PERIODIC TABLE & PERIODICITY

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JEE (Advanced) Syllabus

Inorganic Chemistry –Classification of elements and Periodicity in Properties
Why do we need to Classify Elements? Genesis of Periodic Classifications,Modern Periodic Law and the present from of the Periodic Table, Nomenclature of elements with Atomic Number > 100, Electronic Configurations of elements and the Periodic Table, Electronic Configurations and Types of Elements s,p,d, f-Blocks, Periodic Trends in Properties of Elements.

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PERIODIC TABLE & PERIODICITY

INTRODUCTION :

The arrangement of all the known elements according to their properties in such a way that the elements of similar properties are grouped together in a tabular form is called periodic table.

Development of periodic table :

(A) LAVOISIER CLASSIFICATION OF ELEMENTS

At first Lavoisier classified the elements into two categories.

(i) Metal (ii) Non-metal

Note : This classification cannot categories the metalloid variety.

(B) DOBEREINER'S TRIAD RULE [1817]

(i) He made groups of three elements having similar chemical properties called TRIAD.

(ii) In Dobereiner triad, atomic weight of middle element is nearly equal to the average atomic weight of first and third element.

e.g.

Cl	Br	I	
35.5	$\boxed{80.0}$	127	$\frac{35.5+127}{2}=81.25$

Ca	Sr	Ba	
40	$\boxed{87.5}$	137	$\frac{40+137}{2}=88.5$

Li	Na	K	
7	$\boxed{23}$	39	At. wt of $Na=\frac{7+39}{2}=23$

(iii) Other examples – (K, Rb, Cs), (P, As, Sb), (S, Se, Te), (H, F, Cl), (Sc, Y, La)

(C) NEWLAND'S OCTAVE RULE [1865]

(i) He arranged the elements in the increasing order of their atomic mass and observed that properties of every 8th element was similar to the 1st one, like in the case of musical vowels notation.

Sa	Re	Ga	Ma	Pa	Dha	Ni	Sa
1	2	3	4	5	6	7	8

(ii) At that time inert gases were not known.

						H
Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca					

(iii) The properties of Li are similar to 8th element i.e. Na, Be are similar to Mg and so on.

Drawbacks or Limitations :

(a) This rule is valid only upto Ca. because after Ca due to filling of d-orbitals there is difference of 18 elements instead of 8 element.

(b) After the discovery of inert gas this law had to be dropped out.

(c) He failed in the case of heavier metals as Fe has been placed along with O and S.

(D) LOTHER MEYER'S CURVE [1869]

(i) He plotted a curve between atomic weight and atomic volume of different elements.

(ii) The following observations can be made from the curve –

(a) Most electropositive elements i.e. alkali metals (Li, Na, K, Rb, Cs etc.) occupy the peak portions of the curve.

(b) Less electropositive i.e. alkaline earth metals (Be, Mg, Ca, Sr, Ba) occupy the descending portions of the curve.

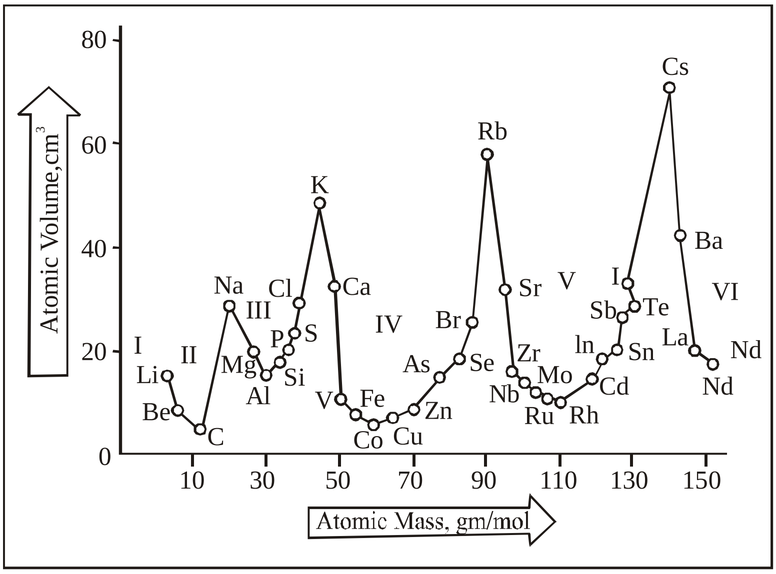
(c) Metalloids (B, Si, As, Te, At etc.) and transition metals occupy bottom part of the curve.

(d) Most electronegative i.e. halogens (F, Cl, Br, I) occupy the ascending portion of the curve.

Note : Elements having similar properties occupy similar portions of the curve.

Conclusion :

On the basis of this curve, Lother Meyer proposed that the physical properties of the elements are periodic function of their atomic weight and this became the basis of Mendeleev's periodic table.



(E) MENDELEEV'S PERIODIC TABLE [1869]

(i) ***Mendeleev's periodic law –*** The physical and chemical properties of elements are the periodic function of their atomic weight

(ii) ***Characteristic of Mendeleev's periodic table –***

(a) It is based on atomic weight

(b) 63 elements were known, noble gases were not discovered.

(c) He was the first scientist to classify the elements in a systamatic manner i.e. in horizontal rows and in vertical columns.

(d) Horizontal rows are called periods and there were 7 periods in Mendeleev's Periodic table.

(e) Vertical columns are called groups and there were 8 groups in Mendeleev's Periodic table.

(f) Each group upto VIIth is divided into A & B subgroups.'A' sub group elements are called normal elements and 'B' sub group elements are called transition elements.

(g) The VIIIth group consists of 9 elements in three rows.

(h) The elements belonging to same group exhibit similar properties.

(iii) Merits or advantages of Mendeleev's periodic table –

(a) Study of elements – For the first time, all known elements were classified in groups according to their similar properties. So study of the properties of elements became easier.

(b) Prediction of new elements – It gave encouragement to the discovery of new elements as some gaps were left in it.

Sc (Scandium), Ga (Gallium), Ge (Germanium), Tc (Technetium)

were the elements whose position and properties were well defined by Mendeleev even before their discoveries and he left the blank spaces for them in his table.

e.g. - Blank space at atomic weight 72 in silicon group was called Eka silicon (means properties like silicon) and element *(discovered later)* was named Germanium .

Similarly other elements discovered after mendeleev periodic table were :

Eka aluminium – Gallium(Ga)
Eka Boron – Scandium (Sc)
Eka Silicon – Germanium (Ge)
Eka Manganese – Technetium (Tc)

(c) Correction of doubtful atomic weights – Corrections were done in atomic weight of some elements.

$$\boxed{\text{Atomic Weight} = \text{Valency} \times \text{Equivalent weight.}}$$

Initially, it was found that equivalent weight of Be is 4.5 and it is trivalent (V = 3), so the weight of Be was 13.5 and there is no space in Mendeleev's table for this element. So, after correction, it was found that Be is actually divalent (V = 2). So, the weight of Be became 2 × 4.5 = 9 and there was a space between Li and B for this element in Mendeleev's table.

Corrections were done in atomic weight of elements are – U, Be, In, Au, Pt.

(iv) Demerits of Mendeleev's periodic table –

(a) Position of hydrogen – Hydrogen resembles both, the alkali metals (IA) and the halogens (VIIA) in properties so Mendeleev could not decide where to place it.

(b) Position of isotopes – As atomic weight of isotopes differs, they should have been placed in different position in Mendeleev's periodic table. But there was no such place for isotopes in Mendeleev's table.

(c) Anomalous pairs of elements – There were some pair of elements which did not follow the increasing order of atomic wts.

eg : Ar and Co were placed before K and Ni respectively in the periodic table, but having higher atomic weights.

$$\underset{\text{in modified mendeleev P.T.}}{\begin{pmatrix}\text{Ar} & \text{K}\\ 39.9 & 39.1\end{pmatrix}} \qquad \begin{pmatrix}\text{Te} & \text{I}\\ 127.5 & 127\end{pmatrix}$$

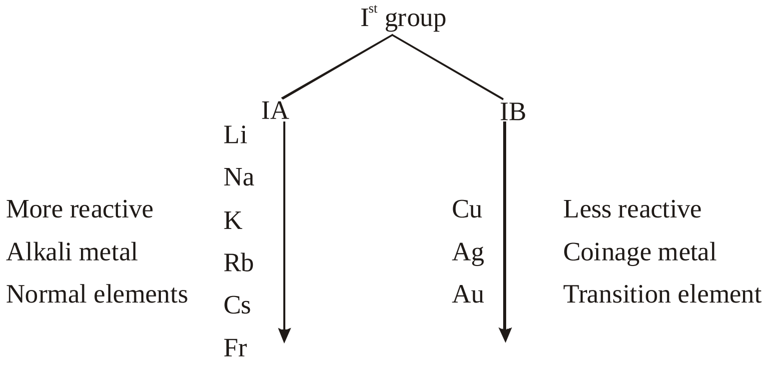
$$\begin{pmatrix}\text{Co} & \text{Ni}\\ 58.9 & 58.6\end{pmatrix} \qquad \begin{pmatrix}\text{Th} & \text{Pa}\\ 232 & 231\end{pmatrix}$$

(d) Like elements were placed in different groups.

There were some elements like Platinum (Pt) and Gold (Au) which have similar properties but were placed in different groups in Mendeleev's table.

Pt	Au
VIII	IB

(e) Unlike elements were placed in same group.



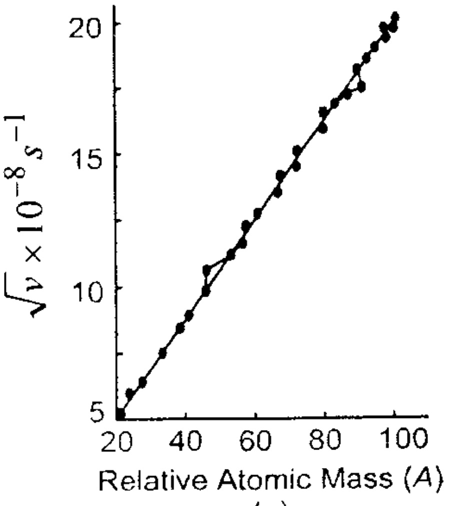
Cu, Ag and Au placed in Ist group along with Na, K etc. While they differ in their properties (Only similar in having ns^1 electronic configuration)

(F) MODERN PERIODIC TABLE (MODIFIED MENDELEEV PERIODIC TABLE)

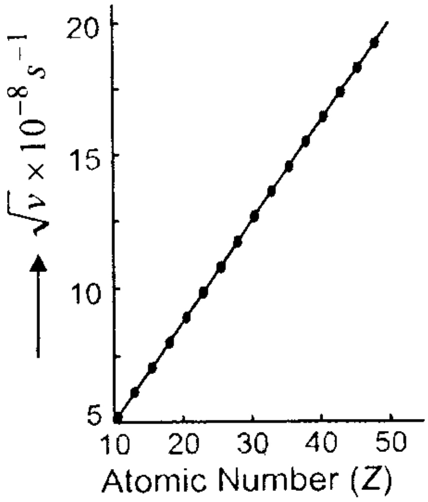
(i) It was proposed by **Moseley (1913)**.

(ii) Modern periodic table is based on atomic number.

(iii) Moseley did an experiment in which he bombarded high speed electron on different metal surfaces and obtained X-rays.



(a)



(b)

He found out that $\sqrt{\nu} \propto Z$ (where ν = frequency) of X-rays from this experiment, Moseley concluded that the physical and chemical properties of the elements are periodic function of their atomic number. It means that when the elements are arranged in the increasing order of their atomic number, elements having similar properties gets repeated after a regular interval. This is also known as **'Modern Periodic Law'**.

(iv) Modern periodic law – The physical & chemical properties of elements are a periodic function of their atomic number.

(v) Characteristics of modern periodic table –

(a) 9 vertical columns called groups.
(b) IA to VIIA, IB to VIIB, VIII and 0
(c) Inert gases were introduced in periodic table by Ramsay.
(d) 7 horizontal series called periods.

(G) LONG FORM / PRESENT FORM OF MODERN PERIODIC TABLE

(It is also called as 'Bohr-Burry & **Rang**, Werner Periodic Table.)

(i) It is based on the Bohr-Burry electronic configuration concept and atomic number.

(ii) This model was proposed by Rang & Werner

(iii) It consists of 7 horizontal periods and 18 vertical columns (groups)

(iv) According to I. U. P. A. C. 18 vertical columns are named as 1st to 18th group.

(v) The co-relation between the groups in long form of periodic table and in modern form of periodic table are given below.

IA,	IIA,	IIIB,	IVB,	VB,	VIB,	VIIB,	VIII,	IB,	IIB,	IIIA,	IVA,	VA,	VIA,	VIIA,	0
1	2	3	4	5	6	7	8, 9, 10	11	12	13	14	15	16	17	18

(vi) Elements belonging to same group have same no. of electrons in the outermost shell so their properties are similar.

Description of periods :

Period	n	Sub shell	No. of elements	Element	Name of Period
1.	1	1s	2	${}_{1}H$ – ${}_{2}He$	Shortest
2.	2	2s, 2p	8	${}_{3}Li$ – ${}_{10}Ne$	Short
3.	3	3s, 3p	8	${}_{11}Na$ – ${}_{18}Ar$	Short
4.	4	4s, 3d, 4p	18	${}_{19}K$ – ${}_{36}Kr$	Long
5.	5	5s, 4d, 5p	18	${}_{37}Rb$ – ${}_{58}Xe$	Long
6.	6	6s, 4f, 5d, 6p	32	${}_{55}Cs$ – ${}_{86}Rn$	Longest
7.	7	7s, 5f, 6d,	26	${}_{87}Fr$ – ${}_{112}Uub$	Incomplete

CLASSIFICATION OF ELEMENTS INTO s , p , d & f BLOCK ELEMENTS :

s – block :

(i) configuration ns^{1-2} **(ii)** last e^- enters in s orbital

(iii) two groups I A or 1 ; II A or 2

p – block :

(i) configuration $ns^2\ np^{1-6}$ **(ii)** last e^- enters in p orbital

(iii) six groups III A , IV A , V A , VI A , VII A , zero or 13 , 14 , 15 , 16 , 17 , 18

d – block : [Transition Elements]

(i) configuration $(n-1)\ d^{1-10}\ ns^{0-2}$ **(ii)** last e^- enters in d orbital

(iii) their two outermost shell are incomplete

(iv) 10 groups III B , IV B , V B , VI B , VII B , VIII (Triad) , I B , II B or
3 , 4 , 5 , 6 , 7 , (8 , 9 , 10) , 11 , 12 .

(v) four series 3 d , 4 d , 5 d , 6 d belong to 4th, 5th, 6th & 7th period respectively in long form of Periodic table.

f – block : [Inner Transition]

(i) configuration $(n-2)\ f^{0-14}\ (n-1)\ d^{0-2}\ ns^2$

(ii) last e^- enters in f orbital

(iii) two series 4 f Lanthanides & 5 f Actinides belong to 6th & 7th period respectively in long form of Periodic table.

Neil Bohr's classification of elements :

Using electronic configuration as the criteria, elements are of four types. The classification of the elements into these groups is dependent on the extent to which the s, p, d and f orbitals are filled.

Inert Gases :

(a) s – and p-orbitals of the outer most shell of these elements are completely filled. The outermost electronic configuration is ns^2np^6.

(b) Helium is also inert gas but its electronic configuration is $1s^2$

Representative or Normal Elements :

(a) Outermost shell of these elements is incomplete. The number of electrons in the outermost shell is less than eight.

(b) s-and p-block elements except inert gases are called normal or representative elements.

Transition Elements :

(a) These metals were placed between s-block metals and p-block elements so, are named transition metals.

(b) Their outermost electronic configuration is similar to d-block elements i.e. (n–1) d^{1-10} ns^{1-2}.

(c) Last two shells of these elements namely outermost and penultimate shells are incomplete.

(d) The last shell contains one or two electrons and the penultimate shell may contain more than eight and up to eighteen electrons.

(e) According to definition of transition elements, those elements which have partly filled d-orbitals in neutral state or in any stable oxidation state are called transition elements. According to this definition Zn, Cd and Hg (IIB group) are d-block elements but not transition elements because these elements have d^{10} configuration in neutral as well as in stable +2 oxidation state.

Inner Transition Elements :

(a) In these elements last three shells i.e. last, penultimate and prepenultimate shells are incomplete.

(b) These are related to IIIB i.e. group 3.

(c) The last shell contains two electrons. Penultimate shell may contain eight or nine electrons and pre-penultimate shell contains more than 18, up to 32 electrons.

(d) Their outemost electronic configuration is similar to f-block element

i.e. $(n-2)\,f^{\,0-14}\,(n-1)\,d^{\,0-2}\,ns^2$

METALS, NON-METALS & METALLOIDS

Apart from classifying elements into s, p, d and f-blocks, there is yet another broad classification of elements based on their properties. The elements can be broadly classified into

(a) Metals:

Majority of the elements in periodic table are metals and appears on the left side of the periodic table.

Properties:

(i) These are usually solid at room temperature [exception - mercury]

(ii) They have high melting and boiling point [exception Gallium & Cesium have very low melting point (303 K and 302 K respectively)]

(iii) They are good conductor of heat and electricity.

(iv) They are malleable (can be flattened into thin sheets by hammering) and ductile (can be drawn into wires)

(b) Non-Metals:

These are placed at the top right hand side of periodic table. As we move horizontally along a period, the property of elements changes from metallic (on left) to non-metallic (on the right).

Properties :

(i) These are usually solids or gases at room temperature.

(ii) They have low melting point and boiling point (exception : Boron, Carbon).

(iii) Most Non-metallic solids are brittle and are neither malleable nor ductile.

(c) Metalloids (Semi-metals):

Properties of these elements show the characteristics of both metals and non-metals. Silicon (Si), Germanium(Ge), Arsenic(As), Antimony(Sb) and Tellurium(Te) are metalloids.

ESTIMATING POSITION OF AN ELEMENT FROM ITS ELECTRONIC CONFIGURATION

The last electron enters in which subshell gives idea of its block.

[☺ ***Think*** :$1s^1$ and $1s^2$ belongs to which block]

Period number = Principal quantum number of valence shell electron in ground state electronic configuration.

Group number for s block = number of valence shell electrons

Group number for p block = 10 + number of valence shell electrons

Group number for d block = number of [ns + (n–1) d] electrons

Group number for f-block = 3

[☺ Use these carefully while locating the position.]

IUPAC NOMENCLATURE OF THE ELEMENT :

The names are derived by using roots for the three digits in the atomic number of the element and adding the ending –ium. The roots for the number are

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

Thus element with atomic number 109 will be named as **une** (**u** for 1, **n** for 0 and **e** for 9). Table summarises the names of the elements with atomic number above 100.

PERIODIC TABLE & PERIODICITY

Diagonal relationship :

Some elements of certain groups of 2nd period resemble much in properties with the elements of third period of next group i.e. elements of second and third period are diagonally related in properties. This phenomenon is known as diagonal relationship. For example, the similarity between lithium (the first member of group 1) and magnesium (the second element in group 2) is called a diagonal relationship. Diagonal relationship also exist between other pairs of elements Be and Al, B and Si as shown in figure ;

$$\begin{tikzcd}
2^{nd}\text{ period} & \text{Li} \arrow[dr] & \text{Be} \arrow[dr] & \text{B} \arrow[dr] & \text{C} \\
3^{rd}\text{ period} & \text{Na} & \text{Mg} & \text{Al} & \text{Si}
\end{tikzcd}$$

Diagonal relationship arises because of ;

(i) on descending a group, the atoms and ions increase in size. On moving from left to right in the periodic table, the size decreases. Thus on moving diagonally, the size remains nearly the same.

($Li = 1.23$ Å & $Mg = 1.36$ Å ; $Li^+ = 0.76$ Å & $Mg^{2+} = 0.72$ Å)

(ii) it is sometimes suggested that the diagonal relationship arises because of diagonal similarity in electronegativity values.

($Li = 1.0$ & $Mg = 1.2$; $Be = 1.5$ & $Al = 1.5$; $B = 2.0$ & $Si = 1.8$)

(iii) Be and Al also show a diagonal relationship. In this case sizes are not so close ($Be^{2+} = 0.45$ Å and $Al^{3+} = 0.535$ Å but the charge per unit area is nearly similar ($Be^{2+} = 2.36$ and $Al^{3+} = 2.50$) because the charges are 2+ and 3+ respectively.

$$\text{Charge per unit area} = \frac{(\text{ionic charge})}{\frac{4}{3}.\pi.(\text{ionic radius})}$$

The periodicity of atomic properties :

(i) Effective nuclear charge :

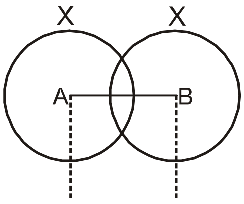
Between the outer most valence electrons and the nucleus of an atom, there exists number of shells containing electrons. Due to the presence of these intervening electrons, the valence electrons are unable to experience the attractive pull of the actual number of protons in the nucleus. These intervening electrons act as shield between the valence electrons and protons in the nucleus. Thus, the presence of intervening (shielding) electrons reduces the electrostatic attraction between the protons in the nucleus and the valence electrons because intervening electrons repel the valence electrons. The concept of effective nuclear charge helps in understanding the effects of shielding on periodic properties. The effective nuclear charge (Z_{eff}) is the charge actually felt by the valence electron Z_{eff} is given by the formula $Z_{eff} = Z - \sigma$, (where Z is the actual nuclear charge (atomic number of the element) and σ is the shielding (screening constant). The value of σ i.e. shielding effect can be determined using the Slater's rule.

(ii) Atomic radius :

Probability of finding the electron is never zero even at large distance from the nucleus. Based on probability concept, an atom does not have well defined boundary. Hence exact value of the atomic radius can't be evaluated. Atomic radius is taken as the effective size which is the distance of the closest approach of one atom to another atom in a given bonding state.

Atomic radius can be

(A) **Covalent radius :** It is one-half of the distance between the centres of two nuclei (of like atoms) bonded by a single covalent bond as shown in figure.

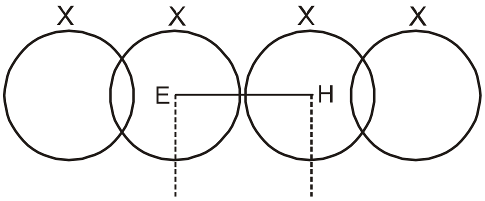


$r_{covalent} = \frac{1}{2}AB$ (of element X)

- Covalent radius is generally used for nonmetals.

(B) **Vander Waal's radius (Collision radius) :**
It is one-half of the internuclear distance between two adjacent atoms in two nearest neighbouring molecules of the substance in solid state as shown in figure.

- Vander Waal's radius does not apply to metal.
- Its magnitude depends upon the packing of the atoms when the element is in the solid state.



$r_{vander\ Waal's} = \frac{1}{2}EH$

Comparison of Covalent radius and Vander Waal's radius

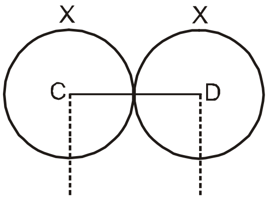
- The Vander Waal's force of attraction is weak, therefore, the internuclear distance in case of atoms held by Van der Waal's force is much larger than those of held covalent bond. Therefore Vander Waal's radii are always larger than covalent radii.

- A covalent bond is formed by the overlapping of two half-filled atomic orbitals and in this a part of the orbital becomes common. Therefore, covalent radii are always smaller than the Van der Waals radii. For example,

Table-5

Elements	H	O	F	S	Br
Covalent radius (Å)	0.37	0.66	0.64	1.04	1.11
van der Waal's radius (Å)	1.20	1.40	1.35	1.85	1.95

(C) **Metallic radius (Crystal radius) :**
It is one-half of the distance between the nuclei of two adjacent metal atoms in the metallic crystal lattice as shown in the figure.



$r_{crystal} = \frac{1}{2}CD$ (of element X)

- Metallic radius of an element is always greater than its covalent radius. It is due to the fact that metallic bond (electrical attraction between positive charge of an atom and mobile electrons) is weaker than covalent bond and hence the internuclear distance between the two adjacent atoms in a metallic crystal is longer than the internuclear distance between the covalently bonded atom.

For example :

	Metallic radius	Covalent radius
K	231 pm	203 pm
Na	186 pm	154 pm

- Thus the magnitude of the above three radii follow the below order.

$$r_{covalent} < r_{crystal} < r_{vander\ Walls}$$

Table-6

Variation in a Period (Left to Right)	**Variation in a Group (Top to Bottom)**
1. Nuclear charge (Z) increases by one unit	**1.** Nuclear charge (Z) increases by more than one unit
2. Effective nuclear charge (Z_{eff}) also increases	**2.** Effective nuclear charge (Z_{eff}) almost remains constant because of increased screening effect of inner shells electrons.
3. As a result, the electrons are pulled closer to the nucleus by the increased Z_{eff}. $r_n \propto \frac{1}{Z^*}$ Hence atomic radii decrease with increase in atomic number in a period from left to right.	**3.** The effect of increased number of atomic shells overweighs the effect of increased nuclear charge. As a result of this the size of atom increases from top to bottom in a given group.
4. But number of orbitals (n) remains constant	**4.** But number of orbitals (n) increases.

- The atomic radius of inert gas es (zero group) is quoted as highest in magnitude in their respective period because it is sine for then Vander Waal's radius is quoted. The Vander Waal's radius of inert gases also increases from top to bottom in a group.

(iii) Ionic radius :

The effective distance from the centre of nucleus of the ion up to which it has an influence in the ionic bond is called ionic radius.

Table-7

Cation	**Anion**
(1) It is formed by the loss of one or more electrons from the valence shell of an atom of an element. **(2)** Cations are smaller than the parent atoms because, (a) the whole of the outer shell of electrons is usually removed. (b) in a cation, the number of positive charges on the nucleus in greater than number of orbital electrons leading to incresed inward pull of remaining electrons causing contraction in size of the ion.	**(1)** It is formed by the gain of one or more electrons from in the valence shell of an atom of an element. **(2)** Anions are larger than the parent atoms because (a) anion is formed by gain of one or more electrons in the neutral atom and thus number of electrons increases but magnitude of nuclear charge remains the same. (b) nuclear charge per electrons is thus reduced and the electrons cloud is held less tightly by the nucleus leading to the expansion of the outer shell. Thus size of anion is increased.

- The sizes of ions increases as we go down a group (considering the ions of same charge). For example :

  Li^+ (0.76) < Na^+ (1.02) < K^+ (1.38) < Rb^+ → (in Å)
  $Be^{2+} < Mg^{2+} < Ca^{2+} < Sr^{2+}$
  $F^- < Cl^- < Br^- < I^-$

- For the cations same element the ionic radius decreases as more electrons are lost (i.e. the valency increases) For e.g. Cr^{2+} = 0.80 Å, Cr^{3+} = 0.615 Å, Cr^{4+} = 0.55 Å, Cr^{5+} = 0.49 Å and Cr^{6+} = 0.44 Å.

- The d and f orbitals do not shield the nuclear charge very effectively .Therefore there is significant reduction in the size of the ions, just after d or f orbitals have been filled completely. This is called a lanthanide contraction. Atomic radii of Hf (Hf^{4+} = 0.71 Å) and Zr (Zr^{4+} = 0.72 Å), and Nb (Nb^{3+} = 0.72 Å) and Ta (Ta^{3+} = 0.72 Å) are almost identical due to lanthanide contraction.

- The species containing the same number of electrons but differ in the magnitude of their nuclear charges are called as isoelectronic species. For example, N^{3-}, O^{2-}, F^{-}, Ne, Na^{+}, Mg^{2+} and Al^{3+} are all isoelectronic species with same number of electrons (i.e 10) but different nuclear charges of +7, +8, +9, +10, +11, +12 and +13 respectively.

  Within a series of isoelectronic species as the nuclear charge increases, the force of attraction by the nucleus on the electrons also increases. As a result, the ionic radii of isoelectronic species decrease with increases in the magnitude of nuclear charges. For example, as shown in figure.

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

  $\longrightarrow$

  Ionic radii increase

  $\longrightarrow$

  As effective nuclear charge decrease.

- Following are the examples of isoelectronic series

  (i) S^{2-}, Cl^{-}, K^{+}, Ca^{+2}, Sc^{+3} (ii) SO_2, NO_3^{-}, CO_3^{2-}, (iii) N_2, CO, CN^{-} (iv) NH_3, H_3O^{+}

- Pauling's empirical formula for ionic radius $\propto \dfrac{1}{\text{nuclear charge}}$ (only for isoelectronic species).

Solved Examples

Ex.6 Consider the following species :

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

(a) What is common in them?

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

Sol. (a) Each one of these ions contains 10 electrons and hence all are isoelectronic ions.

(b) The ionic radii of isoelectronic ions decrease with the increase in the magnitude of the nuclear charge. For example, consider the isoelectronic ions : N^{3-}, O^{2-}, F^{-}, Na^{2+} and Al^{3+} . All these ionse have 10 electrons but their nuclear charges increase in the order :

N^{3-} (+ 7), O^{2-} (+ 8), F^{-} (+ 9), Na^{+} (+ 11), Mg^{2+} (+ 12) and Al^{3+} (+ 13). Therefore, their ionic radii decrease in the order : $N^{3-} > O^{2-} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+}$.

Ex.7 Mg^{2+} is smaller than O^{2-} in size, though both have same electronic configuration. Explain ?

Sol. Mg^{2+} and O^{2-} both are isoelectronic i.e., have same number of electrons. But Mg^{2+} having 12 protons in its nucleus exerts higher effective nuclear charge than O^{2-} having 8 protons and thus valence shell as well as inner shells electrons are more strongly attracted by the nucleus in Mg^{2+} resulting smaller size than O^{2-}.

(iv) Ionisation Energy :

Ionisation energy (IE) is defined as the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom to form a cation.

$M(g) \xrightarrow{(IE_1)} M^{+}(g) + e^{-}$; $M^{+}(g) + IE_2 \longrightarrow M^{2+}(g) + e^{-}$

$M^{2+}(g) + IE_3 \longrightarrow M^{+3}(g) + e^{-}$

IE_1, IE_2 & IE_3 are the I^{st}, II^{nd} & III^{rd} ionization energies to remove electron from a neutral atom, monovalent and divalent cations respectively.

In general for the same element, $(IE)_1 < (IE)_2 < (IE)_3 <$ .... because, as the number of electrons decreases, the attraction between the nucleus and the remaining electrons increases considerably and hence subsequent ionization energies increase.

- Ionization energies are determined from spectra and are measured in kJ mol^{-1}, k Cal mol^{-1}, eV (electron volt).

- **Factors Influencing Ionisation energy**

Variation in ionization energies in a period and group may or may not be regular and can be influenced by the following factors.

(A) Size of the Atom (Dominant factor in case of group) :

Generally ionisation energy decreases with increase in the atomic size.

As the distance between the outer most electrons and the nucleus increases, the force of attraction between the valence shell electrons and the nucleus decreases. As a result, outer most electrons are held less firmly and lesser amount of energy is required to knock them out.

For example, ionisation energy decreases in a group from top to bottom with increase in atomic size.

(B) Nuclear Charge (Dominant factor in case of group) :

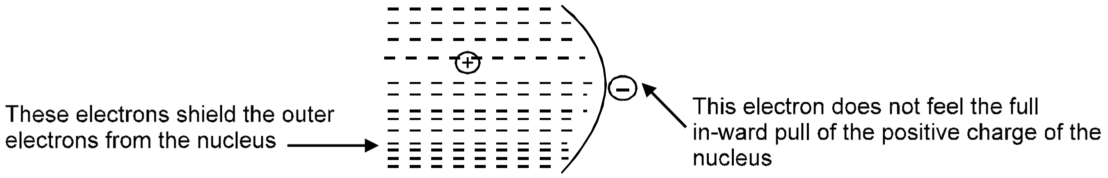
The ionisation energy increases with increase in the nuclear charge.

This is due to the fact that with increase in the nuclear charge, the electrons of the outer most shell are more firmly held by the nucleus and thus greater amount of energy is required to pull out an electron from the atom.

For example, ionisation energy increases as we move from left to right along a period due to increase in nuclear charge.

(C) Shielding or screening effect :

The electrons in the inner shells act as a screen or shield between the nucleus and the electrons in the outer most shell. This is called shielding effect. The larger the number of electrons in the inner shells, greater is the screening effect and smaller the force of attraction and thus ionization energy (IE) decreases.



(D) Penetration effect of the electron :

The ionization energy also depends on the type of the electron which is to be removed.Reason being : s, p, d and f electrons have orbitals with different shapes. An s electron penetrates closer to the nucleus, and is therefore more tightly held than a p electron. Similarly p-orbital electron is more tightly held than a d-orbital electron and a d-orbital electron is more tightly held than an f-orbital electron. If other factors being equal, ionisation energies are in the order s > p > d > f.

For example, ionisation energy of aluminium is comparatively less than magnesium because outer most electron is to be removed from 3p-orbital (having lesser penetration effect) in aluminium where as in magnesium it will be removed from 3s-orbital (having larger penetration effect) of same energy level.

(E) Electronic Configuration :

If an atom has exactly half-filled or completely filled orbitals, then such an arrangement is said to have extra stability.

The removal of an electron from such an atom requires more energy than expected. For example, first ionisation energy of beryllium is greater than boron because beryllium has extra stable completely filled outer most 2s orbital while boron has partially filled less stable outer most 2p-orbital.

Be (Z = 4) $1s^2$, $2s^2$ B (Z = 5) $1s^2$, $2s^2$, $2p^1$

Similarly noble gases have completely filled electronic configurations and hence they have highest ionisation energies in their respective periods.

- Metallic or electropositive character of elements increases as the value of ionisation energy decreases.
- The relative reactivity of the metals in gaseous phase increases with the decrease in ionisation energy.
- The reducing power of elements in the gaseous phase increases as the value of ionisation energy decreases. Among alkali metals, lithium is the strongest reducing agent in aqueous solution.

Solved Examples

Ex.8 Energy of an electron in the ground state of the hydrogen atom is -2.18×10^{-18} J. Calculate the ionization enthalpy of atomic hydrogen in terms of J mol^{-1}.

Sol. Ionization energy is the amount of energy required to remove the electron from the ground state to infinity.
Now, energy of the electron in the ground state = -2.18×10^{-15} J
Energy of the electron at affinity = 0.
The energy required to remove an electron in the ground state of hydrogen atom
= 0 – (its energy in the ground state) = $-(-2.18 \times 10^{-18}$ J) = 2.18×10^{18} J.

$\therefore$ Ionization enthalpy per mole of hydrogen atoms = $\frac{2.18 \times 10^{18} \times 6.02 \times 10^{23}}{1000}$ kJ

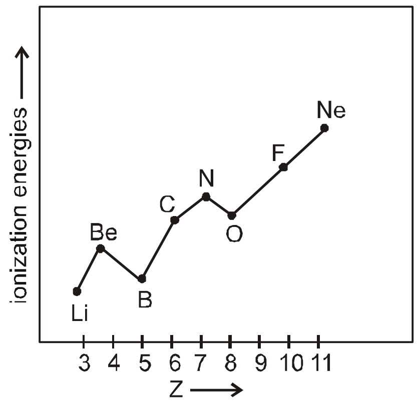
= 1312.36 kJ mol^{-1} = 1312.36×10^{3} J mol^{-1}.

Ex.9 Among the second period elements the actual ionization enthalpies are in the order Li < B < Be < C < O < N < F < Ne.
Explain why :
(i) Be has higher $\Delta_i H$ than B
(ii) O has lower $\Delta_i H$ than N and F?

Sol. (i) The ionization enthalpy, among other things depends upon the type of electron to be removed from the same principal shell. In case of Be ($1s^2\ 2s^2$) the outermost electron is present in 2s-orbital while in B ($1s^2\ 2s^2\ 2p^1$) it is present in 2p- orbital. Since 2s-electrons are more strongly attracted by the nucleus than 2p-electrons, therfore, lesser amount of energy is required to knock out a 2p-electron than a 2s-electron. Consequently, $\Delta_i H$ of Be is higher than that $\Delta_i H$ of B.

(ii) The electronic configration of N ($1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$) in which 2p-orbitals are exactly half-filled is more stable than the electronic configuration of O ($1s^2\ 2s^2\ p_x^2\ 2p_y^1\ 2p_z^1$) in which the 2p- orbitals are neither exactly half-filled nor completely filled. Therefore, it is difficult to remove an electron from N than from O. As result, $\Delta_i H$ of N is higher than that of O. Further, the electronic configuration of F is $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^1$. Because of higher nuclear charge (+9) the first ionization enthalpy of F is higher than that of O. Further, the effect of increased nuclear charge outweights the effect of stability due to exactly half-filled orbitals, therefore, the $\Delta_i H$ of N and O are lower than that of F.

Ex.10 Following graph shows variation of ionization energies with atomic number in second period (Li–Ne). Value of ionization energies of Na(11) will be –



(1) above Ne	(2) below Ne but above O.
(3*) below Li	(4) between N and O.

Sol. Na is 3rd period element and is bigger than Li. The distance between the nucleus and outer most electron is more as compared to Li. Thus the outer most electron is loosely bound with nucleus and removal of electron is easier. So option (3) is correct.

(V) Electron Affinity :

The electron gain enthalpy $\Delta_{eg}H^{\ominus}$, is the change in standard molar enthalpy when a neutral gaseous atom gains an electron to form an anion.

$$X(g) + e^-(g) \longrightarrow X^-(g)$$

Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion. Electron gain may be either exothermic or endothermic depending on the elements.

When an electron is added to the atom and the energy is released, the electron gain enthalpy is negative and when energy is needed to add an electron to the atom, the electron gain enthalpy is positive.

Although the electron gain enthalpy is the thermodynamically appropriate term, much of inorganic chemistry is discussed in terms of a closely related property, the electron affinity, E_a of an element which is the difference in energy between the gaseous atoms and the gaseous ions at T = 0.

$E_a = E(X, g) - E(X^-, g)$,

Electron gain enthalpy is conventionally defined as the energy released when an electron is added to the valence shell of an isolated gaseous atom.

$$F + e^- \longrightarrow F^- \quad E_a = 328 \text{ kJ mol}^{-1} ; \Delta_{eg}H^{\ominus} = -328 \text{ kJ mol}^{-1}$$

Although the precise relation is $\Delta_{eg}H^{\ominus} = -E_a - \frac{5}{2}RT$, the contribution $\frac{5}{2}RT$ is commonly ignored.

A positive electron gain enthalpy indicates that the ion X^- has a lower, more negative energy than the neutral atom X. The second electron gain enthalpy, the enthalpy change for the addition of a second electron to an initially neutral atom, invariably positive because the electron repulsion out weighs the nuclear attraction.

$$O(g) + e^-(g) \xrightarrow{\text{Exothermic}} O^-(g) ; \qquad E_a = 140.9 \text{ kJ mol}^{-1} ; \Delta_{eg}H^{\ominus} = -140.9 \text{ kJ mol}^{-1}$$

$$O^-(g) + e^-(g) \xrightarrow{\text{Endothermic}} O^{2-}(g) ; \qquad E_a = -744 \text{ kJ mol}^{-1} ; \Delta_{eg}H^{\ominus} = 744 \text{ kJ mol}^{-1}$$

An element has a high electron affinity if the additional electron can enter a shell where it experiences a strong effective nuclear charge.

- Group 17 elements (halogens) have very high negative electron gain enthalpies (i.e. high electron affinity) because they can attain stable noble gas electronic configuration by picking up an electron.
- Across a period, with increase in atomic number, electron gain enthalpy becomes more negative because left to right across a period effective nuclear charge increases and consequently it will be easier to add an electron to a small atom.
- As we move in a group from top to bottom, electron gain enthalpy becomes less negative because the size of the atom increases and the added electron would be at larger distance from the nucleus.
- Noble gases have large positive electron gain enthalpies because the electron has to enter the next higher energy level leading to a very unstable electronic configuration.
- Negative electron gain enthalpy of O or F is less than S or Cl. This is due to the fact that when an electron is added to O or F, the added electron goes to the smaller n = 2 energy level and experiences significant repulsion from the other electrons present in this level. In S or Cl, the electron goes to the larger n = 3 energy level and consequently occupies a larger region of space leading to much less electron-electron repulsion.
- Electron gain enthalpies of alkaline earth metals are very less or positive because the extra electron is to be added to completely filled s-orbitals in their valence shells.
- Nitrogen has very low electron affinity because there is high electron repulsion when the incoming electron enters an orbital that is already half filled.
- (i) Electron affinity $\propto \frac{1}{\text{Atomic size}}$ (ii) Electron affinity $\propto$ Effective nuclear charge (z_{eff})

  (iii) Electron affinity $\propto \frac{1}{\text{Screening effect}}$. (iv) Stability of half filled and completely filled orbitals of a subshell is comparatively more and the addition of an extra electron to such an system is difficult and hence the electron affinity value decreases.

Solved Examples

Ex.11 Consider the elements N, P, O and S and arrange them in order of increasing negative electron gain enthalpy.

Sol. Order of increasing negative electron gain enthalpy is $N < P < O < S$. For detail refer text.

Ex.12 Why do halogens have high electron gain enthalpies (i.e. $-\Delta_{eg}H^{\ominus}$) ?

Sol. The valence shell electronic configuration of halogens is ns^2np^5 and thus they require one electron to acquire the stable noble gas configuration ns^2np^6. Because of this they have strong tendency to accept an additional electron and hence have high electron gain enthalpies.

Ex.13 Which will have the maximum value of electron affinity O^x ,O^y, O^z [x,y and z respectively are 0, – 1 and –2]?

(1) O^x (2) O^y (3) O^z (4) All have equal.

Sol. Being neutral atom oxygen will have higher electron affinity as there is electrostatic repulsion between additional electron and negative ion in case of O^- and O^{2-}. So option (A) is correct.

Ex.14 Account for the large decrease in electron affinity between Li and Be despite the increase in nuclear charge.

Sol. The electron configurations of Li and Be are $[He]2s^1$ and $[He]2s^2$, respectively. The additional electron enters the 2s orbital of Li but the 2p orbital of Be and hence is much less tightly bound. In fact, the nuclear charge is so well shielded in Be that electron gain is endothermic.

(VI) Electronegativity :

Electronegativity is a measure of the tendency of an element to attract shared electrons towards itself in a covalently bonded molecules.

The magnitude of electronegativity of an element depends upon its ionisation potential & electron affinity.

Higher ionisation potential & electron affinity values indicate higher electronegativity value.

- With increase in atomic size the distance between nucleus and valence shell electrons increases, therefore, the force of attraction between the nucleus and the valence shell electrons decreases and hence the electronegativity values also decrease.
- With increase in nuclear charge force of attraction between nucleus and the valence shell electrons increases and, therefore, electronegativity value increases
- In higher oxidation state, the element has higher magnitude of positive charge.
  Thus, due to more positive charge on element, it has higher polarising power.
  Hence with increase in the oxidation state of element, its electronegativity also increases.
  Charge on cation α electronegativity of the atom.
- The electronegativity also increases as the s-character in the hybrid orbitals increases.

Hybrid orbital	sp^3	sp^2	sp
s-character	25%	33%	50%

$\longrightarrow$

Electronegativity increases

Table-8

Variation of electronegativity in a group	Variation of electronegativity in a period
On moving down the groups, Z increases but Z_{eff} almost remains constant, number of shells (n) increases, r_n (atomic radius) increases. Therefore, electronegativity decreases moving down the groups.	While moving across a period left to right, Z, Z_{eff} increases & r_n decreases. Therefore, electronegativity increases along a period.

There is no direct method to measure the value of electronegativity, however, there are some scales to measure its value.

(a) Pauling's scale :

Linus Pauling developed a method for calculating relative electronegativities of most elements. According to Pauling

$$\Delta = X_A - X_B = 0.208\sqrt{E._{A-B} - \sqrt{E_{A-A} \times E_{B-B}}}$$

$E_{A\text{-}B}$ = Bond enthalpy/ Bond energy of A – B bond.
E_{A-A} = Bond energy of A – A bond
E_{B-B} = Bond energy of B – B bond

(All bond energies are in kcal / mol)

$$\Delta = X_A - X_B = 0.1017\sqrt{E._{A-B} - \sqrt{E_{A-A} \times E_{B-B}}}$$

All bond energies are in kJ / mol.

(b) Mulliken's scale :

Electronegativity χ (chi) can be regarded as the average of the ionisation energy (IE) and the electron affinity (EA) of an atom (both expressed in electron volts).

$$\chi_M = \frac{IE + EA}{2}$$

Paulings's electronegativity χ_P isrelated to Mulliken's electronegativity χ_M as given below.

$$\chi_P = 1.35\,(\chi_M)^{1/2} - 1.37$$

Mulliken's values were about 2.8 times larger than the Pauling's values.

(c) Allred–Rochow's Electronegativity (χ_{AR}): Allred and Rochow defined electronegativity as the force exerted by the nucleus of an atom on its valence electrons :

$$\chi_{AR} = \frac{0.359\,Z_{\text{effective}}}{r^2} + 0.744$$

where $Z_{\text{effective}}$ is the effective nuclear charge and r the covalent radius in Å.

or $$0.744 + \frac{35.90\,Z_{\text{effective}}}{(r/pm)^2}$$

- According to the Allred-Rochow definition, elements with high electronegativity are those with high effective nuclear charge and the small covalent radius. Such elements lie close to fluorine.

Solved Examples

Ex.15 Calculate the electronegativity of carbon from the following data :

E_{H-H} = 104.2 kcal mol^{-1} , E_{C-C} = 83.1 kcal mol^{-1}
E_{C-H} = 98.8 kcal mol$^{-1|}$, X_H = 2.1

Sol. Let the electronegativity of carbon be X_C, applying Pauling equation,

$$X_C - X_H = 0.208\,[E_{C-H} - \frac{1}{2}\,(E_{C\text{-}C} + E_{H-H})]^{1/2}$$

$$x_C - 2.1 = 0.208\,[98.8 - \frac{1}{2}(83.1 + 104.2)]^{1/2}$$

$\Rightarrow x_C = 2.5$ **Ans.** 2.5

Ex.16 If electronegativity of x be 3.2 and that of y be 2.2, the percentage ionic character of xy is –

(1) 19.5 (2) 18.5 (3) 9.5 (4) 29.5.

Sol. $EN_x - EN_y = 3.2 - 2.2 = 1.$

$\Delta = 1$

[Δ = difference of electronegativity values between x and y].

% ionic character = $16\Delta + 3.5\Delta^2 = 19.5$.

Ans. (1)

(VII) Periodicity of Valence or Oxidation States :

The valence of representative elements is usually (though not necessarily) equal to the number of electrons in the outermost orbitals and / or equal to eight minus the number of outermost electrons. Nowadays the term oxidation state is frequently used for valence.

Consider the two oxygen containing compounds : OF_2 and Na_2O. The order of electronegativity of the three elements involved in these compounds is F > O > Na. Each of the atoms of fluorine, with outer electronic configuration $2s^22p^5$, shares one electron with oxygen in the OF_2 molecule. Being highest electronegative element, fluorine is given oxidation state –1. Since there are two fluorine atoms in this molecule, oxygen with outer electronic configuration $2s^22p^4$ shares two electrons with fluorine atoms and thereby exhibits oxidation state +2. In Na_2O, oxygen being more electronegative accepts two electrons, one from each of the two sodium atoms and thus, shows oxidation state –2. On the other hand sodium with electronic configuration $3s^1$ loses one electron to oxygen and is given oxidation state +1. Thus, the oxidation state of an element in a particular compound can be defined as the charge acquired by its atom on the basis of electronegative consideration from other atoms in the molecule.

There are many elements which exhibit variable valence. This is particularly characteristic of transition elements and actinoids.

MISCELLANEOUS SOLVED PROBLEMS

1. A M^{2+} ion derived from a metal in the first transition metal series has four electrons in 3d subshell. What element might M be ?

Sol. Electron configuration of M^{2+} is : $[Ar]^{18}\,4s^0\,3d^4$

$\therefore$ Electron configuration of M is : $[Ar]^{18}\,4s^1\,3d^5$ (and not $4s^2\,3d^4$)

So total number of electrons = 24.

Hence, metal M is chromium (Cr).

2. Following are the valence shell electronic configurations of some elements.

(i) $3s^2\,3p^5$ (ii) $3d^{10}\,4s^2$ (iii) $2s^2\,3p^6\,4s^1$ (iv) $1s^2\,2s^2$

Find out the blocks to which they belong in the periodic table ?

Ans. (i) p-block (ii) d-block (iii) s-block (iv) s-block

Sol. The block of the elements depend on the type of sub-shell which receive the last electron. In case of (i) it enters in 3p-subshell, (ii) it enters 3d-subshell, (iii) it enters 4s-subshell and (iv) it enters 2s-subshell.

3. Find out the group of the element having the electronic configuration, $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^6\,4s^2$.

Ans. As last electron enters in d-subshell, therefore this belongs to d-block. For d-block element the group number is equal to the number of valence shell electrons + number of electrons in (n-1) d-subshell. So, group number = 6 + 2 = 8.

4. Ionisation energy and electron affinity of fluorine are respectively 17.42 and 3.45 eV. Calculate electronegativity of fluorine atom.

Sol. According to Mulliken's electronegativity $(\chi_M) = \dfrac{\text{Ionisation energy} + \text{Electron affinity}}{2}$

$$= \frac{17.42 + 3.45}{2} = 10.435$$

Therefore, electronegativity on Pauling's scale $(\chi_P) = \dfrac{10.435}{2.8} = 3.726$

Ans. $\chi_P = 3.726$

5. In Column-I, there are given electronic configurations of some elements. Match these with the correct metals given in Column-II :

	Column-I		**Column-II**
(1)	ns^2, np^5	(p)	Chromium
(2)	$(n-1)\, d^{10}$, ns^1	(q)	Copper
(3)	$(n-1)\, d^5$, ns^1	(r)	Krypton
(4)	$(n-1)\, d^{10}$, ns^2 , np^6	(s)	Bromine

Ans. (1) → (s) ; (2) → (q) ; (3) → (p) ; (4) → (r).

Sol. (1) ns^2np^5 is general valence shell electron configuration of halogens. So this configuration belongs to bromine.

(2) $(n-1)\, d^{1\text{-}10}\, ns^{1\text{-}2}$; This is electron configuration of d-block elements. As it contains $(n-1)\, d^{10}ns^1$ configuration it belongs to copper.

(3) $(n-1)\, d^{1\text{-}10}\, ns^{1\text{-}2}$; This is electron configuration of d-block elements. As it contains $(n-1)\, d^5\, ns^1$ configuration it belongs to chromium.

(4) Noble gases has valence shell electron configuration $ns^2\, np^6$, so it belongs to krypton.

6. Match the metals given in Column-II with their type given in Column-I :

	Column-I		**Column-II**
(1)	Metalloid	(p)	Sulphur
(2)	Radioactive	(q)	Gold
(3)	Transition metal	(r)	Arsenic
(4)	Chalcogen	(s)	Uranium

Ans. (1) → (r) ; (2) → (s) ; (3) → (q) ; (4) → (p)

Sol. (1) Arsenic is a metalloid because it behaves as metal (forming cation, As^{3+} -$AsCl_3$) as well as nonmetal (forming anion, As^{3-}-AsH_3).

(2) Uranium is a radioactive element.

(3) Those elements which in their neutral atoms or in most common oxidation state have partially filled d-orbitals are called as transition elements. Gold in its +3 oxidation state has electron configuration $[Xe]^{54}$, $5d^86s^0$.

(4) 16th group elements like oxygen and sulphur are ore forming elements and therefore are called as chalcogens.

7. Match the metals given in Column-II with their type given in Column-I :

	Column-I		Column-II
(1)	Representative element	(p)	Cerium
(2)	Lanthanide	(q)	Aluminium
(3)	Coinage metal	(r)	Thorium
(4)	Actinide	(s)	Gold

Ans. (1) → (q) ; (2) → (p) ; (3) → (s) ; (4) → (r)

Sol. (1) s-block and p-block elements are collectively called as representative elements. As in aluminium last electron enters in p-subshell ($[Ne]^{10}3s^{2}3p^{1}$).

(2) Lanthanide series follows lanthanum (atomic number 57) and starts from cerium (atomic number 58) to lutetium (atomic number 71), fourteen 4f- series elements.

(3) Group 11- transition elements copper, silver & gold are known as coinage metals (used for making the coins).

(4) Actinides series follows actinium (atomic number 89) and starts from thorium (atomic number 90) to lawrencium (atomic number 103), fourteen 5f- series elements.

8. Match the particulars given in Column-I with the process/metal / species given in Column-II.

	Column-I		Column-II
(1)	Isoelectronic species	(p)	$A^{+}(g) + \text{energy} \rightarrow A^{++}(g) + e^{-}(g)$
(2)	Half filled orbital	(q)	Ar, K^{+}, Ca^{++}
(3)	Second ionisation energy	(r)	Lutetium
(4)	Inner transition element	(s)	Antimony

Ans. (1) → (q) ; (2) → (s) ; (3) → (p) ; (4) → (r)

Sol. (1) Species having same number of electrons but different nuclear charge are called isoelectronic species. Ar, K^{+} & Ca^{++} have same number of electrons i.e. 18 but 18, 19 & 20 number of protons respectively.

(2) np^{3}, $(n-1)\,d^{5}$ and $(n-2)\,f^{7}$ represent half filled orbitals. Antimony has ($[Kr]^{36}4d^{10}5s^{2}5p^{3}$).

(3) The energy required to remove an electron from an univalent cation(g) is called second ionisation energy.

(4) 4f and 5f- series elements are called inner transition elements because they have three outer most shells incomplete.

9. Match the type of elements / characteristic of the elements listed in Column-I with the correct element listed in Column-II.

	Column-I		Column-II
(1)	Highest 1st ionisation energy	(p)	Technitium
(2)	Highest electronegativity	(q)	Lithium
(3)	Synthetic element	(r)	Helium
(4)	Strongest reducing agent	(s)	Fluorine

Ans. (1) → (r) ; (2) → (s) ; (3) → (p) ; (4) → (q).

Sol. (1) Helium has highest 1st ionisation energy amongst all the elements of periodic table because of ns^{2} valence electron configuration and its small size of atom.

(2) Fluorine has highest electronegativity i.e. 4.0 on Pauling scale on account of its small size.

(3) Technitium is a man made element.

(4) Lithium is a strongest reducing agent because of its highest negative value of E^{o} due to its higher hydration energy on account of its small size of atom.

Extended or Long Form of the Periodic Ta-

☐ Metals

🟩 Non metals

■

s-Block Elements

d-Block Elements

p-Block Elements

Group →
Period ↓

Period	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
	1 H 1.0079																	2 He 4.0026
	3 Li 6.940	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
	11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.066	17 Cl 35.453	18 Ar 39.948
	19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.941	24 Cr 51.996	25 Mn 54.938	26 Fe 55.847	27 Co 58.933	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 62.723	32 Ge 72.61	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.80
	37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.94	43 Tc 98	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
	55 Cs 132.91	56 Ba 137.33	57 La* 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po 210	85 At 210	86 Rn 222
	87 Fr 223	88 Ra 226	89 Ac** 227	104 Unq 261	105 Unp 262	106 Unh 266	107 Uns 264	108 Uno 269	109 Une 268	110 Uun 269	111 Uuu 272	112 Uub 277						

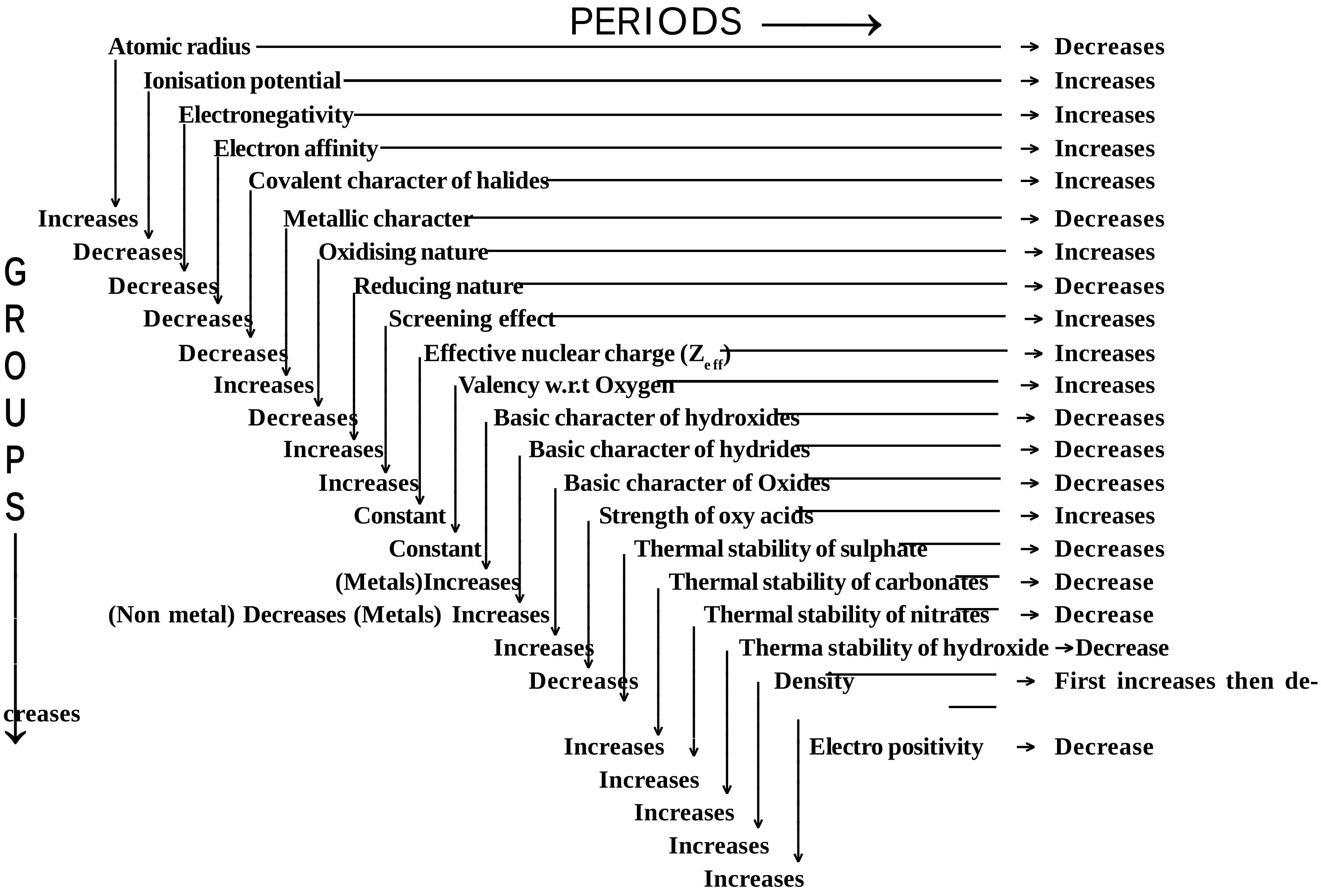
The symbols for elements 104-109 used in this table are those proposed by the American Chemical Society and

f-Block Elements

	*Lanthanide Series	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm 145	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50 Dyspro-	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.07	71 Lu 174.97
7	**Actinide Series	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237	94 Pu 244	95 Am 243	96 Cm 247	97 Bk 247	98 Cf 251	99 Es 252	100 Fm 257	101 Md 258	102 No 259	103 Lr 262

IUPAC designations of groups of elements are given in brackets

GENERAL TRENDS OF DIFFERENT PROPERTIES IN THE PERIOD AND GROUPS



PERIODS ⟶

GROUPS ↓

Atomic radius → Decreases
Ionisation potential → Increases
Electronegativity → Increases
Electron affinity → Increases
Covalent character of halides → Increases
Increases Metallic character → Decreases
Decreases Oxidising nature → Increases
Decreases Reducing nature → Decreases
Decreases Screening effect → Increases
Decreases Effective nuclear charge (Z_{eff}) → Increases
Increases Valency w.r.t Oxygen → Increases
Decreases Basic character of hydroxides → Decreases
Increases Basic character of hydrides → Decreases
Increases Basic character of Oxides → Decreases
Constant Strength of oxy acids → Increases
Constant Thermal stability of sulphate → Decreases
(Metals)Increases Thermal stability of carbonates → Decrease
(Non metal) Decreases (Metals) Increases Thermal stability of nitrates → Decrease
Increases Therma stability of hydroxide → Decrease
Decreases Density → First increases then decreases
Increases Electro positivity → Decrease
Increases
Increases
Increases
Increases

Exercise-1

✍ Marked Questions may have for Revision Questions.

PART - I : SUBJECTIVE QUESTIONS

Section (A) : Shielding Effect & Z_{eff}

A-1. Tell the relation between effective nuclear charge (Z_{eff}), atomic number (Z) and shielding constant (σ). Explain it qualitatively.

A-2.✍ Which orbital electrons are known to shield the nuclear charge improperly ? Does this generate some irregularity in properties of elements ?

Section (B) : Atomic and Ionic radius

B-1.✍ Explain why cations are smaller and anions larger in radii than their parent atoms ?

B-2. Arrange the following ions in the increasing order of their size : Be^{2+}, Cl^-, S^{2-}, Na^+, Mg^{2+}, Br^- ?

B-3.✍ The atomic radii of palladium and platinum are nearly same. Why ?

B-4. In the ionic compound KF, the K^+ and F^- ions are found to have practically identical radii, about 1.34 Å each. What can you predict about the relative atomic radii of K & F ?

Section (C) : Ionisation energy

C-1. Why second ionization enthalpy is always higher than the first ionisation enthalpy for every element ?

C-2.✍ The first ionization enthalpy of carbon is greater than that of boron, whereas the reverse is true for second ionization enthalpy. Explain.

C-3. Among the elements B, Al, C and Si, (i) which element has the highest first ionisation enthalpy ?
(ii) which element has the most metallic character ?
Justify your answer in each case.

C-4. The (IE_1) and the (IE_2) in kJ mol^{-1} of a few elements designated by Roman numerals are shown below:

	I	II	III
IE_1	403	549	1142
IE_2	2640	1060	2080

Which of the above elements is likely to be a
(a) non-metal (b) alkali metal (c) alkaline earth metal ?

Section (D) : Electron gain enthalpy

D-1. Why the electron gain enthalpy values of alkaline earth metals are lower (i.e. less negative) or positive?

D-2. Be and Ne have positive values of electron gain enthalpy against the general trend in their period in Modern periodic table. Explain.

D-3.✍ Nitrogen has positive electron gain enthalpy whereas oxygen has negative. However, oxygen has lower ionisation enthalpy than nitrogen. Explain.

D-4. $A^-(g) \rightarrow A^{2+}(g)$ $\Delta H = 1100$ KJ/mol
$A(g) \rightarrow A^{2+}(g)$ $\Delta H = 1200$ KJ/mol
Electron gain enthalphy of A is P × 10^2 KJ/mol. What is the value of P ?

Section (E) : Electronegativity

E-1. Among alkali metals, which element do you expect to be least electronegative ?

E-2.✍ Explain the following according to Modern periodic table :
(a) Electronegativity of elements increase on moving from left to right in a period.
(b) Ionisation enthalpy decrease in a group from top to bottom.

Section (F) : Oxidation states & Inert pair effect

F-1.✍ Pb^{4+} compounds are very good oxidising agents. Explain.

F-2. Arrange the following in correct order of stability :
(i) Ga^+, In^+, Tl^+ (ii) As^{+5} , Sb^{+5}, Bi^{+5}

PART - II : ONLY ONE OPTION CORRECT TYPE

Section (A) : Shielding Effect & Z_{eff}

A-1. The order of screening effect of electrons of s, p, d and f orbitals of a given shell of an atom on its outer shell electrons is :
(A) s > p > d > f (B) f > d > p > s (C) p < d < s > f (D) f > p > s > d

A-2. Which of the following is generally true regarding effective nuclear charge (Z_{eff}) :
(A) It increases on moving left to right in a period.
(B) It remains almost constant on moving top to bottom in a group.
(C) For isoelectronic species, as Z increases, Z_{eff} decreases.
(D) Both (A) and (B).

A-3.✍ Among following species which of them have maximum Z_{eff}
(A) Sn (B) Sn^{4+} (C) In (D) In^+

A-4. From the given set of species, point out the species from each set having highest Z_{eff}
(a) O^{2-}, F^-, Na^+ (b) Li, Be, Na (c) He, Li^+, H^-

	a	b	c		a	b	c
(A)	Na^+	Be	Li^+	(B)	O^{2-}	Li	H^-
(C)	F^-	Na	He	(D)	Na^+	Be	He

Section (B) : Atomic and Ionic radius

B-1. Which of the following order of radii is correct :
(A) Li < Be < Mg (B) $H^+ < Li^+ < H^-$ (C) O < F < Ne (D) Li < Na < K < Cs < Rb

B-2. The lanthanide contraction refers to :
(A) radius of the series. (B) valence electrons of the series.
(C) the density of the series. (D) electronegativity of the series.

B-3.✍ Select correct statement about radius of an atom :
(A) Values of Vander waal's radii are larger than those of covalent radii because the Vander waal's forces are much weaker than the forces operating between atoms in a covalently bonded molecule.
(B) The metallic radii are smaller than the Vander waal's radii, since the bonding forces in the metallic crystal lattice are much stronger than the Vander waal's forces.
(C) Both (A) & (B)
(D) None of these

B-4. Match the correct atomic radius with the element :

S.No.	Element	Code	Atomic radius (pm)
(i)	Be	(p)	74
(ii)	C	(q)	88
(iii)	O	(r)	111
(iv)	B	(s)	77
(v)	N	(t)	66

(A) (i) – r, (ii) – q, (iii) – t, (iv) – s, (v) – p
(B) (i) – t, (ii) – s, (iii) – r, (iv) – p, (v) – q
(C) (i) – r, (ii) – s, (iii) – t, (iv) – q, (v) – p
(D) (i) – t, (ii) – p, (iii) – r, (iv) – s, (v) – q

B-5. Choose the correct order of atomic radii of Fluorine and Neon (in pm) out of the options given below :
(A) 72, 160 (B) 160, 160 (C) 72, 72 (D) 160, 72

B-6. The size of isoelectronic species O^{-2} , F^- and Na^+ is affected by :
(A) nuclear charge (Z)
(B) valence principal quantum number (n)
(C) electron-electron interaction in the outer orbitals
(D) none of the factors because their size is the same.

Section (C) : Ionisation energy

C-1. Decreasing ionization potential for K, Ca & Ba is
(A) Ba> K > Ca (B) Ca > Ba > K (C) K > Ba > Ca (D) K > Ca > Ba

C-2. The correct order of second I.P.
(A) Na < Mg > Al < Si
(B) Na > Mg < Al > Si
(C) Na > Mg > Al < Si
(D) Na > Mg > Al > Si

C-3. The second ionization enthalpies of elements are always higher than their first ionization enthalpies because:
(A) cation formed always have stable half filled or completely filled valence shell electron configuration.
(B) it is easier to remove electron from cation.
(C) ionization is an endothermic process.
(D) the cation is smaller than its parent atom.

C-4.✍ A large difference between the third and fourth ionization energies indicates the presence of :
(A) 4 valence electrons in an atom
(B) 5 valence electrons in an atom
(C) 3 valence electrons in an atom
(D) 2 valence electrons in an atom

C-5.✍ Which one of the following statements is incorrect in relation to ionisation enthalpy ?
(A) Ionization enthalpy increases for each successive electron.
(B) The greatest increase in ionization enthalpy is experienced on removal of electron from core of noble gas configuration.
(C) End of valence electrons is marked by a big jump in ionization enthalpy.
(D) Removal of electron from orbitals bearing lower n value is easier than from orbitals having higher n value.

C-6. The first ionisation enthalpies (in eV) of N & O are respectively given by :
(A) 14.6, 13.6 (B) 13.6, 14.6 (C) 13.6, 13.6 (D) 14.6, 14.6

C-7. The first ionisation enthalpies of Na, Mg, Al and Si are in the order :
(A) Na < Mg > Al < Si (B) Na > Mg > Al > Si (C) Na < Mg < Al < Si (D) Na > Mg > Al < Si

C-8.✍ Which represents alkali metals (i.e. 1st group metals) based on $(IE)_1$ and $(IE)_2$ values (in kJ/mol) ?

		$(IE)_1$	$(IE)_2$			$(IE)_1$	$(IE)_2$
(A)	X	500	1000	(B)	Y	600	2000
(C)	Z	550	7500	(D)	M	700	1400

C-9.✍ Which of the following relation is correct with respect to first (I) and second (II) ionization enthalpies of potassium and calcium ?
(A) $I_{Ca} > II_{K}$ (B) $I_{K} > I_{Ca}$ (C) $II_{Ca} > II_{K}$ (D) $II_{K} > II_{Ca}$

Section (D) : Electron gain enthalpy

D-1. Which one of the following statement is correct ?
(A) The elements having large negative values of electron gain enthalpy generally act as strong oxidising agents.
(B) The elements having low values of ionisation enthalpies act as strong reducing agents.
(C) The formation of S^{2-}(g) from S(g) is an endothermic process.
(D) All of these.

D-2. The correct order of electron gain enthalpy (most endothermic first and most exothermic last) is :
(A) Be < B < C < N (B) Be < N < B < C (C) N < Be < C < B (D) N < C < B < Be

D-3. Among halogens, the correct order of amount of energy released in electron gain (electron gain enthalpy) is:
(A) F > Cl > Br > I (B) F < Cl < Br < I (C) F < Cl > Br > I (D) Cl > Br > F > I

D-4.✍ Which of the following will have the most negative electron gain enthalpy and which the least negative ? F, P, S, Cl.
(A) P, Cl (B) Cl, F (C) Cl, S (D) Cl, P

D-5.✍ The order of electron gain enthalpy (magnitude) of O, S and Se is :
(A) O > S > Se (B) S > Se > O (C) Se > S > O (D) S > O > Se

D-6. Electronic configurations of four elements A, B, C and D are given below :
(i) $1s^2 2s^2 2p^6$ (ii) $1s^2 2s^2 2p^4$ (iii) $1s^2 2s^2 2p^6 3s^1$ (iv) $1s^2 2s^2 2p^5$
Which of the following is the correct order of increasing tendency to gain electron :
(A) (i) < (iii) < (ii) < (iv) (B) (i) < (ii) < (iii) < (iv) (C) (iv) < (ii) < (iii) < (i) (D) (iv) < (i) < (ii) < (iii)

D-7. Which of the following statement is correct ?
(A) Electron gain enthalpy may be positive for some elements.
(B) Second electron gain enthalpy always remains positive for all the elements.
(C) $\Delta_{eg}H(K^+) = -$ IE (K)
(D) All of these

D-8. Highest electron affinity is shown by
(A) F^- (B) Cl^- (C) Li^+ (D) Na^+

Section (E) : Electronegativity

E-1. The correct set of decreasing order of electronegativity is :
(A) Li, H, Na (B) Na, H,Li (C) H, Li, Na (D) Li, Na, H

E-2. Which of the following is most electronegative in p-block elements
(A) Oxygen (B) Chlorine (C) Fluorine (D) Phosphorus

E-3. Which of the following is affected by the stable electron configuration of an atom ?
(a) Electronegativity (b) Ionisation enthalpy (c) Electron gain enthalpy
Correct answer is :
(A) only electronegativity (B) only ionisation enthalpy
(C) both electron gain enthalpy and ionisation enthalpy (D) all of the above

E-4. The electronegativity values of C,N,O and F on Pauling scale :
(A) decrease from carbon to fluorine.
(B) increase from carbon to fluorine.
(C) increase upto oxygen and then decrease upto fluorine.
(D) decrease from carbon to nitrogen and then increase continuously.

E-5. Correct order of electronegativity of N, P, C and Si on Pauling scale is :
(A) N > P > C > Si (B) C > Si > N > P (C) N < P < C < Si (D) N > C > P > Si

E-6. The correct order of electronegativity on Pauling scale is :
(A) F > Cl > O > S (B) Li > Na > K > Rb > Cs (C) Be < B < N < C (D) Both (A) and (B)

E-7. Which of the following is most electronegative element.
(A) Li (B) Mg (C) H (D) Na

Section (F) : Oxidation states & Inert pair effect

F-1. The atomic number of an element which can not show the oxidation state of +3 is-
(A) 13 (B) 32 (C) 33 (D) 17

F-2. The most common oxidation state of an element is –2. The number of electrons present in its outer most shell is -
(A) 2 (B) 4 (C) 6 (D) 8

F-3. Most stable oxidation state of gold is :
(A) + 1 (B) +3 (C) +2 (D) zero

F-4. Which can have both +ve and –ve oxidation states in their compounds
(A) F (B) I (C) Na (D) Al

F-5. The oxidation state of nitrogen varies from :
(A) –3 to + 5 (B) 0 to +5 (C) –3 to 1 (D) +3 to +5

F-6. Which metal exhibtis more than one oxidation states in their compounds
(A) Na (B) Mg (C) Al (D) Fe

F-7. Electrons of which subshell do not participate in bonding due to inert pair effect ?
(A) 6s (B) 6p (C) 5d (D) 4f

F-8. In which of the following elements, + 3 oxidation state is more stable than + 5 ?
(A) P (B) As (C) N (D) Bi

F-9. Which of the following is correct order of stability :
(A) $Tl^{3+} > Bi^{3+}$ (B) $PbO_2 > PbO$ (C) $BiI_5 < BiF_5$ (D) $Sn^{2+} = Ge^{2+}$

PART - III : MATCH THE COLUMNS

1. The Column-I has certain details about the elements of s-, p- and d-block elements. Match those with the group number of the elements listed in Column-II.

Column-I		Column-II	
(element / elements)		(group number)	
(A)	An element whose fourth shell contains two p-electrons	(p)	8th group
(B)	An element whose valence shell contains one unpaired p-electron	(q)	12th group
(C)	An element which receives last electron in (n – 1) d-subshell	(r)	14th group
(D)	An element with the ground-state electron configuration $[Ar]4s^2 3d^{10}$	(s)	17th group

2. Match the characteristics mentioned in column II with the process in column I.

	Column-I		Column-II
(A)	$O(g) + e^- \longrightarrow O^-(g)$	(p)	Positive Electron gain enthalpy
(B)	$O^-(g) + e^- \longrightarrow O^{2-}(g)$	(q)	Negative Electron gain enthalpy
(C)	$Na^-(g) \longrightarrow Na(g) + e^-$	(r)	Exothermic
(D)	$Mg^+(g) + e^- \longrightarrow Mg(g)$	(s)	Endothermic

Exercise-2

✍ **Marked Questions may have for Revision Questions.**

PART - I : ONLY ONE OPTION CORRECT TYPE

1. Which group of atoms have nearly same atomic radius :
(A) Na, K, Rb, Cs (B) Li, Be, B, C (C) Fe, Co, Mn (D) F, Cl, Br, I

2. The incorrect order of radius is :
(A) $Cu^- > Cu > Cu^+$ (B) $Sc^{3+} > K^+ > S^{2-}$ (C) Ni < Cu < Zn (D) All of these

3. Which of the following order of atomic / ionic radius is not correct ?
(A) F < Cl < Br < I (B) $Y^{3+} > Sr^{2+} > Rb^+$ (C) Nb $\approx$ Ta (D) Li > Be > B

4.✍ 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) $I^-, Se^{2-}, Br^-, F^-, O^{2-}$
(C) $Se^{2-}, I^-, Br^-, F^-, O^{2-}$ (D) $I^-, Se^{2-}, Br^-, O^{2-}, F^-$

5. In which of the following compounds, manganese shows maximum radius ?
(A) MnO_2 (B) $KMnO_4$ (C) MnO (D) $K_3[Mn(CN)_6]$

6. Amongst the following, the incorrect statement is
(A) IE_1 (Al) < IE_1 (Mg) (B) IE_1 (Na) < IE_1(Mg)
(C) IE_2 (Mg) > IE_2 (Na) (D) IE_3 (Mg) > IE_3 (Al)

7.✍ Which of the following is the correct order of ionisation enthalpy ?
(1) $Be^+ > Be$ (2) $Be > Be^+$ (3) C > Be (4) B > Be
(A) 2, 3 (B) 3, 4 (C) 1, 3 (D) 1, 4

8.✍ Considering the elements B, Al, Mg, and K, the correct order of their metallic character is :
(A) B > Al > Mg > K (B) Al > Mg > B > K
(C) Mg > Al > K > B (D) K > Mg > Al > B

9.✍ Which of the following is the correct order of ionisation enthalpy ?
(A) $Te^{2-} < I^- < Cs^+ < Ba^{2+}$ (B) $I^- < Te^{2-} < Cs^+ < Ba^{2+}$
(C) $Te^{2-} < Cs^+ < I^- < Ba^{2+}$ (D) $Ba^{2+} < Cs^+ < I^- < Te^{2-}$

10. Which is true statement(s) ?
(A) Larger is the value of ionisation enthalpy, easier is the formation of cation.
(B) Larger is the value of electron gain enthalpy, easier is the formation of anion.
(C) Larger is the value of ionisation energy as well as electron affinity, smaller is the Mulliken electronegativity of atom.
(D) Larger is the Z_{eff}, larger is the size of atom.

11. 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 utilising 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 in isolated stage.

12.✍ Which one of the following arrangements represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species ?
(A) Cl < F < S < O (B) O < S < F < Cl (C) S < O < Cl < F (D) F < Cl < O < S

13.✍ Which of the following statement is incorrect ?

(A) The tendency to attract bonded pair of electron in case of hybrid orbitals follow the order : $sp > sp^2 > sp^3$

(B) Alkali metals generally have negative value of electron gain enthalpy.

(C) $Cs^+(g)$ releases more energy upon gain of an electron than $Cl(g)$.

(D) The electronegativity values for 2p-series elements is less than that for 3p-series elements on account of small size and high inter electronic repulsions.

14. Which of the following order is incorrect against the property indicated :

(A) Mg < Ar < Na (2nd I.E.) (B) Be < F < Cl $\left[|\Delta H_{EA}|\right]$

(C) Rb < Na < K > Ca (atomic radius) (D) P < S < N (electronegativity)

15. The electronic configuration of an element is $1s^2\,2s^2\,2p^6\,3s^2\,3p^4$. The atomic number and the group number of the element 'X' which is just below the above element in the periodic table are respectively.

(A) 24 & 6 (B) 24 & 15 (C) 34 & 16 (D) 34 & 8

16. Which of the following can show +7 oxidation state?

(A) Mn (B) F (C) In (D) N

17.✍ Which of following does not exist :

(A) TlI_3 (B) PbF_4 (C) Both (A) and (B) (D) None of these

18. Elements of which period show maximum inert pair effect :

(A) 3 (B) 4 (C) 5 (D) 6

PART - II : SINGLE AND DOUBLE VALUE INTEGER TYPE

1. If internuclear distance between A atoms in A_2 is 10Å and between B atoms in B_2 is 6Å, then calculate internuclear distance between A and B in Å. [Electronegativity difference between A and B has negligible value].

2.✍ Report atomic number of the element having largest size among the following :

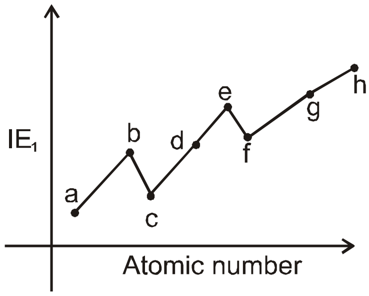
Ni, Cu, Zn

3. Total number of elements which have more ionization energy as compare to their next higher atomic number elements. Li, Be, C, N, O, F, Ne

4. Total number of elements which have less IE_1 than that of 'N'.

Be, B, C, F, P, He

5.



Where a, b, c, d, e, f, g, h are 3rd period elements. If difference between atomic number of elements b and e is x and difference between atomic number of elements c and f is y. What is the value of x – y.

6. For an element the successive ionisation energy values (in eV/atom), are given below.

14.534, 29.601, 47.448, 77.472, 97.888, 552.057, 667.029

Find the number of valence shell electrons in that element.

7. For the gaseous reaction $K + F \rightarrow K^+ + F^-$, ΔH was calculated to be 18.4 kcal/mol under conditions where the cations and anions were preverted from combining with each other. The ionisation enthalpy of K is 4.3 eV/atom. What is the electron gain enthalpy of F (in eV) ?
If your answer is x report it as –2x.

8.✍ The electron gain enthalpy of a hypothetical element 'A' is –3 eV per atom. How much energy in kCal is released when 10 g of 'A' are completely converted to A^- ions in gaseous state ?
(Take : 1 eV per atom = 23 kCal mol^{-1}, Molar mass of A = 30 g)

9. How many of the following elements are more electronegative than Boron.
H, Li, Be, C, N, O, F

10. What is atomic number of element which have maximum electron affinity in Modern Periodic table.

PART - III : ONE OR MORE THAN ONE OPTIONS CORRECT TYPE

1. Select correct order of size :
(A) $Ti^{2+} < Ti < Zr$ (B) $Ti^{2+} < Ti < Hf$ (C) $Zr^{2+} < Zr \approx Hf$ (D) $Hf^{2+} < Hf \approx Zr$

2. Which of the following orders of atomic / Ionic radius is correct ?
(A) $B < Al \approx Ga$ (B) $Sc > Cu < Zn$ (C) $C < O < N$ (D) $Al^{+3} < Al^{+2} < Al^{+}$

3. Which is/are the correct order/s of atomic radius ?
(A) Li < B < Be (C) Be < B < Li (C) Li > Be > B (D) N > O > F

4. Which is/are the correct order/s of atomic radius ?
(A) Mn > Fe > CO (B) $Mn \approx Fe \approx Co$ (C) Sc > Ti > V (D) Zn < Cu < Ni

5. Which of the following orders is(are) correct for size :
(A) $Al \approx Ga$ (B) $Te^{2-} > I^- > Cs^+ > Ba^{2+}$
(C) $Cr^{3+} < Cr^{6+}$ (D) $Pd \approx Pt$

6.✍ The ionic radii depends upon in the following factors :
The ionic radii depends upon in the following factors :
(A) Charge on cation
(B) Charge on anion
(C) Shell number of valence shell electron(s) of the ion.
(D) Effective nuclear charge

7. Which of the following statements is/are correct ?
(A) The second ionization enthalpy of oxygen element is greater than that of fluorine element.
(B) The third ionization enthalpy of phosphorus is greater than that of aluminium.
(C) The first ionization enthalpy of aluminium is slightly greater than that of gallium.
(D) The second ionization enthalpy of copper is greater than that of zinc.

8. Which of the following are correct ?
(A) $IE_2(Mg) < IE_2(Na)$
(B) EA (N) < EA (P)
(C) Atomic size Mg^{+2} > Atomic size (Li^+)
(D) IP of Na < Mg < Al

9. Which of the following elements will gain one electron more readily in comparison to other elements of their group ?
(A) S(g) (B) N(g) (C) O(g) (D) Cl (g)

10.✍ Which of the following is/are correct order/s of electron affinity.
(A) N < C < O < F (B) P < Si < S < Cl (C) Si < P < S < Cl (D) C < N < O < F

11. Order of Electron affinity of the elements or ions shown correctly?
(A) $S > O^-$ (B) $P > N^-$ (C) $O^- > S$ (D) $N^- > P$

12. Which of the following is correct order of electronegativity :
(A) Cs > Rb > Na (B) Li < Be < B (C) C < N < O (D) Cl > F > Br

13. Choose the correct statement(s) :
(A) In general more the ionisation energy more will be electronegativity.
(B) Electronegativity increase means metallic character increases.
(C) In general lower will be the ionisation energy, easier will be to remove electron.
(D) Electron affinity of S is less than that of Cl.

PART - IV : COMPREHENSION

Read the following passage carefully and answer the questions.

Comprehension # 1

It is not possible to measure the atomic radius precisely since the electron cloud surrounding the atom does not have a sharp boundary. One practical approach to estimate the size of an atom of a non-metallic element is to measure the distance between two atoms when they are bound together by a single bond in a covalent molecule and then dividing by two. For metals we define the term “metallic radius” which is taken as half the internuclear distance separating the metal cores in the metallic crystal. The van der waal’s radius represents the over all size of the atoms which includes its valence shell in a non bonded situation. It is the half of the distance between two similar atoms in separate molecules in a solid. The atomic radius decreases across a period and increases down the group. Same trends are observed in case of ionic radius. Ionic radius of the species having same number of electrons depends on the number of protons in their nuclei. Sometimes, atomic and ionic radii give unexpected trends due to poor shielding of nuclear charge by d- and f-orbital electrons.

Now answer the following three questions :

1.✍ Which of the following relations is correct, if considered for the same element :
(A) $r_{Vanderwaal} > r_{Covalent} > r_{Metallic}$ (B) $r_{Covalent} > r_{Metallic} > r_{Vanderwaal}$
(C) $r_{Vanderwaal} > r_{Metallic} > r_{Covalent}$ (D) $r_{Metallic} > r_{Covalent} > r_{Vanderwaa}$

2. K^+, Cl^- , Ca^{2+} , S^{2-} ions are isoelectronic. The decreasing order of their size is :
(A) $Ca^{2+} > K^+ > Cl^- > S^{2-}$ (B) $S^{2-} > Cl^- > K^+ > Ca^{2+}$
(C) $K^+ > Cl^- > Ca^{2+} > S^{2-}$ (D) $S^{2-} > Cl^- > Ca^{2+} > K^+$

3. Select the INCORRECT option regarding atomic/ionic sizes :
(A) Zn > Cu (B) $Pb^{2+} > Pb^{4+}$ (C) Zr $\approx$ Hf (D) $N^{3-} < Al^{3+}$

Paragraph for Questions 4 to 5

The IE_1 and the IE_2 in KJ/mol of a few elements designated by U, V, W, X are shown below.

Atom	IE_1	IE_2
U	2464	6110
V	610	7542
W	928	1810
X	1588	3410

Based on the above information answer the follwoing question :-

4. Which of the elements represent a noble gas.
(A) U (B) V (C) W (D) X

5. Which of the following element belongs to group 1 (IA).
(A) U (B) V (C) W (D) X

Comprehension # 3

The periodicity is related to the electronic configuration. That is, all chemical and physical properties are a manifestation of the electronic configuration of the elements.

The atomic and ionic radii generally decrease in a period from left to right. As a consequence, the ionization enthalpies generally increase and electron gain enthalpies become more negative across a period. In other words, the ionization enthalpy of the extreme left element in a period is the least and the electron gain enthalpy of the element on the extreme right is the highest negative. This results into high chemical reactivity at the two extremes and the lowest in the centre. Similarly down the group, the increase in atomic and ionic radii result in gradual decrease in ionization enthalpies and a regular decrease (with exception in some third period elements) in electron gain enthalpies in the case of main group elements.

The loss and gain of electrons can be co-related with the reducing and oxidising behaviour, and also with metallic and non-metallic character respectively, of the elements.

6. The correct order of the metallic character is :

(A) Al > Mg > Na > Si (B) Na > Mg < Al > Si (C) Na > Mg > Al > Si (D) Al > Mg > Si > Na

7.✍ Considering the elements B, C, N, F, and Si, the correct order of their non-metallic character is :

(A) B > C > Si > N > F (B) Si > C > B > N > F

(C) F > N > C > B > Si (D) F > N > C > Si > B

8. Which of the following statement is correct ?

(A) Ionisation enthalpies of elements decrease along a period and increase along a group in Modern periodic table.

(B) In the 3rd period of Modern periodic table, the two most reactive elements are sodium and fluorine.

(C) Fluorine has the least negative electron gain enthalpy among all halogens.

(D) Ionisation enthalpy of Pb is greater than that of Sn.

Paragraph for Questions 9 to 10

First electron gain enthalpy (in $\frac{\text{kJ}}{\text{mol}}$) of few elements are given below :

Elements	**ΔH_{eg}**
I	–60
II	–45
III	–328
IV	–295
V	+ 48

Answer the following questions on the basis of above data:

9. Which element may be an inert gas

(A) I (B) III (C) IV (D) V

10. Which element is most non-metallic among all the elements -

(A) I (B) II (C) III (D) IV

Exercise-3

* Marked Questions may have more than one correct option.

PART - I : JEE (ADVANCED) / IIT-JEE PROBLEMS (PREVIOUS YEARS)

1. The correct order of radii is : **[JEE-2000, 1/35]**
(A) $N < Be < B$ (B) $F^- < O^{2-} < N^{3-}$ (C) $Na < Li < K$ (D) $Fe^{3+} < Fe^{2+} < Fe^{+4}$

2. **Assertion :** The first ionization energy of Be is greater than that of B.
Reason : 2p orbital is lower in energy than 2s. **[JEE-2000, 1/35]**
(A) Both Assertion and Reason are true and Reason is the correct explanation of Assertion.
(B) Both Assertion and Reason are true but Reason is not correct explanation of Assertion.
(C) Assertion is true but Reason is false.
(D) Assertion is false but Reason is true.

3. The set representing the correct order of first ionization potential is : **[JEE-2001, 1/35]**
(A) $K > Na > Li$ (B) $Be > Mg > Ca$ (C) $B > C > N$ (D) $Ge > Si > C$

4. Identify the least stable ion amongst the following : **[JEE-2002, 3/90]**
(A) Li^- (B) Be^- (C) B^- (D) C^-

5. Among the following, the number of elements showing only one non-zero oxidation state is :
[JEE 2010, 3/163]
O, Cl, F, N, P, Sn, Tl, Na, Ti

6. The increasing order of atomic radii of the following group 13 elements is **[JEE 2016]**
(A) $Al < Ga < In < Tl$ (B) $Ga < Al < In < Tl$
(C) $Al < In < Ga < Tl$ (D) $Al < Ga < Tl < In$

7. The option(s) with only amphoteric oxides is (are): **[JEE 2017]**
(A) Cr_2O_3, CrO, SnO, PbO (B) NO, B_2O_3, PbO, SnO_2
(C) Cr_2O_3, BeO, SnO, SnO_2 (D) ZnO, Al_2O_3, PbO, PbO_2

PART - II : JEE (MAIN) / AIEEE PROBLEMS (PREVIOUS YEARS)

1. The increasing order of the first ionization enthalpies of the elements B, P, S and F (lowest first) is :
[AIEEE-2006, 4/220]
(1) $F < S < P < B$ (2) $P < S < B < F$ (3) $B < P < S < F$ (4) $B < S < P < F$

2. Lanthanoid contraction is caused due to : **[AIEEE-2006, 4/220]**
(1) the appreciable shielding on outer electrons by $4f$ electrons from the nuclear charge
(2) the appreciable shielding on outer electrons by $5f$ electrons from the nuclear charge
(3) the same effective nuclear charge from Ce to Lu
(4) the imperfect shielding on outer electrons by 4f electrons from the nuclear charge

3. The stability of dihalides of Si, Ge, Sn and Pb increases steadily in the sequence. **[AIEEE-2007, 3/120]**
(1) $SiX_2 << GeX_2 << SnX_2 << PbX_2$ (2) $PbX_2 << SnX_2 << GeX_2 << SiX_2$
(3) $GeX_2 << SiX_2 << SnX_2 << PbX_2$ (4) $SiX_2 << GeX_2 << PbX_2 << SnX_2$

4. The set representing the correct order of ionic radius is : **[AIEEE-2009, 4/144]**
(1) $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$ (2) $Li^+ > Na^+ > Mg^{2+} > Be^{2+}$
(3) $Mg^{2+} > Be^{2+} > Li^+ > Na^+$ (4) $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$

5. The correct sequence which shows decreasing order of the ionic radii of the elements is : **[AIEEE-2010, 4/144]**

(1) $Al^{3+} > Mg^{2+} > Na^{+} > F^{-} > O^{2-}$ (2) $Na^{+} > Mg^{2+} > Al^{3+} > O^{2-} > F^{-}$
(3) $Na^{+} > F^{-} > Mg^{2+} > O^{2-} > Al^{3+}$ (4) $O^{2-} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+}$

6. The outer electron configuration of Gd (Atomic No : 64) is : **[AIEEE 2011 (Cancelled), 4/120]**

(1) $4f^3\ 5d^5\ 6s^2$ (2) $4f^8\ 5d^0\ 6s^2$ (3) $4f^4\ 5d^4\ 6s^2$ (4) $4f^7\ 5d^1\ 6s^2$

7. The correct order of electron gain enthalpy with negative sign of F, Cl, Br and I, having atomic number 9, 17, 35 and 53 respectively, is: **[AIEEE 2011, 4/120]**

(1) F > Cl > Br > I (2) Cl > F > Br > I (3) Br > Cl > I > F (4) I > Br > Cl > F

8. The increasing order of the ionic radii of the given isoelectronic species is : **[AIEEE-2012, 4/144]**

(1) $Cl^{-}, Ca^{2+}, K^{+}, S^{2-}$ (2) $S^{2-}, Cl^{-}, Ca^{2+}, K^{+}$ (3) $Ca^{2+}, K^{+}, Cl^{-}, S^{2-}$ (4) $K^{+}, S^{2-}, Ca^{2+}, Cl^{-}$

9. Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar ? **[JEE Mains-2013, 4/120]**

(1) Ca < S < Ba < Se < Ar (2) S < Se < Ca < Ba < Ar
(3) Ba < Ca < Se < S < Ar (4) Ca < Ba < S < Se < Ar

10. The first ionisation potential of Na is 5.1 eV. The value of electron gain enthalpy of Na^{+} will be : **[JEE Mains-2013, 4/120]**

(1) –2.55 eV (2) –5.1 eV (3) –10.2 eV (4) +2.55 eV

11. Which of the following arrangements represents the increasing order (smallest to largest) of ionic radii of the given species $O^{2-}, S^{2-}, N^{3-}, P^{3-}$? **[JEE(Main) 2014 Online (15-04-14), 4/120]**

(1) $O^{2-} < N^{3-} < S^{2-} < P^{3-}$ (2) $O^{2-} < P^{3-} < N^{3-} < S^{2-}$
(3) $N^{3-} < O^{2-} < P^{3-} < S^{2-}$ (4) $N^{3-} < S^{2-} < O^{2-} < P^{3-}$

12. Which one of the following has largest ionic radius ? **[JEE(Main) 2014 Online (19-04-14), 4/120]**

(1) Li^{+} (2) O_2^{2-} (3) B^{3+} (4) F^{-}

13. The ionic radii (in Å) of N^{3-}, O^{2-} and F^{-} are respectively : **[JEE Mains-2015, 4/120]**

(1) 1.36, 1.40 and 1.71 (2) 1.36, 1.71 and 1.40
(3) 1.71, 1.40 and 1.36 (4) 1.71, 1.36 and 1.40

14. In the long form of the periodic table, the valence shell electronic configuration of $5s^2 5p^4$ corresponds to the element present in : **[JEE(Main) 2015 Online (10-04-15), 4/120]**

(1) Group 17 and period 6 (2) Group 17 and period 5
(3) Group 16 and period 6 (4) Group 16 and period 5

15. Which of the following atoms has the highest first ionization energy? **[JEE(Main)-2016, 4/120]**

(1) Na (2) K (3) Sc (4) Rb

16. The group having isoelectronic species is : **[JEE(Main)-2017, 4/120]**

(1) O^{-}, F^{-}, Na, Mg^{+} (2) $O^{2-}, F^{-}, Na, Mg^{2+}$
(3) $O^{-}, F^{-}, Na^{+}, Mg^{2+}$ (4) $O^{2-}, F^{-}, Na^{+}, Mg^{2+}$

17. Consider the following ionization enthalpies of two elements 'A' and 'B'

Element	Ionization enthalpy (kJ/mol)		
	1st	2nd	3rd
A	899	1757	14847
B	737	1450	7731

Which of the following statements is correct ? **[JEE(Main) 2017 Online (08-04-17), 4/120]**

(1) Both 'A' and 'B' belong to group-1 where 'B' comes below 'A'.

(2) Both 'A' and 'B' belong to group-2 where 'A' comes below 'B'.

(3) Both 'A' and 'B' belong to group-2 where 'B' comes below 'A'.

(4) Both 'A' and 'B' belong to group-1 where 'A' comes below 'B'.

18. The electronic configuration with the highest ionization enthalpy is : **[JEE(Main) 2017 Online (09-04-17), 4/120]**

(1) [Ne] $3s^2\ 3p^1$ (2) [Ne] $3s^2\ 3p^2$ (3) [Ne] $3s^2\ 3p^3$ (4) [Ar] $3d^{10}\ 4s^2\ 4p^3$

19. The correct order of electron affinity is :- **[JEE-MAIN 2018]**

(1) Cl > F > O (2) F > O > Cl (3) F > Cl > O (4) O > F > Cl

20. For Na^+, Mg^{2+}, F^- and O^{2-} the correct order of increasing ionic radii is : **[JEE(Main) 2019 Online (15-04-18), 4/120]**

(1) $O^{2-} < F^- < Na^+ < Mg^{2+}$ (2) $Na^+ < Mg^{2+} < F^- < O^{2-}$

(3) $Mg^{2+} < Na^+ < F^- < O^{2-}$ (4) $Mg^{2+} < O^{2-} < Na^+ < F^-$

21. In general, the properties that decrease and increase down a group in the periodic table, respectively, are : **[JEE(Main) 2019 Online (09-01-19), 4/120]**

(1) atomic radius and electronegativity

(2) electronegativity and atomic radius

(3) electron gain enthalpy and electronegativity

(4) electronegativity and electron gain enthalpy

22. The atomic radius of Ag is closest to : **[JEE(Main) Online (January set-1 2020]**

(1) Cu (2) Hg (3) Au (4) Ni

ANSWER KEY

EXERCISE # 1

PART - I

A-1. $Z_{eff} = Z - \sigma$

A-2. d- and f-orbital electrons are known for poor shielding of nuclear charge, because of their scattered structure. This poor shielding generates some irregularities in properties like atomic radii and ionisation enthalpy of d-block elements, f-block elements and group-13 elements.

B-1. The ionic radius of a cation is always smaller than the parent atom because the **loss of one or more electrons increases the effective nuclear charge (Z_{eff})**. As a result, the **force of attraction of nucleus for the remaining electrons increases and hence the electron cloud contracts** and ionic radii decreases. In contrast, the ionic radius of an anion is always larger than its parent atom because the **addition of one or more electrons decreases the effective nuclear charge (Z_{eff}).** As a result, the **force of attraction of the nucleus for the remaining electrons decreases and hence electron cloud expands** and the ionic radii increases.

B-2. $Be^{2+} < Mg^{2+} < Na^{+} < Cl^{-} < S^{2-} < Br^{-}$

B-3. Due to lanthanide contraction (poor shielding of nuclear charge by 4f-electrons), atomic radii of 4d and 5d elements are nearly same.

B-4. Atomic radius of K is larger than F because the size of cation is smaller than its parent atom while size of anion is bigger than its parent atom. Thus, atomic radii of K will be greater than 1.34 Å while atomic radii of F will be less than 1.34 Å.

C-1. Electron is more tightly bound by the nucleus in an cation (i.e. M^{+}) as the number of proton remains the same as in neutral atom whereas number of electron is one less than the proton. This increases the attraction between the valence shell electrons and the nucleus (Z_{eff} increases). So, second ionization enthalpy is always higher than the first ionisation enthalpy for every element.

C-2. Carbon has higher IE_1 because of smaller atomic size and greater Z_{eff}. Removal of second electron from stable $1s^2\ 2s^2$ configuration in case of B^{+} requires greater energy. So, B has greater IE_2.

C-3. (i) C (ii) Al

C-4. (a) non-metal(III) - Due to highest ionisation energy, (IE_1) and (IE_2).
(b) alkali metal(I) - Due to lowest ionisation energy, (IE_1) and there is quite high jump in (IE_2) due to inert gas configuration.
(c) alkaline earth metal (II) - There is little difference in (IE_1) and (IE_2) and the value of (IE_1) is slightly greater than(I) due to stable configuration(ns^2).

D-1. The general valence shell electron configuration of alkaline earth metals is ns^2 (stable configuration).
The extra electron must enter np subshell, which is effectively shielding by the two ns electrons and the inner electrons. Consequently, the alkaline earth metals have little or no tendency to pick up an extra electron.

D-2. In Be, the extra electron is to be added in 2p orbital because 2s orbital is completely filled and in Ne, it is to be added to a noble gas configuration. Since full-filled orbitals and noble gas configuration are more stable, reluctancy in accepting the electron is found. So, they have positive values of electron gain enthalpy.

D-3. Nitrogen has stable half filled configuration $2s^2\ 2p^3$. So removal of one electron will require more energy than oxygen. Similarly, in nitrogen, addition of one electron will require energy (endothermic) while in oxygen, addition of one electron will release energy (exothermic).

D-4. 1

E-1. Caesium (Cs).

E-2. (a) On moving left to right in a period, tendency of an atom to attract the shared electron pair towards itself increases due to increasing Z_{eff}. So, electronegativity of elements increase on moving from left to right in a period.
(b) On moving top to bottom in a group, size increases due to addition of extra shells. So, attraction of nucleus outermost electron decreases. So, ionisation enthalpy decrease in a group from top to bottom.

F-1. Pb^{4+} is less stable than Pb^{2+} due to inert pair effect. So, Pb^{4+} compounds are very good oxidising agents.

F-2. (i) $Ga^{+} < In^{+} < Tl^{+}$ (ii) $As^{+5} > Sb^{+5} > Bi^{+5}$

PART - II

A-1.	(A)	**A-2.**	(D)	**A-3.**	(B)	**A-4.**	(A)	**B-1.**	(B)	**B-2.**	(A)	**B-3.**	(C)
B-4.	(C)	**B-5.**	(A)	**B-6.**	(A)	**C-1.**	(B)	**C-2**.	(B)	**C-3.**	(D)	**C-4.**	(C)
C-5.	(D)	**C-6.**	(A)	**C-7.**	(A)	**C-8.**	(C)	**C-9.**	(D)	**D-1.**	(D)	**D-2.**	(B)
D-3.	(C)	**D-4.**	(D)	**D-5.**	(B)	**D-6.**	(A)	**D-7.**	(D)	**D-8**.	(C)	**E-1.**	(C)
E-2.	(C)	**E-3.**	(C)	**E-4.**	(B)	**E-5.**	(D)	**E-6.**	(B)	**E-7.**	(C)	**F-1.**	(B)
F-2.	(C)	**F-3.**	(D)	**F-4.**	(B)	**F-5.**	(A)	**F-6.**	(D)	**F-7.**	(A)	**F-8.**	(D)
F-9.	(C)												

PART - III

1. (A) → (r) ; (B) → (s) ; (C) → (p, q) ; (D) → (q).

2. (A - q,r) ; (B - p,s) ; (C - s) ; (D - q,r)

EXERCISE # 2

PART - I

1.	(C)	**2.**	(B)	**3.**	(B)	**4.**	(D)	**5.**	(C)	**6.**	(C)	**7.**	(C)
8.	(D)	**9.**	(A)	**10.**	(B)	**11.**	(D)	**12.**	(B)	**13.**	(D)	**14.**	(C)
15.	(C)	**16.**	(A)	**17.**	(D)	**18.**	(D)						

PART - II

1.	8	**2.**	30	**3.**	3	**4.**	4	**5.**	0	**6.**	5	**7.**	7
8.	**23**	**9.**	5	**10.**	17								

PART - III

1.	(ABCD)	**2.**	(ABD)	**3.**	(CD)	**4.**	(BC)	**5.**	(ABD)	**6.**	(ABCD)
7.	(ABD)	**8.**	(AB)	**9.**	(AD)	**10.**	(AB)	**11.**	(AB)	**12.**	(BC)
13.	(ACD)										

PART - IV

1.	(C)	**2.**	(B)	**3.**	(D)	**4.**	(A)	**5.**	(B)	**6.**	(C)	**7.**	(C)
8.	(D)	**9.**	(D)	**10.**	(C)								

EXERCISE # 3

PART - I

1.	(B)	**2.**	(C)	**3.**	(B)	**4.**	(B)	**5.**	**2**	**6.**	(B)	**7.**	(CD)

PART - II

1.	(4)	**2.**	(4)	**3.**	(1)	**4.**	(1)	**5.**	(4)	**6.**	(4)	**7.**	(2)
8.	(3)	**9.**	(3)	**10.**	(2)	**11.**	(1)	**12.**	(2)	**13.**	(3)	**14.**	(4)
15.	(3)	**16.**	(4)	**17.**	(3)	**18.**	(3)	**19.**	(1)	**20.**	(3)	**21.**	(2)
22.	(3)												

✍ Marked Questions may have for Revision Questions.

This Section is not meant for classroom discussion. It is being given to promote self-study and self testing amongst the Reliable students.

Self Assessment Test

PART- 1 : PAPER JEE (MAIN) PATTERN

SECTION–I : (Maximum Marks : 80)

- This section contains **TWENTY** questions.
- Each question has **FOUR** options (A), (B), (C) and (D). **ONLY ONE** of these four options is correct.
- For each question, darken the bubble corresponding to the correct option in the ORS.
- For each question, marks will be awarded in one of the following categories :
  Full Marks : +4 If only the bubble corresponding to the correct option is darkened.
  Zero Marks : 0 If none of the bubbles is darkened.
  Negative Marks : –1 In all other cases

1. The size of the following species increases in the order:
(A) $Mg^{2+} < Na^{+} < F^{-}$ (B) $F^{-} < Na^{+} < Mg^{2+}$
(C) $Mg^{2+} < F^{-} < Na^{+}$ (D) $Na^{+} < F^{-} < Mg^{2+}$

2. Highest size will be of
(A) Br^{-} (B) I (C) I^{-} (D) I^{+}

3. Of the following pairs, the one containing examples of metalloid elements is :
(A) B and Al (B) Ga and Ge (C) Al and Si (D) As and Sb

4. Which of the following is the wrong statement ?
(A) All the actinide elements are radioactive.
(B) Alkali and alkaline earth metals are s-block elements.
(C) Pnicogens and halogens are p-block elements.
(D) The first member of the lanthanide series is lanthanum.

5. Atomic number of 15, 33, 51 represents the following family :
(A) carbon family (B) nitrogen family (C) oxygen family (D) None of these

6. Which of the following is correct order of Z_{eff} :
(A) $I^{-} > I > I^{+}$ (B) $Mg^{2+} > Na^{+} > F^{-}$ (C) $P^{5+} < P^{3+}$ (D) Li > Be > B

7.✍ What is correct order of reducing capacity :
(A) $Ge^{2+} > Sn^{2+} > Pb^{2+}$ (B) $Ge^{2+} < Sn^{2+} < Pb^{2+}$ (C) $Ge^{2+} \approx Sn^{2+} \approx Pb^{2+}$ (D) $Pb^{2+} > Ge^{2+} > Sn^{2+}$

8. The correct order of increasing atomic size of element N,F, Si & P.
(A) N < F < Si < P (B) F > N < P < Si (C) F < N < P < Si (D) F < N < Si < P

9. The correct order of atomic or ionic size

(A) N < Li < B (B) Cl < Mg < Ca (C) $Ca^{+2} < S^{-2} < Cl^{-}$ (D) $Na^{+} < Mg^{+2} < Cl^{-}$

10. The correct order of second ionisation potential of C, N, O and F is:

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

11. The ionization energy will be maximum for which process?

(A) $Ba \rightarrow Ba^{+}$ (B) $Be \rightarrow Be^{+}$ (C) $Cs \rightarrow Cs^{+}$ (D) $Li \rightarrow Li^{+}$

12. The atomic number of Vanadium (V), Chromium (Cr), Manganese (Mn) and Iron (Fe) are respectively 23, 24, 25 and 26 which one of these may be expected to have the highest second Ionization enthalpy.

(A) V (B) Cr (C) Mn (D) Fe

13.✍ With reference to 1st IP which are correct.

(a) Li < C (b) O < N (c) Be < N < Ne

(A) a, b (B) b, c (C) a, c (D) a, b & c

14. Values of 1st four ionisaiton energies (kJ/mol) of an element are respectively 496, 4563, 6913, 9541 ; the electronic configuration of that element can be.

(A) $1s^2, 2s^1$ (B) $1s^2\ 2s^2\ 2p^1$ (C) $1s^2, 2s^2, 2p^6\ 3s^1$ (D) (B) and (C) both

15. Of the following elements, which possesses the highest electron affinity?

(A) As (B) O (C) S (D) Se

16. For magnitude of electron gain enthalpy of chalcogens and halogens, which of the following options is correct?

(A) Br > F (B) S > F (C) O < Cl (D) S < Se

17. The process requires absorption of energy is

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

18. $\frac{N_0}{2}$ atoms of X (g) are converted into X^{+} (g) by absorbing E_1 energy. $2N_0$ atoms of X (g) are converted into X^{-}(g) by releasing E_2 energy. Calculate ionisation enthalpy and electron gain enthalpy of X(g) per atom.

(A) I.E. = $\frac{2E_1}{N_0}$, $\Delta_{eq}H = -\frac{E_2}{2N_0}$ (B) I.E. = $-\frac{E_2}{2N_0}$, $\Delta_{eq}H = \frac{2E_1}{N_0}$

(C) I.E. = $\frac{E_1}{2N_0}$, $\Delta_{eq}H = -\frac{E_2}{2N_0}$ (D) I.E. = $\frac{N_0}{2E_1}$, $\Delta_{eq}H = -\frac{2N_0}{E_2}$

19. The properties which are not common to both groups 1 and 17 elements in the periodic table are :

(A) Elelctropositive character increase down the gorups.

(B) Reactivity decrease from top to bottom in these groups.

(C) Atomic radii increase as the atomic number increase.

(D) Electronegativity decrease on moving down a group.

20. For which of the following species 2nd IE < 1st IE

(A) Be (B) Ne (C) Na^{+} (D) None of these

SECTION-II : (Maximum Marks: 20)

- This section contains **FIVE** questions.
- The answer to each question is a **NUMERICAL VALUE**.
- For each question, enter the correct numerical value (If the numerical value has more than two decimal places, **truncate/round-off** the value to **TWO** decimal places; e.g. 6.25, 7.00, –0.33, –.30, 30.27, –127.30, if answer is 11.36777..... then both 11.36 and 11.37 will be correct) by darken the corresponding bubbles in the ORS.

  For Example : If answer is –77.25, 5.2 then fill the bubbles as follows.
- Answer to each question will be evaluated according to the following marking scheme:

  Full Marks : +4 If ONLY the correct numerical value is entered as answer.

21.✍ Identify the group (in Modern Periodic Table) and valency of a hypothetical element having atomic number 119. If group number is x and valency is y. Give the the value of x + y.

22.^✍ An element belonging to 3d series of modern periodic table has spin magnetic moment = 5.92 B.M. in +3 oxidation state. Determine the atomic number of element.

23. An element has atomic number 29. It belongs to x period and y group. Give value of 2x + y :

24.✍ How many of the following compounds are found to exist?

BiF_5, $T\ell I_3$, PbO_2, $SnCl_2$, $T\ell_2O_3$, PbI_4, As_2O_3

25. The number of unpaired electrons in Ni^{2+} is

PART 2 : PAPER JEE (ADVANCED) PATTERN

SECTION-I : (Maximum Marks : 12)

- This section contains **FOUR** questions.
- Each question has **FOUR** options (A), (B), (C) and (D). **ONLY ONE** of these four options is correct.
- For each question, darken the bubble corresponding to the correct option in the ORS.
- For each question, marks will be awarded in one of the following categories :

Full Marks	:	+3	If only the bubble corresponding to the correct option is darkened.
Zero Marks	:	0	If none of the bubbles is darkened.
Negative Marks	:	–1	In all other cases

1. Which set does not shows correct matching according to Modern periodic table :

(A) Cr = [Ar] $3d^5 4s^1$; element belongs to 6th group.

(B) Fe^{2+} = [Ar] $3d^6$; element belongs to 8th group.

(C) Sc^{3+} = [Ne] $3s^2 3p^6$; element belongs to zero/eighteen group.

(D) All of the above.

2. In which element shielding effect is not possible ?

(A) H (B) Be (C) B (D) N

3. Which of following ions do not exist together in aqueous solution :

(A) Pb^{2+}, F^- (B) Tl^{3+}, I^- (C) Both (A) and (B) (D) None of these

4. Select correct statement(s) :

(A) Across a transition series (from Cr to Cu), there is only a small change in atomic radius from one element to another due to very small change in effective nuclear charge.

(B) The rate of decrease in the size across the lanthanide series is less than the across the first transition series.

(C) Both are correct statements.

(D) None of the statement is correct.

SECTION-II : (Maximum Marks: 32)

- This section contains **EIGHT** questions.
- Each question has **FOUR** options for correct answer(s). **ONE OR MORE THAN ONE** of these four option(s) is (are) correct option(s).
- For each question, choose the correct option(s) to answer the question.
- Answer to each question will be evaluated according to the following marking scheme:

Full Marks : +4 If only (all) the correct option(s) is (are) chosen.
Partial Marks : +3 If all the four options are correct but ONLY three options are chosen.
Partial Marks : +2 If three or more options are correct but ONLY two options are chosen, both of which are correct options.
Partial Marks : +1 If two or more options are correct but ONLY one option is chosen and it is a correct option.
Zero Marks : 0 If none of the options is chosen (i.e. the question is unanswered).
Negative Marks : –1 In all other cases.

- **For Example :** If first, third and fourth are the **ONLY** three correct options for a question with second option being an incorrect option; selecting only all the three correct options will result in +4 marks. Selecting only two of the three correct options (e.g. the first and fourth options), without selecting any incorrect option (second option in this case), will result in +2 marks. Selecting only one of the three correct options (either first or third or fourth option), without selecting any incorrect option (second option in this case), will result in +1 marks. Selecting any incorrect option(s) (second option in this case), with or without selection of any correct option(s) will result in –1 marks.

5.✍ Which of the following statement is correct for the d-block elements :

(A) They have general electronic configuration $(n-1)d^{1-10}\ ns^{0-2}$.

(B) They generally exhibit variable valency.

(C) Last electron enters in (n – 1)d sub-shell in them.

(D) They are placed from 3rd to 6th period in modern periodic table.

6. Poor shielding of nuclear charge by d or f- orbital electrons is responsible for which of the following facts ?

(A) Atomic radius of Nb (4d-series) is comparable to that of Ta (5d-series)

(B) The Ist ionisation enthalpy of copper is less than that of zinc

(C) The value of electron gain enthalpy is more negative for sulphur than for oxygen.

(D) The Ist ionisation energy for gold is greater than that of silver.

7. Which of the following element(s) have only one non-zero oxidation state.

(A) Be (B) O (C) F (D) N

8.✍ Which of the following is/are true order(s) ?

(A) $B^+ < B < B^-$ Size

(B) $I < Br < Cl < F$ Electron gain enthalpy

(C) $O^{--} < O^- < O^+$ Z_{eff}

(D) $Na < Al < Mg < Si$ Ionisation potential

9. Select the endothermic step(s) :

(A) $S^-(g) + e^- \longrightarrow S^{2-}(g)$ (B) $Ne(g) + e^- \longrightarrow Ne^-(g)$

(C) $N(g) + e^- \longrightarrow N^-(g)$ (D) $Al^{2+}(g) \longrightarrow Al^{3+}(g) + e^-$

10. Which of the following has/have no unit ?

(A) Electronegativity (B) Electron gain enthalpy

(C) Ionisation enthalpy (D) Metallic character

11. Those elements impart colour to the flame on heating in it, the atoms of which require low energy for the ionisation (i.e. absorb energy in the visible region of spectrum). The elements of which of the following groups in Modern periodic table will impart colour to the flame ?

(A) 2 (B) 13 (C) 1 (D) 17

12. Amongst the following statements, which is / are correct?

(A) Electronegativity of sulphur is greater than that of oxygen.

(B) Electron affinity of oxygen is smaller than that of sulphur.

(C) Electron gain enthalpy of fluorine is most negative

(D) Electron gain enthalpy of chlorine is most negative

SECTION-III : (Maximum Marks: 18)

- This section contains **SIX** questions.
- The answer to each question is a **NUMERICAL VALUE**.
- For each question, enter the correct numerical value (in decimal notation, truncated/rounded-off to the **second decimal place**; e.g. 6.25, 7.00, –0.33, –.30, 30.27, –127.30, if answer is 11.36777..... then both 11.36 and 11.37 will be correct) by darken the corresponding bubbles in the ORS.

  For Example : If answer is –77.25, 5.2 then fill the bubbles as follows.
- Answer to each question will be evaluated according to the following marking scheme:

  Full Marks : +3 If ONLY the correct numerical value is entered as answer.

  Zero Marks : 0 In all other cases.

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13. Atomic number of Ag is 47. In the same group the atomic numbers of elements placed above and below Ag in long form of periodic table will be x and y respectively. Give the value of (x + y)/12.

14. What is oxidation states of hydrogen in CaH_2 & CH_4.

15. Most stable oxidation state of Thallium is +n. What is the Value of n.

16.✍ How many elements are more electropositive than Cl.

B, N, O, C, S, P, At, H, Li

17. How many of following atoms have maximum ionization energy than boron.

Be, N , P, Ga, S, Mg

18. Values of IE_1, IE_2, IE_3 of an element are 9.3, 18.2 and 553.8 eV. Predict group number in Modern Periodic Table.

PART - 3 : OLYMPIAD (PREVIOUS YEARS)

1. The element whose electronic configuration is $1s^2, 2s^2\ 2p^6\ 3s^2$ is a/an **[NSEC-2000]**

(A) metal (B) inert gas (C) metalloid (D) non-metal

2. Oxygen shows +2 oxidation state in **[NSEC-2000]**

(A) F_2O (B) H_2O (C) K_2O (D) D_2O

3.✍ Which Group IIIA element is expected to have physical and chemical properties least similar to other members of that group ? **[NSEC-2000]**
(A) Ga (B) Al (C) B (D) In

4. Which of the following ions will show highest magnetic moment (Z values for neutral atoms are as follows: N = 7, Cr = 24, Fe = 26 & Co = 27) **[NSEC-2000]**
(A) Fe^{3+} (B) Cr^{3+} (C) N^{3+} (D) Co^{3+}

5. The oxidation state of Cr in $K_2Cr_2O_7$ is : **[NSEC-2000]**
(A) + 3 (B) + 6 (C) + 4 (D) – 4

6. Which of the following is the smallest in size ? **[NSEC-2001]**
(A) N^{3-} (B) F^- (C) O^{2-} (D) Na^+

7. Oxidation Number of Mn in $[MnO_4]^-$ is : **[NSEC-2001]**
(A) –7 (B) + 7 (C) + 2 (D) – 2

8. The formation of anion from a neutral atom X is favoured by : **[NSEC-2001]**
(A) high electron affinity (B) large size of X
(C) low ionisation potential (D) high charge on anion X

9.✍ Which element of 3rd row has biggest atomic size ? **[NSEC-2002]**
(A) chlorine (B) sodium (C) silicon (D) neon.

10. The atom of an element X contains 27 electrons. X is expected to be **[NSEC-2003]**
(A) a non-metal belonginig to p-block (B) paramagnetic belonging to d-block
(C) diamagnetic belonging to d-block (D) an s-block element.

11. The group in the periodic table that contains the elements in all the different physical states at room temperature is **[NSEC-2004]**
(A) V A (B) I A (C) VII A (D) IV A.

12. The ion having a noble gas electronic configuration is **[NSEC-2004]**
(A) Se^{2-} (B) Fe^{3+} (C) Cr^{3+} (D) Cu^+.

13. Element with Z = 83 belongs to which block? **[NSEC-2005]**
(A) s (B) p (C) d (D) f.

14. Inert pair effect plays an important role in the case of **[NSEC-2005]**
(A) P (B) Bi (C) Sb (D) As

15. The element having electronegativity next to that of fluorine is **[NSEC-2005]**
(A) oxygen (B) chlorine (C) iodine (D) sodium.

16. The lithium ion (Li^+) and hydride ion (H^-) are isoelectronic ions. Which statement about these systems is true ? **[NSEC-2006]**
(A) Chemical properties of these ions are identical since they are isoelectronic.
(B) Li^+ is a stronger reducing agent than H^-
(C) More energy is needed to ionize H^- than Li^+
(D) Radius of H^- is larger than that of Li^+.

17.✍ Which of the following sequence of elements is arranged in the order of increasing atomic radii ? **[NSEC-2006]**
(A) Na, Mg, Al, Si (B) C,N,O,F (C) O,S,Se, Te (D) I,Br,Cl,F.

18. The number of unpaired electrons in the scandium atom is **[NSEC-2006]**

(A) 1 (B) 2 (C) 0 (D) 3.

19. Which of the following pairs represents isoelectronic ions? **[NSEC-2006]**

(A) Mn^{3+} and Fe^{2+} (B) Mn^{2+} and Fe^{3+} (C) Cr^{3+} and Mn^{2+} (D) Fe^{2+} and Co^{2+}.

20. Element having (4, 0, 0, + 1/2) as a set of four quantum numbers for its valence electron is- **[NSEC-2007]**

(A) Na (B) Ca (C) K (D) Br

21. Europium forms stable Eu^{2+} ions because **[NSEC-2007]**

(A) Europium is a lanthanide. (B) +2 is the common oxidation state of lanthanides.

(C) Eu^{2+} has f^7 valence electronic configuration. (D) All of the above reasons.

22. How many unpaired electrons are in a Fe^{2+} ion in the ground state ? **[NSEC-2008]**

(A) 0 (B) 2 (C) 4 (D) 6

23. For the atoms Li, Be, B and Na, the correct order of increasing atomic redius is : **[NSEC-2008]**

(A) B, Be, Li, Na (B) Li, Be, B, Na (C) Be, Li, B, Na (D) Be, B, Li, Na

24. The noble gas was first time discovered by **[NSEC-2008]**

(A) Cavendish (B) Willian Ramsay (C) Rayleigh (D) Frankland

25. The correct order of increasing first ionization energy is **[NSEC-2010]**

(A) Ca < K < Ne < P < F (B) F < Ca < Ne < P < K

(C) K < Ca < P < F < Ne (D) Ne < F < P < Ca < K

26. The group that has the species correctly listed in the order of decreasing radius is : **[NSEC-2011]**

(A) Cu^{2+}, Cu^{+}, Cu (B) V, V^{2+}, V^{3+} (C) F^{-}, Br^{-}, I (D) B, Be, Li

27. The number of valence electrons in an atom with the configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$ is : **[NSEC-2011]**

(A) 6 (B) 5 (C) 4 (D) 2

28. The most abundant element in the earth's crust is : **[NSEC-2011]**

(A) aluminium (B) oxygen (C) silicon (D) iron

29. Einsteinium has 11 electrons in the 4f subshell. The number of unpaired electrons in the subshell is : **[NSEC-2011]**

(A) 3 (B) 4 (C) 7 (D) 11

30. Lanthanide contraction is caused due to **[NSEC-2012]**

(A) the appreciable shielding on outer electrons by 4f electrons from the nuclear charge

(B) the appreciable shielding on outer electrons by 5d electrons from the nuclear charge

(C) the same effective nuclear charge from Ce to Lu.

(D) the imperfect shielding on outer electrons by 4f electrons from the nuclear charge

31. Europium (Eu) and Terbium (Tb) attain stable $4f^7$ configuration by exhibiting oxidation states of **[NSEC-2013]**

(A) +2 and +4 (B) +3 and +4 (C) +2 and +3 (D) +1 and +3

32. 4s orbital has lesser energy than 3d orbital because it has **[NSEC-2013]**

(A) Greater value of n (B) Lesser value of ℓ

(C) Lesser value of $(n + \ell)$ (D) $\ell = 0$

33. In the compound $Na_2S_2O_3$, the oxidation state of sulphur is : **[NSEC-2013]**

(A) –2 (B) +2 (C) +4 (D) +6

34. Of the following, the ion with the largest size is **[NSEC-2014]**

(A) O^{2-} (B) Na^{+} (C) F^{-} (D) Al^{3+}

PART - 4 : ADDITIONAL PROBLEMS

Paragraph For Questions 1 to 3

EA_1 value of some group of p-Block elements are given :

	At no. increase →		
At no. increases ↓	– 8(a)	141(e)	328(i)
	72 (b)	200(f)	349(j)
	78 (c)	195(g)	325(k)
	103 (d)	190(h)	295(ℓ)

a, b, c...........ℓ are non radioactive p-Block elements :

1.✍ Select the correct order of atomic radius :
(A) $a < b < c < d$ (B) $a < e < i$ (C) $i > j > k > \ell$ (D) $e > f > g$

2. Select the correct order of 2^{nd} Ionisation energy :
(A) $a < e < i$ (B) $a < e < i$ (C) $e < a < i$ (D) $e > i > a$

3. Choose correct match :
(A) a, b, c, d = Pnictogens (B) e, f, g, h = Chalogens
(C) i, j, k, l = Halogens (D) All of these

4.✍ Match the electronic configurations of the elements given in **List-I** with their correct characteristic(s) (i.e. properties for given configuration) given in **List-II** and select the correct answer using the code given below the lists.

List-I			**List-II**
P.	$1s^2$	1.	Element shows highest negative oxidation state.
Q.	$1s^2\ 2s^2\ 2p^5$	2.	Element shows highest first ionisation enthalpy.
R.	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	3.	Element shows highest electronegativity on Pauling scale.
S.	$1s^2\ 2s^2\ 2p^2$	4.	Element shows maximum electron gain enthalpy (most exothermic).

Code :

	P	Q	R	S
(A)	4	2	3	1
(B)	2	3	4	1
(C)	2	1	3	4
(D)	1	2	3	4

RRP ANSWER KEY

PART - 1

1.	(A)	**2.**	(C)	**3.**	(D)	**4.**	(D)	**5.**	(B)	**6.**	(B)	**7.**	(A)
8.	(C)	**9.**	(B)	**10.**	(C)	**11.**	(B)	**12.**	(B)	**13.**	(D)	**14.**	(C)
15.	(C)	**16.**	(C)	**17.**	(C)	**18.**	(A)	**19.**	(B)	**20.**	(D)	**21.**	2
22.	26	**23.**	19	**24.**	6	**25.**	2						

PART - 2

1.	(C)	**2.**	(A)	**3.**	(B)	**4.**	(C)	**5.**	(ABC)	**6.**	(AD)	**7.**	(AC)
8.	(ACD)	**9.**	(ABCD)	**10.**	(AD)	**11.**	(AC)	**12.**	(BD)	**13.**	9	**14.**	0
15.	1.	**16.**	7	**17.**	4	**18.**	2						

PART - 3

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

PART - 4

1.	(A)	**2.**	(D)	**3.**	(D)	**4.**	(B)

RRP SOLUTIONS

PART- 1

3. As and Sb behave as metals as well as nonmetals because they form cations (M^{3+}) and anions (M^{3-}). Their oxides and hydroxides react with acid as well as base forming corresponding salts.

4. The first member of the lanthanide series is Cerium (Z= 58).

5. $Z = 15 = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$; so element belongs to p-block. Thus its group number will be 10 + 2 + 3 = 15.
$Z = 33 = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^3$; so element belongs to p-block. Thus its group number will be 10 + 2 + 3 = 15.
$Z = 51 = [Kr]^{36}\ 4d^{10}\ 5s^2\ 5p^3$; so element belongs to p-block. Thus its group number will be 10 + 2 + 3 = 15.

16. Order of $\Delta_{eg}H$ for halogens : Cl > F > Br > I & Order of $\Delta_{eg}H$ for chalcogens : S > Se > Te > Po > O.
Cl and F have the highest and IInd highest values in Modern periodic table.

18. $X(g) \longrightarrow X^+(g) + e^-$
If I.E. is ionisation enthalpy, then

$\therefore \quad \frac{N_0}{2}(I.E.) = E_1$; $\quad \therefore \quad I.E. = \frac{2E_1}{N_0}$

$X(g) + e^- \longrightarrow X^-(g)$
If $\Delta_{eg}H$ is electron gain enthalpy, then

$\therefore \quad 2N_0(E.A.) = -E_2 \quad \therefore \quad \Delta_{eg}H = -\frac{E_2}{2N_0}$.

21. $8s^1$
x = 1, y = 1 ; 1 + 1 = 2.

22. Fe

23. x = 4 Period ; y = 11 Group ; 8 + 11 = 19.

24. BiF_5, $T\ell I_3$, PbO_2, $SnCl_2$, $T\ell_2O_3$, As_2O_3

25. $Ni^{2+} : 3d^8$

↑↓	↑↓	↑↓	↑	↑

No. of unpaired electron = 2

PART - 2

1. (A) $_{21}Sc^{3+}$; $[Ar]^{18}\ 3d^0\ 4s^0$ and $_{21}Sc$; $[Ar]^{18}\ 3d^1\ 4s^2$
As last electron enters in d-subshell so it belongs to d-block and thus its group number = 2 + 1 = 3.
Element belong to 3rd group of Modern periodic table, not zero group.

2. It has only one orbital and single electron. So, shielding effect is not possible.

3. Tl^{3+} gets reduced to Tl^+ because of I^- and then it forms the compound TlI.

4. (A) Successive addition of d-electrons screen the outermost electrons (4s) from the inward pull of the nucleus. As a result of this, the size of the atom does not change much from Cr to Cu.
(B) This is due to lanthanide contraction.

6. The d and f orbitals do not shield the nuclear charge very effectively .Therefore there is significant reduction in the size of the ions, just after d or f orbitals have been filled completely. This is called lanthanide contraction.
Atomic radii of Nb (Nb^{3+} = 0.72 Å) and Ta (Ta^{3+} = 0.72 Å) are almost identical due to lanthanide contraction.
This is also the reason for the higher ionisation energy of gold than silver.

8. Consider the factors on which these properties depend :
(A) Cation is smaller while anion is bigger than its parent atom.
(B) Correct order is Cl > F > Br > I.
(C) Cation is smaller as it is formed by the loss of electron(s). The anion is formed by the gain of electron(s). The size of anion increases with increase in charge on anion i.e. as the Z/e ratio decreases the size increases.
(D) Across the period the size decreases and nuclear size increases. So, ionisation energy increases. However, the first ionisation energy of Mg is greater than Al because of high penetration power of $2s^2$ electrons of Mg as compared to that of $2p^1$ electron of Al.

9. (A) $S^-(g) \longrightarrow S^{2-}(g)$; $\Delta H_{eg} = (+)$ ve because of electrostatic repulsion.
(B) $Ne\ (g) + e^-\ (g) \longrightarrow Ne^-(g)$; $\Delta H_{eg} = (+)$ ve because of stable completely filled electron configuration.
(C) $N(g) \longrightarrow N^-\ (g)$; $\Delta H_{eg} = (+)$ ve because of stable half filled electron configuration.
(D) $Al^{2+}(g) \longrightarrow Al^{3+}(g)$; $\Delta H_{IE} = (+)$ ve because of the removal of electron from cation.

11. Elements of Group 1 & 2 have low IE values. So they impart colour to flame. E.g . Li, Na, K, Ca, Ba etc.

13. Atomic number of Cu is 29 = x

Atomic number of Au is 79 = y
x + y = 108

$$\frac{x+y}{12} = \frac{108}{12} = 9.$$

16. B, C, S, P, At, H, Li

17. (Be, N, P, S)

PART - 3

13. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 6p^3$ element is Bi.

26. $V > V^{+2} > V^{+3}$

$$\text{Radius of same atom} \propto \frac{1}{\text{amount of the charge}}$$

27. Valence shell is n = 3. So valence e^- is 4 ($3s^2\ 3p^2$)

28. Most abundant element is earth crust is oxygen

30. 4f electrons does imperfect shielding on outer electrons there fore effective nuclear charge increase.

31. Tb : [Xe] $4f^9 6s^2$; Eu : [Xe] $4f^7 6s^2$

34. The size of isoelectronic atom or ion depends on the e/z ratio. So it is for O^{-2} ion.

PART - 4

3. a is N b is P c is As d is Sb e is O f is S
g is Se h is Te i is F j is Cl k is Br k is I

4. (A) This configuration belongs to He which has highest first ionisation enthalpy amongst all the elements of the periodic table. This is attributed to stable configuration and its small size.
(B) and (C) Group 17th has $ns^2\ np^5$ valence shell electron configuration. They have highest EN values and very high negative electron gain enthalpy because they can attain stable noble gas electronic configuration by picking up an electron. (B) configuration belongs to fluorine and F has highest electronegativity on Pauling scale. (C) configuration belongs to Cl, which has hte maximum negative electron gain enthalpy (even greater than F ; due to its larger size and lesser interelectronic repulsion).
(D) This configuration belongs to C and it shows –4 oxidation state because it attains inert gas configuration of neon by gaining four electrons.