

PERIODIC CLASSIFICATION OF ELEMENTS (PART-I)

It is very difficult to study individually the chemistry of all the elements (about 114 at present) and their innumerable compounds. To ease out this problem, scientists searched for a systematic way to classifying the elements. This is done by classifying the elements into groups and periods of periodic table.

DOBEREINER'S LAW OF TRIADS : This was the classification of elements into groups of three elements each with similar properties such that the atomic weight of the middle element was the arithmetic mean of the other two e.g. Ca, Sr, Ba; Cl, Br, I etc. The difference between atomic weights of any two successive elements is nearly constant.

NEWLAND'S LAW OF OCTAVES: When elements were arranged in the order of increasing atomic weights then it was observed that every eighth element had properties similar to those of the first just like the eight note of an octave of music.

Li Be B C N O F

Na Mg Al Si P S Cl

Prouts Hypothesis: Atomic weight of an element is simple multiple of atomic weight of hydrogen.

MENDELEEV'S PERIODIC TABLE: The physical and chemical properties of the elements are a periodic function of their atomic weights.

Mendeleev named Gallium and Germanium elements as Eka-aluminium and Eka-silicon respectively.

STRUCTURAL FEATURES OF THE MENDELEEV'S PERIODIC TABLE:

(i) Mendeleev's original periodic table consists of eight vertical columns which are called groups I-VIII & seven horizontal rows are called periods 1-7. But modified Mendeleev's periodic table contains nine groups, i.e., I-VIII and zero (of noble gases).

(ii) All groups except VIII and zero have been further divided into two sub-groups called A and B. **A** groups of left hand side consist of normal elements while groups **B** of right hand side consist of transition elements.

(iii) Elements of group IA are called **alkali metals** while those of group IB (i.e, Cu, Ag, Au) are called **coinage metals**.

MODERN PERIODIC LAW : In 1913, the English physicist, Henry **Moseley** observed regularities in the characteristic X-ray spectra of the elements. A plot of $\sqrt{\nu}$ (where ν is frequency of X-rays emitted) against atomic number (Z) gave a straight line and not the plot of $\sqrt{\nu}$ vs atomic mass. He thereby concluded that the atomic number is a more fundamental property of an element than its atomic mass.

Modern Periodic Law and can be stated as : **The physical and chemical properties of the elements are periodic functions of their atomic numbers.**

Numerous forms of Periodic Table have been devised from time to time. Some forms emphasise chemical reactions and valence, whereas others stress the electronic configuration of elements. A modern version, the so-called “**long form**” of the Periodic Table of the elements, is the most convenient and widely used. The horizontal rows (which Mendeleev called series) are called periods and the vertical columns, groups. Elements having similar outer electronic configurations in their atoms are arranged in vertical columns, referred to as groups or families. According to the recommendation of International Union of Pure and Applied Chemistry (IUPAC), the groups are numbered from 1 to 18 replacing the older notation of groups IA ... VIIA, VIII, IB ... VIIB and 0. There are altogether seven periods. The period number corresponds to the highest principal quantum number (n) of the elements in the period. The first period contains 2 elements. The subsequent periods consists of 8, 8, 18, 18 and 32 elements, respectively. The seventh period is incomplete and like the sixth period would have a theoretical maximum (on the basis of quantum numbers) of 32 elements. In this form of the Periodic Table, 14 elements of both sixth and seventh periods (lanthanoids and actinoids, respectively) are placed in separate panels at the bottom. It was observed that the elements with similar properties reoccurred at regular intervals of 2, 8, 8, 18 or 32. These numbers (2, 8, 8, 18 and 32) are called **magic numbers**, and cause of periodicity in properties.

Structural features of the long form of the periodic table

(a) Groups : (i) The 18 vertical columns, of the periodic table, are called groups (ii) Elements of groups 1, 2, 13-17 are called normal or representative elements. (iii) Elements of groups 3-12 are called transition elements (iv) The elements belonging to a particular group is known as a family and is usually named after the first element. For example, Boron family (group 13). In addition to this, some groups have typical names. For example, Elements of group 1 are called alkali metals Elements of group 2 are called alkaline earth metals Elements of group 15 are called pnictogens ,elements of group 16 are

called chalcogens, elements of group 17 are called halogens, elements of group 18 are called noble gases.

(b) Periods :

(i) The 7 horizontal columns (or rows) are called periods. The seven periods of periodic table are

Shortest period - 1st period (${}_1\text{H}$ to ${}_2\text{He}$) contains 2 elements

Short periods - 2nd period (${}_3\text{Li}$ to ${}_{10}\text{Ne}$) and 3rd period (${}_{11}\text{Na}$ to ${}_{18}\text{Ar}$) contains 8 elements each.

Long periods - 4th period (${}_{19}\text{K}$ to ${}_{36}\text{Kr}$) and 5th period (${}_{37}\text{Rb}$ to ${}_{54}\text{Xe}$) contain 18 elements each.

Longest period - 6th period (${}_{55}\text{Cs}$ to ${}_{86}\text{Rn}$) contains 32 elements and is the longest period.

Incomplete period - 7th period (${}_{87}\text{Fr}$ –) is, however, incomplete and contains at present only 26 elements.

(c) Blocks :

The periodic table is divided into four main blocks (s, p, d and f) depending upon the subshell to which the valence electron enters into.

(i) s-Block - Elements of group 1 and 2 constitute s-block.

(ii) p-Block - Elements of group 13 to 18 constitute p-block.

(iii) d-Block - Elements of group 3 to 12 constitute d-block.

There are three complete series and one incomplete series of d-block elements. These are: 1st or 3d-transition series which contains ten elements with atomic number 21-30 (${}_{21}\text{Sc}$ – ${}_{30}\text{Zn}$). 2nd or 4d-transition series which contains ten elements with atomic number 39–48 (${}_{39}\text{Y}$ – ${}_{48}\text{Cd}$). 3rd or 5d-transition series which also contains ten elements with atomic number 57 and 72-80 (${}_{57}\text{La}$, ${}_{72}\text{Hf}$ – ${}_{80}\text{Hg}$). 4th or 6d-transition series which contains only ten elements.

(v) The f-block elements comprise two horizontal rows placed at the bottom of the periodic table to avoid its un-necessary expansion and make the symmetrical nature of periodic table. The two series of f-block elements containing 14 elements each.

Lanthanides - The 14 elements from 58Ce–71Lu in which 4f-subshell is being progressively filled up are called lanthanides or rare earth elements.

Actinides - Similarly, the 14 elements from 90Th –103Lr in which 5f-subshell is being progressively filled up are called actinides.

- Elements of s and p-blocks are called normal or representative elements, those of d-block are called transition elements while the f-block elements are called inner transition elements.

- The 11 elements with $Z = 93-103$ ($93\text{Np} - 103\text{Lr}$) which occur in the periodic table after uranium and have been prepared by artificial means are called transuranics. These are all radioactive elements.

NOMENCLATURE OF ELEMENTS WITH ATOMIC NUMBER > 100:

According to IUPAC the nomenclature can be derived using numerical roots for 0 and numbers 1-9 for atomic numbers of elements. The roots are put together in order of digits which make the atomic number and 'ium' is added at the end. Use the following table.

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

Example :

- Name the IUPAC name of the element of atomic number 108 be Unniloctium and symbol UnO
- Example – Name the element with atomic number 115 will be - Ununpentium and symbol UuP

DIAGONAL RELATIONSHIP: Certain elements of 2nd period i.e., Li, Be, B show similarity in properties with their diagonal elements in the 3rd period i.e., Mg, Al, Si as shown below:

	Group 1	Group 2	Group 13	Group 14
2nd period	Li	Be	B	C
3rd period	Na	Mg	Al	Si

This is called diagonal relationship and is due to the reason that these pairs of elements have almost identical ionic radii and polarizing power. (i.e. charge/size ratio)

PERIODIC PROPERTIES : Properties which show a regular gradation when we move from left to right in a period or from top to bottom in a group are called periodic properties. These properties are atomic size, ionisation energy electron affinity etc.

(A) Atomic size: It refers to the distance between the centre of nucleus of atom to its outermost shell of electrons. The absolute value of atomic radius cannot be determined because (i) It is not possible to locate the exact position of electrons in an atom as an orbital has no sharp boundaries. (ii) It is not possible to isolate an individual atom. (iii) In a group of atoms, the probability distribution of electrons is influenced by the presence of neighbouring atoms. Since absolute value of atomic size cannot be determined, it is expressed in

terms of the operational definitions such as covalent radius, vander waal's radius, ionic radius and metallic radius

(a) **Covalent radius** : It is half of the distance between the nuclei of two like atoms bonded together by a single covalent bond, hence it is also known as single bond covalent radius (SBCR). Thus, covalent radius $r_{\text{cov}} = d/2$, where d = internuclear distance between two covalently bonded like atoms.

(b) **Van-der Waal's radius** : It is one-half of the distance between the nuclei of two adjacent atoms belonging to two neighbouring molecules of an element in the solid state. The covalent radius is always smaller than the corresponding van der waal's radius.

(c) **Metallic radius** : It is half of the distance between two successive nuclei of two adjacent metal atoms in the metallic closed packed crystal lattice. Metallic radius of an element is always greater than its covalent radius.

(d) **Ionic radius** : It is the effective distance from the nucleus of the ion upto the electrons in the outer shell to which it can influence the ionic bond. An atom can be changed into a cation (by loss of electron) which is always much smaller than the corresponding atom, or to an anion (by gaining of electrons) which is always larger than the corresponding atom.

(v) **Periodic variation of atomic radii** :

(a) On moving down the group the valence shells become far away from the nucleus and thus the atomic radius increases.

(b) On moving along the period, the effective nuclear charge increases and thus the electron cloud is attracted more strongly towards the nucleus resulting in the contraction of atomic radius.

(vi) **Isoelectronic ions or species** : These are ions of the different elements which have the same number of electrons but different magnitude of the nuclear charge. The size of isoelectronic ions decreases with the increase in the nuclear charge.

(B) Ionisation energy (I.E.) : The amount of energy (work) required to remove an electron from the outermost shell of an isolated (free) atom in gaseous state is known as ionisation potential or energy or better first ionisation potential of the element, i.e. $M(g) + I.E \longrightarrow M(g) + e^-$

(i) The amount of energies required to remove the subsequent electrons (2nd, 3rd, ...) from the gaseous cation of the element one after the other are collectively called successive ionisation energies. These are designated as $I.E_1$, $I.E_2$, $I.E_3$, $I.E_4$ and so on.

It may be noted that. $I.E_4 > I.E_3 > I.E_2 > I.E_1$ (for a particular element)

IE is expressed in eV/atom or kcal mol⁻¹ or kJ mol⁻¹. Note that 1 eV atom⁻¹ = 23.06 kcal mol⁻¹ = 96.3 kJ mol⁻¹.

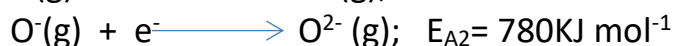
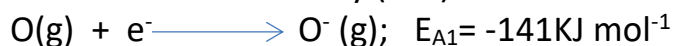
- (ii) In general, the first I.E. increases along the period from left to right due to decrease in size and increase in effective nuclear charge.
However there are some exceptions to the general trend –
 - (a) I.E. decreases from elements of group 2 → 3 (e.g. B < Be, Al < Mg)
 - (b) I.E. decreases from elements of group 15 → 16 (e.g. O < N, S < P)
- (iii) In a group of the periodic table, the ionisation energy decreases from top to bottom.
- (iv) The factors which affect the ionisation energy are :
 - (a) Atomic size or radius : I.E. decreases as the atomic size increases so the attractive force decreases.
 - (b) Number of electrons in the inner shell (screening effect) : On moving down a group, the number of inner shells increases which increase the screening effect and hence the ionisation potential tends to decrease.
 - (c) Nuclear charge : On moving along the period, effective nuclear charge increases due to addition of electrons in same valence shell and hence ionisation energy increases.
 - (d) Stable configuration : Half filled or completely filled subshells possess extra stability and thus it is more difficult to remove electron and hence more is I.E.
 - (e) Penetration effect : More penetrating (i.e. more closes) are subshells of a shell to the nucleus, more tightly the electrons are held towards the nucleus and more is I.E. I.E. : $s > p > d > f$ for a given shell. Penetration power \propto Surface area of a subshell.
 - (v) In second period elements, the correct increasing order of ionisation energies is
 IE1 : $\text{Li} < \text{B} < \text{Be} < \text{C} < \text{O} < \text{N} < \text{F} < \text{Ne}$
 IE2 : $\text{Be} < \text{C} < \text{B} < \text{N} < \text{F} < \text{O} < \text{Ne} < \text{Li}$
 - (vi) In third period elements, the correct increasing order of ionisation is
 IE1 : $\text{Na} < \text{Al} < \text{Mg} < \text{Si} < \text{S} < \text{P} < \text{Cl} < \text{Ar}$
 IE2 : $\text{Mg} < \text{Si} < \text{Al} < \text{P} < \text{S} < \text{Cl} < \text{Ar} < \text{Na}$

(C) Electron affinity (EA): It is the amount of energy released when a gaseous atom accepts the electron to form gaseous anion.



- (i) E_{A} values are expressed in eV/atom or kcal/mol or kJ/mol.
- (ii) The energy change accompanying the addition of first, second, third etc. electrons to neutral isolated gaseous atoms are called successive electron affinities and are designated as E_{A1} , E_{A2} , E_{A3} etc.
- (iii) The first EA is always taken as negative. However, the addition of second electron to the negatively charged ion is opposed by coulombic repulsion and

hence required (absorbed) energy for the addition of second electron. Thus, second electron affinity (E_{A2}) of an element is taken as positive. For example,



(iv) Electron affinity increases in moving along the period from left to right due to increase in charge. But the values are unexpectedly low in elements of group 2, 15 and 18 due to stable electronic configurations of exactly half-filled and completely filled sub orbits.

(v) Within a group of the periodic table the electron affinities decreases from top to bottom with increase in size of elements.

(viii) **Electron affinity depends upon-**

(a) Effective nuclear charge - Greater the effective nuclear charge, more is the attraction of nucleus towards the electron and hence higher will be the value of E_A

(b) Atomic size - Greater the atomic radius of the atom, less will be the attraction of the nucleus to the electron to be added and hence lower will be the value of E_A .

(c) Penetrating power - Due to penetrating power, E_A for addition of electron show the order : $s > p > d > f$

(d) Electronic configuration : Half filled and fully filled subshell are extra stable and thus oppose the addition of electron which leads to lower, E_A value, e.g. E_A of C > E_A of N.

(D) Electronegativity (EN) : It is the tendency of an atom in a molecule to attract the bonded shared pair of electrons, towards itself.

(i) There are several electronegativity scales-

(a) Mulliken scale: On the Mulliken scale, electronegativity χ is taken as average of IE and E_A , i.e.,

$$\chi = \frac{EA + IE}{2} \text{ where IE and EA are expressed in electron volts.}$$

$$\text{Or, } \chi = \frac{EA + IE}{540} \text{ where IE and EA are expressed in kJ mol}^{-1}.$$

$$\text{or, } \chi = \frac{EA + IE}{130} \text{ where IE and EA are expressed in kcal mol}^{-1}$$

(b) Pauling scale: This is the most widely used scale and is based upon bond energy data. According to Pauling, the difference in electronegativity of two atoms A and B is given by the relationship as

$$\chi_A - \chi_B = 0.208 \sqrt{\Delta E}$$

where χ_A and χ_B are electronegativities of the atoms A and B respectively while.

$$\Delta E = E_{A-B} - \sqrt{(E_{A-A} \times E_{B-B})}$$

where E_{A-B} , E_{A-A} and E_{B-B} represent bond dissociation energies of the bonds A-B, A-A and B-B respectively. The Pauling and the Mulliken scales are related to each other by the relation, $\chi_{\text{Pauling}} = \chi_{\text{Mulliken}} / 2.8$

(ii) In a period, EN increases from left to right due to decrease in size and increase in nuclear charge of atoms.

(iii) In a group, EN decreases from top to bottom due to increase in atomic size.

(iv) Electronegativity depends on:

(a) Atomic size

(b) Nuclear charge

(c) Shielding effect

(d) Oxidation state – EN increases as the positive oxidation state increases.

(e) Hybridization - Greater is the s-character in a hybrid orbital more is electronegativity.

(v) If electronegativity difference is greater than 1.7 the bond is ionic otherwise covalent.

(vi) In general, greater is difference of EN between two atoms smaller is the bond length.

(vii) Smaller the electronegativity, larger is the atomic size.

ATOMIC VOLUME : It may be defined as the volume occupied by one mole atoms of the element at its melting point in solid state.

(i) It is obtained by dividing the gram atomic mass of the element by its density.

(ii) It decreases along the period, reaches a minimum in the middle and then starts increasing, because of different packing arrangement of atoms in different elements in the solid state, i.e., P_4 , S_8 etc.

(iii) In moving down the group atomic volume goes on increasing gradually.

ACID-BASE BEHAVIOUR OF OXIDES AND HYDROXIDES : The oxides or hydroxide of an element may act either as base or an acid depending upon its ionization energy.

(i) If the IE is low, it acts as a base and if the IE is high, it acts as an acid.

(ii) The IE of alkali metals is the lowest, therefore, their oxides and hydroxides are the strongest bases. The basic character of their hydroxides increases down the group in the order: $\text{CsOH} > \text{RbOH} > \text{KOH} > \text{NaOH} > \text{LiOH}$

(iii) The IE of halogens is quite high, therefore, their oxides are the strongest acids. The acidic character of their oxides and hydroxides decreases down the group in the order: $\text{HClO}_4 > \text{HBrO}_4 > \text{HIO}_4$

(iv) Within a period, the ionization energies of the elements usually increase and hence their oxides and hydroxides show a gradual variation from strongly basic through amphoteric to strongly acidic character.

PERIODIC TABLE: PART-II (Q/A)

1. Identify the correct order of the size of the following:

- (a) $\text{Ca}^{2+} < \text{K}^+ < \text{Ar} < \text{Cl}^- < \text{S}^{2-}$ (b) $\text{Ar} < \text{Ca}^{2+} < \text{K}^+ < \text{Cl}^- < \text{S}^{2-}$
(c) $\text{Ca}^{2+} < \text{Ar} < \text{K}^+ < \text{Cl}^- < \text{S}^{2-}$ (d) $\text{Ca}^{2+} < \text{K}^+ < \text{Ar} < \text{S}^{2-} < \text{Cl}^-$

Ans. (a) For isoelectronic species, size of anion increases as negative charge increases whereas size of cation decreases with increase in positive charge. Further ionic radii of anions is more than that of cations.

Thus the correct order is $\text{Ca}^{2+} < \text{K}^+ < \text{Ar} < \text{Cl}^- < \text{S}^{2-}$

2. In which of the following electronic configuration an atom has the lowest ionisation enthalpy?

- (a) $1s^2 2s^2 2p^3$ (b) $1s^2 2s^2 2p^5 3s^1$
(c) $1s^2 2s^2 2p^6$ (d) $1s^2 2s^2 2p^5$

Ans. (b) It will release one electron most easily to attain stable electronic configuration.

3. Which one of the following ionic species has the greatest proton affinity to form stable compound?

- (a) NH_2^- (b) F^- (c) I^- (d) HS^-

Ans. (a) Because its conjugated acid (NH_3) is weakest among the conjugated acids of the given species(bases).

4. The stability of + 1 oxidation state increases in the sequence:

- (a) $\text{Tl} < \text{In} < \text{Ga} < \text{Al}$ (b) $\text{In} < \text{Tl} < \text{Ga} < \text{Al}$
(c) $\text{Ga} < \text{In} < \text{Al} < \text{Tl}$ (d) $\text{Al} < \text{Ga} < \text{In} < \text{Tl}$

Ans. (d) The stability of +1 oxidation state increases from aluminium to thallium i.e. $\text{Al} < \text{Ga} < \text{In} < \text{Tl}$ due to Inert pair effect.

5. Amongst the elements with following electronic configurations, which one of them may have the highest ionization energy?

- (a) $[\text{Ne}]3s^2 3p^2$ (b) $[\text{Ar}]3d^{10} 4s^2 4p^3$
(c) $[\text{Ne}]3s^2 3p^1$ (d) $[\text{Ne}]3s^2 3p^3$

Ans. (d) This is due to

(i) Half-filled stable electronic configuration

(ii) Presence of inner lying d electrons having poor shielding effect.

6. Which of the following represents the correct order of increasing electron gain enthalpy with negative sign for the elements O, S, F and Cl

- (a) $\text{Cl} < \text{F} < \text{O} < \text{S}$ (b) $\text{O} < \text{S} < \text{F} < \text{Cl}$
(c) $\text{F} < \text{S} < \text{O} < \text{Cl}$ (d) $\text{S} < \text{O} < \text{Cl} < \text{F}$

Ans.(b) Due to combined effect of size and inter-electronic repulsion.

7. Among the elements Ca, Mg, P and Cl, the order of increasing atomic radii is

- (a) $\text{Ca} < \text{Mg} < \text{P} < \text{Cl}$ (b) $\text{Mg} < \text{Ca} < \text{Cl} < \text{P}$
(c) $\text{Cl} < \text{P} < \text{Mg} < \text{Ca}$ (d) $\text{P} < \text{Cl} < \text{Ca} < \text{Mg}$

Ans. (c) Size of atoms increases down the group and decreases along a period from left to right.

8. What is the value of electron gain enthalpy of Na^+ if IE_1 of Na = 5.1 eV ?

- (a) -5.1 eV (b) -10.2 eV (c) +2.55 eV (d) +10.2 eV

Ans. (a) IE_1 of Na = - Electron gain enthalpy of Na^+

9. According to the Periodic Law of elements, the variation in properties of elements is related to their

- (a) nuclear masses (b) atomic numbers
(c) neutron-proton number ratios (d) atomic masses

Ans. (b) The properties of elements change with a change in atomic number

10. Which is the correct order of ionic sizes (At. No. : Ce = 58, Sn = 50, Yb = 70 and Lu = 71) ?

- (a) $\text{Ce} > \text{Sn} > \text{Yb} > \text{Lu}$ (b) $\text{Sn} > \text{Ce} > \text{Yb} > \text{Lu}$
(c) $\text{Lu} > \text{Yb} > \text{Sn} > \text{Ce}$ (d) $\text{Sn} > \text{Yb} > \text{Ce} > \text{Lu}$

Ans. (b) Due to Lanthanide contraction.

11. Which one of the following ions has the highest value of ionic radius ?

- (a) O^{2-} (b) B^{3+} (c) Li^+ (d) F^-

Ans. (a) Ionic radii follows the relation: $\frac{1}{r} \propto Z/e$

12. Among Al_2O_3 , SiO_2 , P_2O_3 and SO_2 the correct order of acid strength is

- (a) $\text{Al}_2\text{O}_3 < \text{SiO}_2 < \text{SO}_2 < \text{P}_2\text{O}_3$ (b) $\text{SiO}_2 < \text{SO}_2 < \text{Al}_2\text{O}_3 < \text{P}_2\text{O}_3$
(c) $\text{SO}_2 < \text{P}_2\text{O}_3 < \text{SiO}_2 < \text{Al}_2\text{O}_3$ (d) $\text{Al}_2\text{O}_3 < \text{SiO}_2 < \text{P}_2\text{O}_3 < \text{SO}_2$

Ans. (d) From left to right in a period acidic character of oxides increases with increase in their non-metallic character.

13. Which of the following statements is not correct?

(a) $\text{La}(\text{OH})_3$ is less basic than $\text{Lu}(\text{OH})_3$.

(b) In lanthanide series ionic radius of Ln^{3+} ions decreases.

(c) La is actually an element of transition series rather than lanthanide series.

(d) Atomic radii of Zr and Hf are same because of lanthanide contraction.

Ans. (a) : $\text{La}(\text{OH})_3$ is more basic than $\text{Lu}(\text{OH})_3$. This is because ionic size of La^{3+} ion is more than Lu^{3+} ion (Fajan's rule).

14. Which of the following oxides is amphoteric in character?

(a) SnO_2

(b) SiO_2

(c) CO_2

(d) CaO

Ans. (a) SnO_2 is amphoteric as it reacts with both acid and base.



15. In which of the following arrangements, the order is NOT according to the property indicated against it?

(a) $\text{Li} < \text{Na} < \text{K} < \text{Rb}$: Increasing metallic radius

(b) $\text{I} < \text{Br} < \text{F} < \text{Cl}$: Increasing electron gain enthalpy (with negative sign)

(c) $\text{B} < \text{C} < \text{N} < \text{O}$: Increasing first ionization enthalpy

(d) $\text{Al}^{3+} < \text{Mg}^{2+} < \text{Na}^+ < \text{F}^-$: Increasing ionic size

Ans.(c) In a period the value of ionisation potential increases from left to right with breaks where the atoms have some what stable configuration. In this case N has half filled stable orbital. Hence has highest ionisation energy

16. The increasing order of the first ionization enthalpies of the elements B, P, S and F (Lowest first) is

(a) $\text{B} < \text{P} < \text{S} < \text{F}$

(b) $\text{B} < \text{S} < \text{P} < \text{F}$

(c) $\text{F} < \text{S} < \text{P} < \text{B}$

(d) $\text{P} < \text{S} < \text{B} < \text{F}$

Ans.(b) On moving along a period ionization enthalpy increases from left to right and decreases from top to bottom in a group. But this trend breaks in case of atoms having fully or half filled stable orbitals. In this case P has a stable orbitals filled electronic configuration hence its ionisation enthalpy is greater in comparison to S.

17. Following statements regarding the periodic trends of chemical reactivity of the alkali metals and the halogens are given. Which of these statements gives the correct picture?

- (a) Chemical reactivity increases with increase in atomic number down the group in both the alkali metals and halogens
- (b) In alkali metals the reactivity increases but in the halogens it decreases with increase in atomic number down the group
- (c) The reactivity decreases in the alkali metals but increases in the halogens with increase in atomic number down the group
- (d) In both the alkali metals and the halogens the chemical reactivity decreases with increase in atomic number down the group

Ans. (b) On moving down group- I, from Li to Cs ionisation enthalpy decreases hence the reactivity increases. The lighter halogens are most reactive elements due to their low bond dissociation energy, high electron affinity and high enthalpy of hydration of halide ion. However their reactivity decreases with increase in atomic number.

18. The charge/size ratio of a cation determines its polarizing power. Which one of the following sequences represents the increasing order of the polarizing power of the cationic species, K^+ , Ca^{2+} , Mg^{2+} , Be^{2+} ?

- (a) $Ca^{2+} < Mg^{2+} < Be^{2+} < K^+$
- (b) $Mg^{2+} < Be^{2+} < K^+ < Ca^{2+}$
- (c) $Be^{2+} < K^+ < Ca^{2+} < Mg^{2+}$
- (d) $K^+ < Ca^{2+} < Mg^{2+} < Be^{2+}$.

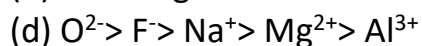
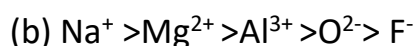
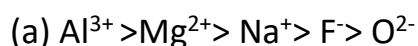
Ans.(d) Smaller the size and higher the charge more will be polarising power of cation.

19. In which of the following arrangements, the sequence is not strictly according to the property written against it?

- (a) $HF < HCl < HBr < HI$: increasing acid strength
- (b) $NH_3 < PH_3 < AsH_3$: increasing basic strength
- (c) $B < C < O < N$: increasing first ionization enthalpy
- (d) $CO_2 < SiO_2 < SnO_2 < PbO_2$: increasing oxidising power

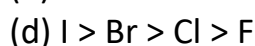
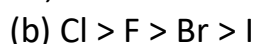
Ans. (b) In hydrides of 15th group elements, basic character decreases on descending the group with decrease in electron density on the central atom.

20. The correct sequence which shows decreasing order of the ionic radii of the elements is



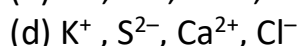
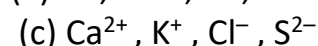
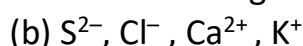
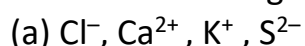
Ans. (d) All the given species are isoelectronic. For isoelectronic species anion having high negative charge is largest in size and the cation having high positive charge is smallest.

21. The correct order of electron gain enthalpy with negative sign of F, Cl, Br and I, having atomic number 9, 17, 35 and 53 respectively, is :



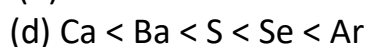
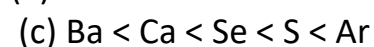
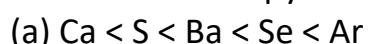
Ans. (b) As we move down in a group electron gain enthalpy becomes less negative because the size of the atom increases. Negative electron gain enthalpy of F is less than Cl due to high electron–electron repulsion in F compared to that in Cl.

22. The increasing order of the ionic radii of the given isoelectronic species is :



Ans. (c) All the given species are isoelectronic. For isoelectronic species anion having high negative charge is largest in size and the cation having high positive charge is smallest

23. Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar ?



Ans. (c) On moving along a period from left to right I.E. increases and on moving down a group I.E. decreases.

24. Element with atomic number 56 belongs to which block?



Ans. (a) $\text{Ba}_{56} \longrightarrow [\text{Xe}]6s^2$. The last electron enters into s-subshell. So it belongs to s-block.

25. The values of electronegativity of atoms A and B are 1.20 and 4.0 respectively. The percentage of ionic character of A-B bond is

(a) 50% (b) 72.24% (c) 55.3% (d) 43%

Ans. (b) **Percentage of ionic character** $= 16(\chi_A - \chi_B) + 3.5(\chi_A - \chi_B)^2$
 $= 16 \times (4 - 1.2) + 3.5 \times (4 - 1.2)^2$
 $= 72.24$