03

Classification of Elements & Periodicity in Properties

More than 100 elements are known today. It is not possible to study the properties of each and every element separately. So, for the ease in study periodic table was developed.

In periodic table the elements are arranged in increasing order of their atomic numbers. Using the position of an element in the periodic table, it is possible to predict its physical properties and chemical behaviour.

These behaviours and properties show gradual charges on moving left to right or top to bottom in periodic table. This regular gradation in the properties is called **periodicity** and such properties are called **periodic properties**.

IN THIS CHAPTER ....

Earlier Attempts to Classify Elements

Several attempts were made to classify the elements into groups purposefully for the sake of better study of their properties. The initial developments in this regard are given below

Prout's Hypothesis

This hypothesis was given in 1815. It is also called **unitary theory** or **unitary hypothesis** as according to it hydrogen atom was considered as the fundamental unit from which all other atoms were made. This hypothesis lost its influence after the discovery of isotopes.

Dobereiner's Triads

Dobereiner, between 1815-1829, gave this law of triads. A triad is a group of three elements with more or less similar properties.

According to him, in the triads the atomic weight of the middle element was approximately the arithmetic mean of the two others. e.g.

Element	**Li**	**Na**	**K**
Atomic weight	7	23	39

$$\text{Mean of atomic masses} = \frac{7 + 39}{2} = 23$$

Similarly,

Element	**Cl**	**Br**	**I**
Atomic weight	35.5	80	127

$$\text{Mean of atomic masses} = \frac{35.5 + 127}{2} = 81.25$$

But in some triads, all the three elements possessed nearly equal atomic masses, e.g. Fe, Co, Ni, Os, Ir, Pt, etc., hence, this law was rejected.

Newland's Law of Octaves

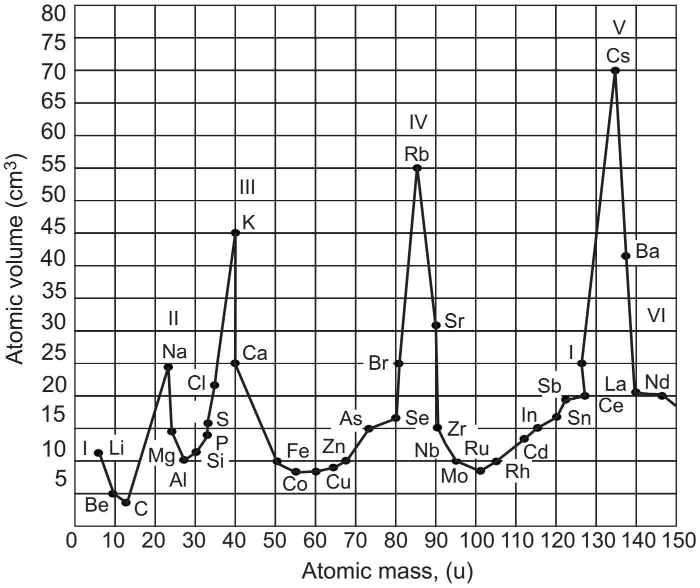
Newland (in 1864) presented the **law of octaves** which was considered as first serious attempt to classify the elements. According to this law, if the known elements were arranged in increasing order of their atomic weights, then every 8th element starting from a given one, was a kind of repetition of the first, like the 8th note in the octave of music, i.e.

Sa	**Re**	**Ga**	**Ma**	**Pa**	**Dha**	**Ne**
H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S

The above law worked quite well for lighter elements but failed in the case of heavier elements. Moreover, after the discovery of noble gases, it is the ninth element that resembles with the first, which is not in accordance to this law.

Lothar Meyer's Atomic Volume Curve

In 1869, Lothar Meyer prepared a curve between atomic volumes and atomic masses of different elements and concluded that, "the properties of the elements were the periodic functions of their atomic volumes". His curve outline is given below



Lothar Meyer's curve

He calculated atomic volume with the help of the formula,

$$\text{Atomic volume} = \frac{\text{Molecular mass}}{\text{Density}}$$

Important findings regarding the curve were

- Alkali metals having the largest atomic volumes occupy the maxima of the curve.
- The alkaline earth metals (Mg, Ca, Sr, Ba) occupy the mid point positions on the descending portions of the curve.
- Halogens occupy positions on ascending portions before inert gases.
- The transition elements occupy minima of the curve.

Remember that **Lothar Meyer** (Germany) and **Mendeleef** (Russia) quite independently evolved identical classification basics for elements in 1869.

Both of them showed the connection between the periodicity of properties and atomic masses of elements, with the difference that Mendeleef considered the chemical properties while Lothar Meyer considered the physical properties of elements in their explanations.

Mendeleef's Periodic Law

It states "*that the physical and chemical properties of elements are the periodic function of their atomic masses*". In other words, if the elements are arranged in the order of their increasing atomic masses then, after a regular interval elements of similar properties are repeated.

His law infact was an impressive version of the law of octaves and formed the basis for the modern classification of elements in the form of a table, called **periodic table.**

Mendeleef's Periodic Table

This table was based on the periodic law of Mendeleef.

In his original table, Mendeleef arranged the elements in horizontal rows but subsequently he modified the original table and rearranged the elements in vertical columns with horizontal rows.

Mendeleef's original table consisted of 8 vertical columns, called **groups** and 12 horizontal rows, called **series** or **periods**.

When Mendeleef presented the periodic table only 63 elements were known and a number of gaps were left for unknown elements (like Eka-aluminium, Eka-silicon,Gallium, Germanium). The properties of those unknown elements were also predicted on the basis of periodic law, which helped future scientists in the discovery of such elements.

Zero group was not known at Mendeleef's time. It was added later on in the modified Mendeleef's periodic table.

Uses of Mendeleef's Periodic Table

- Atomic weights of various elements were determined with the help of periodic table.
  Atomic weight = Valency × Equivalent weight.
- Atomic weight of different elements were corrected e.g. Atomic weight of Be was calculated to be $3 \times 4.5 = 13.5$ by considering its valency 3 but, Mendeleef calculated it $2 \times 4.5 = 9$
- In Mendeleef's periodic table two consecutive members differ by two or three units in the atomic weight. Where this gap was more, the gaps were left in the periodic table. This helped in discovery of new elements.

Defects of Mendeleef's Periodic Table

- Position of hydrogen was not clear. It has been placed in IA and VIIA groups because of its resemblance with elements of both the groups.
- No separate positions were given for isotopes.
- No separate group was formed for lanthanides and actinides.
- Order of increasing atomic weights was not strictly followed in the arrangement of elements with in the periodic table. e.g. Co (At. wt. 58.9) was placed before I (127) and Ar (39.9) before K (39).

Moseley's Work with X-rays

Moseley, while working in Rutherford's team in 1913, studied the wavelength of the characteristic X-rays by using different metals as anti-cathode. He showed that the square root of the frequency of a line of X-ray spectrum is related to the atomic number, Z of the target material as

$$\sqrt{\nu} = a(Z - b)$$

where, a = proportionality constant

b = another constant it has same value for all the lines of X-rays spectrum

Z = atomic number.

Modern Periodic Law

Moseley thereby showed that atomic number is more fundamental property of an element than its atomic mass and on this basis he modified the periodic law of Mendeleef as, "the physical and chemical properties of the elements are periodic functions of their atomic numbers".

This modified law of Moseley is called "modern periodic law" and considered as the basis of modern periodic table.

Modern Periodic Table

The characteristics of this table are

- It consists of 18 vertical columns, called the **groups** and 7 horizontal rows, called the **periods**.
- The groups are marked 0 to VIII out of which group I to VII are subdivided into subgroups A and B.
- Starting from hydrogen to heaviest element uranium, there are 92 elements which are arranged in different periods as follows

1st period	2 elements	Shortest period
2nd period	8 elements	Ist short period
3rd period	8 elements	IInd short period
4th period	18 elements	Ist long period
5th period	18 elements	IInd long period
6th period	32 elements	Longest period
7th period	6 elements	Incomplete period

- The group IA elements (Li, Na, K, Rb, Cs and Fr) are called **alkali metals** and the group II A elements (Be, Mg, Ca, Sr,Ba and Ra) are called **alkaline earth metals**. Elements in group VII A (F, Cl, Br, I and At) are known as **halogens** and elements in group VIII A (He, Ne, Ar, Kr, Xe and Rn) are called **noble gases** or **rare gases**.
- The group VIII has three similar elements placed together in one place. These are called **transition triads**, e.g. Fe, Co and Ni; Ru, Rh and Pd; Os, Ir and Pt etc.
- In the 6th period, 14 elements from lanthanum (Z = 57) to lutetium (Z = 71) are given one position in III B group. These elements are termed as **rare earths** or **lanthanides**.
- Afterwords, 14 elements from actinium (Z = 89) to lawrencium (Z = 103) are also added to same place and called them actinides.
- In this way, every period starts from a member of alkali group and ends with a member of zero group. The first period however starts with hydrogen.

Long Form of Periodic Table

The Modern Periodic Table was further modified to extended form or **long form** of periodic table. This modification was not done by a single person, but it incorporates the modifications suggested by many scientists especially **Rang**, **Werner**, **Bohr**, **Bury** etc.

This table is just a graphical representation of Aufbau principle. The electronic configuration based classification of elements was given by Bohr that's why the table is also called **Bohr's periodic table.**

Features

- In this table, elements are arranged in increasing order of their atomic numbers (Z), which is similar to modern periodic table. The isotopes are grouped together as they have the same atomic number.

- The table contains 7 periods (representing 7 orbits) and 18 groups (1-18). The concept of subgroup A and B is removed and groups are given no. 1 to 18.
- The transitional triad elements of modern periodic table (group 8) are included in group 8, 9 and 10. The name of zero group is changed to group 18.
- A new period begins with the occupancy of new energy shell.

  Thus, all periods begin with the elements having ns^1configuration (n = 1, 2, 3, … for 1st, 2nd and 3rd period respectively).
- A period usually terminates at electronic configuration $ns^2\ np^6$ except 1st period which terminates on $1s^2$ as K orbit lacks p-orbital.
- The 3rd group is the largest group of periodic table containing 32 elements.
- On the basis of electronic configuration, elements are classified into following four groups
  - (i) Inert gases
  - (ii) Representative elements
  - (iii) Transitional elements
  - (iv) Inner-transitional elements

Representative elements (*s*-block elements) GROUP NUMBER

Alkali metal

Alkaline earth metal

Transition elements (*d*-block elements) GROUP NUMBER

Representative elements (*p*-block elements) GROUP NUMBER

Noble gases

1
H
$1s^1$

Period Number	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB ←	8	9 VIII	10 →	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	Chalcogen 16 VIA	Halogen 17 VIIA	18 0
1																		2 He $1s^2$
2	3 Li $2s^1$	4 Be $2s^2$											5 B $2s^22p^1$	6 C $2s^22p^2$	7 N $2s^22p^3$	8 O $2s^22p^4$	9 F $2s^22p^5$	10 Ne $2s^22p^6$
3	11 Na $3s^1$	12 Mg $3s^2$											13 Al $3s^23p^1$	14 Si $3s^23p^2$	15 P $3s^23p^3$	16 S $3s^23p^4$	17 Cl $3s^23p^5$	18 Ar $3s^23p^6$
4	19 K $4s^1$	20 Ca $4s^2$	21 Sc $3d^14s^2$	22 Ti $3d^24s^2$	23 V $3d^34s$	24 Cr $3d^54s^1$	25 Mn $3d^54s^2$	26 Fe $3d^64s^2$	27 Co $3d^74s^2$	28 Ni $3d^84s^2$	29 Cu $3d^{10}4s^1$	30 Zn $3d^{10}4s^2$	31 Ga $4s^24p^1$	32 Ge $4s^24p^2$	33 As $4s^24p^3$	34 Se $4s^24p^4$	35 Br $4s^24p^5$	36 Kr $4s^24p^6$
5	37 Rb $5s^1$	38 Sr $5s^2$	39 Y $4d^15s^2$	40 Zr $4d^25s$	41 Nb $4d^45s^1$	42 Mo $4d^55s^1$	43 Tc $4d^55s^2$	44 Ru $4d^75s^1$	45 Rh $4d^85s^1$	46 Pd $4d^{10}$	47 Ag $4d^{10}5s^1$	48 Cd $4d^{10}5s^2$	49 In $5s^25p^1$	50 Sn $5s^25p^2$	51 Sb $5s^25p^3$	52 Te $5s^25p^4$	53 I $5s^25p^5$	54 Xe $5s^25p^6$
6	55 Cs $6s^1$	56 Ba $6s^2$	57 La* $5d^16s^2$	72 Hf $4f^{14}5d^26s^2$	73 Ta $5d^36s^2$	74 W $5d^46s^2$	75 Re $5d^56s^2$	76 Os $5d^66s^2$	77 Ir $5d^76s^2$	78 Pt $5d^96s^1$	79 Au $5d^{10}6s^1$	80 Hg $5d^{10}6s^2$	81 Tl $6s^26p^1$	82 Pb $6s^26p^2$	83 Bi $6s^26p^3$	84 Po $6s^26p^4$	85 At $6s^26p^5$	86 Rn $6s^26p^6$
7	87 Fr $7s^1$	88 Ra $7s^2$	89 Ac** $6d^17s^2$	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cp	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

Inner-transition elements
(*f*-block elements)

*Lanthanides $4f^n5d^{0-1}6s^2$	58 Ce $4f^25d^06s^2$	59 Pr $4f^35d^06s^2$	60 Nd $4f^45d^06s^2$	61 Pm $4f^55d^06s^2$	62 Sm $4f^65d^06s^2$	63 Eu $4f^75d^06s^2$	64 Gd $4f^75d^16s^2$	65 Tb $4f^95d^06s^2$	66 Dy $4f^{10}5d^06s^2$	67 Ho $4f^{11}5d^06s^2$	68 Er $4f^{12}5d^06s^2$	69 Tm $4f^{13}5d^06s^2$	70 Yb $4f^{14}5d^06s^2$	71 Lu $4f^{14}5d^16s^2$
**Actinides $5f^n6d^{0-2}7s^2$	90 Th $5f^16d^17s^2$	91 Pa $5f^26d^17s^2$	92 U $5f^36d^17s^2$	93 Np $5f^46d^17s^2$	94 Pu $5f^66d^07s^2$	95 Am $5f^76d^07s^2$	96 Cm $5f^76d^17s^2$	97 Bk $5f^96d^07s^2$	98 Cf $5f^{10}6d^07s^2$	99 Es $5f^{11}6d^07s^2$	100 Fm $5f^{12}6d^07s^2$	101 Md $5f^{13}6d^07s$	102 No $5f^{14}6d^07s^2$	103 Lr $5f^{14}6d^17s^2$

Two super heavy elements, element 113 and element 115 have been recently discovered by scientist. Element 113 is produced *via* the alpha decay of element 115. Experiment of this discovery was made between July 14 and August 10, 2003. The results have been published in Physical Review C.

Classification of Elements

Depending upon the valence orbital, the elements are classified as

1. *s*-Block Elements

The *s*-block elements are present on the extreme left in the periodic table. All of these are highly reactive metals except hydrogen which is highly reactive but not considered as a metal.

Electronic Configuration

The generalised electronic configuration is

$$[\text{inert gas}]\ ns^x \quad (x = 1, 2)$$

- ***Case* I** If $x = 1$, elements constitute a group of **alkali metals** placed in group 1 (IA) and
- ***Case* II** If $x = 2$, elements constitute a group of **alkaline earth metals** placed in group 2 (IIA).

The important characteristics of elements of this block are as follows

- They are soft metals with low melting and boiling points. Their reactivity increases down the group.
- They are highly electropositive and having low ionisation enthalpies.
- Most of the metals of this block impart characteristic colour to the flame.
- They are strong reducing agents and good conductor of heat and electricity.
- Na and Mg have extensive three-dimensional atomic networks, which are held together by forces characteristic of the metallic state.

2. *p*-Block Elements

These elements are present in the right portion of periodic table. This is the only block that contains metals, semi-metals and non-metals. A clear demarcation line separates the metals and non-metals.

This line begins from B, runs in a *zig-zag* manner and terminates at Po. Semi-metals are present along the sides of this line.

Electronic Configuration

The general electronic configuration of *p*-block elements is [inert gas] ns^2np^x (where, $x = 1$ to 6).

Group number of an element of this block is $(2 + x)$ based on old system like IIIA, IVA, … etc., and $(12 + x)$ based on new IUPAC system like, 13, 14, 15, …, etc.

Important characteristics of the elements of this block are

- Metals of this block are solids generally.
- Non-metals may be solid, liquid or gases.
- The metalloids of this block show variable characteristics like giant 3-D structures, allotropism etc.
- The elements exist in simple discrete bi or multimolecular units like P_4, S_8, Cl_2 etc.
- Noble gases are monoatomic molecules.
- These elements mostly form covalent compounds.

3. *d*-Block Elements

The elements having incompletely filled *d*-orbitals in the ground state or in excited or chemically combined state are called **transitional elements** (*d*-families) or *d*-block elements. All the elements of *d*-block are present in the middle portion of periodic table, i.e. in between the *s* and *p*-block elements. They are all metals.

Electronic Configuration

The general electronic configuration of the shells of transitional elements is $(n-1)\,s^2p^6d^x\,ns^{1\text{ or }2}$

Here, $x = 1$ to 10 for *d*-1 to *d*-10 families.

or $$(n-1)d^x ns^y$$

where, x = number of electrons in $(n-1)d$-orbital

= 1 to 10 (except 4 and 9)

y = number of electrons in ns-orbital.

= 2 in all cases [except in group 6(VIB) and 2(IB), $y = 1$]

According to IUPAC (new pattern) group $= (x + y)$, but old group number is decided by

- group number is $[(x + y)\text{B}]$, if $x < 6$
- group number is [VIII], if $x = 6, 7, 8$
- group number is $[y\text{B}]$, if $x = 10$

There are three complete series of 10 transition elements, called $3d\,({}_{21}\text{Sc}$ to ${}_{30}\text{Zn})$, $4d\,({}_{39}\text{Y}$ to ${}_{48}\text{Cd})$ and $5d\,({}_{57}\text{La};\ {}_{72}\text{Hf}$ to ${}_{80}\text{Hg})$ whereas, $6d$ series is incomplete. ${}_{21}\text{Sc}$ is the lightest transition metal.

Exception

Following two exception of general trend are visible in this block.

(i) Although by definition zinc (group 12) does not belong to the *d*-family but for chemical reasons it is classified as *d*-block elements with d^{10} configuration.

(ii) By the strict definition, lutetium (Z=71) and lawrencium ($Z = 103$) should be included in the *d*-families, but their properties justify their inclusion in the inner-transitional elements.

Further, as in these elements $4f$ and $5f$ levels are just filled up, these can be termed as the bridging elements between transition (*d*-block) and inner-transition (*f*-block) elements.

The transitional elements show striking resemblances to one another in their characteristics such as melting and boiling points, metallic nature, etc., as they have almost same outermost shell configuration.

Important characteristics of elements of this block are as follows

- These are hard, ductile and malleable metals with high melting and boiling points.
- Their ionisation energy is in between *s*- and *p*-block elements.
- They show variable oxidation states and are good conductors of heat and electricity.
- They form both ionic and covalent compounds.
- These compounds are generally coloured and paramagnetic in nature.
- Most of the transition metals form alloys.

4. *f*-Block Elements

The inner-transition elements or *f*-block elements have their $(n-2)f$ orbitals incompletely filled. There are 2 series of inner-transition elements, namely, $4f$ and $5f$ series. The $4f$ series begins with cerium ($Z = 58$) and ends on lutetium ($Z = 71$) and the $5f$ series contains elements from thorium ($Z = 90$) to lawrencium ($Z = 103$).

The *f*-block elements are positioned at the bottom portion of periodic table and are included in the 3rd group (III B of modern periodic table). This group is considered as the largest group of periodic table.

The members of $4f$ series are called **lanthanons** or **rare earth** or **lanthanides** as these follow *lanthanum* ($Z = 57$) in their properties.

The member of $5f$ series are called **actinons** or **actinides** as these follow actinium ($Z = 89$) in their properties.

Electronic Configuration

The generalised electronic configuration of $4f$ and $5f$ series is $(n-2)f^{1-14}\,(n-1)s^2p^6d^{0-1}ns^2$, i.e.

for $4f$ series is $4f^{1-14}, 5s^2p^6d^{0-1}, 6s^2$

for $5f$ series is $5f^{1-14}, 6s^2p^6d^{0-1}, 7s^2$

The elements of both the series show striking resemblance in their properties due to almost similar configurations of their outer two shells and same size.

Elements with atomic number greater than 92 (U_{92}) are called the **transuranium elements**. All these elements are man-made through artificial nuclear reactions. At present, transuranium elements with atomic numbers up to 118 have been reported.

Important characteristics of the elements of this block are as follows

- The tendency to loose *f*-electrons is more seen in the case of actinons as their electrons are less firmly held by the nuclear charge in comparison with those of $4f$ family.
- Lanthanides show reduction in their sizes with increase in their atomic numbers. This phenomenon is called as **lanthanide contraction**. This is an extra ordinary phenomenon as all the lanthanoids occupy, same position within the periodic table, i.e. 3rd group and 6th period.

  This can be explained on the basis of constant poor screening or shielding effect of *f*- electrons by other *f*-electrons.
- These are heavy metals with high melting and boiling points.
- They show variable oxidation states and compounds are generally coloured.
- Most of the elements of the actinide series are radioactive.

Example 1. *In the sixth period, the orbitals that filled are* ***(JEE Main 2020)***

(a) $6s, 5f, 6d, 6p$ (b) $6s, 4f, 5d, 6p$

(c) $6s, 5d, 5f, 6p$ (d) $6s, 6p, 6d, 6f$

Sol. (*b*) Filling of electrons in orbitals in any period takes place as

$$s \to ns\underset{\text{If possible}}{(n-2)\,f}\underset{\text{If possible}}{(n-l)d}\;np$$

$\therefore$ A per $(n+l)$ rule for sixth period $n = 6$, orbitals that are filled are $6\,s$, $4f$, $5d$ and $6p$.

Prediction of Group, Period and Block of an Element

We can easily predict the group, period and block of an element from its electronic configuration by the following information.

- Principal quantum number or the valence shell corresponds to the period of an element.
- The orbital containing the last electron indicates the block of an element.
- The group of an element is predicted from the number of electrons in the valence shell or a penultimate shell $(n-1)$ in following ways
  - (i) **For *s*-block elements**, group number is equal to the number of valence electrons.
  - (ii) **For *p*-block elements** group number is equal to (10 + number) of electrons in the valence shell.
  - (iii) **For *d*-block elements,** group number is equal to the number of electrons in $(n-1)$ *d*-subshell + number of electrons in valence shell (*n*th shell).

    Alternately, you can calculate group number = Number of electrons in (penultimate shell + valence shell) −8.
  - (iv) **For *f*-block elements,** group number is always III B.

Example 2. *The elements with atomic numbers* 101 *and* 104 *belong to, respectively,* ***(JEE Main 2020)***

(a) actinoids and group 4 (b) actinoids and group 6

(c) group 6 and actinoids (d) group 11 and group 4

Sol. (a) Actinoid series have atomic number from 90 to 103 (*f*-block element).

So, element with atomic number (Z) 101 belongs to actinoid series and elements with atomic number (Z) 104 belongs to 4th group.

Nomenclature of the Elements having Atomic Number >100

CNIC (Commission on Nomenclature of Inorganic Chemistry) appointed by IUPAC in 1994, approved official names for elements with atomic number 104 to 109.

For the elements having $Z > 109$, IUPAC recommended a nomenclature to be followed until their names are officially recognised. This nomenclature is based on the Latin words for the atomic numbers of the elements.

The names are derived by using roots for the three digits in the atomic number of the elements and adding-ium at the end -ium. The roots for the numbers are

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

e.g. The element with atomic number 109 will be named as un-nil-ennium and symbolised as une (u for 1, *n* for 0 and e for 9). Other names from atomic no. 101 to 118 are given below

IUPAC Nomenclature of Elements with Atomic Number Above 100

Atomic number	Name	Symbol	IUPAC official name	IUPAC symbol
101	Un-nil-unnium	Unn	Mendelevium	Md
102	Un-nil-bium	Unb	Nobelium	No
103	Un-nil-trium	Unt	Lawrencium	Lr
104	Un-nil-quadium	Unq	Rutherfordium	Rf
105	Un-nil-pentium	Unp	Dubnium	Db
106	Un-nil-hexium	Unh	Seaborgium	Sg
107	Un-nil-septium	Uns	Bohrium	Bh
108	Un-nil-octium	Uno	Hassnium	Hs
109	Un-nil-ennium	Une	Meitnerium	Mt
110	Un-un-nillium	Uun	Darmstadtium*	Ds
111	Un-un unnium	Uuu	Roentgenium*	Rg
112	Un-un-bium	Uub	**	
113	Un-un-trium	Uut	**	
114	Un-un-quadium	Uuq	**	
115	Un-un-pentium	Uup	**	
116	Un-un-hexium	Unh	**	
117	Un-un-septium	Uus	***	
118	Un-un-octium	Uno	***	

*IUPAC approved the name on August 16, 2003.

IUPAC name announced on 1 Nov, 2004.

**Reported but not fully authenticated.

***Elements yet to be reported.

Example 3. *The* IUPAC *symbol for the element with atomic number* 119 *would be* ***(JEE Main 2019)***

(a) unh (b) uue (c) uun (d) une

Sol. (*b*) Atomic number (119) = $\underset{\text{un}}{\underset{\vee}{1}}\ \underset{\text{un}}{\underset{\vee}{1}}\ \underset{\text{en}}{\underset{\vee}{9}}$

So, symbol of the element = uue

Name of the element = un-un-ennium.

It is expected to be *s*-block element an alkali metal and the first element in eighth period. It will be the lightest element that has not yet been synthesised.

Periodic Properties

Within any period or group of periodic table a regular gradation in properties is visible with increase in atomic number. This regular gradation in properties is called **periodicity**.

The main cause of periodicity in properties is the repetition of similar electronic configuration at regular intervals. Next couple of pages of this chapter are dedicated to these properties only.

However, before discussing the periodic properties, it is essential to know **screening** or **shielding effect**, as it greatly affects the periodic properties.

Screening or Shielding Effect

It is the effect in which, electrons present in the inner orbits of an atom are themselves attracted more towards the nucleus. They thus, decrease the nuclear attractive forces for the electrons of outer orbits.

Technically it is defined as an effect in which the electrons of inner orbits work as a curtain between the nucleus and the electrons of outer orbits. This is due to the reason, that the actual force of attraction on these valence electrons is decreased by the repulsive forces between them as both act in opposite direction.

Magnitude of Screening Effect

The magnitude of screening effect depends upon the number of inner electrons, i.e. higher the number of inner electrons, greater shall be the value of screening effect.

The Slater's Formula

The screening effect and effective nuclear charge, i.e. the charge felt by the electrons of a particular orbit are very closely related.

Slater gave the relation between them which is written as

$$Z' = Z - \sigma$$

where, Z' = effective nuclear charge

Z = atomic number

σ = screening constant or Slater's constant

The value of Slater's constant can be given as

$$\sigma = (0.35 \times \text{no. of } e^- \text{ in } n\text{th orbit} + 0.85 \times \text{no. of } e^- \text{ in } (n-1)\text{th orbit} + 1 \times \text{rest of the } e^-)$$

[Take 0.30 instead of 0.35 for 1st orbit, i.e. $(n = 1)$]

Trends of Screening Constant and Z_{eff}

- The magnitude of σ increases across a period or down the group for the elements of groups 1, 2, 13, 14, 15, 16, 17.
- The magnitude of effective nuclear charge, Z' increases in a period but it remains almost same in the case of elements of groups 1, 2, 13, 14, 15, 16, 17.
- Screening effect of all the inner electrons is not equal. It also depends upon the subshell in which electrons are present. For a given orbit, the screening effect of electron belonging to different subshells decreases in the order
$$s > p > d > f$$
- The screening effect of d and f-orbitals is taken as almost constant.

Remember

(i) For second period with n valence electrons,
$$Z' = 0.65\,(1 + n)$$
(ii) For third period with n valence electrons
$$Z' = 1.55 + 0.65\,n$$
(iii) For higher periods with d and f-orbitals variation may differ.

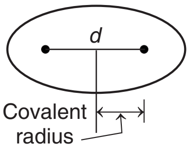
Atomic Radius

It is defined as the distance from the centre of the nucleus to the outermost shell of electrons. This definition is imaginary because according to quantum mechanics there is no certainty about the exact position of an electron in an atom. Thus, for practical purposes, atomic radius can be defined in following types

1. Covalent Radius

It is defined as *the half of the internuclear distance between two identical or almost identical atoms bonded by a single covalent bond*.

So, for a diatomic molecule, A_2



Distance $(A - A)$ = radius of A + radius of A

or $$d = 2r_A \quad \text{or} \quad r_A = \frac{d}{2}$$

Similarly, for a heteronuclear diatomic molecule, AB

Distance $(A\text{-}B)$ = radius of A + radius of B

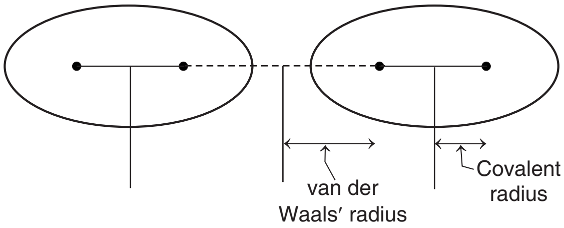
The above relations hold good only when the bond is purely covalent.

2. van der Waals' Radius

It is defined as half of the internuclear distance between the nuclei of two non-bonded adjacent atoms belonging to different molecules having identical atoms. This radius is taken into consideration for those cases where intermolecular attractive forces are less strong, e.g. weak dispersion forces.

The examples of such molecules are those of inert gases, saturated molecules etc.

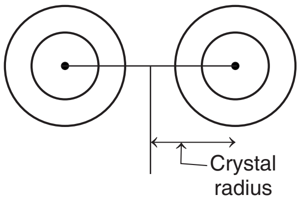
For a given element, the covalent radius is always smaller than van der Waals' radius.



Representation of van der Waals' and covalent radii

3. Crystal Radius

The atomic radius of metals in solid state is called their crystal radius and is defined as *half of the inter-nuclear distance between two non-bonded adjacent atoms in closed packed crystal lattice*.



Representation of crystal radius

A crystal radius is 10-15% greater than the covalent radius of a particular atom but is smaller than its van der Waals' radius.

Trends in Atomic Radius

The variations in atomic radii of various elements show the following trend

- **Atomic radii of elements increases on moving down the group**. It is due to increase in the number of shells by a factor of 1 and increase in screening effect, which further reduces the effective nuclear charge.
- **Atomic radius of elements decreases while moving from left to right in a period**. This is due to the fact that effective nuclear charge increases as number of inner electrons remain the same throughout the period and the change occur only in the number of valence electrons.
- **In case of transition elements, atomic radii remains almost same** because the electrons add to penultimate, i.e. $(n-1)\,d$ subshell. Consequently, the increased nuclear charge is almost compensated by the screening effect.

- **In case of inner transition elements** the remarkable phenomenon of lanthanoid contraction and actinoid contraction are visible.

Example 4. *The increasing order of the atomic radii of the following elements is* **(JEE Main 2020)**

(*A*) C (*B*) O (*C*) F

(*D*) Cl (E) Br

(a) (*A*) < (*B*) < (*C*) < (*D*) < (*E*)

(b) (*C*) < (*B*) < (*A*) < (*D*) < (*E*)

(c) (*D*) < (*C*) < (*B*) < (*A*) < (*E*)

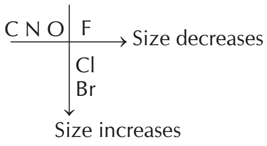
(d) (*B*) < (*C*) < (*D*) < (*A*) < (*E*)

Sol. (*b*) Atomic radius generally decreases as we compare elements in a period from left to right,

$$\therefore \qquad \text{C} > \text{O} > \text{F}$$

but elements present in next period are larger in size,

$$\therefore \qquad \text{Br} > \text{Cl} > \text{C} > \text{O} > \text{F};$$



So, the correct increasing order of the atomic radii

$$C < B < A < D < E$$

4. Ionic Radius

For ions, it is difficult to consider all the above written radii as cations and anions have significant differences in their sizes. The above written methods are used for identical or almost identical atoms.

For ions, thus we have to consider the imaginary definition written in the beginning, i.e. *the distance between the centre of nucleus and its outermost shell the distance between the nucleus and the point at which nucleus exerts its influence*, in an ion **is called ionic radius.** For the determination of cationic and anionic radii, X-ray diffraction techniques are used.

Radii of Cation

The **radius of cation is always smaller** than its neutral atom, because of increase in the effective nuclear charge on electrons left and decrease in the number of orbits for most of the cases.

If various cations have same number of electrons (isoelectronic), the cation having greater number of proton will be smallest in size. e.g.

Size of	Na^+ >	Mg^{2+} >	Al^{3+}
Number of e^-:	10	10	10
Number of p :	11	12	13

Greater is the positive charge on the cation of a given atom, smaller is its ionic radius. e.g.

Size of $Fe^+ > Fe^{2+} > Fe^{3+}$

Similarly, size of $Sn > Sn^{2+} > Sn^{4+}$

Radius of Anion

The radius of anion is always greater than its neutral atom. It is due to decrease in effective nuclear charge.

In case of isoelectronic anions, the anion having greater number of protons is smallest in size. e.g.

Size of	N_3^- >	O^{2-} >	F^-
Number of e^-	10	10	10
Number of p	7	8	9

Greater is the negative charge on an atom, greater is its radius. e.g. Out of O^{2-}, O^- and O the order of size is

$$O^{2-} > O^- > O$$

Trend in Ionic Radius

Following trends are visible for ionic radii

- In case of isoelectronic species, the ionic radii decreases with increase in atomic number.
- In a group, the ionic radius of cation as well as anion increases on moving downwards.

The above written generalised facts for cations and anions can be evaluated with the help of formula,

$$\frac{Z}{e} \propto \frac{1}{r}$$

where, Z = atomic number of element,
e = number of electrons
r = radius of ion

The formula must be used separately for cation and anions.

Example 5. *Consider the isoelectronic species,* Na^+, Mg^{2+}, F^- *and* O^{2-}. *The correct order of increasing length of their radii is* **(NCERT Exemplar, JEE Main 2020)**

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

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

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

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

Sol. (*b*) In case of isoelectronic species

$$\text{ionic radii} \propto \frac{1}{\text{atomic number}}$$

Ion	Mg^{2+} <	Na^+ <	F^- <	O^{2-}
Atomic number	(12)	(11)	(9)	(8)

Thus, ionic radii have the following order

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

Example 6. *The correct order of the ionic radii of* O^{2-}, N^{3-}, F^-, Mg^{2+}, Na^+ *and* Al^{3+} *is* **(JEE Main 2020)**

(a) $N^{3-} < O^{2-} < F^- < Na^+ < Mg^{2+} < Al^{3+}$

(b) $Al^{3+} < Na^+ < Mg^{2+} < O^{2-} < F^- < N^{3-}$

(c) $Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$

(d) $N^{3-} < F^- < O^{2-} < Mg^{2+} < Na^+ < Al^{3+}$

Sol. (c) Size of species $\propto \dfrac{1}{\text{Nuclear charge}}$

Iso-electronic species are those atoms or ions which has the same number of electrons.

Size of species decreases with increasing protons.

More is effective nuclear charge (Z_{eff}) lesser will be ionic size.

Correct order of ionic radii

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

Ionisation Energy or Ionisation Potential or Ionisation Enthalpy

The ionisation energy of an element is defined as "*the amount of energy required to remove an electron from the outermost orbit of an isolated gaseous atom*."

When such a definition is written in the terms of work done. Then, the property is called **ionisation potential**, similarly, when terms of heat evolved, i.e. change in heat is used then property is called **ionisation enthalpy**.

The unit of ionisation energy and ionisation enthalpy are kJ/mol that of ionisation potential are **eV/atom.** These are denoted by word I.

Characteristics of ionisation energy are

- An atom can have as many ionisation enthalpies as number of electrons it has. These are denoted as $I_1, I_2, I_3, \ldots$ etc., for 1st electron, 2nd electron, 3rd electron and so on respectively.
- The value of second ionisation energy (I_2) is always greater than first ionisation energy (I_1). Similarly that of I_3 is greater than I_2 and so on.
- The reason for such an increase is increased effective nuclear charge after removal of every electron from the atom. This result in more difficulty in removing the next electron from the outermost orbit of cation formed.

Factors Affecting Ionisation Enthalpy

Important factors that affect the ionisation enthalpy are as follows

- The ionisation energy of an element depends upon its size as

$$\text{IE} \propto \frac{1}{\text{size of atom / ion}}$$

  i.e. smaller the atom/ion, larger will be the ionisation energy.
- Atoms and ions having inert gas configuration, i.e. stable completely filled or half-filled configurations have high Ist ionisation energies.
- In simpler words, completely filler or it is relatively easy to remove an electron from a partially filled valence shell, i.e. where Z_{eff} is lower but it is relatively difficult to remove an electron from an atom or ion that has a filled valence shell, i.e. when Z_{eff} is higher.
- The order of energy required to remove an electron from different subshells (if rest of the conditions are same) is

$$s > p > d > f$$

- It is due to the closeness of above written subshells with the nucleus. Out of the four given subshells, the s-subshell is nearest to the nucleus while f-subshell is farthest from the nucleus.

Trends in Ionisation Energy

Following generalised trends are visible regarding ionisation energy.

1. In general, the *first ionisation energy decreases in a regular way on descending the main groups.*
2. On moving across a period, the atomic size decreases and nuclear charge increases and therefore, the force of attraction exerted by the nucleus on the electron in outermost shell increases. Hence, (IE) increases along a period from left to right.

Exceptions

Exceptions of above written trend are as follows

- *First ionisation energy of group* 2 (IIA) *elements is greater than the corresponding elements of group* 13(IIIA). It is due to the presence of fully-filled s-orbitals in case of group 2 elements, i.e. their outermost electronic configuration is ns^2. It is always difficult to pull out electron from stable configuration.
- *First ionisation energies of group* 15(VA) *elements is greater than the corresponding elements of group* 16(VIA). It is due to the presence of stable half-filled electronic configuration in group 15 elements. Their outermost electronic configuration is ns^2np^3.
- In case of transition elements opposite trends are observed. Thus, $(\text{IE})_1$ *of the corresponding elements of* 3*d and* 4*d-series are almost similar* but these are smaller than the $(\text{IE})_1$ of the 5d-series elements. Certainly the higher values of ionisation energies of the 5d-transition elements are consistent with the relatively small size of the atoms.

Example 7. *The first ionisation enthalpies of* Na, Mg, Al *and* Si *are in the order* ***(NCERT Exemplar)***

(a) Na < Mg > Al < Si (b) Na > Mg > Al > Si

(c) Na < Mg < Al < Si (d) Na > Mg > Al < Si

Or

The first ionisation energy (in kJ/mol*) of* Na, Mg, Al *and* Si *respectively, are* ***(JEE Main 2020)***

(a) 496, 577, 737, 786

(b) 786, 737, 577, 496

(c) 496, 577, 786, 737

(d) 496, 737, 577, 786

Sol. (*a*) Follow the following steps to solve out such problems.

Steps		Apply
***Step* I**	Write the electronic configuration to find position in the periodic table.	Na_{11} = [Ne]$3s^1$, Mg_{12} = [Ne]$3s^2$, Al_{13} = [Ne]$3s^2 3p^1$, Si_{14} = [Ne]$3s^2, 3p^2$
***Step* II**	Arrange them in the order as they are in the periodic table.	1 2 13 14 Na Mg Al Si
***Step* III**	Follow the general trend and also keep in mind the exception.	The IP increases along a period from left to right but IP of Mg is higher than that of Al due to completey filled 3*s* orbital in Mg.
***Step* IV**	On the above basis find the order.	The order of IP is Na < Mg > Al < Si. Thus, (a) is the correct option.

Option (d) is correct in the or 'part' question.

Important Facts Related to Trends in Ionisation Enthalpy

- Transition elements have almost same ionisation enthalpy. However, *the first ionisation enthalpy of elements of group* 12 (IIB) *is higher than their immediate neighbours*. This is because of the stable configuration, i.e. $(n-1)\,d^{10}ns^2$ of these elements.
- Alkali metals have the lowest and inert gases have the highest IE in a period.
- Among all the known elements. He *has the highest and* Cs *has the lowest ionisation energy*.
- As the ionisation energy increases, basic character of hydroxides increases. Thus, IE gives a rough idea about basic character of hydroxides.
- The IE, of the metalloids generally fall between those of metals and non-metals. The difference in IE suggests why metals always form cations and non-metals form anions.
- Metallic or electropositive character of elements increases as the value of ionisation potential decreases. Thus, both these properties show opposite trends across the period or down the group.
- The relative reactivity of the elements increases with the decrease in IP values. The IP values of inert gases are very high. So they are almost inactive.
- In a group moving from top to bottom the reactivity of elements atoms increases because their IP value decreases.
- The reducing power of elements increases as the values of IP decreases because tendency to lose the electron increases.
- This reducing power increases on going down a group as,

$$\text{Li} > \text{Cs} > \text{Rb} > \text{K} > \text{Na}$$

Example 8. B *has a smaller first ionisation enthalpy than* Be. *Consider the following statements*

I. *It is easier to remove 2p-electron than 2s-electron.*
II. *2p-electron of* B *is more shielded from the nucleus by the inner core of electrons than the 2s-electrons of* Be.
III. *2s-electron has more penetration power than 2p-electron.*
IV. *atomic radius of* B *is more than* Be.
(atomic number B = 5, Be = 4 *)*

The correct statements are **(JEE Main 2020)**

(a) (I), (II) and (III)
(b) (II), (III) and (IV)
(c) (I), (III) and (IV)
(d) (I), (II) and (IV)

Sol. (*a*) Statements I, II and III are correct, whereas statement IV is incorrect. It's correct statement is, atomic radius of B is less than Be.

Example 9. *The element having greatest difference between its first and second ionisation energy, is* **(JEE Main 2019)**

(a) Ca
(b) Se
(c) Ba
(d) K

Sol. (*d*) The electronic configuration of given elements are as follows

$$\text{K}(19) = 1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^1$$
$$\text{Ca}(20) = 1s^2\, 2s^2\, 2p^6\, 3s^2 3p^6\, 4s^2$$
$$\text{Sr}(34) = [\text{Ar}]\, 4s^2 3d^{10}\, 4p^4$$
$$\text{Ba}(56) = [\text{Xe}]\ 6s^2$$

First ionisation enthalpy IE_1 of K is lowest among the given options. Here, the energy required to remove an electron from $4s^1$ is least as only one electron is present in the outermost shell. IE_1 is comparatively high for Ca and Ba as two electrons (fully-filled) are placed in *s*-orbital.

Second ionisation enthalpy of K is highest among the given options. It is because removal of an electron occur from p^6 (fully-filled). So, high energy is required to remove the electron. From the above discussion, it can be concluded that $(IE_2 - IE_1)$ value is maximum for K (potassium).

Electron Gain Enthalpy

It is defined as, "*The energy liberated when an extra electron is added to an isolated gaseous atom* with energy terminology this property is called **electron affinity** (A_e) of that atom while the heat change occurring during this process is called the **electron gain enthalpy** $(\Delta_{e_g} H)$.

The units of electron affinity or electron gain enthalpy are eV/atom or kJ/g-atom.

Features of electron gain enthalpy are as follows

- Electron affinity and electron gain enthalpy have same magnitude but opposite sign. In real sense, electron affinities are defined at absolute zero while electron gain enthalpy at 298 K.
  These two quantities at a given temperature are related as

$$\Delta_{e_g} H = -A_e - \frac{5}{2} RT$$

- For the process, $\mathrm{Cl}(g) + e^- \longrightarrow \mathrm{Cl}^-(g) + E_A$
- Electron affinity is positive but $\Delta_{e_g} H$ is negative.
- Electron affinity describes the tightness with which an atom binds the extra electron added to it, i.e. greater the value of A_e (energy liberated) or $\Delta_{e_g} H$, greater is the binding force.
- Just like IP, an atom possess as many electron gain enthalpies as many electrons it has.
- For second onwards, the value of $\Delta_{e_g} H$ always positive means here for addition of electron, energy is to be absorbed from the system. e.g. The first electron gain enthalpy of oxygen has positive value (141 kJ / mol).

  $$\mathrm{O}(g) + e^- \longrightarrow \mathrm{O}^-(g); \Delta H = -141 \text{ kJ/mol}$$

  which means that the process is exothermic.
- On the other hand, the electron affinity of the O^- ion is highly negative (–780 kJ/mol).

  $$\mathrm{O}^-(g) + e^- \longrightarrow \mathrm{O}^{2-}(g); \Delta H = 780 \text{ kJ/mol}$$

  which means the process is **endothermic** even though O^{2-} ion has a stable noble gas configuration. This process is unfavourable in the gas phase because the resulting increase in electron-electron repulsion outweighs the stability gained by achieving a noble gas configuration.
- In solid phase, the process becomes favoured because O^{2-} ions stabilised by the adjacent cations to form ionic lattice releasing lattice energy.

Factors Affecting Electron Gain Enthalpy

Various factors which affect the $\Delta_{e_g} H$ are

- Magnitude of $\Delta_{e_g} H \propto Z_{\text{eff}}$
- Magnitude of $\Delta_{e_g} H \propto \dfrac{1}{\text{atomic size}}$
- Magnitude of $\Delta_{e_g} H \propto$ penetrating power
  Therefore, the order of $\Delta_{e_g} H$ is $s > p > d > f$.
- Stability of half-filled and completely filled orbitals.

Trends in Electron Gain Enthalpy

Following generalised trends are visible regarding electron gain enthalpy.

- In a period, *the electron gain enthalpy (negative) increases on moving from left to right.* Thus, the tendency to accept electrons in a period increase and the electron gain enthalpy values become more and more negative. The halogens have the highest electron affinity values in a given period.
- The value of electron gain enthalpy (negative) in a group decreases. This variation is related with the size of atom and electron density in the outermost shell of the atom concerned.

Exceptions

Few exception of above written trend are as follows

- ***$\Delta_{e_g} H$ values of the elements of group*** 2 (IIA) ***and group*** 15 (VA) ***are positive.*** It is due to the presence of outer fully filled s-orbital in group 2 and half-filled p-orbital in group 15 because of which these atoms do not have the tendency to accept electron.
- *The noble gases have no or very little tendency to accept electrons*, thus all the noble gases have electron gain enthalpy more than zero. This means that the anions of noble gases (if formed) would be inherently unstable.
- In general, the $\Delta_{e_g} H$ values (negative) of the elements of *third period are higher than the $\Delta_{e_g} H$ values of the elements of second period*. It is due to small radius and high electron density in the outermost orbit of second period elements.
- The small size here signifies limited capacity to release energy, while high electron density signifies greater repulsive forces which extra electron experiences while it is to be added. e.g. The $\Delta_{e_g} H$ value of Cl is higher than that of F as Cl belongs to 3rd period while F belongs to 2nd period.
- The electron gain enthalpy of metals are generally lower (less negative) than those of non-metals.
- It has been found that elements having higher negative values of $\Delta_{e_g} H$, e.g. $\mathrm{F_2}$, $\mathrm{Cl_2}$, $\mathrm{O_2}$, $\mathrm{Br_2}$ etc., are good oxidants.

Example 10. *Among halogens, the correct order of amount of energy released in electron gain (electron gain enthalpy) is* ***(NCERT Exemplar)***

(a) $\mathrm{F > Cl > Br > I}$
(b) $\mathrm{F < Cl < Br < I}$
(c) $\mathrm{F < Cl > Br > I}$
(d) $\mathrm{F > Cl < Br < I}$

Or

The electron gain enthalpy (in kJ/mol) of fluorine, chlorine, bromine and iodine, respectively, are ***(JEE Main 2020)***

(a) −333, −325, −349 and −296
(b) −296, −325, −333 and −349
(c) −333, −349, −325 and −296
(d) −349, −333, −325 and −296

Sol. (c) To solve out such problem, follow the same steps as shown in solved example 7.

From this you get the order $\mathrm{F < Cl > Br > I}$

Option (c) is correct for 'or part' question.

Example 11. *Within each pair of elements* F *and* Cl, S *and* Se, *and* Li *and* Na, *respectively, the elements that release more energy upon an electron gain are* ***(JEE Main 2020)***

(a) F, Se and Na
(b) F, S and Li
(c) Cl, S and Li
(d) Cl, Se and Na

Sol. (c) The first electron gain enthalpy is exothermic (or negative).

Generally, electron gain enthalpy becomes less exothermic (or less negative) when comparing elements of a group from top to bottom.

Therefore, electron gain enthalpy of S > Se and Li > Na.

But there are some exceptions to this.

One of them is the case of a group 17 elements where electron gain is most negative for Cl instead of F, due to extra small size of fluorine.

∴ Upon an electron gain, energy releases in the order

Cl > F, S > Se and Li > Na

Electronegativity and Electronegative Character

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

Remember it is different from electronegative character of an atom which is infact the tendency of an element to form the anion by gaining the electron. The first term is comparative while the second term is used for a single atom.

Factors Affecting the Magnitude of Electronegativity

Various factors affecting the magnitude of electronegativity are as follows

- As the atomic radius of the elements increases, the electronegativity value decreases.

$$\text{Electronegativity} \propto \frac{1}{\text{Atomic radius}}$$

- The electronegativiy value increases as the effective nuclear charge on the atomic nucleus increases.

$$\text{Electronegativity} \propto \text{Effective nuclear charge } (Z_{\text{eff}})$$

- The electronegativity increases as the oxidation state (i.e. the number of positive charge) of the atom increases.
- If the *s*-character in the hybridisation state of the atom increases, electronegativity also increases because *s*-electrons are comparatively near to the nucleus. e.g.

  The approximate electronegativity values of C-atom in various hybridisation states are as under

Hybridisation states	sp^3	sp^2	sp
s-character	25%	33.33%	50%
Electronegativity	2.48	2.75	3.25

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

s-character is increasing

So, the electronegativity value is increasing.

Trends in Electronegativity

Different tends shown by electronegativity along periods and groups are given below

- In a period, on moving from left to right, the electronegativity increases due to the increase in effective nuclear charge/increase in ionisation energy.
- In a group on moving from top to bottom, the electronegativity decreases because atomic radius increases.
- The electronegativity value of F is maximum and that of Cs is minimum in the periodic table.

Diagonal Relationship and Electronegativity

- These are examples of elements that lying diagonally in periodic table. Such elements show almost similar properties. e.g. Li—Mg, Be—Al and B—Si. It is because offollowing reason.
- As we move from left to right across a period, the increased nuclear charge holds the electrons more tightly to the atom.
- On moving downwards one place in a group, the extra shell of electrons decreases the attraction of the nucleus for the outer electrons.
- A diagonal move means that these effects tend to compensate for one another.
- The electronegativity increases as we go from Li to Be but it decreases as we move from Be to Mg.
- As we move diagonally, these two effects partly cancel each other and there is no marked change in electronegativity and thus, Li and Mg have close values of electronegativities.
- Because of this, Li and Mg show similarity in properties—a diagonal relationship. Measurement of electronegativity

Since, electronegativity is a tendency and tendencies cannot be measured accurately, electronegativity of different elements is calculated on different scales.

Some important of these scales are as follows

(i) **Pauling scale** We all know, that F is the most electronegative element of the periodic table. The electronegativity of fluorine is arbitrarily assigned 4.0 by Pauling. The electronegativity of other elements is calculated against this standard.

The relative order of electronegativity of some important elements is

Max. of all the elements known ↙	F	> O	> Cl ≈ N	> Br	> S ≈ C ≈ I	> H
	4.0	3.5	3.0	2.8	2.5	2.1

The calculations are done with the help of bond energies used in the formula,

$$X_A - X_B = 0.208\sqrt{\Delta}$$

$$\text{where,} \quad \Delta = E_{A-B} - \sqrt{E_{A-A} \times E_{B-B}}$$

$X_A - X_B$ = electronegativity difference between two atoms A and B.

E_{A-B}, E_{A-A}, and E_{B-B} are the bond energies of the molecules AB, A_2 and B_2 respectively.

Different other elements have following values of electronegativity based on Pauling's scale.

						H
						2.1
Li	Be	B	C	N	O	F
1.0	1.5	2.0	2.5	3.0	3.5	4.0
Na	Mg	Al	Si	P	S	Cl
0.9	1.2	1.5	2.8	2.1	2.5	3.0
K	Ca	Ga	Ge	As	Se	Br
0.8	1.0	1.6	1.8	2.0	2.4	2.8
Rb						I
0.8						2.5
Cs						
0.7						

(ii) **Mulliken scale** at which electronegativity is determined from the values of A_e and IE of atom. This method is not much in use.

On this scale electronegativity is denoted by word *chi* (χ) which can be calculated as, $\chi = \frac{\text{IE} + A_e}{2}$

- If IE and A_e are determined in eV then
$$\chi = (\text{IE} + A_e) / 5.6 \text{ eV}$$
- If IE and A_e are determined in kJ/g-atom then
$$\chi = (\text{IE} + A_e) / 125 \text{ kcal / g-atom}$$
- If IE and A_e are in kJ/mol then
$$\chi = (\text{IE} + A_e) / 544 \text{ kJ / mol}$$

- The electronegativity values we get on this scale are 2.8 times greater than those of Pauling scale.

Example 12. *In general, the property (magnitudes only) that shows an opposite trend in comparison to other properties across a period is* ***(JEE Main 2020)***

(a) electronegativity (b) electron gain enthalpy

(c) ionisation enthalpy (d) atomic radius

Sol. (*d*) With increase in atomic number (Z) along a period, effective nuclear charge (Z^*) of the elements also increases. We know,

$$Z \propto \text{Electronegativity } (\chi)$$
$$\propto \text{Electron gain enthalpy or } |\Delta_{e_g} H|$$
$$\propto \text{Ionisation enthalpy or } \Delta_i H \propto \frac{1}{\text{Atomic radius } (r)}$$

So, with increase in atomic number across a period. Electronegativity, electron gain enthalpy and ionisation enthalpy increases whereas atomic radius decreases.

Note $Z^* = Z - \sigma$

where, σ = shielding constant which can be calculated using Slatter's rule.

Example 13. *The ionisation energy and electron affinity of an element are* 12.0 eV *and* 3.8 eV *respectively. Its electronegativity on the Pauling scale is*

(a) 3.0 (b) 4.0 (c) 3.5 (d) 2.7

Sol. (*d*) Electronegativity on the Mulliken's scale

$$= \frac{1}{2}(12.0 + 3.8) = 7.9.$$

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 $\frac{7.9}{2.8} = 2.7$.

(iii) **Allred-Rochow's scale** is considered as the real scale as it is based upon covalent radius of an atom.

According to Allred-Rochow, electronegativity is a force which is exerted by the nucleus on the electrons of outermost orbit,

i.e. $$E_n = \frac{0.359 \times Z_{\text{eff}}}{r^2} + 0.744$$

where, r = covalent radius

$Z_{\text{effective}}$ = effective nuclear charge

Applications of Electronegativity

- The electronegativity values are used to predict the nature of A—B bond as
  - (i) Greater is the difference in the electronegativity of elements A and B, greater will be the ionic character of bond between A and B or weaker will be the covalent character.
  - (ii) If electronegativities of A and B are equal or difference is very small, the bond between A and B is supposed to be purely covalent.
- The electronegativity value also gives an idea about the bond lengths. As the electronegativity difference increases, bond length decreases.
- If the electronegativity difference (Δx) of covalently bonded atoms increases, the bond energy of the covalent bond also, increases, e.g.

  The order of the H—X bond strength is
  $\text{H—F} > \text{H—Cl} > \text{H—Br} > \text{H—I}$.

  As the bond strength is decreasing, the acid strength is increasing. So order of increasing acid strength is
  $\text{HF} < \text{HCl} < \text{HBr} < \text{HI}$.

  The acidic nature of the oxides of normal elements increase as we move from left to right in a period.
- In a period from left to right the electronegativity of the elements increases.

  The order of acidic or basic nature of the oxides of third period elements is given below.

$$\xrightarrow{\text{Na}_2\text{O} \quad \text{MgO} \quad \text{SiO}_2 \quad \text{P}_2\text{O}_5 \quad \text{SO}_3 \quad \text{Cl}_2\text{O}_7}$$

Basic nature is decreasing or acidic nature is increasing.

- The metallic character decreases as the electronegativity of the elements increases, i.e. from left to right.

Electropositive or Metallic Character

The tendency of an element to form the cation by the loss of electron is called its electropositive character or metallic character.

Thus, metals are the elements which have the tendency to lose electrons.

Trend in Periodic Table

The variation of electropositive or metallic character in the periodic table is given below.

→ **Decrease** (across the period)
Electropositive or metallic character
↓ **Increase**
(down the group)

This tendency varies inversely with ionisation energy and electronegativity of an element.

Periodic Trends in Chemical Properties

The periodic trends in valency and chemical properties of elements are as follows

1. Valency

It is the **combining capacity of an element** and can never be equal to zero. Remember it is only a number and is not assigned as positive or negative sign. e.g. The valency of C in CH_4, CCl_4, CH_2Br_2 all is 4. Along a period from left to right it increases from 1 to 7 with respect to hydrogen and then decreases to 1 with respect to oxygen.

For the elements of a group it remains the same because the number of valence electrons (electrons of valence or outer shells) remains the same.

2. Oxidation States

It is the residual charge left on an atom of a molecule when other atoms of the molecule are removed as ions. Its value can be zero, positive, negative or fractional. An element can have different oxidation states in different compounds, e.g.

- Oxidation state of N in NO_2, NO, N_2O_3 and N_2O_5 are $+4, +2, +3$ and $+5$ respectively.
- In case of alkali and alkaline earth metals it is fixed, i.e. $+1$ and $+2$ respectively.
- It is generally variable for other elements. Higher *p*-block elements exhibit more stability for lower oxidation state due to inert pair effect, e.g. Pb^{4+} is less stable than Pb^{2+}.
- In case of *d*-block elements, max. oxidation state $= (n-1)\,d\,e^-$ (unpaired) $+\, ns\, e^-$.
- In general, the stability of the higher oxidation states is in the order

$$3d \ll 4d < 5d$$

3. Chemical Reactivity

It is highest at the two extremes of a period and is lowest at the centre. This is because the elements present at left extreme of a period have the lowest ionisation enthalpy and that present at the right extreme have the highest electron affinity.

Highly reactive elements are always found in combined state. In the groups, the reactivity of former elements (that present at left extreme) increases while that of latter decreases.

4. Reducing and Oxidising Properties

Reducing agents provide electrons, thus **metals act as good reducing agents**. The reducing power of elements in the periodic table shows the following trend.

→ **Decrease** (across the period)
Reducing power of elements
↓ **Increase** (down the group)

Oxidising agents accept electrons, thus **non-metals are good oxidising agents**.

The oxidising power of elements shows following trend in the periodic table.

→ **Increase** (across the period)
Oxidising power of elements
↓ **Decrease** (down the group)

Thus, alkali metals are most powerful reducing agents while halogens are most powerful oxidising agents in the periodic table.

5. Nature of Oxides

Metal oxides are basic, these form alkaline solutions when dissolved in water. Similarly, non-metal oxides are acidic while semi-metal oxides are amphoteric in nature. Thus, the basic character of oxides shows following trend in the periodic table.

→ **Decreases** (across the period)
Basic nature of oxide
↓ **Increases** (down the group)

Similarly, acidic nature of oxides shows the following trend

→ **Increases** (across the period)
Acidic nature of oxides
↓ **Decreases** (down the group)

- The oxides CO, N_2O, NO and H_2O are neutral in nature.

  The oxides of Al, Zn, Sn, As and Sb along with metalloids are amphoteric in nature.
- The acidic character of the oxides of a non-metal increases with increase in oxygen content in it, e.g. $N_2O_3 < N_2O_5$; $SO_2 < SO_3$; $Cl_2O < Cl_2O_3 < Cl_2O_7$.

- Non-metal oxides when dissolved in water form oxy-acids, thus, they are called anhydrides of the acid formed.

Example 19. *Three elements X, Y and Z are in the 3rd period of the periodic table. The oxides of X, Y and Z, respectively, are basic, amphoteric and acidic. The correct order of the atomic number of X, Y and Z is* ***(JEE Main 2020)***

(a) $Z < Y < X$ (b) $X < Y < Z$
(c) $X < Z < Y$ (d) $Y < X < Z$

Sol. (*b*) While moving across a period, metallic property of elements changes first into metalloids than into non-metallic nature, therefore metal gives basic oxides, metalloid produces amphoteric oxides and non-metal gives acidic oxides.

So, $X \rightarrow$ metal, $Y \rightarrow$ metalloid and $Z \rightarrow$ non-metal

$\Rightarrow$ Order of atomic number : $X < Y < Z$

$\Rightarrow$ As, X, Y and Z belongs to 3rd period,

X can be Na or Mg

Y can be Al or Si

Z can be P or S or Cl

A summary of variation of some important periodic properties is given below.

Periodic property	**Variation**	
	Along a period	**Along a group**
Atomic radius	Decreases	Increases
Ionic radius	Decreases	Increases
Metallic nature	Decreases	Increases
Electron gain enthalpy	Increases	Decreases
Ionisation energy	Increases	Decreases
Electronegativity	Increases	Decreases
Basic nature of oxides	Decreases	Increases
Acidic nature of oxide	Increases	Decreases
Reducing power	Decreases	Increases

Other Properties of Elements

Some other properties of elements are discussed below

1. Density

There is no regular trend in the variation of density in a group or period. Density of osmium (Os) is the highest ($d = 22.6$ g/cc) among all the known elements. Lightest metal is lithium ($d = 0.54$ g/cc). Density of H_2 is lowest among all the known elements.

2. Melting and Boiling Points

There is no regular trend in the variation of melting point and boiling point in a group or period.

- Tungsten (W) has the highest melting poing among all the metals.
- Mercury has the lowest melting poing (–38°C) among all the metals.
- Carbon in the form of diamond has highest melting point (3277°C) among all the known elements.
- Helium has lowest melting point (–270°C) among all the elements.
- In a period, inert gases have lowest melting point and boiling point
- In a group, the melting and boiling points of metals usually decreases while that of non-metals increases.

3. Atomic Volume

Volume occupied by one gram atom of an element in solid state is called its atomic volume.

Atomic volume = mass of one gram atom/density

Atomic volume of elements increases on moving down the group due to the presence of an extra shell of electrons. On moving from left to right in a period atomic volume first decreases and becomes minimum in the middle and then again begins to increase.

On plotting a graph between atomic number of elements and their atomic volume, a curve is obtained in which alkali metals are present at peaks while elements of 'boron family' Ni, Ru and H etc., are present at troughs.

Elements Showing Highest and Lowest Properties

Many times questions based on highest and lowest of some properties are seen in the JEE.

Here, we are providing some elements showing highest and lowest properties.

Highest and Lowest of Some Properties

Highest melting point	**Carbon** (diamond)
Highest EA	Chlorine
Highest electronegativity	F
Highest ionisation potential	He
Lowest ionisation potential	Cs
Highest density	Os
Lowest density	H_2
Lowest density among metals	Li
Highest melting point among metals	Tungsten
Lowest melting point	He
Smallest anion	H^-
Largest anion	I^-
Smallest cation	H^+
Largest cation	Cs^+

Practice Exercise

ROUND I Topically Divided Problems

Classification of Elements

1. Newlands' law of octave applies to which of the following set of elements?

(a) Be, Mg, Ca (b) As, K, Ca
(c) B, N, C (d) None of these

2. Mendeleef 's periodic table is upset by the fact that

(a) many elements have several isotopes
(b) noble gases do not form compounds
(c) some groups stand divided into two sub groups *A* and *B*
(d) atomic weights of elements are not always whole numbers

3. The basis of keeping the elements in the groups of the periodic table is

(a) ionisation potential
(b) electronegativity
(c) electron affinity
(d) number of valence electrons

4. Which of the following statements related to the modern periodic table is incorrect ?

(a) The *p*-block has 6 columns, because a maximum of 6 electrons can occupy all the orbitals in a *p*-shell
(b) The *d*-block has 8 columns because a maximum of 8 electrons can occupy all the orbitals in a *d*-subshell
(c) Each block contains a number of columns equal to the number of electrons that can occupy that subshell
(d) The block indicates value of azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration

5. According to the periodic law of elements, the variation in properties of elements is related to their

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

6. Who developed the long form of the periodic table?

(a) Niels Bohr (b) Moseley
(c) Mendeleef (d) Lothar Meyer

7. Which of the following pairs has both members of the same period of the periodic table?

(a) Na—Cl (b) Na—Ca
(c) Ca—Cl (d) Cl—Br

8. Why do elements in the same group have similar physical and chemical properties?

(a) Because of same electronic configuration
(b) Because of same number of electrons
(c) Because of same number of protons
(d) Because of same valence electrons

9. The statement that is true for the long form of the periodic table is

(a) it reflects the sequence of filling the electrons in the order of sub-energy levels *s, p, d* and *f*
(b) it helps to predict the stable valency states of the elements
(c) it reflects trends in physical and chemical properties of the elements
(d) All of the above

10. 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

11. Find the element having outer electronic configuration $(n-1)\,d^2ns^2$ for $n = 4$.

(a) Scandium
(b) Vanadium
(c) Titanium
(d) Tungsten

12. The element with outer shell electronic configuration $(n-2)\,f^7\,(n-1)\,d^1ns^2$ for $n = 6$ is … .

(a) uranium
(b) gadolinium
(c) lead
(d) niobium

13. The statement that is not correct for periodic classification of elements is

(a) the properties of elements are periodic functions of their atomic numbers
(b) non-metallic elements are less in number than metallic elements
(c) for transition elements, the $3d$-orbitals are filled with electrons after $3p$-orbitals and before $4s$-orbitals
(d) the first ionisation enthalpies of elements generally increase with increase in atomic number as we go along a period

14. The position of the element having outer shell electronic configuration, ns^2np^4 for $n = 3$ is

(a) II period, fourth group
(b) III period, fourth group
(c) II period, 16 group
(d) III period, 16 group

15. Which one of the following belongs to representative group of elements in the periodic table?

(a) Aluminium (b) Chromium
(c) Argon (d) Lanthanum

16. The characteristics of elements X, Y and Z with atomic numbers, respectively, 33, 53 and 83 are

(JEE Main 2021)

(a) X and Y are metalloids and Z is a metal.
(b) X is a metalloid, Y is a non-metal and Z is a metal.
(c) X, Y and Z are metals.
(d) X and Z are non-metals and Y is a metalloid

17. Write the atomic number of the element present in the third period and seventeenth group of the periodic table.

(a) 15 (b) 17 (c) 21 (d) 9

18. According to IUPAC nomenclature, a newly discovered element has been named as Uun. The atomic number of the element is

(a) 111 (b) 112 (c) 109 (d) 110

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

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

Atomic and Ionic Radii

20. In third row of the periodic table, the atomic radii from Na to Cl

(a) continuously decreases
(b) continuously increases
(c) remains constant
(d) increases but not continuously

21. Which of the following given atoms has the largest atomic radius?

(a) B (b) Li
(c) O (d) Ne

22. Which of the following atoms would be expected to have the largest atomic radii ?

(a) Oxygen (b) Aluminium
(c) Calcium (d) Mg

23. Which of the following has largest ionic radius?

(a) Li^+ (b) K^+
(c) Na^+ (d) Cs^+

24. Which of the following ion is the smallest ion?

(a) O_2 (b) O_2^+
(c) O_2^- (d) O_2^{2-}

25. The radii of F, F^-, O and O^{2-} are in the order of

(a) $F^- > O^{2-} > F > O$ (b) $F > F^- > O > O^{2-}$
(c) $O^{2-} > F^- > O > F$ (d) $F > O > F^- > O^{2-}$

26. Consider the following species

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

The property which is common in them is

(a) number of electrons
(b) number of valence electrons
(c) number of protons
(d) All of the above

27. The correct order of increasing ionic radii is

(a) $Na^+ < Mg^{2+} < Al^{3+} < N^{3-} < O^{2-} < F^-$
(b) $Al^{3+} < Mg^{2+} < Na^+ < N^3 < O^{2-} < F^-$
(c) $Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$
(d) $F^- < O^{2-} < N^{3-} < Al^{3+} < Mg^{2+} < Na^+$

28. Which of the following has the smallest size ?

(a) Na^+ (b) Mg^{2+}
(c) Al^{3+} (d) P^{5+}

29. Which of the following statements is correct?

(a) X^- ion is larger in size than X-atom

(b) X^+ ion is larger in size than X-atom

(c) X^+ ion is larger in size than X^- ion

(d) X^+ and X^- ions are equal in size

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

(a) Cl^-, Ca^{2+}, K^+, S^{2-}

(b) S^{2-}, Cl^-, Ca^{2+}, K^+

(c) Ca^{2+}, K^+, Cl^-, S^{2-}

(d) K^+, S^{2-}, Ca^{2+}, Cl^+

31. The ionic radius of Na^+ ions is 1.02 Å. The ionic radii (in Å) of Mg^{2+} and Al^{3+}, respectively, are *(JEE Main 2021)*

(a) 1.05 and 0.99

(b) 0.72 and 0.54

(c) 0.85 and 0.99

(d) 0.68 and 0.72

32. The set representing the correct order of ionic radius is

(a) $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$

(b) $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$

(c) $Li^+ > Na^+ > Mg^{2+} > Be^{2+}$

(d) $Mg^{2+} > Be^{2+} > Li^+ > Na^+$

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

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

34. Ce^{3+}, La^{3+}, Pm^{3+} and Yb^{3+} have ionic radii in the increasing order as

(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) $Yb^{3+} < Pm^{3+} < La^{3+} < Ce^{3+}$

Ionisation Enthalpy

35. Arrange the elements N, P, O and S in the order of increasing first ionisation enthalpy.

(a) P < S < O < N

(b) S < P < O < N

(c) S < O < P < N

(d) P < S < N < O

36. Among the given elements the highest first ionisation enthalpy is of

(a) B (b) Al (c) C (d) Si

37. As one moves along a given row in the periodic table, ionisation energy

(a) increases from left to right

(b) decreases from left to right

(c) first increases, then decreases

(d) remains the same

38. The first ionisation energy of lithium will be

(a) greater than Be

(b) less than Be

(c) equal to that of Na

(d) equal to that of F

39. When the first ionisation energies are plotted against atomic number, the peaks are occupied by

(a) alkali metals

(b) halogens

(c) transition metals

(d) rare gases

40. The ionisation energy of nitrogen is larger than that of oxygen because of

(a) greater attraction of electrons by the nucleus

(b) the size of nitrogen atom being smaller

(c) the half-filled *p*-orbitals possess extra stability

(d) greater penetration effect

41. The correct order of ionisation energy for comparing carbon, nitrogen and oxygen is

(a) C < N > O

(b) C > N < O

(c) C > N > O

(d) C < N < O

42. Identify the elements X and Y using the ionisation energy values given below : *(JEE Main 2021)*

	Ionisation energy	**(kJ/mol)**
	1st	2nd
X	495	4563
Y	731	1450

(a) X = Na ; Y = Mg

(b) X = Mg ; Y = F

(c) X = Mg; Y = Na

(d) X = F; Y = Mg

43. Which of the following isoelectronic ions has lowest ionisation energy?

(a) Cl^-

(b) Ca^{2+}

(c) K^+

(d) S^{2-}

44. Which of the following relation is correct?

(a) Ist IE of C > Ist IE of B

(b) Ist IE of C < Ist IE of B

(c) IInd IE of C > IInd IE of B

(d) Both (b) and (c)

45. If the IP of Na is 5.48 eV, the ionisation potential of K will be

(a) same as that of Na

(b) 4.34 eV

(c) 5.68 eV

(d) 10.88 eV

46. The first ionisation energy of oxygen is less than that of nitrogen. Which of the following is the most appropriate reason for this observation ?

(a) Lesser effective nuclear charge of oxygen than nitrogen

(b) Lesser atomic size of oxygen than nitrogen

(c) Greater inter-electron repulsion between two electrons in the same *p*-orbital counter balances the increase in effective nuclear charge on moving from nitrogen to oxygen

(d) Greater effective nuclear charge of oxygen than nitrogen

47. Among the second period elements, the actual ionisation enthalpies are increase in the order as

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

Be has higher $\Delta_i H$ than B. This is because

(a) of the stable configuration of B from which electron has to be removed.

(b) of the fact p-orbital being larger attracted more towards nucleus.

(c) $2s$-orbital is closer to nucleus as compared to $2p$-orbital.

(d) $2p$-orbital is of higher stability than $2s$-orbital.

48. O has lower $\Delta_i H$ than N and F because

(a) its size is smaller than F but larger than N

(b) F is smaller but N have stable configuration

(c) N is smaller and F have stable configuration

(d) its size is smaller than N but larger than F

49. The increasing order of the first ionisation enthalpies of the elements B, P, S and F (lowest first) is

(a) F < S < P < B (b) P < S < B < F
(c) B < P < S < F (d) B < S < P < F

50. The ionic mobility of alkali metal ions in aqueous solution is maximum for

(a) K^+ (b) Rb^+
(c) Li^+ (d) Na^+

Electron Affinity and Other Periodic Properties

51. The correct increasing order of non-metallic character is

(a) P < S < O < N (b) S < P < O < N
(c) S < O < P < N (d) P < S < N < O

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

(a) I > Br > Cl > F (b) F > Cl > Br > I
(c) Cl > F > Br > I (d) Br > Cl > I > F

53. Electron affinity is the

(a) energy released when an electron is added to an isolated atom in the gaseous state.

(b) energy absorbed when an electron is added to an isolated atom in the gaseous state.

(c) energy required to take out an electron from an isolated gaseous atom.

(d) power of an atom to attract an electron to itself.

54. Which of the following pairs show reverse properties on moving along a period from left to right and from top to down in a group?

(a) Nuclear charge and electron affinity

(b) Ionisation energy and electron affinity

(c) Atomic radius and electron affinity

(d) None of the above

55. Which of the following has highest electron affinity?

(a) N (b) O (c) F (d) Cl

56. Elements of which group form anions most readily?

(a) Halogens (b) Alkali metals
(c) Oxygen family (d) Nitrogen group

57. The halogen that most easily reduced is

(a) F_2 (b) Cl_2
(c) Br_2 (d) I_2

58. Which one of the following process requiring absorption of energy?

(a) $Cl \longrightarrow Cl^-$ (b) $H \longrightarrow H^-$
(c) $O \longrightarrow O^{2-}$ (d) $F \longrightarrow F^-$

59. The electron gain enthalpy of fluorine is less negative than that of chlorine because

(a) the added electron experience greater force of attraction from the other electron in case of F

(b) the added electron experience greater force of repulsion from the other electrons in case of F

(c) the size of F is smaller than Cl

(d) Both (b) and (c)

60. The values of electronegativity of atom A and B are 1.20 and 4.0 respectively. The percentage of ionic character of A—B bond is

(a) 58.3% (b) 48% (c) 79.6% (d) 73.6%

61. The elements X, Y, Z and T have the indicated electronic configurations. Starting with the innermost shell, which of the following is the most metallic element ?

(a) $X = 2, 8, 4$ (b) $Y = 2, 8, 8$
(c) $Z = 2, 8, 8, 1$ (d) $T = 2, 8, 8, 7$

62. Among alkali metals which element do you expect to be least electronegative?

(a) K (b) Na (c) Cs (d) Fr

63. With respect to chlorine, hydrogen will be

(a) electropositive (b) electronegative
(c) neutral (d) None of these

Periodic Trends in Chemical Properties

64. The set of elements that differ in mutual relationship from those of the other sets is

(a) Li – Mg (b) B – Si ***(JEE Main 2021)***
(c) Be – Al (d) Li – Na

65. As we go from left to right in period two of the periodic table, gram atomic volume of the elements

(a) will change indefinitely
(b) decreases
(c) increase at a constant rate
(d) first increase then decrease

66. Metallic nature and basic nature of the oxides…… as we move along a period.

(a) increases
(b) decreases
(c) remains constant
(d) first increase then decreases

67. The metal having highest melting point is

(a) Cr (b) Ag (c) diamond (d) W

68. Which one of the following order presents the correct sequence of the increasing basic nature of the given oxides?

(a) $Al_2O_3 < MgO < Na_2O < K_2O$
(b) $MgO < K_2O < Al_2O_3 < Na_2O$
(c) $Na_2O < K_2O < MgO < Al_2O_3$
(d) $K_2O < Na_2O < Al_2O_3 < MgO$

69. Which of the following metals exhibits more than one oxidation state?

(a) Na (b) Mg
(c) Al (d) Fe

70. The correct order of reactivity of halogens is

(a) F > Br > Cl > I (b) F > Cl > Br > I
(c) I > Br > Cl > F (d) Cl > I > Br > F

71. The correct order of increasing oxidising power is

(a) $F_2 < Cl_2 < Br_2 < I_2$ (b) $I_2 < F_2 < Cl_2 < Br_2$
(c) $Br_2 < I_2 < F_2 < Cl_2$ (d) $I_2 < Br_2 < Cl_2 < F_2$

72. Which of the following oxides is most basic?

(a) Na_2O
(b) SiO_2
(c) SO_2
(d) All are equally basic

73. Which one of the following sets of ions represents a collection of isoelectronic species ?

(a) K^+, Cl^-, Ca^{2+}, Sc^{3+}
(b) Ba^{2+}, Sr^{2+}, K^+, S^{2-}
(c) N^{3-}, O^{2-}, F^-, S^{2-}
(d) Li^+, Na^+, Mg^{2+}, Ca^{2+}

ROUND II Mixed Bag

1. If Aufbau rule is not followed, K-19 will be placed in

(a) *s*-block (b) *p*-block (c) *d*-block (d) *f*-block

2. Choose the correct order of atomic radii of fluorine and neon (in pm) out of the options given below.

(a) 72, 160 (b) 160, 160
(c) 72, 72 (d) 160, 72

3. The trivalent ion having largest size in lanthanide series is

(a) Ti (b) Zr (c) Hf (d) La

4. Which of the following has the highest second ionisation energy?

(a) Calcium (b) Chromium
(c) Iron (d) Cobalt

5. The bond length of LiF will be

(a) equal to that of KF (b) more than that of KF
(c) equal to that of NaF (d) less than that of NaF

6. A trend common to both group I and VII elements in the periodic table is that as the atomic number increases the

(a) atomic radius also increases
(b) oxidising power also increases
(c) reactivity with water also increases
(d) maxivalency also increases

7. The correct order of ionic radius is

(a) $Ti^{4+} < Mn^{7+}$ (b) $^{35}Cl^- > ^{37}Cl^-$
(c) $K^+ > Cl^-$ (d) $P^{3+} > P^{5+}$

8. The ionic radius of 'Cr' is minimum in which of the following compounds?

(a) CrO_2 (b) K_2CrO_4
(c) CrF_3 (d) $CrCl_3$

9. A sudden large jump between the values of second and third ionisation energies of an element would be associated with the electronic configuration

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

10. $A \rightarrow A^+ + e, E_1$ and $A^+ \rightarrow A^{2+} + e, E_2$. The energy required to pull out the two electrons are E_1 and E_2 respectively. The correct relationship between two energy would be

(a) $E_1 < E_2$ (b) $E_1 > E_2$
(c) $E_1 = E_2$ (d) $E_1 \neq E_2$

11. The correct order of radii is

(a) N < Be < B (b) $F^- < O^{2-} < N^{3-}$
(c) $Fe^{3+} < Fe^{2+} < Fe^{4+}$ (d) Na < Li < K

12. The electronic configuration of four elements are

(i) $[\text{Xe}]5s^1$ (ii) $[\text{Xe}]4f^{14}, 5d^1, 6s^2$

(iii) $[\text{Ar}]4s^2 4p^5$ (iv) $[\text{Ar}]3d^7, 4s^2$

Select the incorrect statement about these elements.

(a) (i) is a strong reducing agent

(b) (ii) is a d-block element

(c) (iii) has high magnitude of $\Delta_{eg}H$

(d) (iv) exhibits variable oxidation states.

13. Among the statements (I-IV), the correct ones are :

I. Be has smaller atomic radius compared to Mg.

II. Be has higher ionisation enthalpy than Al.

III. Charge/radius ratio of Be is greater than that of Al.

IV. Both Be and Al form mainly covalent compounds. *(JEE Main 2020)*

(a) (II), (III) and (IV) (b) (I), (II) and (IV)

(c) (I), (III) and (IV) (d) (I), (II) and (III)

14. Electronic configuration of four elements A, B, C and D are given below

(A) $1s^2, 2s^2, 2p^6$ (B) $1s^2, 2s^2, 2p^4$

(C) $1s^2, 2s^2, 2p^6, 3s^1$ (D) $1s^2, 2s^2, 2p^5$

Which of the following is the correct order of increasing tendency to gain electron?

(a) $A < C < B < D$ (b) $A < B < C < D$

(c) $D < B < C < A$ (d) $D < A < B < C$

15. The formation of the oxide ion, $O^{2-}(g)$, from oxygen atom requires first an exothermic and then an endothemic step as shown below.

$$O(g) + e^- \longrightarrow O^-(g)\,; \quad \Delta H^\ominus = -141 \text{ kJ mol}^{-1}$$

$$O^-(g) + e^- \longrightarrow O^{2-}(g)\,; \Delta H^\ominus = +780 \text{ kJ mol}^{-1}$$

Thus, process of formation of O^{2-} in gas phase is unfavourable even though O^{2-} is isoelectronic with neon. It is due to the fact that

(a) oxygen is more electronegative

(b) addition of electron in oxygen results in larger size of the ion

(c) electron repulsion outweighs the stability gained by achieving noble gas configuration

(d) O^- ion has comparatively smaller size than oxygen atom

16. In which one of the following pairs, the radius of the second species is greater than that of the first?

(a) Na, Mg (b) O^{2-}, N^{3-}

(c) Li^+, Be^{2+} (d) Ba^{2+}, Sr^{2+}

17. The size of isoelectronic species; F^-, Ne and Na^+ is affected by

(a) nuclear charge (Z)

(b) valence principal quantum number (n)

(c) electron-electron interaction in the outer orbitals

(d) none of the factors because their size is the same

18. The correct order of the first ionisation enthalpies is *(NCERT Exemplar)*

(a) Mn < TI < Zn < Ni (b) Ti < Mn < Zn < Ni

(c) Zn < Ni < Mn <Ti (d) Ti < Mn < Ni < Zn

19. The incorrect statement among the following is

(a) the first ionisation potential of Al is less than the first ionisation potential of Mg

(b) the second ionisation potential of Mg is greater than the second ionisation potential of Na

(c) the first ionisation potential of Na is less than the first ionisation potential of Mg

(d) the third ionisation potential of Mg is greater than that of Al

20. Which of the following sequence correctly represents the decreasing acidic nature of oxides?

(a) $Li_2O > BeO > CO_2 > N_2O_3 > B_2O_3$

(b) $CO_2 > N_2O_3 > B_2O_3 > Li_2O > BeO$

(c) $N_2O_3 > CO_2 > B_2O_3 > BeO > Li_2O$

(d) $CO_2 > BeO > Li_2O > B_2O_3 > N_2O_3$

21. Which is not the correct order for the stated property?

(a) Ba > Sr > Mg, atomic radius

(b) F > O > N, first ionisation enthalpy

(c) Cl > F > I, electron affinity

(d) O > Se > Te, electronegativity

22. The electron affinity values (in kg mol^{-1}) of three halogens X, Y and Z are respectively –349, –333 and –325. Then X, Y and Z respectively, are

(a) F_2, Cl_2 and Br_2 (b) Cl_2, F_2 and Br_2

(c) Cl_2, Br_2 and F_2 (d) Br_2, Cl_2 and F_2

23. Which one of the following arrangements represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species?

(a) Cl < F < S < O (b) O < S < F < Cl

(c) S < O < Cl < F (d) F < Cl < O < S

24. For d-block elements the correct first ionisation potential is of the order

(a) Zn > Fe > Cu > Cr

(b) Sc ≈ Ti < V ≈ Cr

(c) Zn < Cu < Ni < Co

(d) V > Cr > Mn > Fe

25. Which of the following is correct order of increasing size?

(a) $Br^{-} > S^{2-} > Cl^{-} > Na^{+} > Mg^{2+} > Be^{2+}$
(b) $Be^{2+} > Mg^{2+} > Na^{+} > S^{2-} > Cl^{-} > Br^{-}$
(c) $S^{2-} > Cl^{-} > Br^{-} > Na^{+} > Mg^{2+} > Be^{2+}$
(d) $Na^{+} > Mg^{2+} > Be^{2+} > Br^{-} > Cl^{-} > S^{2-}$

26. Which of the following is the correct order of atomic sizes?
(At. no. : Ce = 58, Sn = 50, Yb = 70 and Lu = 71)

(a) Ce > Sn > Yb > Lu
(b) Sn > Yb > Ce > Lu
(c) Sn > Ce > Yb > Lu
(d) Lu > Yb > Sn > Ce

27. The atomic size of Ga and Al are almost same because

(a) of poor shielding effect of electrons in d-orbitals, due to which effective nuclear charge increases in gallium
(b) more shielding effect of d-electrons
(c) poor shielding effect of f-orbital in gallium
(d) high shielding effect of $3s, 3p$ and $3d$ electrons

28. An element X which occurs in the first short period has an outer electronic structure s^2p^1. What are the formula and acid-base character of its oxides?

(a) XO_3, basic
(b) X_2O_3, basic
(c) X_2O_3, amphoteric
(d) XO_2, acidic

29. Which of the following does not represent the correct order of the property indicated?

(a) $Sc^{3+} > Cr^{3+} > Fe^{3+} > Mn^{3+}$ — ionic radii
(b) Sc < Ti < Cr < Mn — density
(c) $Mn^{2+} > Ni^{2+} > Co^{2+} < Fe^{2+}$ — ionic radii
(d) FeO < CaO < MnO < CuO — basic nature

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

(a) 2.18×10^{-18}
(b) 13.12×10^{5}
(c) 3.16×10^{-13}
(d) 2.21×10^{6}

31. Match the correct atomic radius with the element.

	Column I Element		Column II Atomic radius (pm)
A.	Be	(i)	74
B.	C	(ii)	88
C.	O	(iii)	111
D.	B	(iv)	77
(E)	N	(v)	66

Codes

	A	B	C	D	E		A	B	C	D	E
(a)	(iii)	(iv)	(v)	(i)	(ii)	(b)	(iii)	(iv)	(v)	(ii)	(i)
(c)	(iv)	(iii)	(v)	(ii)	(i)	(d)	(v)	(iii)	(iv)	(i)	(ii)

32. Match the correct ionisation enthalpies and electron gain enthalpies of the following elements.

	Elements		$\Delta_i H_i$	$\Delta_i H_2$	$\Delta_{e_g} H$
(A)	Most reactive non-metal	(i)	419	3051	– 45
(B)	Most reactive metal	(ii)	1681	3374	– 328
(C)	Least reactive element	(iii)	738	1451	– 40
(D)	Metal forming binary halide	(iv)	2372	5251	+ 48

Codes

	A	B	C	D		A	B	C	D
(a)	(ii)	(i)	(iv)	(iii)	(b)	(i)	(ii)	(iv)	(iii)
(c)	(i)	(iv)	(iii)	(iv)	(d)	(ii)	(i)	(iii)	(iv)

Numeric Value Questions

33. $x \times 10^{y}$ number of chlorine atoms that can be ionised in the process $Cl \longrightarrow Cl^{+} + e^{-}$, by the energy liberated for the process $Cl + e^{-} \longrightarrow Cl^{-}$ for one Avogadro number of atoms is [IE = 13.0 eV, $\Delta_{e_g} H$ = 3.60 eV]. The value of $x + y$ will be ……… .

34. The covalent radius of silicon is 1.175 Å. Calculate the electronegativity of silicon using Allred-Rochow method ……… .

35. The amount of energy when million atoms of iodine are completely converted into I^{-} ions in the vapour state according to the equation
$I(g) + e^{-} \longrightarrow I^{-}(g)$ is 4.9×10^{-13} J. The electron gain enthalpy of iodine (in eV per atom) is ……… .

36. The Ist IE of Li is 5.4 eV and the electron gain enthalpy of Cl is 3.6 eV. If the following reaction,

$$Li(g) + Cl(g) \longrightarrow Li^{+} + Cl^{-},$$

is carried out at such a low pressure that resulting ions do not combine with each other, the ΔH of the reaction (in kcal mol^{-1}) is ……… kcal mol^{-1}.

37. The effective nuclear charge for $4s$ electrons of Zn will be ……… .

38. The ionisation energy of Li is 520 kJ mol^{-1}. The amount of energy required to convert 140 mg of Li atoms into gaseous Li^{+} ion, is ……… kJ.

39. The electron gain enthalpy of chlorine is 3.7 eV. How much energy in kJ is released when 1 g of chlorine is converted completely to $Cl^{\ominus}$ ion in the gaseous state ……… .

40. The ionisation potential of hydrogen is 13.60 eV. If x kJ energy required to produce one mole of H^{+} ion. Then the value of x …… .

Answers

Round I

1. (a)	**2.** (a)	**3.** (d)	**4.** (b)	**5.** (c)	**6.** (a)	**7.** (a)	**8.** (d)	**9.** (c)	**10.** (c)
11. (c)	**12.** (b)	**13.** (c)	**14.** (d)	**15.** (a)	**16.** (b)	**17.** (b)	**18.** (d)	**19.** (d)	**20.** (a)
21. (b)	**22.** (c)	**23.** (d)	**24.** (b)	**25.** (c)	**26.** (a)	**27.** (c)	**28.** (d)	**29.** (a)	**30.** (c)
31. (b)	**32.** (b)	**33.** (c)	**34.** (b)	**35.** (b)	**36.** (c)	**37.** (a)	**38.** (b)	**39.** (d)	**40.** (c)
41. (a)	**42.** (c)	**43.** (d)	**44.** (a)	**45.** (b)	**46.** (c)	**47.** (c)	**48.** (b)	**49.** (d)	**50.** (b)
51. (d)	**52.** (c)	**53.** (a)	**54.** (c)	**55.** (d)	**56.** (a)	**57.** (a)	**58.** (c)	**59.** (d)	**60.** (d)
61. (c)	**62.** (d)	**63.** (a)	**64.** (d)	**65.** (b)	**66.** (b)	**67.** (d)	**68.** (a)	**69.** (d)	**70.** (b)
71. (d)	**72.** (a)	**73.** (a)							

Round II

1. (c)	**2.** (a)	**3.** (d)	**4.** (b)	**5.** (d)	**6.** (a)	**7.** (d)	**8.** (b)	**9.** (a)	**10.** (a)
11. (b)	**12.** (b)	**13.** (b)	**14.** (a)	**15.** (c)	**16.** (b)	**17.** (a)	**18.** (d)	**19.** (b)	**20.** (c)
21. (b)	**22.** (b)	**23.** (b)	**24.** (a)	**25.** (a)	**26.** (c)	**27.** (a)	**28.** (c)	**29.** (a)	**30.** (b)
31. (b)	**32.** (a)	**33.** (24.667)	**34.** (1.73)	**35.** (–3.06)	**36.** (41.508)	**37.** (4.35)	**38.** (10.4)	**39.** (10.04)	**40.** (1309.68)

Solutions

Round I

2. Mendeleef failed to assign positions of the isotopes on the basis of atomic mass, according to his periodic law.

4. Statement (b) is incorrect. The correct statement (b) is; the d-block has 10 columns, because a maximum of 10 electrons can occupy all the orbitals in a d-subshell.

All other given statements are correct.

7. Na—Cl Both belongs to III period.

8. Same group elements have same number of valence electrons therefore, have similar physical and chemical properties.

9. Periodic table reflects trends in physical and chemical properties of the elements.

10. Period number = Maximum n of any element of the respective period

(where , n = principal quantum number)

11. $(n-1)\,d^2\,ns^2$ for $n = 4$

$n = 4$ means the element belongs to fourth period. Since, last electron enters in d-orbital, the given element belongs to d-block.

For d-block elements, group number = number of d-electrons + number of ns electrons = 2 + 2 = 4.

Hence, the element belongs to 4th group.

The complete electronic configuration of the element is as follows,

$$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^2, 4s^2$$

Thus, the element is titanium.

12. $(n-2)\,f^7\,(n-1)\,d^1\,ns^2$ for $n = 6$

$n = 6$ means, the element belongs to sixth period.

Since, last electron enters in f-orbital, the given element belongs to f-block and all f-block elements are the members of third group. Hence, the element belongs to third group.

The complete configuration of the element is as follows

$$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 4d^{10}, 5s^2, 5p^6, 4f^7, 5d^1, 6s^2$$

Thus, the element is gadolinium (Gd).

13. In case of transition elements (or any elements), the order of filling of electrons in various orbitals is $3p < 4s < 3d$.

Thus, $3d$ orbital is filled when $4s$ orbital gets completely filled.

14. ns^2np^4 for $n = 3$

$n = 3$ means element belongs to third period. Since, last electron enters in the p-orbital, it belongs to p-block.

For p-block elements, the group number = 10 + valence shell electrons = 10 + (2 + 4) = 16

Hence, the given element belongs to 16th group.

The complete electronic configuration of the element is as follows,

$$1s^2, 2s^2, 2p^6, 3s^2, 3p^4$$

Thus, the given element is sulphur.

16. $X = {}_{33}\text{As} \rightarrow$ Metalloid $\quad Y = {}_{53}\text{I} \rightarrow$ Non-metal

$Z = {}_{83}\text{Bi} \rightarrow$ Metal

17. General configuration for 17th group elements is ns^2np^5. In the third period, the principal quantum number for valence shell is three, so the electronic configuration of valence shell for the given element is $3s^2, 3p^5$. Third period starts from atomic number, $Z = 11$ and end at $Z = 18$. Hence, the atomic number of the given element is $10 + 7 = 17$.

20. Since, effective nuclear charge increases from left to right in a period thus, atomic radii decreases.

21. All the given elements belongs to 2nd period of the periodic table.

The atomic radius decreases from left to right along a period due to increased in effective nuclear force.

24. Cation has the small size than parent atom and anion has larger size than parent atom.

26. All the given species have same number of electrons ($10e^-$). Therefore, all are isoelectronic species.

27. The ionic radii of isoelectronic species decreases with increase in atomic number (as magnitude of the nuclear charge increases with increase in atomic number). Therefore, their ionic radii increase in the order

Isoelectronic ions =	Al^{3+} <	Mg^{2+} <	Na^+ <	F^- <	O^{2-} <	N^{3-}
Atomic number, Z =	13	12	11	9	8	7

28. As the nuclear charge per electron is maximum in P^{5+}. therefore, its size is smallest among the givne ions.

31. The ionic radii order is

$$Na^+ > Mg^{2+} > Al^{3+}$$

35.

Period	Group 15	Group 16
2nd period	N	O
3rd period	P	S

Ionisation enthalpy of nitrogen ($_7N = 1s^2, 2s^2, 2p^3$) is greater than oxygen ($_8O = 1s^2, 2s^2, 2p^4$) due to presence of extra stable exactly half-filled $2p$-orbitals.

Similarly, ionisation enthalpy of phosphorus ($_{15}P = 1s^2, 2s^2, 2p^6, 3s^2, 3p^3$) is greater than sulphur ($_{16}S = 1s^2, 2s^2, 2p^6, 3s^2, 3p^4$).

On moving down the group, ionisation enthalpy decreases with increasing atomic size. So, the increasing order of first ionisation enthalpy is

$S < P < O < N \rightarrow$ First ionisation enthalpy increases.

36.

Period	Group 13	Group 14
2nd period	Boron	Carbon
3rd period	Aluminium	Silicon

Ionisation enthalpy increases along a period (as we move from left to right in a period) with decrease in atomic size but it decreases down the group with increase in atomic size. Hence, carbon has the highest first ionisation enthalpy.

41. Nitrogen has high ionisation potential than carbon and oxygen because its outermost orbit is half-filled. So, the order is $C < N > O$.

42. $Na \rightarrow [Ne]\, 3s^1$ IE_1 is very low but IE_2 is very high due to stable noble gas configuration of Na^+.

$Mg \rightarrow [Ne]\, 3s^2$ IE_1 and $IE_2 \rightarrow$ Low

IE_3 is very high.

43. S^{2-} has the largest size and hence, has the lowest ionisation energy.

44. $_6C \longrightarrow 1s^2, 2s^2, 2p^2$

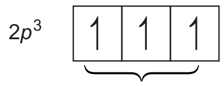
$_5B \longrightarrow 1s^2, 2s^2, 2p^1$

In first case IE_1 of $C > IE_1$ of B. Since, carbon is smaller than B in size. But $IE_2(B) > IE_2(C)$ because electron are paired as well as present in inner s-orbital whereas, for carbon it will be still in $2p$-orbital and in unpaired state.

45. Ionisation potential decreases down the group.

46. The electronic configuration of nitrogen is

$_7N = 1s^2, 2s^2, 2p^3$

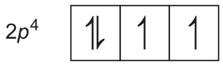


Half-filled p-orbital

Due to the presence of half-filled p-orbitals, a large amount of energy is required to remove an electron from nitrogen. Hence, first ionisation energy of nitrogen is greater than that of oxygen.

The electronic configuration of oxygen is

$$_8O = 1s^2, 2p^2, 2p^4$$



The other reason for the greater IP of nitrogen is that in oxygen, there is a greater inter-electronic repulsions between the electrons present in the same p-orbital which counter balance the increase in effective nuclear charge from nitrogen to oxygen.

47. Be has higher $\Delta_i H$ (ionisation enthalpy) than boron. In both the cases, the electron to be removed from the same principal shell. In $_4Be = (1s^2, 2s^2)$, it is $2s$-electron while in boron $_5B = (1s^2, 2s^2, 2p^1)$ it is $2p$-electron. The penetration of a $2s$-electron to the nucleus is more than that of a $2p$-electron. It means $2s$-electrons are more strongly attracted by the nucleus than $2p$-electrons.

Therefore, higher amount of energy is required to remove a $2s$-electron than a $2p$-electron. Hence, Be has higher $\Delta_i H$ than B.

48. O has lower $\Delta_i H$ than N and F

$$_7N = 1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1;$$

$$_8O = 1s^2, 2s^2, 2p_x^2, 2p_y^1, 2p_z^1;$$

$$_9F = 1s^2, 2s^2, 2p_x^2, 2p_y^2, 2p_z^1$$

Across a period, ionisation enthalpy increases as we move from left to right due to decrease in atomic size. But $\Delta_i H$ of nitrogen is greater than oxygen. It is becasue of the more stable electronic configuration (exactly half filled orbitals are more stable) of nitrogen, so it is difficult to remove an electron from nitrogen than from oxygen. That's why, oxygen has lower ionisation enthalpy than nitrogen and fluorine.

55. The electron affinities of some of the elements of second period (i.e. N, O, F, etc.) are however, lower than the corresponding elements (i.e. P, S, Cl, etc.) of the third period. This is due to the reason that the elements of second period have the smallest atomic size amongst the elements in their respective groups.

As a result, there are considerable electron-electron repulsion within the atom itself and hence, the additional electron is not accepted with the same ease as is the case with the remaining elements in the same group.

57. Fluorine being most electronegative atom, has a high tendency to gain electron. Thus, it readily forms anions.

58. (c) $O \rightarrow O^{2-}$

$$O + e^- \rightarrow O^-; \; EA_1 = -\text{ve}$$
$$O^- + e^- \rightarrow O^{2-}; \; EA_2 = +\text{ ve}$$
$$\therefore \qquad EA_2 > EA_1$$

It become difficult for an atom to take up another electron since already added electron repels existing electrons.

Thus, second electron affinity of oxygen is endothermic that means it require energy and greater than first electron affinity, which is exothermic.

59. Electron gain enthalpy of F is less negative than that of Cl because when an electron is added to F, the added electron goes to the smaller $n = 2$ quantum level and suffers repulsion from other electrons present in this level.

In case of Cl, the added electron goes to the larger $n = 3$ quantum level and suffers much less repulsion from other electrons.

60. If the EN difference is 1.9, then bond is 50% ionic. The difference in electronegativity is 2.8, therefore, percentage ionic character due to EN difference of 2.8 is $\frac{2.8}{1.9} \times 50 = 73.6\%$

61. $Z = 2, 8, 8, 1$. Because it would donate electron more easily.

62. On moving down the group, electronegativity decreases because atomic size increases. Fr has the largest size, therefore it is least electronegative.

64. Li–Mg, B–Si, Be–Al show diagonal relationship but Li and Na do not show diagonal relationship as both belongs to same group and not placed diagonally.

70. Fluorine is more reactive than chlorine, bromine and iodine.

71. Oxidising power decreases in a group.

72. In a period, from left to right basic character of oxides decreases, thus Na_2O is most basic.

Round II

1. If Aufbau rule is not followed then 19th electron in K enters in $3d$ sub-shell, not in $4s$. In that case it is placed in d-block.

2. Atomic radius of F is expressed in terms of covalent radius while atomic radius of neon is usually expressed in terms of van der Waals' radius. van der Waals' radius of an element is always larger than the covalent radius. Therefore, atomic radius of F is smaller than atomic radius of Ne (F = 72 pm, Ne = 160 pm).

3. In lanthanide series, size decreases from La to Lu due to lanthanide contraction. Thus, La is largest among lanthanides.

4. $Cr(24) = [Ar]_{18} \; 3d^5, 4s^1$

Chromium after losing one electron, gain stable configuration due to the presence of half-filled d-orbitals, therefore its second ionisation enthalpy is highest.

5. Down the group, size of atom increases. Therefore, bond length of LiF is less than that of NaF.

6. For both group I and VII elements, atomic radii increases on moving down the group.

7. Ionic radius $\propto \frac{1}{Z_{\text{eff}}}$

Since, P^{5+} has higher Z_{eff} as compared to P^{3+}, it has smaller ionic radii.

8. Among the given compounds, oxidation state of Cr is highest in K_2CrO_4, i.e. + 6. Therefore, Cr has the minimum radius in K_2CrO_4.

9. $1s^2 2s^2 2p^6 3s^2$ configuration after losing two electrons acquire noble gas configuration, thus a large amount of energy is required to remove a third electron, i.e. there is a large jump between the values of second and third IEs.

10. The correct relation is $E_1 < E_2$, there is a strong electrostative attractive force between the unipositive cation and electron or we can say that nuclear charge per electron increases. Due to this reason the removal of second electron is difficult and require a large amount of energy.

11. The given ions in the option (b) are isoelectronic species. The ionic radii of isoelectronic species decreases with increase in atomic number (as magnitude of the nuclear charge increases with increase in atomic number).

Therefore, their ionic radii increase in the order

Isoelectronic ions = $F^- < O^2 < N^{3-}$

Atomic number $Z = 9 \quad 8 \quad 7$

12. (i) Since, in (i) the last electron enters in *s*-orbital, it is a *s*-block element and *s*-block elements are good reducing agents.

(ii) Since, in (ii) the last electron enters in *f*-block, it is a *f*-block element (element of 4*f* series).

(iii) Electronic configuration of (iii) reveals that it contains 7 electrons in its outer shell, so it is a halogen and halogens have high magnitude of $\Delta e_g H$.

(iv) The last electron enters in *d*-orbital, so it is a *d*-block element and variable oxidation states is an important feature of *d*-block elements.

13. Statements I, II and IV are correct whereas statement III is incorrect. Corrected III statement is : charge/radius ratio of Be and Al is same because of diagonal relationship.

14. Element with *A*, *B* and *D* configuration belong to second period. *D* being a halogen (i.e. because of its smaller size in the period) has the highest ΔH_e among the given followed by *B*. *A* is an inert gas, so ΔH_e is zero or negative. Thus, the correct order of ΔH_e is $A < C < B < D$.

15. Although, O^{2-} has noble gas configuration but its formation is unfavourable because here, the electronic repulsion outweigh the stability gained by achieving inert gas configuration.

16. O^{2-} and N^{3-} both are isoelectronic but differ in the charge possessed by them. As the negative charge increases, the electrons are held less and less tightly by the nucleus, therefore, radius increases from O^{2-} to N^{3-}.

17. The size of isoelectronic species; F^-, Ne and Na^+ is affected by nuclear charge (Z). With increase in nuclear charge (atomic number), the size of the isoelectronic species decreases.

18. The 3*d*-transition series is

	Sc	$\boxed{\text{Ti}}$	V	Cr	$\boxed{\text{Mn}}$	Fe	Co	$\boxed{\text{Ni}}$	Cu	$\boxed{\text{Zn}}$
Atomic number	21	22	23	24	25	26	27	28	29	30
Outermost electronic configuration		← $3d^24s^2$			← $3d^54s^2$			← $3d^84s^2$		← $3d^{10}4s^2$

In 1st ionisation, one electron will be removed from $4s^2$ subshell/orbital.

With increase in atomic number (Z), i.e. with increase in number of protons in the nucleus, effective nuclear charge (Z^*) also increases from Sc to Zn.

$$IE \propto Z^*$$

So, IE order of the given elements will be

$$Ti < Mn < Ni < Zn$$

19. IE (II) of Na is higher than that of Mg because in case of Na, the second e^- has to be removed from the noble gas configuration while in case of Mg removal of second e^- gives a noble gas. Mg has high first ionisation potential than Na because of its stable ns^2 configuration.

20. On passing from left to right in a period, acidic character of the normal oxides of the element goes on increasing with increase in electronegativity. Thus, the correct order is

$$N_2O_3 > CO_2 > B_2O_3 > BeO > Li_2O$$

21. On moving along a period, ionisation enthalpy increases. Thus, the expected order of ionisation enthalpy should be as follows

$$F > O > N$$

But N has half-filled configuration, therefore it is more stable than O and has higher ionisation enthalpy.

22. Generally electron affinity decreases on moving down a group but Cl_2 has higher electron affinity than F_2 due to its larger size as compared to F.

So, the correct order of electron affinity will be

$$Cl_2 > F_2 > Br_2$$

23. The correct order of electron gain enthalpy is

$$O < S < F < Cl$$

Element	O	S	F	Cl
Electron gain enthalpy (in eV)	1.48	2.07	3.45	3.61

24. The ionisation energy increases with increase in atomic number in a period. The trend is irregular among *d*-block elements.

Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
IE° (kJ/mol)	631	656	650	652	717	762	758	736	745	905

Hence, correct order is Zn > Fe > Cu > Cr.

25. $\text{Size} \propto \dfrac{1}{Z_{\text{eff}} \text{ or nuclear charge}}$

Thus, the correct order of size is

$$Br^- > S^{2-} > Cl^- > Na^+ > Mg^{2+} > Be^{2+}$$

26. $Ce_{58} = [Xe]\, 4f^1\, 5d^1\, 6s^2$

$Sn_{50} = [Kr]\, 4d^{10}\, 5s^2\, 5p^2$

$Yb_{70} = [Xe]\, 4f^{14}\, 6s^2$

$Lu_{71} = [Xe]\, 4f^{14}\, 5d^1\, 6s^2$

Among lanthanoids size reduces from La to Lu thus, the correct order of size is Ce > Yb > Lu. Size of Sn is greater than Ce.

27. $Al_{13} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^1$

$Ga_{31} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^1$

Electronic configuration reveals that extra 3*d* electron are present in case of Ga which have lower shielding tendency and hence, Z_{eff} is more. Because of this atomic size decreases and becomes almost equal to Al.

28. The outer electronic configuration = s^2p^1

Thus, valency = 2 + 1 = 3

Therefore, the formula of the oxide is X_2O_3. Since, it is an oxide of III group element, its nature is amphoteric.

29. Option (a) does not represent the correct order of the property indicated. The correct order is

$$Cr^{3+} > Mn^{3+} > Fe^{3+} > Sc^{3+}.$$

30. Ionisation energy is the amount of energy required to remove the electron from the ground state (E_1) to infinity (E_∞).

$$E_1 = -2.18 \times 10^{-18} \text{ J}$$
$$E_\infty = 0 \text{ [Energy of an electron at infinity = 0.]}$$
$$\Delta E = E_\infty - E_1$$
$$= 0 - (-2.18 \times 10^{-18} \text{ J})$$
$$= 2.18 \times 10^{-18} \text{ J}$$

Ionisation enthalpy per hydrogen atom = 2.18×10^{-18} J

∴ Ionisation enthalpy per mole of hydrogen atoms

$$= 2.18 \times 10^{-18} \times 6.022 \times 10^{23} \text{ J mol}^{-1}$$
$$= 13.12 \times 10^{5} \text{ J mol}^{-1}$$

31. All the given elements are of same period and along a period atomic radii decreases because effective nuclear charge increases.

Thus, the order of atomic radii is

$$O < N < C < B < Be$$

or Be = 111, O = 66, C = 77, B = 88, N = 74.

32. (ii) is the most reactive non-metal because of its higher IE and $\Delta_{eg}H$.

(i) is the most reactive metal because of its lowest Ist IE. [There is a larger difference in IE_1 and IE_2 in case of (i)]

(iv) is a least reactive element because of its high IEs and positive $\Delta_{eg}H$.

(iii) is a bivalent metal because of less difference in ΔH_1 and ΔH_2.

33. Let n Cl-atoms are ionised.

Thus, $6.02 \times 10^{23} \times \Delta_{eg}H = n \times IE$

$$n = \frac{6.02 \times 10^{23} \times 3.6}{13.0}$$
$$= 1.667 \times 10^{23}$$

The value of $x = 1.667$ and $y = 23$. Thus, $x + y = 24.667$.

34. Allred-Rochow equation is

$$x = 0.359 \times \frac{Z_{eff}}{r^2} + 0.744 \qquad \ldots(i)$$

Z_{eff} is calculated on the basis of Slater's rules taking all the electrons.

Electronic configuration of Si is $1s^2, 2s^2, 2p^6, 3s^2, 3p^2$

$$Z_{eff} = 14 - (0.35 \times 4 + 0.85 \times 8 + 2 \times 1) = 3.80$$

On putting the values of Z_{eff} and r as 1.175Å in Eq(i) we can calculate x

$$x = 1.73.$$

35. Electron gain enthalpy per mol

$$= -\frac{4.9 \times 10^{-13} \times 6.023 \times 10^{23}}{10^6}$$
$$= -29.5 \times 10^4 \text{ J} = -295 \text{ kJ}$$
$$= -\frac{295}{96.49} = -3.06$$

36. Given, $Li(g) \longrightarrow Li^+(g) + e^-; \quad \Delta E_1 = 5.4 \text{ eV}$

$$Cl(g) + e^- \longrightarrow Cl^-(g); \qquad \Delta E_2 = -3.6 \text{ eV}$$
$$\Delta H = 5.4 - 3.6 = 1.8 \text{ eV}$$
$$= 1.8 \times 23.06 \text{ kcal mol}^{-1}$$
$$= 41.508 \text{ kcal mol}^{-1}$$

37. The electronic configuration of Zn is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$.

$$Z_{eff} = Z - r$$

$$r = \left[\begin{array}{l} 0.85 \times \text{Number of electron in } (n-1)\text{th.} \\ \text{Shell} + 1.00 \times \text{Total number of electrons} \\ \text{in the inner shell} \end{array}\right]$$

$$r = 0.35 \times 1 + 18 \times 0.85 + 10 \times 1 = 25.65$$
$$Z_{eff} = 30 - 25.65 = 4.35$$

38. Number of moles of Li $= \dfrac{\text{mass}}{\text{atomic mass}}$

$$= \frac{140 \times 10^{-3}}{7} = 2 \times 10^{-2} \text{ mol}$$

∵ For the conversion of 1 mole atoms of Li into Li^+, the energy required = 520 kJ mol^{-1}

∴ For the conversion of 2×10^{-2} mole atoms of Li into Li^+ ions, the energy required will be $520 \times 2 \times 10^{-2}$

$$= 10.4 \text{ kJ}$$

39. The electron gain enthalphy in kJ is represented as

$$Cl(g) + e^- \longrightarrow Cl^-(g) + 3.7 \text{ eV}$$

Now, 1 eV atom^{-1} = 96.3 kJ mol^{-1}

∴ 3.7 eV = 3.7 eV × 96.3 kJ mol^{-1}

Energy released when 35.5 g of chlorine is completely converted to Cl^- ion = 3.7×96.3 kJ.

Energy released when 1 g of chlorine is completely converted to $Cl^-(g)$ ion $= \dfrac{3.7 \times 96.3}{35.5} = 10.04$ kJ

40. The ionisation potential may be represented as

$$H(g) + 13.60 \text{ eV} \rightarrow H^{\oplus}(g) + e^-$$

We know 1 eV = 96.3 kJ mol^{-1}

13.60 eV = 96.3 × 13.60 = 1309.86 kJ

Thus, energy per mol = 1309.68 kJ