomic spectrum can be viewed;

**(i)** Atomic emission spectrum **(ii)** Atomic absorption spectrum

### (i) Atomic Emission Spectrum

When an element in its gaseous state is heated to high temperatures or subjected to electrical discharge, radiation of certain wavelengths is emitted. Such an atomic spectrum is called atomic emission spectrum.

- This spectrum contains **bright lines against a dark background.**

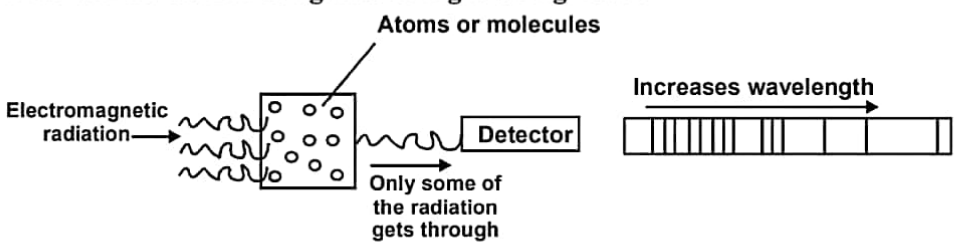



(Atomic emission spectrum)

### (ii) Atomic Absorption Spectrum

When a beam of white light is passed through a gaseous sample of an element in cold state, certain wavelengths are absorbed. Such an atomic spectrum is called atomic absorption spectrum.

- This spectrum contains **dark lines against a bright background.**

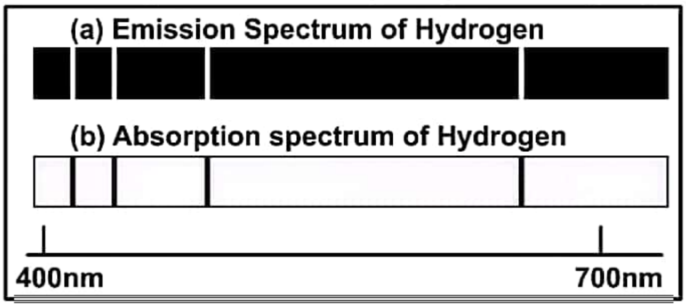



(Atomic absorption spectrum)

### Similarities of emission and absorption spectra

It is interesting to note that the wavelengths of the dark line in the absorption spectrum of a substance are the same as those of the wavelengths of the bright lines in the emission spectrum of the same substance.

In emission spectrum these lines appear bright because the corresponding wavelengths are being emitted by the element, whereas they appear dark in the atomic absorption spectrum because the wavelengths are being absorbed by the element. E.g. The atomic absorption spectrum of hydrogen.




**Q.** Define Atomic Spectrum. Give its two types.
**Q.** Why atomic spectrum is line spectrum?
**Q.** Compare line emission and line absorption spectra.
**Q.** What is the origin of line spectrum?

## RELATION BETWEEN IONIZATION ENERGY AND ENERGY LEVELS (electronic shells)

Electrons in different shells have different energy levels. The electronic configuration of the atoms can be investigated by measuring their ionization energies experimentally. There are two main ways ionization energies help us understand electronic configurations:

**(1)** Successive Ionization Energies of the Same Element
**(2)** First Ionization Energies of Different Atoms

**Ex.Q.5: What do you mean by successive ionization energies? How the electronic shell structure of magnesium (Mg) is derived from the successive ionization energies?**

### (1) Successive Ionization Energies of the Same Element

**"The removal of electrons from an atom continues until only the nucleus is left. This sequence of ionization energies is called successive ionization energies."**

- The successive ionization energies show clearly the arrangement of electrons in shells around the nucleus.
- The successive ionization energies of an element are all endothermic and their values become more and more endothermic. This is not surprising because each successive ionization requires the removal of a negative electron from an ion of increasing positive charge.

**General Example:**

$$X \longrightarrow X^+ + e^- \qquad I_1 = +ive$$
$$X^+ \longrightarrow X^{2+} + e^- \qquad I_2 = +ive$$
$$X^{2+} \longrightarrow X^{3+} + e^- \qquad I_3 = +ive$$
$$\mathbf{I_1 < I_2 < I_3}$$

### Electronic Shell Structure of Mg and Successive Ionization Energies

For magnesium atom, the energy required to remove successively the first electron, the second, the third, and so on are given below;

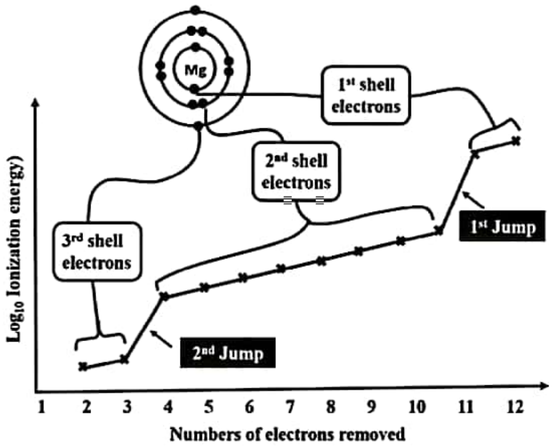
$$Mg_{(g)} \longrightarrow Mg^{+}_{(g)} + 1e^- \qquad I_1 = +738\ kJ\ mol^{-1}$$
$$Mg^{+} \longrightarrow Mg^{2+} + 1e^- \qquad I_2 = +1450\ kJ\ mol^{-1}$$
$$Mg^{2+} \longrightarrow Mg^{3+} + 1e^- \qquad I_3 = +7730\ kJ\ mol^{-1}$$

It is clear from ionization energy values that $\mathbf{I_3 > I_2 > I_1}$

**Graphical Explanation**

- A graph plotted between ionization energies values and number of electrons shows that successive ionization energies increase when we move from the valence shell to the inner shells.

**Valence Shell:** First two electrons are removed from the outermost shell and require lower energy for their removal.




**Logarithms to base 10**
Logarithmic scales are a means of bringing a very wide range of numbers onto the same scale. For a number, x, $\log_{10}x$ is the power to which 10 must be raised to equal x. So,

If $x = 10$, $\quad \log_{10}10 = 1$
If $x = 100$, $\quad \log_{10}100 = 2$
If $x = 10^6$, $\quad \log_{10}10^6 = 6$
If $x = 0.1$, $\quad \log_{10}0.1 = -1$

**Second Shell:** A large increase occurs when the third electron is removed. This is because when two electrons of the outer shell have been removed, the next has to be removed from the shell that is very much closer to the nucleus. The next seven electrons are removed successively from the second shell and a gradual increase in ionization energy is observed.

**First Shell:** A similar but much more enormous jump occurs when the 11th and 12th electrons are removed. These electrons are removed from the first, innermost shell, right next to the nucleus.

## Conclusion

Hence, over all, we observe two large jumps in the successive ionization energies. These two large jumps in the series of successive ionization energies are very good evidence that the electron in the magnesium atoms exist in three different shells.

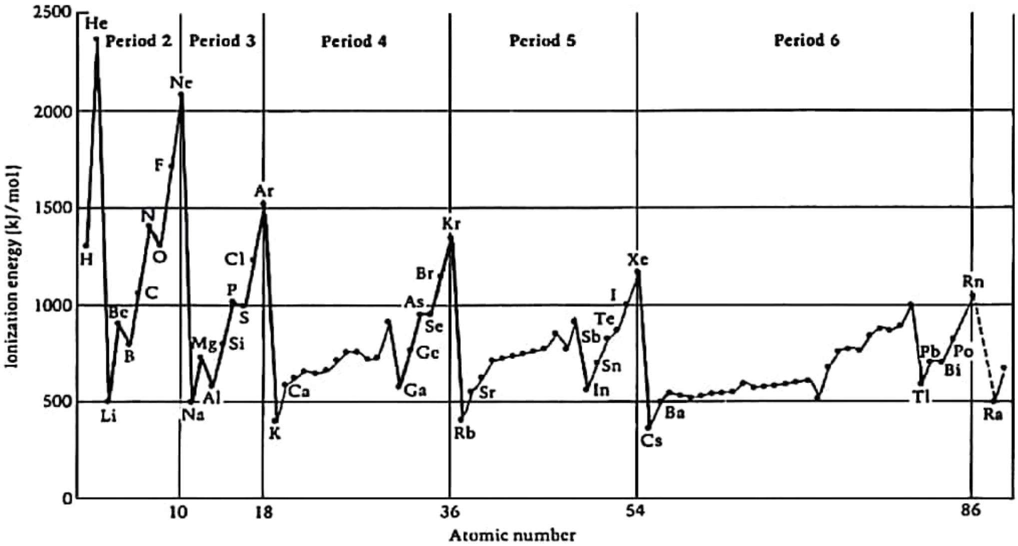
## (2) First Ionization Energies of Different Atoms

### (i) Down the group:

As we go down a particular group, for example, from helium to neon to argon, or from lithium to sodium to potassium, ionization energies decrease. The larger the atom, the easier is to separate an electron from it. Actually, down the group, number of shells increases, hold of nucleus on the valence electrons decreases, hence removal of electrons becomes easier.

### (ii) Across the period:

The ionization energies generally increase on going across a period. The group 1 elements, the alkali metals, have the lowest ionization energy within each period, and the noble gases have the highest. It is due to the reason that across the period shell number remains same. As the proton number increases, electrons are added in the same shell. Therefore, nucleus attracts the valence electrons more strongly. As a result, ionization energy increases across the period.



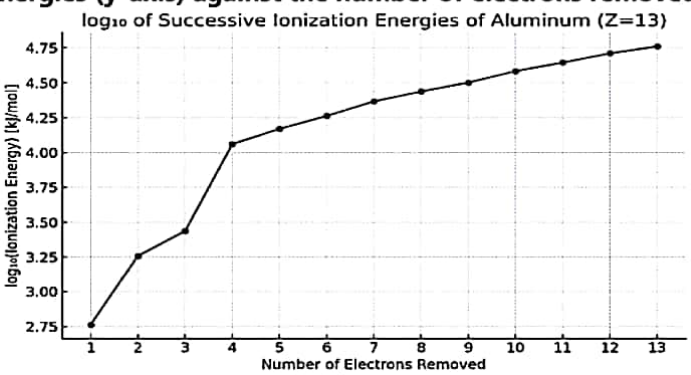

### Quick Check 2.2

**a) Write equations that describe:**

| | |
|---|---|
| **i. the 1st ionization energy of calcium**<br>**Ans:** $Ca_{(g)} \longrightarrow Ca^{+}_{(g)} + e^{-}$ | **ii. the 3rd ionization energy of potassium**<br>**Ans:** $K^{2+}_{(g)} \longrightarrow K^{3+}_{(g)} + e^{-}$ |
| **iii. the 2nd ionization energy of lithium**<br>**Ans:** $Li^{+}_{(g)} \longrightarrow Li^{2+}_{(g)} + e^{-}$ | **iv. the 5th ionization energy of sulfur.**<br>**Ans:** $S^{4+}_{(g)} \longrightarrow S^{5+}_{(g)} + e^{-}$ |

**b) For the element aluminium (Z = 13), draw a sketch graph between the log10 of the successive ionization energies (y-axis) against the number of electrons removed (x-axis).**

**Ans:**



c) The first $\Delta Hi_1$ & the second $\Delta Hi_2$ ionization energies (kJ / mol) of a few elements are given in table.

| Element | $\Delta Hi_1$ | $\Delta Hi_2$ |
|---|---|---|
| I | 2372 | 5251 |
| II | 520 | 7300 |
| III | 900 | 1760 |
| IV | 1680 | 3380 |

**Which of the above element is likely to be:**

| | |
|---|---|
| **i. a reactive metal,**<br>**Ans:** Element II (because $I_2 >> I_1$) | **ii. a reactive non-metal,**<br>**Ans:** Element IV (because both values are very high) |
| **iii. a noble gas**<br>**Ans:** Element I (Highest ionization energies values) | **iv. a metal that forms a stable binary halide of the formula $AX_2$ (X = halogen)**<br>**Ans:** Element III |

## From Bohr to Schrödinger: Understanding Electron Arrangement

- **Bohr's Model** was a one-dimensional model that used one quantum number to describe the distribution of electrons in the atom. The only important information was the size and energy of the orbit, which was described by quantum number denoted by 'n'.
- **Schrodinger's Model** allowed the electron to occupy three-dimensional space. It required three coordinates, or three quantum numbers, to describe the orbitals in which electrons can be found.
- **Atomic Orbital** is the volume of space in which there is 95% chance of finding an electron.
- The term orbital should not be confused with the term orbit as used in the Bohr's theory, the orbital can be regarded as a spread of charge surrounding the nucleus. This is often called the **"electron cloud"**.

**Ex.Q.3: What are quantum numbers? Describe briefly principal and spin quantum numbers.**

## QUANTUM NUMBERS

***"The sets of numerical values which give the acceptable solution to Schrodinger wave equation for hydrogen atom are called quantum numbers."*** ***or***

***"These are number or labels which completely describe an electron in an atom.***

There are four quantum numbers which can describe the electron completely.

**(1)** Principal quantum number (n)
**(2)** Azimuthal (angular) quantum number (l)
**(3)** Magnetic quantum number (m)
**(4)** Spin quantum number (s)

| Location | |
|---|---|
| **Person** | **Electron** |
| City | Shell |
| Town | Sub-shell |
| Block | Orbital |
| House# | Spin |

- The three quantum numbers that come from Schrödinger's wave equations are the principal (n), angular (l), and magnetic (m) quantum numbers.
- These quantum numbers describe the size, shape, and orientation in space of the orbitals in an atom.

### (1) Principal Quantum Number(n)

The different energy levels in Bohr's atom are represented by 'n'. This is called principal quantum number by Schrodinger. Its value is non-zero, positive integers upto infinity.

$$n = 1, 2, 3, 4, \ldots\ldots$$

| n | 1 | 2 | 3 | 4 |
|---|---|---|---|---|
| Shell | K | L | M | N |

- This quantum number, n, describes the size and energy of the orbital.
- The collection of orbitals with the same values of n is called an electron shell.

**Applications**

**(i)** **Size of an orbit:** "When 'n' is larger, the electron stays farther from the nucleus, so the orbital becomes larger."

**(ii)** **Energy of the orbit:** An increase in 'n' also means that the electron has a higher energy and is therefore less tightly bound to the nucleus.

**(iii)** **Number of electrons:** The maximum number of electrons that can be accommodated in any shell is given by the formula $2n^2$.

| **Shell** | **K** | **L** | **M** | **N** |
|---|---|---|---|---|
| **Principal quantum number (n)** | 1 | 2 | 3 | 4 |
| **Maximum number of electrons($2n^2$)** | 2 | 8 | 18 | 32 |

## (2) Azimuthal Quantum Number ($\ell$)

When spectral lines in a spectrum are observed using a high resolving power spectrometer, each individual line is further divided into several very fine lines. This indicates that each main energy level (or shell) is divided into smaller regions called **subshells**.

The subshells are described by the azimuthal quantum number (also called the angular momentum or subsidiary quantum number), which determines the shape of the orbitals within a shell.

**Value:** Its value are $\ell = 0, 1, 2, 3, \ldots\ldots (n - 1)$

Its value depends upon 'n'. These values represent different subshells which are designated by small letters **s, p, d and f** which stand for **sharp, principal, diffused and fundamental** respectively. These are the spectral terms used to describe certain features of spectral lines.

**(i) Shapes of subshells**

The set of orbitals that have the same 'n' and '$\ell$' values is called a subshell. A subshell may have different shapes depending upon the nature of '$\ell$'.

$\ell = 0$ s-subshell spherical shape
$\ell = 1$ p-subshell polar (dumb bell) shape
$\ell = 2$ d-subshell clover leaf (double dumb bell) shape
$\ell = 3$ f-subshell complicated shape

**(ii) Relation between 'n' and '$\ell$'**

The number of subshells in a shell is equal to its shell number. For example, 1st, 2nd, 3rd, and 4th shells have one, two, three and four subshells respectively.

| Shell | N | $\ell$ | Sub-shells | No. of sub-shells in a shell |
|---|---|---|---|---|
| K | 1 | 0 | 1s | 1 |
| L | 2 | 0<br>1 | 2s<br>2p | 2 |
| M | 3 | 0<br>1<br>2 | 3s<br>3p<br>3d | 3 |
| N | 4 | 0<br>1<br>2<br>3 | 4s<br>4p<br>4d<br>4f | 4 |

**(iii) Total number of electrons**

Number of electrons in a subshell can be determined by formula = $\mathbf{2(2\ell + 1)}$ or $\mathbf{4\ell + 2}$

| Value of '$\ell$' | 0 | 1 | 2 | 3 |
|---|---|---|---|---|
| **Orbital designation** | **s** | **p** | **d** | **f** |
| **Max No. of electrons** | **2** | **6** | **10** | **14** |

## (3) Magnetic Quantum Number(m)

This quantum number explains the splitting of spectral lines further when an excited atom is placed in a strong magnetic field (Zeeman Effect). This quantum number describes the orientation of an orbital in space that is why it is also called spacial orientation quantum number.

**Values:** $\mathbf{m = 0, \pm1, \pm2, \pm3, \ldots\ldots}$ (The values of 'm' indicate the number of orbitals in a subshell.)

**(i) Relation between '$\ell$' and 'm'**

$\boxed{\text{Number of orbitals (m)} = 2\ell + 1}$

The value of 'm' depends upon values of '$\ell$' i.e. $m = (2\ell + 1)$

When $\ell = 0$ s-subshell $m = 0$
$\ell = 1$ p-subshell $m = -1, 0, +1$ (p-subshell has three degenerate orbitals)
$\ell = 2$ d-subshell $m = -2, -1, 0, +1, +2$ (d-subshell has five degenerate orbitals)
$\ell = 3$ f-subshell $m = -3, -2, -1, 0, +1, +2, +3$ (f-subshell has seven degenerate orbitals)

| **Value of $\ell$** | 0 | 1 | 2 | 3 | 4 |
|---|---|---|---|---|---|
| **Subshell** | s | p | d | f | g |
| **Number of orbitals** | 1 | 3 | 5 | 7 | 9 |

**(ii) Degeneracy of orbitals**

Orbitals of the same subshell have same energy and are called degenerate orbitals. Actually, the value of 'm' gives us the information of degeneracy of orbitals in space. It tells us the number of different ways in which a given s, p, d or f sub-shell can be arranged along x, y and z axes in the presence of magnetic field.

**(iii) Number of orbitals in a shell**

The number of orbitals in a shell can be determine by using the formula, $n^2$. Where n = principal **Quantum Number:**

K(n = 1) has 1 orbital, L(n = 2) has 4 orbitals, M(n = 3) has 9 orbitals, N(n = 4) has 16 orbitals and so on.

**(iv) Number of electrons in an orbital**

An orbital can accommodate maximum of 2 electrons with opposite spin.

> **Did You Know!** Splitting of small fine lines in the presence of magnetic field and their three-dimensional orientation in space indicate the presence of orbitals in subshells. This splitting is called **"Zeeman effect"**.

## (4) Spin Quantum Number(s)

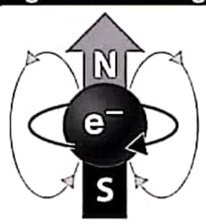
Electrons are thought of spinning around their own axes, as the Earth does. According to electromagnetic theory, a spinning charge generates a magnetic field. It is this motion that causes an electron to behave like a magnet.

There are following two possible spinning motions of an electron;

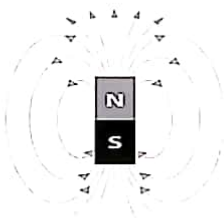
**(i) Clockwise spin:** It is represented by an arrow pointing upwards (↑). Its value is +1/2.

**(ii) Anti-clockwise spin:** It is represented by an arrow pointing downwards (↓). Its value is –1/2.

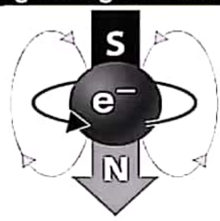
**Electron aligned with magnetic field**




Clockwise with spin up
s = +1/2




**Electron aligned against magnetic field**




Antilock wise with spin down
s = –1/2

- Each orbital can accommodate at the most two electrons provided the two electrons have opposite spins.
- The three quantum numbers i.e.'n', 'l' & 'm' describe the energy, shape and orientation of an orbital, but fourth quantum number (spin quantum number) is used to differentiate between the two electrons that can occupy an orbital.

**Quick Check 2.3**

**(a) What information about an electron in an atom can be obtained from:**

**Ans:**

| Quantum number | Information about electron |
|---|---|
| Principal Quantum Number | Size and energy of the orbital |
| Azimuthal Quantum Number | Shape of the orbital |
| Magnetic Quantum Number | Orientation of an orbital in space |
| Spin Quantum Number | Differentiate between two electrons in an orbital |

**(b) For an electron(s):**

**(i) If n = 2 and $\ell$ = 1, how many orientations in space are possible?**

**(ii) If n = 3 and $\ell$ = 2, which shell and subshell does the electron belong to?**

**(iii) If l = 2, find all possible values of 'm' and maximum number of electrons for 'm'.**

**Ans: Orientation in space, orbitals (m) = $2\ell + 1$**

**(i)** m = 2(1)+1 = 3 (3 orientation in space means three orbitals of p-subshell i.e. $p_x$, $p_y$ and $p_z$)

**(ii)** m = 2(2)+1 = 5 (5 orientation in space means five orbitals of d-subshell i.e. $d_{xy}$, $d_{yz}$, $d_{xz}$, $d_{x^2-y^2}$ & $d_{z^2}$ )

**(iii)** m = 2(2)+1 = –2, –1, 0, +1, +2. Since there are five orbitals and each orbital can accommodate maximum of two electrons so, total number of electrons in these orbitals are 10.

**Ex.Q.4:** **Draw the shapes of s, p and d-orbitals. Justify these by keeping in view the azimuthal and magnetic quantum numbers.**

## SHAPES OF ATOMIC ORBITALS

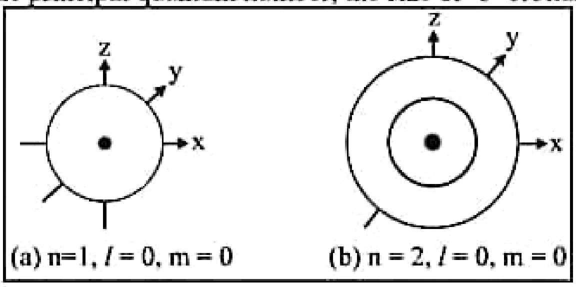
### Atomic Orbital

*"Three dimensional region in space around the nucleus in which the probability of finding the electron is maximum is called atomic orbital."*

- Size and energy of an orbital is influenced by Principal Quantum Number (n) while shape is determined by Azimuthal Quantum Number ($l$).
- An orbital can accommodate maximum of 2 electrons with opposite spin.

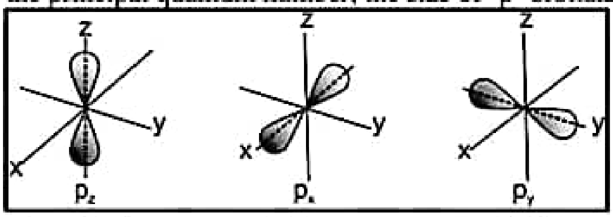
### (i) s-orbital

- s-orbital is spherical in shape. The electronic density around the nucleus in an s orbital is uniformly distributed in all directions.
- It is non-directional.
- It starts from first shell (n = 1)
- With the increase in the principal quantum number, the size of 's' orbital also becomes larger.




### (ii) p-orbitals

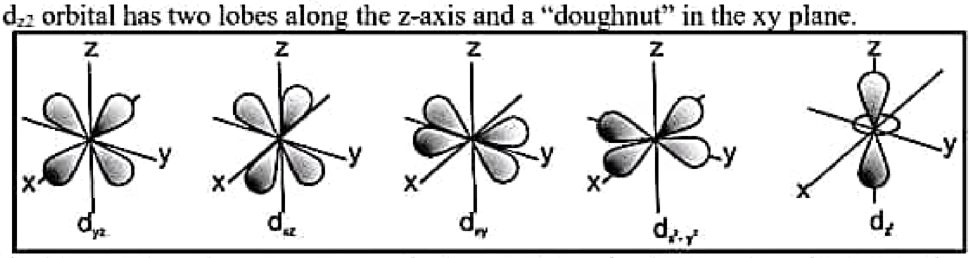
- p-orbitals have dumb-bell shape. The electron density is not distributed in a spherically symmetric fashion as in an s orbital.
- These are directional. A p-orbital has two lobes on any of the axis.
- It starts from second shell (n = 2)
- With the increase in the principal quantum number, the size of 'p' orbitals also becomes larger.




$l = 1, m = 0$ $l = 1, m = -1$ $l = 1, m = +1$

### (iii) d-orbitals

- d-orbitals have double dumb-bell or clover leaf shape.
- It starts from third shell (n = 3)
- In a given shell, 'd' orbitals have different shapes and orientations in space.
- The $d_{xy}$, $d_{xz}$, and $d_{yz}$ lie in the xy, xz, and yz planes, respectively.
- The lobes of the $d_{x^2-y^2}$ lie along the x and y axes.
- The $d_{z^2}$ orbital has two lobes along the z-axis and a "doughnut" in the xy plane.




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

### (iv) f-orbitals

- A f-subshell has seven orientations in space, i.e. there are seven f orbitals.
  ($f_{x^3}$, $f_{y^3}$, $f_{z^3}$, $f_{x(z^2-y^2)}$, $f_{y(z^2-x^2)}$, $f_{z(y^2-x^2)}$, $f_{xyz}$)
- The shapes of f orbitals are very complicated

**Quick Check 2.4**

**(a) What does an orbital represent according to the wave mechanical model of atom?**

**Ans:** In the wave mechanical model of the atom, an orbital represents a three-dimensional region around the nucleus where there is a high probability of finding an electron.

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

**Ans:** The magnetic quantum number (m) determines the orientation of an orbital in space.
For a p-orbital, the azimuthal quantum number (l) = 1.
For any value of l, the possible values of 'm' are from –l to +l, including 0.
So, for l = 1,
m = –1, 0, +1 → this gives three orientations of the p-orbital.
These three orientations are called:
• $p_x$ → along the x-axis • $p_y$ → along the y-axis • $p_z$ → along the z-axis
Hence, the three values of 'm' result in three spatial orientations of the p-orbitals.

## ELECTRONIC CONFIGURATION

**"The distribution of electrons among available shells, subshells, or orbitals of an atom or ion is called electronic configuration."**

### (1) Electronic Configuration in Shells

The electronic configuration of an atom describes the distribution of electrons in its atomic shells. The shells, denoted as K, L, M, N, and so on, correspond to the principal quantum number (n) of the orbitals.

Shell capacities each shell has a specific capacity for electrons:

K shell (n=1): 2 electrons maximum (1s orbital)
L shell (n=2): 8 electrons maximum (2s and 2p orbitals)
M shell (n=3): 18 electrons maximum (3s, 3p, and 3d orbitals)
N shell (n=4): 32 electrons maximum (4s, 4p, 4d, and 4f orbitals)

Examples: Hydrogen (H): $1s^1$ (K shell) , Helium (He): $1s^2$ (K shell)
Sodium (Na): $1s^2\ 2s^2\ 2p^6\ 3s^1$ (K, L, and M shells)

**Arrangement of orbitals according to (n + ℓ) rule**

| Orbital | n | ℓ | n + ℓ |
|---|---|---|---|
| 1s | 1 | 0 | 1 + 0 = 1 |
| 2s | 2 | 0 | 2 + 0 = 2 |
| 2p | 2 | 1 | 2 + 1 = 3 |
| 3s | 3 | 0 | 3 + 0 = 3 |
| 3p | 3 | 1 | 3 + 1 = 4 |
| 3d | 3 | 2 | 3 + 2 = 5 |
| 4s | 4 | 0 | 4 + 0 = 4 |
| 4p | 4 | 1 | 4 + 1 = 5 |
| 4d | 4 | 2 | 4 + 2 = 6 |
| 4f | 4 | 3 | 4 + 3 = 7 |
| 5s | 5 | 0 | 5 + 0 = 5 |
| 5p | 5 | 1 | 5 + 1 = 6 |
| 5d | 5 | 2 | 5 + 2 = 7 |
| 5f | 5 | 3 | 5 + 3 = 8 |
| 6s | 6 | 0 | 6 + 0 = 6 |
| 6p | 6 | 1 | 6 + 1 = 7 |
| 6d | 6 | 2 | 6 + 2 = 8 |
| 6f | 6 | 3 | 6 + 3 = 9 |
| 7s | 7 | 0 | 7 + 0 = 7 |
| 7p | 7 | 1 | 7 + 1 = 8 |
| 7d | 7 | 2 | 7 + 2 = 9 |
| 7f | 7 | 3 | 7 + 3 = 10 |
| 8s | 8 | 0 | 8 + 0 = 8 |

### (2) Distribution of electrons in sub-shells (orbitals)

**Auf bau Principle (Building up Principle)**

(Auf-bau is a German language word which means "building up")

This principle says;

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

Since, the energy of a subshell in the absence of any magnetic field, depends upon the principal quantum number (n) and the azimuthal quantum number (ℓ), hence the order of filling subshells with electrons may be obtained from the summation (n + ℓ).

### (n + ℓ) rule:

The following two points must be kept in mind while arrange the subshell in their increasing energy order:

**(a)** The subshell having lower (n + ℓ) value has lower energy and is filled first. For example, 4s orbital has (n + ℓ) = 4 + 0 = 4 and 3d orbital has (n + ℓ) = 3+2 = 5. Since (n + ℓ) value of 4s orbital is lower than that of 3d, hence 4s subshell has lower energy than 3d and 4s will be filled first.

**(b)** In case there are two subshells are having equal (n + ℓ) values, then the subshell with lower 'n' value will be filled first. For example, both 4p and 3d subshells have n + ℓ value equal to 5 (4p=4+1=5) and (3d = 3+2=5); 3d subshell will be preferred to be filled because of its low n value.
According to this rule the energy wise arrangement of orbitals should be.

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7P < 8s

**WAK Conceptual Corner**

| | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|
| سکول | سکول | پبلک | سکول | پبلک | سکول | ڈویژنل | پبلک | سکول | ڈویژنل |
| s | s | p | s | p | s | d | p | s | d |
| پبلک | سکول | فیڈرل | ڈویژنل | پبلک | سکول | فیڈرل | ڈویژنل | پبلک | سکول |
| p | s | f | d | p | s | f | d | p | s |

How to start numbering? 1s, 2p, 3d, 4f

## Pauli's Exclusion Principle

According to this principle:

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

***"Two electrons in an orbital will always have opposite spins".***

**Example:** Helium ($He = 1s^2$) has two electrons in s-orbital. According to Pauli's Exclusion Principle the value of quantum numbers of these two electrons are;

For First electron $n = 1$ $\ell = 0$, $m = 0$, $s = +1/2(\uparrow)$ clockwise

For second electron $n = 1$ $\ell = 0$, $m = 0$, $s = -1/2(\downarrow)$ anticlockwise

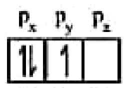
The two electrons having the same values of 'n', '$\ell$' and 'm' can have different values of 's'. It means that their spins are in the opposite directions.

**Pauli's exclusion principle** is called exclusion principle because the set of quantum numbers possessed by one electron in an atom is excluded for other electrons.

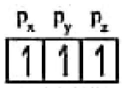
## Hund's Rule

***"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."***

**Explanation:** This rule gives an idea for filling electrons into the orbitals having equal energies. For example, three p-orbitals, i.e., px, py and pz have equal energy. To understand it, let us take an example in which three electrons are to be filled into three p-orbitals. There are two different ways to do this as shown below:




I (wrong)




II (right)

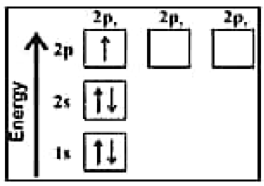
According to the Hund's rule, the correct way of filling three electrons in three p orbitals is that in which each orbital is singly occupied.

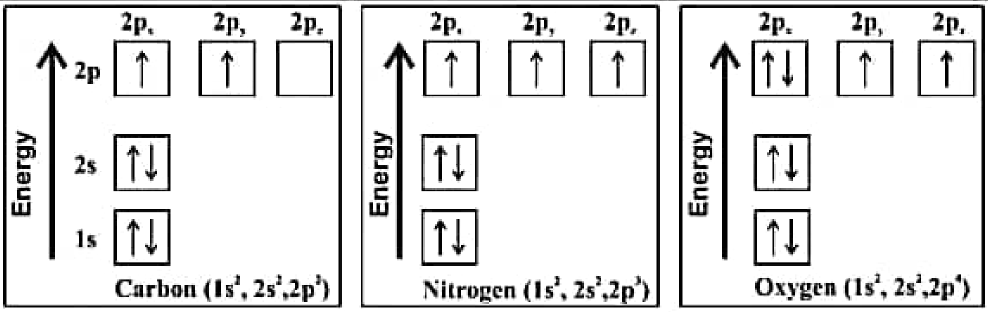
## Filling the Orbitals

A useful way of representing electronic configurations is a diagram that places electrons in boxes. Each box represents an atomic orbital.

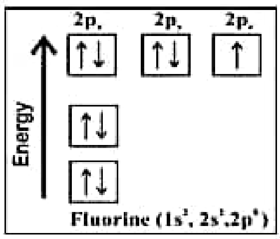
- The boxes (orbitals) can be arranged in order of increasing energy from bottom to top.
  - An electron is represented by an arrow.
- The direction of the arrow represents the 'spin' of the electron.

**Explanation:** When there are two electrons in an orbital, the 'spins' of the electrons are opposite, so the two arrows in this box point in opposite direction. Electrons in the same region of space repel each other because they have the same charge. So wherever possible, electrons will occupy separate orbitals in the same subshell to minimize this repulsion. These electrons have the same 'spins' in different orbitals. Electrons are only paired when there are no more empty orbitals available within a subshell. The electronic structures of carbon, nitrogen and oxygen are illustrated these points in the following.






**Exception:** In case of electronic configuration of fluorine, the Hund's rule is not applicable.




| Z | Element | Configuration | Z | Element | Configuration |
|---|---|---|---|---|---|
| 1 | H | $1s^1$ | 19 | K | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$ |
| 2 | He | $1s^2$ | 20 | Ca | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$ |
| 3 | Li | $1s^1, 2s^1$ | 21 | Sc | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^1, 4s^2$ |
| 4 | Be | $1s^2, 2s^2$ | 22 | Ti | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^2, 4s^2$ |
| 5 | B | $1s^2, 2s^2, 2p^1$ | 23 | V | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^3, 4s^2$ |
| 6 | C | $1s^2, 2s^2, 2p^2$ | 24 | Cr | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5, 4s^1$ |
| 7 | N | $1s^2, 2s^2, 2p^3$ | 25 | Mn | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5, 4s^2$ |
| 8 | O | $1s^2, 2s^2, 2p^4$ | 26 | Fe | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^6, 4s^2$ |
| 9 | F | $1s^2, 2s^2, 2p^5$ | 27 | Co | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^7, 4s^2$ |
| 10 | Ne | $1s^2, 2s^2, 2p^6$ | 28 | Ni | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^8, 4s^2$ |
| 11 | Na | $1s^2, 2s^2, 2p^6, 3s^1$ | 29 | Cu | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^1$ |
| 12 | Mg | $1s^2, 2s^2, 2p^6, 3s^2$ | 30 | Zn | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2$ |
| 13 | Al | $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$ | 31 | Ga | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^1$ |
| 14 | Si | $1s^2, 2s^2, 2p^6, 3s^2, 3p^2$ | 32 | Ge | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^2$ |
| 15 | P | $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$ | 33 | As | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^3$ |
| 16 | S | $1s^2, 2s^2, 2p^6, 3s^2, 3p^4$ | 34 | Se | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^4$ |
| 17 | Cl | $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$ | 35 | Br | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^5$ |
| 18 | Ar | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ | 36 | Kr | $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6$ |

## ELECTRONIC CONFIGURATION AND THE PERIODIC TABLE

### Relating Electron Configuration to Periodic Position

- The electron configurations of elements are related to their position in the periodic table.
- In periodic table, the elements with the same pattern of outer-shell (valence) electron configuration are arranged in same groups.
- The elements can be grouped in terms of the type of orbital into which the electrons are placed.
- On the basis of location in the periodic table, the electron configuration of an element can easily be written.

### Relationship Between Electron Configuration and Groups

(i) **Alkali metals and Alkaline earth metals** (groups 1 and 2), are those in which the outer-shell s orbitals are being filled. The group 1 and 2 elements all have $ns^1$ and $ns^2$ outer configurations respectively.

(ii) **Group 13 elements** have $ns^2np^1$ configuration. On the right is a block of six columns. These are the elements in which the outermost 'p' orbitals are being filled.

(iii) **Transition metals** (a block of 10 columns in the middle of the table) contain the d orbitals which are being filled.

(iv) **f-block elements** are two rows that contain fourteen elements each, below the main portion of the table. These are the ones in which the 'f' orbitals are being filled.

| 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| $ns^1$ | $ns^2$ | $ns^2(n-1)d^1$ | $ns^2(n-1)d^2$ | $ns^2(n-1)d^3$ | $ns^1(n-1)d^5$ | $ns^2(n-1)d^5$ | $ns^2(n-1)d^6$ | $ns^2(n-1)d^7$ | $ns^2(n-1)d^8$ | $ns^1(n-1)d^{10}$ | $ns^2(n-1)d^{10}$ | $ns^2 np^1$ | $ns^2 np^2$ | $ns^2 np^3$ | $ns^2 np^4$ | $ns^2 np^5$ | $ns^2 np^6$ |
| s Block elements | | d-Block elements | | | | | | | | | | p-Block elements | | | | | |
| f block elements | → Lanthanides $4f^{1-14}5d^{0-1}6s^2$ | | | | | | | | | | | | | | | | |
| | → Actinides $5f^{0-14}6d^{0-2}7s^2$ | | | | | | | | | | | | | | | | |

**Quick Check 2.5**

**a) With the help of periodic table, write the electronic configurations for the following elements by giving the appropriate noble-gas inner core plus the electrons beyond it (i) $_{48}Cd$; (ii) $_{57}La$.**

**Ans:** **(i)** $_{48}Cd$, atomic number = 48

Nearest noble gas before Cd = [Kr] (Atomic number 36)

$_{48}Cd$ = [Kr] $4d^{10}5s^2$

**(ii)** $_{57}La$, atomic number = 57

Nearest noble gas before La = [Xe] (Atomic number 54)

$_{57}La$ = [Xe] $5d^1\ 6s^2$

**b) Write the complete electron configuration for antimony (Sb) with atomic number 51.**

**Ans:** $_{51}Sb = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 5s^2, 5p^3$

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

**Ans:** $_{51}Sb = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 5s^2, 5p_x^1, 5p_y^1, 5p_z^1$

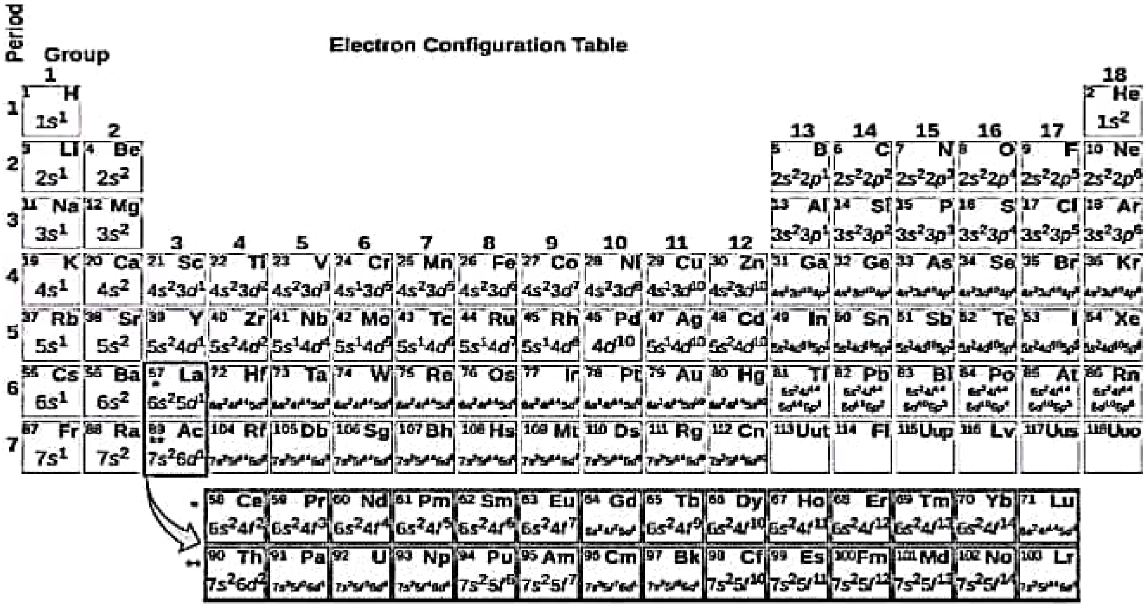
There are three unpaired electrons in p-orbitals

- Q. Compare orbit and orbital.
- Q. What is the function of Principal quantum number?
- Q. Define Principal Quantum Number and give its values.
- Q. When azimuthal quantum number has a value 3, then there are seven values of magnetic quantum number. Give reasons.
- Q. Define spin quantum number.
- Q. Draw the shapes of 's' and 'p' orbitals.
- Q. Distribute electrons in orbitals of $_{57}La$, $_{29}Cu$, $_{79}Au$, $_{24}Cr$, $_{53}I$, $_{86}Rn$.
- Q. State Auf-bau principle. Write electronic configuration of Sodium ($_{11}Na$) following this principle.
- Q. State Pauli's exclusion principle and Hund's rule.
- Q. Distribute electrons in orbitals of $_{29}Cu$ and $_{20}Ca$.
- Q. Write Electronic Configuration of Na = 11 and Cr = 24.

## Valence Electrons

"The electrons in an atom in the outermost shell are called valence electrons."

- Valance electrons are primarily involved in chemical reactions.
- The similarities among the configurations of valence electrons account for similarities of the chemical properties among groups of elements.




## Classification of elements of periodic table

On the basis of electronic configuration and their valence electrons, the elements of periodic table has been classified into;

**(i) Metals and Non-metals**

Metals are typically located on the left and in the middle, while nonmetals are on the right. Between them metalloid exist.

**(ii) Representative and Transition Elements**

- The main-group or representative elements all have valence-shell configurations $ns^a\ np^b$. They have some choice of 'a' and 'b'. In other words, the outer 's' or 'p' subshell is being filled.
- the d-block transition elements, 'd' subshell is being filled. In the f-block transition elements or inner-transition elements, 'f' subshell is being filled.

**(iii) Periods**

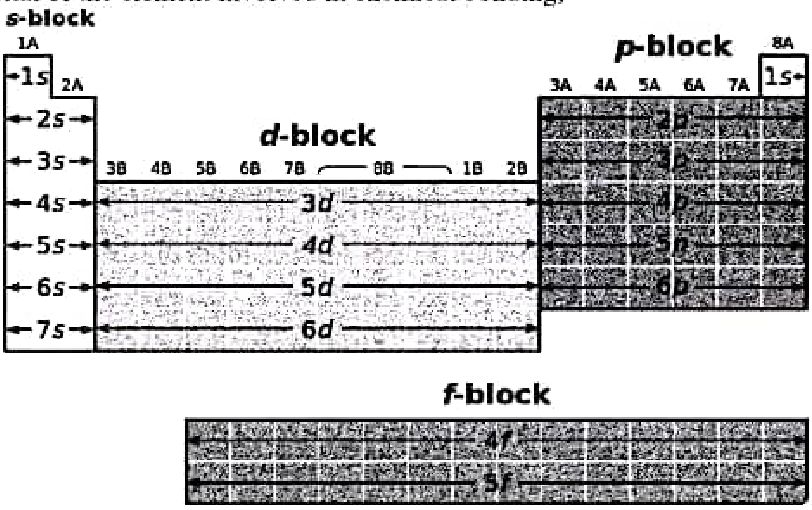
In the periodic table, periods are the horizontal rows of elements. There are seven periods in the periodic table. Each period represents a new principal energy level filling with electrons.

**(iv) Groups**

A vertical column of elements is called group. Elements within the same group tend to have similar chemical properties because they have the same number of valence electrons.

**(v) Blocks**

Elements in the periodic table can be classified into four blocks. This classification is based upon the valence orbital of the element involved in chemical bonding,



**Quick Check 2.6**

**a)** **An element has the electronic configuration: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 5s^2, 5p^5$.**
**(i) Which block in the Periodic Table does this element belong to?**
**(ii) Which group does it belong to?** **(iii) Which period does it belong to?**
**(iv) Identify this element.**

**Ans:** **(i) Block**
The last electron enters the **p-orbital**, so the element belongs to the **p-block**.
**(ii) Group:**
Count the electrons in the outermost shell (n = 5):
$5s^2\ 5p^5$ = **7 valence electrons,** Group = **Group 17** (Halogens)
**(iii) Period:**
The highest principal quantum number (n) is **5**, so it is in the **5th period**.
**(iv) Element:**
Number of electrons = atomic number (Z),
Total electrons = 53 → Atomic number = **53**
The element is **Iodine (I)**

**b)** **Which block in the periodic table does the element with the electronic configuration? $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$ belong to?**

**Ans:** The electronic configuration shows that the last electrons are entering the 3d and 4s orbitals. Among these, the 3d subshell is being actively filled. Since d-orbital is being filled, the element belongs to the d-block.

## ELECTRONIC CONFIGURATION OF IONS

**Positive ions**

- Positive ions are formed when electrons are removed from atoms.
- Electrons in the outer subshell are removed when metal atoms form the positive ions.

**Examples:**

(i) $Na_{(2,8,1)} \longrightarrow Na^+_{(2,8)} + 1e^-$

(ii) $Ca_{(2,8,8,2)} \longrightarrow Ca^{2+}_{(2,8,8)} + 2e^-$

**Positive ions formation in d-block elements**

In general, when metal atoms form the positive ions, electrons in the outer subshell are removed. However, the d-block elements behave slightly differently. Moving across the Periodic Table from potassium to zinc, the 4s subshell fills before the 3d subshell. But when atoms of a d-block element lose electrons to form ions, the 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:** $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^3$

**Negative ions**

- Negative ions are formed when atoms gain electrons.
- Electrons in the outer subshell are added when non-metal atoms form the negative ions.

**Examples:**

(i) $O_{(2,6)} + 2e^- \longrightarrow O^{2-}_{(2,8)}$

(ii) $Cl_{(2,8,7)} + 1e^- \longrightarrow Cl^-_{(2,8,8)}$

## ELECTRONIC CONFIGURATION OF FREE RADICALS

**Free Radicals**

***"A species that has one or more unpaired electrons is called free radical."***

**Examples:**

**(i)** Free radicals from single atoms: $\dot{Cl}$, $\dot{H}$ etc.
**(ii)** Free radicals from group of atoms: $\dot{C}H_3$, $\dot{O}H$ etc.

**Formation of chlorine free radical**

The electron configuration of chlorine free radical is $1s^2, 2s^2, 2p^6, 3\overset{\uparrow\downarrow}{s}, 3\overset{\uparrow\downarrow}{p_x}, 3\overset{\uparrow\downarrow}{p_y}, 3\overset{\uparrow}{p_z}$. In the 2p subshell, two orbitals have paired electrons whereas, the third one contains a single unpaired electron. The unpaired electron is shown by a single dot as in $\dot{Cl}$.

## ELECTRONIC CONFIGURATION AND THE FORMATION OF SEMICONDUCTORS

### Semiconductors

***"A material that conducts electricity under certain conditions is called semiconductor."***

- They are widely used in electronic devices like smartphones, laptops, and cars.
- Common semiconductor elements include Silicon (Si), Germanium (Ge) and Arsenic (As).

**Role of Electronic Configuration in Semiconductors**

The formation of semiconductors is due to the unique electronic configuration of these elements. e.g. Silicon (Si) has the configuration 2, 8, 4. It has 4 valence electrons, allowing it to form four covalent bonds in a crystal structure. In pure silicon crystals, every Si atom is bonded to four others, so there is no possibility of electronic conduction through the Si crystal.

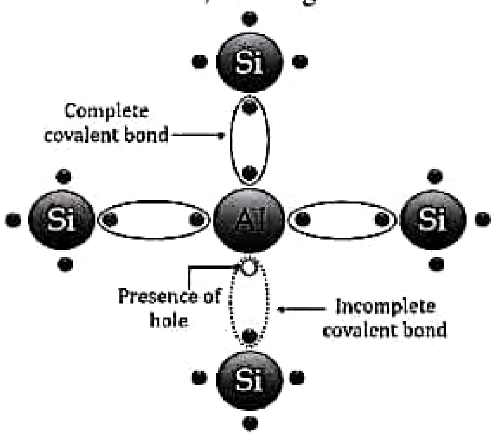
### Doping

***"The process of adding impurity atoms to a pure semiconductor material to change its electrical properties is called doping."***

There are two types of doped semiconductors: (i) P-type (positive-type) (ii) N-type (negative-type)

### P-Type Semiconductor Formation

- This type of semiconductor is formed by adding trivalent impurities (like Boron or aluminium) in a pure semiconductor.
- These atoms replace some silicon atoms but can form only three bonds, leaving one bond incomplete.
- This creates a "hole" (missing electron) in the crystal, which acts as a positive charge carrier.
- Electrons from an external source fill these holes, allowing the material to conduct electricity.

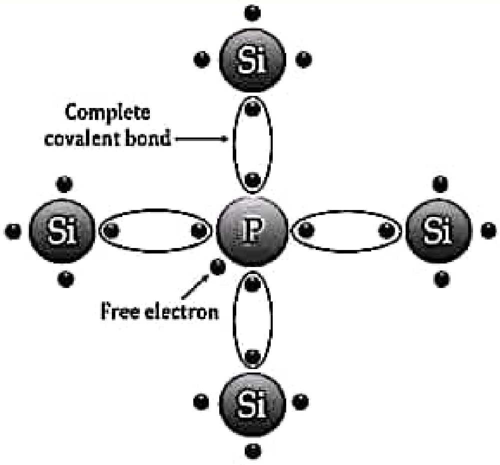



- Si = Intrinsic semiconductor atom
- Al = Trivalent impurity atom

Formation of P type extrinsic semiconductor

### N-Type Semiconductor Formation

- This type of semiconductor is formed by adding pentavalent impurities (like phosphorus) in a pure semiconductor.
- These atoms replace some silicon atoms and form four bonds, leaving one extra electron.
- The extra electron becomes free to move in the crystal lattice as a negative charge carrier.
- These free electrons enable the material to conduct electricity when connected to a power source.




- Si = Intrinsic semiconductor atom
- P = Pentavalent impurity atom

Formation of N type extrinsic semiconductor

## WAK CONCEPTUAL LEARNING

| Orbit | Shell |
|---|---|
| • Circular path around the nucleus. | • Spherical path around the nucleus. |
| • Two dimensional like wheel. | • Three dimensional like football. |
| • has fixed energy throughout within an orbit. | • has variable energy in a shell. |
| • Introduced by Bohr | • Introduced by Schrodinger and Dirac |

### Hund's Rule

- Two electrons in the same orbitals have a higher energy than two electrons in different orbitals because of the electrostatic repulsion.
- Electrons in the same orbital repel each other more than electrons in separate orbitals.
- Each electron is shown as an arrow, indicating its spin: either $\downarrow$ or $\uparrow$.
- Within an orbital, the electrons must have opposite spins.

allowed

$\boxed{\uparrow\downarrow}$ $\boxed{\downarrow\uparrow}$

not allowed

$\boxed{\uparrow\uparrow}$ $\boxed{\downarrow\downarrow}$

### Cause of greater stability of exactly half-filled and completely filled configurations

The greater stability of these configurations is due to the following two reasons:

**(i) Symmetry:**

The half-filled and completely filled configurations are more symmetrical and symmetry leads to greater stability.

**(ii) Exchange energy:**

The electrons present in the different orbitals of the same subshell can exchange their positions. Each such exchange leads to a greater stability which can be explained in terms of exchange energy. As the number of exchanges that can take place is maximum in the exactly half-filled and completely filled arrangements (i.e. more in $d^5$ than in $d^4$ and more in $d^{10}$ than in $d^9$), therefore exchange energy is maximum and hence the stability is maximum.

# MULTIPLE CHOICE QUESTIONS

## EXERCISE MCQs

**(i) The quantum number 'm' of a free gaseous atom is associated with:**
(a) the effective volume of the orbital
(b) the shape of the orbital
(c) the spatial orientation of the orbital
(d) the energy of the orbital in the absence of a magnetic field

**(ii) When 3d subshell is completely filled, the next entering electron goes into:**
(a) 4f (b) 4s
(c) 4p (d) 4d

**(iii) Quantum number values for 2p orbitals are:**
(a) $n = 2, \ell = 1$ (b) $n = 1, \ell = 2$
(c) $n = 1, \ell = 0$ (d) $n = 2, \ell = 0$

**(iv) An electron having the set of values: $n = 4$, $\ell = 0$, $m = 0$ and $s = +1/2$ lies in:**
(a) 2s (b) 3s
(c) 4s (d) 5s

**(v) The quantum number values for the fourth electron of $^{9}_{4}Be$ atom are:**
(a) 1, 0, 0 (b) 2, 0, 0
(c) 2, 1, 0 (d) 1, 1, 1

**(vi) The correct order of first ionization energies is:**
(a) F > He > Mg > N > O (b) He > F > N > O > Mg
(c) He > O > F > N > Mg (d) N > F > He > O > Mg

**(vii) A p-orbital has a characteristic shape with how many lobes?**
(a) 1 (b) 2
(c) 3 (d) 4

**(viii) The three p-orbitals in a given energy level are oriented:**
(a) Along the same axis.
(b) At 45° to each other.
(c) Mutually perpendicular to each other along the x, y, and z axes.
(d) In a complex tetrahedral arrangement.

**(ix) How many d-orbitals are there in a given energy level?**
(a) 1 (b) 3
(c) 5 (d) 7

**(x) Which of the following species is predicted to have the highest bond order?**
(a) $N_2$ (b) $O_2$
(c) $F_2$ (d) $Ne_2$

**(xi) Which of the following species has a zero bond order according to MOT?**
(a) $H_2^{1+}$ (b) $He_2^{1+}$
(c) $He_2^{2+}$ (d) $He_2$

**(xii) Which of the following molecular geometries has bond angles of approximately 120°?**
(a) (V-shaped) (b) Trigonal planar
(c) Tetrahedral Bent (d) Trigonal pyramidal

## ADDITIONAL PRACTICE MCQs

**1. In the ground state of an atom, the electron is present:**
(A) in the nucleus
(B) in the second shell
(C) nearest to the nucleus
(D) farthest from the nucleus

**2. Orbitals having same energy are called:**
(A) hybrid orbitals (B) valence orbitals
(C) degenerate orbitals (D) d-orbitals

**3. When 6d orbital is complete, the entering electron will go to:**
(A) 7f (B) 7s
(C) 7p (D) 7d

**4. A photon knocks a proton out of $^{24}_{12}Mg$ nucleus to form:**
(A) the isotope of parent nucleus
(B) the isobar of parent nucleus
(C) the nuclide $^{23}_{11}Na$
(D) the isobar of $^{23}_{11}Na$

**5. Calculate the number of neutrons in the $^{238}_{90}X$ atom produced after the emission of alpha particle:**
(A) 152 (B) 146
(C) 185 (D) 135

**6. The triad of nuclei that is isotonic is:**
(A) $^{14}_{6}C, ^{15}_{7}N, ^{17}_{9}F$ (B) $^{14}_{6}C, ^{14}_{7}N, ^{17}_{9}F$
(C) $^{12}_{6}C, ^{14}_{7}N, ^{19}_{9}F$ (D) $^{14}_{6}C, ^{14}_{7}N, ^{19}_{9}F$

**7. Which pair of ions has the same total number of electrons?**
(A) $Na^+$ and $Mg^{2+}$ (B) $O^-$ and $O^{2-}$
(C) $F^-$ and $C\ell^-$ (D) $Ga^{3+}$ and $Fe^{3+}$

**8. Which property is the same for the two nuclides $^{66}_{29}Cu$ and $^{66}_{30}Zn$?**
(A) number of electrons (B) number of neutrons
(C) number of nucleons (D) number of protons

**9. The atomic number of an element is 35. What is the total number of electrons present in all the p orbitals of the ground state atom of that element?**
(A) 17 (B) 23
(C) 11 (D) 6

**10. The correct set of four quantum numbers for the valence electrons of rubidium atom (Z = 37) is:**
(A) 5, 0, 0, +1/2 (B) 5, 1, 1, +1/2
(C) 5, 1, 0, +1/2 (D) 5, 0, 1, +1/2

**11. Which of the following sets of quantum numbers is correct for an electron in 4f orbital?**
(A) $n = 4, \ell = 3, m = +4, s = +(1/2)$
(B) $n = 4, \ell = 3, m = +1, s = +(1/2)$

(C) n = 3, $\ell$ = 2, m = –2, s = +(1/2)
(D) n = 4, $\ell$ = 4, m = –4, s = –(1/2)

**12. Degenerate orbitals are those that contain same:**
(A) wave function (B) energy
(C) orientation (D) shape

**13. In a chromium atom, how many electrons have zero azimuthal quantum number?**
(A) 1 (B) 7
(C) 6 (D) 8

**14. Which of the following electronic configuration is not possible?**
(A) [Kr] $4d^8\ 5s^2$ (B) [Xe] $4f^{14}\ 5d^7\ 6s^2$
(C) [Kr] $3d^7\ 4s^2$ (D) [Ar] $3d^5\ 4s^1$

**15. The relative energies of 4s, 4p and 3d orbitals are in the order:**
(A) 4d < 4p < 4s (B) 4s < 3d < 4p
(C) 4p < 4s < 3d (D) 4p < 3d < 4s

**16. Which set has same number of unpaired electrons?**
(A) $Fe^{2+}$, $Mn^{+2}$ (B) $Fe^{3+}$, $Mn^{2+}$
(C) $Cr^{3+}$, $Ni^{2+}$ (D) $Zn^{2+}$, $Cu^{2+}$

**17. Quantum number of an atom can be defined on the basis of:**
(A) Aufbau's Principle
(B) Heisenberg's uncertainty Principle
(C) Hund's rule
(D) Pauli's exclusion Principle

**18. How many electrons can be accommodated in p-orbitals?**
(A) 6 electrons (B) 2 electrons
(C) 4 electrons (D) 10 electrons

**19. The orbital diagram in which Hund's rule is violated is:**
(A) $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\phantom{\uparrow}}$ (B) $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$
(C) $\boxed{\uparrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$ (D) $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$

**20. When l = 2, the number of possible orbitals in sub-shell is:**
(A) 3 (B) 5
(C) 09 (D) 10

**21. Which of the following quantum number explains the shape of orbitals:**
(A) Principal Q No. (B) Azimuthal Q No.
(C) Magnetic Q No. (D) Spin Q No.

**22. What is the electronic configuration for $Fe^{2+}$?**
(A) $[Ar]4s^2 3d^6$ (B) $[Ar]4s^2 3d^4$
(C) $[Ar]4s^0 3d^6$ (D) $[Ar]4s^2 3d^4$

**23. The maximum number of electrons with n = 3 and $\ell$ = 2 is:**
(A) 10 (B) 6
(C) 8 (D) Zero

**24. In which of the following orbitals, lobes lies on the axis:**
(A) $d_{xy}$ (B) $d_{yz}$
(C) $d_{zx}$ (D) $d_{x^2-y^2}$

**25. Which electronic level would allow the hydrogen atom to absorb a photon but not emit a photon?**
(A) 1s (B) 2s
(C) 2p (D) 3s

**26. The maximum number of electrons with clockwise spin that can be accommodated in "N" shell is:**
(A) 16 (B) 32
(C) 18 (D) 8

**27. Hund's rule is not followed in the electronic configuration of which of the following element?**
(A) Carbon (B) Nitrogen
(C) Oxygen (D) Fluorine

**28. For d-subshell, the azimuthal quantum number has a value of:**
(A) 1 (B) 2
(C) 3 (D) 4

**29. The number of electrons in a given sub-shell is given by:**
(A) $2n^2$ (B) $n^2$
(C) $2(2\ell - 1)$ (D) $4\ell + 2$

**30. Electronic configuration of $H^-$ is:**
(A) $1s^2$ (B) $1s^2 2s^1$
(C) $1s^1$ (D) $1s^2 2s^2$

**31. The number of orbitals present in a sublevel is given by the formula:**
(A) $2(2\ell + 1)$ (B) $2\ell + 1$
(C) $n^2$ (D) $2n^2$

**32. In spectrum, the spin motion of electron is responsible for:**
(A) Spectral lines (B) Fine structures
(C) Double line structure (D) Degeneracy

**33. The number of neutrons present in 19K39 is:**
(A) 18 (B) 19
(C) 20 (D) 39

**34. (n + $\ell$) value for 4p orbital is:**
(A) 4 (B) 5
(C) 6 (D) 7

**35. The quantum number which gives information about degeneracy of orbitals in space is:**
(A) Principal quantum number
(B) Azimuthal quantum number
(C) Magnetic quantum number
(D) Spin quantum number

## ANSWERS OF MULTIPLE CHOICE QUESTIONS

### Exercise MCQs

| | | | | | | | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| (i) | c | (ii) | c | (iii) | a | (iv) | c | (v) | b | (vi) | b | (vii) | b | (viii) | c | (ix) | c | (x) | a |
| (xi) | d | (xii) | b | | | | | | | | | | | | | | | | |

### Additional Practice MCQs

| | | | | | | | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 1. | C | 2. | C | 3. | C | 4. | C | 5. | B | 6. | A | 7. | A | 8. | C | 9. | A | 10. | A |
| 11. | B | 12. | B | 13. | B | 14. | C | 15. | B | 16. | B | 17. | D | 18. | C | 19. | B | 20. | B |
| 21. | B | 22. | C | 23. | A | 24. | D | 25. | A | 26. | A | 27. | D | 28. | B | 29. | D | 30. | A |
| 31. | B | 32. | C | 33. | C | 34. | B | 35. | C | | | | | | | | | | |

## SHORT ANSWER QUESTIONS

### Exercise Short Answers Questions

**Q.2a. There are three orientations of p-orbital due to three values of magnetic quantum number. Justify it.**

**Ans:** The relation between azimuthal and magnetic quantum numbers is;

$$m = 2\ell+1$$

When $\ell = 1$ (p-subshell) then m = –1, 0, +1. These represent the three orientation of p-orbitals i.e. $p_x$, $p_y$, and $p_z$.

**Q.2b. Size of Mg is bigger than Al, but ionization energy of Mg is more than that of Al. Why?**

**Ans:** Although size of Mg (160pm) is bigger than Al (143pm) but Mg has a higher ionization energy (738 kJ/mol) than Al (577.5 kJ/mol). The reason is that electron in Mg is removed from a more stable, paired 2s orbital, while electron in Al is removed from a less stable, more extended 3p orbital.

**Q.2c. '$I_3$' of Mg is much bigger than its '$I_2$'. Justify.**

**Ans:** After the removal of second electron, $Mg^{2+}$ ion is formed which has a stable electron configuration ($1s^2$, $2s^2$, $2p^6$) and reduced size. Removing the third electron from this stable configuration requires significantly more energy. Therefore, $I_3$ of Mg is much higher than $I_2$.

$$Mg^+ \longrightarrow Mg^{2+} + e^- \qquad I_2 = 1450\ kJ\ mol^{-1}$$
$$Mg^{2+} \longrightarrow Mg^{3+} + e^- \qquad I_3 = 7730\ kJ\ mol^{-1}$$

**Q.2d. Among the elements Li, K, Ca, S and Kr which one has the lowest first ionization energy? Which has the highest first ionization.**

**Ans:** **Lowest $I_1$:** Potassium (K): It has the largest atomic size and least nuclear attraction.

**Highest $I_1$:** Krypton (Kr): It has a full outer shell and strong nuclear attraction.

**Q.2e. Consider the electronic configuration of the potassium atom (atomic number 19).**

**(i) Write the full electronic configuration of potassium using the s, p, d, f notation.**

**(ii) Explain why the 4s subshell is filled before the 3d subshell in potassium, even though the principal quantum number of the 3d subshell is lower.**

**Ans:** **(i)** Electronic configuration of potassium, $_{19}K$: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

**(ii)** According to Auf bau principle, the 4s subshell is filled before 3d because it has lower energy than 3d at the time of filling.

**Q.2f. i) An atom of element X has an atomic number of 17 and a mass number of 35. Determine the number of protons, neutrons, and electrons in this atom.**

**(ii) If this element forms an ion with a charge of -1, how many protons, neutrons, and electrons will be present in the ion?**

**Ans:** **(i)** $^{35}_{17}X$

Number of protons = 17

Number of electrons = 17

**Number of neutrons = Mass number – Atomic number**

Number of neutrons = 35 – 17 = 18

(ii) $^{35}_{17}X$ atom loses 1 electrons to form $X^-$ ion.

Number of protons = 17

Number of neutrons = 18

Number of electrons = 17 + 1 = 18

**Q.2g. In the ground state of mercury $_{80}Hg$:**

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

**ii. How many electrons occupy 4d atomic orbitals?**

**iii. How many electrons occupy $4p_z$ atomic orbital?**

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

**Ans:** $_{80}$**Hg** = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2\ 3d^{10}, 4p^6, 4d^{10}, 5s^2, 5p^6, 4f^{14}, 5d^{10}, 6s^2$

**(i)** 18 electrons occupy orbitals with n = 3 ($3s^2\ 3p^6\ 3d^{10}$)

**(ii)** 10 electrons occupy 4d orbitals

**(iii)** 2 electrons occupy the $4p_z$ orbital

**(iv)** 40 electrons have spin "up" (s = +1/2) and 40 electrons have spin "down" (s = – 1/2) in Hg (Z = 80)

**Q.2h. The successive ionization energies for an unknown element are:**

**$I_1$ = 896 kJ/mol, $I_2$ = 1752 kJ/mol**

**$I_3$ = 14,807 kJ/mol $I_4$ = 17,948 kJ/mol**

**To which family in the periodic table, does the unknown element most likely belong?**

**Ans:** There is a small increase from $I_1$ to $I_2$, but a very large jump at $I_3$. This indicates the element has 2 valence electrons, and the third electron is being removed from a stable inner shell. Therefore, the unknown element most likely belongs to Group 2 (alkaline earth metals) i.e. Be.

**Q.2i Consider the following ionization energies for aluminum:**

$Al_{(g)} \rightarrow Al^+_{(g)} + e^-$ $I_1 = 580$ kJ/mol

$Al^+_{(g)} \rightarrow Al^{2+}_{(g)} + e^-$ $I_2 = 1815$ kJ/mol

$Al^{2+}_{(g)} \rightarrow Al^{3+}_{(g)} + e^-$ $I_3 = 2740$ kJ/mol

$Al^{3+}_{(g)} \rightarrow Al^{4+}_{(g)} + e^-$ $I_4 = 11{,}600$ kJ/mol

**(i) Account for the trend in the values of the ionization energies.**

**(ii) Explain the large increase between $I_3$ and $I_4$.**

**(iii) List the four aluminum ions given in order of increasing size, and explain your ordering.**

**Ans:** **(i)** Ionization energy increases with each successive removal because the electron is removed from an increasingly positive ion, which holds electrons more tightly.

**(ii)** The large increase from $I_3$ to $I_4$ is due to removing an electron from a stable noble gas core ($1s^2\ 2s^2\ 2p^6$), which requires much more energy.

**(iii)** Order of increasing size: $Al^{4+} < Al^{3+} < Al^{2+} < Al^+$

More positive charge pulls electrons closer, reducing size.

**Q.2j. (a) State the general order of filling orbitals up to the 4p subshell.**

**(b) Explain why the 4s subshell is filled before the 3d subshell, according to the Aufbau principle.**

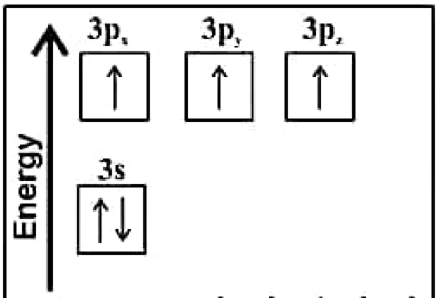
**Ans:** **(a)** General order of filling orbitals up to 4p:

**1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p**

**(b)** The 4s subshell is filled before the 3d subshell because, according to the Aufbau principle, electrons occupy the orbitals in order of increasing energy. Although the principal quantum number of 3d (n = 3) is lower than that of 4s (n = 4), the 4s orbital has a low ($n + \ell$) value than 3d orbital.

**Q.2k. Draw the orbital box diagram for the valence electrons of a phosphorus atom (atomic number 15), ensuring that your diagram adheres to Hund's rule and the Pauli Exclusion Principle.**

**Ans:**




Phosphorus ($1s^2, 2s^2, 2p^6, 3s^2, 3p^3$)

## Additional Short Answer Questions

**1. What is Moseley's law?**

**Ans: Moseley's Law**

"The frequency of a spectral line in the X-rays spectrum varies as the square of atomic number of an element emitting it."

Mathematically, it can be written as:

$$\sqrt{\upsilon} = a\,(Z - b)$$

Where $\upsilon$ = frequency of spectral line, $Z$ = atomic number of target metal,
$a$ = proportionality constant, $b$ = screening constant

**2. Why atomic spectrum is line spectrum?**

**Ans: Atomic Spectrum is Line Spectrum**

- When the atoms of an element are heated in flame or subjected to electric discharge, radiations of certain wavelengths are emitted which appear as bright lines against dark background.
- When white light is passed through vapours of element, it absorbs certain wavelengths which appear as dark lines against bright background.
- In this way, the atomic spectrum has distinct coloured lines separated by dark spaces, and it is line spectrum.

**3. Compare line emission and line absorption spectra.**

**Ans: Difference between emission and absorption spectra**

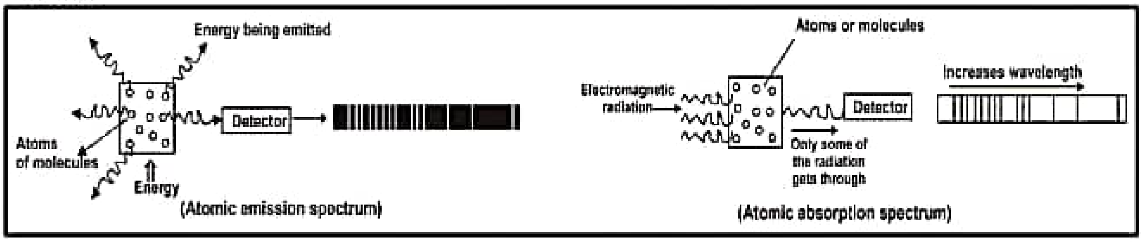
| Line emission spectrum | Line absorption spectrum |
|---|---|
| 1. It is spectrum formed by elements or their compounds when they are heated in flame or subjected to electric discharge. | 1. When beam of white light is passed through vapours of element, it absorbs certain wavelengths, rest of wavelength pass thrown it. |
| 2. Sample is in gaseous and excited state. | 2. Sample is in gaseous, liquid or solid state. |
| 3. Spectrum of this radiation consists of bright lines against a dark background. | 3. Here spectrum appears in the form of dark lines on bright background. |

**Similarities of emission and absorption spectra:**

(i) Absorption spectrum is the photographic negative of emission spectrum.

(ii) The dark lines in absorption spectrum appears exactly at the same position where the bright lines appear in emission spectrum for a same sample.

(iii) Both spectra are actually the finger prints of atoms of various elements and help in identification of elements.



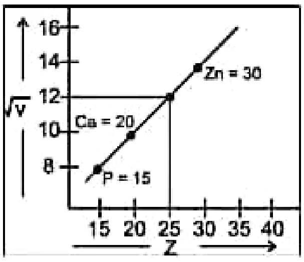
**4. What are X-rays? What is their origin? How was the idea of atomic number derived from the discovery of X-rays?**

**Ans: X-Rays:**

X-rays are the electro-magnetic radiations with high frequency but wavelength short than visible light.

**Origin of X-rays**

When cathode rays are converged on a metal target (taken as anode), then highly penetrating radiations are produced called X-rays. Their frequency is greater but wavelength is shorter than visible light.




**X-rays and Atomic Number (Z):**

The frequency of X-rays depends upon the proton number (Z) of an element and is the characteristic of an element. It is given by Moseley's law as

$$\sqrt{\upsilon} \propto Z$$

When a graph is plotted between $\sqrt{v}$ on y-axis and Z on x-axis, a straight line is obtained for different elements.

**5. Compare orbit and orbital.**

**Ans:**

| Orbit | Orbital |
|---|---|
| 1. It is a definite path in which an electron moves around nucleus as proposed by Bohr. | 1. It is three-dimensional space around nucleus where the probability of finding electron is maximum (95%). |
| 2. It is circular in shape. | 2. Orbitals have different shapes e.g., s-orbital is spherically symmetrical, p-orbitals have dumb bell shape. |
| 3. It does not have a directional character. | 3. All orbitals have directional characters except s-orbital. |
| 4. In an orbit, exact position and momentum of an electron can be measured with certainty. | 4. In an orbital, exact position and momentum of an electron cannot be measured with certainty. |
| 5. It represents planar motion of electrons. | 5. It represents three-dimensional motion of electrons. |
| 6. Maximum no. of electrons in an orbit is given by $2n^2$. | 6. An orbital can have a maximum of 2 electrons. |

**6. What is the function of Principal quantum number?**

**Ans: Function of Principal Quantum Number:**

Principal quantum number tells about:

(i) **Size of an orbit:** "When 'n' is larger, the electron stays farther from the nucleus, so the orbital becomes larger."

(ii) **Energy of the orbit:** An increase in 'n' also means that the electron has a higher energy and is therefore less tightly bound to the nucleus.

(iii) **Number of electrons in a shell** = $2n^2$

**7. Define Principal Quantum Number and give its values.**

**Ans: Principal Quantum Number (n)**

This quantum number, n, describes the size and energy of the orbital.

- Its values are non-zero, positive integers up to infinity, such as n = 1,2,3….so on.
- Letter notations are also used to represent different shells. For example, n =1 (K-shell), n =2 (L-shell), n =3 (M-shall), n = 4 (N-shell), etc.

**8. When azimuthal quantum number has a value 3, then there are seven values of magnetic quantum number. Give reasons.**

**Ans:** Magnetic quantum number values are related to azimuthal quantum number as $m = (2l + 1)$

when $l = 3$, then, $m = (2 \times 3 + 1) = 7$ i.e. $m = 0, \pm 1, \pm 2, \pm 3$

This shows that f-subshell has 7 different ways of orientation in space because it has 7 values of magnetic quantum number.

(i) $f_{x^3}$ (ii) $f_{y^3}$ (iii) $f_{z^3}$ (iv) $f_{xyz}$ (v) $f_{x^2-y^2}$ (vi) $f_{y^2-z^2}$ (vii) $f_{x^2-z^2}$

**9. Define spin quantum number.**

**Ans:** "The quantum number which describes the angular momentum or spin of an electron in an orbital is called spin quantum number."

There are following two possible spinning motions of an electron;

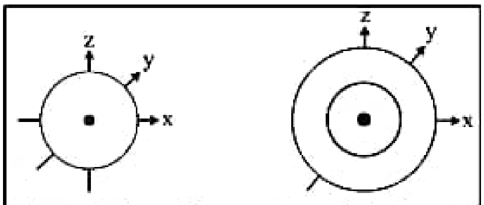
**(i) Clockwise spin:** It is represented by an arrow pointing upwards (↑). Its value is +1/2.

**(ii) Anti-clockwise spin:** It is represented by an arrow pointing downwards (↓). Its value is –1/2.

**10. Draw the shapes of 's' and 'p' orbitals.**

**Ans: (i) s-orbital**

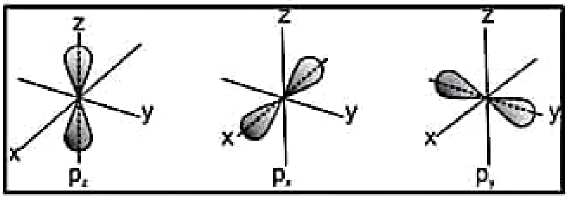
s-orbital is spherical in shape. The electronic density around the nucleus in an s orbital is uniformly distributed in all directions. It is no-directional and starts from first shell (n = 1).




(a) n=1, $l$ = 0, m = 0 (b) n = 2, $l$ = 0, m = 0

**(ii) p-orbitals**

p-orbitals have dumb-bell shape. The electron density is not distributed in a spherically symmetric fashion as in an s orbital. These are directional. A p-orbital has two lobes on any of the axis.




$l$ = 1, m = 0 $l$ = 1, m = –1 $l$ = 1, m = +1

**11. State Auf-bau principle. Write electronic configuration of Sodium ($_{11}$Na) following this principle.**

**Ans: Auf-bau Principle**

This principle says;

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

The electrons are first placed in 1s, 2s, 2p and so on.

**Electronic configuration of Sodium**

$$_{11}Na = 1s^2, 2s^2, 2p^6, 3s^1$$

**12. State Pauli's exclusion principle and Hund's rule.**

**Ans: Pauli's Exclusion Principle**

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

**Hund's Rule:**

"If degenerate orbitals are available and more than one electrons are to be placed in them, they should be placed in separate orbitals with the same spin rather than putting them in the same orbital with opposite spins."

$_6C = \overset{\uparrow\downarrow}{1s}, \overset{\uparrow\downarrow}{2s}, \overset{\uparrow}{2p_x}, \overset{\uparrow}{2p_y}, \overset{\circ}{2p_z}$ , $_7N = \overset{\uparrow\downarrow}{1s}, \overset{\uparrow\downarrow}{2s}, \overset{\uparrow}{2p_x}, \overset{\uparrow}{2p_y}, \overset{\uparrow}{2p_z}$

**13. Distribute electrons in orbitals of $_{29}$Cu and $_{20}$Ca.**

**Ans:** $_{29}Cu = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^2_{xy}, 3d^2_{yz}, 3d^2_{zx}, 3d^2_{x^2-y^2}, 3d^2_{z^2}$

$_{20}Ca = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$

**14. Write Electronic Configuration of Na = 11 and Cr = 24.**

**Ans:** $_{11}Na = 1s^2, 2s^2, 2p^6, 3s^1$

$_{24}Cr = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^1_{xy}, 3d^1_{yz}, 3d^1_{zx}, 3d^1_{x^2-y^2}, 3d^1_{z^2}$

**15. Distribute electrons in orbitals of $_{57}La$, $_{29}Cu$, $_{79}Au$, $_{24}Cr$, $_{53}I$, $_{86}Rn$.**

**Ans:** **Distribution of electron in orbitals**

$_{57}La = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^0\ 5d^1$

$_{29}Cu = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \overset{\uparrow}{4s}\ \overset{\uparrow\downarrow}{3d_{xy}}\ \overset{\uparrow\downarrow}{3d_{yz}}\ \overset{\uparrow\downarrow}{3d_{xz}}\ \overset{\uparrow\downarrow}{3d_{x^2-y^2}}\ \overset{\uparrow\downarrow}{3d_{z^2}}$

$_{79}Au = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ \overset{\uparrow}{6s}\ 4f^{14}\ \overset{\uparrow\downarrow}{5d_{xy}}\ \overset{\uparrow\downarrow}{5d_{yz}}\ \overset{\uparrow\downarrow}{5d_{xz}}\ \overset{\uparrow\downarrow}{5d_{x^2-y^2}}\ \overset{\uparrow\downarrow}{5d_{z^2}}$

$_{24}Cr = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \overset{\uparrow}{4s}\ \overset{\uparrow}{3d_{xy}}\ \overset{\uparrow}{3d_{yz}}\ \overset{\uparrow}{3d_{xz}}\ \overset{\uparrow}{3d_{x^2-y^2}}\ \overset{\uparrow}{3d_{z^2}}$

$_{53}I = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ \overset{\uparrow\downarrow}{5p_x}\ \overset{\uparrow\downarrow}{5p_y}\ \overset{\uparrow}{5p_z}$

$_{86}Rn = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ \overset{\uparrow\downarrow}{6p_x}\ \overset{\uparrow\downarrow}{6p_y}\ \overset{\uparrow\downarrow}{6p_z}$

**CHAPTER END**