c. Both the metals burn in O_2 and form oxides (BeO and Al_2O_3). Both these oxides are amphoteric in nature, i.e., they react with both acids and alkalies to form the salts.

$$BeO + 2HCl \longrightarrow BeCl_2 + H_2O_2$$

$$BeO + 2NaOH \longrightarrow Na_2BeO_2 + 2H_2O$$

$$Al_2O_3 + 6HCl \longrightarrow 4AlCl_3 + 3H_2O$$

$$Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$$

5. Formation of complexes by Be^{+2} and Al^{+3} ions. Both these ions have a strong tendency to combine with ligands and form chelated complexes like

$[Be(H_2O)_4]^{+2}$ $[BeF_4]^{-2}$ $[Be(C_2O_4)_2]^{-2}$

$[Al(H_2O)_6]^{+3}$ $[AlF_6]^{-3}$ $[Al(C_2O_4)_3]^{-3}$ etc.,

The diagonal relationship between Be and Al is due to the reasons :

i. The ionic potentials of Be^{+2} and Al^{+3} ions are almost the same.

ii. The values of standard reduction potentials are almost the same.

iii. The electronegativity values of both the elements are the same.

iv. Both the metals have weakly electropositive character.

v. The covalent radii of both the elements are almost the same (Be = 125 pm, Al = 130 pm).

1.2. Periodic Properties

According to the modern periodic law, the properties of elements are periodic functions of their atomic numbers. This means that when elements are arranged in the increasing order of their atomic numbers there is periodicity in their properties, i.e.: *their physical and chemical properties are repeated at regular intervals. This is referred to as periodicity in property.*

X am

If we analyse the electronic configurations of elements, at regular intervals, elements have similar electronic configurations.

These facts led to the conclusion that periodicity in properties of elements cloud be explained on the basis of their electronic structure.

Elements having similar electronic configurations have similar properties. This is shown below by studying the following properties.

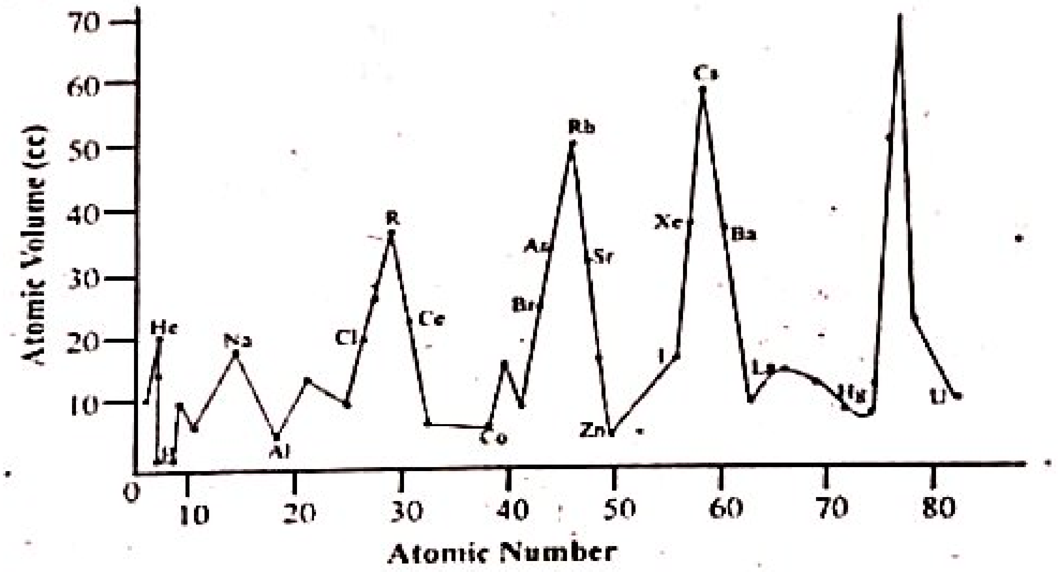
1. Atomic Volume

One year after the publication of Mendeleeff's periodic table. Lother Meyer published a curve showing the relation between atomic weights and the atomic volumes of elements.

Definition : (*The atomic volume is the ratio between atomic weight and specific gravity of the element in solid state.*)

Variation along period and group

The atomic volume curve contains a series of well defined loops. Similar elements occupy, analogous positions. The ascending slopes of these loops contain the electroronegative elements and the decending slopes contain the electoropositive elements. Alkali metals occupy the top positions in the curve.



The atomic volume increases when we go from top to bottom in a group of the periodic table. The members of each period of the periodic table occupy successive positions on the curve. The members of each group occupy similar positions. For example, the elements of the alkali metal group occupy the top points in the curve.

2m * Repeation of Similar proper
called periodisity.

5m 2. Factors affecting atomic volume

i. *Nuclear charge :*

As we move from left to right in a period, the nuclear charge (atomic number) increase by one. This increase in nuclear charge attracts the electronic cloud towards the nucleus. This results in a decrease in the atomic volume.

ii. *Number of valency electrons :*

As the numberof valency electrons increases atomic volume also increases to accommodate the electrons.

These two factors affect the atomic volume in opposite directions. As a result, along the period atomic volume decreases, reaches a minimum and then increases.

iii. *Number of energy levels :*

As the number of energy levels increases the atomic volume also increases.

4Q 2. Atomic radii அணுஆரம்

***Definition :** The radius of an atom may be taken as the distance between the centre of the nucleus and the outer most shell of electrons.* 5m 4Q

Variation along a period and group

There is periodic variation of atomic radii with atomic number. Generally the atomic radii decrease in moving from left to right in any given period. The atomic radii increases in moving from top to bottom in any given group.

Alkali metals has f

The variation of atomic radii in a period can be explained on the basis of electronic configuration of elements. As an example, we shall consider the second period, beginning from lithium and ending with fluorine. The atomic number increases from 3 to 9 and hence the nuclear charge increases from 3 to 9 as we move from lithium to flourine. As the charge of the nucleus increases step by step, the K electrons are attracted more and more strongly towards and nucleus. In otherwords, there is more and more contraction of the K shell in moving from left to right in a period. Thus the atomic radii decreases progressively. The L electrons are also attracted towards the nucleus but the K electrons, lying in between, tend

why decreases

Ionic radius distance from
an influence

Variation along a period of group

The covalent radii of elements decrease from left to right in a period. Thus in a period the alkail metal has the largest covalent radius and the halogen has the smallest covalent radius. This is because, as the atomic number increase across a period the nuclear charge increases and the orbital electrons are pulled in towards the nucleus.

The covalent radii increase down a group in the periodic table. For example, in the alkali metals, the covalent radii increase as we go down from Li to Cs. This is because as we go down a group, extra shell is added.

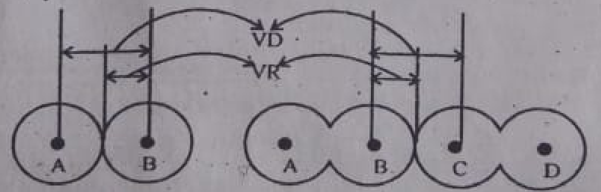
Vander Waals Radii :

When two non - bonded isolated atoms or two non bonded atoms belonging to two adjacent molecules of an element in the solid state approach each other without overlapping their electron clouds to form a bond between them, they cannot come closer than a minimum distance without forming a bond. This minimum distance is called van der waals distance and half this distance is called van der waals radius.

Definition :

The distance between the nuclei of two nonbonded isolated atoms or the distance between two non - bonded atoms belonging to two adjacent molecules of an element in the solid state is called van der waals distance and half this distance is called van der waals radius.

Illustration



Two non - bonded chlorine atoms

Two chlorine molecules

VD = van der waals distance
VR = van der waals radius

properties after a regular interval is 5

to screen them from the nucleus. In spite of the screening effect, the L electrons also experience greater and greater attraction towards the nucleus.

The atomic radii increase as we move from top to bottom in a group. This is because as we move down in a group, the number of shells increases and therefore the size of the atom increases. why increases

Drawbacks in the definition of atomic radii :

The definition of atomic radii as given above is not precise for the following reasons.

i. According to quantum meachanics, there is no certainly with regard to the exact position of the electron. That is, it is impossible to ascertain the boundary of an atom.

ii. We cannot have isolated atom.

Because of these drawbacks we have three operational concepts of atomic radii, viz.,

i. Covalent radii
ii. Van der Waals radii and
iii. Ionic radii.

Covalent radii :

Definition : *Covalent radius is defined as half the distance of separation of nuclei in a covalently bonded molecule.*

The distance between the nuclei of any two neighbouring atoms is called inter nuclear distance.

E.g.,

Molecule	Internuclear distance in A°	Atomic Radius d/2 in A°
H – H	0.74	0.37
Cl – Cl	1.98	0.99
Br – Br	2.28	1.14

For molecules made up of different atoms. e.g., HCl, the internuclear distance will be equal to the sum of the covalent radii of H and Cl.

6

Variation along a period and group

There is a periodic variation of ionisation potential with atomic number. When we move from left to right in a period, the ionisation potential progressively increases. This tendency may be attributed to a rising nuclear charge as the atomic number is increased. The ionisation potential decreases from top to bottom in a group. This is due to the progressive increase in the size of the atom. As the size of the atom increases the outermost electrons are held less firmly. So they can easily be removed. Thus the ionisation potential decreases as we move down the group.

Factors affecting ionisation potential

1. Effective nuclear charge :

Greater the magnitude of effective nuclear charge, higher is the ionisation potential needed to remove the outer most shell electron. Thus on moving from left to right in a period, the effective nuclear charge increases. So ionisation potential also increases from left to right.

2. State of the atom (r) :

Higher the size of an atom (r) lower will be its ionisation energy. Thus as we move down a group the value of r, the atomic radius increases. So the ionisation potential decreases as we move down the group. This is because, as the r of an atom increases, the distance between the nucleus and the electron in the outer most shell increases. So the forces of attraction exerted by the nucleus on them will decrease. So they may be easily removed.

3. Principal quantum number :

The magnitude of the ionisation potential decreases with an increase in the value of the principal quantum number (n) of the orbital from which the electron is to be removed. This factor is also used to explain the variation of ionisation potential of elements in a group. As we move down a group, the value of the principal quantum number (n) of the orbital from which the electron is to be removed increases. So the ionisation potential decreases. For example, the ionisation potential of Group IIA elements decreases in the order.

Be > Mg > Ca > Sr > Ba

ucleus of an ion upto which it [illegible]
to its Electron cloud.

This if we say that the van der waals radius of chlorine is 1.8A° it means that the two chlorine atoms belonging to two different Cl_2 molecules cannot approach each other nearer than 2 x 1.8 = 3.6A° from each otehr without overlapping of their electron clouds to form a bond.

For noble gases in the solid state the crystal radii. i.e., the atomic radii are same as their van dar waals radii.

Ionic radii : atomic radius Increase ionic radius decrease

AQ (Ionic raidus is the distance from the centre of the nucleus of an ion to the edge of its electron cloud.) AQ A cation (positive ion) results from the loss of one or more electrons from an atom. This generally causes the removal of the whole of the outer shell of electrons. Thus in the formation of cations, the outer shell of electrons is removed completely. Therefore the cation is much smaller than the corresponding atom. With the elimination of one or more orbital electrons, the effective nuclear charge increases. So the electrons are pulled in, more inwards the nucleus than before. This effect tends to decrease the radius of the cation.

E.g, $K^+ < K$ gain cation(+), Anion(-) → donate

An anion (negative ion) is formed by the addition of one or more electrons to the atom. The effective nuclear charge is reduced. Therefore the electron cloud expands. As a result of this the radius of a negative ion is larger than the corresponding atom.

E.g, $Cl < Cl^-$.

Variation along a period of group

In going from left to right in a period there is a pronounced and progressive decrease in ionic radii. As we move from top to bottom in a group there is a progressive increase in ionic radii.

Factors affecting atomic and ionic radii

i. ***Effective nuclear charge :***

The magnitude of effective nuclear charge determine the magnitude of force of attraction between the nucleus and the outer most electron cloud. Greater the magnitude of the effective nuclear charge, greater will be the attraction between the nucleus and smaller will be the atomic and ionic radii.

ns^2, np^3, np^6, nd^5, nd^{10} etc. configurations are more stable than elements having other configurations. So the ionisation potential of an atom having half filled or completely filled orbitals in its electronic configuration will be more than that expected normally from its position in the periodic table.

This factor has been used to explain the higher value or ionisation potential of some of the elements as compared to that of their left hand neighbour in a period of the periodic table.

E.g.
i. I.P. of Be > IP of B
ii. I.P. of N > IP of O
iii. I.P. of Mg > IP of Al
iv. I.P. of P. > IP of S

6. *Nature of orbitals :*

The relative order of ease with which the electron occupying these orbtials can be removed will be as follows :

ns < np < nd < nf

This means that it will be the easiest to remove f electron and the most difficult to remove will be s electron in a given shell i.e., the ionisation potential of electron in s, p, d and f orbitals decreases in the order.

ns > np > nd > nf

This factor can be used to explain the successive ionisation potential of an element and to explain the following :

IP of Be > IP of B

IP of Mg > IP of Al

7. *Octet arrangement :*

An atom with octet arrangement i.e., ns^2 np^6 arrangement is highly stable. Hence it is very difficult to remove an electron from this stable octet arrangement. Thus atoms with stable octet arrangement have maximum ionisation potentials. E.g., Inert gases.

⑤ **Electron Affinity :**

Definition : *It is the amount of energy that is released when an electron is added to the neutral atom.*

ii. Number of shells or energy levels :

As the total number of shells or energy levels that are present increase, the atomic and ionic radii also increase, i.e., if n, the principal quantum number increases the atomic and ionic radii also increase.

4. Ionisation potential (Ionisation energy)

The ionisation energy or ionisation potential is the most fundemental property of an atom.

Definition : *The ionisation potential is the energy required to remove an electron from an isolated gaseous atom to form a positive ion.*

$$M_{(g)} \xrightarrow{I_1} M^{+}_{(g)} + e^{-}$$

Successive ionisation potentials :

The energy required to remove the first electron is called the first ionisation potential. The energy required to remove the an electron from gaseous monopositive cation is called the second ionisation potential an so on. The second ionisation potential is usually much higher than the first.

$$M^{+}_{(g)} \xrightarrow{I_2} M^{2+}_{(g)} + e^{-} \qquad I_2 > I_1$$

This can be explained on the basis of an increase in the effective nuclear chrage of the gaseous cation. Thus when a cation is formed it has lesser number of electrons than the parent atom or cation. As the number of electrons decreases, the magnitude of the effective nuclear charge increases as we move from $M_{(g)}$ to $M^{2+}_{(g)}$. With the increase of effective nuclear charge the magnitude of the force of attraction between the nucleus and the outer shell electron cloud also increase.

$$M_{(g)} \xrightarrow{-e^{-}} M^{+}_{(g)} \xrightarrow{-e^{-}} M^{2+}_{(g)} \ldots \xrightarrow{-ne^{-}} M^{n+}_{(g)}$$

So the magnitude of the energy to remove the outermost shell electron, i.e., the ionisation potential also increases.

When we simply say ionisation potential it means the first ionisation potential.

The variation of successive ionisation potentials of an element is also explained on this basis.

E.g., $I_1 < I_2$ for Li where I_1 = first ionisation potential and I_2 = second ionisation potential.

This can be explained as follows : The electronic configuration of Li is $1s^2\ 2s^1$. During first ionisatrion of Li, 2s electron is removed. Here n = 2. In the second ionisation 1s electron is removed. Thus here the magnitude of the ionisation potential increases as there is a decrease in the value of the principal quantum number of the orbtial from which the electron is to be removed.

In a similar fashion we can explain the following order

$I_1 < I_2 << I_3$ of Mg

4. ***Shielding effect :***

In a multi electron atom, the electrons lying between the nuclues and the valence shell are called intervening electrons. These electrons shield the outer - most shell electron from the nucleus. This is called shielding or screening effect. If the shielding effect increases the ionisation potential will decrease. This shielding effect is indicated by σ the shielding constant. This factor can be used to explain the variation of ionisation potential of the elements of the given group in the periodic table. As we move down a group the number of inner shell electrons increases. So the magnitude of σ increases. So ionisation potential decreases.

This factor has also been used to explain the large decrease in the value of ionisation potential when we pass from an inert gas to the next alkali metal. E.g., When we pass from He (I.P. = 2372.3 kJ/mole) to Li (IP = 520.33 kJ/mole) the value of ionisation potential (IP) decreases to a great extent. This is explained as follows : In Li atom ($1s^2\ 2s^1$) the electron to be removed from 2s orbital is effectively shielded by the two electron in 1s orbital. In He atom ($1s^2$) the electron to be removed from the 1s orbital is not effectively shielded by the other electron in 1s orbital, as both of them are in the same 1s orbtial itself.

5. ***Half filled and completely filled orbitals :***

We know that half filled and completely filled orbitals are more stable than partially fitlled orbitals. (Hund's Rule). Thus elements, having

$X_{(g)} + e^- \longrightarrow X_{(g)} + \text{Engery}$

The energy of the above reaction is called electron affinity.

Variation along a period and group

Electron affinity increases as we proceed from left to right along a period. It decreases from top to bottom in a group. This is because as we move along a period, the atomic size decreases. So the force of attraction exerted by the nucleus on the electrons increases. Because of this, the atom has a greater tendency, to attract the newly added electron. In other words, its electron affinity increases.

On moving down a group, the atomic size increases because as we move down, new shells are added. So the nuclear attraction for the electron decreases. Because of this the atom has, less tendency to attract the newly added electron. In other words, its electron affinity decreases.

Second electron affinity :

Definition :

It is the amount of energy requried to add one more electron to a mono negative ion, in the gaseous state to convert it to a dinegative ion.

$M^-_{(g)} + e^- \longrightarrow M^{2-}_{(g)} \quad - \quad \text{Energy}$

Explanation :

In the case of first electron affinity energy is released. In the case of second electron affinity energy has to be supplied. This is because, one electron is added to $M^-_{(g)}$ anion against the electro static repulsion between extra electron, which is added to $M^-_{(g)}$ and the negative charge on $M^-_{(g)}$ anion. Thus the first electron affinity value will have a –ve sign as it is an exothermic reaction. The second electron affinity value will have a +ve sign as it is an endothermic reaction.

$E_2 > E_1$ 68

Where E_1 is the first electron affinity and E_2 is the second electron affinity. E.g, Formation of $O^{2-}_{(g)}$ from $O_{(g)}$.

$O_{(g)} + e^- \longrightarrow O^-_{(g)} ; E_1 = -141 \text{ kJ / mole}$

$O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)} ; E_2 = 921 \text{ kJ/mole}$

$O_{(g)} + 2e^- \longrightarrow O^{2-}_{(g)} ; E_1 + E_2 = 780 \text{ kJ / mole}$

the two atoms are not of the same element, there is unequal sharing of bonding electrons. The electronegativity is the ability of an atom to attract the shared electron pair towards itself.

Variation along a period and group

The electronegativity of an atom depends upon the ionisation energy and the electron affinity of the atom concerned. When the ionisation energy is more, it is difficult to draw out an electron from the atom. The electron affinity will also be greater. Thus higher ionisation energy and greater electron affinity imply greater electronegativity. The elements of Group I have lowest electronegativities and the elements of Group VII have highest values.

Moving in a period of the periodic table from left to right, the electronegativity increases progressively. It is due to the progressive increase in the nuclear charge and the progressive increase in the attractive force between the nucleus and the electrons. In moving down a group of periodic table the electronegativity decreases. When we move down a group the nuclear charge again increases. On this account we expect an increase in electronegativity. Actually it is not so. The electron shielding effect is more than the increase in the nuclear charge.

Factors affecting electronegativity

i. Size of the atom :

Smaller the size of an atom greater is its tendency to attract the bonded pair of electrons towards itself and hence more will be its electronegativity.

ii. Number of inner shells :

The atom with more number of inner shells between the nucleus and the outermost shell will have less electronegativity. This accounts for the decreasing order of the electronegativities of halogens.

iii. Oxidation state :

Elements in higher oxidation states have more electronegativity values. This is because the atom in higher oxidation state will be smaller in size and so will have greater attraction for electrons. This accounts for the following increasing trend of electronegativities.

$Li < Li^+$: $Fe < Fe^{2+} < Fe^{3+}$

$Sn < Sn^{2+} < Sn^{4+}$: $Pb < Pb^{2+} < Pb^{4+}$ etc.

Factors affecting electron affinity

i. *Atomic radius :*

As the atomic radius increases, the tendency of the nucleus of the atom to attract additional electrons towards it decreases. So large atoms have lesser electron affinities than smaller atoms.

ii. ***Effective nuclear charge :***

As the effective nuclear charge increases the tendency of the nuclues of the atom to attract additional electrons towards it increases. Thus atoms with higher nuclear charge have higher electron affinities.

iii. ***Electronic configuration :***

a. *Completely filled orbitals :*

Atoms with completely filled orbitals resist the addition of extra electrons. This accounts for almost zero electron affinity values of IIA group elements and zero group elements which have the outer electronic configuration ns^2 and ns^2p^6 respectively.

b. *Half filled atomic orbitals :*

Atoms with half filled orbitals are stable and they also resist the addition of extra electrons. This accounts for low electron affinity values of N and P which have the electronic configuration of $2s^2p^3$ and $3s^2p^3$ respectively.

c. ***Atoms with electrons one short of the number required for completing their orbitals :***

Such atoms have high electron affinities. E.g., Halogens. They have ns^2p^5 configuration. These have a very high tendency to take up one electron to attain a stable, noble gas, electronic configuration.

6. Electronegativity :

6Q ***Definition :***

The relative ability of a bonded atom in a molecule to attract the shared electron pair towards itself is known as its electronegativity. 6Q

Explanation :

When both the atoms linked by a covalent bond are identical, as in the case of Cl_2 molecule, the bonding electron pair is shared equally between the two atoms and none of them has any net charge on it. When

iv. Number and nature of the other bonded atom :

Electronegativity of an atom is a property which depends on the nature and number atoms bound to it. This is the reason why electronegativity value of an atom is not a constant. E.g., The electronegativity value of P in PCl_3 and that in PF_3 are different

v. Ionisation energy and electron affinity :

Atoms with higher ionisation energies and electron affinities will have higher electronegativity values also. Examples

	Ionisation Energy	**Electrons Affinity**	**Electro Negativity**
Halogens	More	More	More
Alkali metals	Less	Less	Less

iv. Type of hybridisation :

The mangnitude of electronegativity increases with an increase in the s character of the hybrid orbital. This is because the s orbital has the tendency to come closer to the nucleus. Thus the electronegativities of atoms with sp^3, sp^2 and sp orbitals increase in the same order as the percentage of their s character viz., $sp^3 < sp^2 < sp$.

E.g., The electronegativities of C in CH_4, C_2H_4 and C_2H_2 are in the increasing order of $CH_4 < C_2H_4 < C_2H_2$..

Compound	**Type of hybridisation**	**% of s character**
CH_4	sp^3	25%
C_2H_4	sp^2	33.33%
C_2H_2	sp	50%

1.3. Classification of elements on the basis of their electronic configurations : (s, p, d and f block elements)

1. Zero group elements :

Elements with completely filled shells are called zero group elements or inert gases. They have $ns^2\ p^6$ configuration in their ultimate shells.

$He = 2 = 1s^2$

$Ne = 10 = 1s^2\ 2s^2p^6$

$Ar = 18 = 1s^2\ 2s^2\ p^6\ 3s^3p^6$

$Kr = 36 = 1s^2\ 2s^2\ p^6\ 3s^3\ p^6d^{10}\ 4s^2p^6$

$Xe = 54 = 1s^2\ 2s^2\ p^6\ 3s^3\ p^6d^{10}\ 4s^2\ p^6d^{10}\ 5s^2p^6$

$Ro = 36 = 1s^2\ 2s^2\ p^6\ 3s^3\ p^6d^{10}\ 4s^2\ p^6d^{10}f^{14}\ 5s^2p^6d^{10}\ 6s^2p^6$

Characteristics of Zero group elements :

i. The zero group elements possess the highest ionisation potential and zero electron affinity. This is attributed to their stable electronic configuration (ns^2p^6).

ii. The monoatomic molecules of the zero group elements are held by weak ***van dar waal's*** forces. They are therefore low boiling gases.

iii. On account of their stable ns^2p^6 configuration these gases are chemically inert. Hence, they are called inert gases or noble gases.

iv. They are all diatomic and colourless gases due to completely filled electronic shells.

2. Normal Elements :

These elements have incomplete ultimate shells. They are futher classified as

University Questions

1. What are quantum numbers? 1
2. Write short note on
   a. Principle quantum number. 1
   b. Azimuthal quantum number 1
3. Write short note on
   a. Magnetic quantum number 1, 2
   b. Spin quantum number 2
4. What will be the values of other quantum numbers of an element when the principle quantum number is 2?
5. State and explain Pauli's exclusion principle. 3, 4, 5
6. State and explain Aufbau principle. 6, 7
7. State and explain Hund's Rule 5, 6
8. State and explain inert pair effect. 10, 11
9. Explain diagonal relationship between the elements in the periodic table. Explain with examples.
10. Write notes on
    a. Atomic volume 15
    b. Atomic radii 16
    c. Ionic radii 19
11. What are the factors which affect Ionisation energy?
12. Write notes on
    a. Electron affinity 23
    b. Electronegativity 25

Example : IIIA group	IVA group
B = 5 = $[He]2s^2\rho^1$	C = 6 = $[He]2s^2p^2$
Al = 13 = $[Ne]3s^2p^1$	Si = 14 = $[Ne]3s^2p^2$
Ga = 31 = $[Ar]3d^{10}4s^2p^1$	Ge = 32 = $[Ar]3d^{10}4s^2p^2$
n = 49 = $[Kr]4d^{10}5s^2p^1$	Sn = 50 = $[Kr]4d^{10}5s^2p^2$
Tl = 31 = $[Xe]4f^{14}5d^{10}6s^2p^1$) 28	Pd = 82 = $[Xe]4f^{14}5d^{10}6s^2p^1$

Characteristics of p-block elements :

i. The p - block elements are largely electronegative and form covalent compounds, on account of their large ionisation energies and electron affinities.

ii. They all show more than one valency and exist in different oxidation states.

Elements	*Oxidation states*
Ga (Gr, IIIA)	+1, +3
Ge (Gr, IVA)	+2, +4
P (Gr, VA)	+1, +3, +4, +5

iii. The reducing character decreases and the oxidising character increases from Gr, IIIA to VIIA in the p - blocks.

iv. The heavier p - block elements show inert pair effect.

As a consequence of this effect, they form ionic compounds in their lower oxidation states and covalent compounds in their higher oxidation states.

6 (3. d-block elements (Transition elements)

These elements have incompletely filled ultimate and penultimate shells. They are also known as d - block elements since they have incomplete d - orbitals int he penultimate shell. The properties of these elements are midway between those of s - block and p - block elements. Hence the name transition elements. They have the general electronic configuration $(n - 1)\ d^{1-10}ns^2$. There are three transition series of 10 elements each in the 4th, 5th and 6th periods.

First transition series : Scandium (21) to zinc (30)

Second transition series : Yitrium(39) to Cadmium (48)

Third transition series : Lanthanum (57) to Mercury (80).

The outer electronic configurations of the three series are given below:

First series	*Second series*	*Third series*
Sc $3d^1\,4s^2$	Y $4d^1\,5s^2$	La $5d^1\,6s^2$
Ti $3d^2\,4s^2$	Zr $4d^2\,5s^2$	Hf $5d^2\,6s^2$
V $3d^3\,4s^2$	Nb $4d^4\,5s^1$	Ta $5d^3\,6s^2$
Cr $3d^5\,4s^1$	Mo $4d^5\,5s^1$	W $5d^4\,6s^2$
Mn $3d^5\,4s^2$	Tc $4d^6\,5s^1$	Re $5d^5\,6s^2$
Fe $3d^6\,4s^2$	Ru $4d^7\,5s^1$	Os $5d^6\,6s^2$
Co $3d^7\,4s^2$	Rh $4d^8\,5s^1$	Ir $5d^7\,6s^2$
Ni $3d^8\,4s^2$	Pd $4d^{10}\,5s^0$	Pt $5d^9\,6s^1$
Cu $3d^{10}\,4s^1$	Ag $4d^{10}\,5s^1$	Au $5d^{10}\,6s^1$
Zn $3d^{10}\,4s^2$	Cd $4d^{10}\,5s^2$	Hg $5d^{10}\,6s^2$

18a

Characteristics of d - block elements :

i. Since the transition elements have a few valence electrons relative to the number of orbitals available, they are all metals. They are therefore, good heat and electrical conductors.

ii. They have almost the same ionisation potential since the electronic arrangement of their outermost shells remain unaltered. For this reason, they have striking resemblance in physical and chemical properties.

iii. **Both metallic and covalent bonding are possible in transition metals. The former is due to their low ionisation energy and the latter is due to the availability of half filled d - orbitals.**

iv. **They have high melting and boiling points as their atoms are held by strong covalent and metallic bonds.**

8a

a. s - block elements :

They have ns^1 or ns^2 configuration in their ultimate shells.

Example :	IA group	IIA group
	$Li = 3 = [He]2s^1$	$Be = 4 = [He]2s^2$
	$Na = 11 = [Na]3s^1$	$Mg = 12 = [Na]3s^2$
	$K = 19 = [Ar]4s^1$	$Ca = 20 = [Ar]4s^2$
	$Rb = 37 = [Kr]5s^1$	$Sr = 38 = [Kr]5s^2$
	$Cs = 55 = [Xe]6s^1$	$Ba = 56 = [Xe]6s^2$
	$Fr = 87 = [Rn]7s^1$	$Ra = 88 = [Ra]7s^2$ 8a

Characteristics of s-block elements :

i. Most of the s - block elements (except Be and Mg) impart characteristic colour to the flame e.g.,

Sodium - Yellow   Calcium - Brick red

ii. They have low ionisation energy due to their large atomic size.

iii. They are highly reactive electropositive elements. This is because of their tendency to lose electrons readily.

iv. On account of their readiness to lose electrons, they act as good reducing agents.

b. p - block elements :

8a They have a general configuration of ns^2p^{1-5} in the ultimate shell. That is, their p - orbitals are being filled up with electrons.

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v. The unpaired d - electrons get easily excited by the absorption of a small amount of energy in the form of heat or light. The compounds of transition elements are, therefore, coloured e.g.,

$CuSO_4$ - Blue $FeSO_4$ - Pale green

$MnSO_4$ - Pink $CrCl_3$ - Green

vi. Transition elements show variable oxidation states (polyvalency). This unique property is due to the involvement of d - electrons in bond formation along with the usual valence electrons. Example ;

Element	*Oxidation states*
Cu	+1, +2
Fe	+2, +3
V	+2, +3, +4, +5
Cr	+1, +2, +3, +4, +5, +6

vii. The availability of unpaired electrons in the d - orbital makes transition elements paramagnetic.

viii. Transition elements form non - stoichiometric compounds in which this various atoms are not in stoichiomatric ratio. e.g., $VH_{0.56}$, $Fe_{0.95}$, $Cu_{1.7}$, S.

ix. They form co-ordination compounds (complex ions). This is because of the presence of highly charged cations with vacant orbitals in the valence shell. These orbitals can accept ions pairs of electrons donated by ligands.

4. f-block elements (Inner - transition elements)

6Q (These elements contain incompletely filled ultimate, penultimate and antepenultimate shells. Since their f - orbitals in the antepenultimate shells are incompletely filled with electrons, these elements are also known as f - block elements. The general electronic configuration of them is $(n-2)\,f^{1-14}\,(n-1)\,d^{10}ns^2$.

There are two series of inner - transition elements. They are each 14 in number.

4f - series of Lanthanides : Ce (58) to La (71) (i)

5f - series of Actinides : Th(90) to Lw (103) (ii)) 89

Characteristics of f-block elements :

The f - block elements have similar electronic configuration in the outermost (n –1)d and ns orbitals. It is the f - orbitals in the antepenultimate shell that get filled with electrons gradually. Thus, they have striking resemblance in physical and chemical properties.

They have almost the same ionisation energy. The reason for this is their unique property namely lanthanide and a actinide contraction.

They generally have high melting and boiling points.

The common oxidation state of these elements is +3. In a few cases they have +2 and +4 oxidation states.

They are paramagnetic since their f - orbitals contain one or more unpaired electrons.

They form coloured compounds. The colour is due to the excitation of unpaired electrons.

They form co - ordination compounds using their vacant f - orbitals.

University Questions

1. State and explain Pauli's exclusion principle. 3,4,5
2. State and explain Aufbau principle. 6,7
3. State and explain Hund's Rule 5,6
4. Write notes on a. Atomic radii 16 b. Ionic radii 19
5. What are the factors which affect Ionisation energy?
6. Write notes on a. Electron affinity 23 b. Electronegativity 25
7. Discuss the classification of elementry based on their electronic configuration.
8. Write shote note on.
   - i. s - block elements 30
   - ii. b - block elements 30, 31
   - iii. d - block elements 31, 32
   - iv. f - block elements 33, 34