

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

Chapter

02

ATOMIC STRUCTURE

Student Learning Outcomes

After studying this chapter, students will be able to:

- Describe protons, neutrons, and electrons in terms of their relative charge and relative masses. (Understanding)
- Recognize that the terms atomic and proton number represent the same concept. (Understanding)
- Recognize that the terms mass and nucleon number represent the same concept. (Understanding)
- Explain the change in atomic and ionic radius across a period and down a group. (Understanding)
- Describe the behaviour of beams of protons, neutrons and electrons moving at the same velocity in an electric field. (Understanding)
- Determine the number of protons, neutrons, and electrons present in both atoms and ions given atomic or proton number, mass/or nucleon number and charge. (Knowledge)
- Relate Quantum Numbers to electronic distribution of elements. (Understanding)
- Account for the variation in successive ionization energies of an element. (Understanding)
- Define terms related to electronic configuration (some examples include: shells, subshells, orbitals, principal quantum number (n), ground state). (Knowledge)
- Describe the order of increasing energy of the sub-shells (s, p, d, and f). (Understanding)
- Describe that, each atomic shell and sub-shell are further divided into degenerate orbitals having the same energy. (Understanding)
- Apply Aufbau principle, Pauli's exclusion principle and Hund's rule to write the electronic configuration of elements. (Application)
- Describe the number of orbitals making up s, p, d, and f sub-shells, and the number of electrons that can fill s, p, d, and f sub-shells. (Understanding)
- Describe the shapes of s, p, and d orbitals. (Understanding)
- Determine the electronic configuration of elements and their ions with proton numbers. (Knowledge)
- Some examples include:
 - a. simple configuration e.g., 2,8
 - b. Sub-shells e.g., $1s^2, 2s^2, 2p^6, 3s^1$
 - c. students should be able to determine both of these from periodic table and are not required to memorize these
 - d. Students should understand that chemical properties of an atom are governed by valence electrons.
- Explain the electronic configurations to include the number of electrons in each shell, subshell and orbitals. (Understanding)
- Explain the electronic configurations in terms of energy of the electrons and Inter-electron repulsion. (Understanding)
- Determine the electronic configuration of atoms and ions given the proton or electron number and charge. (Understanding)

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- Describe free radical as a species with one or more unpaired electrons. (Understanding)
- Illustrate the importance of electronic configurations and development of new materials for electronic devices. (For example, semiconductors such as silicon has a specific electronic configuration that makes them ideal for their use in electronic devices) (Understanding)
- Deduce the electronic configurations of elements using successive ionization energy data. (Application)
- Deduce the position of an element in the periodic table using successive ionization energy data. (Application)
- Explain that ionization energies are due to the attraction between the nucleus and the outer electrons. (Understanding)
- Explain how ionization energy helps account for the trends across the period and down a group of the periodic table. (Understanding)
- Explain the factors influencing the ionization energies of elements in terms of nuclear charge, atomic/ionic radius, shielding by inner shells and sub-shells and spin pair repulsion. (Understanding)

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

⇒ Atomic Number (Z)

"Atomic number is the number of proton in the nucleus of an atom".

In 1913, Moseley observed that when different elements were bombarded with cathode rays, the X-rays of some characteristic frequencies were produced. It was found that the "the square root of the frequency of the X-rays was directly proportional to the atomic number of an element Z".

$$\sqrt{\text{frequency } (\nu)} \propto Z$$

- Moseley concluded that this number, i.e. the atomic number Z was a fundamental property of an element.
- It is also called proton number.

⇒ Nucleon Number

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

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

$$A = Z + N$$

Example: An atom of an element X having atomic number Z and mass number A is described as ${}_Z^AX$, e.g.

${}_{13}^{27}\text{Al}$. Number of neutrons in an atom can be calculated as

$$N = A - Z$$

For example, ${}_{13}^{27}\text{Al}$

Atomic number/proton number (Z) = 13

Mass number/nucleon number (A) = 27

$$N = 27 - 13 = 14$$

Similarly, the number of electrons, protons, and neutrons can be justified for an ion as in the following example:

${}_{13}^{27}\text{Al}$ atom loses three electrons to form Al^{3+} then, For Al^{3+}

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

Similarly, ${}_{17}^{35}\text{Cl}$ gains an electron to form Cl^- ion, then, For Cl^-

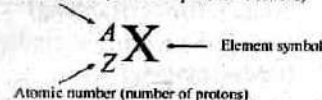


Rock Your Brain!

1. The total number of fundamental particles in an atom of Carbon - 14 is:

- (A) 6 (B) 8
(C) 14 (D) 20

Mass number (number of protons + neutrons)



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

In the cases where electron gain happens by the neutral atoms, say $^{16}_8\text{O}$ to $^{16}_8\text{O}^{2-}$, $^{31}_{15}\text{P}$ to $^{31}_{15}\text{P}^{3-}$ and $^{32}_{16}\text{S}$ to $^{32}_{16}\text{S}^{2-}$, 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

Thus the atomic number and proton number represent the same concept.

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 $\frac{1}{1836}$ times lighter than protons.

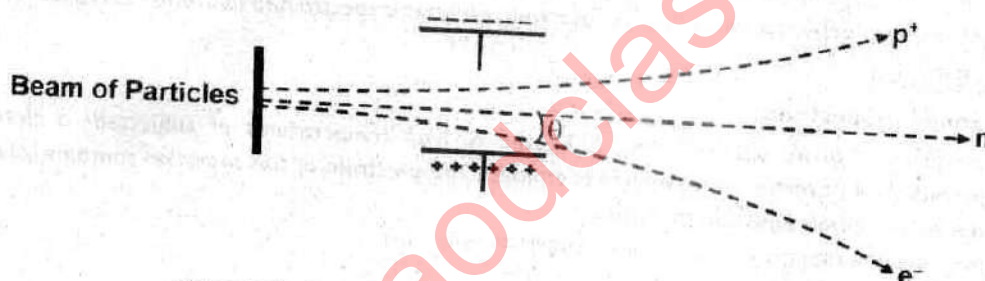


Figure: Behaviour of proton, electron and neutron in the electric field

The amount of deviation from its original direction of movement is measured in two ways.

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

This is possible if we imagine that after deflection, the particle moves in a circular path. Hence, the factors affecting the radius of deflection are reciprocal to that for the angle of deflection.

Properties of Fundamental Particles

The Table shows the properties of three fundamental particles electron, proton and neutron present in an atom.

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}



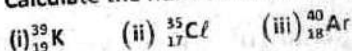
Rack Your Brain!

2. Cathode rays are negatively charged. Why?

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QUICK CHECK 2.1

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



Ans.

Number of Neutrons in:		
(i) $^{39}_{19}\text{K}$	(ii) $^{35}_{17}\text{Cl}$	(iii) $^{40}_{18}\text{Ar}$
Mass number = $A = 39$	Mass number = $A = 35$	Mass number = $A = 40$
Number of Protons = $Z = 19$	Number of Protons = $Z = 17$	Number of protons = $Z = 18$
Number of Neutrons = $N = A - Z$	Number of Neutrons = $N = A - Z$	Number of Neutron = $N = A - Z$
$= 39 - 19$	$= 35 - 17$	$= 40 - 18$
$= 20$	$= 18$	$= 22$

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

Ans. Electron is deflected the most because it is charged and has a very small mass.

Proton is deflected less because it is charged but much heavier.

Neutron is not deflected at all because it has no charge.

EXPERIMENTAL EVIDENCES FOR THE ELECTRONIC CONFIGURATION

The modern theory of electronic structure takes its origin from the Bohr Model of atom. Evidence for this and later models of the atoms derives principally from two sources; atomic spectra and ionization energies.

➤ Atomic Spectra

➤ 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. The spectrum of this radiation contains coloured lines and is called atomic emission spectrum".

The atomic emission spectrum of hydrogen is shown in Figure (a).

➤ 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. The wavelengths of the white light that has been absorbed by the atoms show up as dark lines on the spectrum. The spectrum of this radiation is called an atomic absorption spectrum".

The atomic absorption spectrum of hydrogen is shown in Figure (b).

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

(a) Emission Spectrum of Hydrogen



(b) Absorption Spectrum of Hydrogen

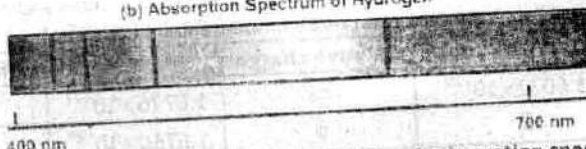


Figure: (a) atomic emission spectrum of hydrogen (b) atomic absorption spectrum of hydrogen

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- 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.
- The atomic spectra are the finger prints of the elements.

Figure shows the emission spectra of some elements.

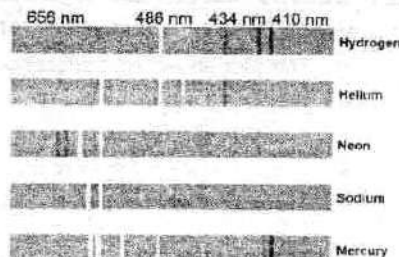


Figure: Spectral series of various elements as plotted by a spectrometer

(Exercise 1: Q. 6)

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

➔ Ionization Energy and Energy Levels (Electronic Shells)

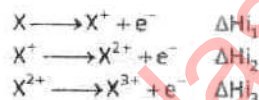
A major difference between electrons in different types of shells is their energy. We can investigate the electronic configuration of the atoms by measuring experimentally the energies of the electrons within them. This can be done by measuring ionization energies.

Ionization energies are used to investigate the electronic configurations of elements in two ways.

- Successive ionization energies of the same element
- First ionization energies for different elements

i) Successive Ionization Energies of the Same Element

We can look at an atom of a particular element and measure the energy required to remove each of its electrons, one by one.



We can continue to remove electrons from an atom until only the nucleus is left. We call this sequence of ionization energies, successive ionization energies. The successive ionization energies show clearly the arrangement of electrons in shells around the nucleus.

Example: If we take the **magnesium atom** as an example, and measure the energy required to remove successively the first electron, the second, the third, and so on. We obtain a graph when the ionization energies values are plotted against number of electrons as in Figure.

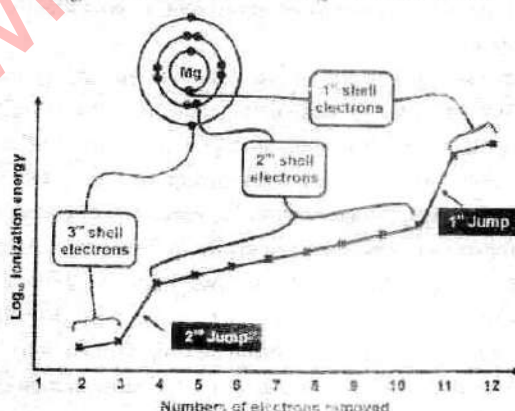
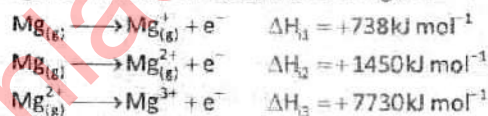


Figure: A plot of the successive ionization energies of Mg

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➤ Graph Depicting Successive Ionization Energies:

- This plot shows that **successive ionization energies increase when we move from the valence shell to the inner shells.**
- First two electrons are removed from the outermost shell and require lower energy for their removal.
- A very 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.
- A similar but much more enormous jump occurs when the eleventh and twelfth electrons are removed.
- These electrons are removed from the first, innermost shell, right next to the nucleus. 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.

ii) First Ionization Energies of Different Elements

The second way in which ionization energies show us the details of electronic configuration is to look at how the first ionization energies of elements vary with atomic numbers. The following figure shows a plot for the first 88 elements.

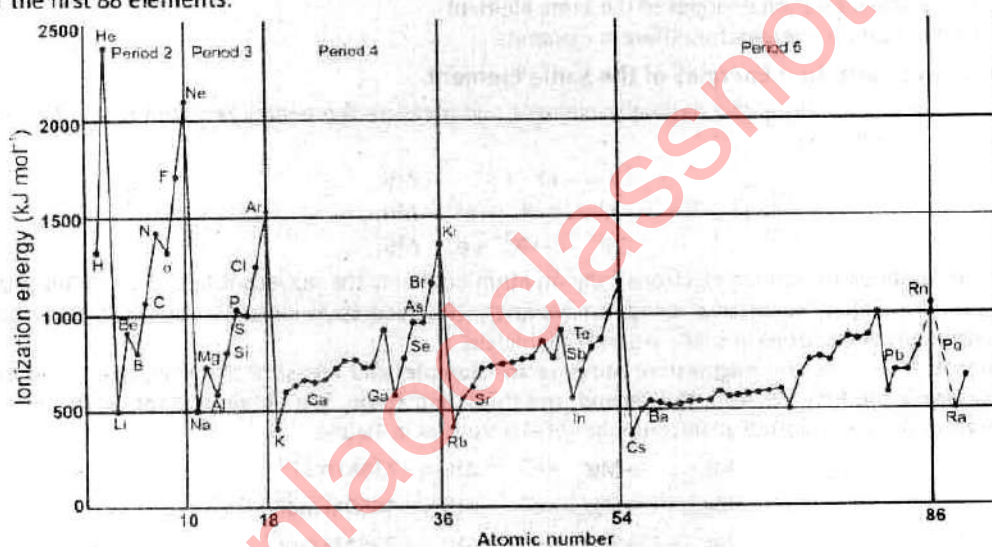


Figure: A plot of the ionization energies of first 88 elements against atomic number

This graph tells us the following:

- **All ionization energies are strongly endothermic;** it takes energy to separate an electron from an atom.
- **Down the Group:** As we go down a particular group, e.g., 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.*
- **Along a 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.

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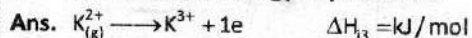
QUICK CHECK 2.2

a) Write equations that describe:

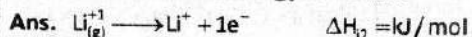
i. 1st ionization energy of calcium



ii. 3rd ionization energy of potassium



iii. 2nd ionization energy of lithium

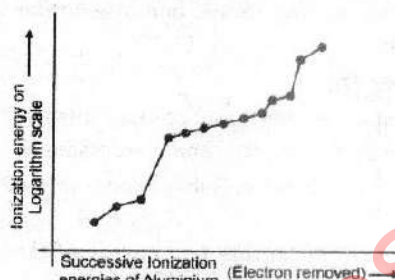


iv. 5th ionization energy of sulphur.

Ans. The 5th ionization energy of Sulphur is 7004.3 kJ/mol.

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

Ans.



c) The first ΔH_{i1} and the second ΔH_{i2} ionization 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,

Ans.

Element	ΔH_{i1}	ΔH_{i2}
II	520	7300

ii. a reactive non-metal,

Ans.

Element	ΔH_{i1}	ΔH_{i2}
IV	1680	3380

iii. a noble gas

Ans.

Element	ΔH_{i1}	ΔH_{i2}
I	2372	5251

iv. a metal that forms a stable binary halide of the formula AX_2 ($X = \text{halogen}$).

Ans.

Element	ΔH_{i1}	ΔH_{i2}
III	900	1760

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QUANTUM NUMBERS

(Exercise L.Q.3)

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

- The Bohr 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 the n -quantum number. Since Schrodinger's model allowed the electron to occupy three-dimensional space, therefore, it required three coordinates, or three quantum numbers, to describe the orbitals in which electrons can be found.



Rack Your Brain!

3. Quantum number values for '3d' orbital will be:
- (A) $n = 3, \ell = 0$ (B) $n = 3, \ell = 1$
 (C) $n = 3, \ell = 2$ (D) $n = 3, \ell = 3$

- 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.*

➤ Principal Quantum Number (n)

- The principal quantum number, n , can have positive integral values 1, 2, 3, 4... designated by K, L, M, N..., and it corresponds the quantum number (number of orbit) in Bohr's model of the hydrogen atom.
- 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.
- 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 and is therefore less tightly bound to the nucleus.*
- The principal Quantum number, n , can also be used to calculate the maximum number of electrons in a shell by the formula $2n^2$.
- Therefore shells K, L, M and N can accommodate maximum electrons. 2, 8, 18, 32 respectively.



Rack Your Brain!

4. Quantum number value for 2s orbitals are:
- (A) $n = 2, \ell = 1$ (B) $n = 1, \ell = 2$
 (C) $n = 1, \ell = 0$ (D) $n = 2, \ell = 0$



Rack Your Brain!

5. What is the value of $(n + \ell)$ for the 3d sub-shell?
- (A) 3 (B) 4
 (C) 5 (D) 6



Rack Your Brain!

6. $(n + \ell)$ value for 5s orbital will be:
- (A) 3 (B) 5
 (C) 7 (D) 9

(Exercise L.Q.4)

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

➤ Azimuthal Quantum Number (ℓ)

- Azimuthal and also called Angular Momentum Quantum Number, ℓ , can have integral values 0 to $(n - 1)$ for each value of n .
- This quantum number describes the shape of the orbital.*
- The values 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., $\ell = 0$ ($n - 1$, 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.
 - If $n = 4$, there are four values of ' ℓ ', i.e. 0, 1, 2, and 3.



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7. For azimuthal quantum number $\ell = 3$, calculate the total values of magnetic quantum number (m).

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- The values of ' ℓ ' are designated by the letters s, p, d, and f, with which stand for sharp, principal, diffused and fundamental, respectively. These are the spectral terms used to describe certain features of spectral lines.
- Subshell:** 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 $2(2\ell + 1)$ as given in the Table.

Table: Shapes of orbital

Value of ℓ	0	1	2	3
Orbital designation	s	p	d	f
shape of orbital	spherical	polar (dumb bell)	cloverleaf (double dumb bell)	complicated
No. of electrons in a subshell, $2(2\ell + 1)$	2	6	10	14

- 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.

Table: Relationship between n, ℓ and subshells

Shell	Principal Quantum number n	(Azimuthal) Quantum Number (ℓ)	Subshells	No. of subshells in a shell
K	1	0	1s	1
L	2	0 1	2s 2p	2
M	3	0 1 2	3s 3p 3d	3
N	4	0 1 2 3	4s 4p 4d 4f	4

➤ Magnetic Quantum Number (m)

- 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 there are $(2\ell + 1)$ integral values of m as follows:

$$-\ell, \dots, 0, \dots, +\ell$$

- The values of m indicate the number of orbitals in a subshell.
- If $\ell = 0$, (s-subshell) there is only one possible value of m , i.e. 0. It means s-subshell has only 1 orbital.
- If $\ell = 1$, (p-subshells) there are three values of m ; -1, 0 and +1. It means p-subshells has three orbitals.
- If $\ell = 2$, (d-subshell), there are five values of m , namely, -2, -1, 0, +1 and +2. It means d-subshells have 5 orbitals.
- If $\ell = 3$, (f-subshell) there are seven values of m ; i.e. -3, -2, -1, 0, +1, +2, +3. It means f-subshells have 7 orbitals.

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.

- Degenerate Orbitals:** Orbitals of the same subshell have same energy and are called degenerate orbitals. These degenerate orbitals are differentiated from each other in the presence of magnetic field; hence the name of this quantum number, i.e. magnetic quantum number. The relationship between the magnetic quantum numbers is provided in Table.

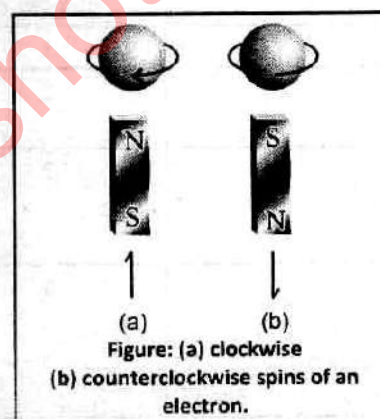
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Table: Relationship between l and m

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

⇒ 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. Figure shows the two possible spinning motions of an electron. One is clockwise and the other is anticlockwise.
- 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 $-\frac{1}{2}$.
- The anti-clockwise spin is represented by an arrow ($\uparrow$) pointing upwards, while the clockwise spin is represented by an arrow ($\downarrow$) pointing downwards.
- Each orbital can accommodate at the most two electrons provided the two electrons have opposite spins.
- Thus it takes three quantum numbers to describe an orbital but fourth quantum number to differentiate between the two electrons that can occupy an orbital.



Did You Know?

In the n th 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(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.
 $n = 1, 2, 3, 4, \dots\dots$ ($n \neq 0, -ve, \text{fractional}$)
- The value of ' n ' represents the shell or energy level in which the electron revolves around the nucleus.

n	1	2	3	4
Shell	K	L	M	N

- These value of ' n ' tells us about the radius of orbit.

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- ii. **Azimuthal Quantum Number (ℓ)**
- The presence of several fine lines in an individual line spectrum can be explained in term of Azimuthal quantum number. Azimuthal quantum number is also known as angular momentum or secondary or subsidiary quantum number.
 - It is represented by ' ℓ '.
 - Its values are $\ell = 0, 1, 2, 3, \dots, (n - 1)$
 - Its value depends upon the value of n .
- iii. **Magnetic Quantum Number (m)**
- To explain one of the defects of Bohr's model i.e. Zeeman effect (splitting of spectral lines due to strong magnetic field) a third quantum number called magnetic quantum number (m) has been proposed.
- This quantum no. is represented by ' m '.
 - Where $m = 0, \pm 1, \pm 2, \pm 3$
- iv. **Spin Quantum Number (s)**
- The doublet line structure of line spectrum can be explained by spin Quantum number.
 - Alkali metals have one electron in their outermost shell. We can record their emission spectra, when the outermost electron jumps from an excited state to a ground state. When the spectra are observed by means of high resolving power spectrometer, each line in the spectrum is found to consist of pair of lines, this is called doublet line structure.
- b) **For an electron(s):**
- i. If $n = 2$ and $\ell = 1$, how many orientations in space are possible?
- Ans. $2p$ has 3 orientations i.e. p_x, p_y, p_z
- ii. If $n = 3$ and $\ell = 2$, which shell and subshell does the electron belong to?
- Ans. $3d$
Shell number 3 (M-shell)
Subshell = d
- iii. If $\ell = 2$, find all possible values of m and maximum number of electrons for m .
- Ans. d subshell
It has five orientation $m = +2, +1, 0, -1, -2$.

SHAPES OF ATOMIC ORBITALS

Atomic Orbital

An atomic orbital is defined as the three dimensional region in space around the nucleus in which the probability of finding the electron is maximum.

Let us discuss the different orbitals one by one.

➤ s-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.

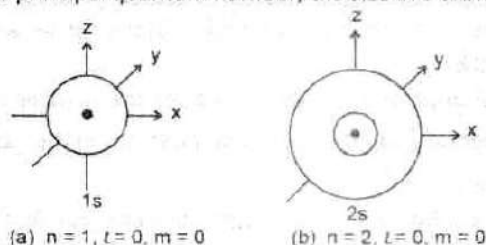


Figure: The boundary surface for, (a) $1s$ (b) $2s$

➤ p-orbitals

- The distribution of electron density for a $2p$ orbital is shown in Figure.

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

- The electron density is not distributed in a spherically symmetric fashion as in an s orbital. Rather, a 'p' orbital has two lobes on any of the axis.
- The p orbitals are named as p_x , p_y and p_z , accordingly to their axes.

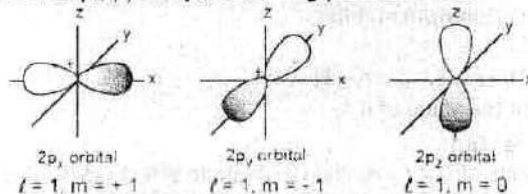


Figure: The 2p orbitals shapes of 2P orbitals

d and f orbitals

- 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.
- The shapes for the five d orbitals ($l = 2$) are shown in Figure.

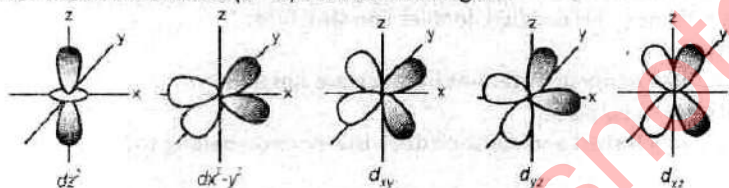


Figure: The shapes of the five 3d orbitals

An f subshell has seven orientations in space, i.e. there are seven f orbitals. However, 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. According to the wave mechanical model of atom, an orbital represents a region of space around the nucleus where there is high probability of finding an electron. It is a 3-D wave function that describes the electrons state and energy.

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

Ans. For p-orbital value of $l = 0$, $m = 0, +1, -1$
So there are three orientations of p-orbital.

ELECTRONIC CONFIGURATION

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

- In case of subshells, the electronic configuration is described by a notation that 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. For example, the configuration of ${}^9_4\text{Be}$ atom with two electrons in the '1s' subshell and two electrons in the '2s' subshell is written $1s^2 2s^2$.
- 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 $s = +1/2$ and downward when $s = -1/2$. The orbital diagram of boron (${}^{11}_5\text{B}$) is as follows

$${}^{11}_5\text{B} = 1s^2, 2s^2, 2p_x^1, 2p_y^0, 2p_z^0$$

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

➤ Distribution of Electrons 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 (1 s orbital)
 - L shell** ($n=2$): 8 electrons maximum (2 s and 2 p orbitals)
 - M shell** ($n=3$): 18 electrons maximum (3s, 3p, and 3d orbitals)
 - N shell** ($n=4$): 32 electrons maximum (4s, 4p, 4d, and 4 f orbitals)

For example:

- 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)

➤ Distribution of Electrons in Subshells

Following rules are applied to fill the orbitals of multi-electron atoms.

Aufbau Principle

Aufbau principle is also known as the building up principle. This principle says *that the subshells in an atom are filled with electrons in an increasing order of their energy values.*

Arrangement of the subshells energy wise:

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).

Thus the order of filling subshells with electrons may be obtained from the summation ($n + \ell$). ($n + \ell$) values are given in Table.

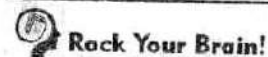
Table: ($n + \ell$) values of various sub-shells.

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

- a) The subshell having lower ($n + \ell$) value has lower energy and is filled first.
For example, 4s orbital has ($n + \ell$) = 4 + 0 = 4 and 3d orbital has ($n + \ell$) = 3 + 2 = 5. Since ($n + \ell$) 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 ($n + \ell$) values, then the subshell with lower 'n' value will be filled first.
For example, both 4p and 3d subshells have $n + \ell$ 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.

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



8. Calculate the number of electrons in s, p, d and f sub shells from the formula and write separately.

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

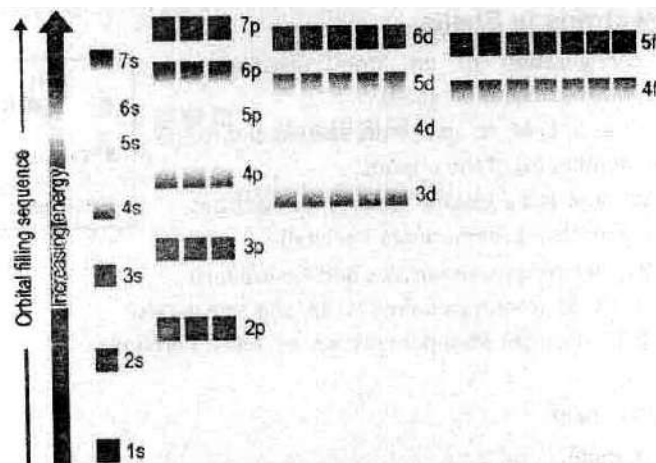


Fig: The diagram shows the energy of each subshell. Each box on the diagram represents an atomic orbital. So, the order of filling of various subshells with electrons obtained by this rule is given below Figure.

➤ 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".

➤ Explanation:

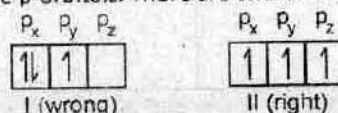
In the first shell of helium (He) atom, there are two electrons. They are present in 1 s orbital. According to the concept of quantum numbers and Pauli's exclusion principle, the values of their quantum numbers are: Table values of quantum numbers of two electrons in the same orbital

Electron	n	l	m	s
Electron 1	1	0	0	$+\frac{1}{2}$ (clockwise)
Electron 2	1	0	0	$-\frac{1}{2}$ (anticlockwise)

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

➤ Hund's Rule

This rule gives an idea for filling electrons into the orbitals having equal energies. For example, three p-orbitals, i.e., p_x , p_y and p_z 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:



The structure II is correct. Hund's rule, states that,

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.

- 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 Figure.

- Each box represents an atomic orbital.
- The boxes (orbitals) can be arranged in order of increasing energy from bottom to top.

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

- An electron is represented by an arrow.
- The direction of the arrow represents the 'spin' of the electron.
- 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.

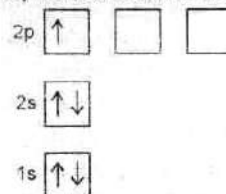


Figure: The electronic configuration of boron

Figure: shows the electronic structures of carbon, nitrogen and oxygen to illustrate these points

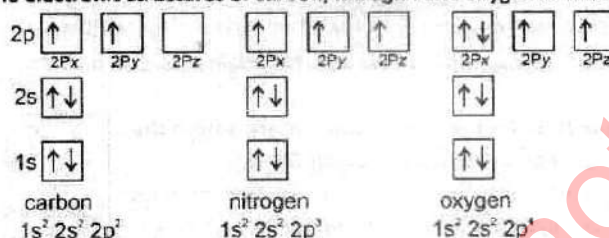


Figure: When adding electrons to a particular subshell, the electrons are only paired when no more empty orbitals are available. The electron configurations of some elements of the periodic table in the light of the abovementioned principles are given in Table.

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

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

ELECTRONIC CONFIGURATION AND THE PERIODIC TABLE

- We have seen that the electron configurations of elements are related to their position in the periodic table. The periodic table is structured so that elements with the same pattern of outer-shell (valence) electron configuration are arranged in same groups.
- We can easily write the electron configuration of an element based on its location in the periodic table. The pattern is summarized in Figure. Notice that the elements can be grouped in terms of the type of orbital into which the electrons are placed.
- On the left are two columns of elements. These elements, known as the alkali metals and alkaline earth metals (groups 1 and 2), are those in which the outer-shell s orbitals are being filled.*
- We see that the group 1 and 2 elements all have ns^1 and ns^2 outer configurations respectively.
- When we jump to group 13 elements, we find that they 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.*
- 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.*
- Below the main portion of the table are two rows that contain fourteen columns. These elements are often referred to as the f-block elements. These are the ones in which the 'f' orbitals are being filled. Recall that the numbers 2, 6, 10, and 14 are precisely the number of electrons that can fill the s, p, d, and f subshells, respectively.



Rack Your Brain!

9. Write electronic configuration of Cu_{29} and Ag_{47} .



Rack Your Brain!

10. Distribute electrons in the orbitals of: (A) Ca_{20} (B) Cs_{55}



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) $_{48}Cd = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}$
 $= [Kr]4d^{10}5s^2$
 (ii) $_{57}La = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 5d^1$
 $= [Xe]5d^16s^2$
- b) Write the complete electron configuration for antimony (Sb) with atomic number 51.
- Ans. $_{51}Sb = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 4d^{10}, 5s^2, 5p^3$
 $= [Kr]4d^{10}5s^25p^3$
- c) How many unpaired electrons are there in each atom of $_{51}Sb$?
- Ans. Number of unpaired electrons in $_{51}Sb = 03$ (three)
 $5p_x^1, 5p_y^1, 5p_z^1$
- d) Write down the electronic configuration of valence shell of the following in terms of orbitals:
- i) $_{13}Al$ ii) $_{14}Si$ iii) $_{28}Ni$
- Ans. The valence shell electronic configurations in terms of orbitals for the given elements:
 i) Aluminium ($_{13}Al$): $3s^2 3p^1$ ii) Silicon ($_{14}Si$): $3s^2 3p^2$ iii) Nickel ($_{28}Ni$): $3d^8 4s^2$

Valence Electrons

The electrons in an atom in the outermost shell are called valence electrons. These are such electrons that 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.

Figure shows a periodic table which include the valence shell configurations included. Note the similarity in electron configuration within any group (column) of elements.

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

The figure shows the modern periodic table with electronic configurations for each element. The configurations are written in the format $ns^x np^y$ for main group elements, and $(n-1)d^x ns^y$ for transition elements, and $(n-2)f^x (n-1)d^y ns^z$ for inner transition elements.

Figure: The electronic distribution of elements of modern periodic table

➤ Classification of Elements of Periodic Table

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. Similarly, 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 as we go through the elements in the periodic table. From this we can write down the building-up order. The overall sketch of periodic table is given in Figure. This shows the blocks of elements.

The diagram shows the periodic table with blocks labeled by their valence subshells: s, p, d, and f. The s-block includes groups 1 and 2. The p-block includes groups 13 to 18. The d-block includes transition elements (groups 3 to 10). The f-block includes lanthanides and actinides.

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

Figure shows a periodic table indicating this pattern. For example, in the blue colored area, an 'ns' subshell is being filled. In the green colored area, an np subshell is being filled. In the yellow area, an $(n-1)d$ subshell is being filled. In the light golden area and $(n-2)f$ subshell is being filled.

We start building-up order by starting with the first period, in which the '1s' subshell is being filled. In the second period, we have '2s'; then staying in the same period but jumping across, we have '2p'. In the third period, we have '3s' and '3p'; in the fourth period, '4s', '3d' and then '4p'. This pattern should become clear enough to visualize with a periodic table.

- The classification of elements of periodic table has been done into metals, non-metals, representative, transition elements, periods, groups and blocks (s, p, d, f). This division is done on the basis of electronic configuration and their valence electrons. The chemical properties of various categories of elements discussed above can be easily assessed.

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

☑ QUICK CHECK 2.6

- a) An element has the electronic configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^5$
- Which block in the Periodic Table does this element belong to?
Ans. p-block
 - Which group does it belong to?
Ans. Halogens
 - Which period does it belong to?
Ans. 5
 - Identify this element.
Ans. Iodine
- 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? Name it.
Ans. d-block and the element is Chromium (Cr).

ELECTRONIC CONFIGURATION OF IONS AND FREE RADICALS

☞ Ions

Positive 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 $1s^2 2s^2 2p^6$. Note that this is the same as the electronic configuration of neon, the element with 10 electrons in each atom.

Negative Ions: *Negative ions are formed when atoms gain electrons.*

The sulfide ion, S^{2-} (proton number = 16), has 18 electrons. Its electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^6$, which is the same as argon, the element with 18 electrons in each atom.

Note that, in general, electrons in the outer subshell are removed when metal atoms form the positive ions. However, the d-block elements behave slightly differently. Reading across the Periodic Table from potassium to zinc, the 4 s subshell fills before the 3 d subshell. But when atoms of a d-block element lose electrons to form ions, the 4 s electrons are lost first.

For example:

- 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".

For example: Free chlorine atom:

$Cl^{\bullet}$: 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 $Cl^{\bullet}$. Apart from single atoms, groups of atoms can also be free radicals. For example, $OH^{\bullet}$, $CH_3^{\bullet}$, etc.

☑ QUICK CHECK 2.7

Write electronic configurations for the following ions and free radicals:

- | | | |
|---|--|---|
| i. $Al (Z = 13)$ | ii. $O^{2-} (Z = 8)$ | iii. $Fe^{3+} (Z = 26)$ |
| Ans. $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$ | Ans. $1s^2, 2s^2, 2p^6$ | Ans. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5$ |
| iv. $Cu^{2+} (Z = 29)$ | v. $Cu (Z = 29)$ | |
| Ans. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^9$ | Ans. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^1$ | |

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

SCHOLAR CHEMISTRY (22)

ELECTRONIC CONFIGURATION AND THE FORMATION OF SEMICONDUCTORS

➤ **Semi-Conductors**

"The materials whose electrical conductivity lies in between that of conductors and insulators are called **semiconductors**".

According to number of elements, these are classification into two classes:

- (i) **Elemental Semiconductors** (ii) **Compound 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.

Silicon as Semiconductor:

Consider the example of Si and explore how it can be converted into a p-type and n-type semiconductors.

The electron configuration of $_{14}\text{Si} = 2, 8, 4$; meaning 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, where adding trivalent impurities (like boron or aluminum) creates a P-type semiconductor, while adding pentavalent impurities (like phosphorus or arsenic) creates an N-type semiconductor.

The difference lies in whether the added impurity creates "holes" (positive charge carriers) in the lattice, leading to P-type, or extra electrons (negative charge carriers) leading to N-type.

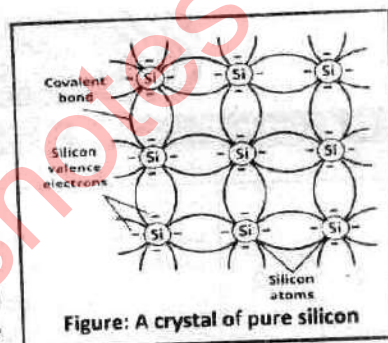


Figure: A crystal of pure silicon

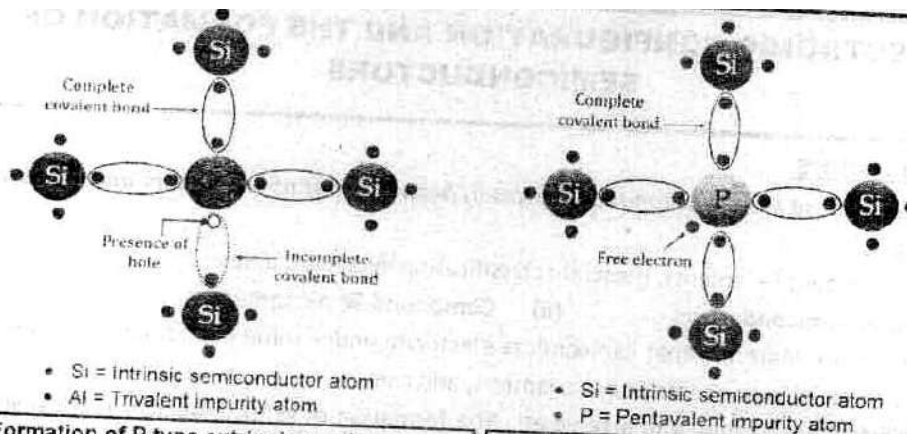
➤ **P-type Semiconductor Formation**

- Impurity atoms with three valence electrons (like Al) are added to the pure semiconductor.
- 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.
- For this reason, there are created holes in the crystal lattice, which act as positive charge carriers as shown in Figure.
- This process creates a positive-type semiconductor or P-type semiconductor.
- Electrons from an external current source can move through the semiconductor and it can act as a conductor.

➤ **N-type Semiconductor Formation**

- When impurity atoms with five valence electrons (like phosphorus) are added to the pure 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 and the fifth electron is an extra electron.
- These impurity atoms contribute extra electrons to the crystal lattice, which become free to move and act as negative charge carriers as in Figure.
- The result is an N-type semiconductor that can conduct electricity when connected to an external source.

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)



Formation of P type extrinsic semiconductor

Formation of N type extrinsic semiconductor

Figure: Doping and the formation of P-type and N-type semiconductor

Solution File

Rack Your Brain!

Sr. #	Option	Explanation						
1.	D	$P = 6, n = 8, e = 6, \Rightarrow \text{total} = 20$ fundamental particles.						
2.	S.Q	In the absence of any electric or magnetic field, cathode rays move in straight line perpendicular to the surface of cathode. In 1895, J. Perrin showed that when cathode rays are passed between the poles of the magnet, the cathode rays was deflected which showed that they have charge. In 1897, Sir J.J. Thomson established their electric charge by the application of electric field. The cathode rays deflected towards positive plate showing their negative nature.						
3.	C	Quantum number values for '3d' orbital will be $n = 3, l = 2$						
4.	D	Quantum number value for 2s orbitals are $n = 2, l = 0$						
5.	C	The value of $(n + l)$ for the 3d sub-shell = $3 + 2 = 5$.						
6.	B	The value of $(n + l)$ for the 5s sub-shell = $5 + 0 = 5$.						
7.	S.Q	For $l = 3$ as $l = 3$ and subshell is 'f'. This shows that f-subshell has 7 different ways of orientation in space because it has 7 values of magnetic quantum number When $l = 3$, then $m = 0, \pm 1, \pm 2, \pm 3 = 7$ values 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. $f_{x^2-z^2}$						
8.	S.Q	The formula for calculating electrons is $2(2l + 1)$. When $l = 0$ s-subshell total electrons = 2 $l = 1$ p-subshell total electrons = 6 $l = 2$ d-subshell total electrons = 10 $l = 3$ f-subshell total electrons = 14						
9.	S.Q	<table border="1"> <thead> <tr> <th>Element</th><th>Electronic Configuration</th></tr> </thead> <tbody> <tr> <td>Copper 29(Cu)</td><td>$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$</td></tr> <tr> <td>Iodine 53(I)</td><td>$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$</td></tr> </tbody> </table>	Element	Electronic Configuration	Copper 29(Cu)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$	Iodine 53(I)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$
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CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

Exercise

MULTIPLE CHOICE QUESTIONS (MCQs)

Q.1 Four choices are given for each question. Select the correct choice.

- 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 ${}^9_4\text{Be}$ 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) $\text{F} > \text{He} > \text{Mg} > \text{N} > \text{O}$
 - b) $\text{He} > \text{F} > \text{N} > \text{O} > \text{Mg}$
 - c) $\text{He} > \text{O} > \text{F} > \text{N} > \text{Mg}$
 - d) $\text{N} > \text{F} > \text{He} > \text{O} > \text{Mg}$
- 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) give 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

Answer Key with Explanations

Sr.No.	Option	Answer	Explanation
I.	c	the spatial orientation of the orbital	The magnetic quantum number (m) determines the orientation of an orbital in space (e.g., p_x, p_y, p_z).
II.	c	4p	After filling 3d, the next electron enters 4p (Aufbau order: $3d \rightarrow 4p$).
III.	a	$n=2, \ell=1$	For 2p orbitals: o Principal quantum number (n) = 2. o Azimuthal quantum number (ℓ) = 1 (for p-subshell).
IV.	c	4s	Quantum numbers $n=4, \ell=0$ indicate a 4s orbital ($\ell=0$ for s-orbitals).

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

V.	b	2, 0, 0	- Beryllium (Be, Z = 4): Electron configuration = $1s^2 2s^2$. - The 4th electron is in 2s: $n=2, \ell=0, m=0$.
VI.	b	He > F > N > O > Mg	Ionization energy trend: Noble gases (He) > halogens (F) > N > O > metals (Mg). - N > O due to half-filled 2p subshell stability.
VII.	b	2	A p orbital has two lobes (dumbbell shape) with a nodal plane at the nucleus.
VIII.	c	Mutually perpendicular to each other along the x, y, and z axes	The three p orbitals (p_x, p_y, p_z) are oriented 90° apart along Cartesian axes.
IX.	c	5	Each energy level has 5d orbitals ($d_{xy}, d_{xz}, d_{yz}, d_{x^2-y^2}, d_{z^2}$).
X.	a	Size	The principal quantum number (n) determines the size of an orbital. As the value of n increases, the size of the orbital also increases. The principal quantum number (n) does not determine the shape of an orbital; that's determined by the azimuthal quantum number (l). Spin of orbital is determined by Magnetic quantum number.
XI.	b	Three	Cobalt's atomic number is 27. Its electronic configuration is $[\text{Ar}] 3d^7 4s^2$. According to Hund's rule, the 3d electrons occupy the orbitals with maximum spin multiplicity, resulting in 3 unpaired electrons.

SHORT ANSWER QUESTIONS

Q.2 Attempt the following short-answer questions:

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

Ans. It is correct. There are three orientations of the p-orbital due to the three values of the magnetic quantum number (m_l). Here's the justification in simple scientific terms:

Quantum Numbers:

Each electron in an atom is described by four quantum numbers:

- Principal quantum number (n) – energy level
- Azimuthal/angular quantum number (l) – shape of orbital
- Magnetic quantum number (m_l) – orientation of orbital
- Spin quantum number (m_s) – spin direction of electron

For a p-orbital:

- The azimuthal quantum number (l) = 1
- (Because p-orbitals correspond to $\ell = 1$)
- The magnetic quantum number (m_l) can have three values:
 $m_l = -1, 0, +1$

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

Ans. The third ionization energy (I_3) of magnesium (Mg) is much bigger than its second ionization energy (I_2).

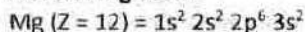
Here's the scientific justification:

- Ionization energy** is the energy needed to remove one electron from an atom or ion.
- I_1 : Remove 1st electron
- I_2 : Remove 2nd electron
- I_3 : Remove 3rd electron

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

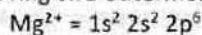
Electronic Configuration of Magnesium:

Neutral Mg atom:



- First ionization (I_1) removes one electron from $3s^2 \rightarrow$ becomes Mg^+
- Second ionization (I_2) removes second $3s$ electron $\rightarrow$ becomes Mg^{2+}
- Third ionization (I_3) would now try to remove an electron from the $2p^6$ (inner, noble gas core)

After removing two outermost $3s$ electrons, Mg^{2+} has the stable configuration of Neon (Ne):



Now, the third electron to be removed is from a completely filled inner shell ($2p$), which:

- Is very stable (noble gas configuration)
- Is closer to the nucleus
- Is more strongly attracted by nuclear charge

Hence, much more energy is needed to remove this tightly bound electron.

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

Ans. Ionization Energy Trend Recap:

- Increases across a period (left $\rightarrow$ right) due to increasing nuclear charge.
- Decreases down a group (top $\rightarrow$ bottom) as outer electrons are farther from the nucleus and more shielded.

Element	Group	Period	Electron Configuration	Notes
Li (Lithium)	1	2	$1s^2 2s^1$	Alkali metal (low IE)
K (Potassium)	1	4	$[\text{Ar}] 4s^1$	Lower IE than Li (further down)
Ca (Calcium)	2	4	$[\text{Ar}] 4s^2$	Slightly higher IE than K
S (Sulphur)	16	3	$[\text{Ne}] 3s^2 3p^4$	Nonmetal (higher IE)
Kr (Krypton)	18	4	$[\text{Ar}] 3d^{10} 4s^2 4p^6$	Noble gas (very high IE)

- Lowest first ionization energy = Potassium (K)
- Reason: It's a Group 1 element in Period 4 $\rightarrow$ outermost electron is loosely held and easily removed.
- Highest first ionization energy = Krypton (Kr)
- Reason: Noble gas with a full outer shell $\rightarrow$ very stable $\rightarrow$ requires lots of energy to remove an electron.

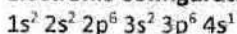
- d. Consider the electronic configuration of the potassium atom (atomic number 19).

Ans. Element: Potassium (K)

Atomic Number: 19

So it has 19 electrons.

Electronic Configuration:



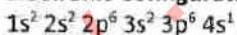
- (i) Write the full electronic configuration of potassium using the s, p, d, f notation.

Ans. Element: Potassium (K)

Atomic Number: 19

So it has 19 electrons.

Electronic Configuration:



- (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. Even though:

- $3d$ has a lower principal quantum number ($n = 3$)
 - $4s$ has a higher principal quantum number ($n = 4$)
- The $4s$ subshell has lower energy than $3d$ at the time potassium's electrons are being placed.

CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

Key Principles:

1. Aufbau Principle

Electrons fill orbitals in order of increasing energy, not just by n value. According to the Aufbau (building-up) principle:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$$

As you can see, **4s comes before 3d**.

2. Energy Level Overlap

- 4s orbital is farther from the nucleus but has less electron-electron repulsion, making it lower in energy than 3d.
- So, the first 19 electrons (in potassium) go into the 4s before starting to fill 3d.

Potassium ($Z = 19$):

- Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
- The 3d subshell remains empty in potassium because it's higher in energy at this point.

Conclusion:

The 4s subshell is filled before 3d in potassium because, despite having a higher principal quantum number, it is lower in energy at the point when electrons are added, following the Aufbau principle.

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.

Ans. Number of Protons = Atomic Number

Protons = 17

Number of Neutrons = Mass Number – Atomic Number

Neutrons = 18

Number of Electrons = Number of Protons (in a neutral atom)

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?

Ans. X element has:

- Atomic number = 17

- Mass number = 35

We previously identified it as chlorine (Cl).

Now, let's consider the Cl^- ion (chloride ion), which has a -1 charge.

Protons

- The number of protons never changes, even when forming ions.

Protons = 17

Neutrons

- Mass number = 35, Atomic number = 17

Neutrons = 18

Electrons (for Cl^- ion)

- Neutral chlorine atom has 17 electrons
- A -1 charge means it has gained one extra electron

Electrons = $17 + 1 = 18$

f. In the ground state of mercury $_{80}\text{Hg}$:

Ans. In the ground state of mercury (Hg), which has an atomic number of 80, we can determine several important features of the atom, such as its electron configuration, and the number of protons, neutrons, and electrons.

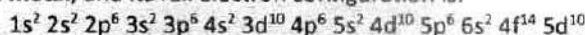
Basic Information for Mercury (Hg):

- Symbol: Hg
- Atomic number (Z): 80 $\rightarrow$ Protons = 80, Electrons = 80 (in a neutral atom)
- Mass number (most common isotope): 200 $\rightarrow$ So, Neutrons = $200 - 80 = 120$

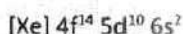
CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

Ground-State Electron Configuration:

Mercury is a transition metal, and its full electron configuration is:



In noble gas short hand:



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

Ans. Electrons can occupy atomic orbitals where the principal quantum number $n = 3$. $3n = 3 \times 3 = 3$.

For $n = 3$, the possible subshells (values of ℓ) are:

- $\ell = 0 \Rightarrow \ell = 0 \rightarrow 3s$
- $\ell = 1 \Rightarrow \ell = 1 \rightarrow 3p$
- $\ell = 2 \Rightarrow \ell = 2 \rightarrow 3d$

Each subshell has a certain number of orbitals, and each orbital holds 2 electrons:

Subshell	Orbitals	Electrons
3s	1	2
3p	3	6
3d	5	10

Total electrons when $n = 3$: $3n = 3 \times 3 = 3$:

$$2(3s) + 6(3p) + 10(3d) = 18 \text{ electrons}$$

So, 18 electrons can occupy atomic orbitals with principal quantum number $n = 3$.

ii. How many electrons occupy 4d atomic orbitals?

Ans. The 4d atomic orbitals can hold a maximum of 10 electrons. This is because the 4d subshell consists of 5 orbitals (each orbital can hold a maximum of 2 electrons), so $5 \text{ orbitals} \times 2 \text{ electrons} = 10 \text{ electrons}$ in total.

iii. How many electrons occupy 4p atomic orbital?

Ans. The 4p atomic orbital can hold a maximum of 6 electrons. Each atomic orbital, including the 4p orbitals, can accommodate up to 2 electrons, with each electron having opposite spins.

iv. How many electrons in the valence shell have spin "up" ($s = +\frac{1}{2}$)?

Ans. In an atomic orbital, each electron can have one of two possible spins: "up" (denoted as $s = +\frac{1}{2}$) or "down" (denoted as $s = -\frac{1}{2}$).

If you are asking how many electrons have a spin of $s = +\frac{1}{2}$ in a particular orbital, it depends on how many electrons are present in that orbital. For example:

- In a single atomic orbital (such as the 4p orbital), a maximum of 2 electrons can occupy it, one with spin $+\frac{1}{2}$ and one with spin $-\frac{1}{2}$.
- If there are multiple orbitals (such as in a subshell like the 4d or 4p), the electrons will fill the orbitals according to the Pauli exclusion principle, with one electron having spin $+\frac{1}{2}$ and the other having spin $-\frac{1}{2}$ in each orbital.

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So, in general, half of the electrons in any set of orbitals would have spin $s = -\frac{1}{2}$, assuming the electrons are paired and fill orbitals according to Hund's rule and the Pauli exclusion principle.
Conclusion: 80 electrons in total and thus 40 orbitals in Hg.

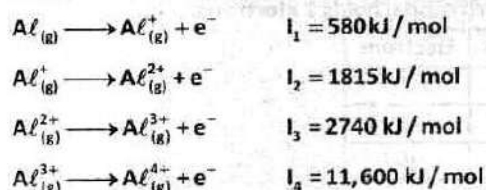
g. The successive ionization energies for an unknown element are

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

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

Ans. As there is greatest difference between 2nd and 3rd ionization energies, it means this element belongs to group 2A (Alkaline Earth's metal).

i. Consider the following ionization energies for aluminum:



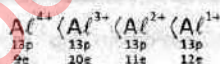
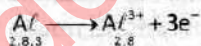
- (i) Account for the trend in the values of the ionization energies.
- (ii) Explain the large increase between I_3 and I_4 .
- (iii) List the four aluminum ions given in order of increasing size, and explain your ordering.

Ans.

(i) $I_4 > I_3 > I_2 > I_1$

(ii) As there is greatest difference between 3rd and 4th ionization energies, it means this element belongs to group 3A (Boron family).

(iii) After losing 3rd electron, Aluminium gets stable electronic configuration of noble gas.



In Al^{1+} there are 13 protons and 12 electrons so, control of nucleus is lesser on outer electrons so its size is greatest among above.

In Al^{4+} there are 13 protons and 9 electrons, there is more imbalance between proton to electron ratio so, control of nucleus is greatest and its size is the least among above.

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

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

Ans. (i) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

(ii) The 4s subshell is filled before the 3d subshell according to the Aufbau principle because the 4s orbital is lower in energy than the 3d orbital when these orbitals are empty or partially filled in neutral atoms.

1. Energy Considerations:

- Even though 4s has a higher principal quantum number ($n = 4$) than 3d ($n = 3$), the energy of an orbital depends on both the principal quantum number n and the azimuthal quantum number ℓ .
- For 4s: $n = 4, \ell = 0$, so $n + \ell = 4$
- For 3d: $n = 3, \ell = 2$, so $n + \ell = 5$

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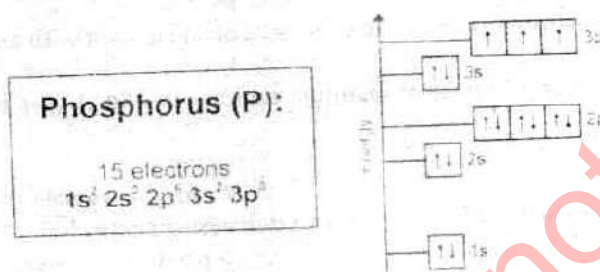
According to the $(n + l)$ rule (used in the Aufbau principle), orbitals with a lower $n + l$ value are filled first.

2. Electron Shielding and Penetration:

- The 4s orbital penetrates closer to the nucleus and experiences less shielding than the 3d orbital.
 - As a result, electrons in the 4s orbital are more tightly held and thus occupy a lower energy level initially.
- However, once electrons are added to the 3d subshell, the energy levels shift slightly. In ions of transition metals, for example, the 4s electrons are usually lost before the 3d electrons, because the 4s orbital becomes higher in energy after the 3d subshell is filled or partially filled.

i. 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.



DESCRIPTIVE QUESTIONS

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

Ans. See Page No. (36)

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

Ans. See Page No. (36)

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?

Ans. See Page No. (33)

ADDITIONAL SLOs BASED MCQs

1. What does the electron configuration $1s^2 2s^2 2p^6$ represent?

A. Carbon (C)
B. Oxygen (O)
C. Neon (Ne)
D. Helium (He)
2. Which subshell can hold a maximum of 10 electrons?

A. s
B. p
C. d
D. f
3. How many unpaired electrons are there in the electron configuration d^5 ?

A. 0
B. 1
C. 2
D. 5
4. Which element has the electron configuration $s^2 2s^2 2p^6 3s^2 3p^4 4s^2 3d^8$?

A. Iron (Fe)
B. Zinc (Zn)
C. Nickel (Ni)
D. Copper (Cu)
5. What is the electron configuration of a chlorine ion (Cl^-)?

A. $1s^2 2s^2 2p^6 3s^2 3p^6$
B. $1s^2 2s^2 2p^6 3s^2 3p^5$
C. $1s^2 2s^2 2p^6 3s^2 3p^7$
D. $1s^2 2s^2 2p^6 3s^2 3p^8$
6. As we move away from nucleus the distance between the adjacent orbits goes on:

A. Increasing
B. Decreasing
C. Remains same
D. May increase or decrease

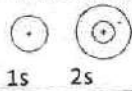
CHEMISTRY NOTES FOR CLASS - 11 (FOR PUNJAB)

- Chapter 2 // ATOMIC STRUCTURES
7. Which of the following has the same number of electrons as that of α -particles:
 A. Li^+ B. H^+ C. He^+ D. Be^{+2}
 8. Which of the following orders with respect to relative energy is incorrect:
 A. $2s < 2p$ B. $2p < 3p$ C. $3p < 3d$ D. $4d < 3s$
 9. After filling the 4s orbital the entering electron goes into:
 A. 4p B. 5s C. 3d D. 4f
 10. Which of the following particles contains 20 neutrons, 19 protons and 18-electrons?
 A. $^{39}_{19}\text{K}^+$ B. $^{39}_{19}\text{K}$ C. $^{40}_{18}\text{Ar}$ D. $^{39}_{20}\text{Ca}$
 11. Which of the following has maximum number of unpaired electrons?
 A. Fe^{+3} B. Zn^{+2} C. Ni^{+2} D. Cu^+
 12. According to $(n + l)$ rule. Which one of the following has the highest energy?
 A. 2s B. 3s C. 3p D. 2p
 13. How many unpaired electrons are in the 3d sub-shells of an element with an $Z = 24$?
 A. 2 B. 3 C. 4 D. 5
 14. With increase in the value of principal quantum number 'n' the shapes of the s-orbital remain same although their sizes.
 A. Increase B. Decrease
 C. Remain the same D. May or may remain the same
 15. The correct order of sub-shells, within a shell w.r.t decreasing energy is:
 A. $f > d > s > p$ B. $f > d > p > s$ C. $s > p > d > f$ D. $p > s > f > d$

Answer Key with Explanations

Sr.No.	Option	Explanations
1.	C	The electron configuration $1s^2 2s^2 2p^6$ corresponds to a total of 10 electrons, which is the electron configuration of Neon (Ne), a noble gas.
2.	C	The d subshell can hold a maximum of 10 electrons, as it has five orbitals, with each orbital capable of holding two electrons.
3.	D	The electron configuration $3d^5$ indicates five electrons in the d subshell. According to Hund's rule, each of these five electrons will occupy a separate orbital in the d subshell, resulting in five unpaired electrons.
4.	C	The electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$ corresponds to Nickel (Ni). The total 28 number of electrons, which shows the atomic number of Nickel.
5.	A	A chlorine ion (Cl^-) has gained one extra electron compared to the neutral chlorine atom, resulting in a full p orbital. The electron configuration of Cl^- thus becomes $1s^2 2s^2 2p^6 3s^2 3p^6$.
6.	A	Distance between shell can be calculated by $0.529n^2\text{Å}$ $r_1 = 0.529\text{Å}$ $r_2 = 2.11\text{Å}$ so Distance $r_3 = 4.761\text{Å}$ b/w consecutive $r_4 = 8.4\text{Å}$ shell is increasing continuously $r_2 - r_1 < r_3 - r_2 < r_4 - r_3$
7.	B	α -particle similarly H^+ also having no e^- .
8.	D	According to $n + l$ rule $4d < 3s$ $4 + 2 < 3 + 0$ $6 < 3$ Incorrect order
9.	C	According to $n + l$ rule after $4s \Rightarrow 4 + 0 = 4$ Entering e^- goes into $3d \Rightarrow 3 + 2 = 5$
10.	A	Among the following potassium contain 20 neutrons, 19 protons and 18-electrons.

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11.	A	$\text{Fe}^{+3} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5, 4s^0$ – 5 unpaired e in 3d $\text{Ni}^{+2} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^8, 4s^0$ – 2 unpaired e in 3d $\text{Cu}^{+1} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^0$ – no unpaired e in 3d $\text{Zn}^{+2} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^0$ – no unpaired e in 3d
12.	C	$2s = 2 + 0 = 2$ $3s = 3 + 0 = 3$ $3p = 3 + 1 = 4$ $2p = 2 + 1 = 3$ Highest $(n + \ell)$ value is for 3P
13.	D	$Z = 24$ is Cr, $\text{Cr}_{24} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5, 4s^1$
14.	A	With the increase n value (principal quantum number), the size of s-orbital increases where as the shape remains the same. e.g. the size of 2s-orbital is greater than 1s-orbital. 
15.	B	The correct order of sub-shells, within a shell w.r.t decreasing energy is $f > d > p > s$.

ADDITIONAL SHORT ANSWER QUESTIONS

Q.1 According to Moseley, the atomic number Z was a fundamental property of an element. How?

Ans. In 1913, Moseley observed that when different elements were bombarded with cathode rays, the X-rays of some characteristic frequencies were produced. It was found that the square root of the frequency of the X-rays was directly proportional to the atomic number of an element Z.

$$\sqrt{\text{frequency}} \quad (\nu) \propto Z$$

He concluded that this number, i.e. the atomic number Z was a fundamental property of an element.

Q.2 What is effect of electric field on the fundamental particles?

Ans. The behaviour of particles in an electric field depends upon their mass and charge. They show their behaviour as follows.

- Neutrons being neutral are not deflected but travel in a straight path perpendicular to the direction of electric field.
- Protons being positively charged are deflected towards the negative plate.
- Electrons being negatively charged are deflected towards the positive plate.

Q.3 Differentiate between atomic spectra and atomic absorption spectrum.

Ans.	Atomic emission spectrum	Atomic absorption spectrum
	When an element in its 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.	When a beam of white light is passed through a gaseous sample of an element in cold state, certain wavelengths are absorbed. The wavelengths of the white light that has been absorbed by the atoms show up as dark lines on the spectrum. The spectrum of this radiation is called an atomic absorption spectrum.

Q.4 What is the trend of ionization energy in the periodic table?

Ans. • As we go down a particular group, 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.

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- The ionization energies generally increase on going across a period. 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.

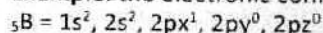
Q.5 Differentiate between shell and subshell.

Ans.	Shell	Subshell
	- The collection of orbitals with the same values of n is called an electron shell - It is denoted by n.	- The set of orbitals that have the same n and l values is called a subshell. - They are denoted by s, p, d and f.

Q.6 Define electronic configuration and give example.

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

Example: the electronic configuration beryllium is given below,



Q.7 State Hund's rule and Pual's exclusion principle.

Ans. **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.

Pual'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".

Q.8 Write the classification of the periodic table on the basis of electronic configuration.

Ans. The classification of elements of periodic table has been done into metals, non-metals, representative, transition elements, periods, groups and blocks (s, p, d, f). This division is done on the basis of electronic configuration and their valence electrons. The chemical properties of various categories of elements discussed above can be easily assessed.

Q.9 Write the electronic configuration of ions with examples.

Ans. **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 $1s^2 2s^2 2p^6$.

Negative ions are formed when atoms gain electrons. The sulfide ion, S^{2-} (proton number = 16), has 18 electrons. Its electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^6$, which is the same as argon, the element with 18 electrons in each atom.

Q.10 What is the trend of ionization energy in the periodic table?

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

An example of a simple free radical is free chlorine atom $\cdot\text{Cl}\cdot$. 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 $\text{Cl}\cdot$.

Q.11 What is the trend of ionization energy in the periodic table?

Ans.	Property	p-type Semiconductor	n-type Semiconductor
	Doping Element	Trivalent (like, Boron, Aluminium).	Pentavalent (like, Phosphorus, Arsenic).
	Charge Carriers	Holes	Electrons
	Type of Conduction	Mainly due to holes.	Mainly due to electrons.
	Electrons in Conduction Band	Few	Many
	Effect of Doping	It creates more holes.	Increases free electrons
	Example	Positive side of a diode (anode).	Negative side of a diode (cathode).

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SELF-ASSESSMENT

Chapter # 02

Total Mark: 30

Q.1 Encircle the correct option.

(1 × 6 = 6)

- (i) When 3d subshell is completely filled, the next entering electron goes into:
A. 4f B. 4s C. 4p D. 4d
- (ii) What is the shape of a p-orbital?
A. Spherical B. Dumbbell C. Circular D. Vertical
- (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) Which principle states that electrons fill the lowest energy orbitals first?
A. Pauli's exclusion principle B. Hund's rule
C. Aufbau principle D. Heisenberg principle
- (v) 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
- (vi) The number of protons, electrons, and neutrons in Na^+ are:
A. 11p, 11e, 12n B. 11p, 10e, 12n C. 11p, 12e, 11n D. 10p, 11e, 12n

Q.2 Write short answers of the following questions.

(2 × 8 = 16)

- (i) Size of Mg is bigger than Al, but ionization energy of Mg is more than that of Al. Why?
- (ii) Define atomic number and how is it related to proton number?
- (iii) How does ionic radius change compared to atomic radius in metals and nonmetals?
- (iv) Define principal quantum number and orbital.
- (v) Write the electronic configuration of Mg ($Z = 12$) using sub-shell notation.
- (vi) Define degenerate orbitals.
- (vii) Apply Hund's rule to explain why nitrogen has three unpaired electrons.
- (viii) Write the behavior of protons and electrons in an electric field.

Q.3 Extensive Questions.

(2 × 4 = 8)

- (a) What are quantum numbers? Describe briefly principal and spin quantum numbers.
- (b) 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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