

13. PERIODIC TABLE AND PERIODICITY

1. INTRODUCTION

You must have visited a library. There are thousands of books in a large library. In spite of this, if you ask for a particular book, the library staff can locate it easily. How is it possible? In a library, the books are classified into various categories and subcategories. They are arranged on the shelves accordingly. Therefore, locating the books becomes easy. Same is the story with chemical elements. A large number of elements and compounds are known today. But a systematic classification of these elements has made their study possible and easy. The well organized and tabulated classification of elements, as we know it today, is called the **Periodic Table**. It not only helps to locate, identify and characterize the element and its properties but also points out the directions in which new investigations are made.

2. GENESIS OF PERIODIC CLASSIFICATION

In 18th century, the number of elements was limited. In 19th century, scientists began to seek ways to classify elements because of their rapidly increasing number. They started recognizing patterns in properties and began to develop classification schemes. Some such early attempts of classification are described below.

2.1 Prout's Hypothesis

The atomic weights of all elements are simple multiples of atomic weight of hydrogen. Prout gave this hypothesis on the basis of Dalton's atomic theory and the atomic weights of some elements known at that time. But this hypothesis could not last longer, because there are some atomic weights which are fractional not in whole number.

2.2 Dobereiner's Triads

It was first attempt towards classification of elements. He arranged similar elements in groups of three elements called triad and the atomic mass of the middle elements of the triad is approximately the arithmetic mean of the other two.

Table 13.1: Example of dobereiner's triad

	Triad			Mean of at. Mass of (I) and (III) element
(i)	Li ⁷	Na ²³	K ³⁹	$\frac{7+39}{2} = 23.00$
(ii)	Cl ^{35.5}	Br ⁸⁰	I ¹²⁷	$\frac{35.5+127}{2} = 81.25$
(iii)	P ³¹	As ⁷⁵	Sb ¹²⁰	$\frac{31+120}{2} = 75.50$
(iv)	S ³²	Se ⁷⁹	Te ¹²⁷	$\frac{127+32}{2} = 79.50$
(v)	Ca ⁴⁰	Sr ⁸⁸	Ba ¹³⁷	$\frac{40+137}{2} = 88.50$

Merits: After Dobereiner, Chemists focused on chemicals in groups having similar physical and chemical properties.

Demerits: All the known elements did not follow this rule. Law of triads was rejected as some triads nearly had same atomic masses, e.g., (Fe, Co, Ni), (Ru, Rh, Pd), (Os, Ir, Pt)

2.3 Newland's Rule of Octaves

When the lighter elements are arranged in order of their increasing atomic weights, then every eighth element is similar to the first element in its properties, similarly as the eighth note of a musical scale is similar to the 1st one. e.g. Na, 8th element resembles in their properties with Li. Similarly K, the 8th element with Na, and so on.

	do	re	mi	pha	sol	la	si
Symbol of element	Li	Be	B	C	N	O	F
	7	9	11	12	14	16	19
Symbol of element	Na	Mg	Al	Si	P	S	Cl
	23	24	27	28	31	32	35.5

It is clear from the above table that sodium is the eighth element from lithium, whose properties resemble that of lithium.

This type of classification was limited up to only 20 elements.

Demerits

- (i) Law of octave worked quite well for lighter elements but failed with heavier elements.
- (ii) Properties of elements were not taken into account and the elements were arranged in the order of their increasing atomic masses.
- (iii) No places were left for unknown elements and so, many elements occupied wrong positions. Thus, resulted in the rejection of the attempt.

2.4 Lothar Meyer's Volume Curve

The graphs of atomic volumes against weights are known as Lothar Meyer's volume curves.

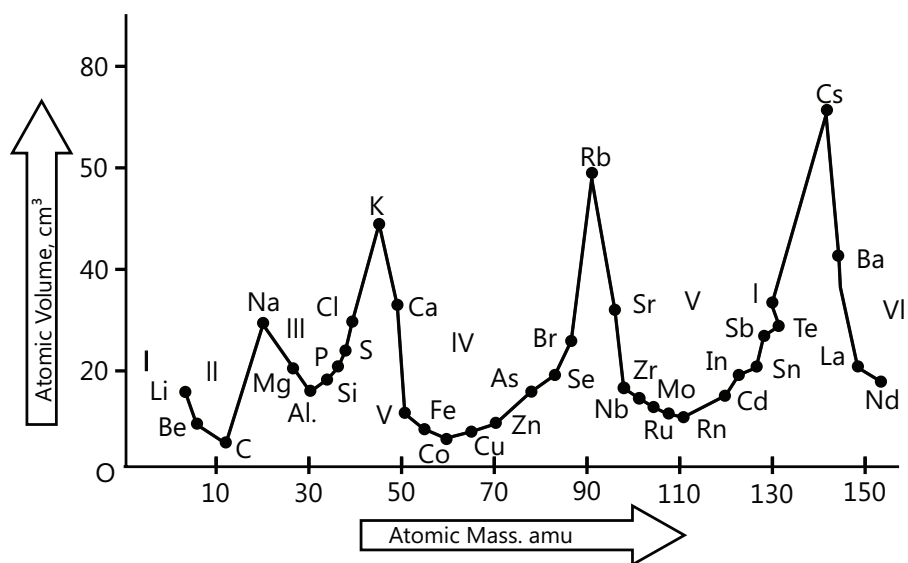


Figure 13.1: Lothar Meyer's volume curve

Features of curve

- (a) Alkali metals having the largest atomic volumes occupy the maxima of the curve.
- (b) Alkaline earth metals lie at about the mid points on the descending portions of the curve.
- (c) The halogens occupy positions on ascending portions of the curve before inert gases.
- (d) The transition metals occupy the minima of the curve.

Demerits: It lacked practical utility as it is not easy to remember the position of different elements on the curve.

2.5 Mendeleev's Periodic Law

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

Merits of Mendeleev's periodic table

- (a) Study of elements and their compounds becomes easy and systematic, as by knowing the property of one element in a group, then the properties of the other elements present in the same group can easily be predicted.
- (b) Helps in the discovery of new elements. As Mendeleev left some blank spaces for some unknown elements and further, predicted the properties of these elements e.g. ekaluminium, ekasilicon.
- (c) Correction of doubtful atomic mass.
- (d) Correction in the valency of some elements.
- (e) Correction in the position of some elements.
- (f) Classification of elements then known, was done for the first time and the elements having similar properties were kept in the same group.
- (g) It encouraged research and led to discovery of newer elements.
- (h) Mendeleev had even predicted the properties of many elements not discovered at that time. This helped in the discovery of these elements.

For example: Mendeleev predicted the properties of the following elements.

- (i) Eka-boron - This was later called scandium (Sc)
- (ii) Eka- aluminium - This was later called gallium (Ga)
- (iii) Eka-silicon - This was later called germanium (Ge)

Limitations of Mendeleev's periodic table

- (a) The position of hydrogen was found to be anomalous due to its resemblance with the 1st group alkali metals and also with the 7th group halogens in their properties.
- (b) Position of isotopes: Isotopes must have different positions but they were placed in the same group.
- (c) Position of isobars: They were placed in different groups.
- (d) Dissimilar elements were placed together in the same group like K and Cu in 1st group.
- (e) Similar elements were placed in different groups.
- (f) Some higher atomic weight elements were placed before the lower atomic weight elements
- (g) e.g. Ar⁴⁰ precedes K³⁹, Co^{58.9} precedes Ni^{58.7}, Te^{127.6} precedes I¹²⁷.
- (h) Position of metals and non-metals: Both were placed together in the same group.
- (i) Diagonal relationship could not be explained.

- (j) Position of lanthanides and actinides was not properly specified.
- (k) No proper position to VIII group elements.
- (l) There was no indication whether lanthanides and actinides were associated with group IIIA or group IIIB.
- (m) Position of Isobars- These elements had different groups when mass remained the same.
- (n) Lot of stress was given to the valence of elements.

2.6 Modern Periodic Law and Modern Periodic Table

Mosley: Proved that the square root of frequency (f) of the rays, which are obtained from a metal on showering high velocity electrons is proportional to the nuclear charge of the atom.

This can be represented by the following expression.

$$\sqrt{f} = a(Z-b) \text{ where } Z \text{ is nuclear charge on the atom and } a \text{ and } b \text{ are constants.}$$

The nuclear charge on an atom is equal to the atomic number.

Modern Periodic Table

According to modern periodic law, "The properties of elements are the periodic function of their atomic numbers".

Period-The details about the seven periods are as follows:-

Period	Atomic number		Number of elements
	From	to	
First	H (1)	He (2)	2
Second	Li (3)	Ne (10)	8
Third	Na (11)	Ar (18)	8
Fourth	K (19)	Kr (36)	18
Fifth	Rb (37)	Xe (54)	18
Sixth	Cs (55)	Rn (86)	32 (including lanthanides)
Seventh	Fr (87)	Ha (105)	19 (including actinides)

Group: The modern periodic table has 18 vertical columns and according to CAS system there are 16 groups having the following number of elements.

Group	Number of Elements
(a) I A group	7 (H, Li, Na, K, Rb, Cs, Fr) Alkali metals
(b) II A group	6 (Be, Mg, Ca, Sr, Ba, Ra) Alkaline earth metals
(c) III A group	5 (B, Al, Ga, In, Tl) Boron family
(d) IV A group	5 (C, Si, Ge, Sn, Pb) Carbon family
(e) V A group	5 (N, P, As, Sb, Bi) Nitrogen family
(f) VI A group	5 (O, S, Se, Te, Po) Oxygen family (Chalcogen)
(g) VII A group	5 (F, Cl, Br, I, At) Halogen family
(h) Zero group	6 (He, Ne, Ar, Kr, Xe, Rn) Inert elements 32 (Sc, Y, La, Ac & 14 lanthanide elements & 14 actinide elements.)
(i) III B group	These are elements of IIIB group, which could not be accommodated in one column and therefore written separately outside the periodic table.

(j)	IV B group	4 (Ti, Zr, Hf, Rf)
(k)	V B group	4 (V, Nb, Ta, Db)
(l)	VI B group	3 (Cr, Mo, W)
(m)	VII B group	3 (Mn, Tc, Re)
(n)	VIII (3) group	9 (Fe, Co, Ni, Ru, Rh, Pd, Os, Ir, Pt)
(o)	I B group	3 (Cu, Ag, Au)
(p)	II B group	3 (Zn, Cd, Hg)

Advantages of the Long Form of the Periodic Table

- (a) The table is based on a more fundamental property i.e. atomic number.
- (b) It correlates the position of elements with their electronic configuration more clearly.
- (c) The completion of each period is more logical. In a period, as the atomic number increases, the energy shells are gradually filled up until an inert gas configuration is reached.
- (d) It eliminates the even and odd series of IV, V and VI periods of Mendeleev's periodic table.
- (e) The position of VIII group is also justified in this table. All the transition elements have been brought to the middle as the properties of transition elements are intermediate between s- and p-block elements.
- (f) Due to the separation of two sub-groups, dissimilar elements do not fall together. One vertical column accommodates elements with same electronic configuration thereby showing same properties.
- (g) The table completely separates metals and non-metals. Non-metals are present in upper right corner of the periodic table.
- (h) There is a gradual change in properties of the elements with increase in their atomic numbers i.e., periodicity of properties can be easily visualized. The same properties occur after the intervals of 2, 8, 8, 18, 18 and 32 elements which indicates the capacity of various periods of the table.
- (i) The greatest advantage of this periodic table is that this can be divided into four blocks namely s-, p-, d- and f-block elements.
- (j) This arrangement of elements is easy to remember and reproduce.

Defects of the Long Form of the Periodic Table

- (a) The position of hydrogen is still disputable as it was there in MENDELEEV periodic table in group IA as well as IVA & VIIA.
- (b) Helium is an inert gas but its configuration is different from that of the other inert gas elements
- (c) Lanthanide and actinide series could not be adjusted in the main periodic table and therefore they had to be provided with a place separately below the table.

To Locate Group and Period if Atomic Number is given

Locate period: Write electronic configuration of each element for which the atomic number is given. The number of outermost shell suggests the period to which it belongs in the periodic table.

Locate it's Block and Group

Group can be located after knowing the block of an element as follows:

For s-block elements,

gp. no. = No. of s-electrons in valence shell

For p-block elements,

gp. no. = No. of s-electrons + p-electrons in valence shell + 10

For d-block elements,

gp. no. = No. of $(n - 1)d + ns$ electrons

Illustration 1: How many elements from the following atomic number are p-block elements?

83, 79, 42, 64, 37, 54 34

Use the following data for predicting answer – (Atomic Number of noble gases are given in the bracket)

(JEE MAIN)

He[2] Ne[10] Ar[18] Kr[36]

Xe[54] Rn[86]

Sol: 83 – p-block $Xe^{54} 6s^2 5d^{10} 4f^{14} 6p^3$

79 – d-block $Xe^{54} 6s^1 4f^{14} 5d^{10}$

42 – d-block $Kr^{36} 5s^1 4d^5$

64 – f-block $Xe^{54} 6s^2 5d^1 4f^7$

37 – s-block $Kr^{36} 5s^1$

54 – p-block Xe

34 – p-block $Ar^{18} 4s^2 3d^{10} 4p^4$

Illustration 2: (a) Write the electronic configuration of the elements given below:

A (At. No. = 9), B (At. No. = 12), C (At. No. = 29), D (At. No. = 54) and E (At. No. = 58)

(b) Also predict the period, group number and block to which they belong.

(JEE MAIN)

Sol: (a) Electronic configuration of the element A, B, C, D and E are as follows:

Element	At. No.	Electronic configuration
A	9	$1s^2 2s^2 2p^5$
B	12	$1s^2 2s^2 2p^6 3s^2$
C	29	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
D	54	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6$
E	58	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^2 5d^1 4f^1$

(b) **Element A:** Receives the last electron in 2p-orbital, therefore, it belongs to **p-block**.

Group number = 10 + No. of electrons in the valence shell = 10 + 7 = **17**.

Period = Principal quantum number of the valence shell = **2nd**.

Element B: Receives the last electron in 3s-orbital, thus, it belongs to **s-block**.

Group number = No. of electrons in the valence shell = **2**

Period = Principal quantum number of the valence shell = **3rd**

Element C: Receives the last electron in the 3d-orbital, thus, it belongs to **d-block**.

Group number = No. of electrons (penultimate shell + valence shell) = 10 + 1 = **11**

Period = Principal quantum number of the valence shell = **4th**

Element D: Receives its last electron in the 5p-orbital, thus, it belongs to **p-block**.

Group number = 10 + No. of electrons in the valence shell = 10 + 8 = **18**

Period of the element = Principal quantum number of the valence shell = **5th**

Element E: Receives its last electron in the 4f-orbital, thus, it belongs to **f-block**

It may be noted here, that, the filling of 4f-orbital occurs only when one electron has already entered the 5d-orbital. Therefore, element E belongs **to f-block** and not to **d-block**. Since it belongs to lanthanide series, there is no such group number of its own but is usually considered to lie in **group 3**.

Period = Principal quantum number of the valence shell = **6th**.

Illustration 3: Elements A, B, C, D and E have the following electronic configurations.

A: $1s^2 2s^2 2p^1$

B: $1s^2 2s^2 2p^6 3s^2 3p^1$

C: $1s^2 2s^2 2p^6 3s^2 3p^3$

D: $1s^2 2s^2 2p^6 3s^2 3p^5$

E: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

Which among these will belong to the same group in the periodic table?

(JEE ADVANCED)

Sol: We know that elements having similar valence electronic configuration belong to the same group of the periodic table. Therefore, elements **A** and **B** having three electrons in the valence shell, i.e. $2s^2 2p^1$ and $3s^2 3p^1$ respectively belong to the same group, i.e., group 13 of the periodic table.

3. NAMING OF ELEMENTS HAVING ATOMIC NUMBER GREATER THAN 100

(a) The name is derived directly from the atomic number of the elements using the following numerical roots:

Table 13.2: Naming of elements having atomic number greater than 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	oct	e

(b) The roots are put together in the order of the digits which make up the atomic number and are terminated by 'ium' to spell out the name. The final 'n' of 'enn' is removed when it occurs before 'nil' and the final 'i' of 'bi' and of 'tri' when it occurs before 'ium'

(c) The symbol of the element is composed of the initial letters of the numerical roots which make up the name.

Illustration 4: Eka-aluminium and eka-silicon were the names given by Mendeleev for the unknown elements gallium and germanium respectively. A recently discovered element was first named as eka-mercury. What is its atomic number? Write its group number, electronic configuration, IUPAC name. **(JEE MAIN)**

Sol: The element which comes after mercury in the periodic table is called eka-mercury. Its various parameters are:

$$Z = 80 + 32 = 112 \quad \text{IUPAC name: Uub}$$

Official name Cn(copernicium) E.C. = [Rn] $5f^{14} 8d^{10} 7s^2$

4. CLASSIFICATION OF PERIODIC TABLE BASED ON BLOCKS

s-block Elements: Elements of groups 1 and 2 including He in which the last electron enters the s-orbitals of the valence shell are called s-block elements. There are only **14 s-block elements** in the periodic table.

Characteristics:

- The electronic configuration of outermost shell of s-block elements is ns^1 (alkali metals; group 1) or ns^2 (alkaline earth metals; group 2)
- The valence of group I elements is +1 and those of group II elements is +2.
- These are soft metals having low melting points and boiling points.
- Most of these form ionic compounds on account of their lower ionization energy.
- Most of these metals (except Be & Mg) and their salts impart characteristic colour to the flame e.g., sodium imparts a golden yellow colour; potassium imparts violet colour to the flame.
- These are highly reactive elements and are strong reducing agents.
- All are good conductors of heat and electricity.

p-block Elements: Elements of groups 13-18 in which the last electron enters the p-orbitals of the valence shell are called p-block elements.

Characteristics:

- The electronic configuration of the outermost shell of p-block elements (group 13, 14, 15, 16, 17 and 18) is $ns^2 np^{1-6}$.
- These elements include metals and non-metals with a few metalloids. The metallic character, however, decreases along the period but increases down the group.
- These possess relatively higher ionization energy which tends to increase along the period but decreases down the group.
- Most of them form covalent compounds.
- Most of these elements show negative (except some metals) as well as positive oxidation states (except F).
- The oxidizing power of these elements increases along the period but decreases down the group.

d-Block Elements: 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 numbers 39-48 ($_{39}\text{Y}$ - $_{48}\text{Cd}$).

3rd or 5d transition series which contains ten elements with atomic numbers 57 and 72-80

($_{57}\text{La}$, $_{72}\text{Hf}$ - $_{80}\text{Hg}$). 4th or 6d transition series which is incomplete at present and contains only nine elements. These are

$_{89}\text{Ac}$, $_{104}\text{Rf}$, $_{105}\text{Ha}$, Unh (Seaborgium, $Z = 106$), $_{107}\text{Bh}$ (Bohrium), $_{108}\text{Hs}$ (Hassium), $_{109}\text{Mt}$ (Meitnerium), Ds (Darmstadtium, $Z = 110$) and Cn (Copernicium, $Z = 112$) or Ekamercury. The element, $Z = 111$ has not been discovered yet. Thus, in all, there are 39 d-block elements.

Characteristics:

- (a) The electronic configuration of outermost shell of d-block elements is ns^{0-2} followed with $(n-1)s^2p^6d^{1-10}$.
- (b) All (except Hg) are hard, ductile metals with high melting and boiling points.
- (c) All of these are good conductors of heat and electricity.
- (d) Their ionization energies are higher than s-block elements but lesser than p-block elements.
- (e) Most of the transition metals form coloured ions (Zn^{2+} , Hg^{2+} , Cd^{2+} are colourless.)
- (f) These elements show variable oxidation states.
- (g) Most of these elements possess catalytic activity.
- (h) Metals and their ions are generally paramagnetic due to the presence of unpaired electrons.
- (i) Most of the transition metal ions possess the tendency to form complex ions.
- (j) Most transition metals form alloys.

f-block Elements: f-Block elements are also called inner-transition elements. In these elements, the f-subshell of the inner-penultimate is progressively filled up. There are two series of f-block elements each containing 14 elements. The fourteen elements from $_{58}\text{Ce}$ - $_{71}\text{Lu}$ in which, 4f-subshell is progressively filled up are called lanthanides or rare elements. Similarly, the fourteen elements from $_{90}\text{Th}$ - $_{103}\text{Lr}$ in which, 5f-subshell is progressively filled up are called actinides.

Characteristics:

- (a) The electronic configuration of outermost shell of f-block elements is ns^2 , followed with $(n-2)f^{1-14}$, $(n-1)d^{0-2}$.
- (b) All are metals.
- (c) Lanthanoids are also known as **rare earth elements** whereas most of the members of actinoid series are known as **transuranic elements** (made artificially).
- (d) These show variable valency.
- (e) These form coloured ions.
- (f) Actinoids are radioactive.
- (g) These also form complexes.

5. POSITION OF METALS AND NON METALS IN PERIODIC TABLE

Metals, Non Metals and Metalloids in Periodic Table

(a) Trends in metallic character in Periodic table

- (i) The metallic character increases down the group and decreases along the period.
- (ii) The non-metallic character decreases down the group and increases along the period.

Note: All the non-metals and metalloids belong to p-block (except H and He).

6. EFFECTIVE NUCLEAR CHARGE AND SHIELDING

In a polyelectronic atom, the internal electrons repel the electrons of the outermost orbit. This results in the decrease in the nuclear attraction on the electrons of the outermost orbit.

Therefore, only a part of the nuclear charge is effective on the electrons of the outermost orbit. Thus, the inner electrons protect or shield the nucleus and thereby decrease the effect of nuclear charge towards the electrons of the outermost orbit.

Thus, part of the nuclear charge works against outer electrons, and is known as effective nuclear charge $Z^* = Z - S$
 Z^* = effective nuclear charge, S = shielding constant and Z = nuclear charge

A scientist named Slater, determined the value of shielding constant and put forward some rules which are listed below:

1. The shielding effect or screening effect of each electron of 1s orbital is 0.30.
2. The shielding effect of each electrons of ns and np i.e. electron of the outermost orbit, is 0.35.
3. The shielding effect of each electron of s, p or d orbitals of the penultimate orbit ($n - 1$) is 0.85.
4. The shielding effect of each electron of s, p, d or f orbital of the inner penultimate orbit ($n - 2$) and below this is 1.0

Table 13.3: Z^* for II period elements

Element	Atomic number	Calculation of			$\frac{349}{35.3}$	$Z^* = Z - S$
		$[0.35 \times \text{No. of nth electrons} - 1]$	$[0.85 \times \text{No. of (n-1)th electrons}]$	$[1.0 \times \text{No. of inner electrons}]$		
Li	3	—	0.85×2	—	1.70	1.30
Be	4	0.35×1	0.85×2	—	2.05	1.95
B	5	0.35×2	0.85×2	—	2.40	2.60
C	6	0.35×3	0.85×2	—	2.75	3.25
N	7	0.35×4	0.85×2	—	3.10	3.90
O	8	0.35×5	0.85×2	—	3.45	4.55
F	9	0.35×6	0.85×2	—	3.80	5.20
Ne	10	0.35×7	0.85×2	—	4.15	5.85

Illustration 5: What is the screening constant for the last electron in Sc?

(JEE MAIN)

Sol: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$

Last electron is in 4s orbital.

The shielding effect for one electron of 4s = 0.35

Electrons of 3rd shell = 9; their contribution = 9×0.85

Contribution of 2nd and 1st shell = 10×1

Total = $0.35 + 9 \times 0.85 + 10 \times 1 = 18$

$$\frac{349}{35.3} = 18$$

$$z^* = z - S = 21 - 18 = 3$$

7. TRENDS IN PHYSICAL PROPERTIES

7.1 Atomic Radius

It refers to the distance between the centre of the nucleus of the atom to the outermost shell containing electrons. Since absolute value of the atomic size cannot be determined, it is usually expressed in terms of the following operational definitions.

(a) Covalent Radius

- (i) Normally, this term is used for non-metals.
 - (ii) It is defined as half of the distance between two successive nuclei of two covalently bonded atoms in a molecule.
 - (iii) bonded atoms in a molecule.
- $$\text{Covalent radius} = \frac{1}{2} \times \text{Internuclear distance between two covalently bonded like atom}(d)$$

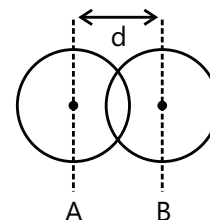


Figure 13.2-a: Diagrammatic representation of covalent radius

In other words, covalent radius is the radius of a spherical atom that leads to observed bond lengths when the spheres are just touching each other.

Covalent radius may be of following types

Single bond covalent radius

Double bond covalent radius

(b) Vander Waals Radius

- (i) Van der Waals radius is defined as half of the internuclear separation of two non-bonded atoms of the same element on their closest possible approach. The term is used for non-metals (covalent compound) and noble gases.
- (ii) It is half of the distance between two successive nuclei of two covalently bonded molecules of like atoms or two successive molecules of inert gases.

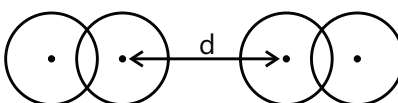


Figure 13.2-b: Diagrammatic representation of vander waal radius

Vander Waals radius = $\frac{1}{2} \times$ Internuclear distance between two successive nuclei of two covalent molecules (d)

(c) Crystal Radius or Metallic Radius

- (i) The term is usually used for metals.
- (ii) It is defined as half of the distance between two successive nuclei of two adjacent metal atoms in the metallic closed packed crystal lattice.

(d) Ionic Radius

- (i) This term is used in case of ions.
- (ii) It is the distance of outermost shell of an anion or cation from its nucleus. In other words, it is defined as the effective distance from the nucleus of the ion which is under influence in an ionic bond.

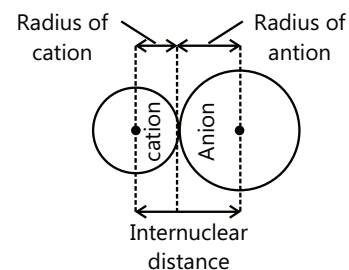


Figure 13.2-c: Diagrammatic representation of ionic radius

PLANCESS CONCEPTS

It is evident that van der Waals radius is greater than Covalent radius of an atom because Vander Waals forces of attraction are weaker than covalent bond forces.

Vaibhav Krishnan JEE 2009, AIR 22

Trends of Atomic Radius

- (a) Along the period: On moving across a period, atomic radii decreases because effective nuclear charge increases.
- (b) Down the group: On moving down a group, atomic radii increases, because number of orbits increases.

Factors Affecting Atomic Radii

- (a) **Effective nuclear charge:** As the effective nuclear charge increases, the attractive force between nucleus and valence electron increases. Thus, across a period, atomic size/atomic radii decreases.

$$\text{So, Atomic radii} \propto \frac{1}{Z_{\text{eff}}}$$

- (b) **Size of valence shell:** Atomic radii is the measure of radius of valence shell. As the value of n (principal quantum no.) increases, for an orbit, its size increases, thus down a group, atomic radii increases.
- (c) **Multiplicity of bond:** Covalent radii decreases, as the multiplicity of bond increases.

For example, in case of carbon

	C – C	C = C	$\Delta_{\text{A-B}}$
$D_{\text{C-C}}$ (Å)	1.54	1.34	1.20
r_{C} (Å)	0.77	0.67	0.60

- (d) **Percentage ionic character in bond:** Covalent radii of an atom in a bond depends upon % of ionic character. Increase in ionic character % leads in shortening of bond, decreasing the atomic radii.
- (e) **Cationic Radii:** Size of cation is always lesser than its parent atom and greater the charge on cation, smaller is its ionic radii. E.g. $\text{Fe} > \text{Fe}^+ > \text{Fe}^{2+} > \text{Fe}^{3+}$ (decreasing ionic radii)
- Formation of cation involves loss of electron. Thus, effective nuclear charge increases, pulling the remaining electrons more tightly towards the nucleus.
- (f) **Anionic Radii:** Size of an anion is always larger than its parent atom. Formation of an anion involves gain of electrons by an atom and so, effective nuclear charge decreases. Thus, the valence shell electrons are less tightly held by the nucleus.

PLANCESS CONCEPTS

Lanthanide contraction also plays an important role in deciding the trends of atomic radii accounting for the similar atomic radii of palladium and Pt.

Nikhil Khandelwal JEE 2009, AIR 94

Illustration 6: Which one of the following pairs would have a large size? Explain.

- (i) K or K⁺ (ii) Br or Br⁻ (iii) O²⁻ or F⁻ (iv) Li⁺ or Na⁺ (v) P or As (vi) Na⁺ or Mg²⁺

(JEE MAIN)

Sol:

- (i) Due to higher effective nuclear charge, K⁺ has smaller atomic size than K.
 (ii) Due to lower effective nuclear charge, the size of Br⁻ is greater than that of Br.
 (iii) O²⁻ and F⁻ are isoelectronic species. Since effective nuclear charge of O²⁻ is lower than that of F⁻, therefore, O²⁻ has a greater atomic size than F⁻.
 (iv) Li⁺ and Na⁺ both belong to group 1. Because of a greater number of shells (2, in case of Na⁺ and 1, in case of Li⁺), Na⁺ has a bigger atomic size than Li⁺.
 (v) As has four shells while P has three. Therefore, atomic size of As is greater than that of P.
 (vi) Na⁺ and Mg²⁺ are isoelectronic cations. Therefore, due to lower effective nuclear charge, the ionic radius of Na⁺ is greater than that of Mg²⁺.

Illustration 7: Arrange the following in order of increasing radii?

- (i) I, I⁺, I⁻ (ii) C, N, Si, P (iii) O²⁻, N³⁻, S²⁻, F⁻

(JEE MAIN)

Sol: The size of a neutral, positive and negative species is compared.

- (i) Size of a cation is always smaller while that of an anion is always bigger than the neutral atom, i.e. I⁺ < I < I⁻
 (ii) C and N lie in 2nd period while Si and P lie below them in the 3rd period. Since elements in the 3rd period have higher atomic size than those in the 2nd period. Therefore, atomic radii of Si and P are higher than those of C and N respectively. Since, atomic radii decreases across a period due to higher nuclear charge, therefore, C has higher atomic radius than N and Si has higher atomic radius than P. Thus, the overall order of increasing atomic radii is: N < C < P < Si.
 (iii) Among isoelectronic ions, the size of anions increases as the nuclear charge decreases: F⁻ < O²⁻ < N³⁻ < S²⁻.

Illustration 8: Select from each group the species which has the smallest radius stating appropriate reasons.

- (i) O, O⁻, O²⁻ (ii) P³⁺, P⁴⁺, P⁵⁺

(JEE ADVANCED)

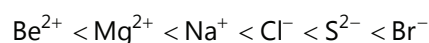
Sol: Ionic radius decrease on loss of electrons since the nuclear charge increases.

- (i) O has the smallest radius. The anion is larger than its parent atom. Also, the anion of the same atom with higher negative charge is bigger in size as compared to the anion with smaller negative charge as proton to electron ratio decreases. Thus, attraction between valence shell electrons and nucleus decreases. Hence, the electron cloud expands.
 (ii) The ionic radius decreases as more electrons are ionized off. Thus, valency increases. So, the correct order is P⁵⁺ < P⁴⁺ < P³⁺.

Illustration 9: Arrange the following ions in the increasing order of their size: Be²⁺, Cl⁻, S²⁻, Na⁺, Mg²⁺, Br⁻?

(JEE ADVANCED)

Sol: Ionic radius $\propto \frac{1}{\text{Nuclear charge}}$



Be²⁺ is smaller than Mg²⁺, as, Be²⁺ has one shell whereas, Mg²⁺ has two shells. Mg²⁺ and Na⁺ are isoelectronic species.

Cl⁻ is smaller than Br⁻ as Cl⁻ has three shells whereas Br⁻ has four shells.

PLANCESS CONCEPTS

- In isoelectronic ions, size decreases with increase in atomic number.

For e.g. $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

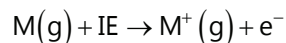
Although, Ne is also isoelectronic to them but its atomic radius is 1.60Å. It should not be compared to other radii because noble gases do not form ionic compounds and their radii are simply Van der Waals radii.

- In the end of the period, the atomic radii of inert gases are exceptionally higher because they do not form molecules and their radii are simply Van der Waals radii.

Saurabh Gupta JEE 2010, AIR 443

7.2 Ionization Potential

It is the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom i.e.



The amount of energies required to remove the first, second, third etc. electrons from an isolated gaseous atom are called successive ionization energies and are designated as $IE_1, IE_2,$

IE_3 etc. It may be noted that IE_2 is always greater than IE_1 . Thus, the order is- $IE_3 > IE_2 > IE_1$

PLANCESS CONCEPTS

The removal of a second electron is relatively more difficult because after the removal of the first electron, remaining electrons in the cation are more effectively pulled by the nucleus due to increased effective nuclear charge, thus, $IE_3 > IE_2 > IE_1$

Neeraj Toshniwal (JEE 2009, AIR 21)

Factors affecting Ionisation Potential

- Number of shells:** With the increase in the number of shells, the atomic radius increases i.e. the distance of outermost shell electron from the nucleus increases and hence the ionization potential decreases.
- Effective Nuclear Charge:** Atomic size decreases with the increase in effective nuclear charge because, higher the effective nuclear charge, stronger will be the attraction of the nucleus towards the electron of the outermost orbit and higher will be the ionization potential.
- Shielding Effect:** The electrons of the inner orbits repel the electrons of the outermost orbit due to which the attraction of the nucleus towards the electrons of the outermost orbit decreases and thus the atomic size increases and the value of ionization potential decreases.
- Stability of half-filled and fully filled orbitals:** The atoms whose orbitals are half-filled (p^3, d^5, f^7) or fully-filled (s^2, p^6, d^{10}, f^{14}) have greater stability than the other. Therefore, they require greater energy to remove an electron. However, stability of fully filled orbitals is greater than that of the half-filled orbitals.
- Penetration power:** In any atom, the s-orbital is nearer to the nucleus in comparison to p, and f orbitals. Therefore, greater energy is required to remove an electron from s-orbital than from p, d and f orbitals. The order is as follows- $s > p > d > f$

Periodic Trends in Ionisation Potential

In a Period: The value of ionization potential normally increase across a period, because effective nuclear charge increases and the atomic size decreases.

Exceptions: In the second period, ionization potential of Be is greater than that of B, and in the third period, ionization potential of Mg is greater than that of Al due to the high stability of fully filled orbitals. In the second period, ionization potential of N is greater than O and in the third period, ionization potential of P is greater than that of S, due to the stability of half-filled orbitals.

In a Group: The value of ionization potential normally decreases down the group because both, atomic size and shielding effect increase.

Exception: The value of ionization potential remains almost constant from Al to Ga in the III A group. ($B > Al$, $Ga > In$).

In IV B group i.e. Ti, Zr and Hf, the I.P. of Hf is higher than that of Zr due to Lanthanide contraction. Thus the I.P. of IV B group varies as $Ti > Zr < Hf$.

Ionisation Potential of Transition Elements: In transition elements, the value of ionization potential has changes very little across a period. This is because, the outermost orbit remains the same but electrons get filled up in the $(n-1)d$ orbitals resulting in very little increase in the values of ionization potential.

In transition element series, the first ionization potential normally increases with increase in atomic number on going from left to right, but this periodicity is not uniform. The value of ionization potential of transition elements depends on two important factors-

- (a) The value of ionization potential increases with increase in effective nuclear charge.
- (b) The value of ionization potential decreases with increase in shielding effect when number of electrons increase in $(n - 1)$ orbitals.
 - (i) In the first transition element series, the first ionization potential normally increases on going from left to right from Sc to Cr because shielding effect is much weaker in comparison to effective nuclear charge. The value of first ionisation potential of Fe, Co and Ni remains constant, because shielding effect and effective nuclear charge balance each other. The value of ionization potential shows a slight increase from Cu to Zn because they have fully filled s and d orbitals. The value of first ionisation potential of Mn is maximum because it has maximum stability due to fully filled s and half-filled orbitals.

Inner Transition Elements: The size of inner transition elements is greater than that of d block elements. Therefore, the value of ionization potential of f- block elements is smaller than that of d- block elements. Due to the almost constant atomic size of f- block elements in a period, the value of their ionisation potential remains more constant than that of d- block elements.

PLANCESS CONCEPTS

Ionization energy and the stable oxidation states of elements:

- (i) When difference in two successive IE values (Δ^2) for an atom is approximately 10-15 eV or less, then, the higher ox. state will be more stable.

E.g.

$$\text{For Al, } I.E_1 = 6.0 \text{ eV. } I.E_2 = 18.8 \text{ eV,}$$

$$I.E_3 = 28.4 \text{ eV and } I.E_4 = 120 \text{ eV,}$$

Then, for Al(II); $\Delta I.E_{(1,2)} = 12.8 \text{ eV}$

$$\text{Al (III); } \Delta I.E_{(2,3)} = 9.6 \text{ eV}$$

$$\text{Al (IV); } \Delta I.E_{(3,4)} = 91.6 \text{ eV}$$

PLANCESS CONCEPTS

So, Al (III) is more stable than Al (I) or Al (II).

(ii) If the value of ∞ is greater than 16.0 eV, then, the lower ox. state will be more stable.

E.g., For Na, I.E.₁ = 5.1 eV; I.E.₂ = 47.3 eV

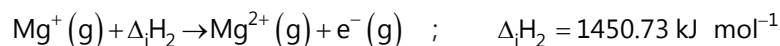
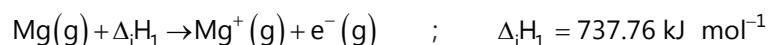
$$\therefore \Delta IE_{(1,2)} = 42.4 \text{ eV}$$

So, Na(I) is formed and not Na(II).

Aman Gour JEE 2012, AIR 203

Illustration 10: Calculate the energy required to convert all the atoms of magnesium to magnesium ions present in 24 mg of magnesium vapours? First and second ionization enthalpies of Mg are 737.76 and 1450.73 kJ mol⁻¹ respectively. **(JEE MAIN)**

Sol: According to the definition of successive ionization enthalpies,



$$\Delta \text{ Total amount of energy needed to convert Mg (g) atom into Mg}^{2+}(\text{g}) \text{ ion} = \Delta_i H_1 + \Delta_i H_2 \\ = (737.76 + 1450.73) \text{ kJ mol}^{-1} = 2188.49 \text{ kJ mol}^{-1}$$

$$24 \text{ mg of Mg} = \frac{24}{1000} \text{ g} = \frac{24}{1000 \times 24} \text{ mole} = 10^{-3} \text{ mole}$$

Therefore, amount of energy needed to ionize 10⁻³ mole of Mg vapours = 2188.49 × 10⁻³ = 2.188 kJ.

Illustration 11: The $\Delta_i H_1$ and $\Delta_i H_2$ of Mg(g) are 740 and 1450 kJ mol⁻¹. Calculate the percentage of Mg⁺(g) and Mg²⁺(g) if 1 g of Mg(g) absorbs 50 kJ of energy. **(JEE MAIN)**

Sol: $\Delta_i H_1$ and $\Delta_i H_2$ is the ionization enthalpy i.e. heat energy used to remove the loosely bound electron.

No. of moles of Mg vapours present in 1 g = 1/24 = 0.0417

Energy absorbed in the ionization of 0.0417 mole of Mg(g) to Mg⁺(g) = 0.0417 × 740

= 30.83 kJ; Energy left unused = 50 – 30.83 = 19.17 kJ

Now, 19.17 kJ will be used to ionize Mg⁺(g) to Mg²⁺(g)

$$\underbrace{\text{Na} \quad \text{Mg} \quad \text{Al}}_{\text{Metals}} \quad \text{No. of moles of Mg}^+(\text{g}) \text{ converted into Mg}^{2+}(\text{g}) = 19.17/1450 = 0.0132$$

No. of moles of magnesium ions left as Mg⁺(g) = 0.0417 – 0.0132 = 0.0285

$$\underbrace{\text{Si}}_{\text{Metalloid}} \quad \% \text{ of Mg}^+(\text{g}) = (0.0285/0.0417) \times 100 = 68.35\% \text{ and } \% \text{ of Mg}^{2+}(\text{g}) = 100 - 68.35 = 31.65\%$$

Illustration 12: The first ($\Delta_i H_1$) and the second ($\Delta_i H_2$) ionization enthalpies (kJ mol⁻¹) of a few elements designated by Roman numerals are shown below: **(JEE ADVANCED)**

Element	$\Delta_i H_1$	$\Delta_i H_2$
I	2372	5251
II	520	7300

III	900	1760
IV	1680	3380

Which of the above elements is likely to be (a) a reactive metal (b) a reactive non-metal (c) a noble gas (d) a metal that forms a stable binary halide of the formula AX_2 ($X = \text{halogen}$).

Sol:

- (i) Since Element II has a very low $\Delta_i H_1$ but a very high $\Delta_i H_2$, therefore, it has only one electron in the valence shell and hence is likely to be a reactive metal (i.e., an alkali metal).
- (ii) Since the $\Delta_i H_1$ of Element IV is very high and its $\Delta_i H_2$ is not so high (actually almost double), IV is likely to be a reactive non-metal (i.e. a halogen).
- (iii) Among the elements listed, $\Delta_i H_1$ of element I is the highest and its, $\Delta_i H_2$ is also not so high, therefore, it must be a noble gas.
- (iv) The $\Delta_i H_1$ of element III is higher than that of element II, but unlike element II, its $\Delta_i H_2$ is only about twice its $\Delta_i H_1$, therefore, it is likely that element III has two electrons in the valence shell (i.e., alkaline earth metal). As such it will form a stable binary halide of the formula AX_2 where A is the metal and X is the halogen.

Illustration 13: From each set, choose the atom which has the largest ionization enthalpy and explain your answer.

- (i) F, O, N (ii) Mg, P, Ar (iii) B, Al, Ga

(JEE MAIN)

Sol: Largest ionization enthalpy is the highest amount of energy needed to remove the valence electron due to a stronger nuclear charge and a smaller atomic size.

- (i) F, O, N-All belong to 2nd period. Among these, F has the highest $\Delta_i H_1$ because of its smallest size and highest nuclear charge.
- (ii) Mg, P, Ar-All lie in the 3rd period. Among these, Ar has the highest $\Delta_i H_1$ because it has stable inert gas configuration.
- (iii) B, Al, Ga-All lie in group 13. B has the highest $\Delta_i H_1$ due to its smallest size.

Illustration 14: Compare qualitatively the first and second ionization potentials of copper and zinc. Explain the observation.

(JEE ADVANCED)

Sol:

	IE_1 kJ mol ⁻¹	IE_2 kJ mol ⁻²
Cu	744	1961
Zn	906	1736

IE_1 of copper is less than that of zinc, because removal of electron takes place from $4s^1$ (attaining a more stable configuration $3d^{10}$) whereas, in case of zinc, it is from completely filled $4s^2$ (attaining the configuration $4s^1$)

IE_2 of copper is higher than zinc, because the removal of 2nd electron from a stable configuration (d^{10}) requires higher energy.

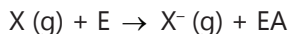
Illustration 15: The first four successive ionization energies for an element are (6.113, 11.871, 50.908, 67.01) respectively. What is the number of valence shell electrons?

(JEE ADVANCED)

Sol: The difference in second and third ionization is very large. Therefore, the no. of valence shell electrons should be 2. The element would attain the noble gas configuration after losing these 2 electrons.

7.3 Electron Affinity

It is the amount of energy released when a neutral isolated gaseous atom accepts an electron to form a gaseous anion.

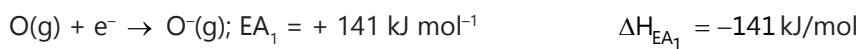


Similarly, second and third electron can be added to form gaseous dinegative and trinegative ions. The energy changes accompanying the addition of first, second, third etc. electrons to neutral isolated gaseous atoms are called successive electron affinities and are designated as EA_1 , EA_2 , EA_3 , etc.

Since an atom has a natural tendency to accept an electron, therefore, the first electron affinity (EA_1) is always taken as positive. However, the addition of second electron to the negatively charged ion is opposed by coulombic repulsion. Hence, energy has to be supplied for the addition of second electron.

Thus, second electron affinity (EA_2) of an element is taken as negative.

For example,



Factors Affecting Electron Affinity

Atomic size or atomic radius: When the atomic size/radius increases, the electrons entering the outermost orbit is more weakly attracted by the nucleus and the value of electron affinity is lower.

Effective Nuclear charge: When effective nuclear charge is more, then, the atomic size is less. Then, the atom can easily gain an electron and possess a higher value of electron affinity.

Stability of Fully-Filled and Half-Filled orbitals: The stability of the configuration having fully-filled orbitals (p^6 , $d^{10}f^{14}$) and half-filled orbitals (p^3 , d^5 , f^7) is relatively higher than that of other configurations. Hence, such type of atoms have a lesser tendency to gain an electron, therefore, their electron affinity values will be very low or zero.

Trends in Electron Affinity: In a period, atomic size decreases with the increase in effective nuclear charge and hence, increases the electron affinity.

Exception:

- (a) Ongoing from C^6 to N^7 in the second period, the values of electron affinity decrease instead of increasing. This is because there are half-filled ($2p^3$) orbitals in the outermost orbit of N, which are more stable. On the other hand, the outermost orbit in C has a $2p^2$ configuration.
- (b) In the third period, the value of electron affinity of Si is greater than that of P. This is because the electronic configuration of the outermost orbit in P atom is $3p^3$, which being half-filled, is relatively more stable. The values of electron affinity of inert gases are zero, because their outermost orbit has fully-filled p orbitals.
- (c) In a period, the value of electron affinity goes on decreasing on going from group IA to group IIA. The value of electron affinity of the elements of group IIA is zero because ns orbitals are fully-filled and such orbitals have no tendency to accept electrons.

In a Group: The values of electron affinity normally decrease down a group because the atomic size increases, decreasing the actual attractive force of the nucleus.

Exceptions:

- (a) The value of the electron affinity of F is lower than that of Cl, because the size of F is very small and compact and the charge density is high on the surface. Therefore, the incoming electron/s experience more repulsion in comparison to Cl accounting for the highest value of Cl in the periodic table.

- (b) The values of electron affinity of alkali metals and alkaline earth metals can be regarded as zero, because they do not have the tendency to form anions by accepting electron/s.

PLANCESS CONCEPTS

- (a) Cl has the highest electron affinity (3.7)
- (b) Higher the EA of an element, easier is the addition of an electron.
- (c) Following are some important observations derived from the general trend of electron affinity:
- (i) More the tendency to gain an electron, more is the non-metallic nature. Therefore, non-metallic nature increases along the period but decreases down the group.
 - (ii) More the tendency to form an anion, more is the tendency to show ionic bonding. Therefore, tendency of non-metals to show ionic bonding increases along the period but decreases down the group.
 - (iii) More the tendency to get reduced, more is the oxidizing nature. Therefore, oxidizing power increases along the period but decreases down the group

Order of Oxidizing power- $F > Cl > Br > I$.

B Rajiv Reddy JEE 2012, AIR 11

Illustration 16: The electron gain enthalpy of chlorine is -349 kJ mol^{-1} . How much energy in kJ is released when 3.55 g of chlorine is converted completely into Cl^- ion in the gaseous state? **(JEE MAIN)**

Sol: According to the definition of electron gain enthalpy, $\text{Cl}(\text{g}) + \text{e}^-(\text{g}) \rightarrow \text{Cl}^-(\text{g}) + 349 \text{ kJ mol}^{-1}$

$\text{C} \equiv \text{C}$ Energy released when 1 mole (=35.5 g) of chlorine atoms change completely into $\text{Cl}^-(\text{g}) = 349 \text{ kJ}$

Energy released when 3.55 g of chlorine atoms change completely into $\text{Cl}^-(\text{g}) = \frac{349}{35.5} \times 3.55 = 34.9 \text{ kJ}$

Illustration 17: Which of the following pairs of elements would have more negative electron gain enthalpy? Explain (i) N or O (ii) S or O (iii) C or Si **(JEE ADVANCED)**

Sol: (i) The electron gain enthalpy of O is highly negative while that of N is slightly positive.

Reason: The electronic configuration of N is quite stable ($1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$) since it has exactly half-filled 2p-orbitals and hence has no tendency to accept an extra electron. In other words, energy has to be supplied to add an extra electron. Thus, electron-gain enthalpy of N is slightly positive. In contrast, the electronic configuration of O ($1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$) is not so stable but it has a higher nuclear charge and smaller atomic size than N and hence, it has a higher tendency to accept an extra electron. In other words, electron gain enthalpy of O is highly negative.

(ii) S has more negative electron gain enthalpy than O.

Reason: The size of O is much smaller than that of S. As a result, the electron-electron repulsions in the smaller 2p-subshell of O are comparatively more than those present in the bigger 3p-subshell of S. Therefore, S has a higher tendency to accept an additional electron than O.

(iii) C has a more negative electron gain enthalpy than Si.

Reason: This is because C-atom has a smaller size than Si-atom. (Note that the electron-electron repulsions in these atoms are not very large because they contain only 4 electrons in the outermost shell.)

Illustration 18: (i) Arrange the following elements in order of decreasing electron gain enthalpy: B, C, N, O.

(JEE ADVANCED)

Sol: N has +ve electron gain enthalpy while, all others have -ve electron gain enthalpies. While moving from B → C → O, size decreases leading to -ve electron gain enthalpy in the same order. Thus, the overall decreasing order of electron gain enthalpies is N, B, C, O.

7.4 Electronegativity

The tendency of an atom to attract the shared pair of electrons of the covalent bond towards itself is called electronegativity of that atom.

Factors Affecting Electronegativity

Atomic size: Electronegativity of a bonded atom decreases with increase in size since the attractive force on the valence electrons decreases and hence electronegativity decreases.

Hybridisation state of atom: Electronegativity increases with increase in the s-character of the hybrid orbital. This is because, the s-orbital is nearer to the nucleus and thus, suffers greater attraction leading to increased electronegativity. The number of covalent bonds present between two bonded atoms is known as its bond order. With increase in the bond order, the bond distance decreases, effective nuclear charge increases and thus electronegativity increases. Increasing order of electronegativity is as follows: C – C < C = C < C ≡ C. When effective nuclear charge is high, the nucleus will attract the shared electrons with greater strength to give high electronegativity.

Oxidation number: The electronegativity value increases with the increase in oxidation number since the radius decreases with the increase in oxidation number.

The increasing order of electronegativity is as follows: Fe < Fe⁺² < Fe⁺³

Electronegativity does not depend on stability of fully-filled or half-filled orbitals because it is simply the capacity of the nucleus to attract a bonded pair of electrons.

Trends in Electronegativity: Atomic size decreases across a period. Thus, electronegativity increases. Atomic size increases down a group decreasing the electronegativity.

F has maximum electronegativity value in the periodic table, while Cs has minimum.

According to the Pauling scale, the electronegativity value of F is 4.0, O is 3.5 N is 3.0 and Cl is 3.1

Exceptions

- (a) The elements of group IIB i.e. Zn, Cd and Hg show increase in electronegativity value down the group.
- (b) The elements of group IIIA, i.e Al to Ga show increase in electronegativity value down the group.
- (c) The elements of group IVA, Si onwards, show no change in electronegativity value down the group.

Measurement of Electronegativity

Pauling Scale: If two atoms, A and B, having different electronegativity values, get bonded to form a molecule, AB, then the bond between A and B in A–B will have both covalent and ionic properties.

$$\Delta_{A-B} = \text{Observed bond energy} - \text{Energy of 100\% covalent} \text{ or } \Delta_{A-B} = D - E_{A-B}$$

Where D = Observed bond energy

E_{A-B} = Bond energy of pure covalent bond of A – B

The value of E_{A-A} and E_{B-B} is $E_{A-B} = \frac{1}{2} [E_{A-A} + E_{B-B}]$

$$= 0.208 \sqrt{\Delta_{A-B}} = X_A - X_B \quad \text{where, } X_A > X_B \text{ or } 0.043 \times \Delta_{A-B} = (X_A - X_B)^2$$

Mulliken Electronegativity Scale: Mulliken suggested that the value of electronegativity of an element as an average of the values of ionization potential and electron affinity of the element.

$$X_M = \frac{\text{I.P.} + \text{E.A.}}{2} \text{ (in eV)}$$

Where X_M = Electronegativity value as given by Mulliken

$$X_P = \frac{X_M}{2.8} = \frac{\text{I.P.} + \text{E.A.}}{5.6}$$

Where X_P = Electronegativity value as given by Pauling or $X_P = 0.336 (X_M - 0.615)$

Allred-Roschow's Scale

$$X_{AR} = \frac{Z_{\text{eff}} \cdot e^2}{r^2}$$

$$X_P = 0.359 \frac{Z_{\text{eff}} \cdot e^2 + 0.744}{r^2}$$

$$Z_{\text{eff}} = Z - s^2, p^6, d^{10}, f^{14}$$

where Z = Nuclear charge

ΔIE = Shielding constant

$$\text{Or } X_P = 0.359 X_{AR} + 0.744$$

Sanderson's Scale: In Sanderson scale, the stability ratio of an atom itself has been regarded as its electronegativity.

$$X_s \text{ or S.R.} = \frac{\text{Average electron density of an atom}}{\text{Electron density of the isoelectronic inert gas}}$$

This is related to Pauling scale as follows-

$$\sqrt{X_P} = \frac{0.2}{\text{S.R. or } X_s} \times 0.77$$

Applications of Electronegativity

(a) Nomenclature: Name of the more electronegative element is written at the end and 'ide' is suffixed to it. The name of the less electronegative element is written before the name of more electronegative element of the formula.

E.g. Correct formula

Name

- | | |
|-----------------------|-------------------|
| (a) IBr | Iodine bromide |
| (b) OF ₂ | Oxygen difluoride |
| (c) Cl ₂ O | Dichlorine oxide |

(b) Nature of bond: If the electronegativity difference between the two elements is 1.7 or more, then, an ionic bond is formed between them. Whereas, if it is less than 1.7, then, covalent bond is formed.

(HF is an exception in which a bond is covalent, although, difference of electronegativity is 1.9)

(c) Metallic and non-metallic nature: Low EN shows metallic nature and high EN shows non-metallic nature

Hydrolysis of AX – where A = Other element
and X = Halogen

If electronegativity of X > Electronegativity of A, then, on hydrolysis, product will be HX.

Example In BCl₃, EN of Cl > EN of B

- (d) **Partial ionic character in covalent bonds:** Partial ionic character is generated in covalent compounds due to the difference in electronegativities.

Hanny and smith: Calculates percentage of ionic character from the electronegativity difference. Percentage of ionic character = $16(X_A - X_B) + 3.5(X_A - X_B)^2$

$$= 16\Delta + 3.5\Delta^2$$

$$= (0.16\Delta + 0.035\Delta^2) \times 100$$

Here, X_A is electronegativity of element A.

X_B is electronegativity of element B.

$$\Delta = X_A - X_B$$

- (e) **Bond length:** $d_{A-B} = r_A + r_B - 0.09(X_A - X_B)$

$$\text{or } d_{A-B} = \frac{1}{2}(D_{A-A} + D_{B+B}) - 0.09(X_A - X_B)$$

Here, $X_A > X_B$

- (f) **Bond energy and stability:** Bond strength and stability of A–B increases on increase in difference of electronegativities of atoms A and B. Therefore, H–F > H–Cl > H–Br > H–I

- (g) **Acidic strength of hydrides:** Bond energy (Strength) \propto stability of molecule.

Order of stability of hydrohalides is HF > HCl > HBr > HI

So, Order of their acidic strength will be – HF < HCl < HBr < HI

Down the 5A group

NH ₃ *Thermal		stability decreases
PH ₃ *Basic	↓	character decreases
AsH ₃ *Acidic	↓	character increases

In PH₃ and AsH₃, the difference in the electronegativities of X_A and X_B is very less, so their bond energy decreases and hence acidic character (losing H⁺ ion) increases.

- (h) **Nature of oxides:** If the difference of the two electronegativities ($X_O - X_A$) is 2.3 or more than 2.3, then the oxide will be basic in nature. Similarly, if value of $X_O - X_A$ is lower than 2.3, then the compound will be first amphoteric and then acidic in nature.

Oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₂ O ₅	SO ₃	Cl ₂ O ₇
($X_O - X_A$)	2.6	2.3	2.1	1.8	1.5	1.1	0.5
Nature	Strongly basic	Basic	Amphoteric	weakly acidic	Acidic	Strong acidic	Strongest acidic

- (i) **Nature of hydroxides:** According to Gallis, if electronegativity of A in a hydroxide (AOH) is more than 1.7, then, it will be acidic in nature whereas, it will be basic in nature, if electronegativity is less than 1.7.

For example	NaOH	and	ClOH
Electronegativity (X_A)	0.9		3.00
Nature	Basic		Acidic

If the value is more than $X_O - X_H$, then that hydroxide will be basic, otherwise, it will be acidic.

Illustration 19: Among the following, how many elements have lower electronegativity than oxygen atom-F, Cl, Br, I, H, S, P, K, Ca **(JEE MAIN)**

Sol: F is the most electronegative element. Order of electronegativity- $F > O > N = Cl > \dots$ Except F, all will have electronegativity less than oxygen.

Illustration 20: What is the difference in the electronegativity of two atoms, when the percentage ionic character is 19.5%? **(JEE ADVANCED)**

Sol: Hanny – Smith equation

$$\text{Percentage of ionic character} = (16\Delta + 3.5\Delta^2)$$

Δ = difference in electronegativity

$$19.5 = 16\Delta + 3.5\Delta^2$$

$$3.5\Delta^2 + 16\Delta - 19.5 = 0$$

Solving quadratic, we get, $\Delta = 1$

8. TRENDS IN CHEMICAL PROPERTIES

Valency/valence, also known as **valence number**, is the number of valence bonds a given atom has formed, or can form, with one or more atoms.

Table 13.4: Variation of valence in a group:

Group	1	2	13	14	15	16	17	18
Number of valence electron/s	1	2	3	4	5	6	7	8
Valence	1	2	3	4	3,5	2,6	1,7	0,8

Variation of valence in a period: On moving along the period, the number of valence electrons increases from 1 to 8. Consequently, the valence of the elements with respect to hydrogen increases from 1 to 4 up to group IV and then decreases to 1 as shown in the table. However, valence with respect to oxygen increases from one to seven along the period.

Variation of valence in group: On moving down the group, the number of valence electrons remains the same. Therefore, all the elements in a group have the same valence. For example, elements of group I have valency 1 and elements of group II have valency 2.

Variation of valence in transition elements: Transition metals show variable valence of 1, 2 or 3 as they can use electrons from their outermost as well as penultimate shell, during chemical reactions as energy difference between them is small.

Some anomalous properties of second period elements

Consider the elements of II period.

II period: Li Be B C N O F Ne

The elements of gp. 1, 2 (Li & Be) and of gp. 13–17 (B to F) differ in many respects from other members of their group. (These points of difference will be later studied in detail.) Some anomalous properties of 2nd period elements are given below with their explanation.

(a) Covalence: The maximum covalency of 2nd period elements is four while, other members may also show higher covalency. e.g., BF_4^- exists but $[\text{BF}_6]^{3-}$ is known. Similarly OF_2 is known but OF_4 , OF_6 are not while, SF_4 are known, N is never pentavalent etc.

Explanation: These elements have only two shells in their atom and the valence shell contain 4 orbitals only (one 2s and three 2p) so, a maximum of four bonds can be formed. In the 3rd period and onwards, the valence shell contain empty d-orbitals also. So, covalency may be more than four.

(b) $p\pi - p\pi$ **Multiple bonding:** Bonds like C=C, C \equiv C, N \equiv N, C=O etc. exist due to $p\pi - p\pi$ multiple bonding. These elements are smaller in size and mostly electronegative in their respective groups thus forming multiple bonds.

(c) **Diagonal relationship:** Diagonal relationship between elements of II & III periods.

Li Be B C

Na Mg Al Si

Diagonal relationship between these elements can be explained on the basis of approximately similar charge/size ratio of diagonally related elements.

9. PERIODIC PROPERTIES OF ELEMENTS

9.1 Periodicity Along the Period

(a) **Ionization enthalpy:** Increases along the period (with exception).

(b) **Electron gain enthalpy:** Increases along the period (with exception).

(c) **Electronegativity:** Increases along the period.

(d) **Atomic radius:** Decreases along the period.

(e) **Ionic radius:** The radii of isoelectronic ions decrease with increase in atomic number.

(f) **Atomic volume:** Volume occupied by 1 g-atom of an element in solid state is Atomic Volume.

Atomic volume decreases up to metals and then increases:

	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic volume (cm ³)	24	14	10	12	17	16	19	23.7

(g) **Melting point, Boiling point:** Increases along the period for metals.

	Na	Mg	Al
M.P (°C)	98	649	660
B.P (°C)	883	1100	1800

(h) **Density:** Increases along the period.

(i) **Reducing behavior:** Decreases along the period.

Na	Mg	Al	Si	P	S	Cl
Strong reductants					Strong oxidants	

(j) **Metallic character:** Decreases along the period.

Na	Mg	Al	Si	P	S	Cl
Metals			Metalloid	Non-metals		

Increasing non-metallic nature
Decreasing metallic nature →

(k) **Electropositive character:** Decreases along the period.

(l) Nature of oxides: The basic character of oxides decreases and acidic character increases along the period.

Elements: Na Mg Al Si P S Cl

Oxides: $\underbrace{\text{Na}_2\text{O} \quad \text{MgO}}_{\text{Basic}} \quad \underbrace{\text{Al}_2\text{O}_3}_{\text{Amphoteric}} \quad \underbrace{\text{SiO}_2 \quad \text{P}_2\text{O}_5 \quad \text{SO}_3 \quad \text{Cl}_2\text{O}_7}_{\text{Acidic}}$

(m) Nature of hydrides: The basic character of hydrides decreases along the period.

(n) Valency: (i) Valency with respect to oxygen increases from one to seven along the period (table 1.9). (ii) Valency with respect to hydrogen increases first from one to four & then decreases to one (table 1.9)

Table 13.5: Valence of Elements

Group	I	II	III	IV	V	VI	VIII
Valency with respect to oxygen	1	2	3	4	5	6	7
Formula of the oxide	R_2O	RO	R_2O_3	RO_2	R_2O_5	RO_3	R_2O_7
Valency with respect to hydrogen	1	2	3	4	3	2	1
Formula of the hydride	RH	RH_2	RH_3	RH_4	RH_3	RH_2	RH

9.2 Periodicity along the Group

- (a) Ionisation enthalpy : Decreases
- (b) Electron gain enthalpy : Decreases
- (c) Electronegativity : Decreases
- (d) Atomic radii : Increases
- (e) Ionic radii : Increases
- (f) Atomic volume : Increases
- (g) M.P./B.P : Decreases
- (h) Density : Increases
- (i) Oxidant-Reductant nature :
- * Reducing nature of metals : Increases
- * Oxidizing nature of non-metals : Decreases
- (i) Metallic character : Increases
- (j) Electro positive character : Increases
- (k) Basic character of oxide : Increases
- (l) Basic character of hydride : Decreases

PLANCESS CONCEPTS

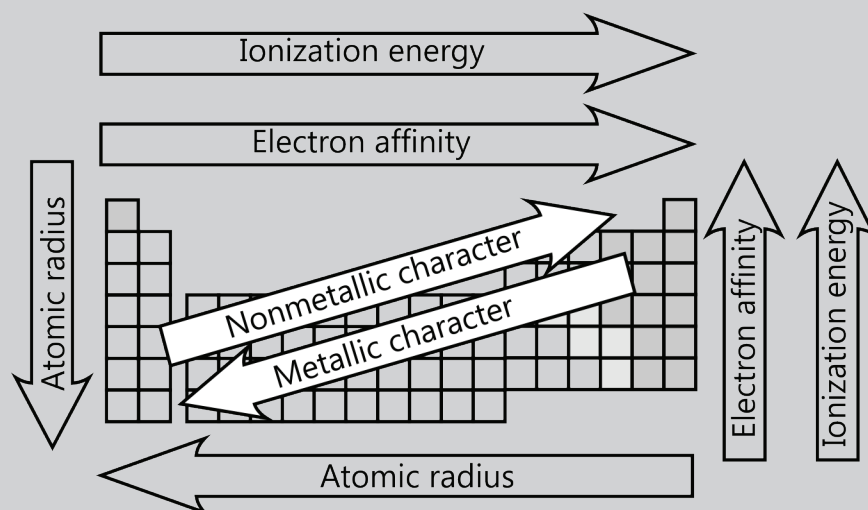


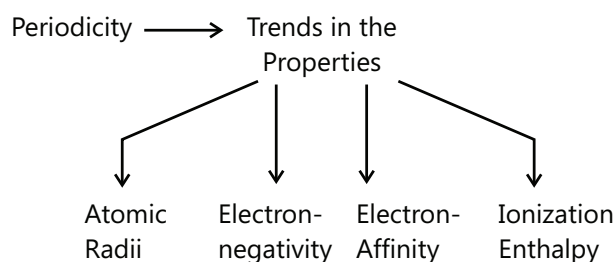
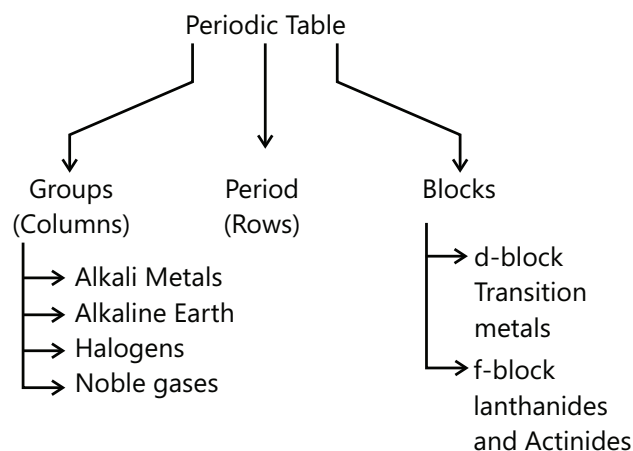
Figure 13.3: Variation of properties across the table

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Some Important Things to Remember

1. Liquid radioactive element: Francium (Fr)
2. Rarest element in earth crust: Astatine (At)
3. Most poisonous metal: Plutonium (Pu)
4. Element with the maximum number of natural isotopes: Tin(Sn)
5. First man-made element: Technetium (Tc)
6. The size of the largest atom (Cs) is approximately 4.4 times to that of the smallest (H) atom.
7. Out of 17 non-metals known, 11 are gases, one is liquid (Br) and 5 are solids (C, P, S, Si and I)
8. Element with highest M.P.: Carbon (C in diamond)
9. Metal with highest M.P.: Tungsten (W)
10. Metal with maximum density: Iridium(Ir) (22.61 g/cc)
11. Metal with minimum density: Lithium(Li)
12. Lightest element: Hydrogen(H)
13. Most acidic oxide: Dichlorine heptoxide (Cl_2O_7)

POINTS TO REMEMBER



JEE Main/Boards

Exercise 1

Q.1 The electron affinity of each one of Be and Mg is zero. Give reason.

Q.2 Explain the following

(i) The process $O(g) + e^- \rightarrow O^-(g) + 141 \text{ kJ mol}^{-1}$ gives a positive electron affinity though

$$\Delta E = -141 \text{ kJ mol}^{-1} \text{ (exothermic).}$$

(ii) The process $O^-(g) + e^- + 700 \text{ kJ mol}^{-1} \rightarrow O^{2-}(g)$ gives a negative electron affinity though ,

$$\Delta E = +700 \text{ kJ mol}^{-1} \text{ (endothermic).}$$

Q.3 There is an irregular trend in the ionization energies of group 13 elements. Explain.

Q.4 Mg^{2+} , O^{2-} , Na^+ , F^- , N^{3-} (Arrange in decreasing order of ionic size)

Q.5 Why Ca^{2+} has a smaller ionic radius than K^+ .

Q.6 Arrange in decreasing order of atomic size: Na, Cs, Mg, Si, Cl.

Q.7 Why the first ionization energy of carbon atom is greater than that of boron atom whereas, the reverse is true for the second ionization energy.

Q.8 (IE_1) of Be is greater than (IE_1) of B but (IE_2) of B is greater than that of Be. Explain.

Q.9 EA of Cu is 123 kJ mol^{-1} but that of Zn is -87 kJ mol^{-1} . Explain.

Q.10 In the preparation of hydrocarbon by Kolbe's electrolysis, generally RCOONa or RCOOK are taken but not RCOOLi. Explain.

Q.11 The IE do not follow a regular trend in II and III periods with increasing atomic number. Why ?

Q.12 Explain why a few elements such as Be (+0.6), N (+0.3) & He(+0.6) have positive electron gain enthalpies while majority of elements have negative values.

Q.13 Which bond in each pair is more polar

(A) P–Cl or P–Br (B) S–Cl or S–O

(C) N–O or N–F

Q.14 LiCl is hydrated but NaCl is always anhydrous. Explain.

Q.15 Explain physical property of third period?

Q.16 Arrange the following according to the instructions given against each.

(i) S, O, Se, C (Increasing order of atomic radius)

(ii) Ca, Al, O, N (Increasing paramagnetism)

(iii) F, Cl⁻, Br⁻, I⁻ (Increasing polarizability)

(iv) Na, Mg and Al (Increasing electropositive character)

(v) Na, P, Si, Al, S (Decreasing order of IE)

(vi) I⁻, Br⁻, Cl⁻ (Decreasing order of ionic size)

(vii) C, F, Li, O, Cs (Increasing order of IE₁)

(viii) O, F, Be, C, N (Decreasing order of electron affinity)

(ix) Mg²⁺, O²⁻, Na⁺, F⁻, N³⁻ (Decreasing order of ionic size)

(x) Na⁺, I⁻, Mg²⁺, Rb⁺, Cl⁻ (Decreasing order of ionic radii)

Q.17 Explain the following:

(i) Fluorine is the most electronegative and Cesium is the most electropositive element.

(ii) The first ionization energy of carbon atom is greater than that of boron atom whereas, the reverse is true for the second ionization energy.

(iii) Electron affinities of halogens are high.

(iv) The formation of F⁻(g) from F(g) is exothermic whereas that of O²⁻(g) from O⁻(g) is endothermic.

(v) Transition and inner transition metals have variable oxidation states.

(vi) Zero group elements are chemically inert.

(vii) The second ionization potential of an element is higher than its first ionization potential.

(viii) The radius of an anion is greater than the parent atom while that of a cation is lesser than the atom.

Q.18 State giving reasons which one will have a higher value

(i) IE₁ of F or Cl

(ii) EA of O or O⁻

(iii) ionic radius of K⁺ or Cl⁻

Q.19 From among the elements, choose the following: Cl, Br, F, Al, C, Li, Cs & Xe.

(i) The element with highest electron affinity.

(ii) The element with lowest ionization potential.

Q.20 In the ionic compound KF, the K⁺ and F⁻ ions are found to have practically radii, about 1.34 Å each. What do you predict about the relative covalent radii of K and F?

Q.21 Explain applications of electronegativity

Q.22 Define atomic size and ionic size?

Q.23 Which oxide is more basic, MgO or BaO? Why?

Q.24 The basic nature of hydroxides of group 13 (III–A) decreases progressively down the group. Comment.

Q.25 Based on location in P.T., which of the following would you expect to be acidic & which basic.

(i) CsOH

(ii) IOH

(iii) Sr(OH)₂

(iv) Se(OH)₂

(v) FrOH

(vi) BrOH

Q.26 Compare the following giving reasons Acidic nature of oxides: CaO, CO, CO₂, N₂O₅, SO₃.

Q.27 Write the drawback of Mendeleev table?

Q.28 Write the postulates of modern periodic table?

Q.29 Explain electron affinity of periodic tables?

Q.30 Explain factors affecting ionization potential?

Q.31 From among the elements, choose the following: Cl, Br, F, Al, C, Li, Cs & Xe.

- (i) The element whose oxide is amphoteric.
- (ii) The element which has smallest radii.
- (iii) The element whose atom has 8 electrons in the outermost shell.

Exercise 2

Single Correct Choice Type

Q.1 When the following five anions are arranged in order of decreasing ionic radius, the correct sequence is:

- (A) Se^{2-} , I^- , Br^- , O^{2-} , F^-
- (B) I^- , Se^{2-} , O^{2-} , Br^- , F^-
- (C) Se^{2-} , I^- , Br^- , F^- , O^{2-}
- (D) I^- , Se^{2-} , Br^- , O^{2-} , F^-

Q.2 Which of the following is wrong regarding the stability of the ions Ge, Sn and Pb

- (A) $\text{Ge}^{2+} < \text{Sn}^{2+} < \text{Pb}^{2+}$
- (B) $\text{Ge}^{4+} > \text{Sn}^{4+} > \text{Pb}^{4+}$
- (C) $\text{Sn}^{4+} > \text{Sn}^{2+}$
- (D) $\text{Pb}^{2+} < \text{Pb}^{4+}$

Q.3 The law of triads is not applicable on

- (A) Cl, Br, I
- (B) Na, K, Rb
- (C) S, Se, Te
- (D) Ca, Sr, Ba

Q.4 The atomic volume was chosen as the basic of periodic classification of elements by

- (A) Niels Bohr
- (B) Mendeleev
- (C) Lothar meyer
- (D) Newlands

Q.5 The majority of gaseous elements in the periodic table are placed

- (A) At bottom left hand side
- (B) At top right hand side
- (C) Below the main table
- (D) Along side d block elements

Q.6 The electronic configuration

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 5s^2$ is for:

- (A) f-block element
- (B) d-block element
- (C) p-block element
- (D) s-block element

Q.7 Fluorine has the highest electronegativity among the $ns^2 np^5$ group on the Pauling scale, but the electron affinity of fluorine is less than that of chlorine because:

- (A) The atomic number of fluorine is less than that of chlorine
- (B) Fluorine being the first member of the family behaves in an unusual manner
- (C) Chlorine can accommodate an electron better than fluorine by utilizing its vacant 3d orbital
- (D) Small size, high electron density and an increased electron repulsion makes addition of an electron to fluorine less favourable than that in the case of chlorine.

Q.8 The last electron in each normal elements of period is filled in

- (A) The same energy sublevel
- (B) The same energy level
- (C) The same orbital
- (D) Relation between I_x and I_z is uncertain

Q.9 The greater stability of the lower oxidation state in heavier p block metals in the consequence of

- (A) Electronic transition within p-orbitals
- (B) Electronic transition from s to p-orbitals
- (C) Inert pair effect
- (D) Expansion of octet

Q.10 Oxidation number of p-block elements is [Excluding inert gases]

- (A) Equal to group number
- (B) Group number +2
- (C) Between the range [Group no...(Group no. 8)]
- (D) Number of unpaired electrons in the valence shell

Q.11 The correct order of second ionization potential of carbon, nitrogen, oxygen and fluorine is:

- (A) $\text{C} > \text{N} > \text{O} > \text{F}$
- (B) $\text{O} > \text{N} > \text{F} > \text{C}$
- (C) $\text{O} > \text{F} > \text{N} > \text{C}$
- (D) $\text{F} > \text{O} > \text{N} > \text{C}$

Q.12 Which statement is wrong:

- (A) 2nd ionization energy shows jump in alkali metals
- (B) 2nd electron affinity for halogens is zero
- (C) Maximum electron affinity exists for F
- (D) Maximum ionization energy exists for He

Q.13 Which of the following is the configuration of second excited state of the element isoelectronic with O_2 or P^- or Cl^+

- (A) $[Ne]3s^2 3p_x^1 3p_y^1 3p_z^1$
 (B) $[Ne]3s^2$
 (C) $[Ne]3s^1 3p_z^1 3p_y^1 3p_x^1 3d_{xy}^1 3d_{yz}^1$
 (D) $[Ne] 3s^1 3p_x^1 3p_z^1 3d_{xy}^1$

Q.14 Metallic radii of transition elements

- (A) First increase, then decrease periodically
 (B) First decrease, then remain almost constant
 (C) First increase, then remaining almost constant
 (D) First increase, then increase periodically

Q.15 For the formation of a covalent bond, the difference in the value of electro negativities should be:

- (A) Equal to or less than 1.7 (B) More than 1.7
 (C) 1.7 or more (D) None of these

Q.16 Which one of the following is the correct order of interactions?

- (A) Covalent < hydrogen bonding < Vander Waal's < dipole-dipole
 (B) Van der Waal's < hydrogen bonding < dipole-dipole < covalent
 (C) Van der Waal's < dipole-dipole < hydrogen bonding < covalent
 (D) Dipole-dipole < Van der Waal's < hydrogen

Q.17 Properties of the elements of which of the following pairs do not resemble?

- (A) Li and Mg (B) Be and Al
 (C) Mg and Al (D) B and Si

Q.18 Which of the following statement is not true?

- (A) The atoms have no tendency to accept electrons in empty higher energy levels
 (B) The atoms have no tendency to accept electrons in empty high energy sublevels
 (C) The alkali metals have no tendency to accept electrons
 (D) The atoms with exactly half-filled electronic configurations have no tendency to accept electrons.

Q.19 IP_1 and IP_2 of Mg are 178 and 348 kcal mol⁻¹. The energy required for the reaction, $Mg \rightarrow Mg^{2+} + 2e^-$

- (A) +170 kcal (B) +526 kcal
 (C) $E - 2E_2$ (D) $(E_1 - E_2)/2$

Q.20 Electronic configuration of an element of atomic weight 40 is 2, 8, 8, 2 which of the following statement regarding this element is not correct

- (A) It belongs to the second group of the P.T
 (B) It has 20 neutrons
 (C) The formula of its oxide is MO_2
 (D) It has 20 Protons

Q.21 There are four elements P, Q, R and S: their configuration are also given. Show that which element will have highest value of IE_2 ?

- (A) $P = [He]2s^2$ (B) $Q = [He] 2s^2 2p^2$
 (C) $R = [He] 2s^2 2p^1$ (D) $S = [He] 2s^1 2p^3$

Q.22 The ionization potential of nitrogen is greater than that of oxygen because

- (A) Nitrogen is an inert element
 (B) The outermost shell of nitrogen has half-filled orbitals
 (C) The radius of nitrogen is more than that of oxygen
 (D) The radius of oxygen is more than that of nitrogen

Q.23 Be and Mg have zero value of electron affinities, because

- (A) Be and Mg have $[He] 2s^2$ and $[Ne] 3s^2$ configuration respectively
 (B) 2s and 3s orbitals are filled to their capacity
 (C) Be and Mg are unable to accept electron
 (D) All the above are correct

Q.24 If I_1 and I_3 etc. represent the successive ionization potentials of an atom then the correct order is.

- (A) $I_1 > I_2 > I_3$ (B) $I_1 < I_2 > I_3$ (C) $I_1 < I_2 < I_3$ (D) $I_2 > I_1 > I_3$

Q.25 For I.P, which order is wrong

- (1) $F > O$ (2) $O > N$ (3) $S > P$ (4) $Be > B$

Code is –

- (A) 1, 2, 3 (B) 2, 3 (C) 1, 4 (D) 1, 2, 4

Q.26 Alkali metals do not form dipositive ions, because

- (A) The difference in the first and second I.P. is more than 16 eV
 (B) The difference in the first and second I.P. is less than 11 eV
 (C) Alkali metals have one electron in their ultimate energy level
 (D) Oxidation state of alkali metals is +1

Q.27 Electronic configuration of X^{+2} and Y^{+3} are: $X^{+2} = [\text{Ar}] 3d^8$, $Y^{+3} = [\text{Ar}] 3d^3$. What are the atomic number of X^0 and Y^0 respectively.

- (A) 28, 24 (B) 28, 25 (C) 28, 26 (D) 28, 27

Q.28 The electronegativity values of C, N, O and F

- (A) Increase from carbon to fluorine
 (B) Decrease from carbon to fluorine
 (C) Increase up to oxygen and minimum at a fluorine
 (D) Is minimum at nitrogen and then increase continuously

Q.29 The valency in the II period from left to right

- (A) Increases
 (B) Decreases
 (C) First increases, then decreases
 (D) First decreases, then increases

Q.30 Which of the following represents incorrect relation of electronegativity

- (A) $C > O > N$ (B) $C < O > N$
 (C) $O > C < N$ (D) $O > N > C$

Q.31 In a period, elements are arranged in strict sequence of

- (A) Decreasing charges in the nucleus
 (B) Increasing charges in the nucleus
 (C) Constant charges in the nucleus
 (D) Equal charges in the nucleus

Q.32 In the periodic table, the metallic character of elements

- (A) Decreases across a period and down a group
 (B) Decreases across a period and increases down a group
 (C) Increases across a period and down a group
 (D) Increases across a period and decreases down a group

Q.33 The screening effect of inert electrons of the nucleus causes

- (A) A decrease in the ionization potential
 (B) An increase in the ionization potential
 (C) No effect on the ionization potential
 (D) An increase in the nuclei attraction of the electrons

Q.34 The statement that is not correct for the periodic classification of elements is

- (A) The properties of elements are the periodic functions of their atomic numbers
 (B) Non-metallic elements are lesser in number than metallic elements
 (C) The first ionization energies along a period vary in a regular manner with increase in the atomic number
 (D) For transition elements, the d-sub-shells are filled with electrons monotonically with an increase in atomic number

Q.35 Match List I with List II and select the correct answer using the codes given below the lists

List I	List II
(i) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	(p) In
(ii) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$	(q) In
(iii) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10}$	(r) Ca
(iv) $1s^2 2p^2 2p^6 3s^2 3d^{10} 4s^2 4s^2 4p^{10} 5s^2 5p^1$	(s) Cu

Codes: i ii iii iv

- (A) p q r s
 (B) p r q s
 (C) r s q p
 (D) p s r q

Q.36 Match List-I with List-II and select the correct answer from the codes given below the lists

List I	List II
(A) ns^2, np^5	(p) Chromium
(B) $(n-1)d^{10}, ns^1$	(q) Copper
(C) $(n-1)d^{10}, ns^2, np^6$	(r) Krypton
(D) $(n-1)d^{10}, ns^2, np^6$	(s) Bromine ($n = 4$)

Q.37 Match List I with List II and select the correct answer from the codes given below the lists

List I	List II
(A) Highest ionization potential	(p) Technitium
(B) Highest electronegativity	(q) Lithium
(C) Artificial element	(r) Helium
(D) High reducing ability	(s) Fluorine

Q.38 The correct order of relative basic character of NaOH, Mg(OH)₂ and Al(OH)₃ is—

- (A) Al(OH)₃ > Mg(OH)₂ > NaOH
 (B) Mg(OH)₂ > NaOH > Al(OH)₃
 (C) NaOH > Mg(OH)₂ > Al(OH)₃
 (D) Al(OH)₃ > NaOH < Mg(OH)₂

Q.39 Match List-I with List-II and select the correct answer from the codes given below the lists

List I	List II
(i) Increasing atomic size	(p) Cl < O < F
(ii) Decreasing atomic radius	(q) Li < Be < B
(iii) Increasing electronegativity	(r) Si < Al < Mg
(iv) Decreasing effective nuclear charge	(s) Na > N > F

Codes: i ii iii iv

- (A) r s p q
 (B) s q r p
 (C) p q r s
 (D) q p s r

Previous Years Questions

Q.1 Which of the following statement is correct with respect to the property of element with an increase in atomic number in the carbon family (group 14) **(2004)**

- (A) Atomic size decrease
 (B) Ionisation energy increase
 (C) Metallic character decrease
 (D) Stability of +2 oxidation state increase

Q.2 The first ionization potentials in electron volts of nitrogen and oxygen atoms are respectively given by **(1987)**

- (A) 14.6, 13.6 (B) 13.6, 14.6
 (C) 13.6, 13.6 (D) 14.6, 14.6

Q.3 The elements which occupy the peaks of ionization energy curve, are **(2000)**

- (A) Na, K, Rb, Cs (B) Na, Mg, Cl, I
 (C) Cl, Br, I, F (D) He, Ne, Ar, Kr

Q.4 Arrange F, Cl, O, N in the decreasing order of electronegativity

- (A) O > F > N > Cl (B) F > N > Cl > O
 (C) Cl > F > N > O (D) F > O > N > Cl

Q.5 Ionic radii of **(1999)**

- (A) Ti⁴⁺ < Mn⁷⁺ (B) ³⁵Cl⁻ < ³⁷Cl⁻
 (C) K⁺ > Cl⁻ (D) P³⁺ > P⁵⁺

Q.6 In which block, 106th element belongs

- (A) s-block (B) p-block
 (C) d-block (D) f-block

Q.7 The first ($\Delta_1 H_1$) and second ($\Delta_1 H_2$) ionization enthalpies (in kJ mol⁻¹) and the electron gain enthalpy ($\Delta_{eg}H$) (in kJ mol⁻¹) of the elements I, II, III, IV and V are given below

Element	$\Delta_1 H_1$	$\Delta_1 H_2$	$\Delta_{eg}H$
I	520	7300	-60
II	419	3051	-48
III	1681	3374	-328
IV	1008	1846	-295
V	2372	5251	+48

The most reactive metal and the least reactive non-metal of these are respectively **(2010)**

- (A) I and V (B) V and II
 (C) II and V (D) V and II
 (E) V and III

Read the assertion and reason carefully to mark the correct option out of the options given below:

- (A) If both assertion and reason are true and the reason is the correct explanation of the assertion.
 (B) If both assertion and reason are true but reason is not the correct explanation of the assertion.
 (C) If assertion is true but reason is false.
 (D) If the assertion and reason both are false.
 (E) If assertion is false but reason is true.

Q.8 Assertion: Positive ions will be wider than parent atoms **(1999)**

Reason: Nuclear charge pulls them closer.

Q.9 Assertion: More is the electron affinity greater is the reducing character. **(2000)**

Reason: Reducing character depends on number of electrons gained.

Q.10 Assertion: Ground state configuration of Cr is $3d^5, 4s^1$. **(2004)**

Reason: A set of half filled orbitals containing one electron each with their spin parallel provides extra stability to the system.

Q.11 Assertion: I.E. Of ${}_7\text{N}$ is more that of ${}_8\text{O}$ as well as ${}_6\text{C}$. **(2005)**

Reason: This is due to difference in reactivity towards oxygen.

Q.12 Which one of the following constitutes a group of the isoelectronic species? **(2008)**

- (A) $\text{C}_2^{2-}, \text{O}_2^-, \text{CO}, \text{NO}$ (B) $\text{NO}^+, \text{C}_2^{2-}, \text{CN}^-, \text{N}_2$
 (C) $\text{CN}^-, \text{N}_2, \text{O}_2^{2-}, \text{C}_2^{2-}$ (D) $\text{N}_2, \text{O}_2^-, \text{NO}^+, \text{CO}$

Q.13 Which one of the following pairs of species have the same bond order? **(2008)**

- (A) CN^- and NO^+ (B) CN^- and CN^+
 (C) O_2^- and CN^- (D) NO^+ and CN^+

Q.14 In which of the following arrangements, the sequence is not strictly according to the property written against it? **(2009)**

- (A) $\text{CO}_2 < \text{SiO}_2 < \text{SnO}_2 < \text{PbO}_2$: increasing oxidising power
 (B) $\text{HF} < \text{HCl} < \text{HBr} < \text{HI}$: increasing acid strength

(C) $\text{NH}_3 < \text{PH}_3 < \text{AsH}_3 < \text{SbH}_3$: increasing basic strength

(D) $\text{B} < \text{C} < \text{O} < \text{N}$: increasing first ionization enthalpy.

Q.15 The set representing the correct order of ionic radius is: **(2009)**

- (A) $\text{Li}^+ > \text{Be}^{2+} > \text{Na}^+ > \text{Mg}^{2+}$
 (B) $\text{Na}^+ > \text{Li}^+ > \text{Mg}^{2+} > \text{Be}^{2+}$
 (C) $\text{Li}^+ > \text{Na}^+ > \text{Mg}^{2+} > \text{Be}^{2+}$
 (D) $\text{Mg}^{2+} > \text{Be}^{2+} > \text{Li}^+ > \text{Na}^+$

Q.16 The bond dissociation energy of B–F in BF_3 is 646 kJ mol^{-1} whereas that of C–F in CF_4 is 515 kJ mol^{-1} . The correct reason for higher B–F bond dissociation energy as compared to that of C–F is: **(2009)**

- (A) Smaller size of B-atom as compared to that of C-atom
 (B) Stronger σ bond between B and F in BF_3 as compared to that between C and F in CF_4
 (C) Significant $p\pi-p\pi$ interaction between B and F in BF_3 whereas there is no possibility of such interaction between C and F in CF_4 .
 (D) Lower degree of $p\pi-p\pi$ interaction between B and F in BF_3 than that between C and F in CF_4 .

Q.17 The correct sequence which shows decreasing order of the ionic radii of the elements is **(2010)**

- (A) $\text{Al}^{3+} > \text{Mg}^{2+} > \text{Na}^+ > \text{F}^- > \text{O}^{2-}$
 (B) $\text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+} > \text{O}^{2-} > \text{F}^-$
 (C) $\text{Na}^+ > \text{F}^- > \text{Mg}^{2+} > \text{O}^{2-} > \text{Al}^{3+}$
 (D) $\text{O}^{2-} > \text{F}^- > \text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+}$

Q.18 Among the following the maximum covalent character is shown by the compound: **(2011)**

- (A) SnCl_2 (B) AlCl_3
 (C) MgCl_2 (D) FeCl_2

Q.19 Which one of the following order represents the correct sequence of the increasing basic nature of the given oxides? **(2011)**

- (A) $\text{MgO} < \text{K}_2\text{O} < \text{Al}_2\text{O}_3 < \text{Na}_2\text{O}$
 (B) $\text{Na}_2\text{O} < \text{K}_2\text{O} < \text{MgO} < \text{Al}_2\text{O}_3$
 (C) $\text{K}_2\text{O} < \text{Na}_2\text{O} < \text{Al}_2\text{O}_3 < \text{MgO}$
 (D) $\text{Al}_2\text{O}_3 < \text{MgO} < \text{Na}_2\text{O} < \text{K}_2\text{O}$

Q.20 The molecule having smallest bond angle is: **(2012)**
 (A) NCl_3 (B) AsCl_3 (C) SbCl_3 (D) PCl_3

Q.21 Which of the following on thermal decomposition yields a basic as well as an acidic oxide? **(2012)**
 (A) NaNO_3 (B) KClO_3 (C) CaCO_3 (D) NH_4NO_3

Q.22 Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar? **(2013)**

- (A) $\text{Ca} < \text{S} < \text{Ba} < \text{Se} < \text{Ar}$
 (B) $\text{S} < \text{Se} < \text{Ca} < \text{Ba} < \text{Ar}$
 (C) $\text{Ba} < \text{Ca} < \text{Se} < \text{S} < \text{Ar}$
 (D) $\text{Ca} < \text{Ba} < \text{S} < \text{Se} < \text{Ar}$

Q.23 The first ionisation potential of Na is 5.1 eV. The value of electron gain enthalpy of Na^+ will be: **(2013)**

- (A) -2.55 eV (B) -5.1 eV
 (C) -10.2 eV (D) +2.55 eV

Q.24 The ionic radii (in Å) N^{3-} , O^{2-} and F^- are respectively: **(2015)**

- (A) 1.36, 1.40 and 1.71
 (B) 1.36, 1.71 and 1.40
 (C) 1.71, 1.40 and 1.36
 (D) 1.71, 1.36 and 1.40

JEE Advanced/Boards

Exercise 1

Q.1 Use the following system of naming elements in which first alphabets of the digits are written collectively,
 0 1 2 3 4 5 6 7 8 9
 nil uni bi tri quad pent hex sept oct enn
 to write the three-letter symbols for the elements with atomic number 101 to 109.

Q.2 Arrange the following in increasing order of the property as indicated:

- (i) Pb, Pb^{2+} and Pb^{4+} (size)
 (ii) Mg, Al, Si and Na (ionisation potential)
 (iii) MgO , SrO , Rb_2O , NiO , Cs_2O (basic character)
 (iv) $\text{Be}(\text{OH})_2$, $\text{Mg}(\text{OH})_2$, $\text{Ca}(\text{OH})_2$, $\text{Ba}(\text{OH})_2$ (basicity)
 (v) Cl^- , K^+ , Ca^{2+} , Ar (ionization energy)
 (vi) As, F, S, Cl (Electronegativity)
 (vii) LiCl, LiBr, LiI (Ionic character)
 (viii) Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ (mobility of hydrated ions)
 (ix) Li, Na, K, Rb and Cs (hydrated radii)
 (x) Be^{2+} , Mg^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+} (hydration of ions)

Q.3 Why are inert gases mono-atomic?

Q.4 Why does the third period contains 8 elements and not 18?

Q.5 Arrange the following in order of increasing ionic radius:

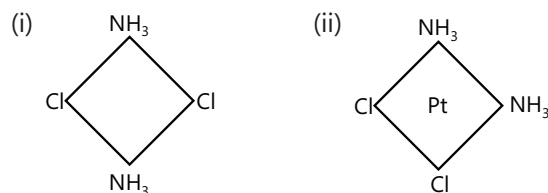
- (i) Cl^- , P^{3-} , S^{2-} , F^- (ii) Al^{3+} , Mg^{2+} , Na^+ , O^{2-} , F^-
 (iii) Na^+ , Mg^{2+} , K^+ .

Q.6 If internuclear distance between Cl atoms in Cl_2 is 10 Å & between H atoms in H_2 is 2 Å, then calculate internuclear distance between H and Cl (Electronegativity of H = 2.1 & Cl = 3.0).

Q.7 The As-Cl bond distance in AsCl_3 is 2.20 Å. Estimate the SBCR (single bond covalent radius) of As.

(Assume EN of both to be same and radius of Cl = 0.99 Å.)

Q.8 The Pt-Cl distance has been found to be 2.32 Å in several crystalline compounds. If this value applies to both of the compounds shown in figure, what is Cl-Cl distance in (i) and (ii)



Q.9 K^+ and F^- have identical radius about 1.34 Å. What should be the atomic radius of K and F atoms.

Q.10 The radii of noble gases are greater than that of the radii of precedent halogens.

Q.11 The covalent radii (Å) do not increases regularly from B to Tl (B = 0.80 Å, Al = 1.25 Å; Ga = 1.25 Å, In = 1.50 Å; Tl = 1.55 Å) How do you account for this anomalous behavior.

Q.12 Calculate E.N. of chlorine atom on Pauling scale if I.E. of Cl^- is 4eV & of E.A. of Cl^+ is +13.0 eV.

Q.13 Why in isoelectronic species, the ionic / atomic radii decrease with increase in atomic numbers?

Q.14 The ionisation energy of the coinage metals fall in the order $Cu > Ag < Au$.

Q.15 The first ionization energy of carbon atom is greater than that of boron where as the reverse is true for second ionization energy. Explain.

Q.16 Define ionization potential?

Q.17 The 1st ionization energy of potassium is less than that of Cu but the reverse is true for 11th ionisation energy.

Q.18 The sums of first and second ionization energies and those of third fourth ionisation energies in (kJ mol⁻¹) of nickel and platinum are

$$(IE)_1 + (IE)_2 \qquad (IE)_3 + (IE)_4$$

$$Ni \ 2.49 \qquad 8.80$$

$$Pt \ 2.66 \qquad 6.70$$

Based on this information; write

(i) The most stable oxidation states of Ni and Pt

(ii) Name of one of the metals which can more easily form compounds in its +4 oxidation state.

Q.19 The first ionization enthalpy of magnesium is higher than of sodium. On the other hand, the second ionization enthalpy of sodium is very much higher than that of magnesium. Explain.

Q.20 Na^+ has higher value of ionization energy than Ne, though both have same electronic configuration. Explain?

Exercise 2

Single Correct Choice Type

Q.1 Two p-block elements x (outer configuration ns^2, np^3) and z (outer configuration $ns^2 np^4$) occupy neighbouring positions in a period. Using this information, which of the following is correct with respect to their ionization potential I_x and I_z .

- (A) $I_x > I_z$ (B) I_z and I_x (C) $I_z = I_x$ (D) $I_z > I_x$

Q.2 True position of lanthanides are:

- (A) After III B group and in the 6th period
(B) After III B group in the 3rd period
(C) After VI B group in the 3rd period
(D) After III B group in the 4th Period

Q.3 The main cause of diagonal relationship between Be and Al is

- (A) Similarity in ionic sizes
(B) Similar ionic potentials
(C) Similar electronegativity
(D) Similar atomic volume

Q.4 Electronic configuration of four elements are: a = $1s^2; 2s^2, 2p^1$; b = $1s^2; 2s^2, 2p^2$; c = $1s^2; 2s^2, 2p^5$; d = $1s^2; 2s^2, 2p^6; 3s^1$. Which one of these would most readily form diatomic molecule?

- (A) a (B) b (C) c (D) d

Q.5 EN of the element A is E_1 and IP is E_2 . Hence EA will be:

- (A) $2E_1 - E_2$ (B) $E_1 - E_2$ (C) $E_1 - 2E_2$ (D) $E_1 + E_2$

Q.6 $\boxed{1\downarrow} \quad \boxed{1\downarrow} \quad \boxed{1} \quad \boxed{1}$

The above configuration would be of the species

- (i) S (ii) Cl^+ (iii) p (iv) Ar^{-2}

Code is –

- (A) i, ii (B) iii, iv (C) i, iv (D) ii, iii

Q.7 Which of the following statements is correct for the addition of an electron to an isolated and gaseous uni-negatively charged oxygen O^- ion?

- (A) The addition of electron cannot occur
- (B) The addition of electron occurs with evolution of energy
- (C) The addition of electron occurs with absorption of energy
- (D) None of the above

Q.8 In the transformation $Na(s) \rightarrow Na^+(g)$, the energies involved are

- (A) Ionization energy
- (B) Sublimation energy
- (C) Ionization energy and sublimation energy both
- (D) None of the above

Q.9 The process of requiring absorption of energy is

- (A) $F \rightarrow F^-$
- (B) $Cl \rightarrow Cl^-$
- (C) $O^- \rightarrow O^{2-}$
- (D) $H \rightarrow H^-$

Q.10 The cyanide CN^- & N_2 are isoelectronic. But in contrast to CN^- , N_2 is chemically inert because of –

- (A) Low bond energy
- (B) Absence of bond polarity
- (C) Unsymmetrical electron distribution
- (D) Presence of more electron in bonding

Previous Year Questions

Q.1 The correct order of radii is **(2000)**

- (A) $N < Be < B$
- (B) $F^- < O^{2-} < N^{3-}$
- (C) $Na < Li < K$
- (D) $Fe^{3+} < Fe^{2+} < Fe^{4+}$

Q.2 The set representing the correct order of first ionization potential is **(2001)**

- (A) $K > Na > Li$
- (B) $Be > Mg > Ca$
- (C) $B > C > N$
- (D) $Ge > Si > C$

Q.3 Identify the least stable ion amongst the following **(2000)**

- (A) Li^+
- (B) Be^-
- (C) B^-
- (D) C^-

Q.4 The statements that is/are true for the long form of the periodic table is/are **(1988)**

- (A) It reflects the sequence of filling the electrons in the order of sub-energy level s, p, d and f
- (B) It helps to predict the stable valency states of the elements
- (C) It reflects trends in physical and chemical properties of the elements
- (D) It helps to predict the relative ionicity of the bond between any two elements

Q.5 Sodium sulphate is soluble in water whereas barium sulphate is sparingly soluble because **(1989)**

- (A) The hydration energy of sodium sulphate is more than its lattice energy
- (B) The lattice energy of barium sulphate is more than its hydration energy
- (C) The lattice energy has no role to play in solubility
- (D) The hydration energy of sodium sulphate is less than its lattice energy

Q.6 The softness of group IA metals increases down the group with increasing atomic number. **(1986)**

Q.7 In group IA of alkali metals, the ionization potential decreases down the group. Therefore, lithium is a poor reducing agent. **(1987)**

Q.8 Arrange the following in order of their **(1985)**

- (i) Decreasing ionic size Mg^{2+} , O^{2-} , Na^+ , F^-
- (ii) Increasing first ionization energy Mg, Al, Si, Na
- (iii) Increasing bond length F_2 , N_2 , Cl_2 , O_2

Q.9 Arrange the following in the order of their increasing size: **(1986)**

Cl^- , S^{2-} , Ca^{2+} , Ar

Q.10 Compare qualitatively the first and second ionization potentials of copper and zinc. Explain the observation. **(1996)**

Q.11 Arrange the following ions in order of their increasing radii **(1997)**

Li^+ , Mg^{2+} , K^+ , Al^{3+} .

Q.12 The species having bond order different from that in CO is (2007)

- (A) NO^- (B) NO^+ (C) CN^- (D) N_2

Q.13 Statement-I: Pb^{4+} compounds are stronger oxidizing agents than Sn^{4+} compounds. (2008)

Statement-II: The higher oxidation states for the group 14 elements are more stable for the heavier members of the group due to 'inert pair effect'.

(A) Statement-I is True, Statement-II is True; Statement-II is correct explanation for Statement-I

(B) Statement-I is True, Statement-II is True; Statement-II is NOT a correct explanation for Statement-I

(C) Statement-I is True, Statement-II is False

(D) Statement-I is False, Statement-II is True

Q.14 Among the following, the number of elements showing only one non-zero oxidation state is: (2010)

O, Cl, F, N, P, Sn, Tl, Na, Ti

Q.15 The increasing order of atomic radii of the following Group 13 elements is (2016)

(A) $\text{Al} < \text{Ga} < \text{In} < \text{Ti}$ (B) $\text{Ga} < \text{Al} < \text{Tl} < \text{In}$

(C) $\text{Al} < \text{In} < \text{Ga} < \text{Ti}$ (D) $\text{Al} < \text{Ga} < \text{Tl} < \text{In}$

PlancEssential Questions

JEE Main/Boards

Exercise 1

Q.1 Q.9 Q.10 Q.14
Q.16 Q.17 (D) Q. 21

Exercise 2

Q.13 Q.16 Q.23 Q.26 Q.33
Q.35 Q.37

Previous Years' Questions

Q.7 Q.10

JEE Advanced/Boards

Exercise 1

Q.9 Q.11 Q.14

Exercise 2

Q.5 Q.10

Previous Years' Questions

Q.5 Q.6

Answer Key

JEE Main/Boards

Exercise 2

Single Correct Choice Type

Q.1 D	Q.2 D	Q.3 B	Q.4 C	Q.5 B	Q.6 D	Q.7 D
Q.8 B	Q.9 C	Q.10 C	Q.11 C	Q.12 C	Q.13 C	Q.14 B
Q.15 A	Q.16 C	Q.17 C	Q.18 D	Q.19 B	Q.20 C	Q.21 C
Q.22 B	Q.23 D	Q.24 C	Q.25 B	Q.26 A	Q.27 A	Q.28 A
Q.29 C	Q.30 A	Q.31 B	Q.32 B	Q.33 A	Q.34 C	Q.35 C
Q.36 C	Q.37 A	Q.38 C	Q.39 A			

Previous Years' Questions

Q.1 D	Q.2 A	Q.3 D	Q.4 D	Q.5 D	Q.6 C	Q.7 C
Q.8 D	Q.9 E	Q.10 A	Q.11 C	Q.12 B	Q.13 A	Q.14 C
Q.15 B	Q.16 C	Q.17 D	Q.18 B	Q.19 D	Q.20 C	Q.21 C
Q.22 C	Q.23 B	Q.24 C				

JEE Advanced/Boards

Exercise 2

Single Correct Choice Type

Q.1 A	Q.2 A	Q.3 B	Q.4 C	Q.5 A	Q.6 A	Q.7 C
Q.8 C	Q.9 C	Q.10 B				

Previous Years' Questions

Q.1 B	Q.2 B	Q.3 B	Q.4 B, C, D	Q.5 A, B	Q.6 True	Q.7 False
Q.12 A	Q.13 C	Q.14 2	Q.15 B			

Solutions

JEE Main/Boards

Exercise 1

Sol 1: As we know, the electronic configuration of Be and Mg are $[\text{He}] 2s^2$ and $[\text{Ne}] 3s^2$ respectively. As we see, both of these elements have a fully filled s-orbital & making their configuration very stable compared to other configuration with incompletely filled orbitals. Therefore, they have very less tendency to gain an electron, which makes their electron affinity equal to zero.

Sol 2: The electronic configuration of oxygen is $[\text{He}] 2s^2 2p^4$, which is not very stable. Since this is neither a half filled or a full filled configuration, therefore it has a natural tendency to accept an electron which makes its first electron affinity positive and the process exothermic but, after an electron is added, the atom becomes negatively charged and the addition of a second electron is opposed by coulombic repulsion. Hence energy has to be supplied for the addition of second electron. Thus, its second electron affinity is negative and the process is endothermic.

Sol 3: We know that the ionisation energy increases while going from left to right in a period. But, there is an exception while going from group 12 to group 13, which means that ionisation energy decreases from group 12 to group 13. This exception is explained by the high stability of group 12 elements due to fully filled orbitals which makes it difficult to remove the outermost electron compared to group 13 where there are no full filled orbitals. Hence, their ionization energy is smaller.

Sol 4: All these ions have the same electronic configuration, i.e. $[\text{Ne}]$. So, difference in size will be due to electronic charge only. Now, we know that the ionic size decreases with increasing positive charge, because a loss of electron results in increased effective nuclear charge and hence the remaining electrons are held more firmly bound with the nucleus, hence decreasing ionic size. Also, ionic size increases with increasing ionic charge, because addition of an electron results in an increase in natural repulsion between electrons which decreases effective nuclear charge and leads to increase in ionic size. Therefore, the order of ionic size will be $\text{N}^{3-} > \text{O}^{2-} > \text{F}^- > \text{Na}^+ > \text{Mg}^{2+}$

Sol 5: Both Ca^{2+} and K^+ have the same electronic configuration, i.e. $[\text{Ar}]$ and hence equal number of shells. The ionic radius decreases with increase in positive charge, because the loss of an electron results in an increase in effective nuclear charge and hence a decrease in radius as the remaining electrons are more firmly bound to the nucleus. If 2 electrons are removed, the effective nuclear charge will increase even more and the size of the ion will decrease further. Therefore, the ionic radius of Ca^{2+} is less than K^+

Sol 6: All the elements except Cs are period 3 elements and we know that atomic size decreases in going from left to right in a period due to increased effective nuclear charge, resulting from the increase in number of protons for same number of electronic shells. Also, the atomic radius of Cs will be bigger than the rest because of more number of shells is Cs. Hence order

$\text{Cs} > \text{Na} > \text{Mg} > \text{Si} > \text{Cl}$

Sol 7: Electronic configuration of carbon: $[\text{He}] 2s^2 2p^2$

Electronic configuration of Boron: $[\text{He}] 2s^2 2p^1$

First ionization energy of carbon is more than Boron because of increased effective nuclear charge in carbon, making removal of electron difficult. After removal of electron:

Electronic configuration of C^+ : $[\text{He}] 2s^2 2p^1$

Electronic configuration of B^+ : $[\text{He}] 2s^2$

Now B^+ has a stable full filled orbital making its configuration very stable compared to incompletely filled configuration of C^+ . Hence second ionization energy of Boron is more than Carbon

Sol 8: Electronic configuration of Be: $[\text{He}] 2s^2$

Electronic configuration of B: $[\text{He}] 2s^2 2p^1$ (IE)₁ of Be > (IE)₁ of B due to stable full filled orbital configuration of Be making removal of electron difficult.

After removal of electron

Electronic configuration of Be^+ : $[\text{He}] 2s^1$

Electronic configuration of B^+ : $[\text{He}] 2s^2$

Here, (IE)₂ of B > (IE)₂ of Be as now B^+ has a more stable configuration than Be^+ due to full filled orbitals which are not present in Be^+

Sol 9: Electronic configuration of Cu: $[\text{Ar}] 3d^2 4s^2$

Electronic configuration of Zn: $[\text{Ar}] 3d^{10} 4s^2$

Zn has a much more stable configuration than Cu because of fully filled d orbital compared to Cu. A stable configuration means that Zn has no tendency to accept an electron which makes its EA negative. Cu on the other hand can attain a stable configuration by adding an electron, hence its EA is positive.

Sol 10: For Kolbe's electrolysis, we want the acid to dissociate into ions easily during the reaction. But since the acidic strength of carbonic acids increases while going from top to bottom in a group, therefore, RCOOLi will be dissociated into ions to a very little extent compared to RCOONa or RCOOK , Hence RCOOLi is not used.

Sol 11: The general trend of IE is that it increases while going from left to right in a period. However IE is more for an element which has a stable configuration due to half-filled or full filled orbitals. Hence, IE of group 7 > IE of group 8 and IE of group 4 > IE of group 5 because group 4 and group 7 have full filled and half-filled orbital respectively.

Sol 12: Electronic configuration of Be: $[\text{He}] 2s^2$

Electronic configuration of N: $[\text{He}] 2s^2 2p^3$

Electronic configuration of He: $[\text{He}]$

We see, that each of those elements have a full filled or half-filled orbital which makes their configuration very stable and because of this, they have no tendency to accept an electron and hence they have negative E_A and hence positive value of electron gain enthalpy.

Sol 13: Polarity of the bond depends upon the electronegativity difference between the two elements. The more electronegativity difference will be the more polar the bond will be

(A) $\text{P} - \text{Cl} > \text{P} - \text{Br}$ as $(\text{EN})_{\text{Cl}} > (\text{EN})_{\text{Br}}$

$= [(\text{EN})_{\text{Cl}} - (\text{EN})_{\text{P}}] > [(\text{EN})_{\text{Br}} - (\text{EN})_{\text{P}}]$

(B) $\text{S} - \text{O} > \text{S} - \text{Cl}$ as $(\text{EN})_{\text{O}} > (\text{EN})_{\text{Cl}}$

$= [(\text{EN})_{\text{O}} - (\text{EN})_{\text{S}}] > [(\text{EN})_{\text{Cl}} - (\text{EN})_{\text{S}}]$

(C) $\text{N} - \text{F} > \text{N} - \text{O}$ as $(\text{EN})_{\text{F}} > (\text{EN})_{\text{O}}$

$= [(\text{EN})_{\text{F}} - (\text{EN})_{\text{N}}] > [(\text{EN})_{\text{O}} - (\text{EN})_{\text{N}}]$

Sol 14: LiCl is hydrated because hydration energy of Li^+ is much higher than Na^+ due to small size of Li^+ ion. Hence, NaCl is amorphous because of low hydration enthalpy of Na^+ .

Sol 15: Atomic Radius: The atomic radius decreases while going from left to right till because of increased effective nuclear charge which pulls the electrons closer to the nucleus and hence decreases atomic radius. Atomic radius of Ar is larger because Vander Waals radius is calculated for Ar as it does not form any strong bonds.

First Ionisation Energy: (IE), generally increases across the period from left to right due to increase in effective nuclear charge and poor screening of outer most electrons. However (IE) of Mg > (IE) of Al or Mg has a stable configuration because of full filled orbital and (IE) of P > (IE) of S as P has half-filled orbital making its configuration stable.

Electron Affinity (EA): Electron affinity increases across the period due to increased nuclear charge and decreasing atomic size. However E_A of S > E_A of P, due to half-filled orbital stable configuration of P which makes addition of electron difficult.

Electronegativity: It increases across the period due to the decrease in atomic size and hence increased effective nuclear charge.

Sol 16: (i) $\text{O} < \text{C} < \text{S} < \text{Se}$

Since, atomic radius increases while going down in a group $\text{O} < \text{S} < \text{Se}$. Also since atomic radius decreases across a period $\text{O} < \text{C}$.

(ii) Paramagnetism will increase with an increase in number of unpaired electrons.

$\therefore \text{Ca}$ (0 unpaired electron) Al (1 unpaired) $< \text{O}$ (2 unpaired) $< \text{N}$ (3 unpaired)

$\text{Ca} < \text{O} < \text{N}$

(iii) The order of polarisability will depend on the order of ionic size. The larger the ionic size, more will be polarizability.

\therefore Order - $\text{F}^- < \text{Cl}^- < \text{Br}^- < \text{I}^-$

(iv) $\text{Al} < \text{Mg} < \text{Na}$

Since electropositive character decreases across a period.

(v) $\text{P} > \text{S} > \text{Si} > \text{Al} > \text{Na}$

I.E. increases across the period

I.E. of P > I.E. of S due to stable half-filled orbital configuration of P

(vi) $\text{I}^- > \text{Br}^- > \text{Cl}^-$

Ionic radius increases down the period.

(vii) $\text{Cs} < \text{Li} < \text{C} < \text{O} < \text{F}$

IE, increases across the period.

(viii) $F > O > C > N > Be$

Electron affinity increases across a period $(EA)_C > (EA)_N$ due to stable configuration of N (half-filled orbital)

(ix) $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+}$

Ionic size decreases with increasing positive charge.

Positive charge means increase in effective nuclear charge and hence decrease in size negative charge lead to increase in repulsion between electrons thus increases atomic size

(x) $I^- > Rb^+ > Cl^- > Na^+ > Mg^{2+}$

Ionic radii decreases with increase in positive charge and with decrease in number of shells.

Sol 17: (i) We know that electronegativity increases across a period from left to right and decreases while going down in a group. Therefore F is the most electronegative and Cs is the least electronegative element.

(ii) Electronic configuration of C: $[He] 2s^2 2p^2$

Electronic configuration of B: $[He] 2s^2 2p^1$

$(IE)_1$ of C $>$ $(IE)_1$ of B because of increased effective nuclear charge in C compared to B

Electronic configuration of C^+ : $[He] 2s^2 2p^1$

Electronic configuration of B^+ : $[He] 2s^2$

$(IE)_2$ of B $>$ $(IE)_2$ of C as B^+ has a stable electronic configuration due to fully filled s orbital which makes removal of electron difficult.

(iii) The outermost electronic configuration of halogens is of the form $ns^2 np^5$. Thus, by adding one electron, they can attain noble gas configuration, which makes addition of electrons very favorable and hence results in high electron affinity.

(iv) The formation of $F^-(g)$ from $F(g)$ enables fluorine to have a stable noble gas configuration. Hence, this process is exothermic.

For formation of $O^{2-}(g)$ from $O^-(g)$, an extra electron needs to be added to an already negative $O^-(g)$ which will be difficult due to coulombic repulsion between electrons making this process endothermic

(v) Transition and inner transition metals have variable oxidation states because ns and $(n-1)d$ orbitals have very similar energy and there is not a huge jump in the amount of energy required to remove the third electron compared with the first and second (which are removed from ns orbitals)

(vi) Zero group elements (noble gases) are chemically inert as these have a very stable electronic configuration

of $ns^2 np^6$ and have no tendency to either accept or donate an electron.

(vii) $(IE)_2$ of an element is always greater than $(IE)_1$ as increased amount of energy is needed to remove an electron from an already positively charged ion due to increased nuclear charge.

(viii) Radius of cation is lesser than atom as removal of an electron leads to increase in effective nuclear charge and hence decrease in size.

Radius of anion is bigger than atom because addition of an electron results in increased coulombic repulsion between electrons causing decrease in effective nuclear charge and hence increase in size.

Sol : 18 (a) IE_1 of F will be higher than that of Cl.

Reason: As we go down in a group, atomic radius increases and consequently the effective nuclear charge decreases and thus ionization decreases.

(b) EA of O^- is greater than O.

Reason: The second electron affinity of oxygen is particularly high because the electron is being forced into a small, very electron-dense space.

(c) Ionic radius of Cl^- will be greater than K^+ .

Reason: Both are isoelectronic species thus ionic radius decreases with increase in nuclear charge.

Sol 19: (i) F because EA increases across the period and decreases down the group

(ii) Cs because IE decreases down the group and increases across the period.

Sol 20 Covalent radius of K $>$ covalent radius of F

Since, size of $K^+ <$ size of K and size of $F^- >$ size of F

Sol 21: (i) Nomenclature:

Name of more electronegative element is written at the end and 'ide' is suffixed to it

Name of less electronegative element is written first in the formula. Eg. I, Br Iodine Bromine

(ii) If difference is $EN > 1.7$ ionic bond

If difference is $EN < 1.7$ covalent bond

(iii) Metallic and non-metallic nature

Low EN: metallic

high EN: nonmetallic

(iv) Percentage ionic character in covalent bonds =

$$16(X_A - X_B)^2 + 3.5(X_A - X_B)$$

(v) Bond length

When electronegativity difference increases, bond length decreases

$$d_{A-B} = r_A + r_B - 0.09 |X_A - X_B|$$

(vi) Bond strength and stability

It increases with increase in electronegativity difference

Eg. $H-F > H-Cl > H-Br > H-I$

(vii) Acidic strength of hydrides

It decreases with increase in electronegativity difference

$H-I > H-Br > H-Cl > H-F$

(ix) Acidic nature of oxides

If $X_0 - X_A \geq 2.3$ basic oxides

$X_0 - X_A < 2.3$ acidic oxides

(x) Nature of hydroxide

In AOH, if $(EN)_A > 1.7$, acidic

$(EN)_A < 1.7$, basic

Sol 22: Atomic size: It is the distance between the centre of the nucleus and the outermost shell containing electrons.

Ionic size: It is the distance between the centre of the nucleus and the outermost shell containing electrons in an ion.

Sol 23: BaO is more basic because $(EN)_{Ba} < (EN)_{Mg}$

Sol 24: This statement is wrong since the electronegativity decreases while going down in a group. Therefore, basic strength should increase.

Sol 25: (i) CsOH: basic ($(EN)_{Cs} < 1.7$)

(ii) IOH: acidic ($(EN)_I > 1.7$)

(iii) $Sr(OH)_2$: basic ($(EN)_{Sr} < 1.7$)

(iv) $Se(OH)_2$: acidic ($(EN)_{Se} > 1.7$)

(v) FrOH: basic ($(EN)_{Fr} < 1.7$)

(vi) BrOH: acidic ($(EN)_{Br} > 1.7$)

Sol 26: Acidic nature: $N_2O_5 > SO_3 > CO_2 < CO > CaO$

Acidic strength decreases with increase in electronegativity difference. Also, CO is neutral and CaO is basic.

Sol 27: (i) Anomalous position of hydrogen.

(ii) Isotopes were placed in same group despite having different properties.

(iii) Isobars are placed in different groups.

(iv) Metals and non-metals both placed together in same group.

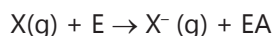
(v) Some similar elements placed differently.

(vi) Position of lanthanides and actinides is not properly specified.

Sol 28: (i) The nuclear charge on an atom is equal to the atomic number.

(ii) The properties of elements are the periodic functions of their atomic number.

Sol 29: Electron affinity (EA) is the amount of energy released when a neutral gaseous atom accepts an electron to form gaseous anion



EA increase across the period in periodic table due to increased effective nuclear charge with the exception of elements having stable configurations (due to half or full filled orbitals) for which EA is high than adjacent elements. EA decreases down the group due to decrease in effective nuclear charge.

Sol 30: (i) **Number of shells:** IE decreases with increase in number of shells.

(ii) **Effective Nuclear Charge:** IE increases with increase in effective nuclear charge

(iii) **Shielding effect:** The more will be the shielding effect. Less will be the IE.

(iv) The elements which have half-filled or fully filled orbitals have higher IE due to their stable electronic configuration.

(v) **Penetrative Power:** In any atom, the s orbital is nearer to the nucleus than p, d or f orbitals. Therefore, greater energy is required to remove an electron from s orbital. Decreasing order of IE is:

$s > p > d > f$

Sol 31: (i) Al, due to its intermediate electronegativity.

(ii) F, due to highest effective nuclear charge and least number of shells.

(iii) Xe, It is a group zero element (a noble gas).

Exercise 2

Single Correct Choice Type

Sol 1: (D) I^- will have the largest ionic radius increase further, we know that ionic radius will down the group and decrease across a period. And for isoelectronic species, the ionic radius will increase with increase in negative charge

Sol 2: (D) In group 13, the stability of +2 oxidation state will increase and stability of +4 oxidation state will decrease down the group due to inert pair effect, which is negligible in Ge and most significant in Pb.

Sol 3: (B) We check the atomic radius of all the elements in the four options.

For option (B), we find that the atomic mass of the middle elements is not equal to the average of atomic mass of other 2 elements.

Sol 4: (C) Lothar Meyer used atomic volume for plotting Lothar Mayer's volume curves.

Sol 5: (B) Majority of low atomic number non metallic elements are gases which are placed at top right hand side.

Sol 6: (D) Since, here s orbital is filled last, it is a s-block element

Sol 7: (D) All the reasons given are correct.

Sol 8: (B) The last electron in each period is filled in np orbital which is the same shell, i.e. same energy level.

Sol 9: (C) This is a consequence of inert pair effect, according to which in heavier p-block elements, the s orbitals become very close to nucleus and thus become relatively inert and are not removed.

Sol 10: (C) p-block elements can display a range of oxidation states due to availability of electrons in both of s and p orbitals and also possibility of attaining noble gas configuration by adding electrons.

Sol 11: (C) We note that order of (IE) is $\text{F} > \text{N} > \text{O} > \text{C}$
Here (IE) of $\text{N} >$ (IE) of O because N has a stable half-filled orbital configuration and also IE increases across a period.

But after removal of an electron O will have a stable, half-filled orbital configuration and therefore, its IE will be more than F^+ . Hence order $\text{O} > \text{F} > \text{N} > \text{C}$

Sol 12: (C) Electron affinity is higher for Cl due to its low electron density.

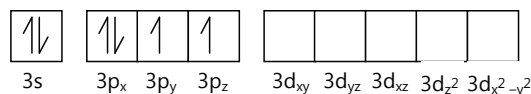
(IE)₂ of alkali metals shows a jump because they have stable noble gas configuration in +1 oxidation state

(IE)₂ of halogens is zero because halogens have noble gas configuration in -1 oxidation state

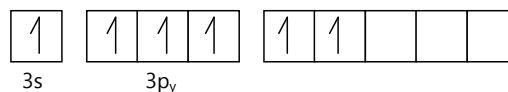
He has maximum ionisation energy due to its small size and fully filled orbital configuration.

Sol 13: (C) The element isoelectronic with O_2 is sulphur (s).
Its configuration in group state $[\text{Ne}] 3s^2 3p^4 3d^0$

Outer shell configuration in ground state



Outer shell configuration in second excited state (all electrons get unpaired)



Sol 14: (B) Metallic radii of transition elements first decreases due to increase in effective nuclear charge, but after since the electrons are filled in $(n-1)d$ orbital, after a few elements, the nuclear charge on ns electrons remains almost constant due to extra stability.

Sol 15: (A) If difference in electronegativity > 1.7 , ionic bond

Sol 16: (C) In covalent bond, there is actual bond formation therefore they are the strongest of allowed by hydrogen bonding. Vander waal's interactions are least strong because they are developed due to momentary dipole interactions.

Sol 17: (C) The properties of other three pairs resemble due to diagonal relationship between them (identical charge 1 size ratio)

Sol 18: (D) The added electron will have no tendency to go to the higher energy level or sublevel as this will further increase the energy of the system. Alkali metals have no tendency to accept electrons due to low effective nuclear charge.

Atoms with exactly half filled configuration will have low (but not zero) tendency to accept electrons.

Sol 19: (B) Energy required = $IP_1 + IP_2$
 $= (178 + 3481 \text{ kcal mol}^{-1}) = 526 \text{ kcal mol}^{-1}$

Sol 20: (C) Configuration = $1s^2 2s^2 2p^2 3s^2 3p^6 4s^2$

Since last filled shell is 4 and it contains 2 electrons, 4th period and 2nd group

Number of neutrons = Atomic weight – Atomic Number
 $= 40 - 20 = 20$

For this element this the most stable oxidation state will be +2 as in +2 oxidation state. It will attain a therefore formula of metal oxide = Mo

Sol 21: (C) (Given answer = a = wrong)

For R, after removal of an electron, a fully filled stable configuration is achieved which increases its $(IE)_2$ above all other 3 elements.

Sol 22: (B) Nitrogen has a configuration of $[\text{He}]2s^2 2p^3$, which is stable because of the half-filled p- orbital and thus, it has more IE than Oxygen.

Sol 23: (D) All the given reasons are correct.

Sol 24: (C) With the removal of an electron, the further removal of electron becomes increasingly difficult due to increase in effective nuclear charge and decrease in size. Hence $I_1 < I_2 < I_3$

Sol 25: (B) Only (B) and (C) order is wrong

F > O, since IP increases across the period

N > O, as N has a stable half filled configuration

P > S, as P has a stable half filled configuration

Be > B, as Be has a stable fully filled configuration.

Sol 26: (A) If difference between $(IP)_1$ and $(IP)_2 > 16 \text{ eV}$, then element will not form a dipositive ion.

Sol 27: (A) Inconsistency in question, taking x^{-2} , no correct option

No. of electrons in $X^{-2} = 18 + 8 = 26$

Atomic number of $X^0 = 26 - 2 = 24$

No. of electrons in $Y^{-3} = 18 + 3 = 21$

Atomic number of Y = $21 + 3 = 24$

Sol 28: (A) Electronegativity increases across a period.

Sol 29: (C) The valency first increases as the number of unpaired electrons increases, but then starts decreasing as electrons begin to get paired.

Sol 30: (A) Electronegativity of carbon is least.

Sol 31: (B) Elements are arranged on the basis of increasing atomic number (i.e. increased nuclear charge)

Sol 32: (B) Metallic character decreases across the period and increases down the group

Sol 33: (A) Due to screening effect of inner electrons, the effective nuclear charge on the valence electrons decreases which lowers the IP.

Sol 34: (C) IE does not vary in a regular manner as IE is higher for elements which have stable half-filled or fully filled configurations compared to their adjacent elements, which violates the general trend that IE increases across a period. All other statements are correct.

Sol 35: (C) Just check the electronic configuration for each elements

In : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^1$ (iv)

Pd : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10}$ (iii)

Ca : $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ (i)

Cu : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$ (ii)

Sol 36: (C) Just check the configuration of each element, taking $n = 4$

Chromium : $[\text{Ar}] 3d^5 4s^1$

Copper : $[\text{Ar}] 3d^{10} 4s^1$

Krypton : $[\text{Ar}] 3d^{10} 4s^2 4p^6$

Bromine : $[\text{Ar}] 3d^{10} 4s^2 4p^5$

Sol 37: (A) Based on facts

Helium has highest ionization potential.

Fluorine has highest electronegativity.

Technitium is an artificial element.

Lithium has a very high reducing ability.

Sol 38: (C) Basic character decreases across a period.

Sol 39: (A) atomic radius, and Atomic size, decreases across a period electronegativity increases across a period. Effective nuclear charge increases across a period

Previous Years' Questions

Sol 1: (D) As we go down the group, inertness of ns^2 pair increase. Hence, tendency to exhibit +2 oxidation state increases and that of +4 oxidation state decreases.

Sol 2: (A) First I.E. of N > First I.E. of O.

Sol 3: (D) All the noble gases occupy the peaks of I.E. curve.

Sol 4: (D) Electronegativity increases on going from left to right in a period. Thus electronegativity of $F > O > N > Cl$.

Sol 5: (D) Nuclear charge per electron is greater in P^{5+} . Therefore its size is smaller.

Sol 6: (C) Element belongs to d-block is unnilhexium (U_{106}).

Sol 7: (C) I represents Li

II represents K

III represents Br

IV represents I

V represents He

So, amongst these, II represents most reactive metal and V represents least reactive non-metal.

Sol 8: (D) Positive ions will be smaller than parent atoms.

Sol 9: (E) Assertion is false but reason is true.

More is the electron affinity, greater is the oxidizing character.

Sol 10: (A) Both assertion and reason are true and reason is the correct explanation of assertion.

I.E. of N is more than that of $_xO$ as well as $_yC$.

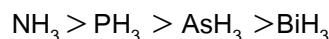
Sol 11: (C) Assertion is true but reason is false.

N is half-filled ($1s^2 2s^2 2p^3$) and therefore more stable and hence energy required to lose electron is greater.

Sol 12: (B) NO^+ , C_2^{2-} , CN^- and N_2 . All have fourteen electrons.

Sol 13: (A) Both are isoelectronic and have same bond order.

Sol 14: (C) Correct basic strength is



Sol 15: (B) Follow the periodic trends

Sol 16: (C) Option itself is the reason

Sol 17: (D) For isoelectronic species higher the $\frac{Z}{e}$ ratio, smaller the ionic radius

$$\frac{Z}{e} \text{ for } O^{2-} = \frac{8}{10} = 0.8$$

$$F^- = \frac{9}{10} = 0.9$$

$$Na^+ = \frac{11}{10} = 1.1$$

$$Mg^{2+} = \frac{12}{10} = 1.2$$

$$Al^{3+} = \frac{13}{10} = 1.3$$

Sol 18: (B) Greater charge and small size of cation cause more polarization and more covalent is that compound.

Sol 19: (D) Across a period metallic strength decreases & down the group it increases.

Sol 20: (C) As the size of central atom increases lone pair bond pair repulsions increases so, bond angle decreases.

Sol 21: (C) $CaCO_3 \rightarrow \underset{\text{Basic}}{CaO} + \underset{\text{Acidic}}{CO_2}$

Sol 22: (C) Order of increasing

$$\Delta H_{IE_1} : Ba < Ca < Se < S < Ar$$

Ba < Ca ; Se < S : On moving top to bottom in a group, size increases. So ionisation energy decreases.

Ar: Maximum value of ionisation energy, since it is an inert gas.

Sol 23: (B) $\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$ 1st I.E.

$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$ Electron gain enthalpy of Na^+

Because reaction is reverse so then.

$$\Delta H_{\text{eg}} = -5.1 \text{ eV.}$$

Sol 24: (C) These are isoelectronic species.

As negative charge increases, ionic radius increases.

(vi) $\text{As} < \text{P} < \text{S} < \text{Cl}$

EN decreases down the group and increases across the period

(vii) $\text{LiI} < \text{LiBr} < \text{LiCl}$

Ionic character decreases with decrease in electronegativity difference.

(viii) $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$

Mobility of ions increases down the group

(ix) $\text{Cs} < \text{Rb} < \text{K} < \text{Na} < \text{Li}$

Size of hydrated radii decreases down the group

(x) $\text{Ba}^{2+} < \text{Sr}^{2+} < \text{Ca}^{2+} < \text{Mg}^{2+} < \text{Be}^{2+}$

Extent of hydration decreases down the group

JEE Advanced/Boards

Exercise 1

Sol 1: 101: Unnilunium

102: Unnilbium

103: Unniltrium

104: Unnilquadium

105: Unnilpentium

106: Unnilhexium

107: Unnilseptium

108: Unniloctium

109: Unnilennium

Sol 2: (i) $\text{Pb}^{4+} < \text{Pb}^{2+} < \text{Pb}$

Size decreases with increasing positive charge.

(ii) $\text{Na} < \text{Al} < \text{Mg} < \text{Si}$

IP increases across a period.

More for Mg because of its stable half filled orbital configuration.

(iii) $\text{NiO} < \text{MgO} < \text{Sr} < \text{Rb}_2\text{O} < \text{Cs}_2\text{O}$

Basic character increases down a group and decreases across a period.

Less for transition metal oxides

(iv) $\text{Be}(\text{OH})_2 < \text{Mg}(\text{OH})_2 < \text{Ca}(\text{OH})_2 < \text{Ba}(\text{OH})_2$

Basic character increases down the group.

(v) $\text{Cl}^- < \text{Ar} < \text{K}^+ < \text{Ca}^{2+}$

Same configuration in all.

IE is decided by effective nuclear charge

Sol 3: Inert gases have an electronic configuration of ns^2np^6 which is an extremely stable configuration due to all fully filled orbitals. Hence, they have no tendency to combine with other atoms and change this configuration, so they are monoatomic.

Sol 4: The third period contains 8 elements as the energy of 3d orbital is more than 4s orbital. 3d orbital is filled during the fourth period making the number of elements in third period equal to 8.

Sol 5: (A) $\text{F}^- < \text{Cl}^- < \text{S}^{2-} < \text{P}^{3-}$

Ionic radius increase with increase in number of shells and negative charge

(B) $\text{Al}^{3+} < \text{Mg}^{2+} < \text{Na}^+ < \text{F}^- < \text{O}^{2-}$

Ionic radius decrease with increasing positive charge

(C) $\text{Mg}^{2+} < \text{Na}^+ < \text{K}^+$

Ionic radius decrease with increase in positive charge and increase with increase in number of shells.

Sol 6: We have,

Internuclear distance

$$d_{\text{H-Cl}} = r_{\text{H}} + r_{\text{Cl}} - 0.09 | r_{\text{H}} - r_{\text{Cl}} |$$

$$\text{Now, } r_{\text{Cl}} = \frac{10}{2} \text{ \AA} = 5 \text{ \AA}$$

$$r_{\text{H}} = \frac{2}{2} \text{ \AA} = 1 \text{ \AA}$$

$$x_{\text{Cl}} - x_{\text{H}} = 3 - 2.1 = 0.9$$

$$\therefore d_{\text{H-Cl}} = 5 + 1 - 0.09 \times 0.9 = 5.919$$

Hence, internuclear distance = 5.919 \AA

Sol 7: We have

$$d_{As-Cl} = r_{As} + r_{Cl} - 0.09 | EN_{Cl} - EN_{As} |$$

Here, $d_{As-Cl} = 2.20 \text{ \AA}$

$$r_{Cl} = 0.99 \text{ \AA}$$

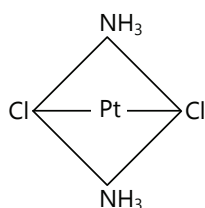
$$EN_{Cl} - EN_e = 0$$

$$\therefore r_{As} = d_{As-Cl} - r_{Cl} = 1.21 \text{ \AA}$$

Hence, SBCR of As is 1.21 \AA

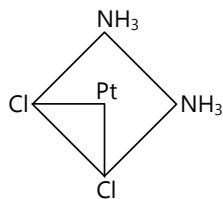
Sol 8: (A) We have Pt-Cl distance = 2.32 \AA from the figure, we have

$$Cl-Cl \text{ distance} = 2 \times (\text{Pt}-Cl \text{ distance}) = 4.64 \text{ \AA}$$



Hence Cl-Cl distance is 4.64 \AA

(B) Here, since this is a square planar geometry. Therefore, Cl-Pt-Cl angle will be equal to 90°



Therefore, Cl-Cl distance = $\sqrt{2}$ Pt-Cl distance is 3.28 \AA

Hence Cl-Cl distance is 3.28 \AA

Sol 9: Since, we know that size of cation is smaller than atomic size and size of anion is larger than atomic size. Therefore atomic radius of K $> 1.34 \text{ \AA}$ and atomic radius of F $< 1.34 \text{ \AA}$

Sol 10: Noble gases are chemically inert and are therefore monoatomic so, it is not possible to calculate covalent radius of noble gases, instead, for noble gases, Vander waals radius is calculated which is always greater than the radii of precedent halogens (which is covalent radius)

Sol 11: While going from Aluminium to Gallium, 3d electrons are filled which poorly screen the outermost electrons from nuclear charge hence, even though the number of shells increases, an increase in effective

nuclear charge will ensure that atomic size of Gallium is equal to the atomic size of Aluminium. Going further down from Ga to In- d-block will be there for all elements.

Hence, atomic radii will increase as usual.

Sol 12: We have $X_p = \frac{X_M}{2.8} = \frac{IE + EA}{5.6} = \frac{4 + 13}{5.6}$

$$\Rightarrow X_p = 3.03$$

Sol 13: In isoelectronic species, the increase in atomic radii means an increase in the effective nuclear charge on outermost shell electrons, as the number of electrons are the same. This increase in effective nuclear charge means that the valence electrons are more firmly bound to the nucleus, hence it results in a decrease in ionic/atomic radii.

Sol 14: Ionisation energy of Cu $>$ Ag is expected as ionization energy decreases down the group. But after Ag the 4f electrons start getting filled and since f orbital is the least penetrative it offers least screening and hence effective nuclear charge increases while going across lanthanoid series. This phenomenon is called lanthanoid contraction and due to this $(IE)_{Ag} < (IE)_{Au}$. As Au has more effective nuclear charge than Ag.

Sol 15: Electronic configuration of C = $[\text{He}] 2s^2 2p^2$

Electronic configuration of B = $[\text{He}] 2s^2 2p^1$

$(IE)_1$ of C $>$ $(IE)_1$ of B as effective nuclear charge of B is greater than that for C

Electronic configuration of B^+ = $[\text{He}] 2s^2$

Electronic configuration of C^+ = $[\text{He}] 2s^2 2p^1$

$(IE)_2$ of B $>$ $(IE)_2$ of C as B have a very stable fully filled orbital configuration which makes removal of electron extremely difficult.

Sol 16: Ionisation potential is the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom $M(g) + IE \rightarrow M^+(g) + e^-$

Sol 17: Electronic configuration of K = $[\text{Ar}] 4s^1$

Electronic configuration of Cu = $[\text{Ar}] 3d^{10} 4s^1$

We see that on both K and Cu have the same outer shell of $4s^1$ but for Cu the effective nuclear charge will be more than K, as the 3d orbital do not provide very good screening from nuclear charge, therefore, its (IE) is high.

Electronic configuration of $K^+ = [Ar]$

Electronic configuration of $Cu^+ = [Ar] 3d^{10}$

We note that K^+ has a very stable noble gas configuration compared to Cu whose configuration is not so stable. Hence, $(IE)_2$ of K is higher than $(IE)_2$ of Cu.

Sol 18: (i) For Ni, sum of four ionisation energies = 11.29 kJ/mol

For Pt, sum of four ionisation energies = 9.36 kJ/mol

Therefore, Ni can only exhibit +2 oxidation state.

Hence, most stable oxidation state for Ni: +2

Most stable oxidation state for Pt: +2, +4

(ii) Pt can easily form compounds in +4 oxidation state as $(IE)_1 + (IE)_2 + (IE)_3 + (IE)_4$ for Pt is not so high.

Sol 19: Electronic configuration of Mg: $[Ne] 3s^2$

Electronic configuration of Na: $[Ne] 3s^1$

Mg has a stable fully filled orbital configuration.

Hence (IE), of Mg is high

Electronic configuration of Mg^+ : $[Ne] 3s^1$

Electronic configuration of Na^+ : $[Ne]$

Na^+ has an extremely stable noble gas configuration which makes removal of electron very difficult.

Hence, $(IE)_2$ of Na is high

Sol 20: Na^+ has higher value of IE because for Na due to removal of an electron, effective nuclear charge is much more compared to Ne, as there are more protons in Na^+ than Ne for same number of p electrons.

Exercise 2

Single Correct Choice Type

Sol 1 (A) X has a stable half-filled orbital configuration which makes removal of electron difficult and increases its IE.

Sol 2: (A) This is the true position of lanthanoids. They are kept below the periodic table to make it more compact.

Sol 3: (B) Similar ionic potentials is the main cause of diagonal relationship.

Sol 4: (C) Will form a diatomic molecule most readily because it has an outer shell configuration of $2s^2 2p^2$ and it needs just 1 electron to attain noble gas configuration, which it can achieve by sharing an electron with another atom to form a diatomic molecule.

Sol 5: (A) We know,

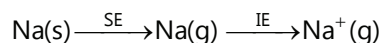
$$EN = \frac{EA + IP}{2}$$

$$\Rightarrow E_1 = \frac{EA + E_2}{2} \Rightarrow EA = 2E_1 - E_2$$

Sol 6: (A) This is $ns^2 np^4$ configuration which is the configuration of ground state of S and Cl^+ .

Sol 7: (C) O^- has no natural tendency to accept electrons because of coulombic repulsion between electrons, thus to add an electron to O^- , we have to give it energy from an external source.

Sol 8: (C) The process is



Therefore, total energy = SE + IE

Sol 9: (C) The addition of an extra electron to O^- requires addition of energy because of the coulombic repulsion offered by the extra electron to the incoming electron. For other processes, energy is evolved. As far, all of them, EA is positive.

Sol 10: (B) N_2 has no bond polarity and is a very neutral molecule compared to CN^- , which has a net negative charge and therefore will be much more reactive.

Previous Years' Questions

Sol 1: (B) Among isoelectronic species, greater the negative charge, greater the ionic size hence $F^- < O^{2-} < N^{3-}$

Sol 2: (B) In a group, ionization energy decreases down the group:
Be > Mg > Ca

Sol 3: (B) Be^- is the least stable ion, $Be(1s^2 2s^2)$ has stable electronic configuration, addition of electron decreases stability.

Sol 4: (B, C, D) Incorrect: Electrons are not filled in sub-energy levels s, p, d and f in the same sequence.

(B) **Correct:** No. of valence shell electrons usually determine the stable valency state of an elements.

(C) **Correct:** Physical and chemical properties of elements are periodic function of atomic number which is the basis of modern, long form, of periodic table.

(D) **Correct:** Relative ionicity of the bond between any two elements is function of electronegativity difference of the bonded atoms which in turn has periodic trend in long form of Periodic table.

Sol 5: (A, B) Correct: For greater solubility, hydration energy must be greater than lattice energy.

(B) **Correct:** Greater lattice energy discourage dissolution of salt.

(C) **Incorrect:** When a salt dissolve, energy is required to break the lattice, which comes from hydration process.

(D) **Incorrect:** Explained in (A).

Sol 6: (True) In a group, size increases from top to bottom.

Sol 7: (False) Ionization potential decreases down the group but this is not the only criteria of reducing power.

Sol 8: (i) Mg^{2+} , O^{2-} , Na^+ and F^- are all isoelectronic, has 10 electrons each. Among isoelectronic species, the order of size is cation < neutral < anion.

Also, between cations, higher the charge, smaller the size and between anions, greater the negative charge, larger the size. Therefore,

Ionic radii = $Mg^{2+} < Na^+ < F^- < O^{2-}$

(ii) First ionization energy increases from left to right in a period. However, exception occurs between group 2 and 13 and group 15 and 16 where trend is reversed on the grounds of stability of completely filled and completely half-filled orbitals. Therefore : Ionization energy (1st) : $Na < Al < Mg < Si$

(iii) If the atoms are from same period, bond length is related directly to atomic radius. Therefore, Bond length : $N_2 < O_2 < F_2 < Cl_2$

Sol 9: Cl^- , S^{2-} , Ca^{2+} , Ar

Size : $Ca^{2+} < Ar < Cl^- < S^{2-}$

Explained in (i), question 1.

Sol 10: $Zn : 3d^{10} 4s^2$, $Cu : 3d^{10}, 4s^1$

The first ionization energy is greater for Zn but reverse is true for 2nd ionization energy.

Sol 11: $Li^+ < Al^{3+} < Mg^{2+} < K^+$

Size decreases from left to right in a period and it increases from top to bottom in a group. Variation is more pronounced in group than in period.

Sol 12: (A) NO^- (16 electron system)

Bond order = 2.

NO^{\oplus} , CN^- and N_2 are isoelectronic with CO therefore all have same bond order (= 3)

Hence (A) is correct.

Sol 13: (C) The lower oxidation states for the group 14 elements are more stable for the heavier member of the group due to inert pair effect.

Sol 14: Na, F show only one non-zero oxidation state.

Sol 15: (B) The increasing order of atomic radii of the following Group 13 elements is $Ga < Al < Tl < In$