3

PERIODIC TABLE

GENESIS OF PERIODIC CLASSIFICATION

* **Dobereiner's Triads :** According to Dobereiner when elements of same properties are kept in the increasing order of their atomic weights, the atomic weight of middle element is equal to the mean atomic weight of remaining two elements. Such a group of elements is called Dobereiner's triad.

Triad of atoms			Mean of first and last element
Li	Na	K	
7	23	39	$\dfrac{7+39}{2} = 23$
Be	Mg	Ca	
8	24	40	$\dfrac{8+40}{2} = 24$

Dobereiner could arrange only a few elements as triads and there are some such elements present in a triad, whose atomic weights are approximately equal, e.g.

    Fe Co Ni
    Ru Rh Pd

Therefore, this hypothesis was not acceptable for all elements.

* **Newland's Rule of Octave :** As in music, the eighth node is same as the first node. If the elements are arranged in the increasing order of atomic weights, on starting with an element, the first element will exhibit similarities with the eighth element e.g.

Symbol of element						
Li	Be	B	C	N	O	F
7	9	11	12	14	16	19

Symbol of element						
Na	Mg	Al	Si	P	S	Cl
23	24	27	28	31	32	35.5

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

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

Lothar Meyer's Volume Curves
* The graphs of atomic volumes against atomic weights are known as Lothar Meyer volume curves.
* The alkali metals have highest atomic volumes.
* Alkaline earth metals (Be, Mg Ca, Sr, Ba, etc.) which are relatively a little less electropositive, occupy positions on the descending part of the curve.
* Halogens and the noble gases (except helium) occupy positions on the ascending part of the curve.
* Transition elements have very small volumes and therefore these are present at the bottoms of the curve.

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

Main features of Mendeleev's periodic table :
Elements are arranged in tabular form in rows and columns.
* The horizontal rows present in the periodic table are called periods. There are seven periods in the periodic table. These are numbered from 1 to 7 (Arabic numerals).
* Properties of elements in a particular period show regular gradation (i.e. increase or decrease) from left to right.
* The vertical columns present are called groups. These are nine in number and are numbered from I to VIII and Zero (Roman numerals).
* Groups I to VII are subdivided into A and B subgroups. Groups Zero and VIII don't have any subgroups.
* All the elements in a particular group are chemically similar in nature. They show regular gradation in their physical properties and chemical reactivities.

Defects in Mendeleev's periodic classification :
In spite of being a historic achievement Mendeleev's periodic table had some defects in it –
* **Position of hydrogen :** Hydrogen resembles alkali metals (forms H^+ ion just like Na^+ ions) as well as halogens (forms H^- ion similar to Cl^- ion). Therefore, it could neither be placed with alkali metals (group I) nor with halogens (group VII).
* **Position of isotopes :** Different isotopes of same elements have different atomic masses, therefore, each one of them should be given a different position in the periodic table. On the other hand, because they are chemically similar, they had to be given same position.
* **Anomalous pairs of elements :** At certain places, an element of higher atomic mass has been placed before an element of lower atomic mass. For example, Argon (39.91) is placed before potassium (39.1)

MODERN CLASSIFICATION
* Moseley, showed that atomic number is more fundamental property of an element than its atomic mass. Therefore, the position of an element in the periodic table depends on its atomic number.

* Moseley observed regularities in the characteristic X-ray spectra of the elements. A plot of $\sqrt{\nu}$ (where ν is frequency of X-rays emitted) against atomic number (Z) gave a straight line and not the plot of $\sqrt{\nu}$ vs atomic mass.

(a) **Modern periodic laws :** The physical and chemical properties of elements are the periodic functions of their atomic numbers.

(b) **Long Form of periodic table:** The arrangement of elements in the long form of periodic table is a perfect matching of electronic configuration of the elements, physical and chemical properties.

Structural Features of the Long Form of Periodic Table:

* There are seven horizontal rows. These rows are called **periods**. Thus the periodic table has seven periods, numbered from 1 to 7.

* First period consists of only two elements (very short period).

* Second and third periods consists of only eight elements each (short periods).

* Fourth and fifth periods consist of 18 elements each (long periods).

* Sixth period consists of 32 elements (long period).

* Seventh period is yet incomplete and more and more elements are likely to be added as the scientific research advances.

* **Period** - The details about the seven periods are as follows.

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

* In this table there are 18 vertical columns called **groups**. They are numbered from 1 to 18. Every group has a unique configuration.

* There are also nick names given to the groups or a cluster of groups on the basis of the similarity of their properties, as given below:

* Group 1 elements except hydrogen, are called **Alkali Metals**

* Group 2 elements are called **Alkaline Earth Metals.**

* Group 3 to 12 elements are called **Transition Metals.**

* Group 15 elements are called **Pnicogens.**

* Group 16 elements are called **Chalcogens.**

* Group 17 elements are called **Halogens.**

* Group 18 elements are called **Noble Gases.**

* Apart from what has been said above elements with atomic numbers 58 to 71 are called **Lanthanoids -** or Inner Transition elements (First series). Elements from atomic numbers 90 to 103 are called **Actinoids** - Inner Transition elements (Second series).

* All elements except transition and inner transition elements are also collectively called **Main Group Elements.**

Position of Metals, Non-Metals and Metalloids:

* In order to locate the position of metals, nonmetals and metalloids in the periodic table, you may draw a diagonal line joining the element boron (At. no. 5) with that of tellurium (At. no. 52) and passing through silicon and arsenic.

* We are in a position to make the following observations.

(i) The elements above the diagonal line and to the far right are non-metals (except selenium which shows slightly metallic character also). The non-metallic character is more marked the farther an element is from the diagonal line and up.

(ii) The elements below the diagonal line and to the left are metals. (Hydrogen is a nonmetal and is an exception)The metallic character is more marked the farther an element is from the diagonal line and down. All lanthanoids and actinoids are metals.

(iii) The elements along the diagonal line are metalloids and possess the characteristics of metals as well as of non-metals. In addition germanium, antimony and selenium also show the characteristics of metalloids.

Merits of modern periodic table over Mendeleev's periodic table:

* The modern periodic table is based on atomic number which is more fundamental property of an atom than atomic mass.

* The long form of modern periodic table is therefore free of main defects of Mendeleev's periodic table.

(a) **Position of isotopes :** All isotopes of the same elements have different atomic masses but same atomic number. Therefore, they occupy the same position in the modern periodic table which they should have because all of them are chemically similar.

(b) **Anomalous pairs of elements :** When elements are arranged in the periodic table according to their atomic numbers the anomaly regarding certain pairs of elements in Mendeleev's periodic table disappears. For example, atomic numbers of argon and potassium are 18 and 19 respectively. Therefore, argon with smaller atomic number comes before potassium although its atomic mass is greater and properties of both the elements match with other elements of their respective groups.

Figure : Long form of the Periodic Table of the Elements with their atomic numbers and ground state outer electronic configurations. The groups are numbered 1-18 in accordance with the 1984 IUPAC recommendations. This notation replaces the old numbering scheme of IA–VIIA, VIII, IB–VIIB and 0 for the elements.

Nomenclature of elements with atomic numbers greater than 100 :

* The naming of the new elements was earlier left entirely to its discoverer.
* The suggested names were then later ratified by IUPAC. But due to certain disputes that arose over the original discoverer of some of the elements of atomic numbers greater than 104, the IUPAC in 1994 appointed a Commission on Nomenclature of Inorganic Chemisty (CNIC).
* After consultation with the Commission and the chemists around the world, the IUPAC in 1997 recommeded a nomenclature to be followed for naming the new elements with atomic numbers greater than 103 until their names are fully recognised.
* The names are derived directly from the atomic number of the element using the following numerical roots for 0 and numbers 1-9.

0 = nil	3 = tri	6 = hex	9 = enn
1 = un	4 = quad	7 = sept	2 = bi
5 = pent	8 = oct		

* The roots are put together in the order of the digits which make up the atomic number and 'ium' is added at the end.

Atomic Number	Name according to IUPAC nomenclature	Symbol	IUPAC Official Name	IUPAC Symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnillium	Uun	Darmstadtium	Ds
111	Unununnium	Uuu	Rontgenium	Rg
112	Ununbium	Uub	Copernicium	Cn
113	Ununtrium	Uut	Nihonium	Nn
114	Ununquadium	Uuq	Flerovium	Fl
115	Ununpentium	Uup	Moscovium	Mc
116	Ununhexium	Uuh	Livermorium	Lv
117	Ununseptium	Uus	Tennessine	Ts
118	Ununoctium	Uuo	Oganesson	Og

CLASSIFICATION OF ELEMENTS ON THE BASIS OF ELECTRONIC CONFIGURATION

1. s-block elements
* Last electron enters in s–orbital.
* Their general formulae are ns^1 and ns^2 respectively, where n = (1 to 7).
* I A group elements are known as alkali metals because they react with water to form alkali.
* II A group elements are known as alkaline earth metals because their oxides react with water to form alkali and these are found in the soil or earth.
* The total number of s block elements are 14.
* They are highly electropositive.
* They have low ionisation energies and good reducing nature.
* Oxides and hydroxides are strongly basic in nature.

2. p - block elements
* Last electron gets filled up in the p-orbital.
* A p-orbital can accommodate a maximum of six electrons. Therefore, p-block elements are divided into six groups which are III A, IV A, V A, VI, A VII A and zero groups.
* The general formulae of p block elements are $ns^2 p^{1-6}$ (where n = 2 to 6)
* The zero group elements having general formula $ns^2 p^6$ are inert, because their energy levels are fully filled.
* The total number of p block elements in the periodic table is 30 (excluding He).
* There are nine gaseous elements (Ne, Ar, Kr, Xe, Rn, F_2, Cl_2, O_2 and N_2) belonging to p-block. Gallium (Ga) and bromine (Br) are liquids. The step-like thick lines drawn in the periodic table in the p-block divides elements into metals nonmetals and metalloids.
* The non-metallic character increases along a period from 13 to 17.
* They form covalent compounds among themselves but electrovalent compounds with s-block elements.
* Their oxides are generally acidic, few are amphoteric also. Example, Al_2O_3.

3. d-block elements
* Last electron gets filled up in the d orbital.
* In d block elements the electron gets filled up in the d orbital of the penultimate shell. That is why these elements are known as **transition elements**.
* The general formula of these elements is $(n–1)s^2, p^6, d^{1-10} ns^{1-2}$ where n = 4 to 7.
* All of these elements are metals. Out of all the d block elements mercury is the only liquid element.
* They are good conductors of heat and electricity.
* They generally form coloured compounds. This is due to the presence of incomplete d-subshell.
* They are generally paramagnetic.

4. f-block elements
* Last electron gets filled up in the f orbital.
* The f block elements are from atomic number 58 to 71 and from 90 to 103.
* The lanthanides occur in nature in low abundance and therefore, these are called rare earth elements.
* There are 28 f block elements in the periodic table.
* The elements from atomic number 58 to 71 are called lanthanides because they come after lanthanum (57).
* The elements from 90 to 103 are called actinides because they come after actinium (89). All the actinide elements are radioactive.
* All the elements after atomic number 92 (i.e. U^{92}) are transuranic elements.
* The general formula of these elements is $(n–2)s^2 p^6 d^{10} f^{(1-14)} (n–1)s^2 p^6 d^{0-1} ns^2$

POSITION OF AN ELEMENT IN THE PERIODIC TABLE
Rules :
1. Period = maximum value of principal quantum no. (n)
2. Groups, for s block = no. of valency e^- in ns subshell.

For p block = no. of valence e^- in ns and np subshell+10

For d block = no. of valence e^- in $[(n - 1) d^{1-10} ns^{1-2}]$.

For example, to know the group, period and block of $_{53}X$ let us see its electronic configuration :

$1s^2, 2s^2, 2p^6, 3s^2 3p^6, 3d^{10}, 4s^2 4p^6, 4d^{10}, 5s^2 5p^5$.
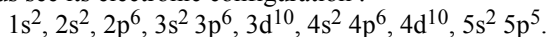
Its group will be 17th (sum of electrons in $(n - 1) d$, ns and np subshells), period 5th (the last orbit having electrons) and p block (the last subshell into which electrons are filled).

Example 1 :

Which alphabet not use in nomenclature of elements.

(1) K   (2) W   (3) V   (4) J & Q

Sol. **(4).** K is potassium (Kalium), W is Tungsten (Wolfram) and V is vanadium. There is no element associated with the letters J and Q.

Example 2 :

Which of the following is the period number of the element whose atomic number is 98 –

(1) 4   (2) 7   (3) 5   (4) 6

Sol. **(2)** The electronic configuration of the element with atomic number 98 is as follow :

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 6p^6, 7s^2, 5f^{10}$
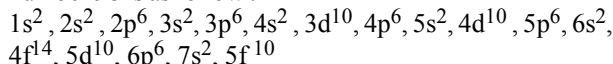
The last electron enters in f orbital, so it belongs to f block in the period.

Example 3 :

The nuclei of elements X, Y and Z have same number of protons, but different numbers of neutrons. According to Mendeleef periodic table, the elements X,Y and Z

(1) belong to same group and same period

(2) belong to different groups and different periods

(3) belong to same group and different periods

(4) are isotopes, which do not have different positions

Sol. **(4).** Isotopes have same number of protons. So they occupy same position in the periodic table. Due to different nos. of neutrons their atomic weights are different.

PERIODIC TRENDS IN PROPERTIES OF ELEMENTS

The recurrence of elements with similar properties after certain interval when these are arranged in the increasing order of their atomic numbers is known as **periodicity**.

1. **ATOMIC RADIUS :** The distance of the outermost orbit from the centre of the nucleus of an atom is called atomic radius.

(i) **Single Bond Covalent Radius, SBCR -**

(a)   For Homoatomic molecules

$$d_{A-A} = r_A + r_A \text{ or } 2r_A \; ; \; r_A = \frac{d_{A-A}}{2}$$

(b)   For hetrodiatomic molecules while electron negativity remains approximately same. $d_{A-B} = r_A + r_B$

For heteronuclear diatomic molecule. A–B, while difference between the electronegativity values of atom A and atom B is relatively larger, $d_{A-B} = r_A + r_B - 0.09 (X_A - X_B)$

where X_A & X_B are electronegativity values of high electronegative element A and less electronegative element B, respectively. This formula is given by Stevenson & Schomaker.

Note: Covalent radius is slightly smaller than actual radius.

Example 4 :

Calculate the bond length of C–X bond, if C–C bond length is 1.54 Å, X–X bond length is 1.00 Å and electronegativity values of C and X are 2.0 and 3.0 respectively

Sol. (1)   C–C bond length = 1.54 Å

$$r_C = \frac{1.54}{2} = 0.77 \text{ Å} \; ; \; r_X = \frac{1.00}{2} = 0.50 \text{ Å}$$

(2)   C–X bond length

$d_{C-X} = r_C + r_X - 0.09 (X_X - X_C)$

$= 0.77 + 0.50 - 0.09 (3-2) = 0.77 + 0.50 - 0.09 \times 1$

$= 1.27 - 0.09 = 1.18 \text{ Å}$. Thus C–X bond length is 1.18 Å

(ii)   **Van der Waals radius :** Half of the distance between the nuclei of two non bonded atoms belonging to two different molecules closest to one another is called van der waals radius. The values of atomic radii in noble gases are always determined as van der Waals radii. Therefore, the value of van der waals radius of a noble gas is always greater than that of the halogen coming before it in the same period.

Covalent radius   vander Waals radius
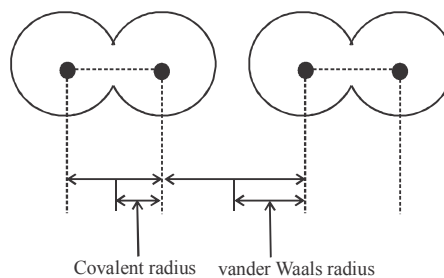
Van der Waals radius > Metallic radius > Covalent radius

Periodicity in atomic radius :

The atomic radius depends upon the following two factors.

(a)   **Number of orbits -** The effect of increase in the number of orbit in an atom increases the atomic size.

(b)   **Effective Nuclear Charge -** In a polyelectronic atom, the internal electrons repel the electrons of the outermost orbit. This results in decrease the nuclear attraction on the electrons of the outermost orbit. Therefore, only a part of the nuclear charge is effective on the electrons of the outermost orbit. Thus, the inner electrons protect or shield the nucleus and thereby decrease the effect of nuclear charge towards the electrons of the outermost orbit.

Thus the part of the nuclear charge works against outer electrons, is known as effective nuclear charge

$$\sigma$$

Z^* = effective nuclear charge, σ = shielding constant and Z = nuclear charge

A scientist named slater, determined the value of shielding constant and put forward some rules as following.

(i) The shielding effect or screening effect of each electron of 1s orbital is 0.30.

(ii) The shielding effect of each electrons of ns and np i.e. electron of the outermost orbit, is 0.35.

(iii) The shielding effect of each electron of s,p or d orbitals of the penultimate orbit $(n-1)$ is 0.85.

(iv) The shielding effect of each electron of s, p, d or f orbital of the pre-penultimate orbit $(n-2)$ and below this is 1.0.

* The effect of increase in the number of protons increases the effective nuclear charge. This results in decrease in the value of atomic radius because protons attract the electronic orbits with greater force.

(c) **Shielding effect/screening effect -** The electrons of inner shell repel the electrons of valence shell from coming closer to the nucleus. Due to this the atomic radius increasing. This is called as shielding. When an atom has more number of shells, the shielding provided will also be greater.

With in a shell the shielding provided by different types of orbitals follows the order $s > p > d > f$. (As 's' orbitals experience greater penetration towards the nucleus, they provide greater shielding to the electrons of p, d and f orbitals).

In the elements of 'd' and 'f' blocks where number of shells remain the same in a particular series along with shielding provided, the nuclear charge is also compared.

In a period - The number of orbits remains same on going from left to right in a period while there is a unit increase in the atomic number. Thus the electron experiences more force of attraction towards nucleus. Hence atomic radius decreases from left to right in a period.

In a group - The atomic radii increase on going downwards in a group because the number of orbits also increase on going from top to bottom in a group.

In a period, the size of an alkali metal (Group I A) is second largest because it has minimum number of proton, while the size of the halogen is smallest.

In the periodic table, Cs is the biggest atom, because Fr is a radioactive element, while H is the smallest atom.

Example 5 :
The increasing order of atomic size of Li, Be, B and Ne is
Sol. $B < Be < Li < Ne$
Inert gas is biggest in a period

Example 6 :
The increasing order of atomic size As, Bi, Sn, Pb and Sb is
Sol. $As < Sb < Sn < Bi < Pb$

2. **IONIC RADIUS :** When a neutral atom loses one or more electrons it forms a cation having one or more number of positive charge. Similarly when a neutral atom acquires one or more electrons it forms an anion having one or more number of negative charge.

Ionic radius is the distance between the nucleus and the limit of the electron cloud scattered around the nucleus in an ion.

Cationic radius : An atom forms a cation on loss of electrons. The cationic radius can be defined as the distance between the nucleus and the limit of the electron cloud scattered around the nucleus. The size of a cation is smaller in comparison to the size of its corresponding atom. Usually a cation has on shell less than the neutral atom hence it has smaller size than the atom. This is because of the fact that an atom on losing electrons/s form a cation, which has lesser number of electrons/s than the number of proton/s. This results in increase in the effective nuclear charge.

Examples (1) $Mn > Mn^{+2} > Mn^{+3} > Mn^{+4} > Mn^{+6} > Mn^{+7}$
(2) $Pb^{+2} > Pb^{+4}$

Anionic radius : When a neutral atom gains electron/s it becomes a negatively charged ion called an anion. The distance between the nucleus of an anion and the limit of the electron cloud scattered around the nucleus, is called its anionic radius. The size of an anion is greater than the size of its corresponding atom, because the number of electrons present in the anion gets larger than the number of protons due to gain of electron/s. This results in decrease in the effective nuclear charge. $O^0 < O^{-1} < O^{-2}$

Size of Isoelectronic Series : The species, which have same number of electrons but different nuclear charges, constitute an isoelectronic series. In the isoelectronic species with the increase in effective nuclear charge, the size of ion goes on decreasing.

Example 7 :
What should be the order of size of H^{-1}, H^{+1} and H ?

H^{-1}	H^{+1}	H
1p	1p	1p
2e	0e	1e

Sol. $H^{+1} < H < H^{-1}$

3. **IONISATION POTENTIAL**
The energy required to remove the most loosely bound electron from the outermost orbit of one mole of isolated gaseous atoms of an element, is called ionisation potential (IP). This ionisation is an endergonic or energy-absorbing process. An electron cannot be removed directly from an atom in solid state. For this purpose, the solid state is converted to gaseous state and the energy required for this is called sublimation energy.

$$A_{(g)} \xrightarrow{I^{st} \ IP} A^{+1}_{(g)} \xrightarrow{II^{nd} \ IP} A^{+2}_{(g)}$$

$$\xrightarrow{III^{rd} \ IP} A^{+3}_{(g)}$$
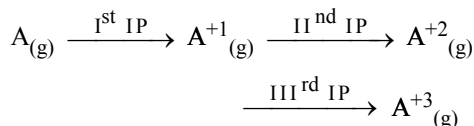
The energy required to remove one electron from a neutral gaseous atom to convert it to monopositive cation, is called first ionisation potential (I^{st} IP). The energy required to convert a monopositive cation to a dipositive cation is called second ionisation potential (IInd IP).

I^{st} IP $< II^{nd}$ IP $< III^{rd}$ IP because as the electrons go out of the atom, the effective nuclear charge increases & the ionic

size goes on decreasing. Thus the nuclear force of attraction on valence shell electrons increases and hence the order.

Factors affecting ionisation potential :

(i) **Number of shells :** With the increase in number of shells the atomic radius increases i.e. the distance of outer most shell electron from the nucleus increases and hence the ionisation potential decreases.

(ii) **Effective Nuclear Charge :** Atomic size decreases with increase in effective nuclear charge because, higher the effective nuclear charge stronger will be the attraction of the nucleus towards the electron of the outermost orbit and higher will be the ionisation potential

(iii) **Shielding Effect :** The electrons of internal orbits repel the electron of the outermost orbit due to which the attraction of the nucleus towards the electron of the outermost orbit decreases and thus atomic size increases and the value of ionisation potential decreases.

(iv) **Stability of half filled and fully filled orbitals :** The atoms whose orbitals are half-filled (p^3, d^5, f^7) or fully-filled (s^2, p^6, d^{10}, f^{14}) have greater stability than the others. Therefore, they require greater energy for removing an electron. However stability of fully filled orbitals is greater than that of the half filled orbitals.

(v) **Penetration power :** In any atom the s orbital is nearer to the nucleus in comparison to p, d and f orbitals. Therefor, greater energy is required to remove an electron from s orbital than from p, d and f orbitals. Thus the decreasing order of ionisation potential of s, p, d and f orbitals is as follows $s > p > d > f$.

Periodic Trends in Ionisation Potential :

(a) **In a Period :** The value of Ionisation potential normally increase on going from left to right in a period, because effective nuclear charge increases and atomic size decreases.

Exceptions : In second period ionisation potential of Be is greater than that of B, and in the third period ionisation potential of Mg is greater than that of Al due to high stability of fully filled orbitals. In second period ionisation potential of N is greater than O and in the third period ionisation potential of P is greater than that of S, due to stability of half filled orbitals. The increasing order of the values of ionisation potential of the second period elements is $Li < B < Be < C < O < N < F < Ne$

The increasing order of the values of ionisation potential of the third period elements is

$Na < Al < Mg < Si < S < P < Cl < Ar$
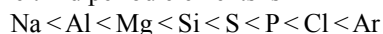
Ionisation potential of transition elements :

In transition elements, the value of ionisation potential has very little increase on going from left to right in a period because the outermost orbit remains the same but electrons get filled up in the (n–1)d orbitals resulting in very little increase in the values of ionisation potential.

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

(i) The value of ionisation potential increases with increase in effective nuclear charge.

(ii) The value of ionisation potential decreases with increase in shielding effect when the number of electrons increases in (n-1)d orbitals

In the first transition element series the first ionisation potential normally increases on going from left to right from Sc to Cr because shielding effect is much weaker in comparison to effective nuclear charge. The value of first ionisation potential of Fe, Co and Ni remains constant, because shielding effect and effective nuclear charge balance one another. The value of ionisation potential shows slight increase from Cu to Zn because they have fully filled s and d orbitals. The value of first ionisation potential of Mn is maximum because it has maximum stability due to fully filled s and d orbitals.

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

(b) **In a Group :** The value of ionisation potential normally decreases on going from top to bottom in a group because both atomic size and shielding effect increase.

Exception : Vlue of ionisation potential remains almost constant from Al to Ga in the III A group. (B>Al , Ga > In) In IV_B group i.e. Ti,, Zr and Hf the I.P. of Hf is higher than that of Zr due to Lanthanide contraction. Thus the I.P. of IV_B group varies as Ti > Zr < Hf.

In the periodic table the element having highest value of ionisation potential is He. The values of ionisation potential of noble gases are extremely high, because the orbitals of outermost orbit are fully-filled (ns^2, np^6) and provide great stability. In a period, the element having least value of ionisation potential is an alkali metal (group I A) and that having highest value is inert gas (Group 0)
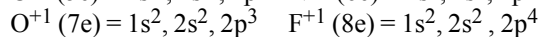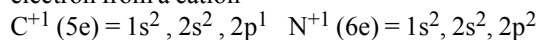
Applications of Ionisation Potential : The elements having high values of ionisation potential have low reactivity, e.g. inert gases. The value of ionisation potential decreases more on going from top to bottom in a group in comparison to a period. Therefore, reactivity increases and the atom forms a cation by loss of electron. The elements having low value of ionisation potential readily lose electron and thus behave as strong reducing agents. The elements having low value of ionization potential readily lose electron and thus exhibit greater metallic property. The elements having low value of ionisation potential readily lose electron and thus oxide and hydroxides of these elements have basic property.

Example 8 :

Which of the following should be the order of increasing values of second ionisation potential of C_6, N_7, O_8 and F_9

(1) $C > N > F > O$ (2) $C < F < N < O$

(3) $C < F < N < O$ (4) $C < N < F < O$

Sol. **(4).** The second ionisation potential means removal of electron from a cation

C^{+1} (5e) $= 1s^2, 2s^2, 2p^1$ N^{+1} (6e) $= 1s^2, 2s^2, 2p^2$

O^{+1} (7e) $= 1s^2, 2s^2, 2p^3$ F^{+1} (8e) $= 1s^2, 2s^2, 2p^4$

Therefore $C < N < F < O$

Example 9 :

Which of the following should be the correct order of the second ionisation potential of Li, Be B and C

(1) $Li < Be > B < C$ (2) $Be < B < C < Li$

(3) $Be < C < B < Li$ (4) $Li < C < B < Be$

Sol. **(3).** Li^{+1} (2e) $= 1s^2$ Be^{+1} (3e) $= 1s^2, 2s^1$

B^{+1} (4e) $= 1s^2 . 2s^2$ C^{+1} (5e) $= 1s^2, 2s^2, 2p^1$

Therefore $Be < C < B < Li$

Here, Li^{+1} has inert configuration and B^{+1} has fully filled s orbital. Hence these will have high value of second ionisation potential.
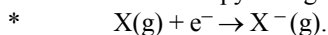
Example 10:

Which of the following should have least value of ionisation potential

(1) Be^{+3} (2) H

(3) Li^{+2} (4) He^{+1}

Sol. **(2).** All the above four species have one electron each, H has least number of proton. Therefore, H will have least value of ionisation potential.

4. ELECTRON GAIN ENTHALPY ($\Delta_{eg}H$)

* It is the enthalpy change for the reaction

* $X(g) + e^- \rightarrow X^-(g)$.

All the elements except noble gases, have negative values of $\Delta_{eg}H$ whereas noble gases have positive values of $\Delta_{eg}H$.

* **Trend in $\Delta_{eg}H$:** Electron gain enthalpy generally becomes more negative with increase in the atomic numbers across a period and within a group becomes less negative down the group.

* Electron gain enthalpies have large negative values toward the upper right of the periodic table preceding the noble gases.

* **Exceptions :** P, S, Cl have more negative $\Delta_{eg}H$ than N, O, F respectively.

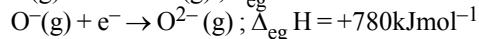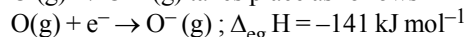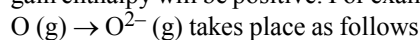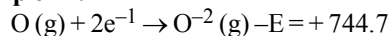The above irregularity can be explained on the basis of small size of elements of II period. Due to small size, they have high electron density $\left(\dfrac{charge}{volume}\right)$ and so repel the electron.

Second electron gain enthalpy of an atom is always positive.

• lowest EA in period = alkali earth metal or noble gas

• highest EA in period = halogen

* **Successive electron gain enthalpy :**

We know that the electrons are lost in succession. Actually when a gaseous atom gains an electron, the anion has a negative charge. In case the anion is to gain another electron, energy will be needed to overcome the repulsion between the anion and the electron. As a result-electron gain enthalpy will be positive. For example,

$O(g) \rightarrow O^{2-}(g)$ takes place as follows

$O(g) + e^- \rightarrow O^-(g)$; $\Delta_{eg}H = -141$ kJ mol^{-1}

$O^-(g) + e^- \rightarrow O^{2-}(g)$; $\Delta_{eg}H = +780$ kJmol^{-1}

Example 11 :

$O(g) + 2e^{-1} \rightarrow O^{-2}(g) -E = +744.7$

The reason for the positive value of E is

(1) endothermic reaction (2) exothermic reaction

(3) both 1 and 2 (4) All of the above are wrong

Sol. **(1).** When electron is brought near O^{-1} there will be repulsion between them, and therefore the energy will be positive i.e there will be absorption of energy during the process.

Example 12 :

The increasing order of electron affinity of N, P and As is

(1) $N < P < As$ (2) $As < P < N$

(3) $As < N < P$ (4) $As < N > P$

Sol. **(3).** Phosphorus have vacant 'd' orbitals due to which it has higher electron affinity than Nitrogen.

5. ELECTRONEGATIVITY(EN)

* It is the qualitative measure of an atom in a chemical compound to attract the shared electrons to itself.

* The most widely used scale is Pauling scale which is based on bond energy data. By giving a reference value of 2.1 to H, maximum electronegativity value of 4 is assigned to F.

Factors affecting electronegativity :

* **Atomic size –** Electronegativity of a bonded atom decreases with increase in size When effective nuclear charge is high, the nucleus will attract the shared electrons with greater strength and the electronegativity will be high. Thus with increasing atomic size electronegativity value decreases.

* **Hybridisation state of atom –** Electronegativity increases with increases in the s character of the hybrid orbital. This is because the s orbital is nearer to the nucleus and thus suffers greater attraction resulting in increase in electronegativity.

The number of covalent bonds present between two bonded atoms is known as bond order. With increases in the bond order, the bond distance decreases, effective nuclear charge increases and thus electronegativity increases.

The increasing order of electronegativity is as follows :

$C–C < C = C < C \equiv C$
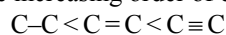
* **Oxidation number –** The electronegativity value increases with increase in oxidation number because radius decreases with increase in oxidation number.

The increasing order of electronegativity is :

$$Fe < Fe^{+2} < Fe^{+3}$$
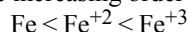
Electronegativity does not depend on stability of fully-filled or half-filled orbitals because it is simply the capacity of nucleus to attract bonded pair of electrons.

* **Periodicity:**

(a) In period- The electro-negativity increases from left to right in a period.

(b) In group- The electro-negativity decreases from top to bottom in a group.

Measurement of Electronegativity

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

Δ_{A-B} = Observed bond energy (D)

$\qquad\qquad$ – Energy of 100% covalent (E_{A-B})

$X_A - X_B \quad = 0.208 \sqrt{\Delta_{A-B}}$

where bond energies are in Kcal mole^{-1}

If Δ is in SI unit i.e. KJ mole^{-1}

$X_A - X_B = 0.1017\sqrt{\Delta_{A-B}}$

where, X_A & X_B are electronegativities of atoms A and B.

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

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

where X_M = Electronegativity value as given by Mulliken

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

where X_p = Electronegativity value as given by pauling.

APPLICATIONS OF ELECTRONEGATIVITY

(i) **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.

Example 13 :

Write the correct formula and name of the following

(a) ICl or CIl $\qquad\qquad$ (b) FCl or ClF

(c) BrCl or ClBr $\qquad\qquad$ (d) BrI or IBr

Sol. Correct formula $\qquad$ Name

(a) $I^+ Cl^-$ $\qquad\qquad$ Iodine chloride

(b) $Cl^+ F^-$ $\qquad\qquad$ Chlorine fluoride

(c) $Br^+ Cl^-$ $\qquad\qquad$ Bromine chloride

(d) IBr $\qquad\qquad\qquad$ Iodine bromide

(ii) **Nature of bond :** If difference of electronegativities of the two elements is 1.7 or more, then ionic bond is formed between them whereas if it is less than 1.7, then covalent bond is formed. (HF is exception in which bond is covalent although difference of electronegativity is 1.9)

(iii) **Metallic and Nonmetallic Nature**

Generally values of electronegativity of metallic elements are low, whereas electronegativity values of nonmetals are high.

(iv) **Partial Ionic Character in Covalent bonds**

Partial ionic characters are generated in covalent compounds by the difference of electronegativities. **Hanny and smith** calculated percentage of ionic character from the difference of electronegativity.

Percentage of ionic character

$= 16(X_A - X_B) + 3.5(X_A - X_B)^2 = 16\Delta + 3.5\Delta^2$

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

X_A is electronegativity of element A

X_B is electronegativity of element B

$\qquad \Delta = X_A - X_B$

(v) **Bond length :** When difference of electronegativities of atoms present in a molecule is increased, then bond length decreases.

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

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

Here $X_A > X_B$

(vi) **Bond Strength & Stability :** Bond strength and stability of A–B increases on increase in difference of electronegativities of atoms A and B bonded A–B.

Therefore H–F > H–Cl > H–Br > H–I

(vii) **Nature of Oxides**

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

Oxide						
Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
$(X_O - X_A)$						
2.6	2.3	2.1	1.8	1.5	1.1	0.5

(viii) **Nature of Hydroxides**

According to Gallis if electronegativity of A in a hydroxide (AOH) is more than 1.7 then it will be acidic in nature whereas it will be basic in nature if electronegativity is less than 1.7

For example $\qquad$ N<u>a</u>OH and $\qquad$ C<u>l</u>OH

Electronegativity (X_A) 0.9 $\qquad\qquad$ 3.00

$\qquad$ Nature $\qquad$ Basic $\qquad\qquad$ Acidic

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

Example 14 :

Increasing order of electronegativities of F, Cl, Br and I is

(1) F < Cl < Br < I $\qquad\qquad$ (2) I < Br < Cl < F

(3) Br < I > Cl > F $\qquad\qquad$ (4) I < Br > Cl < F

Sol. **(2)** Electronegativity decreases in a group on going from top to bottom. Therefore increasing electronegativity order is $\quad$ I < Br < Cl < F.

Example 15 :

CF_3NH_2 is not a base, whereas CH_3NH_2 is a base. What is the reason ?

Sol. Due to high electronegativity of F tendency of donating the lone pair of electrons present on N will be less.

Example 16 :

OF_2 is called oxygen difluoride, whereas Cl_2O is called dichlorine monoxide. Why ?

Sol. Electronegativity of O in OF_2 is less than F. Therefore, there will be positive charge on oxygen and negative charge on fluorine. Whereas in Cl and O, electronegativity of Cl is less than that of O therefore there will be positive charge on Cl and negative charge on O. Positive charge is written first followed by negative charge.

Example 17:

Which of the following is most acidic ?

(A) H_2Te (B) H_2O (C) H_2S (D) H_2Se

Sol. **(A).** Te is biggest which will give H^+ quickly.

6. DENSITY

* Density in solid state changes on change in atomic number of elements.
* Moving left to right in the period densities of solid and liquid elements first increases, it is maximum in the middle and then decreases.
* Moving upwards to downwards in a group, density generally increases but there is irregularity in this order.
* Density decreases instead of increasing on moving towards K in IA group and moving from Mg to Ca in IIA group. Because penultimate value of K and Ca increases, due to which density gets reduces.
* Os (22.6 gm/cc) is most dense in solid elements and in liquids Hg is most dense (13.6 gm/cc).

PERIODIC TRENDS IN CHEMICAL PROPERTIES

(i) **Valence electrons :** The electrons present in outermost shell are called as valence electron. Because the electrons in the outermost shell determine the valency of an element. **Valence of an element :** The number of hydrogen or halogen atom or double the number of oxygen atom, which combine with one atom of the element is taken as its valency. According to the electronic concept of valency, " the number of electrons which an atom loses or gains or shares with other atom to attain the noble gas configuration is termed as its valency."

Periodicity in valency or oxidation state:

(a) In period- The valency first increases then decreases from left to right in a period.

(b) In group- The valency remains constant from top to bottom in a group.

(ii) **Electropositive or metallic character :** The tendency of an element to lose electrons and forms positive ions (cations) is called electropositive or metallic character. The elements having lower ionisation energies have higher tendency to lose electrons, thus they are electropositive or metallic in their behaviour. Alkali metals are the most highly electropositive elements.

Periodicity:

(a) In period- The electropositive or metallic characters decreases from left to right in a period.

(b) In group- The electropositive or metallic characters increases from top to bottom in a group.

The order of increasing metallic character is:

$P < Si < Be < Mg < Na$.
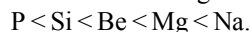
(iii) **Electro-negative or non-metallic characters:** The tendency of an element to accept electrons to form an anion is called its non metallic or electronegative character. The elements having high electro-negativity have higher tendency to gain electrons and forms anion. So, the elements in the upper right hand portion of the periodic table are electro-negative or non-metallic in nature.

Periodicity:

(a) In period- The electro-negative or non- metallic characters increases from left to right in a period.

(b) In group- The electro-negative or non-metallic characters decreases from top to bottom in a group.

(iv) **Reactivity of metals :**

Periodicity:

(a) In period- The tendency of an element to lose electrons decreases in a period. So the reactivity of metals decreases from left to right in a period.

(b) In group- The tendency of an element to lose electrons increases in a period. So the reactivity of metals increases from top to bottom in a group.

(v) **Reactivity of non-metals :**

(a) In period- The tendency of an element to gain electrons increases in a period. So the reactivity of non-metals increases from left to right in a period.

(b) In group- The tendency of an element to gain electrons decreases in a group. So the reactivity of non-metals decreases from top to bottom in a group.

(vi) **Solubility of alkali metals carbonates and bicarbonates :**
Perodicity in group : The solubility of alkali metal carbonates and bicarbonates in water increases down the group (From Lithium to Caesium).

(vii) **Solubility of alkaline earth metal hydroxides and sulphates:**
Perodicity in group : The solubility of alkaline earth metal hydroxide and sulphates in water increases down the group (From Beryllium to Barium).

(viii) **Basic strength of alklanine earth metal hydroxides :**
Perodicity in group : The basic strength of alkaline earth metal hydroxide in water increases down the group (From Beryllium to Barium), i.e.,

$Be(OH)_2 < Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$

$\xrightarrow{\hspace{6cm}}$
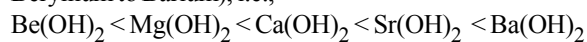
Basic strength increases.

(ix) Thermal stability of carbonates of alkali and alkaline earth metals :

Except lithium carbonate, $(LiCO_3)$, the carbonates of all other alkali metals are stable towards heat, i.e., carbonates of alkali metals (except $LiCO_3$) do not decompose on heating. $LiCO_3$ decomposes on heating to give lithium oxide $(LiCO_3)$. The carbonates of alkaline earth metals are relatively less stable. On heating, they decompose to give corresponding oxide and CO_2 gas. The decomposition temperature for alkaline earth metal carbonates increases as we go down the group.

(x) Acidic and basic nature of hydroxides

Acidic and basic nature of hydroxide of an element AOH depends on ionisation potential of A. If ionisation potential of A is low then it will give its electron to oxygen easily thus AOH will be basic.

Basic character of hydroxides :

$CsOH > RbOH > KOH > NaOH > LiOH$
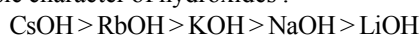
Acidic character of oxides and hydroxides :
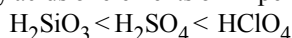
$HClO_4 > HBrO_4 > HIO_4$

(xi) Nature of Oxy acids

* Those inorganic acids in which X – OH group is present are called oxy acids. (Here X = a non-metal) like HOCl is an oxy acid.
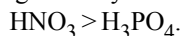* Moving from left to right in a period strength of oxy acids increases. Example,

Oxy acid of elements of II period :

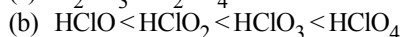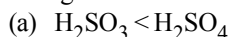$H_3BO_3 < H_2CO_3 < HNO_3$
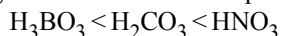
Oxy acids of elements of III period:

$H_2SiO_3 < H_2SO_4 < HClO_4$

In a group : Moving upwards to downward in a group, strength of oxy acids decreases like in group 5.

$HNO_3 > H_3PO_4$.

When amount of oxygen increases in an oxy acid of an element (i.e. on increasing oxidation state of element) strength of acid increases like

(a) $H_2SO_3 < H_2SO_4$

(b) $HClO < HClO_2 < HClO_3 < HClO_4$

Stability of acids also increases in the same order.

(xii) Nature of hydrides

* Moving from left to right in a period, nature of hydrides of non-metal elements becomes basic to acidic. Like

In second period NH_3 H_2O HF
Weak base   Neutral   Acidic
(Amphoteric)

In third period PH_3 H_2S HCl
very weak   weak acid   strong acid
acid

* Moving upwards to downwards in a group, acidic nature and reducing power of hydrides of non-metal elements both increases but stability decreases.

Acid strength $HF < HCl < HBr < HI$

Reducing stability $HF > HCl > HBr > HI$

Exception : Negative N is written first in NH_3 because it became very common from the very beginning. There is no scientific basis for that. Otherwise according to rule it should have been written as H_3N in the form of trihydrogen nitride.

Anomalous Properties of Second Period Elements

* Each element of second period, i.e., first element of each of group 1 and 2 and groups 13-17 shows many properties which are not shown by its cogeners.
* Their anomalous behaviour is attributed to their small size, large charge/radius ratio, high electronegativity, non-availability of d-orbitals in their valence shell.
* Some elements of second period Li, Be, B shows dissimilarities with other elements of this group but shows similarities with elements of third group like Mg,. Al, Si situated diagonally to them. It is called diagonal relationship.

$_3Li$ $_4Be$ $_5B$ $_6C$

$_{11}Na$ $_{12}Mg$ $_{13}Al$ $_{14}Si$
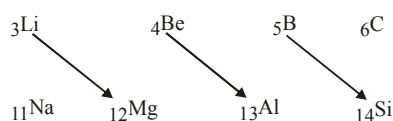
* The first member of each group has only four valence orbitals (one 2s and three 2p orbitals) for bonding, whereas the second member of the group has nine valence orbitals (one 3s, three 3p and five 3d orbitals).
* As a consequence of this, maximum covalency of first member of each group is limited to '4', whereas the other members of the group can expand their valence shell to accommodate more than four pairs of electrons. For example, B from $[BF_4]$ and Al, $[AlF_6]_3$.
* In addition to this, the first member of each group of p-block elements displays greater ability to form $p\pi$ -$p\pi$ multiple bonds to itself (e.g. $C = C, C \equiv C, N = N, N \equiv N$) and to other second period elements (e.g., $C = O, C = N, C \equiv N, N = O$) compared to subsequent members of the same group.

TRY IT YOURSELF

Q.1 The electron configuration of an element is $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$, locate the element in the periodic table.

Q.2 The set representing correct order of IP_1 is –
(A) $K > Na > LI$ (B) $Be > Mg > Ca$
(C) $B > C > N$ (D) $Fe > Si > C$

Q.3 The IE_1 of Be is greater than that of B. **[T/F]**

Q.4 The element with atomic no. 107 and 109 have been prepared artificially. Indicates the groups in which these elements should be placed.

Q.5 Out of Li and Be, which is expected to have higher second ionisation energy?

Q.6 Predict the formulas of the stable binary compound that would be formed by the following pairs of elements.
(i) Lithium and Oxygen (ii) Silicon and Oxygen.

Q.7 The first and second ionization enthalpies in kJ mol^{-1} and the electron gain enthalpy in kJ mol^{-1} of a few elements are given below:

| Element	$\Delta_i H_1$	$\Delta_i H_2$	$\Delta_{eg} H$
(I)	520	7300	-60
(II)	419	3051	-48
(III)	1681	3374	-328
(IV)	2372	5251	$+48$

Which of the above elements is likely to be –
(i)   The least reactive elements?
(ii) The most reactive metal?
(iii) The most reactive non-metal?

Q.8 The correct order of radii is –
(A) $N < Be < B$ (B) $F^- < O^{2-} < N^{3-}$
(C) $Na < Li < K$ (D) $Fe^{3+} < Fe^{2+} < Fe^{4+}$

Q.9 Arrange Co, Co^{2+} and Co^{3+} in increasing order of radii.

Q.10 The periodic table consists of 18 groups. An isotope of copper, on bombardment with protons, undergoes a nuclear reaction yielding element X as shown below. To which group , element X belongs in the periodic table ?

$$^{63}_{29}Cu + {}_1H^1 \rightarrow 6\,{}_0n^1 + \alpha + 2\,{}_1^1H + X$$
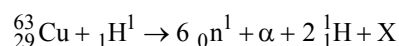
Q.11 The electron gain enthalpy of chlorine is -9.83 kJ/gram. Calculate its electron gain enthalpy per mole.

Q.12 Out of Li^+, Be^{2+} and B^{3+} ions, which has the smallest ionic radius and why ?

ANSWERS

(1) 4th period, s-block, Group 2   **(2)** (B)
(3) True   **(4)** seventh period **(5)** Li
(6) (i) Li_2O, (ii) SiO_2 **(7)** (i) IV, (ii) II, (iii) III
(8) (B)   **(9)** $Co^{3+} < Co^{2+} < Co$
(10) 8   **(11)** 349 kJ   **(12)** B^{3+}

IMPORTANT POINTS

* Mg is bridge element, which joins metals of IIA and II B groups.
* Elements after atomic number 92 are transuranic elements.
* Artificial element is Tc^{43}.
* Liquid non-metal – Br
* Liquid metal – Hg, Ga, Cs, Fr
* Solid volatile non-metal – Iodine
* Lightest metal – Li
* Heaviest metal – Ir
* Hardest metal – W
* Nobel metals – Pc, Pt, Au, Ag
* Element most found on earth – Al
* Gaseous elements – 11
  (He, Ne, Ar, Kr, Xe, Rn, H_2, N_2, O_2, Cl_2, F_2)
* Liquid elements – 5 (Br, Hg, Ga, Cs, Fr)
* Submetals – 5 (B, Si, As, Te, At)
* Inert gases – 6
* Lowest electronegativity : Cs
* Highest electronegativity : F
* Highest ionisation potential : He
* Lowest ionisation potential : Cs
* Highest electron affinity : Chlorine (Cl)

* Least electropositive element : Fluorine (F)
* Most reactive solid element : Li
* Most reactive liquid element : Cs
* Most stable element : Te
* Largest atomic size : Cs

* Ionisation potential $\propto \dfrac{1}{\text{atomic size}}$
* Ionisation potential $\propto$ Effective nuclear charge
* Ionisation potential $\propto \dfrac{1}{\text{shielding effect}}$
* I.P. of fully-filled orbitals > I.P. of half-filled orbitals
* Ionisation potential $\propto$ Penetration power
* I.P. $\propto \dfrac{1}{\text{Reactivity of metal}} \propto \dfrac{1}{\text{Reduing property}}$
* I.P. $\propto \dfrac{1}{\text{Metallic property}} \propto \dfrac{1}{\text{Basic property}}$
* Electron affinity $\propto$ Effective nuclear charge $\propto \dfrac{1}{\text{atomic size}}$
* Electron affinity $\propto \dfrac{1}{\text{shielding effect}}$
* E.A. 2nd period p-block element < E.A. 3rd period p-block element
* Electronegativity $\propto \dfrac{1}{\text{atomic size}} \propto$ Effective nuclear charge
* Electronegativity $\propto \dfrac{1}{\text{shielding effect}}$
* Electronegativity $\propto$ % of s-character of the hybrid orbital.
* Electronegativity $\propto \dfrac{1}{\text{\% of p-character of the hybrid orbital}}$
* Bond order $\propto$ Electronegativity
* Oxidation number $\propto$ Electronegativity
* Polarity of covalent bond $\propto$ Difference of electranegativities
* Basic character of elements $\propto \dfrac{1}{\text{Effective nuclear charge}}$
* Basic character of elements $\propto \dfrac{1}{\text{Ionisation potential of metal}}$
* Basic character of elements $\propto$ Metallic property
* Acidic character of elements $\propto \dfrac{1}{\text{atomic size}}$
* Acidic character of elements $\propto$ Effective nuclear charge
* Acidic character of elements $\propto$ Electronegativity of metal
* Acidic character of elements $\propto$ Oxidation number of metal
* Acidic character of elements $\propto \dfrac{1}{\text{Metallic character}}$
* Acidic character of elements $\propto$ Size of atoms which forms anion
* Basic nature of oxides $\propto$ Metallic nature
* Basic nature of oxides $\propto$ Electropositive nature
* Acidic nature of oxides $\propto$ Non-metallic nature
* Acidic nature of oxides $\propto$ Electronegative nature
* Acidic nature of oxides $\propto$ No. of electrons in outermost shell

GENERAL TRENDS OF DIFFERENT PROPERTIES

	PERIODS →	GROUPS ↓
* Atomic radius	Decreases	Increases
* Ionisation potential	Increases	Decreases
* Electronegativity	Increases	Decreases
* Electron affinity	Increases	Decreases
* Covalent character of halides	Increases	Decreases
* Metallic character	Decreases	Increases
* Oxidising nature	Increases	Decreases
* Reducing nature	Decreases	Increases
* Screening effect	Increases	Increases
* Effective nuclear charge (Zeff)	Increases	Constant
* Valency w.r.t. Oxygen	Increases	Constant
* Basic character of hydroxides	Decreases	Metals Increases
* Basic character of hydrides	Decreases decreases	(Non-metal) (Metal) increases
* Basic character of Oxides	Decreases	Increases
* Strength of oxyacids	Increases	Decreases
* Thermal stability of sulphate	Decreases	Increases
* Thermal stability of carbonate	Decreases	Increases
* Thermal stability of nitrates	Decreases	Increases
* Thermal stability of hydroxides	Decreases	Increases
* Density	First increases then decreases	Increases
* Electropositivity	Decreases	Increases

ADDITIONAL EXAMPLES

Example 1 :

The boiling point of Kr and Rn are –152°C and –62°C respectively. Find the approximate boiling point of Xe.

Sol. According to law of triad, the properties of middle elements are average of rest two

∴ Boiling point of Xe in a triad of Kr, Xe Rn

= (boiling point of Kr + boiling point of Rn) / 2

$= (–152 – 62)/2 = –107°C$

Example 2 :

Which of the following atoms and ions are isoelectronic

(a) Al^{3+} (b) F   (c) Cl^- (d) O^{2-}

(e) Na   (f) Mg^{2+}

Sol. Ion or atom Al^{+3} F Cl^- O^{-2} Na Mg^{+2}

No. of e^- 10   9   18   10   11   10

So, Al^{+3}, O^{-2} & Mg^{+2} are isoelectronic.

Example 3 :

The bond distance between C–Cl in CCl_4 is 1.76 Å. If atomic radius of C is 0.77 Å, find the atomic radius of Cl.

Sol. Bond length of C–Cl = distance between the nuclei of two atoms $= r_C + r_{Cl}$

Given $r_C = 0.77$ and $r_C + r_{Cl} = 1.76$ Å

∴ $r_{Cl} = 1.76 – 0.77 = 0.99$ Å

Example 4 :

How much energy in joules must be needed to convert all the atoms of sodium to sodium ions present in 2.3 mg of sodium vapours ? Ionisation energy of sodium is 495 kJ $mole^{-1}$.

Sol. $Na(g) + IE \rightarrow Na^+(g) + e^-$ I.E. = 495 kJ mol^{-1}

The amount of energy needed to ionise 1 mole of sodium vapours = 495 kJ $mole^{-1}$

Moles of sodium vapours = $\dfrac{2.3 \times 10^{-3}}{23} = 1 \times 10^{-4}$ mol

Amount of energy needed to ionise 1×10^{-4} mol of sodium vapours are 495×10^{-4} kJ/mol.

Example 5 :

The IP_1 , IP_2, IP_3, IP_4 and IP_5 of an element are 7.1 , 14.3, 34.5, 46.8, 162.2 eV respectively. The elements is likely to be

(1) Na   (2) Si

(3) F   (4) Ca

Sol. (2). The jump in IP values exist in IP_5 and thus removal of fifth electron occurs from inner shell. Thus element contains four electrons in its valence shell.

Example 6 :

IP_1 and IP_2 of Mg are 178 and 348 kcal $mole^{-1}$. Find the energy required for the reaction   : $Mg \rightarrow Mg^{2+} + 2e^-$

Sol. Removal of two electrons (one by one) from an atom requires energy = $IP_1 + IP_2 = 178 + 348 = 526$ kcal.

Example 7 :

Calculate the effective nuclear charge (Z*) for valence shell electrons of bromine atom.

Sol. Electronic configuration of bromine atom is

$1s^2 \qquad 2s^2\,2p^6 \qquad 3s^2\,3p^6\,3d^{10} \qquad 4s^2\,4p^5$

K $\qquad$ L $\qquad$ M $\qquad$ N

For valence shell i.e. 4th shell electrons

$\qquad S = 6 \times 0.35 + 18 \times 0.85 + 10 \times 1.0 = 27.40$

$\qquad$ N-shell $\qquad$ M-shell $\qquad$ K and L shell

$Z^* = 35 - 27.40 = 7.60$

Example 8 :

The atomic number of three elements A,B and C are x , x + 1 and x + 2 , C is an alkali metal. In a compound of A and C, the nature of bonding is

(1) Co-ordinate $\qquad\qquad$ (2) Covalent

(3) Ionic $\qquad\qquad\qquad$ (4) Metallic

Sol. (3). If C is alkali metal, A should be halogen nonmetal.

Example 9 :

The ionization energy of lithium is 520 kJ mol^{-1}. Find the amount of energy required to convert 70 mg of lithium atoms in gaseous state into Li$^+$ ions.

Sol. $70\,\text{mg} = 70 \times 10^{-3}\text{g} = \dfrac{70 \times 10^{-3}}{7}\ \text{mol} = 1 \times 10^{-2}\ \text{mol}$

$\therefore$ Amount of energy required $= 1 \times 10^{-2} \times 520\,\text{kJ} = 5.2\,\text{kJ}$

Example 10 :

The ionization potential of X$^-$ ion is numerically equal to :

(1) The electron affinity of X atom

(2) The electronegativity of X atom

(3) The ionization potential of X

(4) None of these

Sol. (1). $X + e^- \rightarrow X^- + EA$; $X^- + IE \rightarrow X + e^-$

QUESTION BANK

CHAPTER 3 : CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

EXERCISE - 1 [LEVEL-1]

Choose one correct response for each question.

PART- 1 : GENESIS OF PERIODIC CLASSIFICATION

Q.1 The law of triads is applicable to -
(A) C, N, O (B) H, O, N
(C) Na, K, Rb (D) Cl, Br, I

Q.2 Periodic classification can be done on the basis of
(A) electronic configuration
(B) electronegativity
(C) valency
(D) Both (A) and (B)

Q.3 Law of octaves stated
(A) Every eighth element had properties similar to the first element.
(B) Every third element had properties similar to the first element.
(C) The properties of the middle element were in between the other two members.
(D) The properties of the elements were repeated after regular intervals of 3, 4 or 8 elements.

Q.4 Formula of oxide of Eka-Aluminium and Eka-Silion are respectively,
(A) E_2O_3, E_3O_2 (B) E_2O_3, EO_2
(C) E_2O_3 (D) Both (B) and (C)

Q.5 According to Mendeleev, iodine with lower atomic weight than that of ____ was placed in group-VII.
(A) bromine (B) chlorine
(C) fluorine (D) tellurium

Q.6 The first periodic law stated by Mendeleev was
(A) There is no correlation in the properties and atomic weights of the elements.
(B) The properties of the elements are a periodic function of their atomic numbers.
(C) The properties of the elements are a periodic function of their atomic weights.
(D) The properties of the elements are a periodic function of their empirical formula.

Q.7 Newlands Law of Octaves seemed to be true only for elements up to ____.
(A) calcium (B) argon
(C) potassium (D) zinc

Q.8 What were the main demerits of Mendeleev's periodic table ?
(i) Hydrogen has been placed in group I though it resembled to group VII as well.
(ii) Position of some elements was not justified.
(iii) Isotopes were not given separated places.
(iv) Lanthanides and actinides were not included in the table.
(A) i, ii and iii (B) i, ii, iii and iv
(C) ii and iv (D) i, iii and iv

PART- 2 : MODERN PERIODIC LAW

Q.9 In the modern periodic table, elements are arranged in
(A) Increasing mass
(B) Increasing volume
(C) Increasing atomic number
(D) Alphabetically

Q.10 In the modern long form of periodic table, elements are arranged in the increasing order of
(A) atomic mass
(B) atomic number
(C) mass number
(D) metallic character.

Q.11 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.

PART- 3 : NOMENCLATURE OF ELEMENTS WITH ATOMIC NUMBERS > 100

Q.12 According to IUPAC nomenclature, the name of element having atomic number 101 is
(A) Unnilbium (B) Unnilunium
(C) Unnilquadium (D) None of these

Q.13 What is the name and symbol of the element with atomic number 112 ?
(A) Ununbium, Uub (B) Unnilbium, Unb
(C) Ununnillum, Uun (D) Ununtrium, Uut

Q.14 The IUPAC official name of an element having atomic number 105
(A) Mendelevium (B) Lawrencium
(C) Dubnium (D) Nobelium

Q.15 According to IUPAC nomenclature, the name of element having atomic number 116 is
(A) Unnilunium (B) Ununoctium
(C) Ununhexium (D) Unnilhexium

PART- 4 : ELECTRONIC CONFIGURATIONS OF ELEMENTS

Q.16 Pd has exceptional valence shell electronic configuration of $4d^{10}5s^0$. It is a member of -
(A) 5th Period, Group 10
(B) 4th Period, Group 12
(C) 6th Period, Group 10
(D) 5th Period, Group 14

Q.17 The electronic configuration of an element is $1s^2, 2s^2 2p^6, 3s^2 3p^4$. The atomic number of the element present just below the above element in the periodic table -
(A) 36   (B) 34
(C) 33   (D) 32

Q.18 The element with configuration $1s^2, 2s^2 p^6, 3s^2$ would be
(A) A metal   (B) A non-metal
(C) An inert gas   (D) A metalloid

Q.19 $1s^2, 2s^2 p^6, 3s^2$ is the electronic configuration of the metal
(A) Na   (B) Mg
(C) Fe   (D) Al

Q.20 Beryllium has four electrons and has the electronic configuration
(A) $1s^2, 2s^2$ (B) $2s^2, 2p^4$
(C) $2s^2, 2p^2$ (D) $1s^2, 2p^2$

Q.21 An element has atomic number 79. Predict the group and period in which the element is placed.
(A) 2nd group, 7th period
(B) 11th group, 6th period
(C) 13th group, 6th period
(D) 12th group, 6th period

Q.22 The period to which an element belongs to in the long form of periodic table represents
(A) atomic mass
(B) atomic number
(C) principal quantum number
(D) azimuthal quantum number.

Q.23 3d transition series of elements starts from ____
(A) scandium   (B) vanadium
(C) calcium   (D) chromium

Q.24 Which of the following have the same number of electrons in outermost shell?
(A) Elements with atomic numbers 30, 48, 80
(B) Elements with atomic numbers 14, 15, 16
(C) Elements with atomic numbers 20, 30, 50
(D) Elements with atomic numbers 10, 18, 26

Q.25 Which of the following elements shown as pairs with their atomic numbers belong to the same period?
(A) Z = 19 and Z = 38   (B) Z = 12 and Z = 17
(C) Z = 11 and Z = 21   (D) Z = 16 and Z = 35

PART- 5 : TYPES OF ELEMENTS
s, p, d, f BLOCKS

Q.26 A metal has electronic configuration $[Ar]^{18} 3d^7 4s^2$. On the basis of this electronic configuration find out the group number.
(A) 9   (B) 2
(C) 6   (D) 4

Q.27 The d-block elements consists mostly of
(A) Monovalent metals
(B) All non-metals
(C) Elements which generally form stoichiometric metal oxide

(D) Many metals with catalytic properties

Q.28 s-block elements are reactive metals, with ____ ionisation enthalpy.
(A) high   (B) low
(C) zero   (D) None of these

Q.29 The non-metallic character of p-block element increases as they move from ____ across a period.
(A) right to left   (B) top to bottom
(C) bottom to top   (D) left to right

Q.30 Few general names are given along with their valence shell configurations. Mark the incorrect name.
(A) $ns^2 np^6$ – Noble gases
(B) $ns^2 np^5$ – Halogens
(C) ns^1 – Alkali metals
(D) $ns^2 np^2$ – Chalcogens

Q.31 The elements in which electrons are progressively filled in 4f-orbital are called
(A) actinoids   (B) transition elements
(C) lanthanoids   (D) halogens

Q.32 The compounds of the s-block elements, with the exception of lithium and __ are predominantly ionic.
(A) hydrogen   (B) helium
(C) magnesium   (D) beryllium

Q.33 General outer electronic configuration of d-block elements is
(A) $(n-1)d^{1-10} ns^3$ (B) $(n+1) d^{1-10} ns^{0-2}$
(C) $(n-1)d^{1-10} ns^{0-2}$ (D) $(n-1) d^0 ns^{0-2}$

Q.34 Which block of the periodic table contains elements with the general electronic configuration $(n-2) f^{1-14} (n-1) d^{0-1} ns^2$?
(A) s–block   (B) p–block
(C) d–block   (D) f–block

Q.35 The last element of the p-block in 6th period is represented by the outermost electronic configuration.
(A) $7s^2 7p^6$ (B) $5f^{14} 6d^{10} 7s^2 7p^0$
(C) $4f^{14} 5d^{10} 6s^2 6p^6$ (D) $4f^{14} 5d^{10} 6s^2 6p^4$

Q.36 Part of the periodic table showing p-block is depicted below. What are the elements shown in the zig–zag boxes called? What is the nature of the elements outside this boundary on the right side of the table?
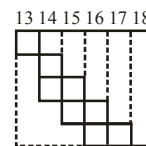

(A) Transition elements, metalloids
(B) Metalloids, non–metals
(C) Metals, non–metals
(D) Non–metals, noble gases

Q.37 Actinoid elements are
(A) radioactive   (B) metals
(C) non-metals   (D) metalloids

Q.38 Atomic numbers of few elements are given below. Which of the pairs belongs to s–block?
(A) 7, 14   (B) 3, 20
(C) 8, 15   (D) 9, 17

PART- 6 : PERIODIC TRENDS IN PROPERTIES OF ELEMENTS

Q.39 Which is correct in the following -
(A) Radius of Cl atoms is 0.99 Å, while that of Cl^+ ion is 1.54 Å.
(B) Radius of Cl atoms is 0.99 Å, while that of Na atom is 1.54 Å.
(C) Radius of Cl atoms is 0.95 Å, while that of Cl^- ion is 0.81Å.
(D) Radius of Na atoms is 0.95 Å, while that of Na^+ ion is 1.54 Å.

Q.40 Beryllium shows diagonal relationship with
(A) Mg   (B) Na
(C) B   (D) Al

Q.41 The radii of F, F^-, O and O^{-2} are in the order of
(A) $O^{-2} > F^- > O > F$ (B) $O^{-2} > F^- > F > O$
(C) $F^- > O^{-2} > F > O$ (D) $O^{-2} > O > F^- > F$

Q.42 Which of the following element has the lowest ionization potential
(A) Fe   (B) H
(C) Li   (D) He

Q.43 The first ionisation potential will be maximum for
(A) Lithium   (B) Hydrogen
(C) Uranium   (D) Iron

Q.44 The electron affinity values for the halogens show the following trend
(A) $F < Cl > Br > I$ (B) $F < Cl < Br < I$
(C) $F > Cl > Br > I$ (D) $F < Cl > Br < I$

Q.45 Which of the following is the most electropositive element
(A) Aluminium   (B) Magnesium
(C) Phosphorus (D) Sulphur

Q.46 Which of the following property displays progressive increase with the rise in atomic number across a period in the periodic table
(A) Electronegativity   (B) Electron affinity
(C) Ionization potential (D) Size of the atom

Q.47 Choose the correct statement
(A) Electronegativity increases down a group.
(B) Electronegativity decreases down a group.
(C) Electronegativity decreases from left to right along a period.
(D) Electronegativity changes along a group but remains constant along a period.

Q.48 Metallic character in periods of the Periodic Table increases from
(A) left to right   (B) right to left
(C) Either (A) and (B) (D) Data insufficient

Q.49 Which is the most non–metallic element among the following?
(A) $1s^2\ 2s^2\ 2p^6\ 3s^1$ (B) $1s^2\ 2s^2\ 2p^5$
(C) $1s^2\ 2s^2\ 2p^6\ 3s^2$ (D) $1s^2\ 2s^2\ 2p^3$

Q.50 In the given graph, a periodic property (R) is plotted against atomic numbers (Z) of the elements. Which property is shown in the graph and how is it correlated with reactivity of the elements?
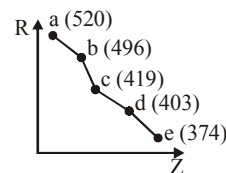
(A) Ionisation enthalpy in a group, reactivity decreases from a → e.
(B) Ionisation enthalpy in a group, reactivity increases from a → e.
(C) Atomic radius in a group, reactivity decreases from a → e.
(D) Metallic character in a group, reactivity increases from a → e.

Q.51 In the Periodic Table metallic character of elements shows one of the following trend
(A) decreases down the group and increases across the period.
(B) increases down the group and decreases across the period
(C) increases across the period and also down the group.
(D) decreases across the period and also down the group.

Q.52 Few values are given in the table in the direction from left to right and top to bottom. Predict the property which could be depicted in the table.

152						
186	160	143	117	110	104	99
231						
244						
262						

(A) Atomic number
(B) Ionisation enthalpy
(C) Electron gain enthalpy
(D) Atomic radius

Q.53 As we move from left to right, the electro negativity increases. An atom which is highly electronegative has
(A) large size
(B) low electron affinity
(C) high ionisation enthalpy
(D) low chemical reactivity.

Q.54 Arrange the following atoms in order of increasing first ionization enthalpy K, Cs, Rb, Ca
(A) Cs, Rb, K, Ca   (B) Rb, Cs, K, Ca
(C) Cs, Rb, Ca, K   (D) Rb, Cs, Ca, K

Q.55 Arrange the following elements in the increasing order of metallic characters. Si, Be, Mg, Na and P. Select the correct answer.
(A) Na > Be > Mg > P > Si   (B) Na > Mg > Be > P > Si
(C) Na > Be > P > Mg > Si   (D) Na > Mg > Be > Si > P

Q.56 Ionization enthalpies of transition metals are –
(A) intermediate between those of s- and p-block elements.
(B) more than p-block elements
(C) highest in all the elements
(D) lower than s-block elements.

Q.57 The first ionisation enthalpy of the elements are in the order of
(A) $C < N < Si < P$　　　(B) $N < Si < C < P$
(C) $Si < P < C < N$　　　(D) $P < Si < N < C$

Q.58 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) $Na > Mg > Al < Si$

Q.59 Which of the following is not correct statement for periodic classification of elements?
(A) The properties of elements are the periodic functions of their atomic number.
(B) Non-metallic elements are less in number than metallic elements.
(C) The first ionisation energies of elements along a period do not vary in regular manner with increase in atomic number.
(D) For transition elements, the last electron enters into $(n – 2)$ d–subshell.

Q.60 The first ionisation enthalpy generally increases as we go across _____.
(A) group　　　　　　(B) period
(C) vertical column　　(D) Both (A) and (B)

Q.61 Why do noble gases have positive electron gain enthalpy?
(A) It is difficult to add an electron due to small size.
(B) It is difficult to add an electron due to high electronegativity.
(C) It is difficult to add an electron due to stable configuration.
(D) It is difficult to add an electron due to high electron affinity.

Q.62 Which of the following statements is not correct about the electron gain enthalpy?
(A) In general, the electron gain enthalpy becomes less negative in going from top to bottom in a group.
(B) The electron gain enthalpy becomes less negative in a period from left to right.
(C) The elements having stable configuration like noble gases have large positive electron gain enthalpies.
(D) Electron gain enthalpy of O or F is less than that of succeeding element.

Q.63 The correct sequence which shows decreasing order of the ionic radii of the elements is
(A) $Al^{3+} > Mg^{2+} > Na^+ > F^- > O^{2-}$
(B) $Na^+ > Mg^{2+} > Al^{3+} > O^{2-} > F^-$
(C) $Na^+ > F^- > Mg^{2+} > O^{2-} > Al^{3+}$
(D) $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

Q.64 Which of the following can most easily form unipositive gaseous ion?
(A) $1s^2\,2s^2\,2p^6\,3s^2$　　(B) $1s^2\,2s^2\,2p^6\,3s^2\,3p^1$
(C) $1s^2\,2s^2\,2p^6\,3s^2\,3p^2$　(D) $1s^2\,2s^2\,2p^6\,3s^2\,3p^3$

EXERCISE - 2 [LEVEL-2]

Choose one correct response for each question.

Q.1 Which of the following relation is correct -
(A) $2\,I.P. – E.A. – E.N. = 0$
(B) $2\,I.P. – E.N. + E.A. = 0$
(C) $2\,E.N. – I.P. – E.A. = 0$
(D) $E.N. – I.P. – E.A. = 0$

Q.2 A given compound A_2 whose total $d_{A–A}$ is 1.4Å. The atomic (covalent) radius of an atom A is –
(A) 0.7 Å　　　　(B) 0.5 Å
(C) 0.2 Å　　　　(D) 0.1 Å

Q.3 Choose the correct statements –
(1) Element with lowest electronegativity is Cs.
(2) Element with highest electronegativity is F.
(3) Element with highest ionisation potential is He.
(4) Element with lowest ionisation potential is Hg.
(A) 1, 2 & 3 are correct (B) 1 & 2 are correct
(C) 2 & 4 are correct　　(D) 1 & 3 are correct

Q.4 Which of the following is the artificial element in the periodic table –
(A) Tc　　　　　(B) Te
(C) Ru　　　　　(D) Os

Q.5 Which of the following is not a transition element
(A) Co　　　　　(B) Ni
(C) Mn　　　　　(D) Zn

Q.6 The elements with atomic number 10, 18, 36, 54 & 86 are all
(A) Light metals　　(B) Inert gases
(C) Halogens　　　(D) Rare-earths

Q.7 Which of the following statement is not correct regarding hydrogen atom –
(A) It resembles halogens in some properties.
(B) It resembles alkali metals in some properties.
(C) It can be placed in 7th group of periodic table.
(D) It can not be placed in first group of periodic table.

Q.8 Which has the smallest size
(A) Na^+　　　　(B) Mg^{2+}
(C) Al^{3+}　　　　(D) P^{5+}

Q.9 Ionization potential is lowest for
(A) Halogens　　　(B) Inert gases
(C) Alkaline earth metals　(D) Alkali metals

Q.10 How many ionisation energies can carbon have
(A) 1　　　　　　(B) 2
(C) 4　　　　　　(D) 6

Q.11 Mg and Li are similar in their properties due to
(A) Same e/m ratio　　(B) Same electron affinity
(C) Same group　　　(D) Same ionic potential

Q.12 Ionic compounds are formed most easily with
(A) Low electron affinity, high ionisation energy
(B) High electron affinity, low ionisation energy
(C) Low electron affinity, low ionisation energy
(D) High electron affinity, high ionisation energy

Q.13 Which one of the elements has the maximum electron affinity
(A) F (B) Cl
(C) Br (D) I

Q.14 Among Al_2O_3, SiO_2, P_2O_3 and SO_2 the correct order of acid strength is –
(A) $Al_2O_3 < SiO_2 < SO_2 < P_2O_3$
(B) $SiO_2 < SO_2 < Al_2O_3 < P_2O_3$
(C) $SO_2 < P_2O_3 < SiO_2 < Al_2O_3$
(D) $Al_2O_3 < SiO_2 < P_2O_3 < SO_2$

Q.15 Ionic radii vary in –
(A) inverse proportion to the effective nuclear charge.
(B) inverse proportion to the square of effective nuclear charge.
(C) inverse proportion to the screening effect.
(D) direct proportion to the square of the screening effect.

Q.16 What is the common property of the oxides CO, NO and N_2O?
(A) All are acidic oxides.
(B) All are basic oxides.
(C) All are neutral oxides.
(D) All are amphoteric oxides.

Q.17 Which group of elements shows lowest ionisation enthalpy?
(A) Alkali metals (B) Halogens
(C) Alkaline earth metals (D) Noble gases

Q.18 Which of the following statements is correct?
(A) Elements of 3d and 4d–series are kept separately in periodic table.
(B) Elements of 4f and 5f–series are kept separately in periodic table.
(C) Elements of 5p and 6p–series are kept separately in periodic table.
(D) All statements are correct.

Q.19 Which of the following is the correct IUPAC name of an element having atomic number 102?
(A) Unnilium (B) Ununnilium
(C) Unnilunium (D) None of the above

Q.20 The sizes of the following species increase in the order
(A) $Mg^{2+} < Na^+ < F^- < Al$ (B) $F^- < Al < Na^+ < Mg^{2+}$
(C) $Al < Mg^{2+} < F^- < Na^+$ (D) $Na^+ < Al < F^- < Mg^{2+}$

Q.21 Examples of elements belonging to s, p, d or f-block are given. Identify the wrong example.
(A) s–block element – Caesium
(B) p–block element – Barium
(C) d–block element – Chromium
(D) f–block element – Thorium

Q.22 Choose the incorrect statement.
(A) Chemical reactivity tends to be high in group 1 metals, lower in elements in middle and increases to maximum in the group 17.

(B) Halogens have very high negative electron gain enthalpy.
(C) Noble gases have large positive electron gain enthalpy.
(D) Decrease in electronegativities across a period is accompanied by an increase in non-metallic properties.

Q.23 In the given compound OF_2, the oxidation state of oxygen is
(A) –2 (B) +1
(C) +2 (D) –1

Q.24 Ionisation enthalpy of nitrogen is more than oxygen because of –
(A) extra stability of half filled orbitals.
(B) more number of energy levels.
(C) less number of valence electrons.
(D) smaller size.

Q.25 The lattice energies of the following have the order
(A) $BaO < SrO < CaO < MgO$
(B) $BaO < CaO < SrO < MgO$
(C) $MgO < CaO < SrO < BaO$
(D) $MgO < BaO < SrO < CaO$

Q.26 The electronic states X and Y of an atom are depicted:
X : $1s^2\, 2s^2\, 2p^6\, 3s^1$
Y : $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^1$
Which of the following statements is not correct?
(A) X represents an alkali metal.
(B) Energy is required to change X into Y.
(C) Y represents ground state of the element.
(D) Less energy is required to remove an electron from X than from Y.

Q.27 Which of the following is correct ?
(A) Ionization energy of noble gas is highest in respective period
(B) Ionization energy of group-I metals is lowest in respective period
(C) Ionization energy of Al and Ga is almost same
(D) All are correct

Q.28 The decreasing order of the ionisation potential of the following elements is
(A) $Ne > Cl > P > S > Al > Mg$
(B) $Ne > Cl > P > S > Mg > Al$
(C) $Ne > Cl > S > P > Mg > Al$
(D) $Ne > Cl > S > P > Al > Mg$

Q.29 The correct order of radii is –
(A) $N < Be < B$ (B) $F^- < O^{2-} < N^{3-}$
(C) $Na < Li < K$ (B) $Fe^{3+} < Fe^{2+} < Fe^{4+}$

Q.30 Of the metals Be, Mg, Ca and Sr of group 2 in the periodic table, the least ionic chloride will be formed by
(A) Be (B) Ca
(C) Mg (D) Sr

Q.31 The ionization energy and electron affinity of an element are 13.0 eV and 3.8 eV respectively. Its electronegativity on the Pauling scale is
(A) 3.0 (B) 3.5
(C) 4.0 (D) 2.8

Q.32 The second ionization enthalpy of an element M is the energy required to
(A) Remove one mole of electrons from one mole of gaseous cations of the element.
(B) Remove one mole of electrons from one mole of gaseous anions.
(C) Remove one mole of electrons from one mole of monovalent gaseous cations of the element.
(D) Remove 2 moles of electrons from one mole of gaseous atoms.

Q.33 Indicate the wrong statement on the basis of the periodic table.
(A) The most electronegative element in the periodic table is fluorine.
(B) Scandium is the first transition element and belongs to fourth period.
(C) There are three transition series in the periodic table each containing 10 elements.
(D) Along a period halogens have maximum negative electron gain enthalpy.

Q.34 There are two rows of inner transition elements in the periodic table each containing 14 elements. The reason for this may be –
(A) f-orbital has seven values for magnetic quantum number, hence total electrons are 14.
(B) in the periodic table there is space to accommodate 14 electrons only.
(C) only 28 inner transition elements have been discovered till date.
(D) 28 is the maximum number of elements that any block can accommodate.

EXERCISE - 3 (NUMERICAL VALUE BASED QUESTIONS)

NOTE : The answer to each question is a NUMERICAL VALUE.

Q.1 Number of Lanthanides is –

Q.2 If there were 10 periods in the periodic table then how many elements would this period can maximum comprise of.

Q.3 Integral value of E.N of fluorine if $(r_F)_{covalent} = 0.72$ Å is

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

Q.5 Integral value of E.N. of Cl from the bond energy ClF (61 KCal/gm). Given that bond energies of F_2 and Cl_2 are 38 and 58 KCal/mol respectively is –

Q.6 I.P. of Be^{+x} is found to be 217.6 electron volt. What is the value of x.

Q.7 Integral value of the screening constant of Ca. (at. number 20) is –

Q.8 Outer configuration of X is $3d^5$, $4s^1$. It belongs to group number –

Q.9 The atomic number of the third alkali metal is

Q.10 What is the period number of the element whose atomic number is 98.

EXERCISE - 4 [PREVIOUS YEARS AIEEE / JEE MAIN QUESTIONS]

Q.1 The correct order of ionic radius is -   **[AIEEE-2002]**
(A) $Ce > Sm > Tb > Lu$ (B) $Lu > Tb > Sm > Ce$
(C) $Tb > Lu > Sm > Ce$ (D) $Sm > Tb > Lu > Ce$

Q.2 Ce^{3+}, La^{3+}, Pm^{3+} and Yb^{3+} have ionic radii in the increasing order as –   **[AIEEE-2002]**
(A) $La^{3+} < Ce^{3+} < Pm^{3+} < Yb^{3+}$
(B) $Yb^{3+} < Pm^{3+} < Ce^{3+} < La^{3+}$
(C) $La^{3+} = Ce^{3+} < Pm^{3+} < Yb^{3+}$
(D) $Yb^{3+} < Pm^{3+} < La^{3+} < Ce^{3+}$

Q.3 According to the Periodic Law of elements, the Variation in properties of elements is related to their **[AIEEE-2003]**
(A) Nuclear masses
(B) Atomic numbers
(C) Nuclear neutron-proton number ratio
(D) Atomic masses

Q.4 The reduction in atomic size with increase in atomic number is a characteristic of elements of -**[AIEEE-2003]**
(A) d-block   (B) f-block
(C) Radioactive series   (D) High atomic masses

Q.5 Which one of the following groups represent a collection of isoelectronic species ?
(At. no. Cs = 55, Br = 35)   **[AIEEE-2003]**
(A) N^{3-}, F^-, Na^+ (B) Be, Al^{3+}, Cl^-
(C) Ca^{2+}, Cs^+, Br (D) Na^+, Ca^{2+}, Mg^{2+}

Q.6 The atomic numbers of vanadium (V). Chromium (Cr), manganese (Mn) and iron (Fe) respectively 23, 24, 25 and 26. Which one of these may be expected to have the higher second ionization enthalpy ?   **[AIEEE-2003]**
(A) Cr   (B) Mn
(C) Fe   (D) V

Q.7 Which one of the following sets of ions represents the collection of isoelectronic species ?   **[AIEEE-2004]**
(A) K^+, Cl^-, Mg^{2+}, Sc^{3+} (B) Na^+, Ca^{2+}, Sc^{3+}, F^-
(C) K^+, Ca^{2+}, Sc^{3+}, Cl^- (D) Na^+, Mg^{2+}, Al^{3+}, Cl^-

Q.8 Which one of the following ions has the highest value of ionic radius ?   **[AIEEE-2004]**
(A) O^{2-} (B) B^{3+}
(C) Li^+ (D) F^-

Q.9 Among Al_2O_3, SiO_2, P_2O_3 and SO_2 the correct order of acid strength is :   **[AIEEE-2004]**
(A) $Al_2O_3 < SiO_2 < SO_2 < P_2O_3$
(B) $SiO_2 < SO_2 < Al_2O_3 < P_2O_3$
(C) $SO_2 < P_2O_3 < SiO_2 < Al_2O_3$
(D) $Al_2O_3 < SiO_2 < P_2O_3 < SO_2$

Q.10 The formation of the oxide ion $O^{2-}_{(g)}$ requires first an exothermic and then an endothermic step as shown below
$O_{(g)} + e^- = O^-_{(g)}$ $\Delta H° = - 142$ kJ mol^{-1}
$O^-_{(g)} + e^- = O^{2-}_{(g)}$ $\Delta H° = 844$ kJ mol^{-1}
This is because of :   **[AIEEE-2004]**
(A) O^- ion will tend to resist the addition of another e^-
(B) Oxygen has high electron affinity
(C) Oxygen is more electronegative
(D) O^- ion has comparatively larger size than oxygen atom

Q.11 In which of the following arrangements the order is NOT according to the property indicated against it ?   **[AIEEE-2005]**
(A) $Al^{3+} < Mg^{2+} < Na < F^-$ – increasing ionic size
(B) $B < C < N < O$ – increasing first ionization enthalpy
(C) $I < Br < F < Cl$ – increasing electron gain enthalpy (with negative sign)
(D) $Li < Na < K < Rb$ – increasing metallic radius

Q.12 Which of the following oxides is amphoteric in character
(A) SnO_2 (B) SiO_2 **[AIEEE-2005]**
(C) CO_2 (D) CaO

Q.13 Pick out the isoelectronic structure from the following :   **[AIEEE-2005]**
I. $^+CH_3$ II. H_3O^+ III. NH_3 IV. CH_3^-
(A) I and II   (B) III and IV
(C) I and III   (D) II, III and IV

Q.14 The lanthanide contraction is responsible for the fact that   **[AIEEE-2005]**
(A) Zr and Y have about the same radius
(B) Zr and Nb have similar oxidation state
(C) Zr and Hf have about the same radius
(D) Zr and Zn have the same oxidation state

Q.15 Which of the following factors may be regarded as the main cause the lanthanide contraction ? **[AIEEE-2005]**
(A) Poor shielding of one of 4f electron by another in the subshell
(B) Effective shielding of one of 4f electrons by another in the subshell
(C) Poorer shielding of 5d electrons by 4f electrons
(D) greater shielding of 5d electrons by 4f electrons

Q.16 The increasing order of the first ionization enthalpies of the elements B, P, S and F (lowest first) is **[AIEEE-2006]**
(A) $F < S < P < B$ (B) $P < S < B < F$
(C) $B < P < S < F$ (D) $B < S < P < F$

Q.17 Which one of the following sets of ions represents a collection of isoelectronic species ?   **[AIEEE-2006]**
(A) N^{3-}, O^{2-}, F^-, S^{2-} (B) Li^+, Na^+, Mg^{2+}, Ca^{2+}
(C) K^+, Cl^-, Ca^{2+}, Sc^{3+} (D) Ba^{2+}, Sr^{2+}, K^+, Ca^{2+}

Q.18 Lanthanoid contraction is caused due to **[AIEEE-2006]**
(A) The same effective nuclear charge from Ce to Lu.
(B) The imperfect shielding on outer electrons by 4f electrons from the nuclear charge.
(C) The appreciable shielding on outer electrons by 4f electrons from the nuclear charge.
(D) The appreciable shielding on outer electrons by 5d electrons from the nuclear charge.

Q.19 The charge/size ratio of a cation determines its polarzing power. Which one of the following sequences represents the increasing order of the polarizing power of the cationic species, K^+, Ca^{+2} Mg^{+2}, Be^{+2}
(A) $Be^{+2} < K^+ < Ca^{+2} < Mg^{+2}$ **[AIEEE-2007]**
(B) $K^+ < Ca^{+2} < Mg^{+2} < Be^{+2}$
(C) $Ca^{+2} < Mg^{+2} < Be^{+2} < K^+$
(D) $Mg^{+2} < Be^{+2} < K^+ < Ca^{+2}$

Q.20 The set representing the correct order of ionic radius is:
(A) $Mg^{2+} > Be^{2+} > Li^+ > Na^+$ **[AIEEE-2009]**
(B) $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$
(C) $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$
(D) $Li^+ > Na^+ > Mg^{2+} > Be^{2+}$

Q.21 The correct sequence which shows decreasing order of the ionic radii of the elements is –   **[AIEEE 2010]**
(A) $Al^{3+} > Mg^{2+} > Na^+ > F^- > O^{2-}$
(B) $Na^+ > Mg^{2+} > Al^{3+} > O^{2-} > F^-$
(C) $Na^+ > F^- > Mg^{2+} > O^{2-} > Al^{3+}$
(D) $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

Q.22 The outer electron configuration of Gd (Atomic No : 64) is   **[AIEEE 2011]**
(A) $4f^3\, 5d^5\, 6s^2$ (B) $4f^8\, 5d^0\, 6s^2$
(C) $4f^4\, 5d^4\, 6s^2$ (D) $4f^7\, 5d^1\, 6s^2$

Q.23 Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar?   **[JEE MAIN 2013]**
(A) $Ca < S < Ba < Se < Ar$ (B) $S < Se < Ca < Ba < Ar$
(C) $Ba < Ca < Se < S < Ar$ (D) $Ca < Ba < S < Se < Ar$

Q.24 The ionic radii (in Å) of N^{3-}, O^{2-} and F^- are respectively   **[JEE MAIN 2015]**
(A) 1.36, 1.71 and 1.40   (B) 1.71, 1.40 and 1.36
(C) 1.71, 1.36 and 1.40   (D) 1.36, 1.40 and 1.71

Q.25 Which of the following atoms has the highest first ionization energy?   **[JEE MAIN 2016]**
(A) Na   (B) K
(C) Sc   (D) Rb

Q.26 The group having isoelectronic species is:
    [JEE MAIN 2017]
(A) O^-, F^-, Na^+, Mg^{2+} (B) O^{2-}, F^-, Na^+, Mg^{2+}
(C) O^-, F^-, Na, Mg^+ (D) O^{2-}, F^-, Na, Mg^{2+}

Q.27 In general, the properties that decrease and increase down a group in the periodic table, respectively, are :
    [JEE MAIN 2019 (JAN)]
(A) electronegativity and electron gain enthalpy.
(B) electronegativity and atomic radius.
(C) atomic radius and electronegativity.
(D) electron gain enthalpy and electronegativity.

Q.28 The IUPAC symbol for the element with atomic number 119 would be :   **[JEE MAIN 2019 (APRIL)]**
(A) unh   (B) uun
(C) une   (D) uue

Q.29 The correct order of catenation is :
    [JEE MAIN 2019 (APRIL)]
(A) $C > Si > Ge \approx Sn$ (B) $C > Sn > Si \approx Ge$
(C) $Ge > Sn > Si > C$ (D) $Si > Sn > C > Ge$

Q.30 The group number, number of valence electrons, and valency of an element with atomic number 15, respectively, are   **[JEE MAIN 2019 (APRIL)]**
(A) 16, 5 and 2   (B) 16, 6 and 3
(C) 15, 5 and 3   (D) 15, 6 and 2

Q.31 The atomic radius of Ag is closest to :
    [JEE MAIN 2020 (JAN)]
(A) Cu   (B) Hg
(C) Au   (D) Ni

Q.32 Which one of the following amongs each pair will release maximum energy on gaining one electron
(a = F, Cl), (b = S, Se), (c = Li, Na)
    [JEE MAIN 2020 (JAN)]
(A) a = Cl, b = S, c = Li   (B) a = F, b = Se, c = Li
(C) a= F, b = Se, c = Na   (D) a = Cl, b = S, c = Na

Q.33 Correct order of first ionization energy of the following metals Na, Mg, Al, Si in KJ mol^{-1} respectively are:
    [JEE MAIN 2020 (JAN)]
(A) 496, 737, 577, 786   (B) 496, 577, 737, 786
(C) 786, 739, 577, 497   (D) 739, 577, 786, 487

Q.34 Among the following least 3rd ionization energy is for
    [JEE MAIN 2020 (JAN)]
(A) Mn   (B) Co
(C) Fe   (D) Ni

Q.35 Correct order of radius of elements is: C, O, F, Cl, Br
    [JEE MAIN 2020 (JAN)]
(A) $Br > Cl > C > O > F$ (B) $Br < Cl < C < O < F$
(C) $Cl < C < O < F < Br$ (D) $C > F > O > Br > Cl$

EXERCISE - 5 (PREVIOUS YEARS AIPMT/NEET EXAM QUESTIONS)

Choose one correct response for each question.

Q.1 Which one of the following arrangements represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species ? **[AIPMT 2005]**
(A) $S < O < Cl < F$
(B) $Cl < F < S < O$
(C) $F < Cl < O < S$
(D) $O < S < F < Cl$

Q.2 Four successive members of the first row transition elements are listed below with their atomic numbers. Which one of them is expected to have the highest third ionization enthalpy – **[AIPMT 2005]**
(A) Vanadium $(Z = 23)$
(B) Chromium $(Z = 24)$
(C) Manganese $(Z = 25)$
(D) Iron $(Z = 26)$

Q.3 Which orders is not in in accordance with the property stated against is – **[AIPMT 2006]**
(A) Acidic property in water : $HI > HBr > HCl > HF$
(B) Electronegativity : $F_2 > Cl_2 > Br_2 > I_2$
(C) Bond dissociation energy : $F_2 > Cl_2 > Br_2 > I_2$
(D) Oxidising power : $F_2 > Cl_2 > Br_2 > I_2$

Q.4 Identify the correct order of the size of the following –
(A) $Ca^{2+} < K^+ < Ar < Cl^- < S^{2-}$ **[AIPMT 2007]**
(B) $Ar < Ca^{2+} < K^+ < Cl^- < S^{2-}$
(C) $Ca^{2+} < Ar < K^+ < Cl^- < S^{2-}$
(D) $Ca^{2+} < K^+ < Ar < S^{2-} < Cl^-$

Q.5 Which of the following electronic configuration an atom has the lowest ionisation enthalphy – **[AIPMT 2007]**
(A) $1s^2\ 2s^2\ sp^3$
(B) $1s^2\ 2s^2\ 2p^5\ 3s^1$
(C) $1s^2\ 2s^2\ 2p^6$
(D) $1s^2\ 2s^2\ 2p^5$

Q.6 The correct order of decreasing second ionisation enthalpy of Ti (22), V (23), Cr (24) and Mn (25), is: **[AIPMT 2008]**
(A) $Ti > V > Cr > Mn$
(B) $Cr > Mn > V > Ti$
(C) $V > Mn > Cr > Ti$
(D) $Mn > Cr > Ti > V$

Q.7 Which one of the following arrangement does not give the correct pictures of the trends indicated against it?
(A) $F_2 > Cl_2 > Br_2 > I_2$: Electronegativity
(B) $F_2 > Cl_2 > Br_2 > I_2$: Oxidizing power **[AIPMT 2008]**
(C) $F_2 > Cl_2 > Br_2 > I_2$: Electron gain enthalpy
(D) $F_2 > Cl_2 > Br_2 > I_2$: Bond dissociation energy

Q.8 Which one of the elements with the following outer orbital configurations may exhibit the largest number of oxidation states? **[AIPMT 2009]**
(A) $3d^5 4s^1$
(B) $3d^5 4s^2$
(C) $3d^2 4s^2$
(D) $3d^3 4s^2$

Q.9 Amongst the elements with following electronic configurations, which one of them may have the highest ionization energy? **[AIPMT 2009]**
(A) Ne $[3s^2 3p^2]$
(B) Ar $[3d^{10} 4s^2 4p^3]$
(C) Ne $[3s^2 3p^1]$
(D) Ne $[3s^2 3p^3]$

Q.10 The correct order of the decreasing ionic radii among the following electronic species **[AIPMT (PRE) 2010]**
(A) $Ca^{2+} > K^+ > S^{2-} > Cl^-$
(B) $Cl^- > S^{2-} > Ca^{2+} > K^+$
(C) $S^{2-} > Cl^- > K^+ > Ca^{2+}$
(D) $K^+ > Ca^{2+} > Cl^- > S^{2-}$

Q.11 Which of the following represents the correct order of increasing electron gain enthalpy with negative sign for the elements O, S, F and Cl ? **[AIPMT (PRE) 2010]**
(A) $Cl < F < O < S$
(B) $O < S < F < Cl$
(C) $F < S < O < Cl$
(D) $S < O < Cl < F$

Q.12 Among the elements Ca, Mg, P and Cl, the order of increasing atomic radii is : **[AIPMT (MAIN) 2010]**
(A) $Mg < Ca < Cl < P$
(B) $Cl < P < Mg < Ca$
(C) $P < Cl < Ca < Mg$
(D) $Ca < Mg < P < Cl$

Q.13 Among the following which one has the highest cation to anion size ratio? **[AIPMT (MAIN) 2010]**
(A) CsI
(B) CsF
(C) LiF
(D) NaF

Q.14 What is the value of electron gain enthalpy of Na^+ if IE_1 of Na = 5.1 eV ? **[AIPMT (MAIN) 2011]**
(A) -5.1 eV
(B) -10.2 eV
(C) $+2.55$ eV
(D) $+10.2$ eV

Q.15 Identify the wrong statement in the following : **[AIPMT (PRE) 2012]**
(A) Amongst isoelectronic species, smaller the positive charge on the cation, smaller is the ionic radius.
(B) Amongst isoelectronic species, greater the negative charge on the anion, larger is the ionic radius.
(C) Atomic radius of the elements increases as one moves down the first group of the periodic table.
(D) Atomic radius of the elements decreases as one moves across from left to right in the 2nd period of the periodic table.

Q.16 Which of the following orders of ionic radii is correctly represented? **[AIPMT 2014]**
(A) $H^- > H > H^+$
(B) $O^{2-} > F^- > Na^+$
(C) $N^{-3} > Mg^{2+} > Al^{+3}$
(D) All of these

Q.17 The species Ar, K^+ and Ca^{2+} contain the same number of electrons. In which order do their radii increase? **[AIPMT 2015]**
(A) $Ca^{2+} < Ar < K^+$
(B) $Ca^{2+} < K^+ < Ar$
(C) $K^+ < Ar < Ca^{2+}$
(D) $Ar < K^+ < Ca^{2+}$

Q.18 In which of the following options, the order of arrangement does not agree with the variation of property indicated against it? **[NEET 2016 PHASE 1]**
(A) $Al^{3+} < Mg^{2+} < Na^+ < F^-$ (increasing ionic size)
(B) $B < C < N < O$ (increasing first ionisation enthalpy)
(C) $I < Br < Cl < F$ (increasing electron gain enthalpy)
(D) $Li < Na < K < Rb$ (increasing metallic radius)

Q.19 The element Z = 114 has been discovered recently. It will belong to which of the following family/group and electronic configuration ? **[NEET 2017]**
(A) Carbon family, $[Rn]\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^2$
(B) Oxygen family, $[Rn]\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^4$
(C) Nitrogen family, $[Rn]\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^6$
(D) Halogen family, $[Rn]\ 5f^{14}\ 6d^{10}\ 7s^2\ 7p^5$

Q.20 Magnesium reacts with an element (X) to form an ionic compound. If the ground state electronic configuration of (X) is $1s^2\ 2s^2\ 2p^3$, the simplest formula for this compound is **[NEET 2018]**
(A) Mg_2X
(B) MgX_2
(C) Mg_2X_3
(D) Mg_3X_2

Q.21 Which is the correct thermal stability order for H_2E (E = O, S, Se, Te and Po)? **[NEET 2019]**
(A) $H_2S < H_2O < H_2Se < H_2Te < H_2Po$
(B) $H_2O < H_2S < H_2Se < H_2Te < H_2Po$
(C) $H_2Po < H_2Te < H_2Se < H_2S < H_2O$
(D) $H_2Se < H_2Te < H_2Po < H_2O < H_2S$

Q.22 For the second period elements the correct increasing order of first ionisation enthalpy is: **[NEET 2019]**
(A) $Li < Be < B < C < N < O < F < Ne$
(B) $Li < B < Be < C < O < N < F < Ne$
(C) $Li < B < Be < C < N < O < F < Ne$
(D) $Li < Be < B < C < O < N < F < Ne$

ANSWER KEY

EXERCISE - 1

Q	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25
A	D	A	A	B	D	C	A	B	C	B	C	B	A	C	C	A	B	A	B	A	B	C	A	A	B
Q	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45	46	47	48	49	50
A	A	D	B	D	D	C	D	C	D	C	B	A	B	B	D	A	C	B	A	B	A	B	B	B	B
Q	51	52	53	54	55	56	57	58	59	60	61	62	63	64											
A	B	D	C	A	D	A	C	A	D	B	C	B	D	B											

EXERCISE - 2

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

EXERCISE - 3

Q	1	2	3	4	5	6	7	8	9	10
A	14	72	4	3	3	3	17	6	19	7

EXERCISE - 4

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

EXERCISE - 5

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

PERIODIC TABLE
TRY IT YOURSELF

(1) (i) As the principal quantum number for the valence shell is 4, the element is present in the 4th period.

(ii) Since the last electron has been filled in 4s-subshell, the element belongs to s-block.

(iii) As there is two electrons in s-subshell, the element is present in group-2.

(2) **(B).** In a group, ionization energy decreases down the group: $Be > Mg > Ca$

(3) **True.**

(4)

Atomic no.	E.C.	Group
107	[Rn] 5f^{14}6d^5 7s^2	$5 + 2 = 7$
109	[Rn] 5f^{14}6d^7 7s^2	$7 + 2 = 9$

Valence shell is 7, so these elements belong to seventh period.

(5) The electronic configuration of the two elements are :

$Li (Z = 3)$ $Be (Z = 4)$

$1s^2 \, 2s^1$ $1s^2 \, 2s^2$

$\downarrow -e^-$ $\downarrow -e^-$

$Li^+ : 1s^2$ $Be^+ : 1s^2 \, 2s^1$

The configuration of two ions suggest that Li^+ ion has already acquired a noble gas configuration while Be^+ ion is yet to acheive the same. Therefore, ionisation energy for Li is expected to be more as compared to Be.

(6) (i)

	Lithium	Oxygen
Valency :	1	2
Formula :	Li_2O or Li_2O	

(ii)

	Silicon	Oxygen
Valency :	4	2
Formula :	Si_2O_4 or SiO_2	

(7) (i) The least reactive element is (IV) : It has very high ionization energy and $\Delta_{eg}H$ is positive which indicates that it is noble gas.

(ii) The most reactive metal is (II) : It has very low Δ_iH and also quite small $\Delta_{eg}H$. This values corresponds to the alkali metal.

(iii) The most reactive non-metal is (III) : Since the element has very high ionization enthalpy and very high negative electron gain enthalpy it is expected to be halogen.

(8) **(B).** Effective nuclear charge (i.e., Z/e ratio) decreases from F^- to N^{3-} hence the radii follows the order :

$F^- < O^{2-} < N^{3-}$.

Z/e for $F^- = 9/10 = 0.9$,

For $O^{2-} = 8/10 = 0.8$, For $N^{3-} = 0.7$

(9) Higher the magnitude of positive charge smaller will be size. Hence, $Co^{3+} < Co^{2+} < Co$

(10) **8.** $^{63}_{29}Cu + {}^1_1H \rightarrow 6 \, {}^1_0n + \alpha + 2 \, {}^1_1H + {}_{26}X^{52}$

$\Rightarrow$ X has atomic number $= 26 =$ atomic number of iron

$\Rightarrow$ group number $= 8$.

(11) $Cl(g) + e^- \rightarrow Cl^-(g)$; $\Delta H = -9.83 \, kJ$

1g chlorine requires energy 9.83 kJ

35.5g chlorine releases energy $= 9.83 \times 35.5 \, kJ = 349 \, kJ$

(12) All the three element belong to same peirod and three ions are isoelectronic and their size decreases with increase in the magnitude of positive charge $Li^+ > Be^{2+} > B^{3+}$.

CHAPTER-3 :
PERIODIC TABLE
EXERCISE-1

(1) (D)   **(2)** (A)   **(3)** (A)

(4) **(B).** Oxide of Eka-Aluminium $= E_2O_3$.
and oxide of Eka-Silicon $= EO_2$
because, oxidation state of $Al = + 3$.
oxidation state of $O = – 2$
oxidation state of $Si = + 4$
oxidation state of $O = –2$

(5) (D)   **(6)** (C)   **(7)** (A)

(8) (B)   **(9)** (C)   **(10)** (B)

(11) (C)

(12) **(B).** Atomic number $= 101$
Notation for IUPAC nomenclature
$0 = $ nil ; $1 = $ un
"ium" is added at the end.
IUPAC name $=$ Unnilunium.

(13) **(A).** $Z = 112$, name of the element is ununbium and its symbol is Uub.

(14) **(C).** Atomic number $= 105$
IUPAC official name $=$ Dubnium

(15) **(C).** Notation for IUPAC nomenclature
$1 = $ un, $6 = $ hex.
IUPAC name $=$ Ununhexium

(16) **(A).** $4d^{10}5s^0$ is the exceptional configuration of Pd. Its electronic configuration should be $[_{36}Kr]4d^8, 5s^2$.
Thus its Period $= 5$th
Group $= ns + (n – 1)d$ electrons$=2+8 = 10$

(17) **(B).** At. No. $= 16(S)$. Next element below this element has atomic number$= 16 + 18 = 34$

(18) **(A).** The configuration represents on alkaline earth metals.

(19) **(B).** On equating no. of e^-'s atomic no. is 12 which is for Mg.

(20) **(A).** Beryllium has four electrons and has the electronic configuration $1s^2 2s^2$. Starting from the next element boron, the 2p orbitals are filled with electrons when the L-shell is completed at neon $(2s^2 2p^6)$. Thus, there are 8 elements in the second period.

(21) **(B).** $Z = 79$ $[Xe] 4f^{14} 5d^{10} 6s^1 (n = 6)$
As $n = 6$, the element belongs to 6th period.
Group $= ns + (n – 1)d = 1 + 10 = 11$

(22) **(C).** n is the principal quantum number or number of shells. It represents the number of period to which an element belongs to.

(23) **(A).** Before the 4p orbital is filled, filling up of 3d-orbitals becomes energetically favourable and we come across the so called 3d transition series of elements. This starts from scandium $(Z = 21)$ which has electronic configuration- $(Ar) 3d^1 4s^2$.

(24) **(A).** Element with atomic number
$30 – [Ar] 3d^{10} 4s^2$
Element with atomic number
$48 – [Kr] 4d^{10} 5s^2$
Element with atomic number

(25) $80 – [Xe] 4f^{14} 5d^{10} 6s^2$

(25) **(B).** $Z = 12, 1s^2 2s^2 2p^6 3s^2$ $(n = 3)$
$Z = 17, 1s^2 2s^2 2p^6 3s^2 3p^5$ $(n = 3)$

(26) **(A).** Last electron enters in d-subshell so it belongs to d-block elements.
For d-block elements group number
$= $ no of electrons in $(n–1)$ d subshell
$+ $ number of electrons in valence shell (i.e. ns)
$= 7 + 2 = 9$.

(27) **(D).** Many metals with catalytic properties because
(i) They provide surface area for reaction to occur
(ii) They decreases the ionisation energy.
(iii) They have vacant d-orbitals.

(28) **(B).** s-block elements are reactive metals, with low ionisation enthalpy.
$+1 \rightarrow$ Alkali metals.
$+2 \rightarrow$ Alkaline earth metals

(29) **(D).** The non-metallic character of group-16 and group-17 elements increases as we move from left to right across a period and metallic character increases as we go down the group.

(30) **(D).** $ns^2 np^2$ are members of carbon family. Chalcogens is the general name of 16th group elements.

(31) **(C).** In lanthanoids, electrons are progressively filled in 4f orbital.

(32) **(D).** With the exception of lithium and beryllium compounds of s-block elements are predominantly ionic.

(33) **(C).** d-block elements are elements of group 3 to 12 in the centre of the Periodic Table. So, general outer electronic configuration $= (n –1) d^{1-10} ns^{0-2}$.

(34) **(D).** $(n – 2)f^{1–14} (n – 1) d^{0–1} ns^2$ is the general electronic configuration of f-block or inner-transition elements in which 2 shells prior to the outermost shell are incomplete.

(35) **(C).** Last element (Radon) of p-block in 6th period is represented as $4f^{14} 5d^{10} 6s^2 6p^6$.

(36) **(B).** Metalloids are shown by the zig-zag boxes and the elements present on the right side of the boundary are non-metals. Metals are present on the left side of the periodic table.

(37) **(A).** Actinoid elements are radioactive in nature.

(38) **(B).** $Z = 3; 1s^2 2s^1$
$Z = 20; 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

(39) **(B).** The atomic radius decreases along the period. Also cations are always smaller than their parent atoms and anions are always larger than their parent atom.

(40) **(D).** Due to identical ionic radii and polarising power.

(41) **(A).** Atomic radius decreases on going from left to right in a period. Thus size of $O > F$. As $O^{2–}$ and F^- are isoelectronic, therefore, size of $O^{2-} > F^-$

(42) **(C).** Li belongs to I^{st} group. There is $1e^-$ in outermost shell. Thus low I.E.

(43) **(B).** The first I.P. is maximum for hydrogen due to its small size.

(44) **(A).** Electron affinity value of Cl is greater the F and then

decreases down the group.

(45) (B). Electropositive character decreases across the period as metallic character decreases.

(46) (A). Both electronegativity and electron affinity increases. This is because decrease in the size and increase in the nuclear charge. But electronegativity increases continuously.

(47) (B). Electronegativity decreases down the group as atomic radius increases.

(48) (B). Metallic character in periods increases from right to left, while non-metallic character increases from left to right in periods.

(49) (B) It contains maximum number of electrons in the outermost shell. An element with more than 3 electrons in outermost shell is non-metallic.

(50) (B). I.E. in a group decreases and reactivity increases.

(51) (B). The metallic character increases down the group, while decreases from left to right in a period because atomic size decreases in moving from left to right in period.

Metallic character $\propto$ atomic size $\rightarrow$ (For group)

Metallic character $\propto \dfrac{1}{\text{nuclear charge}} \rightarrow$ (For period)

(52) (D). Atomic radius decreases in a period from left to right and increases in a group from top to bottom.

(53) (C). It is difficult to remove an electron from a highly electronegative element.

(54) (A). Within a group, IE_1 decreases as the atomic size increases, therefore, the IE_1 of K, Rb and Cs follow the order : Cs < Rb < K. Within period, IE increases as the nuclear charge increases, therefore. Ca has higher IE_1 than K. Thus, the overall sequence of increasing IE_1 is Cs, Rb, K, Ca.

(55) (D). Down the group metallic character increases, while in period from left to right metallic character decreases, but non-metallic character increases.
So, Na > Mg > Be > Si > P
Alkali metals have more metallic character.

(56) (A). Ionisation enthalpies of transition metals are intermediate between those of s & p-block elements.

(57) (C) Ionisation energy increases as we move from left to right in a period.

(58) (A). As we move across the period, atomic size decreases hence, ionization enthalpy increases. But for Al ($3s^2$ $3p^1$), electron has to be removed from outer 3p orbital whereas in Mg ($3s^2$), electron has to be removed from stable fully filled 3s orbital. Thus, I.E. for Mg > Al. So, the correct order of first ionization enthalpies is : Na < Mg > Al < Si.

(59) (D). The last electron in transition elements enters into (n – 1)d-subshell.

(60) (B). The first ionisation enthalpy increases as we move across a period & decrease as we descend in a group.

(61) (C). Due to stable configuration they do not accept an electron.

(62) (B) The electron gain enthalpy becomes more negative in a period from left to right.

(63) (D). For cations and anions

Effective Nuclear Charge $\propto \dfrac{1}{\text{ionic radii}}$

Thus, the decreasing order of the ionic radii is
$O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$

(64) (B) Removal of electron from p-orbital is easier than from s-orbital because of less penetrating power of p-orbitals.

EXERCISE-2

(1) (C). Where E.N. stands for electronegativity E.A. stands for electron affinity and I.P. stands for Ionization potential.
It is observed that for an element, E.A. and E.N. and I.P. usually vary in the same direction. Hence when E.A. and E.N. increase the I.P. also increases E.N. has the mean valued of I.P. and E.A.

$E.N. = \dfrac{I.P. + E.A.}{2} \therefore 2E.N. = I.P. + E.A.$

or $2E.N. - I.P. - E.A. = 0$

(2) (A). $r_A = \dfrac{d_{A-A}}{2} = 1.4/2 = 0.7$ Å

(3) (A).
(i) Element with lowest electronegativity is Cs.
(ii) Element with highest electronegativity is F
(iii) Element with highest ionisation potential is He.
(iv) Element with lowest ionisation potential is Cs.

(4) (A). Tc^{43} is the first artificial element.

(5) (D). There is only one incomplete orbit in Zn^{+2} and its stable oxidation state is (+2) does not have incomplete d orbital. Therefore it is not a transition element.

(6) (B). Inert gases, these have ns^2np^6 configuration.

(7) (D). Hydrogen resembles alkali metals in some properties so it can be placed in the first group of periodic table.

(8) (D). As the nuclear charge per electron is maximum in P^{5+}. Therefore its size is smallest.

(9) (D). Alkali metals, lower the no. of valence e^-, lower is the value of ionization potential.

(10) (C). The energy required to remove an electron from outermost orbit of an isolated gaseous atom is called I.E. Now carbon has $4e^-$ in outermost shell. Thus it has 4 ionization energies.

(11) (D). Li^+ and Mg^{+2} ions have similar polarising power or ionic potential and therefore have similar properties. This type of relationship of the first element of a group with the second element of the next group is known as diagonal relationship.

(12) (B). The formation of ionic bond depends upon easy formation of cation and anion, therefore the ionisation energy value of the metal atom should be low, so that it can easily form cation. On the other hand, the electron affinity value of the non-metal atom should be high so that it can easily form anion.

(13) (B). Flourine although have highest electronegativity due

to its very small size, effective inter electronic repulsions are observed which brings down its electron affinity.

(14) (D). As electronegativity difference between element and oxygen decreases the acidic character of oxides increases. The electronegativity also increases with increasing oxidation states. In general, as non-metallic character increases across the period, the acidic character of their oxides increases.

(15) (A). Ionic radii $\propto$ screening effect $\propto \dfrac{1}{ENC}$

(Effective Nuclear Charge)

(16) (C). CO, NO and N_2O are neutral oxides.

(17) (A) Ionisation energy increases along a period. Alkali metals can most easily lose ns^1 electron to attain noble gas configuration.

(18) (B). Element of 4f- and 5f-series are kept separately as lanthanoids and actinoids.

(19) (D). Atomic number = 102
Notation (0) $\rightarrow$ nil, (1) $\rightarrow$ un, (2) $\rightarrow$ bi
suffix $\rightarrow$ "ium" ; IUPAC name = Unnilbium

(20) (A). For isoelectronic ions more the nuclear charge, lesser is the ionic radius. Out of $Mg^{2+}(Z = 12)$, $Na^+(Z = 11)$ & $F^-(Z = 9)$. F^- has largest size among Mg^{2+}, Na^+ and F^-.

(21) (B). Barium belongs to s-block with an outer configuration $6s^2$.

(22) (D). Electronegativity increases across a period.

(23) (C). OF_2, here EN of fluorine is more than oxygen. So, oxidation state of fluorine = –1
and oxidation state of oxygen = + 2

(24) (A). Ionisation enthalpy of nitrogen is more than oxygen due to extra stability of half filled orbitals.

(25) (A). The stronger the charge on an ion the stronger the attractive forces that will result in an ionic lattice. Larger the ionic radius the lower the lattice enthalpy as the amount of interaction between the ions is smaller and the packing of ions is less efficient.

MgO	–3791 kJ mol^{-1}
CaO	–3401 kJ mol^{-1}
SrO	–3223 kJ mol^{-1}
BaO	–3054 kJ mol^{-1}

(26) (D). Given $X = 1s^2\ 2s^2\ 2p^6\ 3s^1$

$Y = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

Less energy is required to remove an electron from X than Y, because the distance between the nucleus and valence shell in Y is greater as compared to X.

(27) (D). Noble gas has stable electronic configuration and group 1 metals are biggest in the period while the size of Al and Ga is almost same.

(28) (B). Ionisation energy increases from left to right in a period but ionisation energy$_{group\ 2}$ > ionisation energy$_{group-13}$ and ionisation energy$_{group\ 15}$ > ionisation energy$_{group-16}$
This is because of stable electronic configuration of

group 2 is (ns^2) and group 15 is (ns^2np^3). Thus, Ne > Cl > P > S > Mg > Al : Ionisation potential.

(29) (B). Nitrogen has half-filled p-orbital (less ENC) ($3e^-$) Fluorine (F) has more interelectronic repulsion which increases effective nuclear charge (ENC).
As ENC $\propto$ 1/ionic radii, therefore, the correct order is $N^{3-} > O^{2-} > F^-$

(30) (A) Beryllium halides are covalent in nature. This is due to small size and high charge of Be^{2+} ion i.e., it has high polarising power. However, the halides of the other alkaline earth metals (fluorides, chlorides, bromides and iodides) are ionic solids.

(31) (A). Electronegativity on the Mulliken's scale

$= \dfrac{1}{2}\ (13.0 + 3.8) = 8.4.$

But since Mulliken's values are nearly 2.8 times as large as Pauling values, therefore, electronegativity of the element on the more commonly used Pauling

scale is $\dfrac{8.4}{2.8} = 3.0$

(32) (C). Second I.P. of an element will be the energy required to remove one mole of electron from one mole of monovalent gaseous cations of the element.

(33) (C) There are four transition series, three contain 10 elements each and fourth is incomplete.

(34) (A). For f-orbital, $l = 3$
$m_l = 2l + 1 = 7$
There are 7 electrons with +1/2 spin and 7 electrons with –1/2 spin.
Hence, total electrons are 14.

EXERCISE-3

(1) 14   **(2)** 72
(3) 4   **(4)** 3
(5) 3   **(6)** 3
(7) 17
(8) 6. Outer configuration $3d^5$, $4s^1$ belongs to group number 6.
(9) 19. The atomic number of the third alkali metal is 19.
(10) 7. The period number of the element (atomic number 98) is 7.

EXERCISE-4

(1) (A). The correct order of ionic radius
Ce > Sm > Tb > Lu
———————————→
$\downarrow$ order of ionic radius

(2) (B). Generally ionic radius $\downarrow$ from La^{+3} to Yb^{+3}.
So, the $\uparrow$ order is $Yb^{3+} < Pm^{3+} < Ce^{3+} < La^{3+}$

(3) (B). Atomic numbers

(4) (B). The reduction in atomic size with increase in atomic number is a characteristic of elements of f block
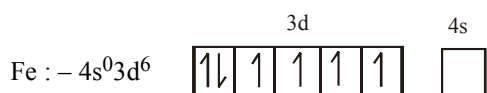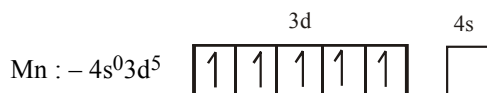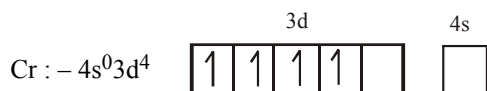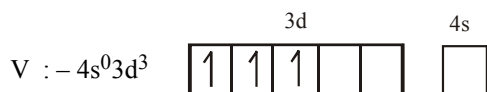
(5) (A). N^{3-}, F^-, Na^+

(6) (A). $_{23}V = [Ar]\ 3d^3\ 4s^2$
$_{24}Cr = [Ar]\ 3d^5\ 4s^1 \rightarrow$ have highest IInd ionizatio

enthalpy

(7) **(C).** Isoelectronic species $\rightarrow K^+,\ Cl^-,\ Ca^{+2},\ Sc^{+3}$
$$18e^-,\ 18e^-,\ 18e^-,\ 18e^-$$

(8) **(A).** $(-)$ ve charge $\uparrow$
Ionic radius $\uparrow$
So, maximum ionic radius O^{-2}.

(9) **(D).** Acidic strength $\propto$ I.P.
$$Al_2O_3 < SiO_2 < P_2O_3 < SO_2$$
$$\xrightarrow{\hspace{4cm}}$$
I.P. $\uparrow$, acidic strength $\uparrow$

(10) **(A).** $-$ve ion always tend to resist the addition of another e^-.

(11) **(B).** $B < C < N < O$ – increasing first ionization enthalpy

(12) **(A).** $\underset{\text{Base}}{SnO_2} + \underset{\text{Acid}}{4HCl} \longrightarrow SnCl_4 + 2H_2O$

$\underset{\text{Acid}}{SnO_2} + \underset{\text{Base}}{NaOH} \longrightarrow \underset{\substack{\text{Sodium}\\\text{stannate}}}{NO_2SnO_3} + H_2O$

(13) **(D).** II, III and IV

(14) **(C).** Zr and Hf have about the same radius

(15) **(C).** Poorer shielding of 5d electrons by 4f electrons

(16) **(D).** I.P. of F is maximum from P & S, P has high I.P. than S (p^3-stable configuration) therefore, the sequence is :
$$B < S < P < F$$

(17) **(C).** Isoelectronic species : $K^+, Cl^-, Ca^{2+}, Sc^{3+}$

(18) **(B).** Lanthanoid contraction is caused due to the imperfect shielding on outer electrons by 4f electrons from the nuclear charge.

(19) **(B).** Higher the charge/size ratio, more is the polarizing power. $\therefore\ K^+ < Ca^{+2} < Mg^{+2} < Be^{+2}$

(20) **(C).** On moving down the group, the ionic radius increases. In case of isoelectronic ions, more the charge on the ion, smaller will be its size.

(21) **(D).** For isoelectronic species higher the Z/e ratio, smaller the ionic radius.

(22) **(D).** Gadolinium $(_{64}Gd) = [Xe]^{54}\ 4f^7 5d^1 6s^2$

(23) **(C).** Order of increasing: $\Delta H_{IE_1} :\ Ba < Ca < Se < S < Ar$
$Ba < Ca$; $Se < S$: On moving top to bottom in a group, size increases. So ionisation energy decreases.
Ar : Max. value of ionisation energy, since it is an inert gas.

(24) **(B).** Radius of N^{3-}, O^{2-} and F^- order : $N^{3-} > O^{2-} > F^-$
As per inequality only option (B) is correct
that is 1.71 Å, 1.40 Å and 1.36 Å

(25) **(C).** $IE_{Na} = 496$ kJ /mol ; $IE_{Sc} = 633$ kJ /mol
It is relatively difficult to remove on e^- from 4s orbital of Sc as compared to 3S of Na due to poor shielding of d-orbital.

(26) **(B).**
ions	O^{2-}	F^-	Na^+	Mg^{2+}
Atomic number =	8	9	11	12
No. of e^- =	10	10	10	10

Isoelectronic species have same no. of electrons. All these species contains 10 electrons.

(27) **(B).** Electronegativity decreases as we go down the group

and atomic radius increases as we go down the group.

(28) **(D).**
Symbol	Atomic number
unh	106
uun	110
une	109
uue	119

(29) **(A).** As we move down the group, bond strength decreases, thereby decreasing the catenation tendency. Hence the order is as expected: $C > Si > Ge \approx Sn$

(30) **(C).** Atomic number $(Z) = 15 \Rightarrow P \rightarrow [Ne]3s^2\ 3p^3$
Phosphorus belongs to 15th group
Number of valence electrons = 5
and valency = 3 in ground state.

(31) **(C).** Atomic radius of Ag and Au is nearly same due to lanthanide contraction.

(32) **(A).**
(i) Electron affinity of second period p-block element is less than third period p-block element due to small size of second period p-block element.
E.A. order : $F < Cl$
(ii) Down the group electron affinity decreases due to size increases.
EA. order : $S > Se$
$$Li > Na$$

(33) **(A).** Correct order of ionisation energy will be :
$$Na < Al < Mg < Si$$

(34) **(C).** $_{26}Fe = [Ar]\ 3d^6\ 4s^2$

(35) **(A).**

EXERCISE-5

(1) **(D).** The amount of energy released when an electron is added to an isolated gaseous atom to produce a monovalent anion is called electron gain enthalpy. Electron affinity value generally increase on moving from left to right in a period however there are exceptions of this rule in the case of those atoms which have stable configuration. These atoms resist the addition of extra electron, therefore the low value of electron affinity.
$$O\ <\ S\ <\ F\ <\ Cl$$
$$-1.48\ \ -2.0\ \ -3.6\ \ -3.8$$
On the other hand Cl because of its comparitivey bigger size than F, allow the addition of an extra electron more easily.

(2) **(C).** For third ionization enthalpy last configuration of

V : – $4s^0 3d^3$

3d   4s

| ↑ | ↑ | ↑ | | | | |

Cr : – $4s^0 3d^4$

3d   4s

| ↑ | ↑ | ↑ | ↑ | | | |

Mn : – $4s^0 3d^5$

3d   4s

| ↑ | ↑ | ↑ | ↑ | ↑ | | |

Fe : – $4s^0 3d^6$

3d   4s

| ↑↓ | ↑ | ↑ | ↑ | ↑ | | |

For third ionization enthalpy Mn has stable configuration due to half filled d-orbital.

(3) **(C).** Bond dissociation energy of fluorine is less because of its small size and repulsion between electrons of two atoms. So option (C) is wrong order. The correct order is $Cl_2 > Br_2 > F_2 > I_2$

(4) **(A).** For isoelectronic species, size of anion increases as negative charge increases whereas size of cation decreases with increase in positive charge.

Further ionic radii of anions is more than that of cations. Thus the correct order is :

$Ca^{++} < K^+ < Ar < Cl^- < S^{--}$

(5) **(B).** The factors on which the ionisation energy depend are :

(i) Atomic size – I.E. decreases as the atomic size increases. So the attractive force decreases.

(ii) Number of electrons in the inner shell (screening effect). As the number of shells increases screening effect also increases and hence the I.E. tends to decrease.

(6) **(B).** The amount of energy required to remove an electron from a unipositive ion is called second IP which generally increases in a period from left to right.

$Cr > Mn > V > Ti$

(7) **(D).** The correct bond dissociation energy is

$Cl_2 > Br_2 > F_2 > I_2$

(8) **(B).** The configuration $3d^5 4s^2$ can have various oxidation states upto +7.

(9) **(D).** Smallest atom having half filled p-subshell has highest I_0 value.

(10) **(C).** Ionic radii $\propto$ charge on anion

$\propto$ 1/charge on cation.

During the formation of a cation, the electrons are lost from the outer shell and the remaining electrons experience a great force of attraction by the nucleus, ie, attracted more towards the nucleus.

In other words, nucleus hold the remaining electrons more tightly and this results in decreased radii. However, in case of anion formation, the addition of electron(s) takes place in the same outer shell, thus the hold of nucleus on the electrons of outer shell decreases and this results in increased ionic radii. Thus, the correct order of ionic radii is :

$S^{2-} > Cl^- > K^+ > Ca^{2+}$
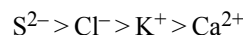
(11) **(B).** Electron gain enthalpy, generally, increases in a period from left to right and decreases in a group on moving downwards.

However, members of III period have somewhat higher electron gain enthalpy as compared to the corresponding members of second period, because of their small size.

O and S belong to VI A (16) group and Cl and F belong to VII A (17) group. Thus, the electron gain enthalpy of Cl and F is higher as compared to O and S.   Cl and F > O and S

Between Cl and F, Cl has higher electron gain enthalpy as in F, the incoming electron experiences a greater force of repulsion because of small size of F atom.

Similar is true in case of O and S i.e, the electron gain enthalpy of S is higher as compared to O due to its small size. Thus, the correct order of electron gain enthalpy of given elements is O < S < F< Cl

(12) **(B).** The atomic radii decreases on moving from left to right in a period, thus order of sizes for Cl, P and Mg is   Cl < P < Mg.

Down the group size increases. Thus overall order is : Cl < P < Mg < Ca.

(13) **(B).** The order of size of given cations is $Li^+ < Na^+ < Cs^+$

and the order of size of given anions is $I^- > F^-$

Thus, when the cation is large and anion is smallest, the cation to anion size ratio is maximum. Hence, cation to anion size ratio is maximum for CsF.

(14) **(A).** IE_1 of Na = – Electron gain enthalpy of Na^+ = – 5.1 Volt.

(15) **(A).** As the positive charge increases on metal cation, radius decreases.

(16) **(D).** Radius of a cation is always smaller than that of a neutral atom due to decrease in the number of orbits.

Ionic radii $\propto \dfrac{1}{atomic\ number}$

Ion	Al^{3+}	Mg^{2+}	N^{3-}
Atomic number	13	12	7

(17) **(B).** In case of isoelectronic species the radius decrease with increase in nuclear charge

$Ca^{2+} < K^+ < Ar$

electrons	18	18	18
Proton	20	19	18

$$\frac{z}{e} \qquad \frac{20}{18} = 1.11 \qquad \frac{19}{18} = 1.05 \qquad \frac{18}{18} = 1$$
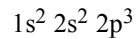
Ionic radius $\propto \dfrac{1}{(Z/e)}$

(18) (BC).

For option (B) : The correct order for 1st ionisation energy is $\quad$ B < C < O < N.

For option (C) : The correct order for magnitude of electron gain enthalpy is I < Br < F < Cl

(19) (A). $Z = 114$ $[Rn]^{86}$ $7s^2$ $5f^{14}$ $6d^{10}$ $7p^2$

14th group (carbon family)

(20) (D). Element (X) electronic configuration
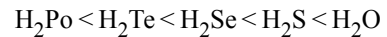
$1s^2$ $2s^2$ $2p^3$

So, valency of X will be 3.

Valency of Mg is 2.

Formula of compound formed by Mg and X will be Mg_3X_2.

(21) (C). On going down the group thermal stability

order for H_2E decreases because H–E bond energy decreases.

$\therefore$ $\quad$ Order of stability would be

$H_2Po < H_2Te < H_2Se < H_2S < H_2O$

(22) (B). 'Be' and 'N' have comparatively more stable valence sub-shell than 'B' and 'O'.

$\therefore$ $\quad$ Correct order of first ionisation enthalpy is:

$Li < B < Be < C < O < N < F < Ne$