

# Classification of Elements and Periodicity in Properties

## Topic 1

1. The basic theme of organisation in the Periodic table is to simplify and systematise the study of Physical and chemical Properties of all the elements and their innumerable compounds.

2. Mendeleev arranged elements in horizontal rows and vertical columns of a table in order of their increasing atomic weights in such a way that the elements with similar properties occupied the same vertical column or group. He realized that some of the elements did not fit with his scheme of classification if the order of atomic weight was strictly followed. So, he did not stick to his criteria, he ignored the order of atomic weights, thinking that the atomic measurements might be incorrect, and placed the elements with similar properties together.

3. According to Mendeleev's periodic law, the physical and chemical properties of the elements are periodic functions of their atomic weights, but according to modern periodic law, the physical and chemical properties of the elements are periodic functions of their atomic numbers.

4. (i) Lawrencium  $\rightarrow$   ${}_{103}\text{Lr}$  and Berkelium  $\rightarrow$   ${}_{97}\text{Bk}$   
(ii) Seaborgium  $\rightarrow$   ${}_{106}\text{Sg}$

## Topic 2

1. In the sixth period of periodic table, the orbitals to be filled are  $6s$ ,  $4f$ ,  $5d$  and  $6p$ . The complete filling of these orbitals require  $2 + 14 + 10 + 6 = 32$  electrons, hence, the sixth period of the periodic table should have 32 elements.

2. The outermost electronic configuration of element  $Z = 114$  is  $[\text{Rn}] 5f^{14} 6d^{10} 7s^2 7p^2$ . It has  $n = 7$ , so period  $\rightarrow 7$  It belongs to  $p$ -block so, group number =  $10 + 4 = 14$ .

3.  ${}_{17}\text{Cl} \rightarrow$  It belongs to the third period. So, outermost shell is  $n = 3$ . Its configuration is  $[\text{Ne}] 3s^2 3p^5$ . Therefore, its atomic number = 17.

4. In a group, the chemical properties of the elements remain nearly the same due to same valence shell electronic configuration.

5.

	Metals	Non-metals
1.	They are usually solids at room temperature. Mercury is an exception.	They may be solids, liquids or gases at room temperature.
2.	They have high melting and boiling points.	They have low melting and boiling points.
3.	They are good conductors of heat and electricity.	They are not good conductors of heat and electricity.
4.	They are malleable and ductile.	They are neither malleable nor ductile.
5.	They have 1 to 3 electrons in their valence shells.	They have 4 to 7 electrons in their valence shells.
6.	They have low ionisation energies.	They have high ionisation energies.
7.	They have low electron affinities.	They have high electron affinities.
8.	They are electropositive.	They are electronegative.
9.	They are hard and lustrous.	They are brittle and non-lustrous.
10.	They are sonorous.	They are non-sonorous.

6. (a) Fluorine (b) Magnesium

(c) Oxygen

(d) Group 17 (Halogens) :

F, Cl – Non-metals and gases

Br – Non-metal and liquid

I – Shows metallic lustre.

7. General outer electronic configuration :

$s$ -block elements :  $ns^{1-2}$ ; e.g., Na :  $3s^1$

$p$ -block elements :  $ns^2 np^{1-6}$ ; e.g., F :  $2s^2 2p^5$

$d$ -block elements :  $(n-1)d^{1-10} ns^{0-2}$ ;

e.g., Zn :  $3d^{10} 4s^2$

$f$ -block elements :  $(n-2)f^{1-14}(n-1)d^{0-1} ns^2$ ;

e.g., Ce :  $4f^2 5d^0 6s^2$ .

8. (i)  $ns^2 np^4$

For  $n = 3$  ; outer electronic configuration =  $3s^2 3p^4$

$\therefore$  The element will be in 3rd period, 16th group and  $p$ -block (Sulphur).

(ii)  $(n-1)d^2 ns^2$

For  $n = 4$  ; outer electronic configuration =  $3d^2 4s^2$

∴ The element will be in 4th period, 4th group and *d*-block (Titanium).

(iii)  $(n - 2)f^7 (n - 1)d^1 ns^2$

For  $n = 6$ ; the outer electronic configuration is  $4f^7 5d^1 6s^2$ .

∴ The element will be in 6th period, 3rd group and *f*-block (Gadolinium).

9. (c) : Principal quantum number

10. (b) : The *d*-block has 10 columns, because a maximum of 10 electrons can occupy all the orbitals in a *d*-subshell.

## Topic 3

1. **Atomic radius** : It is the distance between the centre of the nucleus and outermost shell of an atom where electrons are present.

**Ionic radius** : It is the distance between the nucleus and outermost shell of an ion.

2. **Variation of atomic radii in a period** : As we move from left to right across a period, there is regular decrease in atomic radii of representative elements. This can be explained on the basis of effective nuclear charge which increases gradually in a period. Electron cloud is attracted more strongly towards nucleus as the effective nuclear charge becomes more and more strong along the period. The increased force of attraction brings contraction in size.

**Variation of atomic radii in a group** : Atomic radii in a group increase as the atomic number increases. The increase in size is due to extra energy shell which outweighs the effect of increased nuclear charge.

3. Isoelectronic species are those which have same number of electrons.

(i)  $F^-$  has 10 electrons. Therefore, the species  $N^{3-}$ ,  $O^{2-}$ , Ne,  $Na^+$ ,  $Mg^{2+}$ , etc., are isoelectronic with  $F^-$ .

(ii) Ar has 18 electrons. Therefore, the species  $P^{3-}$ ,  $S^{2-}$ ,  $Cl^-$ ,  $K^+$ ,  $Ca^{2+}$ , etc., are isoelectronic with Ar.

(iii)  $Mg^{2+}$  has 10 electrons. Therefore, the species  $N^{3-}$ ,  $O^{2-}$ ,  $F^-$ , Ne,  $Na^+$ , etc., are isoelectronic with  $Mg^{2+}$ .

(iv)  $Rb^+$  has 36 electrons. Therefore, the species  $Br^-$ , Kr,  $Sr^{2+}$ , etc., are isoelectronic with  $Rb^+$ .

4. (a) All these are isoelectronic species as they are having same number of electrons *i.e.*, 10.

(b) As  $Z/e$  decreases, size increases so, increasing order should be,  
 $Al^{3+} > Mg^{2+} > Na^+ > F^- > O^{2-} > N^{3-}$ .

5. A cation is smaller than its parent atom because it has less number of electrons than the parent atom while its nuclear charge remains the same. The size of an anion will be 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.

6. In the definition of ionisation enthalpy and electron gain enthalpy, isolated gaseous atom is required for comparison purposes.

Ionisation enthalpy is the minimum amount of energy required to remove most loosely bound electron from an isolated atom in the gaseous state of an element so as to convert it into gaseous monovalent positive ion.

Electron gain enthalpy is the energy change accompanying the process of adding an electron to a gaseous isolated atom to convert it into a negative ion, *i.e.*, a monovalent anion. Both the above mentioned processes are carried out for an isolated gaseous atom, which in turn is obtained from either the excitation of a ground state atom (in case the element is monoatomic) or atomisation of polyatomic elements.

The force with which an electron is attracted by the nucleus is appreciably affected by presence of other atoms in the neighbourhood. Since in the gaseous state the atoms are widely separated, therefore these interatomic forces are minimum.

The term ground state means that the atom must be present in the most stable state.

7. Energy of an electron in the ground state of H-atom,  
 $E_1 = -2.18 \times 10^{-18} \text{ J}$

Ionisation enthalpy =  $E_\infty - E_n$

Ionisation enthalpy per mole of atomic hydrogen =  $(E_\infty - E_1) N_A$   
 $= [0 - (-2.18 \times 10^{-18})] \times 6.023 \times 10^{23}$

$= 2.18 \times 6.023 \times 10^5 \text{ J/mol} = 13.13 \times 10^5 \text{ J/mol}$   
 $= 1.313 \times 10^6 \text{ J/mol}$

8. (i) An *s*-electron is attracted to the nucleus more than a *p*-electron. In beryllium, the electron removed during the ionisation is a *s*-electron whereas the electron removed during ionisation of boron is a *p*-electron. The penetration of a *2s*-electron to the nucleus is more than that of a *2p*-electron. Therefore, it is easier to remove the *2p*-electron from boron as compared to the removal of a *2s*-electron from beryllium. Thus, boron has a smaller first ionisation enthalpy than beryllium.

(ii) O has lower ionisation energy than N because N ( $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$ ) has extra stable electronic configuration whereas O ( $1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$ ) does not. O has lower ionisation energy than F because O has larger size than F.

9. The electronic configurations of Na and Mg are :

Na :  $1s^2 2s^2 2p^6 3s^1$  ;  $Na^+$  :  $1s^2 2s^2 2p^6$

Mg :  $1s^2 2s^2 2p^6 3s^2$  ;  $Mg^+$  :  $1s^2 2s^2 2p^6 3s^1$  ;  $Mg^{2+}$  :  $1s^2 2s^2 2p^6$

The 1<sup>st</sup> ionisation enthalpy of Na is lesser than that of Mg because Mg has extra stable configuration ( $3s^2$ ) and smaller size, so, a larger amount of energy would be required to remove an electron from the *3s* orbital, which has a pair of electrons.

The 2<sup>nd</sup> ionisation enthalpy of Na is more than that of Mg because  $Na^+$  has extra stable configuration (complete octet), whereas  $Mg^+$  does not have extra stable configuration.

10. We have to consider two factors : (i) atomic size, (ii) screening effect of electrons from each other. On moving down the group, the atomic size increases gradually due

to addition of new principal energy shell, consequently the force of attraction of the nuclei for the valence electrons decreases and ionisation enthalpy decreases. The effective nuclear charge experienced by a valence electron in an atom will be less than the actual charge on the nucleus because of shielding or screening of the valence electron from the nucleus by the intervening core electrons. As we descend the group, the outermost electron being increasingly farther from the nucleus, there is an increased shielding of the nuclear charge by the electrons in the inner levels. The increase in shielding outweighs the increasing nuclear charge and the removal of the outermost electron requires less energy down a group.

**11.** (i) Al has lower ionisation enthalpy than B due to larger size.

(ii) Ga has slightly higher ionisation enthalpy than Al due to ineffective shielding by  $3d$  electrons.

(iii) In has lower ionisation enthalpy than Ga due to larger size.

(iv) Tl has higher ionisation enthalpy than In due to ineffective shielding by  $4f$  electrons.

**12.** (i) F has more negative electron gain enthalpy than O due to smaller size, higher nuclear charge and greater possibility of attaining the nearest stable noble gas configuration by gaining one electron.

(ii) Cl has more negative electron gain enthalpy because in F the incoming electron is added to the smaller  $n = 2$  quantum level and suffers significant repulsion from the other electrons present in this level. In Cl, the added electron goes to  $n = 3$  quantum level and occupies a larger region of space and electron-electron repulsion experienced is far less.

**13.** The second electron gain enthalpy of O is +ve. This is because energy has to be supplied to convert  $O_{(g)}^-$  to  $O_{(g)}^{2-}$  in order to overcome the repulsive forces.

**14.**

	Electron gain enthalpy	Electronegativity
1.	It provides a measure of the ease with which an atom adds an electron to form an anion.	It is a qualitative measure of the ability of an atom in a chemical compound to attract shared electrons towards itself.
2.	It has an absolute value.	It is not a measurable quantity.
3.	Its periodicity is not regular in a period or in a group.	The periodicity is regular in a period but not so regular in a group.
4.	Its units are electron volts per atom or kilojoules per mole or kilocalories per mole.	It has no units but is merely a number.

**15.** Electronegativity of an atom is not a constant quantity. It depends on state of hybridisation and oxidation state. So,

the given statement is not correct as N cannot have constant value of electronegativity for all of its compounds.

**16.** The distance between the nucleus and the outermost shell of an ion is known as ionic radius.

(a) The gain of an electron leads to the formation of an anion. The radius of the anion is larger than the atomic radius of its 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 so, each electron is held less tightly and thereby the electron cloud expands.

(b) The removal of an electron from an atom results in the formation of a cation. A cation is smaller than its parent atom because it has fewer electrons while its nuclear charge remains the same and since it is now acting on lesser number of electrons and pulls them closer, the ion is smaller.

**17.** Two isotopes of the same element have the same first ionisation enthalpies because of same effective nuclear charges. *e.g.*,  $^{35}_{17}\text{Cl}$  and  $^{37}_{17}\text{Cl}$  have same ionisation enthalpies.

**18.** The trend  $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$  is observed for chemical reactivity because upon descending the group the ionisation energy of alkali metals decreases *i.e.*, it is easy for them to lose an electron from their valence shell and attain the nearest stable noble gas configuration.

The trend  $\text{F} > \text{Cl} > \text{Br} > \text{I}$  is observed for chemical reactivity in halogens because the standard reduction potential decreases as we descend the group. F being the most electronegative readily accepts an  $e^-$  to complete its octet.

**19.** (a) V : The element V has highest first ionisation enthalpy and positive electron gain enthalpy hence, it is least reactive.

(b) II : The element II has the least first ionisation enthalpy hence, it is most reactive metal.

(c) III : The element III has very high negative electron gain enthalpy hence, it is most reactive non-metal.

(d) IV : Element IV has high negative electron gain enthalpy but ionisation enthalpy is not that high hence, it is least reactive non-metal.

(e) VI : The first and second ionisation enthalpies of element VI indicate that it can form a stable binary halide.

(f) I : The element I has very low value of first ionisation enthalpy but very high second ionisation enthalpy. Hence, it will form a stable covalent halide of the formula  $MX$ .

**20.** (a)  $\text{Li}_2\text{O}$  (b)  $\text{Mg}_3\text{N}_2$

(c)  $\text{AlI}_3$  (d)  $\text{SiO}_2$

(e)  $\text{PF}_5$  (f)  $\text{LuF}_3$

**21. (c) :** Nuclear mass does not affect the valence electrons.

**22. (a) :** As nuclear charge ( $Z$ ) increases, size decreases.

**23. (d) :** Removal of an electron from orbitals bearing lower  $n$  value is difficult than from orbital having higher  $n$  value.

**24. (d) :** Along a period metallic character decreases while down a group metallic character increases.

**25. (c)**

**26. (b) :** Oxidising character increases along a period while decreases down a group.



