SESSION -1

AIM

✓ **To introduce history of development of periodic table**

INTRODUCTION

There are more than 115 elements known at present and it is very difficult to study the properties of all these elements separately. The basic object of classification is to arrange the facts regarding elements and their compounds in such a way so that we may have greatest control over their characteristics with least possible effort. The best classification would be one which puts together those elements which resemble in most respects and separates other. Attempts were made to classify the elements according to one property or the other.

HISTORICAL DEVELOPMENT OF PERIODIC TABLE

a) *Classification into metals and non-metals:*

Early chemist realized that there were two kinds of elements, i.e., **metals** and **non-metals** which can be distinguished from each other on the basis of a set of physical and chemical properties.

However, as there are a large number of metals and non-metals, whose properties are not related to one another in a particular set, such a classification is quite inadequate.

A certain elements exhibited metallic as well as non-metallic characters & were named metalloids.

b) Classification on the basis of Dobereiner's Triads:

J.W. Dobereiner, in 1817, attempted to classify elements in triad. A triad is a group of three elements arranged in the order of increasing atomic masses, such that the elements in the group have similar chemical and physical properties.

Statement: It states that when three chemically similar elements are arranged in increasing order of their atomic mass, the atomic mass of middle element of a triad is almost arithmetic mean of the other two elements.

From the Table, it is clear that the atomic mass of middle element is almost equal to the arithmetic mean of the atomic masses of first and last element of triads.

Limitation: A large number of elements cannot be grouped into triads

EXAMPLES OF DOBEREINER'S LAW OF TRIADS

Triad	**Calcium**	**Strontium**	**Barium**
Experimental atomic mass	40	88	137
Mean atomic mass		$\frac{40 + 137}{2} = \frac{177}{2} = 88.5$	
Triad	**Lithium**	**Sodium**	**Potassium**
Experimental atomic mass	7	23	39
Mean atomic mass		$\frac{7 + 39}{2} = \frac{46}{2} = 23$	

Triad	Chlorine	Bromine	Iodine
Experimental atomic mass	35.5	80	127
Mean atomic mass		$\frac{35.5 + 127}{2} = \frac{162.5}{2} = 81.25$	

c) *Classification on the basis of Newland's law of octaves:*

In 1864, Newland arranged the known elements in the increasing order of their atomic masses. He found, the elements with similar properties recurred each time, after every seven elements, like the repetition of musical note in an octave.

Law of octaves:

When the elements are arranged in the order of increasing atomic mass, the properties of eighth element (starting from a given element) are the repetition of the properties of first element.

Table shows the 17 elements arranged by Newland on the basis of his law.

In this table, if we start from Lithium (Li) then the eight elements is sodium (Na) and further eight elements is (K). The above elements have similar chemical and physical properties. Similarly, Be and Mg, B and Al, C and Si, etc., have similar chemical and physical properties.

1	**2**	**3**	**4**	**5**	**6**	**7**
H	**Li**	**Be**	**B**	**C**	**N**	**O**
8	**9**	**10**	**11**	**12**	**13**	**14**
F	**Na**	**Mg**	**Al**	**Si**	**P**	**S**
15	**16**	**17**				
Cl	**K**	**Ca**				

d) *Lothar- Meyer arrangement:*

In 1869, Lothar Meyer, a German Chemist, studied the physical properties such as atomic values, melting point and boiling point of various elements.

He plotted a graph between the atomic volumes (gram atomic weight divided by density) and atomic weights of the elements and observed that the elements with similar properties occupied similar position on the curve.

For example,

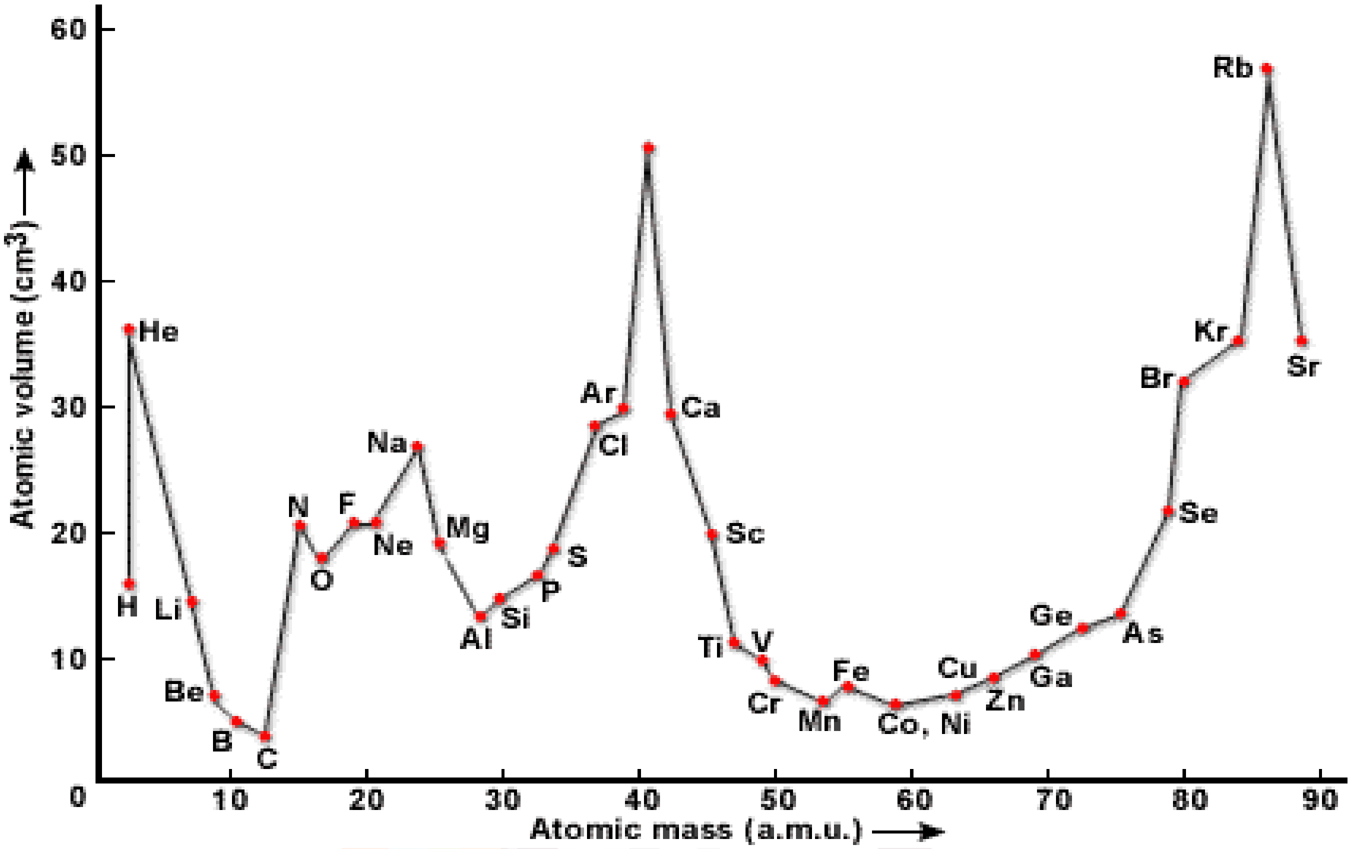
(i) The most strongly electropositive alkali metals (Li, Na, K, Rb and Cs) occupy the peaks on the curve.

(ii) The less strongly electropositive alkaline earth metals (Be, Mg, Ca, Sr and Ba) occupy the descending positions on the curve.

(iii) The most electronegative elements i.e., halogens (F, Cl, Br and I) occupy the ascending positions on the curve.

On the basis of these observations, Lothar Meyer proposed that the physical properties of the elements are a periodic function of their atomic weights. He arranged the then known elements in the tabular form in order of their increasing atomic weights.

It may be pointed out that the general pattern of the curve remained the same when atomic numbers were plotted in place of atomic weights as a result of later developments



Groups are designated as 0, I, II, III, IV, V, VI, VII and VIII. Except for groups 0 and VIII, each group is further divided into two sub-groups designated as A and B. The elements which lie on the left hand side of each group constitute sub-group A while those placed on the right hand side form sub-group B. This sub-division is made on the basis of the difference in their properties. Group VIII contains nine elements in three sets each containing three elements. However, group zero has no sub groups. It consists of only one vertical column of inert gases.

e) *Classification on the basis of Mendeleev's law:*

Mendeleev's periodic law:

According to this classification, the physical and chemical properties of elements are periodic functions of their atomic weights.

When the elements are arranged in order of their increasing atomic weight, elements with similar properties are repeated after certain regular intervals. While designing the periodic table he predicted the properties of some elements and left some gaps. The gaps were named as eka-boron, eka-aluminium and eka-silicon. Later on discovery these gaps are filled with the elements Scandium, Gallium and Germanium respectively. In his time inert gases were not known and zero group was added later on it the modified Mendeleev's periodic table. In this table there were 7 periods and 9 groups.

Groups are designated as 0, I, II, III, IV, V, VI, VII and VIII. Except for groups 0 and VIII, each group is further divided into two sub-groups designated as A and B. The elements which lie on the left hand side of each group constitute sub-group A while those placed on the right hand side form sub-group B. This sub-division is made on the basis of the difference in their properties. Group VIII contains nine elements in three sets each containing three elements. However, group zero has no sub groups. It consists of only one vertical column of inert gases.

Defects in Mendeleev's periodic table:

1) Anomalous position of hydrogen

Position of H is controversial in periodic table because it is placed in 1A whereas it resembles the properties of both groups IA and VII A.

2) Anomalous pair of elements

Anomalous pairs in atomic masses i.e. inversion of atomic masses

1) Ar^{40} K^{39} 2) Co^{60} Ni^{59} 3) $Te^{127.7}$ I^{127} 4) Th^{232} Pa^{231}

3) Position of isotopes

Isotopes of hydrogen i.e. protium, deuterium, and tritium with atomic weight 1, 2 and 3 should be placed separately.

4) Position of Lanthanides and actinides

No justification is given for their position.

CLASS EXERCISE

1] Eka-aluminium and Eka-silicon are known as

a) Gallium and germanium b) Aluminum and silicon

c) Hydrogen and silicon d) Nitrogen and magnesium

2] Which of the following group of elements is a Dobereiner's triad?

a) Os, Ir, Pd b) Ru, Rh, Pd c) Ru, Rh, Pt d) S, As, Te

SESSION - 2

AIM

✓ **To introduce modern periodic table**

THE MODERN PERIODIC TABLE

Moseley gave the modern periodic law which states that physical and chemical properties of the elements are the periodic functions of their atomic numbers. If the elements are arranged in order of their increasing atomic number, the elements with similar properties are repeated after certain regular intervals.

Salient feature of the long form of periodic table:

1) ***Periods:***

There are seven periods. The numbers of elements present in each period are as follows

1st Period – 2 elements (H and He) – Very short period

2nd Period – 8 elements (Li to Ne) – Short period

3rd Period – 8 elements (Na to Ar) – Short period

4th Period – 18 elements(K to Kr) – Long period

5th Period – 18 elements (Rb to Xe) – Long period

6th Period – 32 elements (Cs to Rn) – Very long period

7th Period – incomplete

2) Groups:

A vertical column in the periodic table is known as group. There are 18 groups in the long form of the periodic table. According to the new recommendations of International Union of Pure and Applied Chemistry (IUPAC), the groups are numbered from 1 to 18. Previously, these were numbered only from I to VIII as A and B groups. The elements with similar electronic configurations have similar properties and hence they have been placed together in the same group. The order of various groups in the periodic table; IA, IIA, IIIB, IVB, VB, VIB, VIIB, 3 columns making VIIIB, IB, IIB, IIIA, IVA, VA, VIA, VIIA and zero groups.

The first two groups on the extreme left and last six groups on the extreme right involve the filling of s-orbitals and p-orbitals, respectively. These groups represent the main groups of the periodic table and are numbered as 1, 2, 13, 14, 15, 16, 17 and 18. In the elements present in groups from 3 to 12, the electrons are filled in d-orbitals.

There are two more rows at the bottom of the periodic table. These rows consist of fourteen elements after lanthanum (Z = 57) and fourteen elements which follow actinium (Z = 89)

The elements in the first row, starting from cerium are called lanthanodis (or lanthanides) and the elements present in the second row starting from thorium are called actinoids (or actinides).

Advantages of Long form of periodic table:

1) The position of elements in the periodic table is governed by their electronic configuration.
2) This shows a gradiation in physical and chemical properties of elements and their compounds in a group and in a period.
3) The division of elements into *s, p, d*, and *f* - blocks has made the study simpler and has a logical explanation.
4) Based on their electronic configurations, the elements can be classified into inert gases, representative elements, transition elements and inner transition elements.
5) The properties of new elements can be predicted even before their actual discoveries.

Drawbacks of long form of periodic table:

1) **Position of hydrogen:** Hydrogen shows similarities with both alkali metals and halogens but it is placed with alkali metals.
2) **Position of He:** It is kept along with *p* - block elements.
3) **Position of lanthanides and actinides:** They are kept outside the main body of the periodic table.
4) The designation of sub groups as A and B has got no significance.

CLASS EXERCISE

1] The electronic configuration of an element is $1s^2 2s^2 2p^6 3s^2 3p^3$. What is the atomic number of the elements which is just below the above, element in the periodic table?

2] A transition element 'X' has the configuration $[Ar]3d^5$ in its +3 oxidation state. It is belonging to the same group.

a) VII B b) VI B c) I B d) VIII B

3] Which one of the following pairs of atomic numbers represents elements belonging to the same group?

a) 11, 20 b) 13, 31 c)10, 30 d) 14, 3

4] An element with atomic number greater than 16 and chemically similar to the element in the element with atomic number 10 is........

SESSION - 3

AIM

✓ **To understand classification of elements into various blocks**

CLASSIFICATION OF ELEMENTS INTO FOUR BLOCKS

The division is based upon the name of orbital which receives the last electron.

1) *s-Block Elements:*

Elements in which the last electron enters the s-orbital of their respective outermost shell are called s block elements. s-block has two groups namely group I (alkali metals) and group II (alkaline earth metal). Group I (Li, Na, K, Rb, Cs and Fr) has outer electronic configuration ns^1. Group II (Be, Mg, Ca, Sr, Ba and Ra) has outer general configuration ns^2.

General characteristics of s-block elements:

(i) They are soft metals with low melting point and boiling points.

(ii) They lose the valence (outermost) electrons readily to form +1 (alkali metals) and +2 ions (in case if alkaline earth metals).

(iii) They are very reactive metals.

(iv) The compounds of s block elements with the exception of Be are ionic predominantly

(v) They are strong reducing agents.

(vi) All are good conductors of heat and electricity.

2) *p-Block Elements:*

Elements in which the last electron enters any one of the three p orbitals of their respective outermost shell are called p block elements. p-subshell has three orbitals and each orbital can accommodate two electrons. There are six groups in p-block elements namely 13, 14, 15, 16, 17 and 18 (excluding He) group. The general outer electronic configuration is $ns^2\ np^{1-6}$ (excluding He), since Helium has $1s^2$ electronic configuration. The elements of s and p block collectively called as representative, elements or normal or main group elements.

General characteristics of p-block elements:

(i) p-block includes both metals and non-metals but the number of non-metals is much higher than that of metals.

(ii) They mostly form covalent compounds.

(iii) Most of them are electronegative and non-metals.

(iv) These are bad conductors of heat and electricity (exception: metals).

(v) Most of the non-metals have polyatomic molecules.

3) *d-Block Elements:*

Elements in which the last electron enters any one of the five d orbitals of their respective penultimate shells are called d block elements. Since d subshell has five orbitals and each one of them can accommodate two electrons. Ten vertical columns namely 3, 4, 5, 6, 7, 8, 9, 10, 11, 12 comprise ten groups of d block elements.

General outer shell electronic configuration of d block elements is $(n-1)d^{1-10}\ ns^{1-2}$. Since the properties of these elements are midway between s and p block elements, they are called as transition elements. All these elements are further classified into four series:

a) ***The first transition series:*** it forms a part of fourth period of long form of periodic table. It contains ten elements from $_{21}Sc$ to $_{30}Zn$ in which 3d orbitals are being progressively filled in.

b) ***The second transition series:*** It forms a part of fifth period also contain ten elements from yttrium to cadmium ($_{39}Y$ to $_{48}Cd$) in which 4d orbitals are being progressively filled in.

c) ***Third transition series:*** It forms a part of the sixth period also contain ten elements i.e. Lanthanum ($_{57}La$) and form Hafnium to mercury ($_{72}Hf$ – $_{80}Hg$) in which 5d orbitals are progressively filled in.

d) ***The fourth transition series:*** It forms a part of seventh period also contain ten elements i.e. actinium ($_{89}Ac$) and elements from rutherfordium ($_{104}Rf$), to ekamercury (Z = 112). In all these elements 6d orbitals are progressively filled in,

General characteristics of d block elements:

(i) They are hard, malleable and ductile metals with high melting point and boiling points.

(ii) They are good conductor of heat and electricity.

(iii) They show variable valency.

(iv) They form both ionic and covalent compound.

(v) They form colored compounds.

(vi) Their compounds are generally paramagnetic.

(vii) Most of the transition metals form alloys.

(viii) The metals and their compounds are used as catalysts.

(ix) They have high tendency to form complexes.

4) *f-Block Elements*:

Elements in which the last electron enters any one of the seven f-orbitals of their respective ante-penulatimate shells are called f-block elements. These are called rare earth elements since they occur scarcely in the earth's crust. General electronic configuration of f- block elements i.e. $(n-2)f^{1-14}\ (n-1)d^{0-1}ns^{2}$. They are true series of f block elements. Each containing 14 elements.

First series: ($_{58}Ce$ – $_{71}Lu$) which forms a part of the sixth period and collectively known as (Lanthanides).

Second series: ($_{90}Th$ – $_{103}Lr$) forms a part of incomplete seventh period are called collectively known as actinides.

General characteristics of f-block elements:

(i) They are heavy metals.

(ii) They have generally high melting points and boiling point.

(iii) They show variable valency.

(iv) Their compounds are generally colored.

(v) They have high tendency to form complexes.

(vi) The common oxidation state is +3

(vii) All the actinides are radioactive elements.

CLASS EXERCISE

1] The element having electronic configuration [Kr] $4d^{10}$, $4f^{14}$, $5s^{2}$, $5p^{6}$, $5d^{1}$, $6s^{2}$ belongs to

a) s-Block b) p-Block c) d-Block d) f-Block

2] Which of the following represents the electronic configuration of d-block elements?

a) $(n-1)\ s^{2}\ nd^{1-10}$

b) $(n-1)d^{1-10}\ ns^{1-2}$

c) $(n-1)\ d^{1-10}ns^{2}p^{4}$

d) $(n-1)\ p^{4}\ ns^{2}$

3] The element whose electronic configuration is, $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $3d^{10}$, $4s^2$ is a

a) Metal

b) Non-metal

c) Noble gas

d) Metalloid

4] The element with Z = 106, will be an element of

a) s-element b) p-block c) d-block d) f-block

SESSION – 5

AIM

- ✓ **To understand classification of elements into various types based on electronic configuration and chemical properties**

PERIODIC PROPERTIES:

When elements are arranged in the increasing order of their atomic numbers, similar properties are repeated at regular intervals. Such properties are called periodic properties and this phenomenon is called periodicity and is due to the repetition of similar outer electronic configuration at regular intervals. These properties include atomic size, atomic volume, ionization potential, electron affinity electro positivity, melting and boiling points, nature of oxides etc.

ATOMIC RADIUS (ATOMIC SIZE):

Concept:

The size of an atom is also known as atomic size. The size of an atom (or atomic size) is the distance between the centre of the nucleus and outermost electron shell of an isolated atom. In other words, the size of atom (or atomic size) refers to the radius of atom. Thus, the size of an atom is indicated by writing its radius called "atomic radius". It is expressed in Å, nm, pm.

The exact value of atomic size cannot be determined as the probability of finding an electron even at large distances from the nucleus never becomes zero i.e. orbitals do not have sharp boundaries. Based on nature of bonding, atomic radius is expressed in three different ways i.e. crystal radius, covalent radius and Vander Waals radius.

Crystal radius:

The half of the distance between the centers of the closest adjacent nuclei of atoms in metallic crystal is called crystal radius or atomic radius.

Ex: The distance between two adjacent potassium atoms in solid potassium is 4.62 Å themetallic radius (Crystal radius) of potassium is = 2.31 Å.

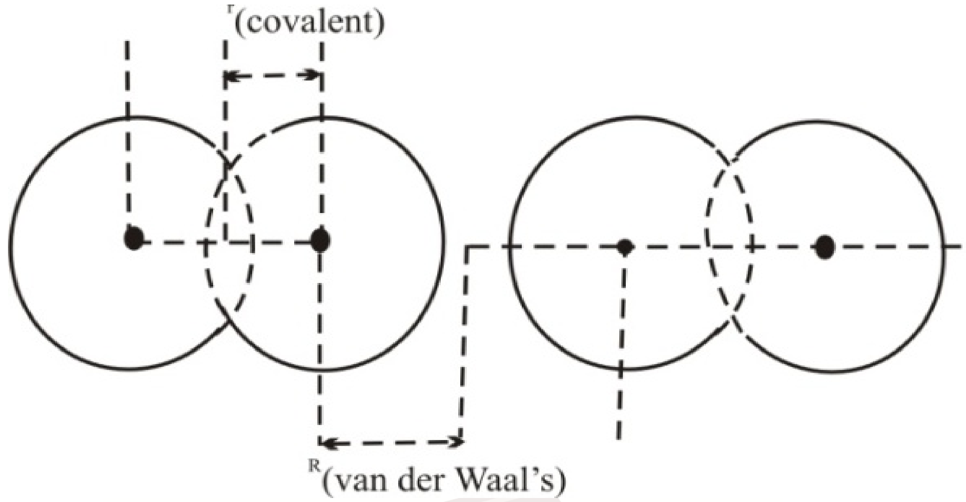
Covalent radius:

The half of the distance between the nuclei of two atoms held together by a covalent bond is called covalent radius.

Ex: The intern clear distance (bond length) of chlorine molecule is 1.98 Å. Therefore, covalent radius of chlorine is = 0.99Å.

Vander Wall's Radius:

The half of the distance between the centers of two adjacent atoms of two different molecules which are very close to each other is called Vander walls radius.



The magnitude of the Vander Waal's radius depends upon the packing of molecules of the element in solid state. For example the distance between adjacent chlorine atoms of the two neighboring molecules in the solid state is 3.6Å. Therefore the Vander Waals' radius of chlorine is $\frac{3.6}{2}$ = 1.8 Å.

Vander Waal's radius of an element is always larger than its covalent radius.

A comparison of three types of radius shows that covalent radius of an atom is shortest while

Vander Waal's radius is the longest.

Vander Waal's radius > Crystal radius > Covalent radius

Vander Waals' forces are weak, thus the distances between the atoms held by these forces are much larger than held by covalent bonding. A covalent bond is formed by overlap of two half-filled atomic orbitals, thus a part of the electron cloud becomes common between two atoms. Due to these reasons Vander Waals' radius is always higher than covalent radius. In case of a crystal, the valence electrons of the atoms are mobile; therefore, they are only weakly attracted. The metallic forces are thus weaker than covalent forces. On account of this, crystal (metallic) radius is larger than covalent radius.

[Vander Waal's forces < Metallic forces < Covalent forces]

Note: For simplicity, we may use the term atomic radius for both covalent and crystal (metallic) radius depending on whether the element is a non-metal or a metal. However atomic radii of inert gases are usually expressed in terms of Vander Waals radii as they do not form chemical compounds.

Factor affecting atomic radius:

1) ***Nuclear charge:*** Greater the nuclear charge more is the attraction for electrons and hence less is the atomic size.
2) ***Number of shell:*** Greater the number of inner shell more is the size of the atom.

Periodic trends in atomic size

In period: In a period atomic size decreases from left to right because the electrons are added to same energy level and the nuclear attraction on the outermost electrons increases.

In group: In a group atomic size increases from top to bottom because the electrons are added to new shells and thus the nuclear attraction on the outermost electrons decreases.

In each period, the highest atomic size element is alkali metal and the lowest atomic size element is halogen. The atomic size of zero group elements is higher than halogen because the atoms of inert gases are attracted by Vander Waals forces.

In a transition series, the atomic size decreases slightly as we move from left to right due to the differentiating electron enters into (n–1) *d*-sub shell and *d*-electrons screens ns electrons from the nucleus.

The very low decrease of atomic size in lanthanides is called lanthanide contraction. It is due to the differentiating electron enters into (n–2) f sub shell and poor shielding effect of f orbitals.

Note: The variation of atomic size in d & f - block elements will be discussed to a greater extent in 12th standard.

Ionic Radius:

It is the radius of the cationic or anionic sphere, which is present in the crystal of an ionic compound, surrounded by oppositely charged ions.

The positive ion is smaller in size than the neutral atom because the number of electrons decreases and nuclear attraction increases on the outermost electrons in positive ion.

The ionic size decreases as the number of positive charge increases

$M > M^{+} > M^{+2} > M^{+3}$ (order of the size of the ion)

The negative ion is larger in size than neutral atom because the number of electrons increases and nuclear attraction decreases on the outermost electrons in negative ion.

The ionic size increases as the number of negative charge increases

$X < X < X^{-2} < X^{-3}$ (order of the size of the ions)

Ex: $I^- > I > I^+$

Isoelectronic series:

The series of ions having same number of electrons but different nuclear charge is called electronic series. In an iso-electronic series higher the nuclear charge, smaller is the size.

Ex: $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

CLASS EXERCISE

1] Which of the following sequence contains atomic numbers of only representative elements?

a) 55, 12, 48, 53 b)13, 93, 33, 83

c) 3, 33, 53, 87 d) 22, 33, 55, 66

2] Predict the position of the element in the periodic table satisfying the configuration $(n-1)d^1ns^2$ for n = 4 and $(n-2)f^7$ $(n-1)d^1ns^2$ for n = 6?

3] **Assertion:** Lanthanides are rare earths

Reason: Lanthanides belongs to 6th period and IIIB group

4] **Assertion:** Inner transition elements have similar physical and chemical properties

Reason: The last two shells in then have similar electronic configuration

SESSION – 6

AIM

✓ **To introduce Ionization Potential**

IONIZATION POTENTIAL:

Concept:

The minimum amount of energy required to remove an electron from the outer most orbit of a gaseous, neutral, isolated atom present in its ground state is called first Ionization potential.

$M_{(g)} + 1P_1 \rightarrow M^+_{(g)} + \bar{e}$

The amount of energy required to remove an electron from a gaseous unipositive ion is called second ionization potential.

$M^+(g) + 1P_2 \rightarrow M^{+2}(g) + \bar{e}$

The units of ionization potential are K.cal/ more or ev/atom or kJ/mole.

I ev/atom = 23.06 K.cal /mole = 96.45 KJ/mole.

For a particular atom, $IE_1 < IE_2 < IE_3 <$ ....

This is because the removal of one electron from a neutral gaseous atom makes the atom positively charged. The positively charged ion contains one electron less than the number of protons in the nucleus. This results in the increase of electrostatic force of attraction between the nucleus and the remaining electrons in the ion. Thus, the nucleus now holds the outer electrons more strongly. Hence, more energy is required to knock out the second electron.

Factors affecting the magnitude of ionisation potential:

Ionization potential of an element depends on the following factors:

1) ***Nuclear charge:*** The greater the charge on nucleus of an atom, the more difficult it is to remove an electron from the atom and hence greater is the value of ionization potential.
2) ***Atomic Size:*** With the increase in atomic size, the ionization potential decreases. This is because of the less attraction between the nucleus and outer-most electron.
3) ***Shielding Effect of the Electrons:*** In a multi-electron system, the valence electron gets shielded by the inner electrons; as a result, it experiences less attraction from the nucleus. This is known as shielding effect. The larger the number of inner-shell electrons greater is the shielding effect and smaller is the force of attraction between the nucleus and valence electron and thus ionization potential decreases.
4) ***Penetration Effect:*** For a given shell, the degree of penetration of orbitals decreases in the order s > p > d > f. Thus other factors being equal, s-electron will be the most tightly held and the ease of removal of the electrons increase in the order s < p < d < f. The ionization potential will also decrease in this order.
5) ***Electronic Configuration:*** Because of the relatively more stability of completely filled (p^6, d^{10} and f^{14}) and half-filled (p^3, d^5 and f^7) configuration, removal of an electron from these configurations is more difficult than from other configurations. Hence atoms or ions' having any of these stable configurations has an exceptionally high value of ionization potential.

Example:

Be($2s^2$) and N($2s^2 2p^3$) having completely filled and half-filled orbitals respectively have higher ionization potential than expected.

Trends in ionisation potential:

In Period: In a period, ionization potential in decreases from left to right, due to decrease of atomic size and increase of nuclear attraction on the outer most electrons.

In Group: In a group, ionization potential decreases from top to bottom, due to increase of atomic size and decrease of nuclear attraction on the outer most electrons.

Note: In 2nd period L i < Be > B < C < N > O < F < Ne

The ionization potential of Beryllium (Be) is greater than Boron (B). It is due to stable configuration ($1s^2 2s^2$) and s- sub shell is more penetrate in Beryllium.

The ionization potential of nitrogen is greater than oxygen. It is due to stable half-filled configuration in nitrogen ($1s^2 2s^2 2p^3$).

In each period, the lowest ionisation potential element is alkali metal and the highest ionization potential element is noble gas or inert gas.

CLASS EXERCISE

1] Identify the correct order in which the I.P. of the following element increases

I) Be II) B III) N IV) O

a) II, I, IV, III b) I, II, III, IV c) II, I, III, IV d) III, IV, I, II

2] The first I.P. value is probably smallest for the atom with electronic configuration

a) ns^2np^3 b) ns^2np^1 c) ns^2np^2 d) ns^2np^4

3] The order of 2nd I.P. of C, N, O,F is ....

4] Which of the following electronic configurations is expected to have maximum difference in 2nd and 3rd I.P?

a) $1s^22s^2\ 2p^6$ b) $1s^22s^2\ 2p^63s^23p^1$

c) $1s^22s^2\ 2p^63s^2$ d) $1s^22s^2\ 2p^63s^23p^3$

5] The correct order of I.P, of the following element is

a) S > P > N > O b) N > O > S > P

c) N > P > O > S d) N > O > P > S

6] The I.P. in 'ev' of nitrogen and oxygen atoms are respectively

a) 14.6, 13.6 b) 13.6, 14.6 c) 13.6, 14.6 d) 14.6, 14.6

7] Which of the following is correct w.r.t.1st and 2nd I.P. of Na and Mg?

a) $I_{Na} = I_{Mg}$ b) $I_{Mg} = II_{Na}$ c) $II_{Mg} > II_{Na}$ d) $II_{Na} > II_{Mg}$

8] The first four I.P. of four consecutive elements in the second period of the periodic table are 8.3, 11.3, 14.5 and 13.6 eV respectively. Which one of the following is the $I.P_1$ of nitrogen?

a) 13.6 b) 11.3 c) 14.5 d) 8.3

9] The first four ionization potentials of an element are 300, 620, 4300, 6000 kJ respectively. The most probable outer electronic configuration of the element could be

a) ns^2np^3 b) ns^2 c) ns^2np^1 d) ns^2np^6

10] One mole of Mg in vapour phase absorbs 1250 kJ of energy. If the ionization energies of Mg are I_1 = 750 kJ and I_2 = 1500 kJ, the final composition of the mixture will be

a) 59% Mg^+ and 41% Mg^{2+} b) 49% Mg^+ and 51% Mg^{2+}

c) 67% Mg^+ and 33% Mg^{2+} d) 29% Mg^+ and 71% Mg^{2+}

SESSION – 7

AIM

✓ **To introduce electron affinity**

ELECTRON AFFINITY

Concept:

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

Since the electron adds up in the outermost orbit, energy is given out. Therefore, **electron affinity** is associated with an exothermic process.

$A(g) + e^- \rightarrow A^-(g); \Delta H = -E_n$

When one electron adds up to a neutral atom, it gets converted to a uni negative ion and energy is released. On adding one more electron to the mono negative anion, there is repulsion between the negatively charged electron and anion. In order to counteract the repulsive forces, energy has to be provided to the system. Therefore, the value of the second electron affinity is positive.

$A^-(g) + e^- \rightarrow A^{-2}(g) ; \Delta H = + E_n$

The units of electron affinity are K.cal/mole or KJ/mole or eV/atom.

Thus instead of electron affinity, electron gain enthalpy is more appropriate word.

The electron gain enthalpy is the energy change that occurs for the process of adding an electron to a gaseous isolated atom to convert it into a negative ion, i.e., to form a monovalent anion. It is denoted by Δ_{eg} H or E_A. The process may be represented as:

A (g) + e → A^- (g); Δ_{eg} H = Δ H

Factors influencing the magnitude of electron affinity:

1) ***Atomic Size or Atomic Radius:*** When the size or radius of an atom increases, the electron entering the outermost orbit is more weakly attracted by the nucleus and the value of electron affinity is lower.
2) ***Effective Nuclear charge:*** When effective nuclear charge is more then. Then the atom can easily gain an electron and higher the electron affinity.
3) ***Electronic Configuration:*** The stability of the configuration having fully-filled orbitals (p^6, d^{10}, f^{14}) and half-filled orbital (p^3, d^5, f^7) is relatively higher than that of other configurations. Hence such type of atoms have less tendency to gain an electron, therefore their electron affinity values will be very low.

Periodic trends in electron affinity:

In a period**:** In a period, electron affinity increases from left to right, due to decrease of atomic size and increase of nuclear attraction on the outermost electrons

In a group: In a group, electron affinity decreases from top to bottom due to increase of atomic size and decreases of nuclear attraction on the outer most electrons.

Exceptions:

(i) The electron affinities of second period elements are lower than the corresponding elements of the third period. This is due to their smaller size and high electron density which makes the addition of electron slightly less favorable due to increased electron-electron repulsions.

e.g., the electron affinity of fluorine is less than chlorine due to small size and high electron repulsions in 2p sub shell of fluorine. Similarly, the electron affinity of nitrogen is less than that of phosphorus and the electron affinity of oxygen is less than that of sulphur.

(ii) The lower electron affinities of nitrogen and phosphorous than carbon and silicon respectively are due to the former having a stable half-filled system and addition of an electron will give a less stable arrangement.

(iii) Be and Mg has positive electron gain enthalpy since both have completely filled *s*-orbitals (Be - $2s^2$, Mg - $3s^2$). The additional electron will have to enter the high energy 2p and 3p orbitals respectively.

(iv) In case of inert gases in which the ns and *np* orbitals are completely filled, the electron gain enthalpy is positive.

CLASS EXERCISE

1] I.P. of K would be numerically equal to

a) E.A. of Ar
b) electro negativity of K
c) E.A. of K^+
d) E.A. of K

2] Which of the following process requires absorption of energy?

a) $F + e^- \rightarrow F^-$
b) $Cl + e^- \rightarrow Cl^-$
c) $O^- + e^- \rightarrow O^{2-}$
d) $H + e^- \rightarrow H^-$

3] The correct order of electron affinities of B, C, N, O is

a) O > C > N > B
b) B > N > C > O
c) O > C > B > N
d) O > B > C > N

4] Point out the wrong statement, in a given period of the periodic table, the s-block elements has, in general a lower value of

a) Electronegativity
b) Atomic radius
c) Ionization energy
d) Electron affinity

5] The formula of the oxide ion, O2– (g), from oxygen atom requires first an exothermic and then
an endothermic step as shown below:

$O (g) + e^- \longrightarrow O^- (g); \Delta H^0 = -141 kJ\ mol^{-1}$

$O^- (g) + e^- \longrightarrow O^{2-} (g); \Delta H^0 = +780 kJ\ mol^{-1}$

Thus process of formation ofO^{2-} in gas phase is unfavourble even though O^{2-} is isoelectronic with neon. It is due to the fact that,

a) Oxygen is more electronegative.

b) Addition of electron in oxygen results in larger size of the ion.

c) Electron repulsion outweighs the stability gained by achieving noble gas configuration.

d) O^- ion has comparatively smaller size than oxygen atom

SESSION – 9

AIM

- ✓ **To introduce remaining periodic properties and diagonal relationship.**
- ✓ **To discuss additional objectives**

METALLIC AND NONMETALLIC CHARACTER:

Variation of metallic and non–metallic character:

In a group, metallic character increases and non-metallic character decreases from top to bottom, due to increase of electropositive character and decrease of electronegative character.

In a period, metallic character decreases and non-metallic character increases from left to right, due to decrease of electropositive character and increase of electronegative character.

VALENCY OR OXIDATION STATE:

Valence is the most characteristic property of the elements and can be understood in terms of their electronic configurations or more precisely on the number of electrons present in the outmost shell of the atom.

The electrons present in the outermost shell of an atom are called valence electrons and the number of these electrons determine the valence or the valency of the atom. It is because of this reason that the outermost shell is also called the valence shell of the atom and the orbitals present in the valence shell are called valence orbitals

In case of representative elements, the valence of an atom is generally equal to either the number of valence electrons (s- and p-block elements) or equal to eight minus the number of valence electrons.

Group	1	2	13	14	15	16	17	18
Number of valence electrons	1	2	3	4	5	6	7	8
Valency	1	2	3	4	3,5	2,6	1,7	0

In contrast, transition and inner transition elements, exhibit variable valence due to involvement of not only the valence electrons but d- or f- electrons as well. However, their most common valency are 2 and 3.

Variation of Valency:

Variation along a period: As we move across a period from left to right , the number of valence electrons increases from 1 to 8. But the valency of elements, *w.r.t.* H first increases from 1 to 4 and then decreases to zero.

Compound	NaH	MgH_2	AlH_3	SiH_4	PH_3	H_2S	HCl
Valency *w.r.t* 'H'	1	2	3	4	3	2	1

With respect to oxygen, valency increases from 1 to 7 across the period.

Compound	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Valency *w.r.t* 'O'	1	2	3	4	5	6	7

Variation within a group: when we move down the group, the number of valence electrons remain the same, therefore, all the elements in a group exhibit the same valence. For example, all the elements of group 1 (alkali metals) have valence one while all the elements of group 2 (alkaline earth metals exhibit a valence of *two.* Noble gases present in group 18 are *zero valent*, i.e., their valence is *zero* since these elements are chemically inert.

BASIC NATURE OF OXIDES:

Concept:

Metallic oxides dissolve in water give basic solutions

Ex. $Na_2O + H_2O \rightarrow 2\,NaOH$; $CaO + H_2O \rightarrow Ca(OH)_2$

Non-metallic oxides dissolve in water give acidic solutions.

Ex. $CO_2 + H_2O \rightarrow H_2CO_3$; $SO_3 + H_2O \rightarrow H_2SO_4$

The oxides which possess both acidic and basic properties are called amphoteric oxides.

Ex. Al_2O_3BeO, ZnO, SnO_2, As_2O_3, Sb_2O_3, etc.

Variation:

In a period the metallic nature decreases and non-metallic nature increases. Hence the basic nature of oxides decreases and acidic nature of oxides increases.

Ex: In third period:

Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Strongly basic	Weakly basic	Amphoteric	Weakly acidic	Weakly acidic	Strongly acidic	Strongly acidic

In a group, the metallic nature increases and non-metallic nature decreases. Hence the basic nature of oxides increases and acidic nature of oxides decreases.

DIAGONAL RELATIONSHIP:

In the periodic table the element present in a group of second period resembles with the element present in the third period of the next higher group This is called diagonal relationship Li and Mg, Be and Al, B and Si exhibit diagonal relationship

	IA	IIA	IIIA	IVA
2nd Period	Li	Be	B	C
3rd Period	Na	Mg	Al	Si

The cause of the diagonal relationship is the identical radii, polarizing power (charge/radius^2) electro-negativity and electropositive character.

CLASS EXERCISE

1] Which of the following is the most basic oxide?

a) SeO_2 b) Al_2O_3 c) Sb_2O_3 d) Bi_2O_3

2] Chloride of an element A gives neutral solution in water. In the periodic table A belongs to

a) First group b) Third group

c) Fifth group d) First transition series

3] Which of the following pairs shows diagonal relationship?

a) Li and Mg b) Na and K c) Zn and Cd d) Li and Be

4] Which of the following pairs of atomic numbers can have diagonal relationship?

a) 3, 11 b) 3, 12 c) 6, 15 d) 11, 15

5] Electro negativity of beryllium is approximately equal to that of

a) aluminum b) boron c) magnesiumd) sodium

KEY

SESSION 1

CLASS EXERCISE:

1) a 2) b

SESSION 2

CLASS EXERCISE:

1) 33 2) d 3) b 4) Any noble gas after Ne

SESSION 3

CLASS EXERCISE:

1) c 2) b 3) a 4) c

SESSION 4

CLASS EXERCISE:

1) c

2) 4^{th} period, 3^{rd} group; 6^{th} period, 3^{rd} group

3) b 4) a

SESSION 5

CLASS EXERCISE:

1) c 2) a 3) d 4) b

SESSION 6

CLASS EXERCISE:

1) a 2) a 3) O > F > N > C 4) c 5) d 6) a

7) d 8) c 9) b 10) c

SESSION 7

CLASS EXERCISE

1) c 2) c 3) d

SESSION 8

CLASS EXERCISE

1) 3 2) b 3) b 4) c 5) b

SESSION 9

CLASS EXERCISE

1) d 2) a 3) a 4) b 5) a