ATOMIC NUMBER, PROTON NUMBER AND NUCLEON NUMBER IDENTITY OF AN ELEMENT

Moseley Law

It states

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

Mathematically

$$\sqrt{frequency(\upsilon)}\ \alpha\ Z$$

> **Short Question**
> **What is Moseley law? Give its equation.**

<u>*Moseley experiment*</u>

- In 1913, Moseley, performed X-rays experiments with different elements.
- When different elements were bombarded with cathode rays, the X-rays of some characteristic frequencies were produced.
- He concluded that the atomic number Z was a fundamental property of an element. It 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.

- Atomic number (Z) and mass number (A) are related as $A = Z + N$
- 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 ${}^{A}_{Z}X$, e.g. ${}^{27}_{13}X$ ${}^{27}_{13}Cl$

<u>Example</u>

Consider ${}^{27}_{13}Al$. Thus, Atomic number/ proton number (Z) = 13

Mass number/ nucleon number (A) = 27

Hence, Neutron number = N = 27 – 13 = 14

> **Short Question**
> **Calculate the number of protons, neutrons and electrons in ${}^{27}_{13}Al$.**

The number of electrons, protons, and neutrons can be calculate for an ion.

- ${}^{27}_{13}Al$ atom loses three electrons to form Al^{+3}, then;

  No. of protons = 13
  No. of neutrons = 14
  No of electrons = 13 – 10 = 3

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

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

- When the electrons are gained by the neutral atoms, ${}_{8}O$ to O^{2-}, ${}_{15}P$ to ${}_{15}P^{3-}$ and ${}_{16}S$ to ${}_{16}S^{2-}$. The number of neutrons, protons and electrons are as follows.

Table: 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

EFFECT OF ELECTRIC FIELD ON FUNDAMENTAL PARTICLES:

- The behaviour of particles in an electric field depends upon their mass and charge.
- If the beams of electrons, protons and neutrons are passed through an electric field at same speed, following behaviour is observed.

1. Neutrons are neutral, so they are not deflected. They travel in a straight path perpendicular to the direction of electric field.
2. Protons are positively charged. They are deflected towards the negative plate.
3. Electrons are negatively charged. They are deflected towards the positive plate. However, electrons are deflected to greater extent than proton. It is because they are 1/1836 times lighter than protons.

Short Question
Why electrons are deflected more than proton in an electric filed?

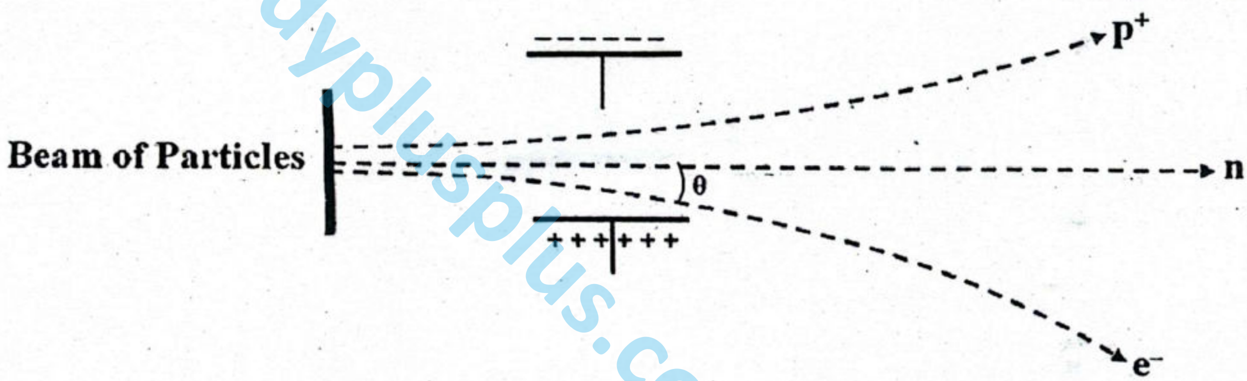


Figure: Behaviour of proton, electron and neutron in the electric field

The amount of deflection from original direction is measured in two ways.

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

After deflection, the particle moves in a circular path.
Hence, the factors affecting the radius of deflection are inversely proportional to the angle of deflection. So, from eq. (i) and (ii)

$$\text{Angle of deflection} \propto \frac{1}{\text{Radius of deflection}}$$

Thus, more the angle of deflection, less will be the radius of deflection

Short Question
How deflection of particles in an electric field is measured?

Table: 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}

ADDITIONAL MCQs

(Answers on Page 50)

1. The property atomic number of an element was discovered by
   (A) Moseley (B) Bohr (C) Schrodinger (D) Medeleeve
2. The mass of an electron is how many times less than that of a proton
   (A) 1836 (B) 7000 (C) 500 (D) 100
3. The atomic number and mass number of S are 16 and 32 respectively. The total number of fundamental particles in S^{2-} is
   (A) 32 (B) 16 (C) 48 (D) 50
4. The number of neutrons in $^{23}_{11}Na$ is
   (A) 11 (B) 22 (C) 23 (D) 12

Quick Check 2.1

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

$^{39}_{19}K$

Thus, Atomic number/ proton number = Z = 19
Mass number/ nucleon number = A = 39
Hence, Neutron number = N = 39 – 19 = 20

$^{35}_{17}Cl$

Thus, Atomic number/ proton number = Z = 17
Mass number/ nucleon number = A = 35
Hence, Neutron number = N = 35 – 17 = 18

$^{40}_{18}Cl$

Thus, Atomic number/ proton number = Z = 18
Mass number/ nucleon number = A = 40
Hence, Neutron number = N = 40 – 18 = 22

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

In a magnetic field, angle of deflection $\propto \frac{charge}{mass}$. Since, electron has the least mass of all, therefore, its angle of deflection will be highest. Thus, electron is deflected the most in magnetic field.

EXPERIMENTAL EVIDENCES FOR THE ELECTRONIC CONFIGURATION

- The modern theory of electronic structure originates from the Bohr Model of atom.
- Evidences for Bohr's model and other theories are derived from
  (i) atomic spectra (ii) ionization energies.

(i) ATOMIC SPECTRA (OR LINE SPECTRA)

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

Atomic spectrum is of two types

(a) Atomic emission spectrum (b) Atomic absorption spectrum

(a) Atomic Emission Spectrum

The spectrum formed by emission of radiations by an atom is called atomic emission spectrum

Formation of Atomic Emission Spectrum

- 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.

> **Short Question**
> **What is the origin of atomic emission spectrum?**

(b) Atomic Absorption Spectrum

The spectrum formed by *absorption of radiations by an atom is called atomic absorption spectrum*

Formation of 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 by atoms.
- The wavelength of the radiation which are absorbed show up as dark lines on the spectrum. The spectrum of this radiation is called an atomic absorption spectrum.

> **Short Question**
> **What is the origin of atomic absorption spectrum?**

Example

The atomic emission and absorption spectrums are shown in **Fig.**

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

(a)

(b)

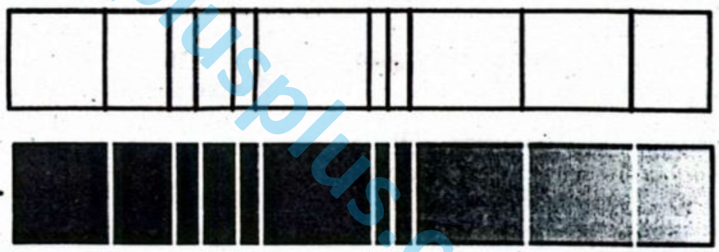
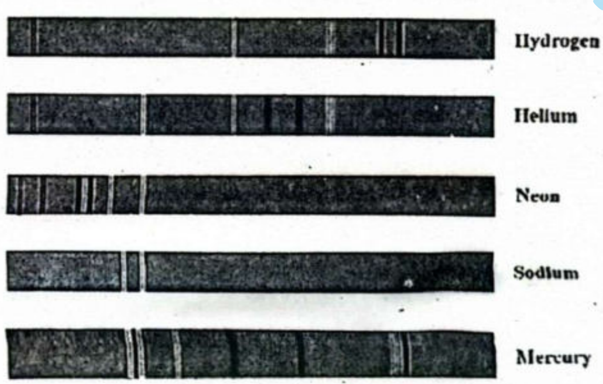


Fig. (a) atomic absorption spectrum (b) atomic emission spectrum

Why atomic spectra is called the finger prints of an element?

- Radiations is emitted or absorbed when electrons jump from one energy level to another in an atom.
- Each element has a unique arrangement of electrons. So, it has a unique range of fixed energy levels.
- Thus, every element absorbs or emits unique wavelengths of the radiation.
- Hence, every element has its own characteristic spectrum. Therefore, every element is identified by its characteristic spectrum. Thus, atomic spectra are the finger prints of the elements.

  Examples: The fig shows the emission spectra of some elements.



ADDITIONAL MCQs

(Answers on Page 50)

5. Atomic spectrum is of how many types?
(A) two (B) three (C) four (D) seven

6. Atoms form which type of spectrum?
(A) continuous spectrum (B) line absorption spectrum
(C) line emission spectrum (D) Both A and B

7. Which of the following is considered as the fingerprint of the elements?
(A) atomic spectra (B) atomic mass (C) number of valence electrons (D) colour

8. The evidences for Bohr's theory and other atomic theories were obtained from
(A) ionization energy (B) atomic spectra (C) Both A and B (D) none of these

(ii) IONIZATION ENERGY

Relation between ionization energy and energy levels (electronic shells)

- Different electrons in different shells have different energy.
- *The electronic configuration* of the atoms can be investigated by measuring the energies of the electrons experimentally. This can be done by measuring ionization energies.
- Ionization energies are used to investigate electronic configurations in two ways.
  (i) Successive ionization energies of the same atom **(ii)** First ionization energies for different atoms

Exercise Q5. What do you mean by successive ionization energies? How the electronic shell structure of magnesium (Mg) is derived from the successive ionization energies?

(i) Successive Ionization Energies of the Same Atom

The energy required to remove each electrons, one by one, from an atom of a an element is measured.

$$X \rightarrow X^{+} + e^{-} \quad \Delta H_{i_1}$$
$$X^{+} \rightarrow X^{2+} + e^{-} \quad \Delta H_{i_2}$$
$$X^{2+} \rightarrow X^{3+} + e^{-} \quad \Delta H_{i_3}$$

- Electrons are removed continuously from an atom of a particular element 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.

Example: Consider magnesium atom.

- The energy required to remove first, second, third and so on electrons is measured.
- A graph is obtained for successive removal of electrons. In this graph the ionization energies values are plotted against number of electrons as shown in Fig.
- This plot shows that successive ionization energies increase from the valence shell to the inner shells.
- First two electrons are removed from the outermost shell and require less energy for removal.
- A large increase occurs when the third electron is removed from Mg. This is because when two electrons of the outer shell have been removed. The 3rd electrons has to be removed from the inner shell that is very much closer to the nucleus.
- The next seven *electrons* are removed successively from the second shell. A gradual *increase* in ionization energy is observed.
- A similar but much more enormous jump occurs when the eleventh and twelfth electrons are removed. It is because, the 11th and 12th electrons are removed from the first shell. This shell is the innermost shell just next to the nucleus.

> **Short Question**
> **Why there is a large increase in ionization energy of 3rd electron in Mg.**

- Over all, two large jumps are observed in the successive ionization energies. These two large jumps give an evidence that the electron in the magnesium atoms exist in three different shells.

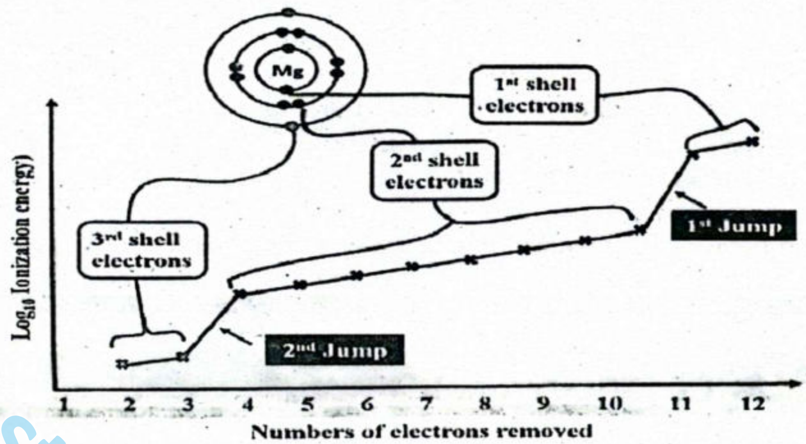


Fig: A plot of the successive ionization energies of Mg

First Ionization Energies of Different Atoms

The details of electronic configuration can be obtained by measuring the first ionization energies of different atoms. The figure shows a plot for the first 88 elements. This graph shows:

- *All ionization energies are strongly endothermic.*
  Thus energy is absorbed to remove an electron from an atom.

- *In a group of periodic table, ionization energies decreases.*
  e.g. from helium to neon to argon, or from lithium to sodium to potassium.
  It is because, *down* the group, number of shells and shielding effect increases. So, hold of nucleus on the valence electrons decreases. Thus, removal of electrons becomes easier. Hence, larger the atom, lower is the ionization energy.

> **Short Question**
> **Why ionization energy decreases in a group in periodic table?**

- *The ionization energies generally increases across a period.*
  e.g. the group 1 elements (alkali metals) have the lowest ionization energy within each period.
  the noble gases have the highest ionization energies in each period.
  It is because across the period shell number remains same but the proton number increases. Electrons are added in the same shell. Therefore, nucleus attracts the valence electrons more strongly. Hence, ionization energy increases across the period.

> **Short Question**
> **Why ionization energy increases across period in periodic table?**

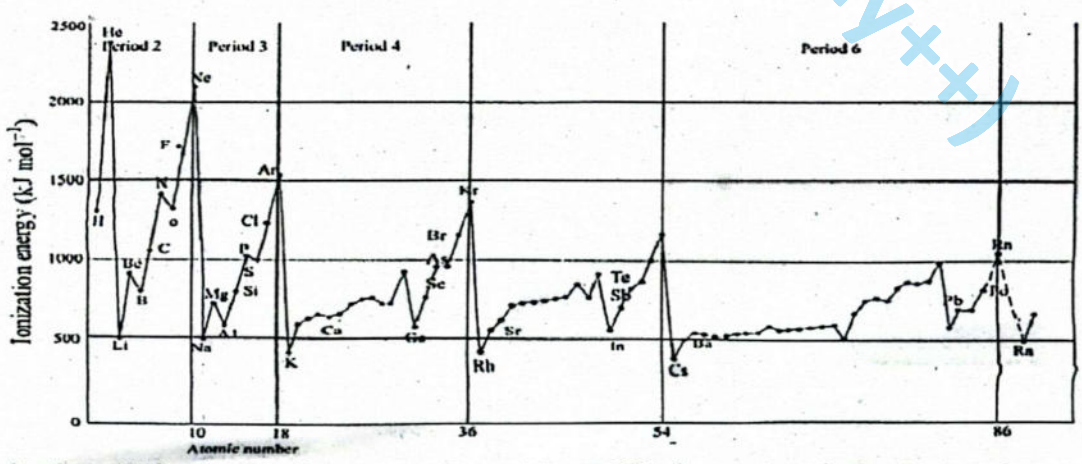


Fig: A plot of the ionization energies of first 88 elements against atomic number

ADDITIONAL MCQs

(Answers on Page 50)

9. Minimum amount of energy required to remove an electron from its gaseous atom is called:
(A) ionization energy (B) electron affinity (C) oxidation (D) reduction

10. The graph between the $\log_{10}$ of the successive ionization energies against the number of electrons removed for Mg shows two large jumps. This shows that Mg has ---------- shells?
(A) two (B) three (C) four (D) one

11. Which element has highest ionization potential
(A) Li (B) Be (C) B (D) C

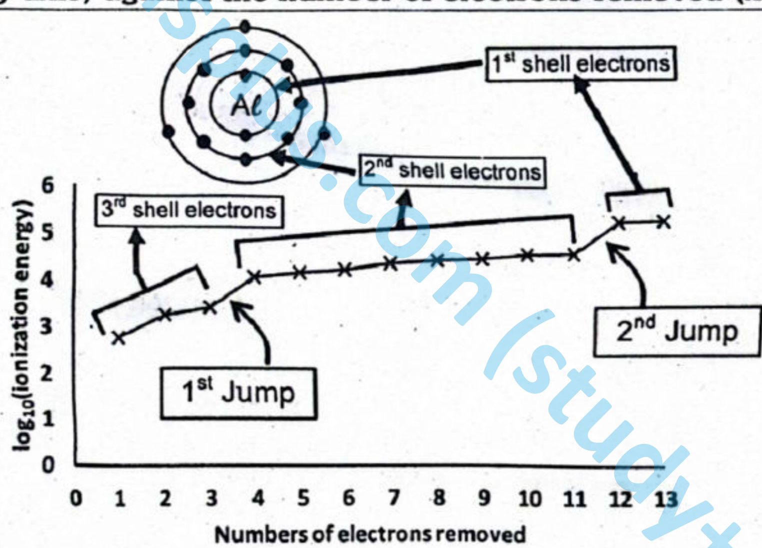
12. The elements which have highest ionization energy in each period are?
(A) halogens (B) noble gases (C) alkali metals (D) chalcogens

Quick Check 2.2

a) Write equations that describe:

i. the 1st ionisation energy of calcium: $Ca_{(g)} \rightarrow Ca^{+}_{(g)} + e^-$ ΔHi_1
ii. the 3rd ionisation energy of potassium: $K^{2+}_{(g)} \rightarrow K^{3+}_{(g)} + e^-$ ΔHi_3
iii. the 2nd ionisation energy of lithium: $Li^{+}_{(g)} \rightarrow Li^{2+}_{(g)} + e^-$ ΔHi_2
iv. the 5th ionisation energy of sulfur: $S^{4+}_{(g)} \rightarrow S^{5+}_{(g)} + e^-$ ΔHi_4

b) For the element aluminium (Z = 13), draw a sketch graph between the $\log_{10}$ of the successive ionisation energies (y-axis) against the number of electrons removed (x-axis).



c) The first ΔHi_1 and the second ΔHi_2 ionisation energies (kJ/mol) of a few elements are given in table

Element	ΔH_{i1}	ΔH_{i2}
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.

Element II. It is because, the element with least first ionization energy (ΔH_{i1}) will be most reactive.

ii. a reactive non-metal,

Element II and III are metals because they have low ionization energies

Element I and IV are non-metals because they have high ionization energies. Among these, element IV is a reactive non-metal since it has relatively lower ionization energy than that of element I

iii. a noble gas

Element I. The element with highest first ionization energy (ΔH_{i1}) will be noble gas.

iv. a metal that forms a stable binary halide of the formula AX_2 (X = halogen).

Element I has one valence electron (monovalent) since ΔH_{i2} is exceptionally high.

Element II has two valence electrons (divalent) since ΔH_{i2} is not exceptionally high. Thus, element II is a metal that forms a stable halide of the formula AX_2

QUANTUM NUMBERS

The Bohr's Model

- The Bohr model was a one-dimensional model.
- It 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. This was described by the n-quantum number.

Exercise Q3. What are quantum numbers? Describe briefly principal and spin quantum numbers.

The Schrodinger's Model

- In Schrodinger's model, the electron occupy three-dimensional space. Therefore, it requires three coordinates, or three quantum numbers, to describe the orbitals of electrons.
- The three quantum numbers from Schrödinger's wave equations are:
  (i) Principal (n) quantum number
  (ii) Azimuthal (ℓ) quantum number or Angular quantum number
  (iii) Magnetic (m) quantum number.
- These quantum numbers describe the size, shape, and orientation in space of the orbitals in an atom.

PRINCIPAL QUANTUM NUMBER (n)

- The principal quantum number is denoted by **'n'**
- It can have positive integral values 1,2,3,4... designated by K, L, M, N...,
- It is similar to the quantum number (number of orbit) in Bohr's model of the hydrogen atom.

Significance

- The orbitals with same **'n'** values is called an electron shell. Therefore, the principal quantum number, n, describes the size of the orbital.
- The larger 'n' is, the greater the average distance of an electron in the orbital from the nucleus and therefore the larger the orbital.
- An increase in 'n' also means that the electron has a higher energy. Therefore, it is less tightly bound to the nucleus. For a one electron system (atom or ion), the radius and energy of a shell can be calculated as:

$$r_{(n)} = 0.529\,(n^2)\ A^\circ$$

$$E_n = -2.18 \times 10^{-18} / n^2\ J$$

- The principal Quantum number, n, can also be used to calculate the maximum number of electrons in a shell by the formula $2n^2$.

AZIMUTHAL QUANTUM NUMBER (ℓ)

- It is denoted by 'ℓ'.
- It can have integral values 0 to (n – 1) for each value of n.
- This quantum number describes the shape of the orbital.
- The value of 'ℓ' are integers that depend on the value of the principal quantum number.
  - ✓ If n= 1, there is only one possible value of 'ℓ', i.e, ℓ =0 (n-1, where n = 1).
  - ✓ If n=2, there are two values of 'ℓ' i.e 0 and 1.
  - ✓ If n=3 there are, three values of 'ℓ' i.e 0,1 and 2.
- The value of 'ℓ' is designated by small letters s,p,d, and f.
- The s,p,d, and f stand for **s**harp, **p**rincipal, **d**iffused and **f**undamental, respectively. These are the spectral terms used to describe certain features of spectral lines before quantum mechanics was developed.
- The set of orbitals that have the same **n** and ℓ values is called a **subshell.**
- The number of electrons in a subshell can be calculated by the formula $2(2\ell+1)$.

Thus

$\ell = 0$ s-subshell, $2(2\times 0 + 1) = 2$ electrons
$\ell = 1$ p-subshell, $2(2\times 1 + 1) = 6$ electrons
$\ell = 2$ d-subshell, $2(2\times 2 + 1) = 10$ electrons
$\ell = 3$ f-subshell, $2(2\times 3 + 1) = 14$ electrons

<u>*Summary*</u>

"ℓ"	*Subshell*	*Maximum number of Electrons*
0	s	2
1	p	6
2	d	10
3	f	14

Table: Relationship between n, ℓ and subshells

Shell	Principal Quantum number n	Angular Momentum Quantum Number (Azimuthal)	Subshells
K	1	0	1s
L	2	0	2s
		1	2p
		0	3s
		1	3p
		2	2p
N	4	0	4s
		1	4p
		2	4d
		3	4f

MAGNETIC QUANTUM NUMBER (m)

- The magnetic quantum number is denoted by '**m**'
- It describes the orientation of an orbital in space.
- Within a subshell, the value of '**m**' depends on the value of ℓ.
- For a certain value of ℓ there are $(2\ell + 1)$ integral values of m as follows:

  $-\ell$............0............$+\ell$
- If $\ell = 0$, there is only one possible value of m, i.e. 0.
- If $\ell = 1$, then there are three values of m; –1,0 and –1.
- If $\ell = 2$, there are five values of m, namely, – 2, – 1, 0, + 1 and + 2.
- The values of m indicate the number of orbitals in a subshell.
- Orbitals of the same subshell have same energy and are called degenerate orbitals.
- These degenerate orbitals are differentiated from each other in the presences of magnetic field. Hence the name of this quantum number, i.e. magnetic quantum number.
- The relationship between the angular momentum and magnetic quantum numbers is provided in Table.

Did you know?

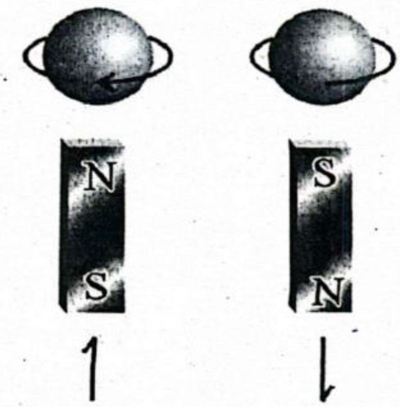
Splitting of small fine lines in the presence of magnetic field and there three dimensional orientation in space indicate the presence of orbitals in subshells.

Table: Relationship between n, ℓ and subshells

Subshell	Azimuthal / angular Quantum number (ℓ)	Magnetic Quantum number (m) ($-\ell$ — 0 — $+\ell$)	Number of degenerate orbitals $(2\ell + 1)$
s	0	0	One orbital
p	1	-1 0 +1	Three degenerate p-orbitals
d	2	+2 +1 0 -1 -2	Five degenerate d-orbitals
f	3	+3 +2 +1 0 -1 -2 -3	Seven degenerate f-orbitals

SPIN QUANTUM NUMBER (s)

- Spin Quantum Number is denoted by **'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.
- The Figure shows the two possible spinning motions of an electron. One is clockwise and the other is counterclockwise.
- To take the electron spin into account, it is necessary to introduce a fourth quantum number called the electron spin quantum number (s).
- Its values are $+\frac{1}{2}$ and $-\frac{1}{2}$ as in Fig.
- The clockwise spin is represented by an arrow (↿) pointing upwards, while the anti-clockwise spin is represented by an arrow (⇂) pointing downwards.



(a) (b)

- Each *orbital* can accommodate at the most two electrons provided the two electrons have opposite spins.

Extend your knowledge

In the n principal quantum number, **there are n subshells, consisting of n^2 orbitals, with a maximum number of $2n^2$ electrons.**

QUICK CHECK 2.3

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

i. Principal quantum number: It describes the shells, their size, energy and number of electrons in them.

ii. Azimuthal quantum number: It describes the sub-shells, their size, energy, shapes and number of electrons in them.

iii. Magnetic quantum number: It describes the directions (orientations) of orbitals in space.

iv. Spin quantum number: It describes the spinning of electron (clockwise or anti-clockwise) about its axis.

b) For an electron(s):

i. If n=2 and ℓ = 1, how many orientations in space are possible?
When n = 2, it is the 2^{nd} shell. When $\ell = 1$, it is the p-subshell.
This is 2p subshell.
So, for $\ell = 1$, m = −1, 0, +1. Hence, it has three orientations. $2p_x$, $2p_y$ and $2p_z$

ii. If n=3 and ℓ= 2, which shell and subshell does the electron belong to?
When n = 3, *it is the* 3^{rd} shell. When $\ell = 2$, it is the d-subshell. Thus, this is 3d subshell.

iii. If ℓ=2, find all possible values of m and maximum number of electrons for m.
For $\ell = 2$, m = −2, −1, 0, +1, +2. Hence, it has five orientations. d_{xy}, d_{xz}, and d_{yz}, $d_{x^2-y^2}$, d_{z^2}

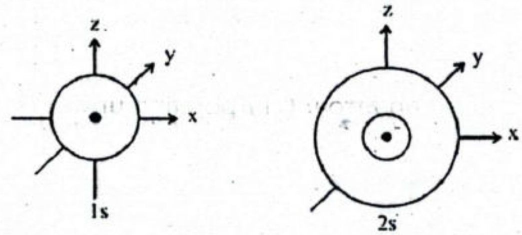
SHAPES OF ATOMIC ORBITALS

Exercise Q4. Draw the shapes of s, p and d-orbitals. Justify these by keeping in view the azimuthal and magnetic quantum numbers.

An atomic orbital is defined as the region is space around the nucleus in which the probability of finding the electron is maximum (90 to 95%).

s-orbitals

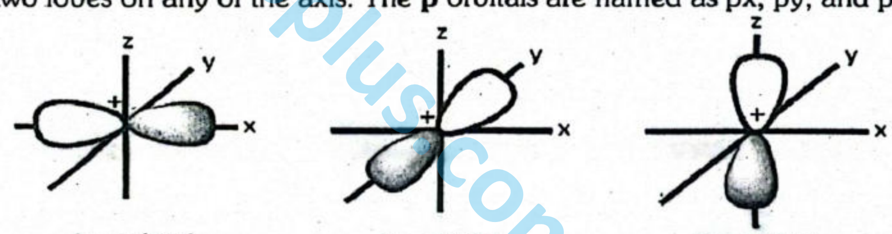
- For s-subshell, ℓ=0, m=0. Hence, it has only one orbital.
- The shape of an **'s'** orbital is spherical.
- The electronic density around the nucleus in an s orbital is uniformly distributed in all directions.
- With the increase in the principal quantum number, the size of **s** orbital also becomes larger.



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

p-orbitals:

- For p-subshell, ℓ=1, m= – 1, 0, +1. Hence, it has three orbitals.
- The distribution of electron density for a 2p orbital is shown in Fig.
- The electron density is not distributed in a spherically symmetric fashion as in an s orbital.
- A **p** orbital has two lobes on any of the axis. The **p** orbitals are named as px, py, and pz accordingly.

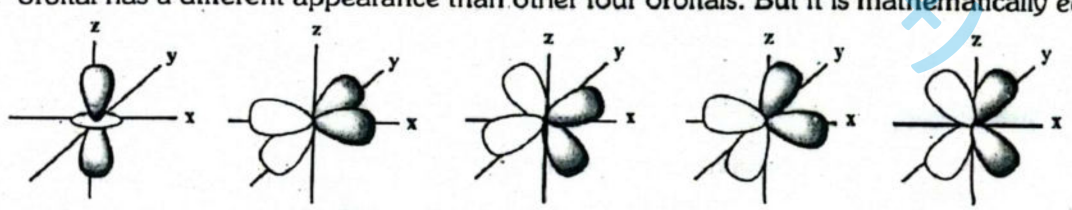


$2p_x$ orbital $2p_y$ orbital $2p_z$ orbital

n=2, ℓ=1, m=+1 **n=2, ℓ=1, m=0** **n=2, ℓ=1, m= –1**

d-orbitals:

- For d-subshell, ℓ=2, m= – 2, – 1, 0, +1, +2. Hence, it has five orbitals.
- The 'd' orbitals in a given shell 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 are oriented between the axes.
  - ✓ The lobes of the $d_{x^2-y^2}$ lie along the x and y axes.
  - ✓ The dz^2 orbital has two lobes along the z-axis and a "doughnut" in the xy plane.
  - ✓ The dz^2 orbital has a different appearance than other four orbitals. But it is mathematically equivalent.



d_{z^2} $d_{x^2-y^2}$ d_{xy} d_{yz} d_{xz}

f-orbitals:

- For f-subshell, ℓ=3, m= – 3, – 2, – 1, 0, +1, +2, +3. Hence, it has seven orbitals.
- 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?

- In wave-mechanical model of an atom, an atomic orbital is the region is space around the nucleus in which the probability of finding the electron is maximum (90 to 95%).
- It is actually a mathematical description of the electron's distribution is space based on its wave-like nature.

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

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 a certain value of ℓ the number of possible orientations are given as

$$-\ell \ldots\ldots\ldots 0 \ldots\ldots\ldots +\ell$$

For p-subshell, $\ell=1$, then there are three values of m; –1, 0 and +1. Thus there are three orientation of p-orbitals. The p orbitals are named as px, py, and pz. These are present perpendicular to each other.

ADDITIONAL MCQs

(Answers on Page 50)

13. Which of the following quantum number is not directly derived from Schrodinger's theory?
(A) principle (B) azimuthal (C) magnetic (D) spin

14. Quantum number values for 4f orbital are
(A) n=4 and $\ell = 2$ (B) n=3 and $\ell = 4$ (C) n=4 and $\ell = 3$ (D) n=4 and $\ell = 4$

15. If n=2 and $\ell = 2$, how many orientations in space are possible?
(A) 1 (B) 3 (C) 5 (D) 7

16. The number of electrons in a sub-shell can be obtained by
(A) $2n^2$ (B) $(2\ell+1)$ (C) $2(2n+1)$ (D) $2(2\ell+1)$

17. An orbital which is spherical and symmetrical is:
(A) s-orbital (B) p-orbital (C) d-orbital (D) f-orbital

ELECTRON CONFIGURATION

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

Each group of orbitals in a subshell is labeled by its subshell notation. An electron in an orbital is shown by an arrow. The arrow points upward, when s = +½ and downward when s = +½. The orbital diagram of boron $^{11}_{5}B$ is as follows

$$^{11}_{5}B = 1s^{\uparrow\downarrow}\ 2s^{\uparrow\downarrow}\ 2p_x^{1}\ 2p_y^{0}\ 2p_z^{0}$$

Distribution of Electrons in Shells

- The electronic configuration of an atom describes the distribution of electrons in its atomic shells.
- The shells are identified by the principal quantum number **(n)**.
- The **'n'** values are 1,2,3,4... denoted as K, L, M, N shells and so on.
- Each shell has specific capacity of electrons determined by the formula $2n^2$.

Thus

n	shell	No. of maximum electrons
1	K shell	$2n^2 = 2\times1^2 = \mathbf{2}$
2	L shell	$2n^2 = 2\times2^2 = \mathbf{8}$
3	M shell	$2n^2 = 2\times3^2 = \mathbf{18}$
4	N shell	$2n^2 = 2\times4^2 = \mathbf{32}$

Example: Distribution of electrons for Na and Cl are

	K	L	M
$_{11}$Na	2	8	1
$_{17}$Cl	2	8	7

Distribution of Electrons in Subshells

In case of subshells, the electronic configuration is described by a notation. This lists the subshell symbols, one after the other. Each symbol has a superscript on the right. This gives the number of electrons in the subshell.

Example: Two electrons in 1s orbital are written as $1s^2$

Aufbau Principle

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

- Aufbau principle is also known as the building up principle.
- According to this principle, 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 shells with electrons may be obtained from the summation ($n + \ell$) values.

($n+\ell$) Rule

- Lower the ($n+\ell$) value lower will be the energy of subshell and is filled first.
- If two subshells have same ($n+\ell$) value then lower the value of '**n**' lower will be the energy of subshell and is filled first.

Short Question

Why 3d orbital has greater energy than 4s orbital?

Examples

Subshell	($n+\ell$) value	n value	Remarks
4s	4+0=4	4	Thus, 4s orbital has
3d	3+2=5	3	lower energy than 3d.
4p	4+1=5	4	Thus, 3d orbitals has
3d	3+2=5	3	lower energy than 4p.

Table: ($n + \ell$) values of various sub-shells

Principal quantum no. (n)	Azimuthal quantum no. (ℓ)	Subshell	($n+\ell$) Value
1	0	1s	(1+0)=1
2	0	2s	(2+0)=2
	1	2p	(2+1)=3
3	0	3s	(2+0)=3
	1	3p	(2+1)=4
	2	9d	(2+2)=5
4	0	4s	(4+0)=4
	1	4p	(4+1)=5
	2	4d	(4+2)=6
		4d	(4+3)=7
5	0	5s	(5+0)=4
	1	5p	(5+1)=6
	2	5d	(5+2)=7
	3	5f	(5+3)=8
6	0	6s	(6+0)=6
	1	6p	(6+1)=7

Fig: The diagram shows the energy of each sublevel. Each box on the diagram represents an atomic orbital.

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

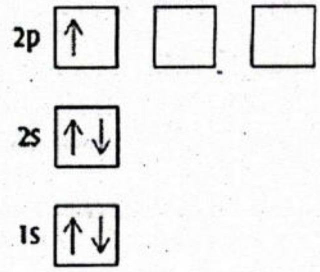
Thus, the configuration of $_4Be$ in subshells is: $1s^2\ 2s^2$

the electronic configuration of $_{11}Na$ in subshells is: $1s^2\ 2s^2\ 2p^6\ 3s^1$

Distribution of Electrons in Orbitals

The filling of orbitals is shown by a box diagram.

- Each box represents an atomic orbital.
- Each orbital occupies a maximum of two electrons
- An electron is represented by an arrow.
- The direction of the arrow represents the 'spin' of the electron.
- The boxes (orbitals) can be arranged in order of increasing energy from bottom to top.



The following rules are obeyed while filling orbitals with electrons.

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

> **Short Question**
> **Define Pauli's exclusion principle. Give an example?**

Example

In the first shell of helium (**He**) atom, there are two electrons. They are present in 1s orbital. According to the concept of quantum numbers and Pauli's rule, the values of their quantum numbers are:

Electron	n	ℓ	m	s
Electron 1	**1**	**0**	**0**	**+½ (clockwise)**
Electron 2	**1**	**0**	**0**	**-½ (anticlockwise)**

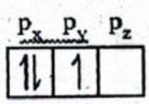
The two electrons having the same values of **'n'**, **'l'** and **'m'** can have different values of **'s'**. Thus, they have opposite spins.

Hund's Rule

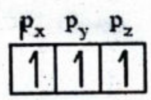
When degenerate orbitals are available and more than two electrons are to be placed in them, they should be placed in separate orbitals with the same spin rather than in the same orbital with opposite spins.

Example

This rule tells about filling electrons into the orbitals having equal energies. The three p-orbtials, *i.e.*, px, py and pz have equal energy. The three electrons in p-orbitals can be filled in two different ways.



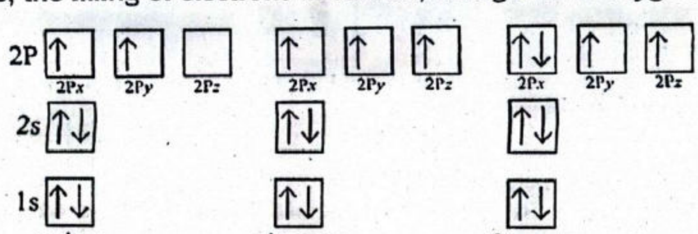
I (wrong)



II (right)

> **Short Question**
> **Define Hund's rule. Give an example?**

Thus, according to this rule, the filling of electrons in carbon, nitrogen and oxygen is:



carbon $1s^2\,2s^2\,2p^2$ nitrogen $1s^2\,2s^2\,2p^3$ Oxygen $1s^2\,2s^2\,2p^4$

Table: Electronic configuration of ground states of elements Z = 1-36.

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

ADDITIONAL MCQs

(Answers on Page 50)

18. Maximum number of electrons in an orbital is:
(A) 6 (B) 10 (C) 14 (D) 2

19. $n+\ell$ value of 6d orbital is
(A) 8 (B) 9 (C) 10 (D) 11

20. The electron in a subshell is filled according to formula:
(A) $2n^2$ (B) $2(2\ell+1)$ (C) $(2\ell+1)$ (D) None of these

21. Maximum number electrons in f-subshell is:
(A) 2 (B) 6 (C) 10 (D) 14

22. The element which has maximum number of unpaired electrons is
(A) Cr_{24} (B) Ca_{20} (C) Fe_{26} (D) Cu_{29}

23. The electronic configuration of an atom is $1s^2 2s^2 2p^4$. The number of unpaired electrons in this atom is:
(A) 0 (B) 2 (C) 4 (D) 6

24. After filling of 4*f*, the entering electron goes into:
(A) 5d (B) 6p (C) 6s (D) 4d

25. When 4s orbital is complete, the electron goes into
(A) 4p orbital (B) 3d (C) 4d (D) 4f

26. When 5d orbital is completed then entering electron goes into:
(A) 6s (B) 6p (C) 6d (D) 6f

ELECTRON CONFIGURATION AND THE PERIODIC TABLE

- The electron configuration of an element can be written from its location in the periodic table.
- The elements can be grouped in terms of the type of orbital into which the electrons are placed.
  - ✓ On the left is a block of two columns. These are alkali metals and alkaline earth metals. In these, **'s'** orbitals are being filled. Group 1 outer configuration is ns^1 and that of group 2 is ns^2. These are s-block elements.
  - ✓ On the right is a block of six columns. These are the elements in which the outermost **'p'** orbitals are being filled. These are p-block elements. e.g. group 13 elements have $ns^2\ np^1$ configuration.
  - ✓ In the middle of the table is a block of ten columns that contain the transition metals. In these, the **'d '** orbitals are being filled. These are d-block elements.
  - ✓ Below the main portion of the table are two rows that contain fourteen columns. In these, the **' f '** orbitals are being filled. These are f-block elements.
- The numbers 2, 6, 10, and 14 are the number of electrons that can fill s, p, d, and f subshells, respectively. Thus, s-block consists of two columns, p-block six columns, d-block ten columns and f-block fourteen columns.

Quick Check 2.5

(a) With the help of periodic table write the electron configurations for the following atoms by giving the appropriate noble-gas inner core plus the electrons beyond it:

(i) $_{48}$Cd $\underline{1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6}\ 5s^2\ 4d^{10}$.
Since, nearest previous noble gas is $_{36}$Kr.
So, the electronic configuration becomes [Kr] $5s^2\ 4d^{10}$

(ii) $_{57}$La $\underline{1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6}\ 6s^2\ 4f^0\ 5d^1$
Since, nearest previous noble gas is $_{54}$Xe.
So, the electronic configuration becomes [Xe] $6s^2\ 4f^0\ 5d^1$

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

$_{51}$Sb $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^3$

(c) How many unpaired electrons are there in each atom of $_{51}$Sb?

Since 5p orbital has 3 electrons ($5p^3$). Thus, all these three electrons are unpaired. $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

d) Write down the electronic configuration of valence shell of the following in terms of orbitals:

i) $_{13}$Al [Ne] $3s^2\ 3p^1$
ii) $_{14}$Si [Ne] $3s^2\ 3p^2$
iii) $_{28}$Ni [Ar] $4s^2\ 3d^8$

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.
- The similarities in valence electrons configurations means similarities in chemical properties of group of elements.

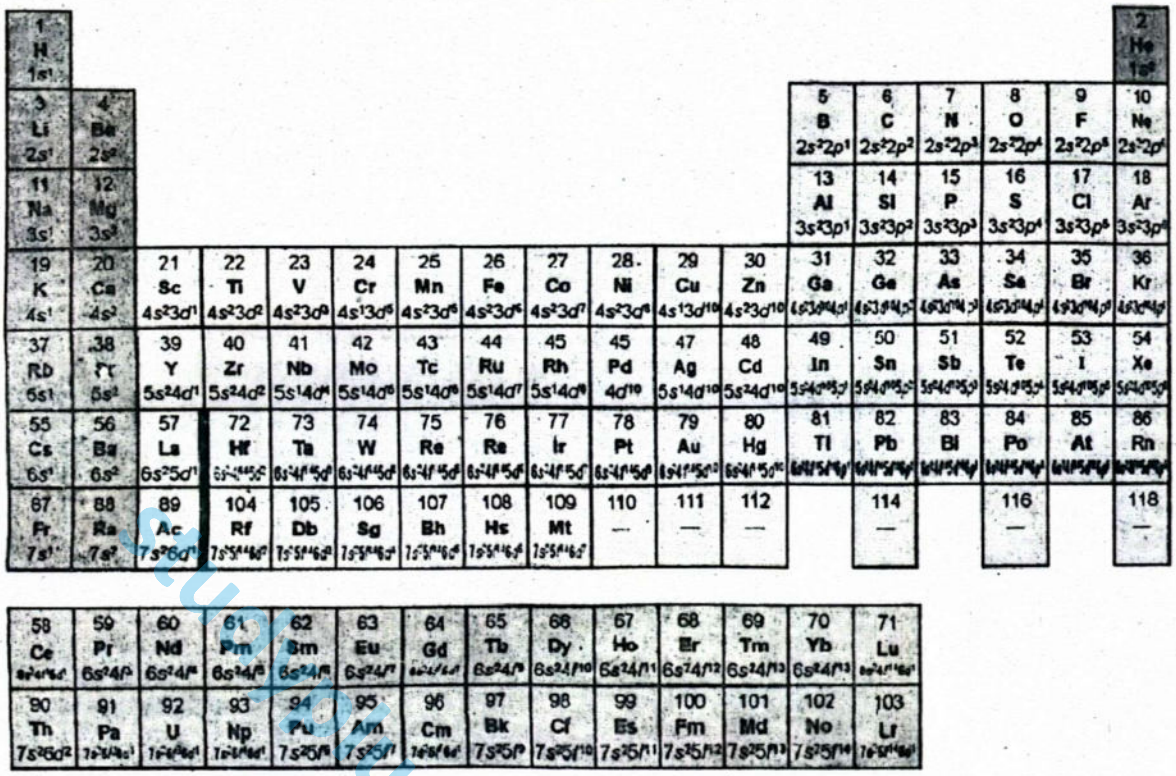


Fig. The electronic distribution of elements of modern periodic table in periods and groups.

Classification of elements of periodic table:

- The main-group or representative elements all have valence-shell configurations $ns^a np^b$. The 'a' and 'b' *vary in different* groups. Thus, the outer 's' or 'p' subshell is being filled.
- In the d-block transition elements, a 'd' subshell is being filled.
- In the f-block transition elements or inner-transition elements, an 'f' subshell is being filled.
- There is a definite pattern to the order of filling of the subshells for the elements in the periodic table. From this the building-up order can be obtianed.
- The overall sketch of periodic table is given in Fig. This shows the blocks of elements.

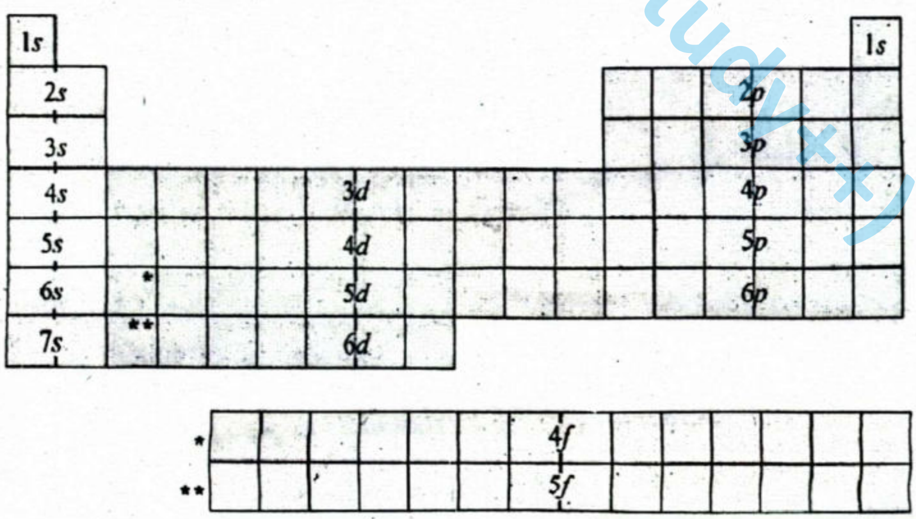


Fig. Modern periodic table showing s, p, d and f-block elements.

- The Fig. shows a periodic table indicating this pattern of filling electrons.
  - ✓ In the first period, '1s' subshell is being filled.
  - ✓ In the second period, '2s' and '2p' subshells are being filled.
  - ✓ In the third period, '3s' and '3p' subshells are being filled.
  - ✓ In the fourth period, '4s', '3d' and then '4p' subshells are being filled.
- The classification of elements of periodic table has also been done into metals, non-metals, representative, transition elements, periods, groups and blocks (s, p, d, f). This division is done on the bases of electronic configuration and their valence electrons.
- The chemical properties of various categories of elements can be easily judged.

Quick Check 2.6

1. 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?***
Since the last electron is in 5p orbital. Thus, it belongs to p-block elements.

ii. ***Which group does it belong to?***
The valence shell electronic configuration is **$5s^2\ 5p^5$. Thus, it has seven electrons in its valence** shell. This shows that it belongs to group 17 of periodic table.

iii. ***Which period does it belong to?***
The valence shell electronic configuration is **$5s^2\ 5p^5$**. Here n = 5, so it belongs to period 5.

iv. ***Identify this element.***
This element has total number of electrons = 53. This is also its proton number or atomic number.
The element with atomic number 53 is iodine (**$_{53}I$**)

2. 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? Name it.

Since d-orbital is in the process of being filled. Thus, this element belongs to d-block elements. The element is Chromium (Cr).

ELECTRONIC CONFIGURATION OF IONS AND FREE RADICALS

IONS

- Positive ions are formed when electrons are removed from atoms.
  The sodium ion, Na^+ (proton number = 11), has 10 electrons. So, its electronic configuration is $1s^2\ 2s^2\ 2p^6$. This is the same as the electronic configuration of neon which is the element with 10 electrons.
- Negative ions are formed when atoms gain electrons.
  The sulfide ion, S^{2-} (proton number= 16), has 18 electrons. Its electronic configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$. This is the same as argon, the element with 18 electrons in each atom.
- In general, electrons in the outer subshell are removed when metals form their positive ions.
- However, the d-block elements behave slightly differently. Reading 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: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3$

FREE RADICALS

A free radical is a species that has one or more unpaired electrons

Free radicals are very reactive.

They readily react with other atoms to give molecules as H_2, Cl_2 and HCl.

Examples:

- A simple free radical is free chlorine atom ($\ddot{Cl}\cdot$). The electron configuration of this radical is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$. 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 $Cl^{\cdot}$.
- Groups of atoms can also be free radicals. e.g., $OH^{\cdot}$, $CH_3^{\cdot}$, etc.

Quick Check 2.7

Write electronic configurations for the following ions and free radicals:

a. Al (Z = 13)
Al atom: $1s^2\ 2s^2\ 2p^3$

b. O^{2-} (Z = 8)
O atom: $1s^2\ 2s^2\ 2p^4$
O^{2-} ion: $1s^2\ 2s^2\ 2p^6$

c. Fe^{3+} (Z = 26)
Fe atom: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$
Fe^{3+} ion: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^0$

d. Cu^{2+} (Z = 29)
Cu atom: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$
Cu^{2+} ion: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9\ 4s^0$

e. Cu (Z = 29)
Cu atom: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

ELECTRONIC CONFIGURATION AND THE FORMATION OF SEMICONDUCTORS

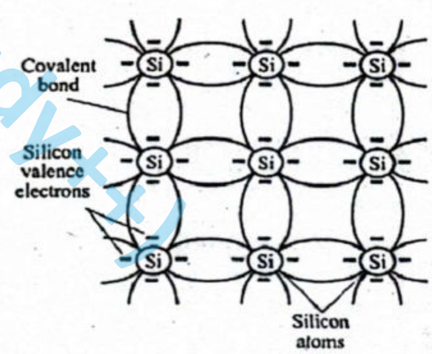
Semiconductors are materials that can conduct electricity under some conditions.

They are used in many electronic devices, including smartphones, laptops, and cars.

Examples:

Silicon, Germanium and Arsenic etc.

- The formation of semiconductors is possible because of a unique electronic configuration of these elements.
- Consider the example of Si. It can be converted into a P-type and N-type semiconductors.
- The electron configuration of $_{14}Si$ = 2, 8, 4. This shows that it has 4 electrons in its valence shell. In the pure crystalline form, each Si atom is bonded to four other Si atoms. In this form, there is no possibility of electronic conduction through the Si crystal.
- P-type and N-type semiconductors are formed by "doping" a pure semiconductor material with impurity atoms.
- If the added impurity creates "holes" (positive charge carriers) in the lattice, it gives P-type semiconductor.



- If the added impurity adds extra electrons (negative charge carriers), it gives N-type semiconductor.

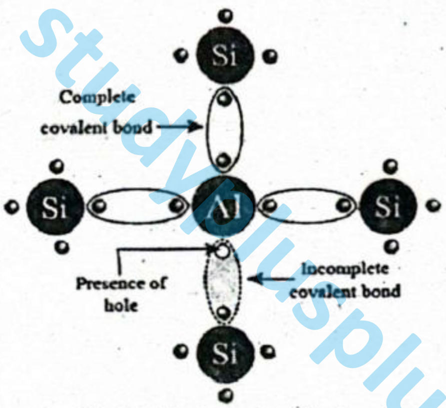
P-type semiconductor formation:

- Adding trivalent impurities (like boron or aluminum) to a pure semiconductor creates a P-type semiconductor.
- Trivalent impurity has three valence electrons.

- Some of the trivalent atoms take place of the Si atoms in the crystals.
- The silicon atoms cannot make four bonds due to the lack of electrons.
- Thus, holes are created in the crystal lattice. These act as positive charge carriers as shown in Fig.
- The result is a P-type semiconductor tat conduct electronicity when connected to an external source.

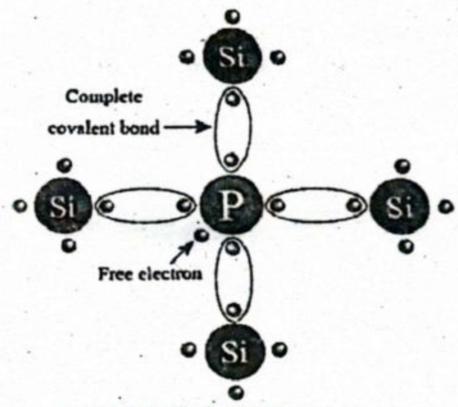
N-type semiconductor formation:

- Adding pentavalent impurities (like phosphorus or arsenic) to the pure semiconductor creates an N-type semiconductor.
- Some of Si atoms are replaced with the pentavalent phosphorus atoms.
- The Si atoms in the vicinity of these atoms can make four bonds. The fifth electron is an extra electron.
- These impurity atoms provide extra electrons to the crystal lattice.
- These free electrons become free to move and act as negative charge carriers as in fig.
- The result is an N-type semiconductor that can conduct electricity when connected to an external source.



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

Formation of P type extrinsic semiconductor



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

Formation of N type extrinsic semiconductor

ADDITIONAL MCQs

(Answers on Page 50)

27. An N-type semiconductor is made by mixing Si-crystal with which impurity?
(A) trivalent (B) tetravalent (C) pentavalent (D) divalent

28. Which of the following is a free radical?
(A) Cl (B) Cl_2 (C) HO^- (D) Mg

29. A certain element has electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$. The block of the element is
(A) s-block (B) p-block (C) d-block (D) f-block

30. The electronic configuration of $_{22}Ti^{2+}$ ion is
(A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ (B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
(C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2$ (D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3$

ANSWERS TO ADDITIONAL MCQS

Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans
1	A	2	A	3	D	4	D	5	A	6	D	7	A	8	C
9	A	10	B	11	D	12	B	13	D	14	C	15	C	16	D
17	A	18	D	19	A	20	B	21	D	22	A	23	B	24	A
25	A	26	B	27	C	28	A	29	C	30	C				

Q1 MULTIPLE CHOICE QUESTIONS

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 orbital is complete, the entering electron goes into:
a) 4f b) 4s c) 4p d) 4d

III. Quantum number values for 2p orbitals are:
a) n = 2, l = 1 b) n = 1, l = 2 c) n = 1, l = 0 d) n = 2, l = 0

IV. An electrons having set of values, n = 4, ℓ = 0, m = 0 and s = +½ , lies in
a) 2s b) 3s c) 4s d) 4p

V. The quantum number values (n, ℓ, m) of 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 energy 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. What information does the principal quantum number (n) gives us about orbitals:
a) Size b) Shape c) Size and shape d) Spin

XI. How many unpaired electrons are present in an atom of cobalt?
a) Two b) Three c) Four d) Five

ANSWERS TO MULTIPLE CHOICE QUESTIONS		
No.	**Ans**	**EXPLANATION**
I.	c	The quantum number 'm' is called magnetic quantum number. It is also called orientation quantum number as it gives the spatial orientation of the orbital.
II.	b	According to (n+ℓ) rule, orbitals are filled energy wise. Thus orbitals are filled in following order 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p… Hence, when 3d orbital is complete, the entering electron will goes into 4s orbital.
III.	a	For 2p orbital, 2 stands for principle quantum number. Hence, for 2p orbital, n=2. According to Azimuthal quantum number, when n=2, l=0,1. Here '0' stands for 's-orbital' and '1' stands for 'p' orbital. Hence, fro 2p orbital, l=1. Thus, for 2p orbital, n=2, l=1
IV.	c	n=4 shows the valence electron is in fourth shell. ℓ=0 & m=0 shows that the electron is in s-orbital. s=+1/2 shows that there is only one valence electron. So, it is $_{19}K$ and the valence electron is present in 4s orbital. The electronic configuration is $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$

V.	**b**	The electronic configuration ^9_4Be is $1s^2$, $2s^2$. Hence, the fourth electron is valence electron in 2s orbital. It has quantum numbers: n=2, ℓ=0 and m=0 **(2 , 0 , 0)**
VI.	**b**	He is noble gas. It has high ionization energy. Ionization energy increases left to right. However, N has higher ionization energy than O due to its half-filled p-orbitals. Mg is a metal. Metal tends to lose electron quickly. So, it has least ionization energy. Thus, order is He > F > N > O > Mg
VII.	**b**	p-orbital has two lobes 
VIII.	**c**	All three p-orbitals are perpendicular to each other on X, Y and Z axis. (p_x , p_y , p_z)
IX.	**c**	For d-subshell, l=2. So, m = –2, –1, 0, +1, +2. Hence, there are five d-orbitals.
X.	**a**	The principle quantum number (n) gives information about shells. Thus, it shows the size of atom.
XI.	**b**	The electronic configuration of $_{27}Co$ is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$. So, d-orbitals has seven electrons. Thus, four electrons are paired and **three are unpaired.** ↑↓ | ↑↓ | ↑ | ↑ | ↑

Q2 SHORT QUESTIONS

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

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 a certain value of ℓ the number of possible orientations are given as
  $-\ell$...........0............$+\ell$
- For p-subshell, ℓ=1, then there are three values of m; –1, 0 and +1. Thus there are three orientation of p-orbitals. The p orbitals are named as p_x, p_y, and p_z. These are present perpendicular to each other.

(b) 'I_3' of Mg is much bigger than its 'I_2'. Justify.

In Mg, there are two electrons in the valence shell.

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_3 is much bigger than I_2

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

Generally, more the atomic size, lesser will be electrostatic attraction between electrons and nucleus so lower will be ionization energy and vice versa.

The order of atomic size of the given elements is K > Ca > Li > S > Kr.

Hence, the order of ionization energy will be K < Ca < Li < S < Kr.

(d) 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.

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

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

According to Aufbau principle: the subshells in an atom are filled with electrons in an increasing order of their increasing energy values. The order of filling shells with electrons is obtained by $(n + \ell)$ values. Lower the $(n+\ell)$ value lower will be the energy of subshell and is filled first

Subshell	$(n+\ell)$ value	*n* value	Remarks
4s	4+0=4	4	Thus, 4s orbital has lower energy than 3d.
3d	3+2=5	3	

Hence, 4s orbitals is filled before than 3d orbital

(e) (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.

Atomic number/ proton number = Z = 17

Mass number/ nucleon number = A = 35

So, No of neutrons = N = A – Z = 35 – 17 = 18

Hence,

No. of protons = 17 No of neutrons = 18 No of electrons = 17

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

The element X gains an electron to form X^- ion. Thus,

Hence,

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

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

The electronic configuration is

$_{80}Hg$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10}$

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

n = 3 means 3rd shell.

The electronic configuration shows that 3rd shell has 18 electrons. i.e., $3s^2\ 3p^6\ 3d^{10}$

ii. How many electrons occupy 4d atomic orbitals?

The electronic configuration shows that it 4d contains 10 electrons. i.e. $4d^{10}$

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

Six electrons are present in 4p orbital ($4p^6$). So, each p-orbital has two electrons. $4p_x^2\ 4p_y^2\ 4p_z^2$

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

The electronic configuration shows that all orbitals are completely filled. Hence, out of 80 electrons, half of electrons i.e. 40 electrons will have spin up.

(g) 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?

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

(h) Consider the following ionization energies for aluminum:

$Al_{(g)} \longrightarrow Al^{+}_{(g)} + e^-$	**I_1 = 580 kJ/mol**
$Al^{+}_{(g)} \longrightarrow Al^{2+}_{(g)} + e^-$	**I_2 = 1815 kJ/mol**
$Al^{2+}_{(g)} \longrightarrow Al^{3+}_{(g)} + e^-$	**I_3 = 2740 kJ/mol**
$Al^{3+}_{(g)} \longrightarrow Al^{4+}_{(g)} + e^-$	**I_4 = 11600 kJ/mol**

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

When electrons are removed successively, then due to removal of electron, hold of the nucleus on the remaining electrons increases. So, it becomes difficult to remove next electron. Thus ionization energy increases. Hence, the order of ionization energies is

$I_1 < I_2 < I_3 < I_4$

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

This shows that there are three electrons in the valence shell of $A\ell$. The fourth electron has to be removed from the inner shell that is closer to the nucleus. So, more energy is required. Hence, I_4 is much bigger than I_3.

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

When an atom loses one or more electrons, it becomes a positive ion. The positive ion becomes smaller than the neutral atom. This is because the loss of electrons reduces electronic repulsions. Thus remaining electrons are pulled closer to the nucleus and size decreases. Hence, higher the positive charge smaller is the size. Therefore, the order of increasing size of four aluminium ions is.

$$A\ell^{4+} < A\ell^{3+} < A\ell^{2+} < A\ell^{+}$$

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

According to Aufbau principle, the energy wise general order of filling orbitals up to 4p is

1s 2s 2p 3s 3p 4s 3d 4p

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

According to Aufbau principle: the subshells are filled energy. The order of filling is obtained by $(n + \ell)$ values. Lower the $(n+\ell)$ value lower will be the energy of subshell and is filled first

Subshell	$(n+\ell)$ value	n value	Remarks
4s	4+0=4	4	Thus, 4s orbital has lower energy than 3d.
3d	3+2=5	3	

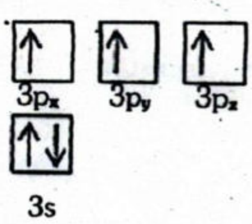
Hence, 4s orbitals is filled before than 3d orbital

(j) 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.

The electronic configuration is

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

Thus, P has 5 valence electrons ($3s^2\ 3p^3$). So, its orbital box diagram for the *valence electrons* is



DESCRIPTIVE QUESTIONS

Q3. What are quantum numbers? Describe briefly principal and spin quantum numbers.

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Q4. Draw the shapes of s, p and d-orbitals. Justify these by keeping in view the azimuthal and magnetic quantum numbers.

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Q5. What do you mean by successive ionization energies? How the electronic shell structure of magnesium (Mg) is derived from the successive ionization energies?

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ENTRANCE TEST MCQs

1. **In atomic particles:** MCAT 2009
   (A) Mass of neutron is almost equal to mass of electron
   (B) e/m of a proton is almost equal to e/m of electron
   (C) Mass of proton is almost equal to mass of electron
   (D) Charge of proton is almost equal to charge of electron
2. **What is the proton or atomic number of an element with four unpaired electrons in ground state?** MDCAT 2022
   (A) 6 (B) 14 (C) 22 (D) 26
3. **Which quantum number helps to study the orientation of an orbital in space?** MCAT 2010
   (A) Principal Quantum Number
   (C) Magnetic Quantum Number
   (B) Spin Quantum Number
   (D) Azimuthal Quantum Number
4. **Quantum number values for 2p orbitals are** MDCAT 2020
   (A) $n = 2, \ell = 1$ (B) $n = 1, \ell = 2$ (C) $n = 1, \ell = 0$ (D) $n = 2, \ell = 0$
5. **Which pair of transition elements shows abnormal electronic configuration?** MCAT 2012
   (A) Sc and Zn (B) Cu and Sc (C) Zn and Cu (D) Cu and Cr
6. **Correct order of energy in the given subshells is:** MCAT 2013
   (A) 5s > 3d > 3p > 4s (B) 5s > 3d > 4s > 3p (C) 3p > 3d > 5s > 4s (D) 3p > 3d > 4s > 5s
7. **Number of electrons in the outermost shell of chloride ion (Cl^-) is:** MCAT 2013
   (A) 17 (B) 3 (C) 1 (D) 8
8. **The relative energies of 4s, 4p and 3d orbitals are in the order** MCAT 2012
   (A) 3d < 4p < 4s (B) 4s < 3d < 4p (C) 4p < 4s < 3d (D) 4p < 3d < 4s
9. **Electronic configuration of Manganese (Mn) is** MCAT 2014
   (A) [Ar] $4s^2\,3d^5$ (B) [Ar] $4s^2\,3d^4$ (C) [Ar] $4s^1\,3d^5$ (D) [Ar] $4s^2\,3d^7$
10. **Which one of the following pairs has the same electronic configuration as possessed by Neon (Ne-10)?** MCAT 2015
    (A) Na^+, Cl^- (B) K^+, Cl^- (C) Na^+, Mg^+ (D) Na^+, F^-
11. **There are four orbitals s, p, d and f. Which order is correct with respect to the increasing energy of the orbitals?** MCAT 2015
    (A) 4s < 4p < 4d < 4f (B) 4p < 4s < 4f < 4d (C) 4s < 4f < 4p < 4d (D) 4f < 4s < 4d < 4p
12. **Which element has the ground state electronic configuration $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$?** MDCAT 2022
    (A) Ar (B) He (C) Ne (D) H_2
13. **Which of the electronic configuration of nitrogen is correct** ECAT 2016 MDCAT 2019
    (A) $1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$ (B) $1s^2, 2s^2, 2p_x^2, 2p_y^1, 2p_z^1$ (C) $1s^2, 2s^2, 2p_x^2, 2p_y^2, 2p_z^2$ (D) $1s^2, 2s^2, 2p_x^2, 2p_y^2, 2p_z^1$

14. **Which of the following electronic configurations represents an element that forms simple ion with a charge of –3** ECAT 2016
(A) $1s^2,2s^2,2p^6,3s^2,3p^1$ (B) $1s^2,2s^2,2p^6,3s^2,3p^3$ (C) $1s^2,2s^2,2p^6,3s^2,3p^6,3d^1,4s^2$ (D) $1s^2,2s^2,2p^6,3s^2,3p^6$

15. **The correct electronic configuration of $_{24}Cr$** MDCAT 2018 MDCAT 2019
(A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$ (B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4$
(C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$ (D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4f^6$

16. **Which two elements are isotopes?** MDCAT 2019
(A) $^{16}_{8}X$ and $^{16}_{8}Y$ (B) $^{18}_{9}X$ and $^{20}_{10}Y$ (C) $^{14}_{8}X$ and $^{15}_{8}Y$ (D) $^{12}_{6}X$ and $^{12}_{7}Y$

17. **According to which scientist, the probability of finding electron at a certain position is possible?** MDCAT 2022
(A) Bohr's (B) de-broglie (C) Hund's (D) Schrodinger

18. **When 6d orbital is filled, the entering electron goes into?** MDCAT 2023
(A) 7f (B) 7d (C) 7p (D) 7s

19. **Which element has the ground state electronic configuration of $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$?** MDCAT 2023
(A) Ar (B) Cl (C) Na (D) S

20. **The p-orbital has?** MDCAT 2024
(A) 2 lobes (B) 3 lobes (C) 4 lobes (D) 5 lobes

21. **Which of the following electronic configuration is correct for carbon?** MDCAT 2024
(A) $1s^2\,2s^2\,2p^3$ (B) $1s^2\,2s^2\,2p^4$ (C) $1s^2\,2s^2\,2p^2$ (D) $1s^2\,2s^2\,2p^1$

Answers

Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans
1	D	2	D	3	C	4	A	5	D	6	B	7	D	8	B
9	A	10	D	11	A	12	A	13	A	14	B	15	A	16	C
17	D	18	C	19	A	20	A	21	C						

1. What is the difference between an orbit and orbital?

	Orbit		Orbital
1	It is the circular path on which electron revolves around the nucleus.	1	It is the region in space in which probability of finding electron is maximum (about 95%)
2	This term is used in the Bohr's theory of atomic structure.	2	This term is used in quantum mechanical model.
3	It is two dimensional	3	It is three dimensional
4	Number of electrons in an orbit is given by the formula, $2n^2$. Where 'n' is the number of orbit.	4	Each orbital contains maximum two electrons.
5	In this exact position of electron is indicated.	5	Only probability of electron is given in an orbital.

2. Differentiate between atomic emission and atomic absorption spectrum

	Line Emission Spectrum Or (Atomic Emission Spectrum)		Line Absorption Spectrum Or (Atomic Absorption Spectrum)
1	In this bright lines are separated by dark bands.	1	In this dark lines are separated by bright bands.
2	It is formed when the substance is in excited state.	2	It is formed when the substance is in unexcited state.
3	It is formed when the substance is excited to vapour state	3	It is formed by transparent gases, transparent liquids and solids.
4	For its formation electron jumps from higher energy level to lower energy level and emit energy as light. The emitted radiations are indicated by coloured lines.	4	For its formation electron jumps from lower energy level to higher energy level by absorbing energy. The absorbed radiations are indicated by dark lines.
5	Emission spectrum of sodium has two yellow lines separated by dark bands.	5	Absorption spectrum of sodium has two dark lines separated by bright bands.

3. *Write electronic configuration of the following elements*

$_{24}Cr$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

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

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

$_{26}Fe$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

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

OBJECTIVE:Time: 10 Minutes: Marks: 08

Q1. Choose the correct answer and encircle it.

1. How many electrons can be accommodated in a sub-shell with n=3, ℓ=1?
(A) 8 (B) 6 (C) 18 (D) 32

2. The electronic configuration of an atom is $1s^2$, $2s^2$, $2p^4$. The number of unpaired electrons in this atom is
(A) 2 (B) 0 (C) 4 (D) 6

3. When atoms are volatilized, they form
(A) continuous spectrum (B) line spectrum
(C) electromagnetic spectrum (D) none

4. Splitting of spectral lines of the hydrogen atom under the influence of electric field and magnetic field show the presence of
(A) shells (B) sub-shells
(C) orbitals (D) none of these

5. Quantum number values for 2p orbitals are
(A) $n = 2, \ell = 1$ (B) $n = 1, \ell = 2$
(C) $n = 1, \ell = 0$ (D) $n = 2, \ell = 0$

6. The electron in a shell is filled according to formula:
(A) $2n^2$ (B) $2(2\ell+1)$ (C) $(2\ell+1)$ (D) None of these

7. Which of the following element can be doped with Si to prepare n-type semiconductor.
(A) B (B) Ga (C) As (D) Ge

8. Free radical among the following is
(A) Cl (B) Be (C) Ne (D) HO^-

Fill in the correct option					Write Correct option here
1.	Ⓐ	Ⓑ	Ⓒ	Ⓓ	
2.	Ⓐ	Ⓑ	Ⓒ	Ⓓ	
3.	Ⓐ	Ⓑ	Ⓒ	Ⓓ	
4.	Ⓐ	Ⓑ	Ⓒ	Ⓓ	
5.	Ⓐ	Ⓑ	Ⓒ	Ⓓ	
6.	Ⓐ	Ⓑ	Ⓒ	Ⓓ	
7.	Ⓐ	Ⓑ	Ⓒ	Ⓓ	
8.	Ⓐ	Ⓑ	Ⓒ	Ⓓ	

SUBJECTIVE: Time: 60 minutes Marks: 32

Section – I

Q2. Answer the following short questions. **(2×12=24)**

(i) Differentiate between line spectrum and continuous spectrum.
(ii) Calculate the number of protons, neutrons and electrons in $^{56}_{26}Fe$?
(iii) How an emission spectrum is obtained?
(iv) Define Moseley's law. Give its mathematical expression.
(v) Distribute the electrons in $_{29}Cu$ and $_{35}Br$
(vi) Define spectrum. Name its two types
(vii) Differentiate between orbit and orbital?
(viii) State Paulis Exclusion principle with an example
(ix) In an electric field, electron is deflected to a greater extent than proton. Explain?
(x) How p-type semiconductors are designed?
(xi) What are free radicals? Give examples.
(xii) Write equations that describe:
(a) the 2nd ionization energy of Be (b) the 5th ionization energy of sulfur.

Section – II

(8×1=08)

Q3. **(a)** What do you mean by successive ionization energies? How the electronic shell structure of aluminium ($A\ell$) is derived from the successive ionization energies? **(04)**
(b) What are quantum numbers? Explain the significance of azimuthal quantum number? **(04)**

ANSWERS TO MCQs: TEST YOUR SKILLS

Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans	Q#	Ans
1	B	2	A	3	B	4	B	5	A	6	B	7	C	8	A