CHAPTER-3

CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES OF ELEMENTS

Mandeleev's Periodic Law:- The properties of the elements are the periodic function of their atomic masses.

Moseley, the English physicist showed that atomic number is more fundamental property of an element than its atomic mass. Therefore, the position of an element in the periodic table depends on its atomic number than its atomic mass.

Modern Periodic Law: The physical and chemical properties of elements are the periodic functions of their atomic numbers.

Types of Elements: s-, p-, d- and f- blocks.

MAIN GROUP ELEMENTS/ REPRESENTATIVE ELEMENTS:

The s- and p- block elements are called main group elements or representative elements.

s- block elements: Group-1 (Alkali metals) and Group-2 elements (Alkaline earth metals) which respectively have ns^1 and ns^2 outermost electronic configurations.

p- Block elements: They belongs to group- 13 to 18. The outer most electronic configuration is ns^2 np^{1-6}. He ($1s^2$) is a s- block element but is positioned with the group 18 elements (ns^2 np^6) because it has completely filled valence shell and as a result, exhibits properties characteristic of other noble gases.

d- block elements (Transition elements) are the elements of group 3 to 12 having outer electronic configuration (n-1) d^{1-10} ns^{1-2}. Four transition series are 3d, 4d, 5d and 6d. The 6d- series is incomplete. Atomic radius generally decreases across a period and increases as we descend the group.

f-**Block elements (Inner- transition Series)**

Lanthanoids charecterised by the filling of4 f-orbitals, are the elements following lanthanum from $_{58}Ce$ to $_{71}Lu$. Actinoids characterised by filling of 5f-orbitals, are the elements following actinium from $_{70}{}^{Th}$ to $_{103}Lr$. Characteristic outer electronic configuration is (n-2) f^{1-14} (n-1) d^{0-1} ns^2.

Noble Gases: The gaseous elements of group 18 are called noble gases. The general outermost electronic configuration of noble gases (except He) is ns^2 np^6. He exceptionally has $1s^2$ configuration. Thus the outermost shell of noble gases is completely filled.

PERIODICITY: The repetition of similar properties after regular intervals is called periodicity.

Cause of Periodicity: The properties of elements are the periodic repetition of similar electronic configuration of elements as the atomic number increases.

ATOMIC PROPERTIES: The physical characteristics of the atom of an element are called atomic properties. The properties such as atomic radius, ionic radius, ionisation energy, electro-negativity, electron affinity and valence etc., called atomic properties.

ATOMIC RADIUS- The distance from the centre of the nucleus to the outermost shell of the electrons in the atom of any element is called its atomic radius.

Periodicity- (a) In period- Atomic radius of elements decreases from left to right in a period.

(b) In Group- Atomic radius of elements increases on moving top to bottom in a group.

COVALENT RADIUS- Half the inter-nuclear distance between two similar atoms of any element which are covalently bonded to each other by a single covalent bond is called covalent radius.

VAN DER WAALS' RADIUS: Half the inter-nuclear separation between two similar adjacent atoms belonging to the two neighbouring molecules of the same substance in the solid state is called the van der waals'radius of that atom.

METALLIC RADIUS: Half the distance between the nuclei of the two adjacent metal atoms in a close packed lattice of the metal is called its metallic radius.

Van der Waals'radius > Metallic radius > Covalent radius

IONIC RADIUS: The effective distance from the centre of the nucleus of an ion upto which it has an influence on its electron cloud is called its ionic radius.

A cation is smaller but the anion is larger than the parent atom. In case of iso-electronic species, the cation with greater positive charge has smaller radius but anion with greater negative charge has the larger radii.

IONISATION ENTHALPY: The ionisation enthalpy is the molar enthalpy change accompanying the removal of an electron from a gaseous phase atom or ion in its ground state. Thus enthalpy change for the reaction; $M_{(g)} \rightarrow M^{+}_{(g)} + e^{-}$

Is the ionisation enthalpy of the element M. Like ionisation energies for successive ionisation, the successive ionisation enthalpy may also be termed as 2^{nd} ionisation enthalpy ($\Delta_r H_2$), third ionisation enthalpy ($\Delta_r H_3$) etc. The term ionisation enthalpy is taken for the first ionisation enthalpy, ($\Delta_r H_1$) is expressed in kg mol^{-} or in eV.

Periodicity:

i) Generally the ionisation enthalpies follow the order (there are few exceptions):

$(\Delta_r H_1) < (\Delta_r H_2) < (\Delta_r H_3)$

ii) The ionisation enthalpy decreases on moving top to bottom in a group.

iii) The ionisation enthalpy increases on moving from left to right in a period.

ELECTRON GAIN ENTHALPY: The electron gain enthalpy (($\Delta_{eg}H$) is the molar enthalpy change when an isolated gaseous atom or ion in its ground state adds an electron to form the corresponding anion thus the enthalpy change for the reaction; $X_{(g)} + e^{-} \rightarrow X^{-}_{(g)}$

Is called the electron gain enthalpy (Δ_{eg} H) of the element X. TheΔ_{eg} H may be positive or negative.

The successive values for the addition of second, third etc. Electron, these are called second, third etc. electron gain enthalpies. For example,

$X_{(g)} + e^{-} \rightarrow X^{-}_{(g)}$ $\Delta H = \Delta_{eg} H_1$ is called first electron gain enthalpy

$X^{-}_{(g)} + e^{-} \rightarrow X^{2-}_{(g)}$ $\Delta H = \Delta_{eg} H_2$ is called second electron gain enthalpy

$X^{2-}_{(g)} + e^{-} \rightarrow X^{3-}_{(g)}$ $\Delta H = \Delta_{eg} H_3$ is called third electron gain enthalpy Usually the term electron gain enthalpy ($\Delta_{eg}H$) means the first electron gain enthalpy.

Periodicity:

(i) In period- The electron gain enthalpy increases from left to right in a period.
(ii) In group- The electron gain enthalpy decreases from top to bottom in a group.

ELECTRONEGATIVITY: "The relative tendency of an atom in a molecule to attract the shared pair of electrons towards itself is termed as its electro-negativity."

Periodicity:

(i) In period- The electro-negativity increases from left to right in a period.
(ii) In group- The electro-negativity decreases from top to bottom in a group.

VALENCE ELECTRONS: The electrons present in outermost shell are called as valence electron. Because the electrons in the outermost shell determine the valency of an element.

VALENCY OF AN ELEMENT: The number of hydrogen or halogen atom or double the number of oxygen atom, which combin with one atom of the element is taken as its valency. According to the electronic concept of valency, " the number of electrons which an atom loses or gains or shares with other atom to attain the noble gas configuration is termed as its valency."

Periodicity:

(i) In period- The valency first increases then decreases from left to right in a period.
(ii) In group- The valency remains constant from top to bottom in a group.

ELECTROPOSITIVE OR METALLIC CHARACTER: The tendency of an element to lose electrons and forms positive ions (cations) is called electropositive or metallic character. The elements having lower ionisation energies have higher tendency to lose electrons, thus they are electropositive or metallic in their behaviour.

Alkali metals are the most highly electropositive elements.

Periodicity: In period- The electropositive or metallic characters decreases from left to right in a period.

In group- The electropositive or metallic characters increases from top to bottom in a group.

ELECTRO-NEGATIVE OR NON- METALLIC CHARACTERS: the tendency of an element to accept electrons to form an anion is called its non metallic or electronegative character. The elements having high electro-negativity have higher tendency to gain electrons and forms anion. So, the elements in the upper right hand portion of the periodic table are electro-negative or non-metallic in nature.

Periodicity:

(i) In period- The electro-negative or non- metallic characters increases from left to right in a period.
(ii) In group- The electro-negative or non-metallic characters decreases from top to bottom in a group.

REACTIVITY OF METALS:

Periodicity:

(i) In period- The tendency of an element to lose electrons decreases in a period. So the reactivity of metals decreases from left to right in a period.
(ii) In group- The tendency of an element to lose electrons increases in a period. So the reactivity of metals increases from top to bottom in a group.

REACTIVITY OF NON- METALS:

(i) In period- The tendency of an element to gain electrons increases in a period. So the reactivity of non-metals increases from left to right in a period.
(ii) In group- The tendency of an element to gain electrons decreases in a group. So the reactivity of non-metals increases from top to bottom in a group.

SOLUBILITY OF ALKALI METALS CARBONATES AND BICARBONATES:

PERIODICITY IN GROUP: The solubility of alkali metal carbonates and bicarbonates in water increases down the group (From Lithium to Caesium).

SOLUBILITY OF ALKALINE EARTH METAL HYDROXIDES AND SULPHATES:

PERIODICITY IN GROUP: The solubility of alkaline earth metal hydroxide and sulphates in water increases down the group (From Beryllium to Barium).

BASIC STRENGTH OF ALKALINE EARTH METAL HYDROXIDES:

PERIODICITY IN GROUP: The basic strength of alkaline earth metal hydroxide in water increases down the group (From Beryllium to Barium), i.e.,

$$Be(OH)_2 < Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$$

$$\xrightarrow{\hspace{8em}}$$

Basic strength increases

THERMAL STABILITY OF CARBONATES OF ALKALI AND ALKALINE EARTH METALS:

Except lithium carbonate, ($LiCO_3$), the carbonates of all other alkali metals are stable towards heat, i.e., carbonates of alkali metals (except $LiCO_3$) do not decompose on heating. $LiCO_3$ decomposes on heating to give lithium oxide ($LiCO_3$).

The carbonates of alkaline earth metals are relatively less stable. On heating, they decompose to give corresponding oxide and CO_2 gas. The decomposition temperature for alkaline earth metal carbonates increases as we go down the group.

Anomalous Properties of Second Period Elements
Their anomalous behaviour is attributed to their small size, large charge/radius ratio, high electro negativity, non- availability of d- orbitals in their valence shell. the first member of each group of p-Block elements displays greater ability to form pp-pp multiple bonds to itself (e.g. C=C, C≡C O=O, N≡N) and to other second period elements (e.g. C=O, C≡N, N=O) compared to subsequent member of the group.

ONE MARK QUESTIONS

Q1. Select the species which are iso-electronic (same number of electron) with each other.

(1) Ne (2) Cl^- (3) Ca^{2+} (4) Rb^+

Ans-The Cl^- and Ca^{2+}. Both have 18 e^- each.

Q.2. What the elements of a group have common among them?

Ans- They have same number of electrons in the valence shell.

Q.3. What the s- and p- block elements are collectively called?

Ans- Representative elements.

Q.4. Define atomic radius.

Ans- The one-half the distance between the nuclei of two covalently bonded atoms of the same element in a molecule is called as atomic radius.

Q.5. State the modern periodic law.

Ans- The physical and chemical properties of the elements are the periodic function of their atomic numbers.

Q.6. Name the groups of elements classified as s-, p- and d- blocks.

Ans- s- block= 1,2 (including He), p- block= 13 to 18 (except He), d- block= 3 to 12.

Q.7. Define the term ionisation enthalpy.

Ans- The energy required to remove the outer most electron from the valence shell of an isolated gaseous atom is called as ionisation enthalpy.

Q.8.In how many groups and periods the elements in modern periodic table are classified?

Ans- In 18 groups and 7 periods.

Q.9. What do you mean by electronic configuration of the elements?

Ans- The systematic distribution of the electrons among the orbitals of an atom of an element according to increasing order of their energies is called as electronic configuration of that element.

TWO MARKS QUESTIONS

Q.1. Describe the two merits of long form periodic table over the Mendeleev's periodic table?

Ans- 1. It removed the anomalies about the position of isotopes which existed in the Mendeleev's table.

2. It relates the position of an element in the periodic table with its electronic configuration.

Q.2. What is a period in the periodic table? How do atomic sizes change in a period with an increase in atomic number?

Ans- The horizontal rows in periodic table are called as periods. The atomic sizes decrease in a period with an increase in atomic number.

Q.3. The outer electronic configuration of some elements are:

(a) $3s^2\ 3p^4$ (b) $3d^{10}4s^2$ (c) $3s^2\ 3p^6\ 4s^2$ (d) $6s^2\ 4f^3$

To which block of elements in the periodic table each of these belongs?

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

Q.4. What is meant by periodicity in properties of elements? What is the reason behind this?

Ans- The repetition of similar properties after regular intervals is called as periodicity. It is due to the similarity in the outer electronic configurations which gives rise to the periodic properties of the elements.

Q.5. How do atomic radii vary in a group and a period?

Ans- In group- Atomic size increases on moving from top to bottom.

In period- Atomic size decreases on moving left to right in a period.

Q.6. Arrange the following in the order of increasing radii:

(a) I, I^+, I^- (b) F, Cl, Br

Ans- (a) $I^+ < I < I^+$ (b) O < N< P

Q.7. Name the factors which affect the ionisation enthalpy of an element.

Ans- (i) Size of atom or ion (ii) Nuclear charge (iii) Electronic configuration

(iv) Screening effect (v) Penetration effect of the electrons

Q.8. How does ionisation enthalpy vary in a group and a period?
Ans- In Period- It increases from left to right
In group- It decreases down the group.
Q.9. Noble gases have zero electron gain enthalpy values. Explain.
Ans- Because the outer most shell of noble gases is completely filled and no more electrons can be added.

Q.10. Elements in the same group have equal valency. Comment on it.

Ans- Because the general outer most electronic configurations of the elements of a group remain same and they contain equal number of electrons in their respective outer most shells.

THREE MARKS QUESTIONS

Q.1. The first ionisation enthalpy of magnesium is higher than that of sodium. On the other hand, the second ionisation enthalpy of sodium is very much higher than that of magnesium. Explain.

Ans- The 1^{st} ionisation enthalpy of magnesium is higher than that of Na due to higher nuclear charge and slightly smaller atomic radius of Mg than Na. After the loss of first electron, Na^+ formed has the electronic configuration of neon (2,8). The higher stability of the completely filled noble gas configuration leads to very high second ionisation enthalpy for sodium. On the other hand, Mg^+ formed after losing first electron still has one more electron in its outermost (3s) orbital. As a result, the second ionisation enthalpy of magnesium is much smaller than that of sodium.

Q.2. What are the major differences between metals and non- metals?
Ans-

Property	Metal	Non- Metal
Nature	Electropositive	Electronegative
Type of ion formed	Cation (Positively Charged)	Anion (Negatively Charged)
Reaction with acids	Active metals displace hydrogen	Do not displace hydrogen
Oxides	Basic	Acidic

Q.3. Among the elements of the second period Li to Ne pick out the element:

(i) with the highest first ionisation energy(ii) with the highest electronegativity

(iii) with the largest atomic radius Give the reason for your choice.

Ans- (i) The ionisation energy increases on going from left to right. Therefore, the element with the highest ionisation energy is Ne.

(ii) The electro negativity is electron- accepting tendency. This increases on going from left to right and decreases down the group. Therefore, the element with the highest electro- negativity is F.

(iii) The atomic radius decreases across a period on going from left to right. Thus, the first element of any period should have the largest atomic radii. Here, Li has the largest atomic radii.

Q.4. Arrange the following as stated:

(i) N_2, O_2, F_2, Cl_2 (Increasing order of bond dissociation energy)
(ii) F, Cl, Br, I (Increasing order of electron gain enthalpy)
(iii) F_2, N_2, Cl_2, O_2 (Increasing order of bond length)

Ans-	(i)	F_2	<	Cl_2	<	O_2	<	N_2
	(ii)	I	<	Br	<	F	<	Cl
	(iii)	N_2	<	O_2	<	F_2	<	Cl_2

Q.5. Why does the first ionisation enthalpy increase as we go from left to right through a given period of the periodic table?

Ans- In a period, the nuclear charge (the number of protons) increases on going from left to right. The electron added to each element from left to right enters the same shell. This results in an increase of the effective nuclear charge across the period on moving from left to right. As a result, the electron get more firmly bound to the nucleus. This causes an increase in the first ionisation enthalpy across the period.

Q.6. Use the periodic table to answer the following questions.

(i) Identify the element with five electrons in the outer sub-shell.

(ii) Identify an element that would tend to lose two electrons.

(iii) Identify an element that would tend to gain two electrons.

Ans- (i) Chlorine (ii) Magnesium (iii) Oxygen

Q.7. Explain why are cations smaller and anions larger in size than their parent atoms?

Ans- (a) The cations are smaller than their parent atoms due to the following reasons:

(i) Disappearance of the valence shell.

(ii) Increase of effective nuclear charge

(b) The anions are larger than their parent atoms due to the following reason:

An increase in the number of electrons in the valence shell reduces the effective nuclear charge due to greater mutual shielding by the electrons. As a result, electron cloud expands, i.e., the ionic radius increases.

Q.8. Describe the theory associated with the radius of an atom as it

(a) gains an electron (b) loses an electron

Ans- **(a)** When an atom gains an electron, its size increases. When an electron is added, the number of electrons goes up by one. This results in an increase in repulsion among the electrons. However, the number of protons remains the same. As a result, the effective nuclear charge of the atom decreases and the radius of the atom increases.

(b) When an atom loses an electron, the number of electrons decreases by one while the nuclear charge remains the same. Therefore, the interelectronic repulsions in the atom decrease. As a result, the effective nuclear charge increases. Hence, the radius of the atom decreases.

Q.9. How does atomic radius vary in a period and in a group? How do you explain the variation?

Ans- Atomic radius generally decreases from left to right across a period. This is because within a period, the outer electrons are present in the same valence shell and the atomic number increases from left to right across a period, resulting in an increased effective nuclear charge. As a result, the attraction of electrons to the nucleus increases.

On the other hand, the atomic radius generally increases down a group. This is because down a group, the principal quantum number (*n*) increases which results in an increase of the distance between the nucleus and valence electrons.

Q.10. Consider the following species:

N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+} and Al^{3+}

(a) What is common in them?

(b) Arrange them in the order of increasing ionic radii.

Ans- (a) the same number of electrons (10 electrons). Hence, the given species are isoelectronic.

(b) $Al^{3+} < Mg^{2+} < Na^{+} < F^{-} < O^{2-} < N^{3-}$

FIVE MARKS QUESTIONS

Q.1. What is the cause of the periodicity in the properties of the elements? How do the following properties vary in (a) a group and (b)in a period

(i) electronegativity (ii) ionisation enthalpy (iii) Atomic size

Ans- It is due to the similarity in the outer electronic configurations which gives rise to the periodic properties of the elements.

(a) In a group:
 - (i) Electronegativity- It decreases down the group.
 - (ii) Ionisation enthalpy- It decreases down the group.
 - (iii) Atomic size- It increases down the group.

(b) In a period:
 - (i) Electronegativity- Increases
 - (ii) Ionisation enthalpy- Increases
 - (iii) Atomic size- Dereases.

Q.2. The first ($\Delta_i H1$) and the second ($\Delta_i H$) ionization enthalpies (in kJ mol^{-1}) and the ($\Delta_{eg} H$) electron gain enthalpy (in kJ mol^{-1}) of a few elements are given below:

Elements	$\Delta_i H$	$\Delta_i H$	$\Delta_{eg} H$
I	520	7300	–60
II	419	3051	–48
III	1681	3374	–328
IV	1008	1846	–295
V	2372	5251	+48
VI	738	1451	–40

Which of the above elements is likely to be :

(a) the least reactive element. (b) the most reactive metal.
(c) the most reactive non-metal. (d) the least reactive non-metal.
(e) the metal which can form a stable binary halide of the formula MX_2, (X=halogen).
(f) the metal which can form a predominantly stable covalent halide of the formula MX (X=halogen)?

Ans- **(a)** Element V is likely to be the least reactive element. This is because it has the highest first ionization enthalpy ($\Delta_i H_1$) and a positive electron gain enthalpy ($\Delta_{eg}H$).
(b) Element II is likely to be the most reactive metal as it has the lowest first ionization enthalpy ($\Delta_i H_1$) and a low negative electron gain enthalpy ($\Delta_{eg}H$).
(c) Element III is likely to be the most reactive non–metal as it has a high first ionization enthalpy ($\Delta_i H_1$) and the highest negative electron gain enthalpy ($\Delta_{eg}H$).
(d) Element V is likely to be the least reactive non–metal since it has a very high first ionization enthalpy ($\Delta_i H_2$) and a positive electron gain enthalpy ($\Delta_{eg}H$).
(e) Element VI has a low negative electron gain enthalpy ($\Delta_{eg}H$). Thus, it is a metal. Further, it has the lowest second ionization enthalpy ($\Delta_i H_2$). Hence, it can form a stable binary halide of the formula MX_2(X=halogen).
(f) Element I has low first ionization energy and high second ionization energy. Therefore, it can form a predominantly stable covalent halide of the formula MX (X=halogen).