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Atomic Structure

Atom: An atom is the smallest particle of an element that participates in chemical combinations i.e. in chemical bond formations.

Sub-atomic Particles of Atom: An atom consists of several elementary particles such as electron, proton, neutron, positron, meson, neutrino, antineutrino etc. These elementary particles are called sub-atomic particles.

Fundamental particles of atom: The total mass and total charge of the atom depends upon three elementary particles electron, proton and neutron. These three particles are the major component of any atom and are called as fundamental particles of atom.

Characteristics of fundamental particles of atom: The characteristics of

fundamental particles of atom can be summarised as below.

	Electron	Proton	Neutron
Discoverer	J.J. Thomson	Eugen Goldstein	James Chadwick
Symbol	$e, -1e^0$	$P, 1P^1, H^+$	$n, 0n^1$
Location	In orbits outside the nucleus	In the nucleus	In the nucleus
Relative charge	-1 unit	+1 unit	Charge less or neutral
absolute charge	-4.8×10^{-10} esu (CGS) -1.6×10^{-19} C (SI)	$+4.8 \times 10^{-10}$ esu (CGS) $+1.6 \times 10^{-19}$ C (SI)	0
Relative mass	1/1837 amu or 0.00055 amu	1 amu, or 1.00728 amu	1 amu, or 1.00866 amu
Absolute mass	9.109×10^{-28} g or 9.109×10^{-31} kg	1.673×10^{-24} g or 1.673×10^{-27} kg	1.673×10^{-24} g or 1.673×10^{-27} kg

Atomic Number (Z) : The total number of protons present in the nucleus of an atom is called atomic number and is designated by Z. i.e.

Atomic Number = No. of protons in the nucleus

or, $Z = P$

For a neutral atom,

No. of protons = No. of protons in the nucleus = No. of electrons

or. $\boxed{Z = P = e}$

Mass Number (A) : The total number of protons and neutrons present in the nucleus of an atom is called as mass number. In other words the total number of nucleons is called mass number. Mass number is always a whole number and is different from atomic mass. It is designated by 'A' i.e.

mass No. = Total number of ~~electrons~~ nucleons

or mass No. = (Total no. of protons + Total no. of neutrons)

$$A = p + n$$

also, Total No. of protons = Atomic No. (Z)

$$\therefore \boxed{A = Z + n}$$

Representation of an Atom : An atom is generally represented as $^{A}_{Z}X$ or $_{Z}X^{A}$.

where,

Z = Atomic Number = Total no. of protons

A = mass Number = Total no. of protons and neutrons

X = Symbol of the element.

Thus, carbon atom having atomic number 6 and mass number 12 can be represented as $^{12}_{6}C$ or $_{6}C^{12}$.

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Isotope : Isotopes are the atoms of same elements having same atomic number but different mass number. In other words, two or more atoms of same element having same number of protons but different number of neutrons are called isotopes. For eg:-

(1) Isotopes of Hydrogen

Name of isotopes	Ordinary Hydrogen or Protium	Heavy hydrogen or Deuterium	Radioactive Hydrogen or Tritium
Symbol	$_1H^1$	$_1H^2$ or $_1D^2$	$_1H^3$ or $_1T^3$
At. No. (Z)	1	1	1
Mass No. (A)	1	2	3
No. of Protons (p)	1	1	1
No. of neutrons (n)	0	1	2
No. of electrons (e)	1	1	1

(2) Isotopes of Carbon

Name of Isotope ~~Symbol~~	Ordinary Carbon	Radioactive Carbon
Symbol	$_6C^{12}$	$_6C^{14}$
Atomic no. (Z)	6	6
mass no. (A)	12	14
No. of protons (p)	6	6
No. of neutrons (n)	6	8
No. of electrons (e)	6	6

Characteristics of Isotopes

Following are the important characteristics of isotopes:

1. Isotopes are the atoms of same element.
2. They are placed at the same place in the periodic table.
3. Isotopes have same electronic configuration.

4. They have same number of valence electrons.

5. Because of same number of electrons they have identical chemical properties.

6. They have different physical properties such as melting point, boiling point, density etc.

<u>Isobar</u> : Isobars are the atoms of different elements having same mass number but different atomic number. In other words, two or more atoms of different element having same number of nucleons but different number of protons are called isobars. For eg:-

1. $_{19}K^{40}$, $_{20}Ca^{40}$

Symbol	$_{19}K^{40}$	$_{20}Ca^{40}$
At. No. (Z)	19	20
Mass No. (A)	40	40
No. of protons (p)	19	20
No. of neutrons (n)	21	20
No. of electrons (e)	19	20

2. $_6C^{14}$, $_7N^{14}$

Symbol	$_6C^{14}$	$_7N^{14}$
Atomic No. (Z)	6	7
Mass No. (A)	14	14
No. of Protons (P)	6	7
No. of neutrons (n)	8	7
No. of electrons (e)	6	7

Characteristics of Isobars

Following are the important characteristics of isobars:

1. Isobars are the atoms of different elements.
2. They are placed at the different places in the periodic table.
3. They have different electronic configuration.
4. They have different number of valence electrons.
5. Isobars have entirely different physical and chemical properties.

Difference between Isotopes and Isobars

The difference between Istopes and Isobars are as follows :

Isotopes	Isobars
1. Two or more atoms having same atomic number but different mass number are called Istopes.	Two or more atoms having same mass number but different atomic number are called isobars.
2. Istopes are the atoms of same element.	Isobars are the atoms of different elements.
3. Isotopes occupy same place in periodic table.	Isobars occupy different places in periodic table.
4. They have same valence electron	They have different valence electrons.
5. They have same electronic configuration.	They have different electronic configuration.

6.	They have identical chemical properties.	They have entirely different chemical properties.

10.03.21 **Planck's Quantum Theory**

Max Planck in 1900 forwared a theory regarding black body radiations which later amended by Einstein. The amended theory are following :

1. Black body radiates energy discontinuosly.

2. The emission or absorption of energy takes place in the form of small packets.

3. These small energy packets are called 'quanta'.

4. Each quantum is associated with a fixed amount of energy.

5. The energy associated with one

quantum is given as :

$$E = h\nu$$

where, E = Energy associated with one quantum
ν = Frequency of radiation
h = Planck's constant
$= 6.625 \times 10^{-27}$ erg-sec
$= 6.625 \times 10^{-34}$ J.sec

From wave motion we have,

$$v = f\lambda$$

where v = Velocity of the wave
f = Frequency of the wave
λ = wave length

For electromagnetic radions we can have,

$$c = \nu\lambda$$

or $$\nu = \frac{c}{\lambda}$$

$$\therefore \quad E = \frac{hc}{\lambda}$$

Also, the energy associated with 'n' quanta can be given as:

$$E = nh\nu$$

$$E = \frac{nhc}{\lambda}$$

Bohr's Atomic model

In order to overcome the limitations of Rutherford's atomic model, Bohr forwarded his model of atom in 1913 making use of planck's quantum theary. According to this model of atom

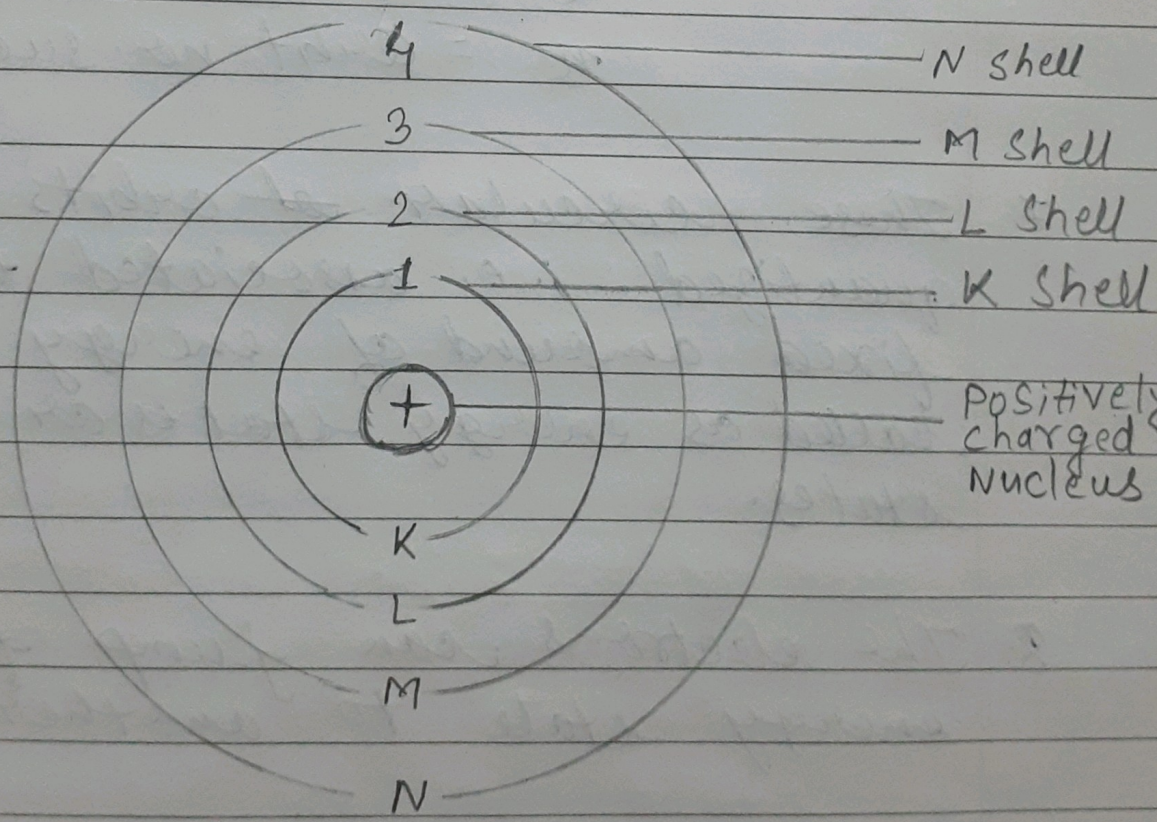
1. Electron revolve around the nucleus in certain selected fixed circular orbits only.

2. During revolution in these fixed orbits, electrons neither emit nor absorb energy.

3. The circular orbits are called stationary states.

4. Stationary does not mean that the electrons are stationary rather the orbits are stationary

5. These circular orbits are represented by K, L, M, N ---- or by numerals 1, 2, 3, 4



6. Only those orbits are permissable for which the angular momentum of the electron is an integral multiple of $h/2\pi$.

i.e, $mvr = \frac{nh}{2\pi}$

where, m = mass of electron
v = Velocity of electron
r = Radius of orbit
h = Planck's constant
n = Orbit no. such as 1, 2, 3, 4..

7. These circular ~~of~~ orbits are quantized i.e. associated with a ~~fixed~~ fixed amount of energy and are called as energy states or quantum states.

8. The electrons can jump from one energy state to another.

9. During these jumping energy is either released or absorbed.

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10. The amount of energy absorbed or released during these jumping can be given from Planck's quantum theory as:

$$E_2 \sim E_1 = h\nu$$

where, E_2 = Energy of orbit where the electron jumps.

E_1 = Energy of orbit from where electron jumps.

h = Planck's constant

ν = Frequency of radiation

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Limitations of Bohr's Atomic Model

The important limitations of Bohr's atomic model are following:

1. The Bohr's model of atom is applicable to the atoms or ions having one electron only i.e. the model is applicable to hydrogen atom and hydrogen like ions.

2. The model fails to explain the spectra of multi-electron systems.

3. When atomic spectra is seen under high resolving power instruments one thick line was found to be sub-divided into several fine lines. The model fails to explain this fact.

4. The model also fails to explain the Stark effect and Zeeman effect.

5. The model also could not explain the de' Broglie wave-matter duality and Heisenberg's uncertainty principle.

<u>Stark Effect</u>: The splitting of atomic spectra into fine lines by the application of strong electric field is called stark effect.

Zeeman Effect: The splitting of atomic spectra into fine lines by the application of strong magnetic field is called Zeeman effect.

- **Quantum numbers:** Quantum numbers are the pure numbers that characterise about the position, energy condition, angular momentum, orientation in space, spin etc. of the electron in an atom. In other words quantum number is the complete address of the electrons in an atom. The different types of quantum numbers are following:

i) Principal quantum number (n)
ii) Azimuthal quantum number (l)
iii) magnetic quantum number (m)
iv) Spin quantum number (s)

Principal quantum number (n): This quantum number provides idea about the orbits and is represented by n.

Following are the important features of principal quantum number.

1. This quantum number represents the orbits or energy levels to which an electron belongs.
i.e., $n = 1$ represents K-Shell
$n = 2$ represents L-Shell
$n = 3$ represents M-Shell
$n = 4$ represents N-Shell

2. This quantum number gives the maximum number of electrons that an orbit can accomodate which is given by $2n^2$. For eg:-

The maximum no. of electrons in K-Shell $= 2 \times 1^2 = 2$
The maximum no. of electrons in L-Shell $= 2 \times 2^2 = 8$
The maximum no. of electrons in M-Shell $= 2 \times 3^2 = 18$
The maximum no. of electrons in N-Shell $= 2 \times 4^2 = 32$

3. This quantum number also provides idea about the radius of an orbit.

4. This quantum number gives the information about the energy of electron in an orbit.

<u>Azimuthal quantum number (l)</u>: This quantum number provides idea about the sub-shells and is represented by l. the important features of azimuthal quantum number are following.

1. The values of azimuthal quantum number ~~ranges~~ depends upon the values of principal quantum number.

2. The values of azimuthal quantum number ranges from 0 to (n-1) i.e., $l = 0 \ldots (n-1)$, where n is the principal quantum number.

Thus, For K (n=1) shell ; $l = 0$
For L (n=2) shell ; $l = 0, 1$
For M (n=3) shell ; $l = 0, 1, 2$
For N (n=4) shell ; $l = 0, 1, 2, 3$

3. The total number of values of azimuthal quantum number is equal to the value of principal quantum number.

4. Each value of azimuthal quantum number represents a sub-shell. i.e

$l = 0$ represents s sub-shell
$l = 1$ represents p sub-shell
$l = 2$ represents d sub-shell
$l = 3$ represents f sub-shell

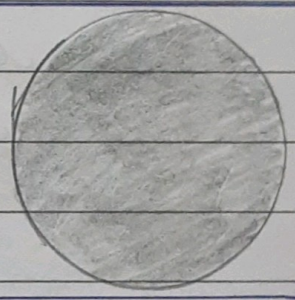
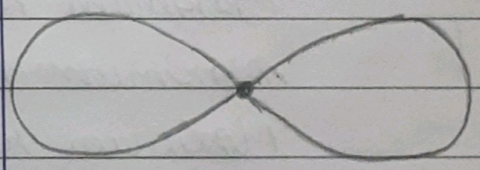
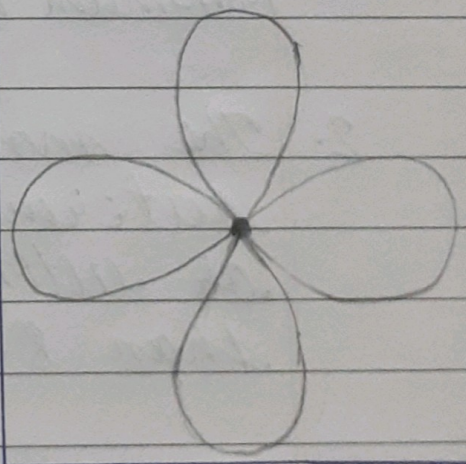
5. The total number of sub-shells in a particular shell is equal to the value of principal quantum number. i.e.

K $(n=1)$ shell has s sub-shell only
L $(n=2)$ shell has s and p sub-shells
M $(n=3)$ shell has s, p & d sub-shells
N $(n=4)$ shell has s, p, d & f sub-shells

6. The maximum number of electrons in these sub-shells is given by $(2l+1)$. i.e.

The maximum no. of electrons in s sub-shell = 2
The maximum no. of electrons in p sub-shell = 6
The maximum no. of electrons in d sub-shell = 10
The maximum no. of electrons in f sub-shell = 14

7. This quantum number also provides idea about the shapes of orbitals.

s-orbitals are spherical in shape	
p-orbitals are dumbbell in shape	
d-orbials are double dumbbell in shape	
f-orbits are complicated in shape.	---

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Bohr - Bury Scheme

Bohr and Bury together forwarded an idea about the distribution of electrons in different orbits. According to this scheme;

1. The maximum number of electrons in an orbit is given by $2n^2$ (where, n is the orbit number i.e. 1, 2, 3, 4, 5, ---) i.e.

Maximum no. of electrons in K (n=1) shell $= 2 \times 1^2 = 2$
Maximum no. of electrons in L (n=2) shell $= 2 \times 2^2 = 8$
Maximum no. of electrons in M (n=3) shell $= 2 \times 3^2 = 18$
Maximum no. of electrons in N (n=1) Shell $= 2 \times 4^2 = 32$

2. The maximum no. of electron in the outermost orbit or valence orbit or ultimate orbit will not be more than 8.

3. The maximum no. of electron in the penultimate orbit will not be more than 18.

4. The maximum no. of electron in the ante-penultimate orbit will not be more than 32.

5. This is not necessary for an orbit to be filled completely before starting to the next higher orbit. In fact electrons enter into the next higher orbit as soon as orbit acquires 8 electrons.

Pauli's Exclusion Principle

The principle states that "no two electrons in an atom can simultaneously have the same set of all the four quantum numbers". For eg:- $_2He - 1s^2$

1s

[↑↓]

$n = 1$	$n = 1$
$l = 0$	$l = 0$
$m = 0$	$m = 0$
$s = +\frac{1}{2}$	$s = -\frac{1}{2}$

It is thus seen that the three quantum numbers are same for the two electrons of He atom, whereas it differs in the spin quantum number.

Hund's Rule of Maximum Multiplicity

The rule states that electron first singly occupied in the orbitals of equal energy i.e. degenerate orbitals ~~of equal~~ and then it is paired up with opposite spin. In ~~other~~ other words, the pairing of electrons in the orbitals of equal energy i.e., degenerate orbitals takes place only when there is no availability of vacant orbitals for the incoming electron. For eg:-

p^4	↑↓	↑	↑

d^7	↑↓	↑↓	↑	↑	↑

Aufbau Principle

'Aufbau' is a German word means 'construction' or 'building up'. The principle states that the "electrons enter into the orbitals on the basis of their increasing order of energy i.e. lower energy orbitals are filled first".

The above principle can be well understood on the basis of (n+l) rule also called as Medelung rule or Janet rule. The rule states that:

1. Electron enters in that orbital first for which the value of (n+l) is minimum.

2. If the value of (n+l) is same for two or more orbitals, then electron enters in that orbital first for which the value of 'n' is minimum.

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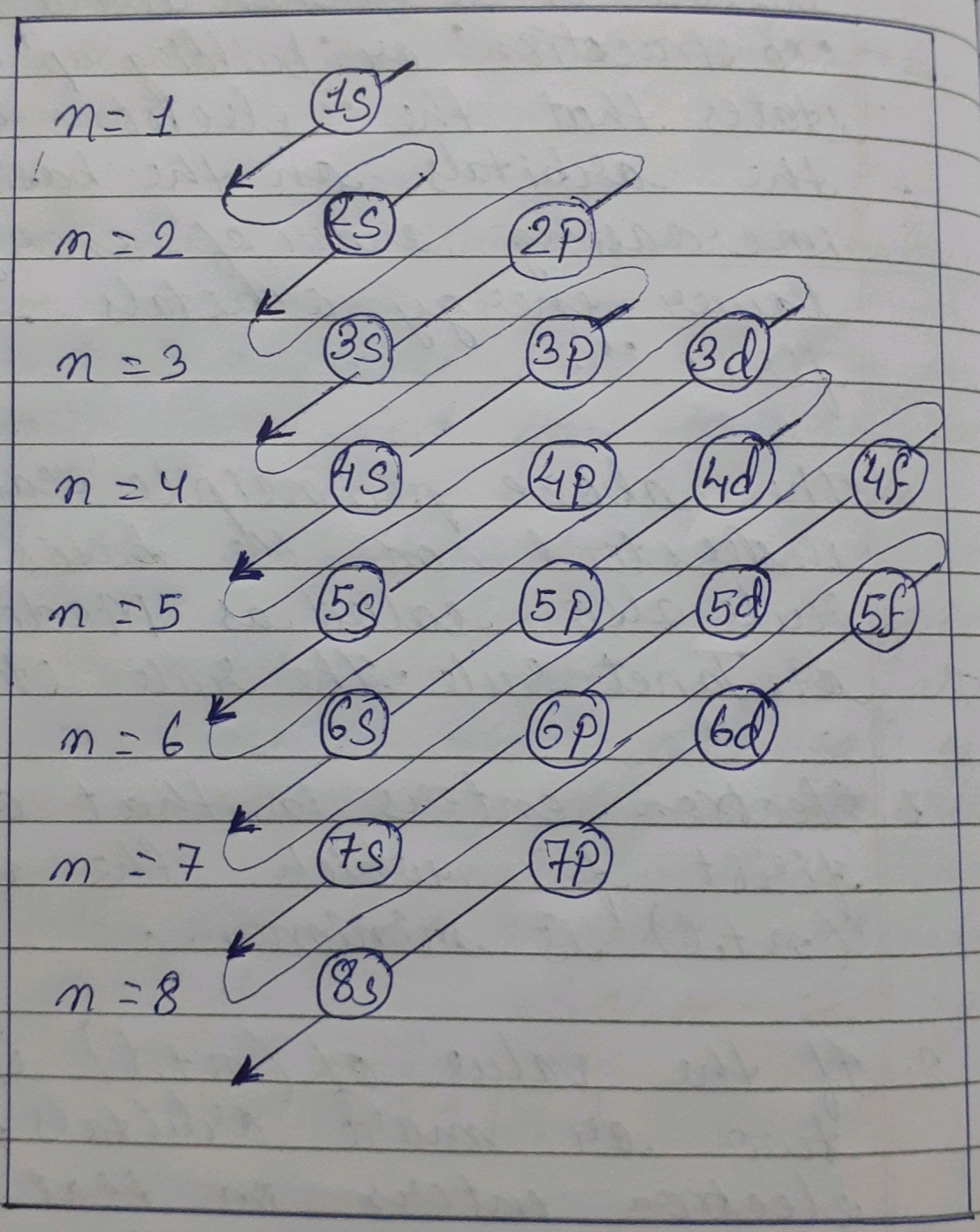
The (n+l) rule can be schematically represented as below.



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Electronic Configuration of Elements

The electronic configurations of the atoms of first thirty elements having atomic number 1-30 are given below.

Name of element	Symbol	Atomic No. (Z)	Electronic Configuration	
Hydrogen	H	1	$1s^1$	$1s^1$
Helium	He	2	$1s^2$	$1s^2$
Lithium	Li	3	$1s^2, 2s^1$	$[He], 2s^1$
Beryllium	Be	4	$1s^2, 2s^2, \cancel{2p^1}$	$[He], 2s^2$
Boron	B	5	$1s^2, 2s^2, 2p^1$	$[He], 2s^2, 2p^1$
Carbon	C	6	$1s^2, 2s^2, 2p^2$	$[He], 2s^2, 2p^2$
Nitrogen	N	7	$1s^2, 2s^2, 2p^3$	$[He], 2s^2, 2p^3$
Oxygen	O	8	$1s^2, 2s^2, 2p^4$	$[He], 2s^2, 2p^4$
Fluorine	F	9	$1s^2, 2s^2, 2p^5$	$[He], 2s^2, 2p^5$
Neon	Ne	10	$1s^2, 2s^2, 2p^6$	$[He], 2s^2, 2p^6$
Sodium	Na	11	$1s^2, 2s^2, 2p^6, 3s^1$	$[Ne], 3s^1$
Magnesium	Mg	12	$1s^2, 2s^2, 2p^6, 3s^2$	$[Ne], 3s^2$
Aluminium	Al	13	$1s^2, 2s^2, 2p^6, 3s^2, 3p^1$	$[Ne], 3s^2, 3p^1$
Silicon	Si	14	$1s^2, 2s^2, 2p^6, 3s^2, 3p^2$	$[Ne], 3s^2, 3p^2$
Phosphorus	P	15	$1s^2, 2s^2, 2p^6, 3s^2, 3p^3$	$[Ne], 3s^2, 3p^3$

Sulphur	S	16	$1s^2, 2s^2, 2p^6, 3s^2, 3p^4$	$[Ne], 3s^2, 3p^4$
Chlorine	Cl	17	$1s^2, 2s^2, 2p^6, 3s^2, 3p^5$	$[Ne], 3s^2, 3p^5$
Argon	Ar	18	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6$	$[Ne], 3s^2, 3p^6$
Potassium	K	19	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$	$[Ar], 4s^1$
Calcium	Ca	20	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$	$[Ar], 4s^2$
Scandium	Sc	21	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^1$	$[Ar], 4s^2, 3d^1$
Titanium	Ti	22	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^2$	$[Ar], 4s^2, 3d^2$
Vanadium	V	23	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^3$	$[Ar], 4s^2, 3d^3$
Chromium	Cr	24	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^4$	$[Ar], 4s^1, 3d^5$
Manganese	Mn	25	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^5$	$[Ar], 4s^2, 3d^5$
Iron	Fe	26	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^6$	$[Ar], 4s^2, 3d^6$
Cobalt	Co	27	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^7$	$[Ar], 4s^2, 3d^7$
Nickel	Ni	28	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^8$	$[Ar], 4s^2, 3d^8$
Copper	Cu	29	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^9$	$[Ar], 4s^1, 3d^9$
Zinc	Zn	30	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}$	$[Ar], 4s^2, 3d^{10}$

Electronic Configuration of Ions

The electronic configuration of some of the cations and anions are given below.

Ions	No. of Protons in the Ion	No. of electrons in the Ion	Electronic Configuration
Al^{3+}	13	10	$1s^2, 2s^2, 2p^6, 3s^0, 3p^0,$
K^{+}	19	18	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0,$
Ca^{++}	20	18	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0,$
Sc^{3+}	21	18	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^0$
Cr^{3+}	24	21	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^3$
Mn^{2+}	25	23	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^5$
Fe^{2+}	26	24	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^6$
Fe^{3+}	26	23	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^5$
Co^{2+}	27	25	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^7$
Ni^{2+}	28	26	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^8$
Cu^{+}	29	28	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^{10}$
Cu^{2+}	29	27	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^9$
Zn^{2+}	30	28	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^0, 3d^{10}$
N^{3-}	7	10	$1s^2, 2s^2, 2p^6,$
O^{2-}	8	10	$1s^2, 2s^2, 2p^6,$
F^{-}	9	10	$1s^2, 2s^2, 2p^6,$
Cl^{-}	17	18	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6,$
S^{2-}	16	18	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6,$
P^{3-}	15	18	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6,$

Electronic Configuration and Number of Unpaired Electrons

The electrical, magnetic and dielectric properties etc. of the atoms and ions deepens upon the electronic arrangement. The number of unpaired electrons in the atoms of first thirty elements and their magnetic properties are given below.

Element	At No. (Z)	Electronic Configuration	Unpaired electrons	Magnetic Nature
H	1	1s: [↑]	1	Paramagnetic
He	2	1s: [↑↓]	0	Diamagnetic
Li	3	1s: [↑↓] 2s: [↑]	1	Paramagnetic
Be	4	1s: [↑↓] 2s: [↑↓]	0	Diamagnetic
B	5	1s: [↑↓] 2s: [↑↓] 2p: [↑] [ ] [ ]	1	Paramagnetic
C	6	1s: [↑↓] 2s: [↑↓] 2p: [↑] [↑] [ ]	2	Paramagnetic
N	7	1s: [↑↓] 2s: [↑↓] 2p: [↑] [↑] [↑]	3	Paramagnetic
O	8	1s: [↑↓] 2s: [↑↓] 2p: [↑↓] [↑] [↑]	2	Paramagnetic
F	9	1s: [↑↓] 2s: [↑↓] 2p: [↑↓] [↑↓] [↑]	1	Paramagnetic
Ne	10	1s: [↑↓] 2s: [↑↓] 2p: [↑↓] [↑↓] [↑↓]	0	Diamagnetic

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Na	11	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑ 3s: ↑	1	Paramagnetic
Mg	12	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓	0	Diamagnetic
Al	13	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑	1	Paramagnetic
Si	14	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑ ↑	2	Paramagnetic
P	15	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑ ↑ ↑	3	Paramagnetic
S	16	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑ ↑	2	Paramagnetic
Cl	17	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑	1	Paramagnetic
Ar	18	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑↓	0	Diamagnetic
K	19	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑↓ 4s: ↑	1	Paramagnetic
Ca	20	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑↓ 4s: ↑↓	0	Diamagnetic
Sc	21	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑↓ 4s: ↑↓ 3d: ↑	1	Paramagnetic
Ti	22	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑↓ 4s: ↑↓ 3d: ↑ ↑	2	Paramagnetic
V	23	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑↓ 4s: ↑↓ 3d: ↑ ↑ ↑	3	Paramagnetic
Cr	24	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑↓ 4s: ↑↓ 3d: ↑ ↑ ↑ ↑ ↑	6	Paramagnetic
Mn	25	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑↓ 4s: ↑↓ 3d: ↑ ↑ ↑ ↑ ↑	5	Paramagnetic
Fe	26	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑↓ 4s: ↑↓ 3d: ↑↓ ↑ ↑ ↑ ↑	4	Paramagnetic
Co	27	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓ 3s: ↑↓ 3p: ↑↓ ↑↓ ↑↓ 4s: ↑↓ 3d: ↑↓ ↑↓ ↑ ↑ ↑	3	Paramagnetic

Ni	28	1s [↑↓] 2s [↑↓] 2p [↑↓ ↑↓ ↑↓] 3s [↑↓] 3p [↑↓ ↑↓ ↑↓] 4s [↑↓] 3d [↑↓ ↑↓ ↑↓ ↑ ↑]	2	Paramagnetic
Cu	29	1s [↑↓] 2s [↑↓] 2p [↑↓ ↑↓ ↑↓] 3s [↑↓] 3p [↑↓ ↑↓ ↑↓] 4s [↑↓] 3d [↑↓ ↑↓ ↑↓ ↑↓ ↑↓]	1	Paramagnetic
Zn	30	1s [↑↓] 2s [↑↓] 2p [↑↓ ↑↓ ↑↓] 3s [↑↓] 3p [↑↓ ↑↓ ↑↓] 4s [↑↓] 3d [↑↓ ↑↓ ↑↓ ↑↓ ↑↓]	0	Diamagnetic

IMPORTANT QUESTIONS

1. Define atom and write down the important characteristics of fundamental particles of the atom.