



INDIAN SCHOOL AL WADI AL KABIR

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HANDOUT	TOPIC: CHAPTER 3 CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES	NOTE: A4 FILE FORMAT
NAME OF THE STUDENT:	CLASS & SEC:	ROLL NO:

MODERN PERIODIC TABLE

- Modern periodic law states that **the physical and chemical properties of the elements are the periodic function of their atomic numbers.**
- In the modern periodic table, the horizontal rows are called periods and the vertical columns are called groups.
- Elements having similar outer electronic configurations in their atoms are arranged in groups.
- There are 7 periods and 18 groups in the modern periodic table.

NOMENCLATURE OF ELEMENTS WITH ATOMIC NUMBER > 100

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	p
6	hex	h
7	sept	s
8	oct	o
9	enn	e

- The roots are put together in order of digits and **“ium”** is added at the end.

IUPAC nomenclature and symbol of elements with atomic number above 100

Atomic Number	Name according to IUPAC nomenclature	Symbol
101	Unnilunium	Unu
102	Unnilbium	Unb
103	Unniltrium	Unt
104	Unnilquadium	Unq
105	Unnilpentium	Unp
106	Unnilhexium	Unh
107	Unnilseptium	Uns
108	Unniloctium	Uno
109	Unnilennium	Une
110	Ununillium	Uun
111	Unununnium	Uuu
112	Ununbium	Uub
113	Ununtrium	Uut
114	Ununquadium	Uuq
115	Ununpentium	Uup
116	Ununhexium	Uuh
117	Ununseptium	Uus
118	Ununoctium	Uuo

ELECTRONIC CONFIGURATIONS OF ELEMENTS AND THE PERIODIC TABLE

- The number of elements in each period is twice the number of atomic orbitals available in the energy level that is being filled.
- Elements in the same group have similar valence shell electronic configurations, the same number of electrons in the outer orbitals, and similar properties.
- Elements are classified into four blocks; **s-block**, **p-block**, **d-block** and **f-block** depending on the type of atomic orbitals that are being filled with electrons.

The s-block elements

- The elements of group 1 (alkali metals) and group 2 (alkaline earth metals) which have ns^1 and ns^2 outermost electronic configuration belong to the s-block elements.
- General outer electronic configuration is $ns^{1 \text{ to } 2}$
- They are all reactive metals with low ionization enthalpies. They form 1+ ion (alkali metals) or 2+ ion (alkaline earth metals).
- The metallic character and the reactivity increase as we go down the group. The compounds of the s-block elements, except lithium and beryllium, are predominantly ionic.

The p-block elements

- The p-block elements comprise those belonging to group 13 to 18.
- s-block and p-block elements are collectively called the representative elements or main group elements.
- General outer electronic configuration is $ns^2 np^{1 \text{ to } 6}$
- Group 18 elements are called noble gases or inert gases. They exhibit very low chemical reactivity as all the orbitals in the valence shell of the noble gases are completely filled.
- Group 17 elements are called **halogens** and group 16 elements are called **chalcogens**. These two groups of elements have highly negative electron gain enthalpies.
- p - block elements consist of metals, non-metals and metalloids.
- The non-metallic character increases as we move from left to right across a period and metallic character increases as we go down the group.

The d-block elements

- These are the elements of group 3 to 12 in the centre of the modern periodic table.
- General outer electronic configuration is $(n-1) d^{1-10} ns^{0-2}$
- d-block elements, except group 12, are also known as transition elements.
- Transition elements are those elements which contain incompletely filled d orbitals in the ground state or in their common oxidation states.
- Group 12 elements such as Zn, Cd and Hg do not contain incompletely filled d orbitals in the ground state or in their common oxidation state (+2).

- Transition elements are all metals. They mostly form coloured ions and exhibit variable valence (oxidation states).
- They are paramagnetic and also used as catalysts.

Eg: - Fe, Ni, Cu etc.

The f-block elements

- The two rows of elements at the bottom of the Periodic Table, Lanthanoids, Ce (Z = 58) – Lu (Z = 71) and Actinoids, Th (Z = 90) – Lr (Z = 103) are called *f*-block elements or inner-transition elements.
- General outer electronic configuration is $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$
- They are all metals. The chemistry of the early actinoids is more complicated than the corresponding lanthanoids, due to the large number of oxidation states possible for these actinoid elements.
- All Actinoid are radioactive. The elements after uranium are called **Transuranium** Elements.

Metals, Non-metals and Metalloids

- Metals comprise more than 78% of all known elements. Metals usually have high melting and boiling points. They are good conductors of heat and electricity. They are malleable and ductile.
- Non-metals are located at the top right-hand side of the periodic table.
- They are poor conductors of heat and electricity. Most non metallic solids are brittle and are neither malleable nor ductile.
- The elements become more metallic as we go down a group and non-metallic character increases across the period.
- Some elements show properties that are characteristic of both metals and non-metals. These elements are called **semi-metals or metalloids**.

Eg:- Silicon, germanium, arsenic, antimony, tellurium etc.

PERIODIC TRENDS IN PHYSICAL PROPERTIES OF ELEMENTS

(a) Atomic Radius

- Atomic radius can be expressed in two ways; Covalent radius and metallic radius.
- Covalent radius is the distance between two atoms when they are bound together by a single bond in a covalent molecule.
- Metallic Radius is half the internuclear distance separating the metal cores in the metallic crystal.

Variation of atomic radius across the period

- The atomic size generally decreases across a period.
- It is because within the period the outer electrons are in the same valence shell and the effective nuclear charge increases as the atomic number increases resulting in the increased attraction of electrons to the nucleus.

Variation of atomic radius on moving down the group

- On moving down the group, atomic radius increases.
- This is because on moving down the group, number of shells increases and the valence electrons are farther from the nucleus. Also, shielding effect increases and thus, outermost electrons are weakly held by the nucleus.

(b) Ionic Radius

- The ionic radius is estimated by measuring the distances between cations and anions in ionic crystals.
- A cation is smaller than its parent atom because it has fewer electrons while its nuclear charge remains the same.

Eg: - Na^+ is smaller than Na

- The size of an anion is larger than that of the parent atom because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge.

Eg: - Cl^- is larger than Cl

Isoelectronic species

- Atoms and ions which contain same number of electrons are called isoelectronic species.

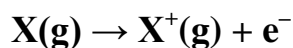
Eg: - O^{2-} , F^- , Na^+ , Mg^{2+} etc.

- Among the isoelectronic ions, the cation with the greater positive charge will have a smaller radius because of the greater attraction of the electrons to the nucleus.
- Anion with the greater negative charge will have the larger radius. In this case, the net repulsion of the electrons will outweigh the nuclear charge and the ion will expand in size.

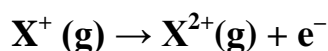
Eg: - Ionic radius decreases in the order $O^{2-} > F^- > Na^+ > Mg^{2+}$

(c) Ionization enthalpy

- Ionization enthalpy is the energy required to remove an electron from an isolated gaseous atom (X) in its ground state.



- The ionization enthalpy is expressed in kJ mol^{-1} .
- The second ionization enthalpy is the energy required to remove the second most loosely bound electron from a unipositive ion.



- Ionization is an endothermic process as energy is always required to remove electrons from an atom.
- The second ionization enthalpy is higher than the first ionization enthalpy because it is more difficult to remove an electron from a positively charged ion than from a neutral atom. In the same way the third ionization enthalpy is higher than the second and so on.
- Noble gases have very high first ionisation energy because they have closed electron shells and very stable electron configurations.

Variation of ionisation energy across the period

- First ionization enthalpy generally **increases** across the period.
- Across the period, successive electrons are added to the same shell and the shielding effect of inner electrons does not increase much to compensate for the increased attraction of the electron to the nucleus.
- Thus, across a period, increasing nuclear charge outweighs the shielding. Consequently, the outermost electrons are held more and more tightly and the ionization enthalpy increases across a period.

Variation of ionisation energy on moving down the group

- First ionization enthalpy **decreases** on moving down the group.
- On moving down the group, number of shells increases and outermost electron goes farther from the nucleus. In this case, increase in shielding outweighs the increasing nuclear charge and the removal of the outermost electron requires less energy down a group.

Successive ionisation energy

- Second ionization enthalpy of alkali metals (group 1 elements) is exceptionally high, which is in fact greater than the second ionisation enthalpy of respective group two elements.
- In the case of group 1 elements, the second electron is removed from an inner shell which is closer to the nucleus and experiences less shielding effect. As a result, large amount of energy is needed to remove the second electron.
- Similarly, third ionisation enthalpy of group 2 elements is exceptionally high and so on.

Exceptions

1) First ionization enthalpy of boron ($Z = 5$) is less than that of beryllium ($Z = 4$).

- In beryllium, the electron removed during the ionization is an s-electron whereas the electron removed during ionization of boron is a p-electron. The penetration of a 2s-electron to the nucleus is more than that of a 2p-electron; hence the 2p electron of boron is more shielded from the nucleus by the inner core of electrons than the 2s electrons of beryllium.

2) First ionization enthalpy of Al ($Z = 13$) is less than that of Mg ($Z = 12$).

- In Mg, the electron removed during the ionization is an s-electron whereas the electron removed during ionization of Al is a p-electron. The penetration of a 3s-electron to the nucleus is more than that of a 3p-electron; hence the 3p electron of Al is more shielded from the nucleus by the inner core of electrons than the 3s electrons of Mg.

3) First ionization enthalpy of oxygen ($Z = 8$) is less than that of nitrogen ($Z = 7$).

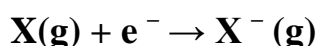
- In N atom, three 2p-electrons reside in different atomic orbitals whereas in oxygen atom, two of the four 2p-electrons occupy the same 2p-orbital resulting in an increased electron-electron repulsion. Hence, it is easier to remove the fourth electron from oxygen.

4) First ionization enthalpy of S (Z = 16) is less than that of P (Z = 15).

- In P atom, three 3p-electrons reside in different atomic orbitals whereas in sulphur atom, two of the four 3p-electrons occupy the same 3p-orbital resulting in an increased electron-electron repulsion. Hence, it is easier to remove the fourth electron from sulphur.

(d) Electron gain enthalpy

- When an electron is added to a neutral gaseous atom (X) to convert it into a negative ion, the enthalpy change accompanying the process is called electron gain enthalpy.



- For many elements, energy is released when an electron is added to the atom and the electron gain enthalpy is negative.
- Group 17 elements have very high negative electron gain enthalpies because they can attain stable noble gas electronic configurations by picking up an electron.
- Noble gases have large **positive** electron gain enthalpies because the electron has to enter the next higher principal quantum level leading to a very unstable electronic configuration.
- Electron gain enthalpies have large negative values toward the upper right of the periodic table preceding the noble gases.
- In general, electron gain enthalpy becomes more negative across a period due to decrease size of atoms.
- Also, electron gain enthalpy becomes less negative on moving down the group because the size of the atom increases and the added electron would be farther from the nucleus.

Exceptions

- Electron gain enthalpy of F is less negative than that of Cl.
- This is because when an electron is added to F, the added electron goes to the smaller $n = 2$ quantum level and suffers significant **repulsion** from the other electrons present in this level. For the $n = 3$ quantum level (Cl), the added electron occupies a larger space and the electron-electron repulsion is much less.
- For the same reason, electron gain enthalpy of O is less negative than that of S.

(e) Electronegativity

- The ability of an atom in a chemical compound to attract shared electrons to itself is called electronegativity.
- Fluorine is the most electronegative element. It has been given an arbitrarily value of 4.0 in the electronegativity scale developed by Linus Pauling.
- Electronegativity generally increases across a period because the attraction between the outer (or valence) electrons and the nucleus increases as the atomic radius decreases in a period.
- On the other hand, electronegativity decreases down the group in the periodic table. On moving down the group, atomic size increases and attraction between the outer electrons and the nucleus decreases.
- In general, non-metals are highly electronegative and metals are electropositive.

Periodic trends in chemical properties

(a) Periodicity of valence or oxidation states

- The valence of representative elements is usually equal to the number of electrons in the outermost orbitals and / or equal to eight minus the number of outermost electrons.
- Oxidation state of an element in a particular compound is defined as the charge acquired by its atom on the basis of electronegative consideration from other atoms in the molecule.

Eg: - In OF_2 oxidation state of oxygen is +2 and in Na_2O , it is -2

(b) Anomalous properties of second period elements

- The first element of each of the groups 1 (lithium) and 2 (beryllium) and groups 13-17 (boron to fluorine) differs in many respects from the other members of their respective group.
- The behaviour of lithium and beryllium is more similar with the second element of the following group i.e., magnesium and aluminium, respectively. This is called **diagonal relationship** in the periodic properties.
- The anomalous behaviour is attributed to their small size, high electronegativity of the elements and large charge/ radius ratio.
- The first member of each group has only four valence orbitals (2s and 2p) available for bonding, whereas the second member of the groups have nine valence orbitals (3s, 3p, 3d).

- Hence, the maximum covalency of the first member of each group is 4, whereas the other members of the groups can expand their valence shell to accommodate more than four.

Eg: - $[\text{AlF}_6]^{3-}$ is known but $[\text{BF}_6]^{3-}$ is unknown.

- The first member of p-block elements displays greater ability to form $p_\pi - p_\pi$ multiple bonds to itself (Eg: - $\text{C} = \text{C}$, $\text{C} \equiv \text{C}$, $\text{N} = \text{N}$) and to other second period elements (Eg: - $\text{C} = \text{O}$, $\text{C} = \text{N}$, $\text{N} = \text{O}$) compared to subsequent members of the same group.

Periodic trends and chemical reactivity

- Ionization enthalpy of the extreme left element in a period is the least and the electron gain enthalpy of the element on the extreme right is the highest negative. This results into high chemical reactivity at the two extremes and the lowest in the centre.
- Thus, the maximum chemical reactivity at the extreme left is exhibited by the loss of an electron leading to the formation of a cation and at the extreme right shown by the gain of an electron forming an anion.
- The normal oxide formed by the element on extreme left is the most basic (Eg: - Na_2O) whereas that formed by the element on extreme right is the most acidic (Eg: - Cl_2O_7).
- Oxides of elements in the centre are amphoteric (e.g., Al_2O_3 , As_2O_3) or neutral (e.g., CO , NO , N_2O).
- Amphoteric oxides behave as acidic with bases and as basic with acids, whereas neutral oxides have no acidic or basic properties.

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