{CHAPTER----2}
CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

UNIT : 1

1. The concept of grouping elements in accordance to their properties
2. The Periodic Law – Mendeleef's classification of elements
3. Significance of atomic number and electronic configuration as the basis for periodic classification
4. Classification of elements into s,p,d,f – blocks and their main characteristics

VERY SHORT ANSWERS :

1. What are the representative elements ? Give their valence shell configuration.

A. S and P block elements excluding zero group elements are called representative element.

Their general valence shell electronic configuration is ns^{1-2} and np^{0-5}

2. Justify the position of lanthanides in periodic table.

A. If the lanthanides are placed in the mainbody of periodic table , the symmetry of the whole arrangement would be disturbed due to abnormal electronic config.. For that reason elements are kept away from the main part of the table and are placed at bottom.

3. If an element X has an atomic number 34, give its position in the periodic table.

A. This element is selenium. Its valence shell configuration is[Ar] $4s^2\,4p^4$. So, it belongs to VIth group and 4th period.

4. What are representative elements ? Why are they called like that ?

A. Elements which have incompletely filled outer shells are called representative elements. Because of the more abundantance of occurence and active nature , these elements are called representative elements.

5. What factors impart characteristic properties to the transition elements ?

A. Small atomic size, high nuclear charge variable valency and unpaired electrons in 'd' subshells impart characteristic properties to the transition elements.

Short Answer questions :

1. What is the necessity to classify the elements ?

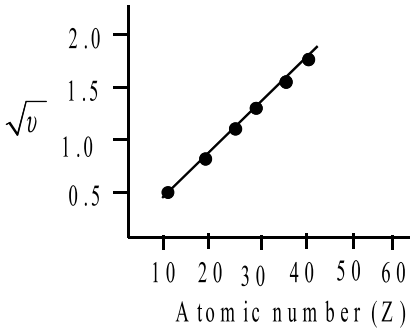
A. To study the elements and their chemistry separately is a very difficult task. The classification simplifies the chemistry of all the known elements.

The classification of elements would not only give the knowledge of known chemical facts but also assist in even predicting the elements to be discovered .

2. How did Moseley's work on atomic numbers show that atomic number is a fundamental property better than atomic weight ?

A. Moseley bombarded various elements with cathode rays in a discharge tube,.which resulted the formation of X – rays with characteristic frequencies. He showed that the highest

frequency radiations of he X – rays (k_α and k_β lines) is related to the the charge present in the nucleus of an atom of the element which is used as anode in the discharge tube. The equation is represented as $\sqrt{\upsilon} = a(Z - b)$, where $\sqrt{\upsilon}$ = frequency of he X – ray line, Z = atomic number of the element .



(Linear relationship between Z and $\sqrt{\upsilon}$)

a, b are constants for an selected type of the line

A plot of $\sqrt{\upsilon}$ against 'Z' gives a straight line. This can be shown in the graph

However, no such relationship was obtained when the plot was drawn between frequency and the atomic mass.. . Moseley concluded that the Z is a fundamental quantity of an atom which increases in regular steps with increasing atomic number. . The correlation between X – rays spectra and atomic number indicated that an element is characterized by its atomic number and not by atomic mass.

3. Define Modern periodic law. How many groups and periods are present in the long form of the periodic table ?

A. Modern form of periodic table : The physical and chemical properties of elements are periodic functions of their electronic configurations.

It is graphical representation of Aufbau's principle

The long form of periodic table consists of 18 groups and 7 periods

4. How many groups and periods, the elements are arranged in the long form of the periodic table ? Why the f – block elements are placed below the main table ?

A. In the long form periodic table, the elements are arranged in the increasing order of their atomic numbers

There are 18 groups and 7 periods in he long form periodic able.

The f – blocks are not given an individual place in the mainbody of periodic table and are placed at the bottom of the periodic table., because these elements have the same configurations in their n and (n-1)shells and are chemically posses simlar proporities.

5. Mention the number of elements present in each of the periods in the long form of periodic table

A. In the longform of periodic table there are 7 periods.The distribution of elements in the periods are as follows

1st period ---------2

2nd period----------8

3rd period ----------8

4th period ----------18

5th period ----------18

6th period ----------32

7th period ----------20 (incomplete)

6. Give any five of the characteristic properties of transition elements.

A. Definition : Transition elements is defined as element with possess atleast one unpaired electron in its free state or in its sable oxidation states.

All d – block elements except IIB group (Zn, Cd, Hg) are called transition elements.

Electronic configuration is $(n-1)d^{1-10}\ ns^{1-2}$ where n = 4, 5,6 (or) 7

Properties of transition elements :

1) They are hard and heavy metals

2) They have high melting points, boiling points and densities

3) They are good conductors of heat and electricity

4) They form alloys like brass, bronze, german silver, etc

5) They are paramagnetic in nature due o he presence of unpaired electrons in their (n-1)sub levels

6) Transition elements and their oxides are good catalyst in various chemical reactions

e.g :

1) Nickel is used as catalyst in hydrogenation of oils

2) Iron is used as catalyst and molybdenum as promoter in the synthesis of NH_3 by Haber's process.

7. Classify the elements on he basis of their oxides.

A i) Nature of oxides of elements : Metallic oxides are basic in nature and non – metallic oxides are acidic in nature. Metalloid oxides are amphoteric (both acidic and basic)

IA group elements are basic due tohighly electropositivenature . Down he group metallic natureand basic nature increases..

e.g : Cs_2O is most basic as caesium is highly electropositive

e.g;3rd period elements oxides :

Na_2O	MgO	Al_2O_3	SiO_2	N_4O_{10}	SO_3	Cl_2O_7
Srong	Weak	Ampho	Weak	Weak	Srong	Srong
Base	base	eric	acid	acid	acid	acid

From left to right periodic tablenon – metallic nature and acidic nature increases ..

8. Write a note on the variation of metallic nature in group and periods.

A. In an group, metallic nature increases and non – metallic nature decreases.

Down the group, metallic nature increases,due to the increase in electropositivity. Across a period, metallic nature decreases and non – metallic nature increases.

In the 3rd period elements Chlorine is more non – metallic

Na Mg Al Si P S Cl

Long Answer Questions :

1. Discuss in detail about the classification of elements by Mendeleeff.

A. The law can be stated as : The physical and chemical properties of the elements and their compounds are the periodic functions of their atomic weights (atomic masses).

Characteristics of Mendeleeff's periodic table :

- The elements are arranged in the increasing order of their atomic weights.
- There are seven periods.and nine groups.
- The vertical columns are called groups.which are subdivided intoA and B subgroups.
- The first three periods are called short periods. In the first period, only hydrogen and helium are present The elements with low atomic weights (typical elements) are present in these short periods
- The fourth, fifth and sixth periods are long periods
- The elements present in zero group are noble gases
- VIII group consists of Fe ,Co.Ni, Ru, Rh.Pd, Os, Ir, Pt and are called traids
- He observed that the elements with similar properties are

  a) having almost the same atomic weights

eg : Fe (56), Co (59), Ni (59)

Os (191) ; Ir (193) ; Pt (195)

b) having a constant increase in atomic weight in the group

K (390, Rb (85), Cs (133)

Ca (40), Sr (88), Ba (137)

Advantages of Mendeleeff's table :

1) Discovery of new elements : Mendeleeff predicted the properties of some elements, which were not known at that time and left gaps for them,that helped.the discoverof these elements.

Mendeleeff's name	**New name**
e.g : Eka aluminium	Gallium
Eka silicon	Germanium
Eka boron	Scandium

2) Correction of atomic weights : He corrected the atomicweight of the corrected that atomic weight of beryllium. Similarly he corrected the atomic weight of indium and uranium also.

Drawbacks of Mendeleef's table :

1**) Position of hydrogen :** Hydrogen resembles with alkali metals (IA) and halogens (VIIA) in properties. So it should have been placed in both IA and VIIA groups in the Mendeleeff's table. It has actually been placed in IA group.

2) Position of isotopes : In Mendeleeff's table, isotopes of an element should be placed in different places due to their different atomic masses. This will be anomalous because isotopes possess similar chemical properties.

3) Position of lanthanides and actinides : The position of lanthanides (14 elements following lanthanum, atomic number 58 – 71) and actinides (14 elements following actinium, atomic number 90 – 103) is also anomalous in Mendeleeff's periodic table. In this table, all these elements arenot placed in mainbody of the periodic table.

4) Anomalous pairs of elements : Argon (atomic mass = 39.9) is placed before potassium (atomic mass = 39.1). If the criterion of atomic weight is strictly followed, then potassium should be placed before argon, i.e, potassium should be placed in zero group while argon in IA group. The other examples are Tellurium and iodine , cobalt and nickel

5) Similar elements placed in different groups and dissimilar elements are placed in the same groups.

The elements like Ag and Tl, Ba and Pb, Cu and Hg show similar properties. but they are placed in different groups.

2. From a study of the properties of neighbouring elements, the properties of an unknown element can be predicted – Justify with an example.

A. Mendeleeff's was able to predict new elements and their properties with the help of properities of neighbouring elements .

Predicted element	**Name of that element**
Eka aluminium	Gallium
Eka silicon	Germanium
Eka boron	Scandium

Zero group elements were not known at the time of Mendeleeff. After the discovery of the inert gas argon Rayleigh and Ramsay, the other inert gases of the group would be predicted and were isolated.

3. What are the limitations of the Mendeleeff's periodic law ? Give any four of them.

A. **Drawbacks of Mendeleeff's table** :

1) Position of hydrogen : Hydrogen should be placed in both IA and VIIA groups in the Mendeleeff's table,as it resembles with alkali metals (IA) and halogens (VIIA) in properties,but is placed in IA group.

2) Position of isotopes : All isotopes did not find the place in Mendeleeff's table. Isotopes of an element should be placed in different places due to their different atomic masses. This will be anomalous because isotopes possess similar chemical properties.

3) Position of lanthanides and actinides : The position of lanthanides (Z=58 – 71) and actinides (Z=90 – 103) did not find the place in the mainbody of the periodic table.

4) Anomalous pairs of elements : More atomicweight element preceeds the less atomicweigh. Argon (atomic mass = 39.9) is placed before potassium (atomic mass = 39.1). If the criterion of atomic weight is strictly followed, then potassium should be placed before argon, i.e, potassium should be placed in zero group while argon in IA group. The other examples are Tellurium and iodine , cobalt and nickel

5) Similar elements placed in different groups :

The elements like Ag and Tl, Ba and Pb, Cu and Hg show similar properties. Yet they are placed in different groups.

6) Dissimilar elements placed in the same groups .

4. Discuss the construction of the long form of the periodic table. (or)
Define modern periodic law. How is the long form of the periodic table constructed on the basis of this law ?

A. **Modern form of periodic table** : The physical and chemical properties of elements are periodic functions of their electronic configurations.It is graphical representation of Aufbau's principle.

The vertical column are called "groups" and the horizontal rows are called "periods"

In the long form of periodic table 18 groups and 7 periods are present and the elements are arranged in the increasing order of atomic numbers.

Groups : It represents number of valency electrons

Elements present in the same group have similar chemical properties due to similar outer electronic configuration.

Groups in this table are IA, IIA, IIIB, IVB, VB, VIB, VIIB, VIIIB (3 subgroups), IB, IIB, IIIA, IVA, VA, VIA, VIIA, VIIIA (or) according to IUPAC system 1 – 18

e.g : Alkali metals (or) IA group :

Hydrogen (H)	:	$1s^1$	(Z = 1)
Lithium (Li)	:	$[He]2s^1$	(Z = 3)
Sodium (Na)	:	$[Ne]3s^1$	(Z = 11)
Potassium (K)	:	$[Ar]4s^1$	(Z = 19)
Rubedium (Rb)	:	$[Kr]5s^1$	(Z = 37)
Caesium (Cs)	:	$[Xe]6s^1$	(Z = 55)
Francium (Fr)	:	$[Rn]7s^1$	(Z = 87)

- The general outer electronic configuration is ns^1
- These elements show similar chemical properties i.e, they show common oxidation state of +1.They are all monovalent. They form similar types of oxides, hydroxides, halides, etc.

Periods:

- Each period starts with alkali metal and ends with noble gas element.
- **1st period :** It contains only two elements. H and He. Hence, this period is called shortest period .
- **2nd and 3rd periods** : These periods are called short periods and each period contains 8 elements only
- **4th and 5th periods** : These periods are called long periods and continuous periods. Each period contains 18 elements.

  4th period starts with K (19) ends with Kr (36), 5th period starts with Rb (37) ends with Xe (54)
- **6th period :** It is the longest period contains 32 elements including 14 lanthanides [Ce (58) to Lu (71)]. This period starts with Cs (55) ends with Rn (86)
- **Lanthanides (14 elements) :** These are called rare earths. In these elements, the differentiating electron enters into 4f – sub shell.

General electronic configuration :

$4f^{-1-14}\ 5d^{0-1}\ 6s^2$ Lanthanides are also called first inner transition elements,

- **7th period :** It is an incomplete period and contains 23 elements including 14 actinides [Th (90) to Lr (103)]. This period starts with Fr (87).
- **Actinides :** In these elements, the differentiating electron enters into 5f – subshell.

  General electronic configuration :

  $5f^{1-14}\ 6d^{0-1}\ 7s^2$

  Actinides are called second inner transition elements

Periods and energy levels in period table :

Period	Energy levels	Number of elements the period
1st period	1s	2
2nd period	2s 2p	8
3rd period	3s 3p	8
4th period	4s 3d 4p	18
5th period	5s 4d 5p	18
6th period	6s 4f 5d 6p	32
7th period	7s 5f 6d	23

6. Discuss the relation between the number of electrons filled into the sub – energy levels of an orbit and the maximum number of elements present in a period.

A. In the long form periodic table, there are seven periods. Based on the number of electrons filled into various sub energy levels in the formation of a period the number of elements in period can be calculated.

1) First method : The first period starts with the filling '1s' is two. So a total number of two elements are present in the first period.

Periods. Energy levels and total number of elements in a period

S.No.	Number of the periods	Main and subenergy levels (as (nl)	Total number of electrons filled in all the sub-levels	Total number of electrons in the period
1	1	1s	2	2
2	2	2s 2p	8	8
3	3	3s 3p	8	8
4	4	4s 3d 4p	18	18
5	5	5s 4f 5d 6p	18	18

6	6	6s 4f 5d 6p	32	32	
7	7	7s 5f 6d	___		incomplete

2) Second period : The second period starts with the filling of '2s' and ends with '2p' is complete. The total number of electrons filled in 2s and 2p will be 2 + 6 = 8. So in the second period there are 8 elements.

3) Third period : The third period starts with the filling of '3s' and ends with '3p' is complete. The total number of electrons filled in 3s and 3p will be equal to 2 + 6 = 8. So in the third period, there are 8 elements.

4) Fourth period : In the fourth period, 4s, 3d and 4p sublevels are filled. The total number of electrons filled in 4s, 3d and 4p sublevels is 2 + 10 + 6 = 18. So in the fourth period, there are 8 elements.

5) Fifth period : In the fifth period, 5s, 4d and 5p sublevels are filled. The total number of electrons filled in 5s, 4d and 5p sublevels is 2 + 10 + 6 = 18. So in the fifth period, there are 8 elements.

6) Sixth period : In the sixth period, 6s, 4f, 5d and 6p sublevels are filled. The total number of electrons filled in 4s, 3d and 4p sublevels is 2 + 14 + 10 + 6 = 32. there are 32 elements in the sixth period.

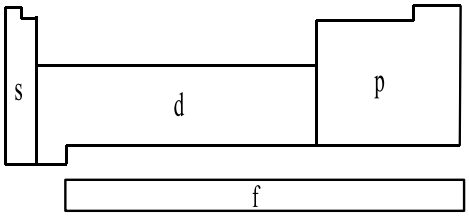
7) Seventh period : In the seventh period, the electrons are filled in the order 7s, 5f, 6d and 7p. This can also accommodate 32 elements. But actually this period is incomplete.

7. Write an essay on the division of elements into s, p, d and f – blocks.

A. Based on electronic configurations of elements, element are classified into four blocks in the periodic table. They are

1) s – block elements p – block elements

3) d – block elements 4) f – block elements



s, p, d and f block in periodic table

1) s – block : s – block has two groups IA and IIA

Group – I (H, Alkali metals) : Electronic configuration of outer shell is ns^1

Group – II (Alkaline earth metals) : Electronic configuration of outer shell is ns^2

$\therefore$ Electronic configuration of s – block elements in ns^1 (or) ns^2

In s – block elements, the differentiating electron enters into s - sub shell of valency shell.

All s – block elements are highly electropositive metals.

2) p – block : P – block has six groups

IIIA group : (Boron family)

Electronic configuration of outer shell : $ns^2 - np^1$

IVA group : (Caron family) : Electronic configuration of outer shell : $ns^2\, np^2$

VA group : (Nitrogen family) or (Pnicogens)

Electronic configuration of outer shell : $ns^2\, np^3$

VIA group : (Oxygen family) or (chalcogens)

Electronic configuration of outer shell : $ns^2\, np^4$

VIIA group : (Halogens) :

Electronic configuration of outer shell : $ns^2\, np^5$

Zero group : (Noble gases) :

Electronic configuration of outer shell : $ns^2\, np^6$ except for He which has $1s^2$

Electronic configuration of p – block elements is $ns^2\ np^{1-6}$ (or) $ns^2\ np^1$ to $ns^2\ np^6$

In p – block elements, the differentiating electron enters upto p – sub shell of valency shell.

All non – metals, metalloids and few metals belongs to p – block.

3) d – block : d – block has 10 groups.

The general electronic configuration in $(n-1)d^{1-10}\ ns^{1-2}$ where n = 4, 5, 6 or 7

In this block elements, the differentiating electron enters into d – sub shell of penultimate (n – 1) shell

In this block, elements are divided into four series and each series contain 10 elements.

3d – series : From scandium (Z = 21) to zinc Z = 30), in the fourth period

4d – series : From yttrium (Z = 39) to cadmium (Z = 48) in the fifth period

5d – series : From lanthanum (Z = 57), hafnium (Z = 72) to mercury (80) in 6^{th} period

6d – series : It is an incomplete series from actinium (Z = 89) in the seventh period. d – block elements are hard and heavy metals with high melting point and boiling point.

4) f – block : This block has two series and each series has 14 elements. The general electronic configuration is $(n-2)f^{1-14}\ (n-1)d^{0-1}ns^2$, where n = 6 or 7

In f – block elements, the differentiating electron enters into f – sub shell of antepenultimate (n – 2) f shell

4f – series : From cerium (Z = 58) to lutetium (Z = 71). The elements in this series are called lanthanides (or) rare earths.

5f – series : From thorium (Z – 90) to law Trencium (Z = 103). The elements of this series are called actinides.

Thorium and uranium are naturally occurring, remaining elements are man – made and all are radioactive elements, f – block elements are also called inner transition elements.

8. Relate the electronic configurations of elements and their properties in the classification elements.

A. On the basis of electronic configurations and properties, the elements are classified into four types. They are

1) Noble gas elements

2) Representative elements

3) Transition elements

4) Inner transition elements

1) Noble gas elements :

a) In the elements, the outermost orbit is completely filled The configuration of other elements is $ns^2\,np^6$. these are chemically inert. So these are also known as "inert gases"

b) The configuration of Helium = $1s^2$

2) Representative elements :

a) In these elements, only the outermost orbits (n) is incomplete

b) The s and p – block elements, except zero group elements, belong to this type

c) The valence configurations of these elements is $ns^{1-2}\,np^{0-5}$

d) These are chemical active. They take part in reactions by losing (or) gaining (or) sharing of electrons

e) The representative elements include metals, non metals and metalloids

3) Transition elements :

a) In these elements, the outermost two orbits [n and (n – 1)] are incompletely filled

b) These are d – block elements

c) The differentiating electron enters into the (n – 1) d – orbital

d) The outer electronic configuration of these elements is $(n-1)\,d^{1\ or\ 10}\ ns^{1-2}$ where $n \geq 4$

Due to small size of atom high nuclear charge and unpaired electrons in 'd' orbitals impart characteristic properties to the transition elements. These properties can be listed as follows

i) They are hard and heavy metals

ii) They have high melting point, boiling point and densities

iii) They are good conductors of heat and electricity

iv) They exhibit more than one oxidation state

e.g : Fe exhibits +II and +III states

v) They form coloured ions in solutions due to d – d transition

e.g : Fe^{+2} - Pale green

Fe^{+3} - Brown colour

vi) Transition metals and their compounds are good catalysts for various chemical processes

Process	Catalyst	Product
1) Haber's catalyst	Fe/Mo (Promoter)	NH_3
2) Ostwald's	Pt	HNO_3
3) Contact	Ni	Dalda

vii) These metals and their ions generally exhibit paramagnetic properties due to the presence of unpaired electrons in their s (or) d – orbitals.

viii) They form useful alloys eg : Brass, Bronze, German silver etc. Zn, Cd and Hg present in IIB group are not strictly d – block elements as per the definition. But they are studied along with transition elements to rationalize their chemistry.

4) Inner transition elements :

a) In these elements, the outermost three orbits are incomplete

b) The outer electronic configuration of these elements is $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$

c) The differentiating electron enters into (n – 2) f – orbitals

d) These are f – block elements. These include lanthanides and actinides

e) In these elements, the configuration in the outer two shells is same, so they have closer similar physical and chemical properties

f) The common oxidation state of these elements is + 3

g) Lanthanides – Ce (Z = 58) to Lu (Z = 71)

Actinides – Th (Z = 90) to Lr (Z = 103)

h) Lanthanides – 4f series – 6^{th} period and IIIB group

Actinides – 5f series – 7^{th} period and IIIB group

g) Most of these elements are synthetic

h) These are radioactive

CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

UNIT: 2

1 Periodic trends in physical and chemical properties of elements
2 Periodic trends of the elements with respect to Atomic radii, Ionic radii, Inert gas Radii

VERY SHORT ANSWER QUESTIONS:

1. The atomic radius of an inert gas atom appears to be the highest among the p – block elements. Why?

A. In noble gases all the electrons in their atoms are paired. The repulsive forces become stronger than the attractive forces between the nucleus and the electrons. Hence, radii increase.

2. Name the catalyst used in Haber's synthesis of ammonia and hydrogenation of oil?

A. Catalyst used in Haber's process is Fe. And in hydrogenation of oils is Ni.

3. Which is the biggest among the ions Na^{+}, F^{-}, O^{-2} and Mg^{+2} ? Why?

A. O^{-2}, the above four are iso electronic ions. The ion with lower nuclear charge has greater size.

4. Why is the order of radii given as $Fe > Fe^{+2} > Fe^{+3}$?

A. As the positive charge increases on metal atom, its effective nuclear charge increases and size decreases

Short Answer questions:

1. Why does the covalent radius increase in a group?

A. As the atomic number increases down the gouphe radius also increases due to the addition of a new shell. Due to this increase of nuclear attraction and also increase of atomic size takes place. The later outweighs hence the atomic radius or covalent radius increases down the group.

2. How does the atomic radius vary in a period of elements?

A. Across a period nuclear charge increases as the atomic number of the element increases. The electrons enter into the same orbit. Due to this nuclear attraction on the valence electrons increases thus the atomic radius decreases.

In every period, the alkali metal atom is the largest, the halogen atom is the smallest Vander Waals radius in considered for inert gases and for halogens, covalent radius is considered. Since the Vader Waals radius is greater than covalent, radius inert gases have larger size than the halogens.

3. The atomic radii of transition elements decrease slowly as the atomic number increases in a period. Account for it.

A. In the transition elements, the differentiating electron enters the (n – 1)d orbitals. The (n – 1)d electrons screen the outer electrons from the attraction of the nucleus. Inspite of increase in nuclear charge due to shielding effect of (n – 1)d electrons the attraction of the nucleus over the outer electrons is feeble. As a result atomic radius decreases slightly from left to right in da transition series. IIB group elements have larger size than IB elements due to the repulsions among the completely filled d electrons.

4. Among the groups of transition elements, the ionic radius increases with an increase in atomic number. Justify.

A. In the transition elements, the differentiating electron enters into (n – 1)d subshell. Down the group, the number of shells increases. As the number of shells increases the shielding effect also increases. Due to this ionic radius increases in a group among the transition elements.

5. As atomic number increases, atomic size increases in a group and decreases along a period. Give reasons for this behaviour.

A. On moving down the group the atomic size increases with an increase in atomic number the number of inner shells and shielding effect increases though there is an increase in nuclear charge, the size of the atom increases.

With an increase in atomic number the size of the atom decreases from left to right in a period. This is because of the increase in effective nuclear charge over the outermost electrons; the electrons are pulled towards the nucleus causing the decrease in the size of the atom.

6. What is lanthanide contraction? What are its consequences?

A. **Lanthanide contraction:** The steady decrease of atomic (or) ionic radii from left to right in lanthanides is called lanthanide contraction.

It is due to poor shielding effect and peculiar shape of f – orbitals.

Consequences: Due to lanthanide contraction

i) The hardness,M.P,B.P of the elements inceases from Ce to Lu.

ii) The atomic radii of 5d and 4d transition elements are very close to each other when compared to those of 3d and 4d transition elements.

iii) The effect of lanthanide contraction is also observed in the 5d transition series.

iv) Zr and Hf resemble each other due to this there isolation becomes difficult.

Long Answer Questions:

1. Write a note on the following:

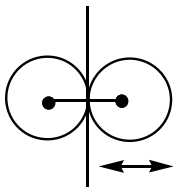
A) Crystal radius B) Vander Waal's radius and C) Covalent radius of elements

A. A) Crystal radius: It is used to metal atoms.

Definition: Half the distance between the centres of the nuclei of two adjacent metal atoms is called crystal radius or atomic radius.

e.g : the internuclear distance between two sodium atoms is 3.72A

$\therefore$ Crystal radius of sodium = $\frac{3.72}{2} = 1.86\,\overset{o}{A}$



Metallic radius

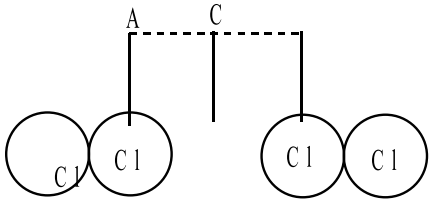
Vander Waal's radius: It is used for molecules in he solid state only

Definition: Half the distance between the centres of the nuclei of two atoms of different molecules which are close to each other

E.g : the distance molecules is $3.6\overset{o}{A}$

Vander Waal's radius of chlorine = $\frac{3.6}{2} = 1.8\overset{o}{A}$

Vander Waal's radius is 40% greater than covalent radius

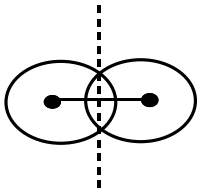


Covalent radius: It is used for non metals .

Definition: Half the distance between he centres of nuclei of two atoms, held b a covalent bond in a homadiatomic molecule.

e.g : the internuclear distance between chlorine atoms in Cl_2 molecule is $1.98\overset{o}{A}$

$$\therefore \quad \text{Covalent radius} = \frac{1.98}{2} = 0.99\overset{o}{A}$$



Covalent radius

Order of radius: Vanderwaals radius > crystal radius > covalent radius

CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

UNIT: 3

1 Ionization Energy, Electron Gain Energy, Electro Negativity.
2 Variation of atomic radii in inner transition elements

VERY SHORT ANSWER QUESTIONS:

1. What is electro – negativity? How is it calculated on Pauling's scale? How is EN useful in understanding the nature of elements?

A. The tendency of an atom to attract the shared pair of electrons towards itself is called electro negativity. It is calculated on Pauling's scale based on bond polarity.

Nonmetals have greater E.N metals have less E.N

2. Write the relation between the magnitude of screening effect and the ionization enthalpy.

A. As the number of electrons in the inner shell increases, shielding effect also increases. So I.P decreases

3. Write the electronic configuration of Chromium (z = 24). What is its importance?

A. Electronic configuration is $[Ar]3d^5\ 4s^1$. The importance is its half-filled stable configuration is against to Aufbau principle.

4. What is meant by penetration of orbitals? Write the sequence of the penetration of the atomic orbitals.

A. The diffusibility of an orbital more towards the nucleus is called penetration of orbitals.

Penetrating power of orbitals is in the order s > p > d > f

5. Write the relation between Pauling scale and Mullikan's scale of electro negativities.

A. E.N. values measured on the Mullikan scale are 2.8 times greater than the values measured on Pauling scale.

When I.P. and E.A. values are in $K.J.mol^{-1}$ then

6. What valency is possible to Arsenic with respect to oxygen?

A. Arsenic can form two oxides i.e, As_4O_6 and As_4O_{10}. In which it exhibits 3 and 5 respectively

7. Name the catalyst used in Haber's synthesis of ammonia and hydrogenation of oil?

A. Catalyst used in Haber's process is Fe. And in hydrogenation of oils is Ni.

8. What is an amphoteric oxide? Give the formula of an amphoteric oxide formed by the element in 3rd period.

A. The oxide having both basic and acidic nature is called amphoteric oxide.

Ex: Al_2O_3

9. Name the most E.N element and the element. Is it also having the highest E.A value? Why?

A. Most electronegative element is 'F'. No, it is not the element with highest electron affinity due to very small size.

10. What is shielding effect how does it affect I.P values?

A. Decreasing the force of attraction between the nucleus and other most electrons by the inner core electrons is called shielding effect. As shielding effect increases I.P decreases and vice – versa.

11. What is diagonal relationship? Give one example?

A. The resemblance in properties between the element present in a group of second period and the element present in next group in third period is called diagonal relationship.

	IA	IIA	IIIA	IVA
Second period:	Li	Be	B	C
Third period:	Na	Mg	Al	Si

12. How does the nature change from Na_2O to Cl_2O_7 in the period?

A. From Na_2O to Cl_2O_7 basic character decreases and acidic character increases

Na_2O MgO Al_2O_3 SiO_2 P_2O_5 SO_3 Cl_2O_7

Basic amphoteric acidic

13. Define electro-negativity of an element on Mullikan scale?

A. According to Mullikan scale E.N. of an element is the average of its $I.P_1$ and Ea_1

$$E.N = \frac{I.P_1 + E.A_1}{2}$$

14. B has less IP than the Be why?

A. IP Be > B

1) $[He]2s^2$ $[He]2s^2 2p^1$

Stable full filled electronic configuration incompletely filled electronic configuration.

Penetrating power of 2S electrons, in Be greater than 2P electron of B

15. 'N' has more I.P than 'O" why?

A. IP N > O

1) $[He]2s^2 2p^3$ $[He]2s^2 2p^4$

Stable electronic configuration due to less stable electronic configuration, relatively when completely filled (or) half-filled compared to 'N'

16. Why is the 2nd IP of an element greater than its first IP?

A. After removal of one electron, on the remaining electrons the effective nuclear charge increases because of attraction between nucleus and outer electron increases so $IP_2 > IP_1$

17. What is the nature of a bond between two atoms predicted based on their electro negativity values?

A. If E.N. difference is 1.7 (or) more than it then the bond formed between these two elements is more ionic if its less than 1.7, bond is more covalent is nature.

Short Answer questions:

1. Arrange the elements Be, B, N and O in the increasing order of their ionization potentials.

A. Increasing order of the ionization potentials:

B < Be < O < N

Be has completely filled $2s^2$ orbital hence has greater IPE than B with incompletely filled $2s^2\ 2p^1$. Due to the half-filled p sub shell $2p^1_x\ 2p^1_y\ 2p^1_z$ N has greater IPE than O $2p^2_x\ 2p^1_y\ 2p^1_z$.

2. Which elements of the third period have highest ionisation energy? Explain variation of ionisation energy in his period.

A. Argon has the highest IE among the III period elements.

In each period, the last element inert gas element has the highest IPE because all orbitals are completely filled.

Variation of IPE along III period: among III period elements from left to right IPE increases due to decrease in size. Al & S have less IPE values than expected due to greater tendency to loose electron.

3. Write short note on electron affinity.

A. **Electron affinity**: The amount of energy released when an electron is added to the neutral isolated gaseous atom is called electron affinity.

Units: eV (or) kcal/mole (or) kJ/mole

$$x_{(g)} + e^- \rightarrow X^-_{(g)} + E\left(E.A_1\right)$$

The released energy is also called first electron affinity $\left(E.A_1\right)$. Bu the $E.A_2$ (second electron affinity) value has a +ve value, because energy is needed to overcome the repulsive forces between uni negative ion and newly entering electron.

Factors: Electron affinity depends on size and the effective nuclear charge.

Variation of electron affinity: Electron affinity decreases down the group as the atomic size increases.

Exception:

1) In every group, second element has greater electron affinity than first element

e.g: In halogens, 'C' has greater electron affinity than fluorine due o small size of 'F' and more electronic repulsions is F.

Similarly, S > O ; P > N ; Si > C ; Al > B

2) In IIA group elements, electron affinity value is +ve due to stable ns^2 configuration.

Across a period electron affinity increases as he atomic size decreases

e.g:

1) In a period halogens have greater electron affinity due to small size.

2) noble gas elements have zero electron affinity due to the completely filled outermost orbit.

Electron affinity values are obtained indirectly by Born – Haber cycle.

4. What is electropositive nature? Explain is variation in a group and along a period.

A. **Electro positivity**: The tendency of an element to lose on electron is called electropositive and the element is electropositive in nature.

$M \rightarrow M^{n+} + ne^{-}$

Meals are electropositive in nature

Variation of elecropositivity: As atomic size increases, down the group, elecropositivity increases.

IA group elements are highly electropositive.

Highest electropositivity possessed by an element is caesium (Cs)

As atomic size decreases, across a period, elecropositivity decreases.

5. What is electro-negativity? How is it calculated on Pauling's scale? How is EN useful in understanding the nature of elements?

A. **Elecronegativity**:

Definition: The tendency of an atom to attract the shared pair of electrons towards itself is called electro negativity; It is he property of bonded atom.

Electro negativity depends on

1) Nature of bonded atoms

2) Atomic size

3) Nuclear charge

Electro negativity scales:

1) **Pauling scale:** his scale is based on bond energy values. He bond energy of compound "A – B" is the average of bond energies of A – A and B – B molecules.

$$E_{A-B} = \frac{1}{2}\left[E_{A-A} + E_{B-B}\right]$$

But the experiment value E^{1}_{A-B} is found to exceed the theoretical value E_{A-B}.The difference is "Δ"

$$\Delta = E^{1}_{A-B} - E_{A-B}$$

Δ indicates the polarity of the covalent bond. According to pauling

$X_A - X_B = 0.208\sqrt{\Delta}$ (where X_A and X_B are E.N. of A and B)

Uses of Elecronegativity:

1. In the identification of nature of bond if E.N. difference is 1.7 or greater than 1.7 hen he bond is ionic, if I is less than 1.7 Then he nature of bond is covalent.

2. Non metals have more EN and metals have less EN.

6. What is valency? How does the valency with respect to hydrogen vary in a period of elements?

A. Valency is the combining capacity of an element

Valency of an element is the number of hydrogen atoms, with which one atom of the element combines

In a period, the valency, with respect to hydrogen, increases upto IVA group and then decreases.

7. What is inert pair effect? What type of elements exhibit this effect?

A. The reluctance of the electron pair present in the outer shall to get unpaired and to get involved in bonding is called inert pair effect.

The effect is predominantly seen in the heavier elements of IIIA, IVA and VA group

Group element	**Valence configuration**	**Common O.S. of the elements in the group**	**O.S. of the last in the group**
IIIA	$ns^2\,np^1$	+1, +3	+1
IVA	$ns^2\,np^2$	+2, +4	+2
VA	$ns^2\,np^3$	+3, +5	+3

e.g: 1: Thallium is more stable in +1 state than in +3 state

TlCl is more stable than $TlCl_3$

2. Lead is more stable in +2 state than in +4 state

$PbCl_2$ is more stable than $PbCl_4$

3. Bismuth is more stable in +3 state than in +5 state

4. $BiCl_3$ is more stable than $BiCl_5$

8. What is diagonal relationship? Give a pair of elements, having diagonal relationship.

A. In the periodic table, an element of a group in the second period is similar in properties with the second element of the next group. This type of relationship is called diagonal relationship.

	IA	IIA	IIIA	IVA
Second period:	Li	Be	B	C
Third period:	Na	Mg	Al	Si

The diagonal relationship is due to similar atomic or ionic size, electro negativity, polarizing power.

$$\text{Polarising power} = \frac{\text{Ionic charge}}{(\text{Ionic radius})^2}$$

Diagonally related elements have similar properties like:

a) BeO and Al_2CO_3 are amphoteric

b) Be_2C and Al_4C_3 give methane on hydrolysis

$$Be_2C + 4H_2O \longrightarrow 2Be(OH)_2 + CH_4$$

$$A_4C_3 + 12H_2O \longrightarrow 4Al(OH)_2 + 3CH_4$$

9. What are "rare earth" and "transuranic" elements?

A. 1) Lanthanides are called rare earths.

Lanthanides: Cerium (58) to Lutetium (71) and belongs to f – block

2) Elements after uranium are called transuranic elements and these elements are man – made. All are radioactive elements.

Long Answer Questions:

1. What are periodic properties? How do the following properties change in a (i) group and (iii) period? Explain.

a) Atomic radius b) Electron gain enthalpy c) Electronegativity

A. Periodic property: The repetition of similar chemical properties of elements at regular intervals with increasing atomic number is called periodic property or periodicity.

a) Atomic radius:

1) It increases down the group as the electrons enter into new shells

2) It decreases in period because the nuclear charge increases, but electrons enter into same outer shell

In a given period, alkali metal, is the largest atom and halogen is the smallest atom

The atomic radius of zero group elements is large due to the consideration of Vander Waals radius.

b) Electron affinity:

1) It decreases down the group due to increase in atomic size. But the second element in a group has greater electron affinity than the first.

Eg: Cl has greater EA than F.

2) Electron affinity increases across a period because the atomic size decreases and non-metallic nature increases. Inert gas elements have zero electron affinity values due to stable octet configuration in their outermost orbit

c) Electronegativity: It decreases down the group as the atomic size increases. It increases from left to right in a period as the atomic size decreases.Inert gas elements have zero electro negativity values due to stable outer configuration.

2. Define first and second ionization potentials. Why is second ionisation potential greater than the first ionization potential? Discuss three factors affecting IPE values of elements.

A. **Ionization energy:** The minimum energy required to remove loosely bound electron from an isolated, gaseous atom is called ionization energy.

It is also called the first ionization energy $\left(I_1\right)$

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

The minimum energy required to remove an other atom from the unipositive ion is called second ionization energy $\left(I_2\right)$

$$M_{(g)} + I_2 \longrightarrow M^{2+}_{(g)} + e^{-}$$

The second ionization energy (I_2) is greater than the first ionization energy (I_1) on removing he electron from an atom, the unipositive ion formed. In the ion effective nuclear charge increases i.e, he nuclear attraction increases on he remaining electrons.

$\therefore$ More energy is required o remove an electron from unipositive ion

Hence, he second ionization energy (I_2) is greater than he firs ionization energy (I_1)

Units of ionization energy: K cal./mole (or) KJ/mole

$1\ ev/atom = 1.602 \times 10^{-19} J$

$= 96.45\ kJ/mole$

$= 23.05\ K\,cal/mole$

Factors influencing ionization potential:

1) Atomic radius 2) Nuclear charge

3) Screening effect

1) Atomic radius: As the size of atom increases the distance between the nucleus and the outermost electrons increases hence the effective nuclear charge on the outermost electrons decreases due to this the energy required to remove the electrons also decreases.

$$\boxed{Atomic\ radius \propto \frac{1}{I.P}}$$

2) Nuclear charge: As the positive charge increases, valence electrons are strongly held by the nucleus, so more energy is required for the removal of valence electron.

$$\boxed{Nuclear\ charge \propto I.P}$$

3) Screening effect: In multi electron atoms, the electrons present in the inner orbits, screen the electrons present in the outermost orbit from the attraction of the nucleus is called screening effect.

As the number of electrons in the inner orbit increases, he screening effect also increases and hence nuclear attraction on the outermost electrons decreases.

$$\boxed{Screening\ effect \propto \frac{1}{I.P}}$$

Trend in a group:

In an group, atomic radius and screening effect increases, Ionization energy decreases.

Trend in a period:

In a period, as nuclear charge increases ionization energy increases.

3. Write an essay on the various factors, which effect the ionization enthalpy.

A. **1) Atomic radius:** As the size of atom increases the distance between the nucleus and the outermost electrons increases hence the effective nuclear charge on the outermost electrons decreases due to this the energy required to remove the electrons also decreases .

$$\boxed{Atomic\ radius \propto \frac{1}{I.P}}$$

2) **Nuclear charge**: As the positive charge increases, valence electrons are strongly held by the nucleus, so more energy is required for the removal of valence electron.

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3) **Screening effect**: In multi electron atoms, the electrons present in the inner orbits, screen the electrons present in he outermost orbit from the attraction of the nucleus is called screening effect.

As the number of electrons in the inner orbit increases, he screening effect also increases and hence nuclear attraction on the outermost electrons decreases.

$$\boxed{Screening\ effect \propto \frac{1}{I.P}}$$

4**) Extent of penetration of orbitals of valence electrons:** As the penetration power of orbitals increases, the IP of electrons in them increases. Order of penetration of orbitals: s > p > d > f

Order of IP is also he same order $\boxed{s > p > d > f}$

5) Half filled and completely filled orbitals:

Atoms with half-filled and completely filled sub shells are relatively more stable. So, the IP of these atoms are high.

e.g I: IP of N > IP of O

$2s^2\ 2p^1\ \ 2s^2\ 2p^4$

e.g II: IP of Mg > IP of Al

$3s^2$ $3s^2$ $3p^1$

4. Ionization enthalpy of elements is a periodic property. Explain.

A. Ionization enthalpy is the minimum energy, required to remove the most loosely bound electrons from an isolated gaseous atom to convert it into a gaseous ion.

$M_{(g)} + I_1 \xrightarrow{+} M^+_{(g)} + e^-$ where I_1 is he firs ionisation energy

Group: In a group, as atomic size increases from top to bottom, the nuclear attraction on the outermost electrons decreases. So ionisation enthalpy decreases down the group.

Period: In a period, as the atomic size decreases from left to right, the nuclear attraction increases on the outermost electrons. So ionisation enthalpy increases, from left to right in a period. However, there are some anomalies.

IP of Boron is less than that of Beryllium (in second period):

Be - $1s^2$ $2s^2$

B - $1s^2$ $2s^2$ $2p^1$

In beryllium, the outermost electron is paired up and is present in a s – orbitals, which can highly penetrate. Thus, it requires more amount of energy for removal.

The outermost electron in boron is loosely held and is present in a less penetrating p – orbital. Therefore, boron requires relatively less energy than that in beryllium.

IP of oxygen is less than that of nitrogen (in second period):

N - $1s^2$ $2s^2$ $2p^1{}_x$, $2p^1{}_y$, $2p^1{}_z$

O - $1s^2$ $2s^2$ $2p^2{}_x$, $2p^1{}_y$, $2p^1{}_z$

Nitrogen atom has a half-filled p – orbitals, which impart greater stability. This requires more energy for the removal of the loosely bound electron.

Oxygen atom has p – orbitals with paired up electrons and two orbitals have one electron each. The repulsions between the paired electrons results in lowering of ionisation energy. Thus, the IP of oxygen is less than that of nitrogen.

IP in transition elements: In a transitions series, the differentiating electron enters the

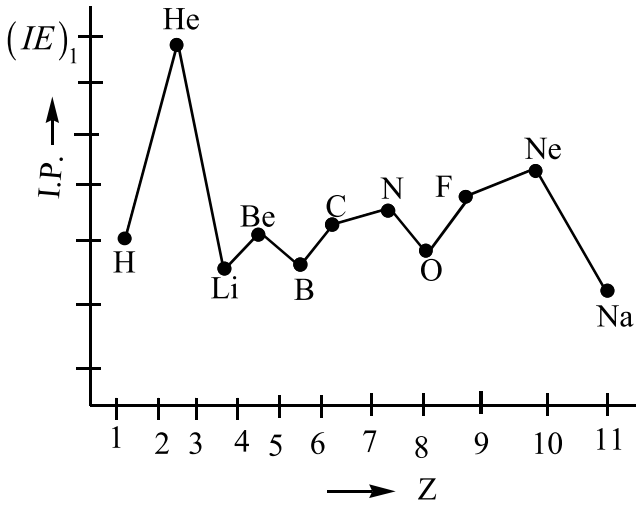
(n – 1)d sublevel. The outer electronic configuration is ns^2 or ns^1 is similar in all of them.

The size of the atoms in a series is almost the same. So the IP values are nearly the same.

Actually, IP increases very slowly in the series.

Thus in each period, the alkali metal has the lowest IP and the last element, inert gas has the highest IP.

Among all the elements, He has the highest IP while Cs has the least.



Atomic number Z (vs) Ionization potential (l)
For the first eleven elements (Drawn not to scale)