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.

JEE(Main) Syllabus

Modem periodic law and present form of the periodic table, s, p, d and f block elements, periodic trends in properties of elements atomic and ionic radii, ionization enthalpy, electron gain enthalpy, valence, oxidation states and chemical reactivity.

Periodic Table & Periodicity

Section (A) : Development of Periodic Table & Modern Periodic Table

Need to classify Elements:

- At present 118 elements are known. Of them the recently discovered elements are man-made.
- With such a large no. of elements it is very difficult to study individually the chemistry of all these elements and their innumerable compounds individually.
- To ease out this problem, scientists searched for a systematic way to organize their knowledge by classifying the elements.
- It would rationalize known chemical facts about elements, but even predict new ones for undertaking further study.

Development of Modern Periodic Table:

(a) Dobereiner's Triads: He arranged similar elements in the groups of three elements called as triads, in which the atomic mass of the central element was merely the arithmetic mean of atomic masses of other two elements or all the three elements possessed nearly the same atomic masses.

Li	Na	K	
7	23	39	$\frac{7+39}{2} = 23$
Fe	Co	Ni	
55.85	58.93	58.71	nearly same atomic masses

It was restricted to few elements, therefore discarded.

(b) Newland's Law of Octave: He was the first to correlate the chemical properties of the elements with their atomic masses.
According to him if the elements are arranged in the order of their increasing atomic masses the eighth element starting from given one is similar in properties to the first one.
This arrangement of elements is called as Newland's Octave.

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

This classification worked quite well for the lighter elements but it failed in case of heavier elements and therefore, discarded

(c) Lother Meyer's Classification: He determined the atomic volumes by dividing atomic masses with their densities in solid states.
He plotted a graph between atomic masses against their respective atomic volumes for a number of elements. He found the following observations.

- Elements with similar properties occupied similar positions on the curve.
- Alkali metals having larger atomic volumes occupied the crests.
- Transitions elements occupied the troughs.
- The halogens occupied the ascending portions of the curve before the inert gases.
- Alkaline earth metals occupied the positions at about the mid points of the descending portions of the curve.

On the basis of these observations he concluded that the atomic volumes (a physical property) of the elements are the periodic functions of their atomic masses.
It was discarded as it lacks practical utility.

(d) Mendeleev's Periodic Table:
Mendeleev's Periodic's Law
According to him the physical and chemical properties of the elements are the periodic functions of their atomic masses.
He arranged then known elements in order of their increasing atomic masses considering the facts that elements with similar properties should fall in the same vertical columns and leaving out blank spaces where necessary.

This table was divided into nine vertical columns called groups and seven horizontal rows called periods.
The groups were numbered as I, II, III, IV, V, VI, VII, VIII and Zero group

Merits of Mendeleev Periodic table:

- It has simplified and systematised the study of elements and their compounds.
- It has helped in predicting the discovery of new elements on the basis of the blank spaces given in its periodic table.
  Mendeleev predicted the properties of those missing elements from the known properties of the other elements in the same group. Eka-aluminium and Eka-silicon names were given for gallium and germanium (not discovered at the time of Mendeleev). Later on it was found that properties predicted by Mendeleev for these elements and those found experimentally were almost similar.

Table-1

Proeprty	eka-aluminium (predicted)	Gallium (found)	eka-silicon (predicted)	Germanium (found)
Atomic Mass	68	70	72	72.6
Density / (g/cm^3)	5.9	5.94	5.5	5.36
Melting point (K)	Low	30.2	High	1231
Formula of oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
Formula of chloride	ECl_3	$GaCl_3$	ECl_4	$GeCl_4$

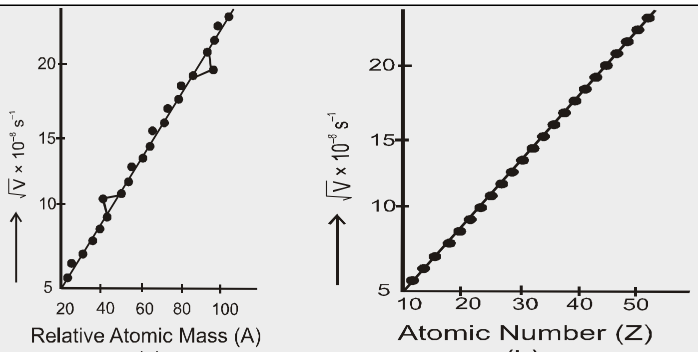
- Atomic weights of elements were corrected. Atomic weight of Be was calculated to be 3 × 4.5 = 13.5 by considering its valency 3, was correctly calculated considering its valency 2 (2 × 4.5 = 9)

Demerits in Mendeleev's Periodic Table :

- Position of hydrogen is uncertain. It has been placed in IA and VIIA groups because of its resemblance with both the groups.
- No separate positions were given to isotopes.
- Anomalous positions of lanthanides and actinides in periodic table.
- Order of increasing atomic weights is not strictly followed in the arrangement of elements in the periodic table. For example Ar(39.94) is placed before K(39.08) and Te (127.6) is placed before I (126.9).
- Similar elements were placed in different groups e.g. Cu in IB and Hg in IIB and similarly the elements with different properties were placed in same groups e.g. alkali metals in IA and coinage metals in IB.
- It didn't explained the cause of periodicity.

(e) Long form of the Periodic Table or Moseley's Periodic Table or Modern Periodic Table :

S.No.	Introduction	DISCRIPTION
1.	**Proposed by**	**Moseley**
2.	**Contribution**	(i) In the long form of periodic table there is contribution of Ramsey, Werner, Bohr and Bury. (ii)This table is also referred to as Bohr's table since it follows Bohr's scheme of the arrangements of elements into four types based on electronic configuration of elements The modern periodic table consits of horizontal rows (periods) and vertical column (groups).
3.	**Based on**	Atomic number
4.	**Experiment**	(i) Moseley did an experiment in which he bombarded high speed electrons on different Metal surfaces and obtained X-rays(electromagnetic rays). He observed regularities in the characteristic X-ray spectra of the elements and found that plot $\sqrt{\nu}$ *vs.* Z (atomic number) is straight line while $\sqrt{\nu}$ *vs. A* (atomic weight) is not, and $\sqrt{\nu} = a(Z - b)$, where *a* and *b* are constants that are same for all elements and ν is frequency of X-rays. Thus he concluded that atomic number is more fundamental property than atomic weight.

		 (a) (b) (ii) Moseley concluded that the physical and chemical properties of the elements are periodic function of their atomic number.
5.	**Modern Periodic Law**	The physical and chemical properties of elements are periodic function of their **atomic number**. So the elements are arranged in order of increasing atomic number, the elements with similar properties comes after regular intervals.
6.	**Periodicity**	The repetition of the properties of elements after regular intervals when the elements are arranged in the order of increasing atomic number is called periodicity. (a) In a period, the ultimate orbit remain same, but the number of electrons gradually increases. (b) In a group, the number of electrons in the ultimate orbit remains same, but the values of n increases.
7.	**Cause of Periodicty**	The periodic repetition of the properties of the elements is due to the recurrence of similar valence shell electronic configuration after certain regular intervals. For example, alkail metals have same electronic configuration ns^1, therefore, have similar properties. In the periodic table, elements with similar properties occur at intervals of 2, 8, 8, 18, 18 and 32. These numbers are called as magic numbers.

The modern periodic table consists of horizontal rows (periods) and vertical column (groups).

Periods :

There are seven periods numbered as 1, 2, 3, 4, 5, 6 and 7.

- Each period consists of a series of elements having same valence shell.
- Each period corresponds to a particular principal quantum number of the valence shell present in it.
- Each period starts with an alkali metal having outermost electronic configuration as ns^1.
- Each period ends with a noble gas with outermost electronic configuration ns^2np^6 except helium having outermost electronic configuration as $1s^2$.
- Each period starts with the filling of new shell.
- The number of elements in each period is twice the number of atomic orbitals available in shell that is being filled. For illustration–
  - Ist period shortest period having only two elements. Filling of electrons takes place in the first shell, for which, n = 1, ℓ = 0 (s-subshell) and m = 0.
    Only one orbital (1s) is available and thus it contains only two elements.
  - 3rd period (short period) having only eight elements. Filling of electrons takes place in the third shell. For which,

    n = 3, ℓ = 0, 1, 2 and number of orbitals 1 (3s) 3 (3p) 5 (3d)

    Total number of orbitals 9

    But the energy of 3d orbitals are higher than 4s orbitals. Therefore, four orbitals (one 3s and three 3p orbitals) corresponding to n = 3 are filled before filling in 4s orbital (next energy level). Hence 3rd period contains eight elements not eighteen elements.

Groups :

There are eighteen groups numbered as 1, 2, 3, 4, 5, ... 13, 14, 15, 16, 17, 18.

Group consists of a series of elements having similar valence shell electronic configuration.

Table-2

Periods	Number of elements	Called as
$(1)^{st}$ n = 1	2	Very short period
$(2)^{nd}$ n = 2	8	Short period
$(3)^{rd}$ n = 3	8	Short period
$(4)^{th}$ n = 4	18	Long period
$(5)^{th}$ n = 5	18	Long period
$(6)^{th}$ n = 6	32	Very long period
$(7)^{th}$ n = 7	32	Very long period

S–Block Elements (groups 1–2) | **d –Block Elements** (groups 3–12) | **p–Block Elements** (groups 13–18)

1 IA	2 II A	3 III B	4 IV B	5 V B	6 VI B	7 VII B	8 VIII	9 VIII	10 VIII	11 I B	12 II B	13 III A	14 IV A	15 V A	16 VI A	17 VII A	18 VIII A
1 H 1.007																	2 He 4.002
3 Li 6.941	4 Be 9.012											5 B 10.811	6 C 12.011	7 N 14.006	8 O 15.999	9 F 18.998	10 Ne 20.179
11 Na 22.98	12 Mg 24.30											13 Al 26.981	14 Si 28.085	15 P 30.973	16 S 32.006	17 Cl 35.452	18 Ar 39.948
19 K 39.08	20 Ca 40.078	21 Sc 44.959	22 Ti 47.88	23 V 50.9415	24 Cr 51.996	25 Mn 54.938	26 Fe 55.84	27 Co 55.933	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.921	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.46	38 Sr 87.62	39 Y 88.905	40 Zr 91.224	41 Nb 92.906	42 Mo 95.94	43 Tc 98	44 Ru 101.07	45 Rh 102.905	46 Pd 106.42	47 Ag 107.868	48 Cd 112.411	49 In 114.82	50 Sn 118.710	51 Sb 121.757	52 Te 127.60	53 I 126.904	54 Xe 132.29
55 Cs 132.90	56 Ba 137.27	57 La* 138.905	72 Hf 178.49	73 Ta 180.947	74 W 183.85	75 Re 186.207	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.666	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 207.980	84 Po 209	85 At 210	86 Rn 222
87 Fr 223	88 Ra 226	89 Ac** 227	104 Rf 261.11	105 Ha 262.114	106 Sg 263.118	107 Bh 262.12	108 Hs 265	109 Mt 266	110 Ds 269	111 Rg 272	112 Cn 285	113 Uut 284	114 Fl 289	115 Uup 288	116 Lv 292	117 Uus	118 Uuo 294

Inner - Transition Metals (f-Block elements)

***Lanthanides**	58 Ce 140.115	59 Pr 140.907	60 Nd 144.24	61 Pm 145	62 Sm 150.36	63 Eu 151.965	64 Gd 157.25	65 Tb 158.925	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.967
****Actinides**	90 Th 232.038	91 Pa 231	92 U 238.028	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 260

Classification of the Elements :

It is based on the type of orbitals which receives the differentiating electron (i.e., last electron).

(a) s-block elements

When shells upto (n – 1) are completely filled and the last electron enters the s-orbital of the outermost (n^{th}) shell, the elements of this class are called s-block elements.

- Group 1 & 2 elements constitute the s-block.
- General electronic configuration is [inert gas] ns^{1-2}
- s-block elements lie on the extreme left of the periodic table.
- This block includes metals.

(b) p-block elements

When shells upto (n – 1) are completely filled and differentiating electron enters the p-orbital of the n^{th} orbit, elements of this class are called p-block elements.

- Group 13 to 18 elements constitute the p-block.
- General electronic configuration is [inert gas] $ns^2\,np^{1-6}$
- p-block elements lie on the extreme right of the periodic table.
- This block includes some metals, all nonmetals and metalloids.

(c) d-Block elements

When outermost (n^{th}) and penultimate shells $(n-1)^{th}$ shells are incompletely filled and differentiating electron enters the (n – 1) d orbitals (i.e., d-orbital of penultimate shell) then elements of this class are called d-block elements.

- Group 3 to 12 elements constitute the d-block.
- General electronic configuration is [inert gas] $(n-1)d^{1-10}ns^{1-2}$ (except, palladium which has valence shell electron configuration $4d^{10}5s^0$).
- All the transition elements are metals and most of them form coloured complexes or ions.
- d-block elements are classified into four series as given below.

(1) I^{st} transition series i.e. 3d series contains 10 elements and starts from $_{21}Sc–_{30}Zn$. Filling of electrons takes place in 3d sub-shell.

(2) II^{nd} transition series i.e. 4d series contains 10 elements and starts from $_{39}Y–_{48}Cd$. Filling of electrons takes place in 4d sub-shell.

(3) III^{rd} transition series i.e. 5d series contains 10 elements and starts from $_{57}La$, $_{72}Hf–_{80}Hg$. Filling of electrons takes place in 5d sub-shell.

(4) IV^{th} transition series i.e. 6d series contains 10 elements and starts from $_{89}Ac$, $_{104}Rf–_{112}Uub$. Filling of electrons takes place in 6d sub-shell (incomplete series).

- Those elements which have partially filled d-orbitals in neutral state or in any stable oxidation state are called transition elements. (Zn, Cd, Hg, Uub are not transition elements)

(d) f-Block elements

When n, (n – 1) and (n – 2) shells are incompletely filled and last electron enters into f-orbital of antepenultimate i.e., $(n-2)^{th}$ shell, elements of this class are called f-block elements. General electronic configuration is $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$

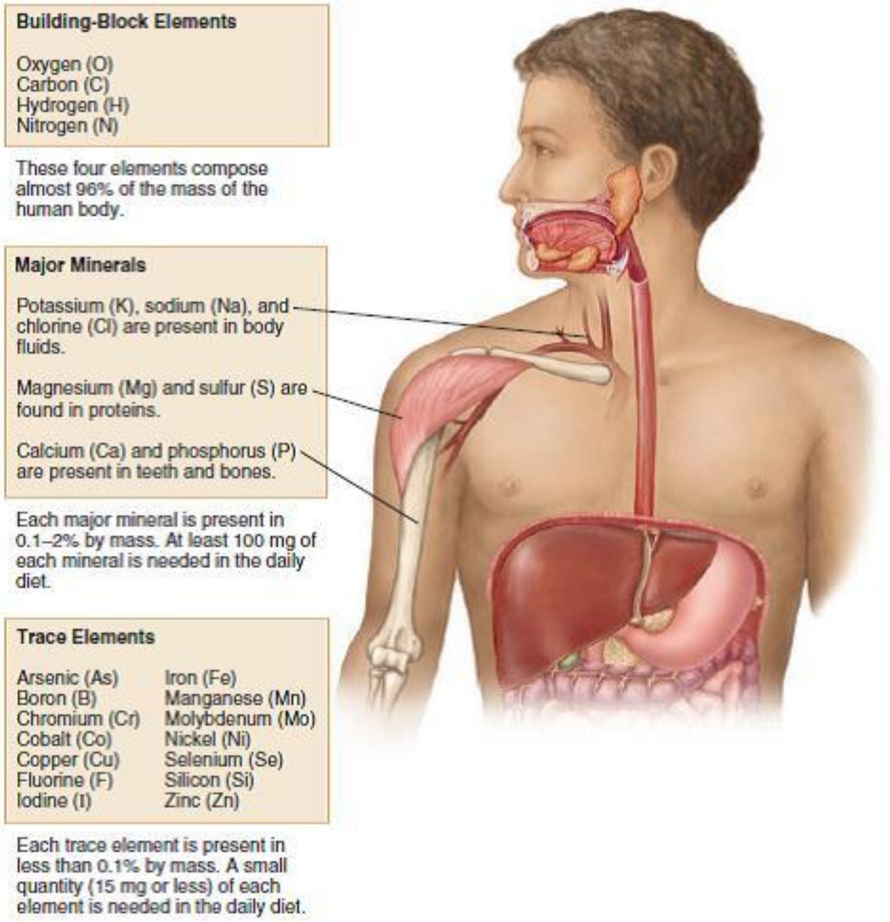
- All f-block elements belong to 3^{rd} group.
- They are metals
- Within each series, the properties of the elements are quite similar.
- The elements coming after uranium are called transuranium elements.
- They are also called as inner-transition elements as they contain three outer most shell incomplete and were also referred to as rare earth elements since their oxides were rare in earlier days.

The elements of f-blocks have been classified into two series.

1. I^{st} inner transition or 4 f-series, contains 14 elements $_{58}Ce$ to $_{71}Lu$. Filling of electrons takes place in 4f subshell.

2. IInd inner transition or 5 f-series, contains 14 elements $_{90}Th$ to $_{103}Lr$. Filling of electrons takes place in 5f subshell.

- The actinides and lanthanides have been placed at the bottom of the periodic table to avoid the undue expansion of the periodic table.



Solved Examples

Ex-1 Elements A, B, C, D and E have the following electronic configurations :

A : $1s^2\ 2s^2\ 2p^1$ B : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$ C : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

D : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ E : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

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

Sol. Out of these, elements A and B will belong to the same group of the periodic table because they have same outer electronic configuration, $ns^2\ np^1$.

Prediction of period, group and block:

- Period of an element corresponds to the principal quantum number of the valence shell.
- The block of an element corresponds to the type of subshell which receives the last electron.
- The group is predicted from the number of electrons in the valence shell or/and penultimate shell as follows.

(a) For s-block elements, Group number = the number of valence electrons

(b) For p-block elements, Group number = 10 + number of valence electrons

(c) For d-block elements, Group number = number of electrons in (n – 1) d sub shell + number of electrons in valence shell.

(d) For f block element – $\left.\begin{matrix}(58\text{-}71) & 4f \\ (92\text{-}103) & 5f\end{matrix}\right]$ - All f block elements are belongs to group number 3.

Two exceptions to this categorization:

Strictly, helium belongs to the s-block but its positioning in the p-block along with other group 18 elements in justified because it has a completely filled valence shell ($1s^2$) and as a result, exhibits properties & characteristics of other noble gases.

Hydrogen has only one s-electron and hence can be placed in group 1 (alkali metals). It can also gain an electron to achieve a noble gas arrangement and hence it can behave similar to a group 17 (halogen family) elements. Because it is a special case, we shell place hydrogen separately at the top of the periodic table.

Isotopes in modern periodic table:

As isotopes of a elements have same atomic number and in modern periodic table the elements have been allotted places based on their atomic number. All the isotopes an elements have been assigned the same position in the modern periodic table.

Solved Examples

Ex-2 An element X with Z = 112 has been recently discovered. What is the electronic configuration of the element ? To which group and period will it belong ?

Sol. (a) The electronic configuration of element X is $[Rn]^{86}\ 5f^{14}\ 6d^{10}7s^2$

(b) It belongs to d-block as last electron enters in d subshell.

(c) As number of electrons in (n – 1)d subshell and valence shell is equal to twelve i.e. 10 + 2. So it belongs to group 12.

(d) It belongs to period 7 of the periodic table as principal quantum number of valence shell is 7 (i.e., $7s^2$).

Metals and nonmetals :

- The metals are characterised by their nature of readily giving up the electron(s) and from shining lustre. Metals comprises more than 78% of all known elements and appear on the left hand side of the periodic table. Metals are usually solids at room temperature (except mercury, gallium). They have high melting and boiling points and are good conductors of heat and electricity. Oxides of metals are generally basic in nature (some metals in their higher oxidation state form acid oxides e.g. CrO_3).
- Nonmetals do not lose electrons but take up electrons to form corresponding anions. Nonmetals are located at the top right hand side of the periodic table. Nonmetals are usually solids, liquids or gases at room temperature with low melting and boiling points. They are poor conductors of heat and electricity. Oxides of nonmetals are generally acidic in nature.

Metalloids (Semi metals) :

- It can be understood from the periodic table that nonmetallic character increases as we move from left to right across a row. It has been found that some elements which lie at the border of metallic and nonmetallic behavior, possess the properties that are characteristic of both metals and nonmetals. These elements are called semi metals or metalloids.
- The metalloids comprise of the elements Si, Ge, As, Sb, Se and Te.
- Oxides of metalloids are generally amphoteric in nature.

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 / ionic radius ratio is nearly similar because the charges are 2+ and 3+ respectively.

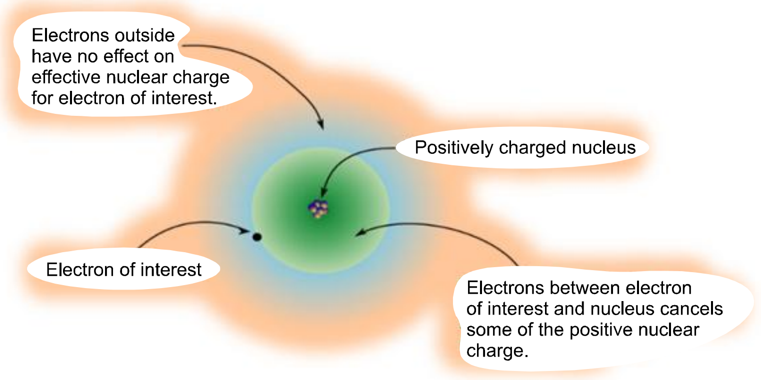
$$\text{Charge /Ionic radius ratio} = \frac{(\text{Ionic charge})}{(\text{Ionic radius})}$$

Section (B) : Shielding Effect and Z_{eff}

The periodicity of atomic properties :

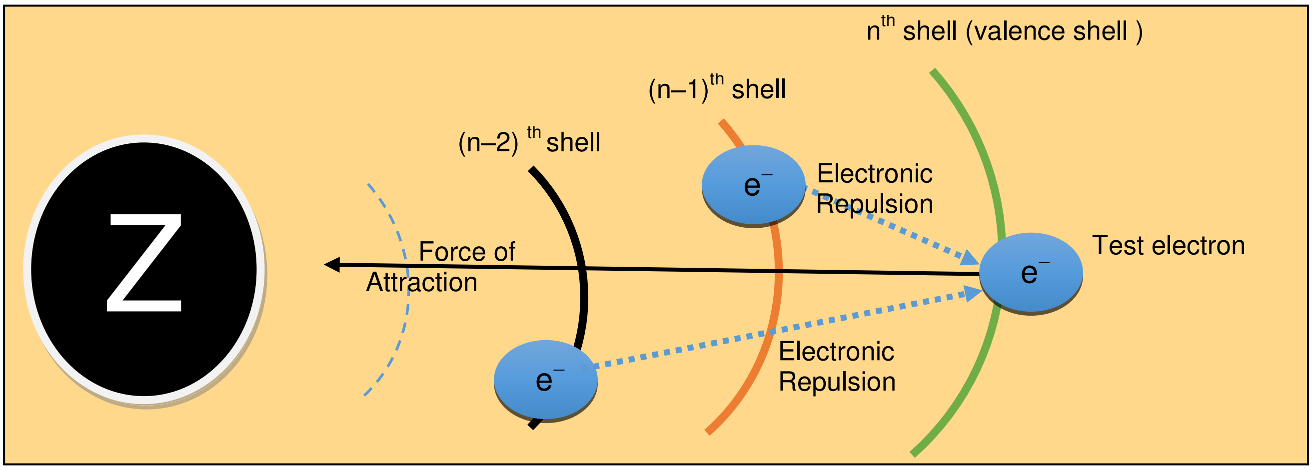
(1) Effective nuclear charge (Z_{eff}):

The outer electron experiences two opposing force : -
(i) Force of attraction from nucleus.
(ii) Force of repulsion from inner electrons.
Suppose inner electrons are not present there then force of attraction experience by test electron must be greater than it actually experienced because electronic repulsion weakens the force of attraction. **This reduced nuclear charge or nuclear actually experienced by an electron is termed as effective nuclear charge.**



$Z_{eff} = Z - \sigma$
Z_{eff} = Effective nuclear charge
Z = Atomic number
σ = Screening constant or shielding constant or slater constant

Effective nuclear charge is not same for all the electrons present in an atom and it varies with distance between electron and nucleus.
This effect in which inner electrons shield valence electron from the attraction due to nucleus is called shielding effect.
It is also known as screening effect as inner electron create a protective screen between nucleus and valence electrons.
Due to shielding effect valence electron experiences less attraction from the side of nucleus as if inner electrons are absent.
s-orbitals have the largest screening effect for a given n value since s electrons are closer to the nucleus.
Screening effect decreases as s-orbital > p-orbital > d-orbital > f-orbital.



Section (C) : Oxidation states & Inert pair effect

(2) Oxidation States:

Definition: The oxidation states (O.S) of an atom in a molecule or in an ion is the (imaginary) charge the atom would have if the electron in each bond were located on the more electronegative atom.
Rule for determining the oxidation state
Fundamental Rule: Sum of oxidation no. of all the atoms of a species is equal to the charge on that species.
e.g. For SO_4^{2-} ion ; (O.N. of S atom) + (O.N. of O atom) × 4 = –2

For C_{60} molecule ; (O.N. of C atom) × 60 = 0
For NO_2^+ ion ; (O.N. of N atom) + (O.N. of O atom) × 2 = +1

Corollary 1: All atoms have oxidation state zero in their elemental state
Eg. Phosphorus atoms in P_4, C atoms in graphite or diamond or fullerenes, O atoms in nascent oxygen [O], diatomic O_2.
Remember: Oxidation state of uncombined element is zero.
Corollary 2: Charge on any monoatomic ion is equal to its oxidation state.

Note : Oxidation state is often imagined as charge on the atom. In an ionic species this idea work very well. In a covalent molecule, the electrons involved in bonding are considered to belong entirely to the more electronegative species (i.e. it is assumed to be entirely ionic).
Caution: This is just an imagination, not the real picture.
Other Rules: In species when the following are combined with dissimilar atom(s) their O.No. is
(i) F atom: –1 ; alkali metals : +1 ; alkaline earth metals : +2; Al, Ga : +3 No exception.
(ii) H is mostly: +1, Exception: Hydrides of metals mentioned in (i) Where it's O.N is –1.
(iii) O is mostly: –2, exceptions, (a) in peroxides (b) when in conflict with rules (i) and (ii) e.g. in peroxides like BaO_2 , Ba is fixed as + 2. If we assume O as –2, Ba will have to assume a charge of +4 which is not permissible [Since Rule (i) has more weightage than rule (ii) or (iii)]. Hence, Ba must have a charge of +2. Accordingly, we must adjust oxygen. Naturally we have O_2^{2-}, this is immediately identified as our familiar peroxide ion. Naturally oxidation number of each atom is –1 in this case. Similarly O atom is +1 or +2 with fluorine in O_2F_2 and OF_2.
(iv) N is mostly: –3 e.g. in ammonia or ammonium ion e.g. in CN^-, we can simply assume N as –3 and move a head. Accordingly, C will be +2 (Check it for yourself by fundamental rule). Exceptions : When in conflict with rule (i) (iii) eq. N +3 in NF_3 (since F must be –1) and +5 in NO_3^- (since O must be –2).
(v) Cl, Br and I are usually –1 e.g. in halides.
Exceptions: When is conflict with rules (i)-(iv). Amongst halogens, the reactivity order is Cl > Br > I e.g. in ICl, Cl is regarded as –1. Naturally, I must adjust itself as +1. $BrCl_3$ Br is + 3 since Cl must be –1. However, is Cl_2O, Cl must adjust to +1 state. Since O should preferably be –2.
Interestingly, in a case like $POCl_3$, both O and Cl can have their desirable state of –2 and –1 respectively. In this case, P must adjust as per the convenience of O and Cl atoms (which have higher priority i.e. higher electonegativity).
(vi) S is –2 in sulphides. If however, can have higher oxidation state, highest being equal to the group no. (i.e. 6)
(vii) Metal usually do not exhibit negative oxidation states. **Exceptions:** Gold is sometimes found in –1.
Remember: Non-stoichiometric compounds always have metal cation in at least two different oxidation state.
Average oxidation number: A given element may be present in more than one oxidation states within a given molecule if more than one atoms of the element are present.
e.g. NH_4NO_2 contain N atom in NH_4^+ ion as well as NO_2^- ion. Try finding their oxidation state separately. In NH_4^+ ion, N is present in –3 oxidation state and in NO_2^- ion N is present in +3 oxidation state. Else, even if you consider the average state of N atom in the compound, it comes out to be [{(+3) + (–3)}/2] that is 0.

(3) Inert Pair Effect :

The outer shell 's' electrons (ns^2) penetrate to (n–1)d electrons and thus become closer to nucleus and are more effectively pulled towards the nucleus. This results in less availability of ns^2 electron pair for bonding or ns^2 electron pair becomes inert. This is called Inert Pair Effect. The inert pair effect begins after $n \geq 4$ and increases with increasing value of n. This can be thought as a reason for two different oxidation states, normal and (normal – 2), among elements of Boron, Carbon and Nitrogen family.
In these families, the stability of higher oxidation state decreases while that of lower oxidation state increases on moving down the group.
Eg. $Ga^{3+} > In^{3+} > Tl^{3+}$; $Ga^+ < In^+ < Tl^+$
$Ge^{4+} > Sn^{4+} > Pb^{4+}$; $Ge^{2+} < Sn^{2+} < Pb^{2+}$
$As^{5+} > Sb^{5+} > Bi^{5+}$; $As^{3+} < Sb^{3+} < Bi^{3+}$
Thats why, Pb^{4+} and Bi^{5+} compounds are strong oxidising agents. However the last elements of the group do form compounds in oxidation state, but only with highly electronegative higher elements like F and O.
So compounds like TlF_3, Tl_2O_3, PbF_4, $PbCl_4$, BiF_5, $NaBiO_3$ exist in nature while compounds like PbI_4, $PbBr_4$, BiI_5, $BiBr_5$, $BiCl_5$ etc do not exist. TlI_3 exists exceptionally as $Tl^+(I_3)^-$.

Table-3

List of common oxidation sate of an element in Periodic Table

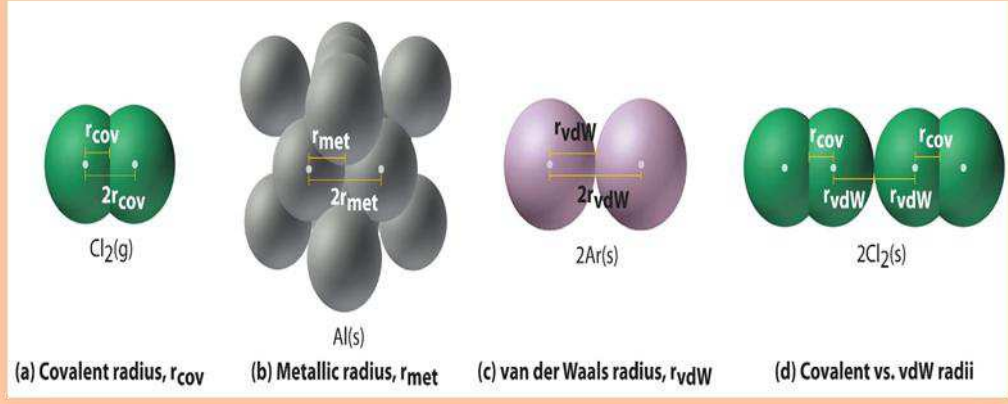
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H **+1** –1																	2 **He**
3 Li +1	3 Be +2											5 **B** **+3** –3	6 **C** **+4** +2 –4 etc.	7 **N** +5 +4 +3 +1 –3 **0** etc.	8 **O** +2 –1/2 –1 **–2**	9 **F** **–1**	10 **Ne**
11 Na +1	12 Mg +2											13 **Al** +3	14 **Si** **+4** –4	15 **P** **+5** +3 +1 –3	16 **S** **+6** +4 +2 –2	17 **Cl** +7 +5 +3 +1 0 **–1**	18 **Ar** **0**
19 **K** +1	20 **Ca** +2	21 **Sc** +2 **+3**	22 **Ti** +2 **+3** **+4**	23 **V** +2 **+3** +4 +5	24 **Cr** +2 **+3** +4 +5 +6	25 **Mn** **+2** +3 **+4** +5 +6 +7	26 **Fe** **+2** +3 +4 +5 +6	27 **Co** **+2** +3 +4 +5	28 **Ni** **+2** +3 +4	29 **Cu** **+1** +2	30 **Zn** **+2**	31 **Ga** +3	32 **Ge** **+4** –4	33 **As** **+5** **+3** –3	34 **Se** **+6** +4 –2	35 **Br** +7 +5 +3 +1 **–1**	36 **Kr** +4 +2 **0**
37 **Rb** +1	38 **Sr** +2											49 **In** **+3** +1	50 **Sn** **+4** +2	51 **Sb** **+5** **+3** –3	52 **Te** **+6** +4 –2	53 **I** +7 +5 +3 +1 **0** **–1**	54 **Xe** +8 +6 +4 +2 **0**
55 **Cs** +1	56 **Ba** +2											81 **Tl** +3 **+1**	82 **Pb** +4 **+2**	83 **Bi** +5 **+3**	84 **Po**	85 **At**	86 **Rn**

* **Bold mark oxidation number are general stable oxidation number of an element in compound state.**

Section (D) : Atomic and Ionic radius

(4) Atomic radius :

It is distance between outermost electron and nucleus.

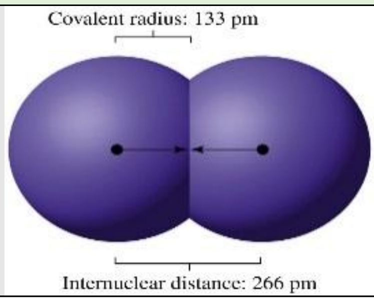


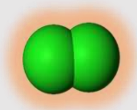
(a) Covalent radius, r_{cov} (b) Metallic radius, r_{met} (c) van der Waals radius, r_{vdW} (d) Covalent vs. vdW radii

X-ray diffraction, electron diffraction method and nuclear magnetic resonance (NMR) spectrum methods are used to determine internuclear distance or bond length.

Atomic radius depends on the type of chemical bond between atoms in a molecule. These are :

1. Covalent radius **2. Ionic radius** **3. Metallic radius** **4. Vander waal's radius**

S.No.	Type of radius	DISCRIPTION
1.	**Covalent radius** SBCR(Single Bonded Covalent Radius)	 Covalent radius is calculated in covalent compounds. Covalent compound's form covalent bonds. Covalent bonds are formed by overlapping of atomic orbitals. Type of Covalent radius.

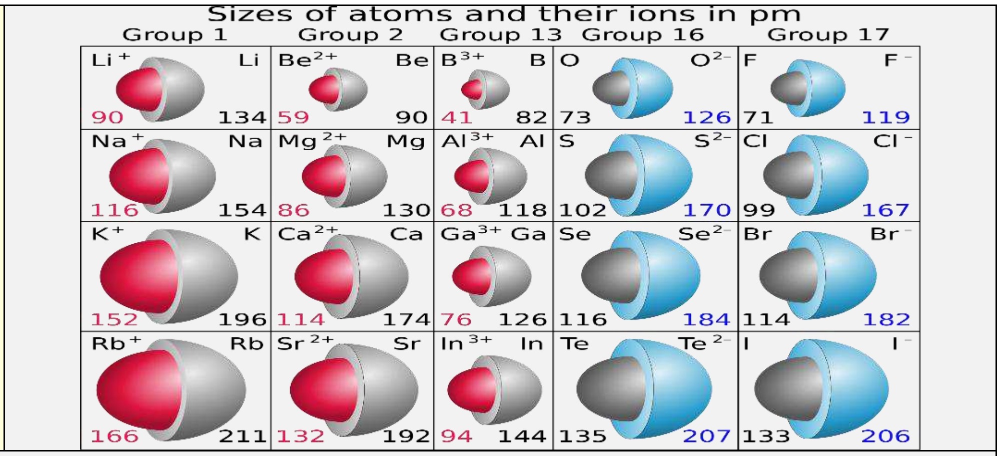
Halogen	Molecule	Structure	Model	*d*(X–X) / pm (gas phase)	*d*(X–X) / pm (solid phase)
Fluorine	F_2	F — F ⟷ 143pm		143	149
Chlorine	Cl_2	Cl — Cl ⟷ 199pm		199	198
Bromine	Br_2	Br — Br ⟷ 228pm		228	227
Iodine	I_2	I — I ⟷ 266pm		266	272

	(i) In Homo atomic Molecules	Covalent radius is the one-half of the internuclear distance between two singly bonded two nuclei (of like atoms).
		Internuclear distance of A–A(A_2) molecule is (d_{A-A}) and covalent radius is r_A then $d_{A-A} = r_A + r_A \Rightarrow d_{A-A} = 2r_A$ $\boxed{r_A = \frac{d_{A-A}}{2}}$
		e.g. – In Cl_2 molecule, internuclear distance is 1.98 Å so $r_{Cl} = \frac{1.98}{2} = 0.99$ Å
		SBCR of O, N and C etc. elements can be determined by taking H_2O_2, N_2H_4, C_2H_6 respectively.
	(ii) Hetero atomic molecule	**(a) For Hetero atomic molecule with not E.N. difference.** $\boxed{d_{A-B} = r_A + r_B}$
		Example : For A–B molecule Electronegativities of A and B are approximately equal e.g. C–I E.N. of C and I are approx equal (2.5) internuclear distance of C–I is 2.13Å and r_C is 0.77Å. **Solution :** $d_{C-I} = r_C + r_I \;\therefore\; r_I = 2.13 - 0.77 = 1.36$Å
		(b) Heteroatomic molecule with Δ E.N. difference more
		If in a diatomic molecule electronegativities of A–B have more difference. Then actual bond length will be reduced.
2.	**Ionic Radius**	
	(i) Cationic radius	When an neutral gaseous atom loses electron it converts into cation. **Radius = Atomic radius > Cationic radius**
		Reason: After loosing electron number of electron reduces, but number of protons remains same, due to this Z_{eff} increases, hence electrons pulls towards nucleus and atomic radius decreases, moreover after losing all the electrons from the outer most shell, penultimate shell becomes ultimate shell which is nearer to nucleus so size decreases.
		$\boxed{\text{Size of cation} \propto \frac{1}{\text{Magnitude of +ve charge or } Z_{eff}}}$
		Example Size of cation = Fe > Fe^{+2} > Fe^{+3} Size of cation = Pb^{+2} > Pb^{+4} Size of cation = Mn > Mn^{+2} > Mn^{+3} > Mn^{+4} > Mn^{+5} > Mn^{+6} > Mn^{+7}
	(ii) Anionic radius	When neutral gaseous atom gains electron it converts into anion. **Radius = Anionic radius > Atomic radius**
		Reason: In an anion electron are more than protons so effective nuclear charge reduces, and inter electronic repulsion increases, which also increases screening effect. So distance between electron and nucleus increases and size of anion also increases.
		$\boxed{\text{Size of anion} \propto \text{Magnitude of} - \text{ve charge}}$
		Example : Size of F^- > F why ? Sol.

	F	F^-
Proton	9	9
electron	9	10
$\frac{Z}{e^{\ominus}}$	$\frac{9}{9} = 1$	$\frac{9}{10} = 0.9$

As Z_{eff} of F^- is less than F so size of F^- > F

(iii) Size of iso electronic species



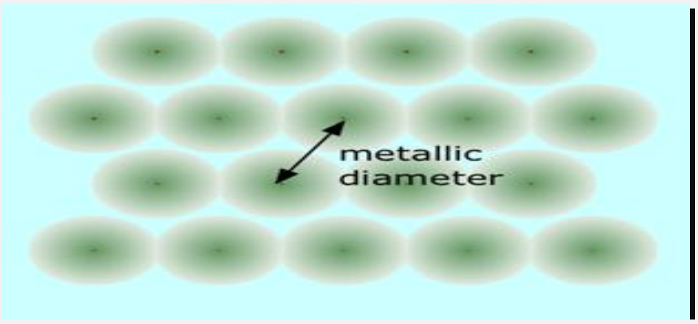
					H^-	He
Li^+	Be^{+2}		N^{3-}	O^{2-}	F^-	Ne
Na^+	Mg^{+2}	Al^{+3}		S^{2-}	Cl^-	Ar
K^+	Ca^{+2}	Sc^{+3}		Se^{2-}	Br^-	Kr
Rb^+	Sr^{+2}	Y^{+3}		Te^{2-}	I^-	Xe
Cs^+	Ba^{+2}	La^{+3}				

Those species having same number of electron but different nuclear charge forms isoelectronic series.
For isoelectronic species the atomic radius increases with decrease in nuclear charge.

Ions	K^+	Ca^{2+}	S^{2-}	Cl^-
Atomic number	19	20	16	17
electron	18	18	18	18

Example
Order of radius = $S^{2-} > Cl^- > K^+ > Ca^{2+}$
Order of radius = $P^{3-} > S^{2-} > Cl^-$
Order of radius = $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$
Order of radius = $Au^+ > Hg^{2+} > Tl^{3+} > Pb^{4+}$

3. **Metallic radius**



It is one-half of the nuclear distance between two adjacent metallic atoms in crystalline lattice structure.
so metallic radius > Covalent radius.

$$\text{Metallic radius} \propto \frac{1}{\text{Metallic bond strength}}$$

More metallic radius → loose crystal packing → less bond strength (body centered packing)

Less metallic radius → tight crystal packing FCC → High bond strength (Hexagonal close packing)

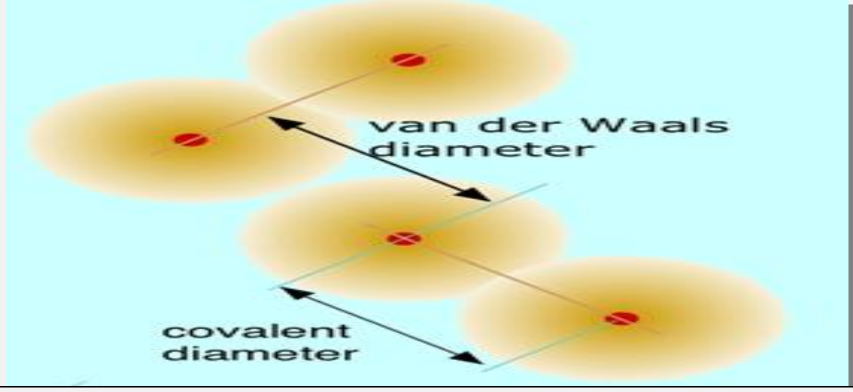
4.	**Van der Waal's radius**	Those atoms (like noble gases) which are not bonded with each other, experiences a weak attractive force to come nearer.
		The half of the distance between the nuclei of adjacently placed atoms in solid state of a noble gas is Vander Waal's radius.
		Vander Waal radius > Covalent radius.
		Inert gas have only Vander Waal radius
		In molecules of nonmetals both covalent and Vander Waal radius exists. 
		Radius = Vander Waal radius > Metallic Radius > Covalent Radius
		Van der waal force of attraction $\propto$ Molecular weight OR atomic weight (in inert gases)
		In a period from left to right Vander Waal radius decreses. In a group from top to bottom its values increases.
5.	**Factors affecting atomic size**	
	(i) Effective nuclear charge	Atomic radius $\propto \dfrac{1}{\text{Effective nuclear charge}}$ — **Example** Li > Be > B > C > N > O > F
	(ii) Number of shells	Atomic radius $\propto$ No. of shells — **Example** Li < Na < K < Rb < Cs
	(iii) screening effect	Atomic radius $\propto$ Screening effect
	(iv) Magnitude of +ve charge	Atomic radius $\propto \dfrac{1}{\text{Magnitude of + ve charge}}$ — **Example** Mn > Mn^{+2} > Mn^{+3} > Mn^{+4}
	(v) Magnitude of -ve charge	Atomic radius $\propto$ Magnitude of – ve charge — **Example** O < O^{-} < O^{-2}
	(vi) Bond order	Atomic radius $\propto \dfrac{1}{\text{Bond order}}$ — **Example** N–N > N=N > N≡N

Table-4

Variation in a Period (left to right) :	Variation in a Group (top to bottom) :
Nuclear charge (Z) increases by one unit	Nuclear charge (Z) increases by more than one unit
Effective nuclear charge (Z_{eff}) also increases	Effective nuclear charge (Z_{eff}) almost remains constant because of increased screening effect of inner shells electrons.
But number of orbitals (n) remains constant	But number of orbitals (n) increases.
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.	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.

Irregularties

- In the transition series (e.g. in first transition series), the covalent radii of the elements decrease from left to right across a row until near the end when the size increases slightly. On passing from left to right, extra protons are placed in the nucleus and the extra electrons are added. The orbital electrons shield the nuclear charge incompletely. Thus the nuclear charge attracts all the electrons more strongly, hence a contraction in size occurs. The radii of the elements from Cr to Cu, are very close to one another because the successive addition of d-electrons screen the outer electrons (4s) from the inward pull of the nucleus. As a result of this, the size of the atom does not change much in moving from Cr to Cu.

Table-5

Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Atomic radius (pm)	144	132	122	117	117	117	116	115	117	125

- The elements of group 3 belonging to d-block show the expected increase in size as that found in case of s-block and p-block elements. However, in the subsequent groups that is 4^{th} onwards (upto 12^{th} group), there is increase in size between the first and second member, but hardly any increase between the second and third elements of the given group. There are 14 lanthanide elements between lanthanum and hafnium, in which the antepenultimate 4f shell of electrons (exert very poor shielding effect) is filled. There is a gradual decrease in size of the 14 lanthanide elements from cerium to lutetium. This is called lanthanide contraction. This lanthanide contraction cancels out the normal size increase on descending a group in case of transition elements.

Solved Examples

Ex-3 Atomic radius of Li is 1.23 Å and ionic radius of Li^+ is 0.76 Å. Calculate the percentage of volume occupied by single valence electron in Li.

Sol. Volume of Li = $\frac{4}{3} \times 3.14 \times (1.23)^3 = 7.79$ Å^3 (–Li = $1s^2 2s^1$)

Volume of Li^+ = $\frac{4}{3} \times 3.14 \times (0.76)^3 = 1.84$ Å^3 (–Li^+ = $1s^2$)

$\therefore$ Volume occupied by 2s subshell = 7.79 – 1.84 = 5.95 Å^3.

$\therefore$ % Volume occupied by single valence electron i.e., 2s electron = $\frac{5.95}{7.79} \times 100 = 76.4\%$

Ex-4 Select from each group the species which has the smallest radius stating appropriate reason.

(a) O, O^-, O^{2-} (b) P^{3+}, P^{4+}, P^{5+}

Sol. (a) O is having smallest radius. Anion is larger than its parent atom. Also the anion of the same atom with higher negative charge is bigger in size as compared to anion with smaller negative charge as proton to electron ratio decreases thus attraction between valence shell electrons and nucleus decreases. Hence electron cloud expands.

(b) The ionic radius decreases as more electrons are ionized off that is as the valency increases. So the correct order is $P^{5+} < P^{4+} < P^{3+}$.

Ex-5 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-}.

Section (E) : Ionisation energy

(5) 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.

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

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

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, $(IE)_1 < (IE)_2 < (IE)_3 <$ .... 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 not be regular and can be influenced by the following factors.

(A) Size of the Atom : Ionisation energy decreases with increase in 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 continuously in a group from top to bottom with increase in atomic size. But in 13th & 14th group the ionisation energy does not decrease continuously and order is as follows:

For 13th group: B > Tl > Ga > Al > In

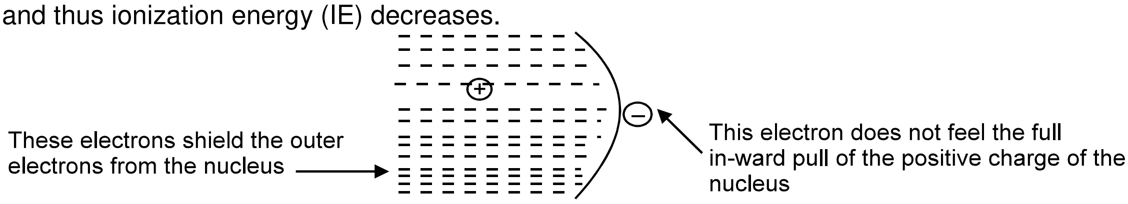
For 14th group: C > Si > Ge > Pb > Sn

(B) Nuclear Charge : 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 electron which is removed. 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 has extra stability.

The removal of an electron from such an atom requires more energy then expected. For example, first ionisation energy of beryllium is greater than boron because beryllium has extra stable outer most 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 gaseous phase increases as the value of ionisation energy decreases. Amongst alkali metals, the lithium is strongest reducing agent in aqueous solution.

Solved Examples

Ex-6 From each set, choose the atom which has the largest ionization enthalpy and explain your answer with suitable reasons. (a) F, O, N (b) Mg, P, Ar

Sol. (a) Fluorine (F) has the largest ionization enthalpy because in moving from left to right in a period, atomic size decreases and electrons are held more tightly. Since F has the smallest size and maximum nuclear charge. It has the largest ionization enthalpy among these elements.
(b) Argon (Ar) has the largest ionization enthalpy as argon has extra stable fully filled configuration.

Ex-7 First and second ionisation energies of magnesium are 7.646 eV and 15.035 eV respectively. The amount of energy in kJ needed to convert all the atoms of magnesium into Mg^{2+} ions present in 12 mg of magnesium vapour will be ? [Given 1 eV = 96.5 kJ ml^{-1}].

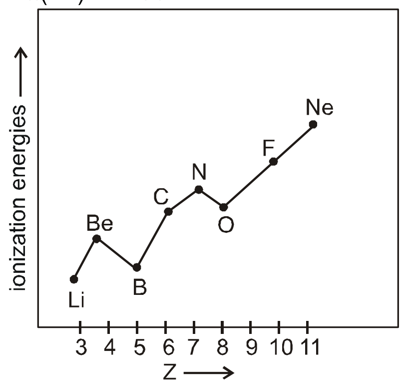
Sol. Total energy needed to convert one Mg atom into Mg^{2+} gas ion,

$= IE_I + IE_{II} = 22.681$ eV $= 2188.6$ kJ mol^{-1}.

$\Rightarrow$ 12 mg of Mg $= 0.5 \times 10^{-3}$ mole.

$\therefore$ Total energy $= 0.5 \times 10^{-3} \times 2188.6 = 1.0943$ kJ **Ans.**

Ex-8 Following graph shows variation of ionization energies with atomic number in second period (Li–Ne). Value of ionization energies of Na(11) will be :



(A) above Ne (B) below Ne but above O.
(C) below Li (D) 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 (C) is correct.

Ex-9 $M(g) \longrightarrow M^+(g) + e^-$; $\Delta H = 100$ eV.
$M(g) \longrightarrow M^{2+}(g) + 2e^-$; $\Delta H = 250$ eV.
Which is/are correct statement(s) ?
(A) IE_1 of M(g) is 100 eV (B) IE_1 of M^+ (g) is 150 eV.
(C) IE_2 of M(g) is 250 eV. (D) IE_2 of M(g) is 150 eV.

Sol. $M(g) \longrightarrow M^+(g) + e^-$ IE_1 of M
$M^+ \longrightarrow M^{2+} + e^-$ IE_2 of M but IE_1 of M^+
$M \longrightarrow M^{2+} + 2e^-$ $(IE_1 + IE_2)$. **Ans.** (A,B,D)

Section (F) : Electron gain enthalpy

(6) Electron Affinity:

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

$$F + e^- \longrightarrow F^- \quad E_a = 328 \text{ kJ mol}^{-1}$$

A positive electron affinity 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.

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.

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

Now a days both electron affinity and electron gain enthalpy terms are used.

Both are same in magnitude but opposite in sign.

E.A. = +ve (Exothermic):– $\Delta_{eg}H^\ominus < 0$.

E.A. = –ve (Endothermic):– $\Delta_{eg}H^\ominus > 0$.

The units of EA and Electron gain enthalpy are ev/atom or Kcal / mole or KJ / Mole.

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.

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

- 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.

In general the electron gain enthalpies for some third period element (e.g. P, S, Cl) are more negative than the corresponding second period members (e.g. N, O, F). This is due to the smaller size of the atoms of the second period elements which would produced larger electron-electron repulsions for the additional electron.

Group 1	$\Delta_{e.g.}$ H	Group 16	$\Delta_{e.g.}$ H	Group 17	$\Delta_{e.g.}$ H	Group 0	$\Delta_{e.g.}$ H
H	–73					He	+48
Li	–60	O	–141	F	–328	Ne	+116
Na	–53	S	–200	Cl	–349	Ar	+96
K	–48	Se	–195	Br	–325	Kr	+96
Rb	–47	Te	–190	I	–295	Xe	+77
Cs	–46	Po	–174	At	–270	Rn	+68

- 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 $\propto \dfrac{1}{\text{Atomic size}}$

  (ii) Electron affinity $\propto$ Effective nuclear charge (z_{eff})

  (iii) Electron affinity $\propto \dfrac{1}{\text{Screening effect}}$

  (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-10 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-11 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-12 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] ?

(A) O^x (B) O^y (C) O^z (D) 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-13 The amount of energy when million atoms of iodine are completely converted into I^- ions in the vapour state according to the equation, I (g) + e^- (g) $\rightarrow$ I^- (g) is 5.0×10^{-13} J.

Calculate the electron gain enthalpy of iodine in terms of kJ mol^{-1} and eV per atom.

Sol. The electron gain enthalpy of iodine is equal to the energy released when 1 mole of iodine atoms in vapour state are converted into I^- ions.

$$= -\frac{5.0\times10^{-13}\times6.023\times10^{23}}{10^6} = -30.1 \times 10^4 \text{ J} = -301 \text{ kJ.}$$

Electron gain enthalpy of iodine in eV per atom = $\dfrac{-301}{96.5} = -3.12.$

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.

Section (G) : Electronegativity

(7) 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%

$\xrightarrow{\hspace{6cm}}$

Electronegativity increases

Electronegativity of some elements according to pauling scale.

Elements	H	Li	Be	B	C	N	O	F	Ne	P	S	Cl	Br	I
Electronegativity	2.1	1.0	1.5	2.0	2.5	3.0	3.5	4.0	0.9	2.1	2.5	3.0	2.8	2.5

Table-6

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-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 is related 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.]*

* Only for reference, student need not memorize it.

APPLICATION OF ELECTRONEGATIVITY :

(a) Nomenclature :

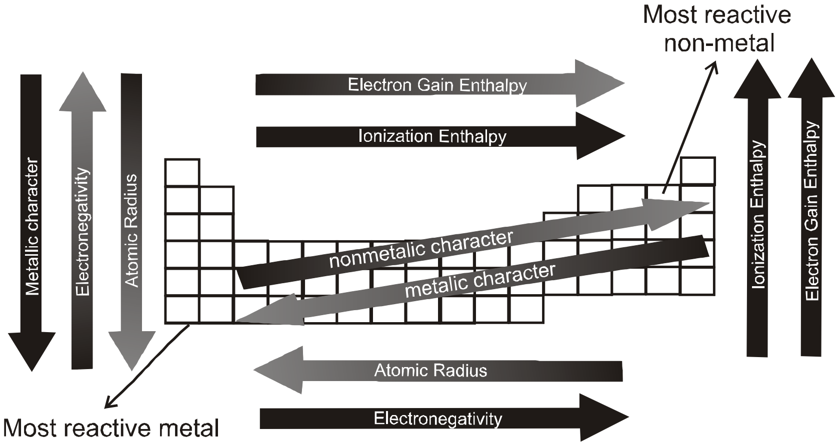
Compounds formed from two nonmetals are called binary compounds. Name of more electronegative element is written at the end and 'ide' is suffixed to it. The name of less electronegative element is written before the name of more electronegative element of the formula. For example -

Correct formula	Name
(a) $I^+ Cl^-$	Iodine chloride
(b) $Cl^+ F^-$	Chlorine fluoride
(c) $Br^+ Cl^-$	Bromine chloride
(d) IBr	Iodine bromide
(e) OF_2	Oxygen difluoride
(f) Cl_2O	Dichlorine oxide

(b) Effect on bond length: When the Δ EN increases between two atoms in a molecule then bond length between that atoms decreases.

Order of Δ EN:– HF > HCl > HBr > HI

Order of bond length:– HF < HCl < HBr < HI



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 ?

Ans. Chromium

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

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

Sol. Be^{2+} is smaller than Mg^{2+} as Be^{2+} has one shell where as Mg^{2+} has two shells.

Mg^{2+} and Na^+ are isoelectronic species : Ionic radius $\propto$ 1/nuclear charge.

Cl^- and S^{2-} are isoelectronic species : Ionic radius $\propto$ 1/nuclear charge.

Cl^- is smaller than Br^- as Cl^- has three shells where as Br^- has four shells.

5. 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 ?

Ans. (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).

6. 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 (χ_M) = $\dfrac{\text{Ionisation energy} + \text{Electron affinity}}{2}$

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

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

Ans. $\chi_P = 3.726$

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

Sol. 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.

8. 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	
(A)	ns^2, np^5	(p)	Chromium
(B)	$(n-1)\ d^{10}, ns^1$	(q)	Copper
(C)	$(n-1)\ d^5, ns^1$	(r)	Krypton
(D)	$(n-1)\ d^{10}, ns^2, np^6$	(s)	Bromine

Ans. (A) $\rightarrow$ (s) ; (B) $\rightarrow$ (q) ; (C) $\rightarrow$ (p) ; (D) $\rightarrow$ (r).

Sol. (A) ns^2np^5 is general valence shell electron configuration of halogens. So this configuration belongs to bromine.
(B) $(n-1)\ d^{1-10}\ ns^{1-2}$; This is electron configuration of d-block elements. As it contains $(n-1)\ d^{10}ns^1$ configuration it belongs to copper.
(C) $(n-1)\ d^{1-10}\ ns^{1-2}$; This is electron configuration of d-block elements. As it contains $(n-1)\ d^5ns^1$ configuration it belongs to chromium.
(D) Noble gases has valence shell electron configuration $ns^2\ np^6$, so it belongs to krypton.

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

Column-I		Column-II	
(A)	Metalloid	(p)	Sulphur
(B)	Radioactive	(q)	Gold
(C)	Transition metal	(r)	Arsenic
(D)	Chalcogen	(s)	Uranium

Ans. (A) $\rightarrow$ (r) ; (B) $\rightarrow$ (s) ; (C) $\rightarrow$ (q) ; (D) $\rightarrow$ (p)

Sol. (A) 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).
(B) Uranium is a radioactive element.
(C) 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 $[Xe]^{54}, 5d^86s^0$.
(D) 16th group elements like oxygen and sulphur are ore forming elements and therefore are called as chalcogens.

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

Column-I		Column-II	
(A)	Representative element	(p)	Cerium
(B)	Lanthanide	(q)	Aluminum
(C)	Coinage metal	(r)	Thorium
(D)	Actinide	(s)	Gold

Ans. (A) $\rightarrow$ (q) ; (B) $\rightarrow$ (p) ; (C) $\rightarrow$ (s) ; (D) $\rightarrow$ (r)

Sol. (A) s-block and p-block elements are collectively called as representative elements. As in aluminium last electron enters in p-subshell ($[Ne]^{10}3s^23p^1$).
(B) Lanthanide series follows lanthanum (atomic number 57) and starts from cerium (atomic number 58) to lutetium (atomic number 71), fourteen 4f- series elements.
(C) Group 11- transition elements copper, silver & gold are known as coinage metals (used for making the coins).
(D) Actinides series follows actinium (atomic number 89) and starts from thorium (atomic number 90) to lawrencium (atomic number 103), fourteen 5f- series elements.

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

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

Ans. (A) $\rightarrow$ (q) ; (B) $\rightarrow$ (s) ; (C) $\rightarrow$ (p) ; (D) $\rightarrow$ (r)

Sol. (A) 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.
(B) np^3, $(n–1)\ d^5$ and $(n–2)\ f^7$ represent half filled orbitals. Antimony has ($[Kr]^{36}\ 4d^{10}5s^25p^3$).
(C) The energy required to remove an electron from an univalent cation(g) is called second ionisation energy.
(D) 4f and 5f- series elements are called inner transition elements because they have three outer most shells incomplete.

12. 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**	
(A)	Highest 1st ionisation energy	(p)	Technitium
(B)	Highest electronegativity	(q)	Lithium
(C)	Synthetic element	(r)	Helium
(D)	Strongest reducing agent	(s)	Fluorine

Ans. (A) $\rightarrow$ (r) ; (B) $\rightarrow$ (s) ; (C) $\rightarrow$ (p) ; (D) $\rightarrow$ (q).

Sol. (A) 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.
(B) Fluorine has highest electronegativity i.e. 4.0 on Pauling scale on account of its small size.
(C) Technitium is a man made element.
(D) Lithium is a strongest reducing agent because of its highest negative value of E° due to its higher hydration energy on account of its small size of atom.

13. 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 $[Ar]4s^23d^{10}$	(s)	17th group

Ans. (A) $\rightarrow$ (r) ; (B) $\rightarrow$ (s) ; (C) $\rightarrow$ (p, q) ; (D) $\rightarrow$ (q).

Sol. (A) $[Ar]3d^{10}4s^24p^2$: Fourth shell contains two electron in 4p-sub shell i.e., $4p^2$. Therefore, group number = 10 + 4 = 14.
(B) Halogens (i.e. group number 17) have valence shell electronic configuration ns^2np^5 and there is one unpaired electron in p-subshell i.e., $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$
(C) The element in which last electron enters in d-subshell belongs to d-block. For d-block elements the group number = number of electrons in valence shell + number of electrons in (n – 1) d-subshell.
Group number 8. Valence shell electronic configuration is $ns^2(n-1)d^6$. Therefore, group number = 2 + 6 = 8.
Like wise, group 12 is $ns^2(n-1)d^{10}$. Therefore, group number = 2 + 10 = 12.
So in group 8 and 12 last electron enters in d-subshell.
(D) For electronic configuration. [Ar] $4s^23d^{10}$ the group number = 2 + 10 = 12.

Exercise-1

Marked questions are recommended for Revision.

PART - I : SUBJECTIVE QUESTIONS

Section (A) : Development of Periodic Table & Modern Periodic Table

A-1. Explain the following :
(i) Why argon (atomic mass = 39.94) has been placed before potassium (atomic mass = 39.10) in the Modern periodic table ?
(ii) There are only 14 lanthanides and only 14 actinides in Modern periodic table.

A-2. Why the third period of Modern periodic table contains 8 elements and not 18?

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

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

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

Section (C) : Oxidation states & Inert pair effect

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

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

Section (D) : Atomic and Ionic radius

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

D-2. The atomic radii of palladium and platinum are nearly same. Why ?

D-3. 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 (E) : Ionisation energy

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

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

E-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.

Section (F) : Electron gain enthalpy

F-1. Be and Ne have positive values of electron gain enthalpy against the general trend in their period in Modern periodic table. Explain.

F-2. Nitrogen has positive electron gain enthalpy whereas oxygen has negative. However, oxygen has lower ionisation enthalpy than nitrogen. Explain.

Section (G) : Electronegativity

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

G-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.

PART - II : ONLY ONE OPTION CORRECT TYPE

Section (A) : Development of Periodic Table & Modern Periodic Table

A-1. The period number in the long form of the periodic table is equal to :
(A) magnetic quantum number of any element of the period.
(B) atomic number of any element of the period.
(C) maximum Principal quantum number of any element of the period.
(D) maximum Azimuthal quantum number of any element of the period.

A-2. Which one of the following statements related to the modern periodic table is **incorrect** :
(A) The p-block has 6 columns, because a maximum of 6 electrons can occupy all the orbitals in a p-subshell.
(B) The d-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a d-subshell.
(C) Each block contains a number of columns equal to the number of electrons that can occupy that subshell.
(D) The block indicates value of Azimuthal quantum number (ℓ) for the last subshell that received electrons in building up the electronic configuration.

A-3. The elements in which electrons are progressively filled in 4f-orbital are called :
(A) actinoids (B) transition elements
(C) lanthanoids (D) halogens

A-4. Which of the following statements is not correct regarding hydrogen :
(A) It resembles halogens in some properties.
(B) It resembles alkali metals in some properties.
(C) It can be placed in 17th group of Modern periodic table.
(D) It cannot be placed in 1st group of Modern periodic table.

A-5. 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 :
(A) 29, 65 (B) 39, 79 (C) 29, 79 (D) 39, 65

A-6. In modern periodic table, the element with atomic number Z = 118 will be :
(A) Uuo ; Ununoctium ; alkaline earth metal
(B) Uno ; Unniloctium ; transition metal
(C) Uno ; Unniloctium ; alkali metal
(D) Uuo ; Ununoctium ; noble gas

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

B-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

B-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).

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

B-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 (C) : Oxidation states & Inert pair effect

C-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

C-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

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

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

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

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

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

C-8. Thallium shows different oxidation states because :
(A) of its high reactivity (B) of inert pair of electrons
(C) of its amphoteric nature (D) its is a transition metal

C-9. In which of the following elements, + 3 oxidation state is more stable than + 5 ?
(A) P (B) As (C) N (D) Bi

C-10. Which of the following is correct order of stability :
(A) $Tl^{3+} > Bi^{3+}$ (B) $PbO_2 > PbO$ (C) $BiI_5 < BiF_5$ (D) $Sn^{2+} = Ge^{2+}$

Section (D) : Atomic and Ionic radius

D-1. Select correct statement about radius of an atom :
(A) Values of Vander waal's radii is 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 is 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

D-2. 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

D-3. 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

D-4. 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.

D-5. 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 ≈ Ta (D) Li > Be > B

Section (E) : Ionisation energy

E-1. 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.

E-2. 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

E-3. 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

E-4. 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

E-5. 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 (F) : Electron gain enthalpy

F-1. 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

F-2. 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

F-3. 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

F-4. 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)

F-5. 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

Section (G) : Electronegativity

G-1. 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

G-2. 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.

G-3. 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

G-4. 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)

G-5.✍ Which of the following is most electronegative element.
(A) Li (B) Mg (C) H (D) Na

PART - III : MATCH THE COLUMNS

1. Match the column.

	Column-I (Atomic number)		Column-II
(A)	57	(p)	is d-Block or p-Block element
(B)	17	(q)	is 4th period element
(C)	19	(r)	is violates Aufbou's principle element
(D)	29	(s)	is non metal
		(t)	is s-Block element

2. Match the column.

	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 are recommended for Revision.

PART - I : ONLY ONE OPTION CORRECT TYPE

1. The statement that is **not** correct for periodic classification of elements in Modern periodic table is :
(A) The properties of elements are periodic function of their atomic numbers.
(B) Non-metallic elements are less in number than metallic elements.
(C) For transition elements, the 3d-orbitals are filled with electrons after 3p-orbitals and before 4s-orbitals.
(D) The first ionisation enthalpies of elements generally increase with increase in atomic number as we go along a period.

2. Which of the following is true about the element $_{33}As$ according to Modern periodic table :
(A) It is a 5th period element. (B) It is a p-block element.
(C) It belongs to 16th group. (D) It is one among typical elements.

3.✍ Which of the following contains atomic number of only s-block
(A) 55,12,18,53 (B) 13,33,54,83 (C) 3, 20, 55, 87 (D) 22,33,55,66

4.✍ Screening effect is not observed in :
(A) He^+ (B) Li^{2+} (C) Be^{3+} (D) In all cases

5. Which of the following have higher Z_{eff} than Fluorine.
(A) Cl (B) O (C) F^- (D) none of these

6. The oxidation number that iron does not exhibit in its common compounds or in its elemental state is :
(A) 0 (B) +1 (C) +2 (D) +3

7. Which of the following can show +7 oxidation state?
(A) Mn (B) F (C) In (D) N

8. Which of following does not exist :
(A) TlI_3 (B) PbF_4 (C) Both (A) and (B) (D) None of these

9. Elements of which period show maximum inert pair effect :
(A) 3 (B) 4 (C) 5 (D) 6

10. 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^-

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

12. 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

13. 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

14. Fluorine has the highest electronegativity among the ns^2np^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.

15. 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

16. 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.

PART - II : SINGLE AND DOUBLE VALUE INTEGER TYPE

1. 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.

2. 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.

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

4. How many of the following have greater Z_{eff} than Silicon atom :
(i) Na (ii) Mg (iii) Al (iv) P (v) Cl
(vi) S (vii) N (viii) O (ix) F

5. The most stable oxidation state of chromium is +n, Give the value of 'n'.

6. How many of the following compounds are found to exist?
(i) BiF_5 (ii) $T\ell I_3$ (iii) PbO_2 (iv) $SnCl_2$
(v) $T\ell_2O_3$ (vi) PbI_4 (vii) As_2O_3

7. The Lanthanides are characterized by the uniform [+n] oxidation state shown by all the Lanthanides. What is the value of 'n'.

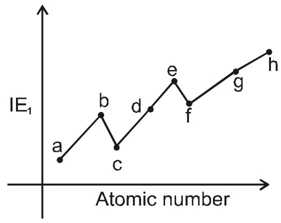
8. Highest oxidation states shown by Chromium & Manganese are +x & +y respectively. Give the value of x + y ?

9. 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].

10. Report atomic number of the element having largest size among the following :
Ni, Cu, Zn

11. How many of following atoms have maximum ionization energy than boron.
(i) Be (ii) N (iii) P (iv) Ga (v) S (vi) Mg

12. 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.



13. 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.

14. $A^-(g) \rightarrow A^{2+}(g)$ $\Delta H = 1100$ 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 ?

15. 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)

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

17. How many of the following elements are more electronegative than Boron.
(i) H (ii) Li (iii) Be (iv) C (v) N
(vi) O (vii) F

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

1. The group in modern periodic table in which all the elements do not have same number of electrons in their outermost shell is (considering upto 6th period) :
(A) 13th (B) 11th (C) 9th (D) 18th

2. Element corresponding to which of these/this atomic number belongs to p-block in Modern Periodic Table :
(A) 19 (B) 35 (C) 53 (D) 83

3. Which of the following have greater Z_{eff} than Zn :
(A) Cu^+ (B) Cu^{2+} (C) Fe^{3+} (D) Zn^{2+}

4. Which of the following is/are correct regarding oxidation state of elements in their compounds :
(A) All d-Block elements show multiple oxidation state.
(B) All p-Block elements show multiple oxidation state.
(C) All s-Block elements show single oxidation state.
(D) Some of 18 group elements can show multiple oxidation state.

5. Which of the following elements have + 3 as most popular oxidation state?
(A) Al (B) Xe (C) Cu (D) Sc

6. Which of the following show non-zero multiple oxidation state ?
(A) S (B) O (C) Zn (D) H

7.✍ Which of the following pairs of elements show similar set of oxidation state ?
(A) O^{16}, O^{18} (B) Na, K (C) C, Be (D) Zn, Rb

8. Which of the following elements have their lower oxidation state as more stable oxidation state.
(A) O (B) Pb (C) $T\ell$ (D) Bi

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

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

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

12.✍ 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

13. 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.

14. 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)

15.✍ 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

16. 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

17. 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

In the modern periodic table, elements are arranged in order of increasing atomic numbers which is related to the electronic configuration. Depending upon the type of orbitals receiving the last electron, the elements in the periodic table have been divided into four blocks, viz, s, p, d and f. The modern periodic table consists of 7 periods and 18 groups. Each period begins with the filling of a new energy shell. In accordance with the Arfbau principle, the seven periods (1 to 7) have 2, 8, 8, 18, 18, 32 and 32 elements respectively. The seventh period is still incomplete. To avoid the periodic table being too long, the two series of f-block elements, called lanthanoids and actinoids are placed at the bottom of the main body of the periodic table.
Now answer the following five questions :

1. The element with atomic number 57 belongs to :
(A) s-block (B) p-block (C) d-block (D) f-block

2. The last element of the p-block in 6th period is represented by the outermost electronic configuration :
(A) $7s^27p^6$ (B) $5f^{14}6d^{10}7s^27p^0$ (C) $4f^{14}5d^{10}6s^26p^6$ (D) $4f^{14}5d^{10}6s^26p^4$

3. Which of the elements, whose atomic numbers are given below, cannot be accommodated in the present set up of the long form of the periodic table ?
(A) 107 (B) 118 (C) 126 (D) 102

4. The electronic configuration of the element which is just above the element with atomic number 43 in the same group is ________:
(A) $1s^22s^22p^63s^23p^63d^54s^2$ (B) $1s^22s^22p^63s^23p^63d^54s^34p^6$
(C) $1s^22s^22p^63s^23p^63d^64s^2$ (D) $1s^22s^22p^63s^23p^63d^74s^2$

5. The elements with atomic numbers 35, 53 and 85 are all ________ :
(A) noble gases (B) halogens (C) heavy metals (D) light metals

Comprehension # 2

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 :

6.✍ 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) $r_{Metallic} > r_{Covalent} > r_{Vanderwaa}$

7. 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) $S^{2–} > Cl^– > Ca^{2+} > K^+$

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

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.

9. 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

10.✍ 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

11. 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.

Comprehension # 4
Answer Q.12, Q.13 and Q.14 by appropriately matching the information given in the three columns of the following table.

Column-1		Column-2		Column-3	
(I)	Graphite	(i)	d-block elements	(P)	Liquid
(II)	Transition elements	(ii)	Group-16	(Q)	$6s^2 6p^4$
(III)	Amalgam	(iii)	Allotropicity	(R)	Lubricant
(IV)	Polonium	(iv)	Mercury	(S)	Variable oxidation number.

12. For given content is column-1, the correct combination is :
(A) (I), (iii), R (B) (II), (iv), R (C) (II), (iii), S (D) (IV), (iv), Q

13. For iron the correct combination is :
(A) (III), (iv), Q (B) (II), (i), S (C) (IV), (i), Q (D) (I), (ii), P

14. The incorrect combination is :
(A) (III), (iv), P (B) (III), (i), S (C) (II), (ii), S (D) (IV), (ii), Q

Exercise-3

* **Marked Questions may have more than one correct option.**

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

1. The incorrect statement among the following is : **[JEE- 1997(Cancelled), 2/200]**
(A) the first ionization energy of Al is less than first ionization energy of Mg.
(B) the second ionization energy of Mg is greater than second ionization energy of Na.
(C) the first ionization energy of Na is less than first ionization energy of Mg.
(D) the third ionization energy of Mg is greater than third ionization energy of Al.

2. Arrange the following ions in order of their increasing size : Li^+, Mg^{2+}, K^+, Al^{3+}. **[JEE-1997, 1/100]**

3. Fill in the blanks :
Compounds that formally contain Pb^{4+} are easily reduced to Pb^{2+}. The stability of the lower oxidation state is due to ....

4. **Assertion :** F atom has a less negative electron affinity than Cl atom. **[JEE-1998, 2/200]**
Reason : Additional electrons are repelled more effectively by 3p electrons in Cl atom than by 2p electrons in F atom.
(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.

5. Ionic radii of : **[JEE-1999, 3/200]**
(A) $Ti^{4+} < Mn^{7+}$ (B) $^{35}Cl^- < {}^{37}Cl^-$ (C) $K^+ > Cl^-$ (D) $P^{3+} > P^{5+}$

6. 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) $Fe^{3+} < Fe^{2+} < Fe^{+4}$

7. **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.

8. 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

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

10. **Statement-1 :** Pb^{4+} compounds are stronger oxidizing agents than Sn^{4+} compounds **[JEE-2008, 3/163]**
Statement-2 : The higher oxidation states for the group 14 elements are more stable for the heavier members of the group due to 'inert pair effect'.
(A) Statement-1 is True, Statement-2 is True; Statement-2 is a correct explanation for Statement-1.
(B) Statement-1 is True, Statement-2 is True; Statement-2 is NOT a correct explanation for Statement-1
(C) Statement-1 is True, Statement-2 is False
(D) Statement-1 is False, Statement-2 is True.

11. 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, T i

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

JEE(MAIN) OFFLINE PROBLEMS

1. Which one of the following ions has the highest value of ionic radius ? **[AIEEE-2004, 3/225]**
(1) Li^+ (2) B^{3+} (3) O^{2-} (4) F^-

2. The formation of the oxide ion $O^{2-}_{(g)}$ requires first an exothermic and then an endothermic step as shown below :
$O_{(g)} + e^- = O^-_{(g)}$; $\Delta H° = -142$ kJmol^{-1}
$O^-_{(g)} + e^- = O^{2-}_{(g)}$; $\Delta H° = 844$ kJmol^{-1}
This is because : **[AIEEE-2004, 3/225]**
(1) oxygen is more electronegative.
(2) oxygen has high electron affinity.
(3) O^- ion will tend to resist the addition of another electron.
(4) O^- ion has comparatively larger size than oxygen atom.

3. In which of the following arrangements the order is NOT according to the property indicated against it ?
[AIEEE-2005, 3/225]
(1) $Al^{3+} < Mg^{2+} < Na^+ < F^-$ – increasing ionic size
(2) B < C < N < O – increasing first ionisation enthalpy
(3) I < Br < F < Cl – increasing electron gain enthalpy (with negative sign)
(4) Li < Na < K < Rb – increasing metallic radius

4. Which of the following factors may be regarded as the main cause of lanthanide contraction ?
(1) Greater shielding of 5d electrons by 4f electrons. **[AIEEE 2005, 4½ / 225]**
(2) Poorer shielding of 5d electron by 4f electrons.
(3) Effective shielding of one of 4f electrons by another in the sub-shell.
(4) Poor shielding of one of 4f electron by another in the sub-shell.

5. The lanthanide contraction is responsible for the fact that : **[AIEEE-2005, 3/225]**
(1) Zr and Y have about the same radius (2) Zr and Nb have similar oxidation state
(3) Zr and Hf have about the same radius (4) Zr and Zn have same oxidation state.

6. 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

7. 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

8. 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) $SiX_2 << GeX_2 << PbX_2 << SnX_2$

9. 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) $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$

10. 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) $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

11. 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) $4f^7\ 5d^1\ 6s^2$

12. 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:
(1) F > Cl > Br > I (2) Cl > F > Br > I (3) Br > Cl > I > F (4) I > Br > Cl > F

13. 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) $K^+, S^{2-}, Ca^{2+}, Cl^-$

14. Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar ?
(1) Ca < S < Ba < Se < Ar (2) S < Se < Ca < Ba < Ar **[JEE(Main)-2013, 4/120]**
(3) Ba < Ca < Se < S < Ar (4) Ca < Ba < S < Se < Ar

15. The first ionisation potential of Na is 5.1 eV. The value of electron gain enthalpy of Na^+ will be :
[JEE(Main)-2013, 4/120]
(1) –2.55 eV (2) –5.1 eV (3) –10.2 eV (4) +2.55 eV

16. The ionic radii (in Å) of N^{3-}, O^{2-} and F^- are respectively : **[JEE(Main)-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

17. Which of the following atoms has the highest first ionization energy? **[JEE(Main)-2016, 4/120]**
(1) Na (2) K (3) Sc (4) Rb

18. 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) $O^{2-}, F^-, Na^+, Mg^{2+}$

JEE(MAIN) ONLINE PROBLEMS

1. Which of the following series correctly represents relation between the elements from X to Y ?
$X \rightarrow Y$ **[JEE(Main) 2014 Online (11-04-14), 4/120]**
(1) $_3Li \rightarrow {}_{19}K$ Ionization enthalpy increses (2) $_9F \rightarrow {}_{35}Br$ Electron gain enthalpy
(3) $_6C \rightarrow {}_{32}Ge$ Atomic radii increases (4) $_{18}Ar \rightarrow {}_{54}Xe$ Noble character increases

2. Similarity in chemical properties of the atoms of elements in a group of the periodic table is most closely related to : **[JEE(Main) 2014 Online (12-04-14), 4/120]**
(1) atomic numbers (2) atomic masses
(3) number of principal energy levels (4) number of valence electrons

3. 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) $N^{3-} < S^{2-} < O^{2-} < P^{3-}$

4. 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^{-}

5. In the long form of the periodic table, the valence shell electronic configuration of $5s^25p^4$ 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

6. The following statements concern elements in the periodic table. Which of the following is true? **[JEE(Main) 2016 Online (10-04-16), 4/120]**
(1) The Group 13 elements are all metals.
(2) All the elements in Group 17 are gases.
(3) Elements of Group 16 have lower ionization enthalpy values compared to those of Group 15 in the corresponding periods.
(4) For Group 15 elements, the stability of +5 oxidation state increases down the group.

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

Element	Ionization enthalpy (kJ/mol)		
	1^{st}	2^{nd}	3^{rd}
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'.

8. 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] $3d^{10} 4s^2 4p^3$

9. For Na^{+}, Mg^{2+}, F^{-} and O^{2-} ; the correct order of increasing ionic radii is : **[JEE(Main) 2018 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) $Mg^{2+} < O^{2-} < Na^{+} < F^{-}$

10. The correct order of electron affinity is : **[JEE(Main) 2018 Online (15-04-18), 4/120]**
(1) F > Cl > O (2) F > O > Cl (3) Cl > F > O (4) O > F > Cl

11. Aluminium is usually found in +3 oxidation state. In contrast, thalium exists in +1 and +3 oxidation states. This is due to : **[JEE(Main) 2019 Online (09-01-19), 4/120]**
(1) inert pair effect (2) lanthanoid contraction
(3) diagonal relationship (4) lattice effect

12. 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

13. When the first electron gain enthalpy ($\Delta_{eg}H$) of oxygen is –141 kJ/mol, its second electron gain enthalpy is : **[JEE(Main) 2019 Online (09-01-19), 4/120]**
(1) almost the same as that of the first (2) negative, but less negative than the first
(3) a more negative value than the first (4) a positive value

14. The effect of lanthanoid contraction in the lanthanoid series of elements by an and large means :
[JEE(Main) 2019 Online (10-01-19), 4/120]
(1) increase in atomic radii and decrease in ionic radii
(2) decrease in both atomic and ionic radii
(3) increase in both atomic and ionic radii
(4) decrease in atomic radii and increase in ionic radii

15. The electronegativity of aluminium is similar to : **[JEE(Main) 2019 Online (10-01-19), 4/120]**
(1) Lithium (2) Carbon (3) Boron (4) Beryllium

16. The correct order of the atomic radii of C, Cs, Al, and S is:
[JEE(Main) 2019 Online (11-01-19), 4/120]
(1) C < S < Al <Cs (2) S < C < Al < Cs
(3) S < C < Cs < Al (4) C < S < Cs < Al

17. The correct option with respect to the Pauling, electronegativity values of the elements is :
[JEE(Main) 2019 Online (11-01-19), 4/120]
(1) Te > Se (2) Ga < Ge (3) Si < Al (4) P > S

18. The element with Z = 120 (not yet discovered) will be an/a :
[JEE(Main) 2019 Online (12-01-19), 4/120]
(1) transition metal (2) alkali metal
(3) alkaline earth metal (4) inner-transition metal

Answers

EXERCISE - 1

PART - I

A-1. (i) In modern periodic table, elements have been placed in order of their increasing atomic numbers. The atomic number of argon is 18 and that of potassium is 19. Thus, argon has been placed before potassium.
(ii) In lanthanides and actinides, the differentiating electron enters to (n – 2) f-subshell. The maximum capacity of f-subshell is of 14 electrons. Thus, there are only 14 lanthanides ($4f^{1-14}$) and only 14 actinides ($5f^{1-14}$).

A-2. In the modern periodic table, each period starts with the filling of a new principal energy level. Thus, the third period begins with the filling of principal quantum number, n = 3. When n = 3, ℓ = 0, 1, 2. But according to Aufbau principle, the electrons are added to different orbitals in order of their increasing energies. Now, the energy of 3d-subshell is higher than that of 4s-subshell. Therefore, in third period, electrons can be filled in only 3s & 3p-subshells, whose energies increase in the order: 3s < 3p. Now, s-subshell has one and p-subshell has three orbitals. Hence, in all, there are 4 (1 + 3) orbitals that can be filled in this period. Since according to Pauli's exclusion principle, each orbital, at the maximum, can accommodate two electrons. Therefore, 4 orbitals, at the maximum, can have 8 electrons and hence, fourth period has 8 elements.

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

B-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.

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

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

D-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.

D-2. Due to lanthanide contraction (poor shielding of nuclear charge by 4f-electrons), atomic radii of 4d and 5d elements are nearly same.

D-3. 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.

E-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.

E-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.

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

F-1. 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.

F-2. 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).

G-1. Caesium (Cs).

G-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.

PART - II

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

PART - III

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

EXERCISE - 2

PART - I

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

PART - II

1.	2	**2.**	26	**3.**	19	**4.**	6 (except i, ii, iii)		
5.	3	**6.**	6 (except (vi))	**7.**	3	**8.**	13	**9.**	8
10.	30	**11.**	2 (i, ii)	**12.**	0	**13.**	2	**14.**	1
15.	23	**16.**	17	**17.**	5 (except ii, iii)				

PART - III

1.	(CD)	**2.**	(BCD)	**3.**	(ABCD)	**4.**	(CD)	**5.**	(AD)
6.	(ABD)	**7.**	(AB)	**8.**	(ABCD)	**9.**	(CD)	**10.**	(BC)
11.	(ABD)	**12.**	(ABCD)	**13.**	(ABD)	**14.**	(AD)	**15.**	(AB)
16.	(BC)	**17.**	(ACD)						

PART - IV

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

EXERCISE - 3

PART - I

1.	(B)	**2.**	$Al^{3+} < Mg^{2+} < Li^{+} < K^{+}$			**3.**	Inert Pair Effect		
4.	(C)	**5.**	(D)	**6.**	(B)	**7.**	(C)	**8.**	(B)
9.	(B)	**10.**	(C)	**11.**	2				

PART - II

JEE(MAIN) OFFLINE PROBLEMS

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

JEE(MAIN) ONLINE PROBLEMS

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

Additional Problems for Self Practice (APSP)

Marked questions are recommended for Revision.
This Section is not meant for classroom discussion. It is being given to promote self-study and self testing amongst the Resonance students.

PART - I : PRACTICE TEST-1 (IIT-JEE (MAINS Pattern))

Max. Time : 1 Hr. **Max. Marks : 120**

Important Instructions
1. The test is of **1 hour** duration.
2. The Test Booklet consists of **30** questions. The maximum marks are **120.**
3. Each question is allotted **4 (four)** marks for correct response.
4. Candidates will be awarded marks as stated above in Instructions No. 3 for correct response of each question. ¼ **(one fourth)** marks will be deducted for indicating incorrect response of each question. No deduction from the total score will be made if no response is indicated for an item in the answer sheet.
5. There is only one correct response for each question. Filling up more than one response in any question will be treated as wrong response and marks for wrong response will be deducted accordingly as per instructions 4 above.

1. The elements which exhibit both vertical and horizontal similarities are :
(1) inert gas elements (2) representative elements
(3) transition elements (4) none of these

2. Of the following pairs, the one containing examples of metalloid elements is :
(1) B and Al (2) Ga and Ge (3) Al and Si (4) As and Sb

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

4. Atomic number of 15, 33, 51 represents the following family :
(1) carbon family (2) nitrogen family (3) oxygen family (4) None of these

5. In a given energy level, the order of penetration effect of different orbitals is :
(1) $f < p < d < s$ (2) $s < p < d < f$ (3) $f < d < p < s$ (4) $s = p = d = f$

6. Which of the following is correct order of Z_{eff} :
(1) $I^- > I > I^+$ (2) $Mg^{2+} > Na^+ > F^-$ (3) $P^{5+} < P^{3+}$ (4) Li > Be >B

7. In Sodium atom on 3s electron the screening is due to :
(1) $3s^2, 3p^6$ (2) $4s^1$ (3) $1s^2, 2s^2, 2p^6$ (4) $3s^1$

8. Which of the following elements can have negative oxidation states.
(1) Al (2) Ca (3) Fe (4) B

9. What is correct order of reducing capacity :
(1) $Ge^{2+} > Sn^{2+} > Pb^{2+}$ (2) $Ge^{2+} < Sn^{2+} < Pb^{2+}$ (3) $Ge^{2+} \approx Sn^{2+} \approx Pb^{2+}$ (4) $Pb^{2+} > Ge^{2+} > Sn^{2+}$

10. Inert pair effect is observed in elements of which block :
(1) s (2) p (3) d (4) f

11. Which of the following order of radii is correct :
(1) Li < Be < Mg (2) $H^+ < Li^+ < H^-$ (3) O < F < Ne (4) Li < Na < K < Cs < Rb

12. The lanthanide contraction refers to :
(1) radius of the series. (2) valence electrons of the series.
(3) the density of the series. (4) electronegativity of the series.

13. Which of the following statement is correct with respect to the property of elements in the carbon family with an increase in atomic number, their :
(1) Atomic size increases
(2) Ionization energy increases
(3) Metallic character decreases
(4) Stability of +4 oxidation state increases

14. Which group of atoms have nearly same atomic radius :
(1) Na, K, Rb, Cs (2) Li, Be, B, C (3) Fe, Co, Mn (4) F, Cl, Br, I

15. The incorrect order of radius is :
(1) $Cu^- > Cu > Cu^+$ (2) $Sc^{3+} > K^+ > S^{2-}$ (3) Ni < Cu < Zn (4) All of these

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

17. A large difference between the third and fourth ionization energies indicates the presence of :
(1) 4 valence electrons in an atom
(2) 5 valence electrons in an atom
(3) 3 valence electrons in an atom
(4) 2 valence electrons in an atom

18. The ionization enthalpy will be highest when the electron is to be removed from .... if other factors are equal :
(1) s-orbital (2) p-orbital (3) d-orbital (4) f-orbital

19. 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.
(1) V (2) Cr (3) Mn (4) Fe

20. For which of the following species 2^{nd} IE < 1^{st} IE
(1) Be (2) Ne (3) Na^+ (4) None of these

21. With reference to 1^{st} IP which are correct.
(a) Li < C (b) O < N (c) Be < N < Ne
(1) a, b (2) b, c (3) a, c (4) a, b & व c

22. Values of 1^{st} four ionisaiton energies (kJ/mol) of an element are respectively 496, 4563, 6913, 9541; the electronic configuration of that element can be.
(1) $1s^2, 2s^1$ (2) $1s^2\ 2s^2\ 2p^1$ (3) $1s^2, 2s^2, 2p^6\ 3s^1$ (4) (2) and (3) both

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

24. For magnitude of electron gain enthalpy of chalcogens and halogens, which of the following options is correct?
(1) Br > F (2) S > F (3) O < Cl (4) S < Se

25. The correct order of electron gain enthalpy (most endothermic first and most exothermic last) is :
(1) Be < B < C < N (2) Be < N < B < C (3) N < Be < C < B (4) N < C < B < Be

26. $\frac{N_0}{2}$ 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.
(1) I.E. = $\frac{2E_1}{N_0}$, $\Delta_{eq}H = -\frac{E_2}{2N_0}$
(2) I.E. = $-\frac{E_2}{2N_0}$, $\Delta_{eq}H = \frac{2E_1}{N_0}$
(3) I.E. = $\frac{E_1}{2N_0}$, $\Delta_{eq}H = -\frac{E_2}{2N_0}$
(4) I.E. = $\frac{N_0}{2E_1}$, $\Delta_{eq}H = -\frac{2N_0}{E_2}$

27. The formation of the oxide ion, $O^{2-}(g)$, from oxygen atom requires first an exothermic and then an endothermic step as shown below :
$O(g) + e^- \longrightarrow O^-(g)$; $\Delta_{eg}H = -141$ kJmol^{-1}
$O^-(g) + e^- \longrightarrow O^{2-}(g)$; $\Delta_{eg}H = +780$ kJmol^{-1}
Thus process of formation of O^{2-} in gas phase is unfavourable even though O^{2-} is isoelectronic with neon. It is due to the fact that :
(1) oxygen is more electronegative.
(2) addition of electron in oxygen results in larger size of the ion.
(3) electron repulsion outweighs the stability gained by achieving noble gas configuration.
(4) O^- ion has comparatively smaller size than oxygen atom.

28. The properties which are not common to both groups 1 and 17 elements in the periodic table are :
(1) Elelctropositive character increase down the gorups.
(2) Reactivity decrease from top to bottom in these groups.
(3) Atomic radii increase as the atomic number increase.
(4) Electronegativity decrease on moving down a group.

29. The correct set of decreasing order of electronegativity is :
(1) Li, H, Na (2) Na, H, Li (3) H, Li, Na (4) Li, Na, H

30. Which of the following is most electronegative in p-block elements
(1) Oxygen (2) Chlorine (3) Fluorine (4) Phosphorus

Practice Test-1 (IIT-JEE (Main Pattern))

OBJECTIVE RESPONSE SHEET (ORS)

Que.	1	2	3	4	5	6	7	8	9	10
Ans.										
Que.	11	12	13	14	15	16	17	18	19	20
Ans.										
Que.	21	22	23	24	25	26	27	28	29	30
Ans.										

PART-II : NATIONAL STANDARD EXAMINATION IN CHEMISTRY (NSEC) STAGE-I

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_2 (C) K_2O_2 (D) D_2O_2

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

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

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

6. From the electronic configuration of the given element K, L, M and N, which one has the highest ionisation potential : **[NSEC-2001]**
(A) M = [Ne] $3s^2,3p^2$ (B) L = [Ne]$3s^1,3p^3$
(C) K = [Ne]$3s^2,3p^1$ (D) N = [Ar]$3d^{10},4s^2,4p^3$

7. 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

8. The outermost electron configuration of one of the element is $5f^2$, $6d^1$, $7s^2$. This element belongs to : **[NSEC-2002]**
(A) s-block (B) transition series (C) lanthanide series (D) actinide series.

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

10. Which oxyacid of chlorine shows oxidation state of + 5 ? **[NSEC-2002]**
(A) hypochlorous acid (B) chloric acid (C) chlorous acid (D) perchloric acid

11. Which element does not show positive oxidation state ? **[NSEC-2002]**
(A) fluorine (B) chlorine (C) oxygen (D) iodine.

12. Due to addition of electrons in d orbital for transition element, the screening effect **[NSEC-2002]**
(A) increases (B) decreases (C) no effect (D) slightly decreases.

13. The diagonal relationship of elements in the periodic table arises because of similarity in **[NSEC-2003]**
(A) ionic radius (B) electronic configuration
(C) crystal structure (D) charge/radius ratio of the corresponding ions.

14. 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.

15. 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.

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

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

18. Which of the following has the highest electron affinity ? **[NSEC-2005]**
(A) F (B) Br (C) Cl (D) I.

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

20. The group in the long form of periodic table having three elements together is **[NSEC-2005]**
(A) zero group (B) IIIrd group (C) IVth group (D) VIIIth group.

21.✍ Atom with the largest electron affinity is **[NSEC-2006]**
(A) Na (B) Cl (C) I (D) P.

22.✍ 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.

23. As the number of electrons in d-orbitals of transition elements increases, the screening effect on the valence electrons- **[NSEC-2007]**
(A) increases (B) decreases greatly (C) is not observed (D) decreases slightly

24. 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

25. The ion which has 18 electrons in the outermost shell is – **[NSEC-2009]**
(A) Cu^{+} (Z = 29) (B) Al^{3+} (Z = 13) (C) K^{+} (Z = 19) (D) Th^{4+} (Z = 90)

26. The correct order of the size of the species is **[NSEC-2010]**
(A) $Ca^{2+} < Ar < K^+ < Cl^-$ (B) $Ca^{2+} < K^+ < Ar < S^{2-}$
(C) $K^+ < Ar < Cl^- < S^{2-}$ (D) $Ar < Ca^{2+} < K^+ < Cl^-$

27. 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

28. 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

29. 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

30. The element with the lowest electronegativity is : **[NSEC-2011]**
(A) S (B) I (C) Ba (D) Al

31. 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

32. The outer most electronic configuration of the most electronegative element is : **[NSEC-2012]**
(A) ns^2, np^3 (B) ns^2,np^6 (n– 1) d^2 (C) ns^2, np^5 (D) ns^2,np^6

33. The first ionisation potential of Na, Mg, Al and Si are in the order : **[NSEC-2012]**
(A) Na < Mg > Al < Si (B) Na > Mg > Al > Si (C) Na < Mg < Al > Si (D) Na > Mg > Al < Si

34. The first four ionization energy values of a metal are 191,587,872 and 5962 kcal/mol respectively. The number of valence electrons in the element is : **[NSEC-2012]**
(A) 1 (B) 2 (C) 3 (D) 5

35. Of the following, the ion with the largest size is **[NSEC-2014]**
(A) O^{2-} (B) Na^+ (C) F^- (D) Al^{3+}

36. Which of the following accounts best for the fact that F^- is smaller than O^{2-} ? **[NSEC-2018]**
(A) F^- has a larger nuclear mass than O^{2-} (B) F^- has a larger nuclear charge than O^{2-}
(C) F^- is more polarizable than O^{2-} (D) F is more electronegative than O

PART - III : PRACTICE TEST-2 (IIT-JEE (ADVANCED Pattern))

Max. Time : 1 Hr. **Max. Marks : 69**

Important Instructions

A. General :
1. The test is of 1 hour duration.
2. The Test Booklet consists of 23 questions. The maximum marks are 69.

B. Question Paper Format :
3. Each part consists of five sections.
4. Section-1 contains 7 multiple choice questions. Each question has four choices (A), (B), (C) and (D) out of which ONE is correct.
5. Section-2 contains 6 multiple choice questions. Each question has four choices (A), (B), (C) and (D) out of which ONE OR MORE THAN ONE are correct.
6. Section-3 contains 6 questions. The answer to each of the questions is a single-digit integer, ranging from 0 to 9 (both inclusive).
7. Section-4 contains 1 paragraphs each describing theory, experiment and data etc. 3 questions relate to paragraph. Each question pertaining to a partcular passage should have only one correct answer among the four given choices (A), (B), (C) and (D).
8. Section-5 contains 1 multiple choice questions. Question has two lists (list-1 : P, Q, R and S; List-2 : 1, 2, 3 and 4). The options for the correct match are provided as (A), (B), (C) and (D) out of which ONLY ONE is correct.

C. Marking Scheme :

9. For each question in Section-1, 4 and 5 you will be awarded 3 marks if you darken the bubble corresponding to the correct answer and zero mark if no bubble is darkened. In all other cases, minus one (– 1) mark will be awarded.
10. For each question in Section-2, you will be awarded 3 marks. If you darken all the bubble(s) corresponding to the correct answer(s) and zero mark. If no bubbles are darkened. No negative marks will be answered for incorrect answer in this section.
11. For each question in Section-3, you will be awarded 3 marks if you darken only the bubble corresponding to the correct answer and zero mark if no bubble is darkened. No negative marks will be awarded for incorrect answer in this section.

SECTION-1 : (Only One option correct Type)

This section contains 7 Single correct questions. Each questions has four choices (A), (B), (C) and (D) out of which Only ONE option is correct.

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. Elements of which block in modern periodic table cannot have –ve oxidation state?
(A) s (B) d (C) p (D) None of these

4. 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

5. 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.

6. 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) $Ba^{2+} < Cs^+ < I^- < Te^{2-}$

7. 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.

Section-2 : (One or More than one options correct Type)

This section contains 6 multipole choice questions. Each questions has four choices (A), (B), (C) and (D) out of which ONE or MORE THAN ONE are correct.

8. 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.

9. 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.

10. Which of the following element(s) have only one non-zero oxidation state.
(A) Be (B) O (C) F (D) N

11. 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

12. 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) $Al^{2+}(g) \longrightarrow Al^{3+}(g) + e^-$

13. Which of the following has/have no unit ?
(A) Electronegativity
(B) Electron gain enthalpy
(C) Ionisation enthalpy
(D) Metallic character

Section-3 : (One Integer Value Correct Type.)
This section contains 6 questions. Each question, when worked out will result in one integer from 0 to 9 (both inclusive)

14. 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.

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

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

17. 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

18. 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.

19. How many elements are more electropositive than Cl.
B, N, O, C, S, P, At, H, Li

SECTION-4 : Comprehension Type (Only One options correct)
This section contains 1 paragraphs, each describing theory, experiments, data etc. 3 questions relate to the paragraph. Each question has only one correct answer among the four given options (A), (B), (C) and (D)

Paragraph for Questions 20 to 22

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

$\xrightarrow{\text{At no. increase}}$

–8(a)	141(e)	328(i)
–72(b)	200(f)	349(j)
–78(c)	193(g)	325(k)
–103(d)	190(h)	295(l)

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

20. Select the correct order of atomic radius :
(A) a < b < c < d (B) a < e < i (C) i > j > k > ℓ (D) e > f > g

21. 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

22. 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

SECTION-5 : Matching List Type (Only One options correct)
This section contains 1 questions, each having two matching lists. Choices for the correct combination of elements from List-I and List-II are given as options (A), (B), (C) and (D) out of which one is correct

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

Practice Test-2 ((IIT-JEE (ADVANCED Pattern))

OBJECTIVE RESPONSE SHEET (ORS)

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
Ans.										
Que.	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
Ans.										
Que.	**21**	**22**	**23**							
Ans.										

APSP Answers

PART - I

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

PART - II

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

PART - III

1.	(C)	**2.**	(A)	**3.**	(A)	**4.**	(B)	**5.**	(C)
6.	(A)	**7.**	(B)	**8.**	(ABC)	**9.**	(AD)	**10.**	(AC)
11.	(ACD)	**12.**	(ABCD)	**13.**	(AD)	**14.**	9	**15.**	0
16.	1.	**17.**	3	**18.**	7	**19.**	7	**20.**	(A)
21.	(D)	**22.**	(D)	**23.**	(B)				

APSP Solutions

1. This is a characteristic feature of transition metals.

2. 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.

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

4. 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.
Hence, all these elements belongs to 15th group i.e. nitrogen family.

5. The order of penetration effect of different orbitals depends upon the different energies of the various sub-shells for the same energy level, e.g., electrons in s-subshell will have lowest energy and thus will be closest to the nucleus and will have highest penetration power, while p-subshell electrons will penetrate the electron cloud to lesser extent and so on.

11. Atomic radius increases on moving top to bottom in a group due to increasing number of shells. However, it decreasing on moving left to right in a period due to increasing Z_{eff} and addition of electrons to the same shell.
For H; cation is smaller than parent atom while anion is bigger than parent atom. H^- and Li^+ are isoelectronic species. So, ionic size $\propto \frac{1}{\text{nuclear charge}}$. Hence the correct order is $H^+ < Li^+ < H^-$.

12. Due to 4f-orbital electrons (poor shielding effect), there is increase in effective nuclear charge which leads to the contraction of the size of atoms. This is called lanthanide contraction.

16. As elements are ionized, the proton to electron ratio increases, so the attraction between valence shell electron and nucleus increases and as a result the size decreases. Therefore, the removal of electron from smaller cation requires higher energy. Hence the second ionisation enthalpy is greater than its first ionisation enthalpy.

17. (3) For possible $ns^2\ np^1$ configuration, the removal of fourth electron will be possibly from an inert gas electron configuration. So there will be high jump in the fourth ionisation enthalpy than the third ionisation enthalpy which will take place from ns^1 electron configuration.

18. The increasing order of 1st ionisation energy is f < d < p < s because of the increasing order of the penetration of the electrons as f < d < p < s if all other factors are same.

23. (1) The elements having large negative values of electron gain enthalpy generally act as strong oxidising agents. E.g. Halogens.
(2) The elements having low values of ionisation enthalpies act as strong reducing agents.E.g. Alkali metals.
(3) The formation of S^{2-}(g) from S(g) is an endothermic process. ($\Delta_{eg}H_1$ = small negative value, $\Delta_{eg}H_2$ = large positive value).

24. 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.

25. Be and N has $1s^2\ 2s^2$ and $1s^2\ 2s^2\ 2p^3$ stable configurations respectively. So addition of extra electron is difficult in their valence shell. The atomic size of C is smaller than B and also C has higher nuclear charge; so addition of electron will be easier in C than B.

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

If I.E. is ionisation enthalpy, then

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

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

If $\Delta_{eg}H$ is electron gain enthalpy, then

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

27. There is electrostatic repulsion between the two species having same type of charge. So energy has to be given for the addition of additional electron to O^-.

PART - III

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.**

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

5. (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. All are isoelectronic species but as number of protons i.e. atomic number increases, the attraction between electron (to be removed) and nucleus increases and thus ionisation enthalpies increase.

Order of Z : Te^{2-} (52) < I^- (53) < Cs^+ (55) < Ba^{2+} (56). So same will be the order of IE.

7. (A) Larger the value of ionisation enthalpy, more difficult will be the removal of electron to form cation.

(B) Electron gain enthalpy is the measure of the ease with which an atom receives the additional electron in its valence shell in gaseous phase. So, larger is the value of electron gain enthalpy, easier is the formation of anion.

(C) Electronegativity (Mulliken) = $\frac{\text{Ionisation energy} + \text{Electron affinity}}{2}$.

(D) As Z_{eff} increases, the valence shell as well as inner shells electrons are more strongly attracted by the nucleus. This causes the contraction in atomic size.

9. 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.

11. 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.

12. (A) $S^-(g) \longrightarrow S^{2-}(g)$; ΔH_{eg} = (+)ve because of electrostatic repulsion.

(B) $Ne(g) + e^-(g) \longrightarrow Ne^-(g)$; ΔH_{eg} = (+)ve because of stable completely filled electron configuration.

(C) $N(g) \longrightarrow N^-(g)$; ΔH_{eg} = (+)ve because of stable half filled electron configuration.

(D) $Al^{2+}(g) \longrightarrow Al^{3+}(g)$; ΔH_{IE} = (+)ve because of the removal of electron from cation.

14. Atomic number of Cu is 29 = x

Atomic number of Au is 79 = y

x + y = 108

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

17. Be, N, Ne

18. $K(g) + F(g) \rightarrow F^-(g) + K^+(g)$ $\Delta H = 18.4$ kCal = 0.8 eV

$K(g) \rightarrow K^+(g) + e^-$ IE = 4.3 eV

$F(g) + e^- \rightarrow F^-(g)$ $\Delta_{eg}H = IE - \Delta H = 0.8 - 4.3 = -3.5$ eV

x = –3.5

2x = 7.

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

22.

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

23. (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.