CHAPTER 3

Classification of Elements and Periodicity in Properties

We know by now that the elements are the basic units of all types of matter. In 1800, only 31 elements were known. By 1865, the number of identified elements had more than doubled to 63. At present 114 elements are known. Of them, the recently discovered elements are man-made. Efforts to synthesize new elements are continuing. With such a large number of elements it is very difficult to study individually the chemistry of all these elements and their innumerable compounds individually. To ease out this problem, scientists searched for a systematic way to organise their knowledge by classifying the elements. Not only that it would rationalize known chemical facts about elements, but even predict new ones for undertaking further study. Earlier altempts of classifications included Dobereinerís triads, Newland octaves and Lothar Meyerís atomic volume curves, etc.

The periodic table displays all chemical elements systematically in order of increasing atomic number *i.e.* the number of protons in the nucleus.

Mendeleevís periodic Table

Mendeleev arranged the elements on the basis of law which stated, the physical and chemical properties of elements are periodic functions of their atomic weights. If the elements are arranged in the order of their increasing atomic weights, after a regular interval elements with similar properties are repeated. The table is divided into nine vertical columns called groups and seven horizontal rows called periods.

Characteristics of periods

(a) First period is called shortest period and contains only two elements. Second and third periods are called short periods containing eight elements each. Fourth and fifth periods are long periods containing eighteen elements each. Sixth period is longest period with thirty-two elements. Seventh period is an incomplete period containing nineteen elements. Numbers 2, 8, 8, 18, 18, 32 are called magic numbers.

(b) Lanthanide and actinide series containing 14 elements each are placed separately under the main periodic table. These are related to sixth and seventh periods of IIIrd group respectively.

(c) Elements of third period from sodium (Na) to chlorine (Cl) are called representative or typical elements.

(d) Valency of an element in a period increases from 1 to 7 with respect to oxygen.

Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
1	2	3	4	5	6	7

(e) From left to right in a period generally

(i) Atomic weight, effective nuclear charge, ionisation potential, electronegativity and electron affinity of an element increase.

(ii) Atomic radius, electropositive character and metallic character of an element decrease.

(f) Diagonal relationship : Elements of second period Li, Be and B resemble closely with the elements Mg, Al and Si of third period in the next higher group.

$$\begin{tikzcd}
\textbf{Second period} & \text{Li} \arrow[dr] & \text{Be} \arrow[dr] & \text{B} \arrow[dr] & \text{C} \\
\textbf{Third period} & \text{Na} & \text{Mg} & \text{Al} & \text{Si}
\end{tikzcd}$$

(g) Elements of second period are called bridge elements.

Characteristic of groups

(a) Mendeleevís periodic table contains nine groups. These are represented by Roman numerals I, II, III, IV, V, VI, VII, VIII and zero. Groups I to VII are divided into subgroups A and B, group VIII consists of three sets, each one containing three elements.

(b) Inert gases are present in zero group. These were not discovered till that time.

(c) The valency of an element in a group is equal to the group number.

(d) There is no resemblance in the elements of sub-groups A and B of same group, except valency.

(e) The elements of the groups which resemble with typical elements are called normal elements. For example - IA, IIA, IIIA, IVA, VA, VIA, VIIA group elements are normal elements.

(f) Those elements of the groups which do not resemble with typical elements are called transition elements. For example- IB, IIB, IIIB, IVB, VB, VIB, VIIB and VIII group elements are transition elements.

(g) Hydrogen is placed in both IA and VIIA groups.

(h) In a group, from top to bottom in general,

(i) Atomic weight, atomic size, electropositive character and metallic character of an element increase.

(ii) Ionisation potential, electron affinity and electronegativity of an element decrease.

Moseleyís periodic table

Moseley (1909) studied the frequency of X-rays produced by the bombardment of a strong beam of electrons on a metal target. He found that the square root of the frequency of X-rays ($\sqrt{\upsilon}$) is directly proportional to the number of nuclear charge (Z) of metal. $\sqrt{\upsilon}$ = a (Z ñ b) where a and b are constants. Nuclear charge of metal is equal to the atomic number. So Moseley related the properties of elements with their atomic number and gave the new periodic law.

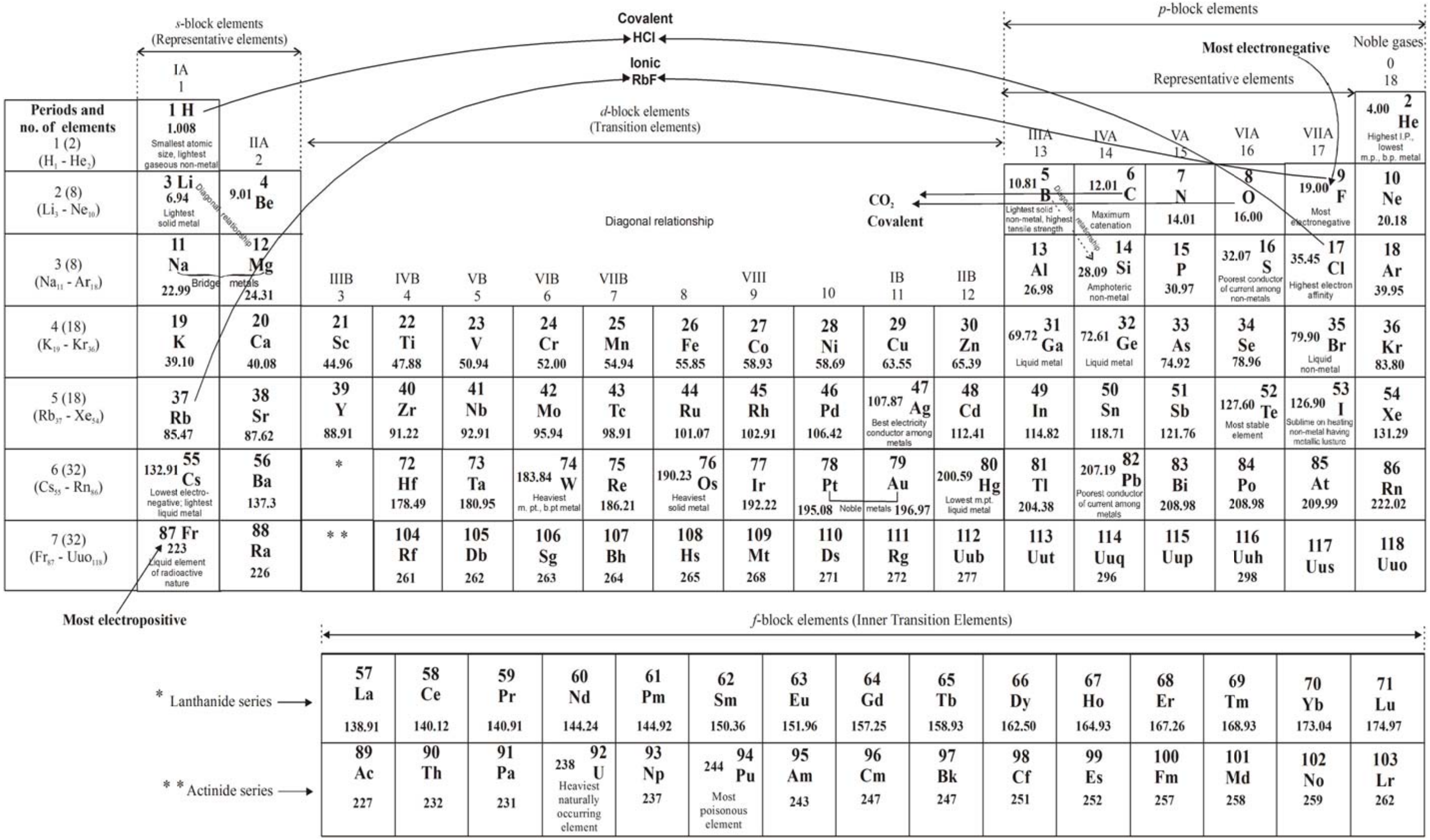
According to him, physical and chemical properties of elements are the periodic functions of their atomic number. If the elements are arranged in order of their increasing atomic number, after a regular interval, element with similar properties are repeated.

Bohrís periodic table

With better understanding of the role of electrons in the properties of elements and the development of the nature of the electrons in atoms, a better understanding of the periodic properties of elements or the periodic table was possible. The long form of periodic table is also known as Bohrís periodic table.

The long form of periodic table offers the following advantages over the Mendeleevís classification.

(i) Sub-group A and sub-group B elements are placed separately. There is clear demarcation between metals and non-metals.



Periods and no. of elements	IA 1	IIA 2	IIIB 3	IVB 4	VB 5	VIB 6	VIIB 7	8	VIII 9	10	IB 11	IIB 12	IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	0 18
1 (2) (H_1 - He_2)	1 H 1.008 Smallest atomic size, lightest gaseous non-metal																	2 He 4.00 Highest I.P., lowest m.p., b.p. metal
2 (8) (Li_3 - Ne_{10})	3 Li 6.94 Lightest solid metal	4 Be 9.01											5 B 10.81 Lightest solid non-metal, highest tansile strength	6 C 12.01 Maximum catenation	7 N 14.01	8 O 16.00	9 F 19.00 Most electronegative	10 Ne 20.18
3 (8) (Na_{11} - Ar_{18})	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09 Amphoteric non-metal	15 P 30.97	16 S 32.07 Poorest conductor of current among non-metals	17 Cl 35.45 Highest electron affinity	18 Ar 39.95
4 (18) (K_{19} - Kr_{36})	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72 Liquid metal	32 Ge 72.61 Liquid metal	33 As 74.92	34 Se 78.96	35 Br 79.90 Liquid non-metal	36 Kr 83.80
5 (18) (Rb_{37} - Xe_{54})	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98.91	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87 Best electricity conductor among metals	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60 Most stable element	53 I 126.90 Sublime on heating non-metal having metallic lusture	54 Xe 131.29
6 (32) (Cs_{55} - Rn_{86})	55 Cs 132.91 Lowest electro-negative; lightest liquid metal	56 Ba 137.3	*	72 Hf 178.49	73 Ta 180.95	74 W 183.84 Heaviest m. pt., b.pt metal	75 Re 186.21	76 Os 190.23 Heaviest solid metal	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59 Lowest m.pt. liquid metal	81 Tl 204.38	82 Pb 207.19 Poorest conductor of current among metals	83 Bi 208.98	84 Po 208.98	85 At 209.99	86 Rn 222.02
7 (32) (Fr_{87} - Uuo_{118})	87 Fr 223 Liquid element of radioactive nature	88 Ra 226	* *	104 Rf 261	105 Db 262	106 Sg 263	107 Bh 264	108 Hs 265	109 Mt 268	110 Ds 271	111 Rg 272	112 Uub 277	113 Uut	114 Uuq 296	115 Uup	116 Uuh 298	117 Uus	118 Uuo

* Lanthanide series →	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm 144.92	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
* * Actinide series →	89 Ac 227	90 Th 232	91 Pa 231	92 U 238 Heaviest naturally occurring element	93 Np 237	94 Pu 244 Most poisonous element	95 Am 243	96 Cm 247	97 Bk 247	98 Cf 251	99 Es 252	100 Fm 257	101 Md 258	102 No 259	103 Lr 262

(ii) The nine elements of group VIII have been placed in separate groups corresponding to d^6, d^7 and d^8 configurations.

(iii) Fourteen lanthanides are not placed together, but are assigned a separate group for each lanthanon in the *f*-block of elements.

(iv) The uniform trivalent state of the lanthanides can be explained due to the availability of only three electrons in the outer, high energy levels, the differentiating electrons going to the inner chemically inert orbitals.

(v) Uniform bivalence for the transition elements is due to the presence of outer ns^2 electrons, which makes them electropositive in nature.

(vi) The change from a highly electronegative to electropositive character through inert gas structure has been explained on the basis of the long form of periodic table.

Nomenclature of Elements with Atomic number more than 100

Enrico fermi and coworkers (1934), first prepared the elements beyond uranium. They irradiated uranium with slow neutrons and found several radioactive products. These were called transuranic or transuranium elements. The elements with, Z = 104 ñ 112, 114 and 116 have been reported recently and are called transactinides or superheavy elements. These are synthetic, *i.e.*, man-made elements. Recently out of competitive spirit dispute have arisen over the names of these newly discovered elements. To overcome this problem, scientists proposed a systematic nomenclature based on the atomic number of the element. Afterwards in 1997, IUPAC gave the approved official names for elements with atomic numbers 104 to 109. A systematic nomenclature based on the atomic number of the element having atomic numbers > 100 is given as follows:

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

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

- In few cases the names are shortened. *e.g*, ëbiiumí and ëtriiumí are shortened to ëbiumí and ëtriumí.
- The symbol for the element is made up from the first letters from the roots which make up the names. The mixture of latin and Greek roots has been chosen to ensure that the symbols are all different.

Illustration - 1 : What would be the IUPAC name and symbol of the element with atomic number 120?

Soln.: The roots for 1, 2 and 0 are un, bi and nil respectively.

∴ Name of element : Unbinilium

Symbol : Ubn

Illustration - 2 : The elements $Z = 107$ and $Z = 109$ have been made recently; element $Z = 108$ has not yet been made. Indicate the group in which you will place the above elements.

Soln.: The electronic configurations of these elements are :

$Z = 107$ $[Rn]5f^{14}\,6d^{5}\,7s^{2}$
$Z = 108$ $[Rn]5f^{14}\,6d^{6}\,7s^{2}$
$Z = 109$ $[Rn]5f^{14}\,6d^{7}\,7s^{2}$

These elements will be placed in *d*-block in groups 7th, 8th and 9th respectively.

Illustration - 3 :

(a) Elements *A, B, C* and *D* have atomic numbers 12, 19, 29 and 36 respectively. On the basis of electronic configuration, write to which group of the periodic table each element belongs.

(b) Predict the blocks to which these elements can be classified. Also predict their periods and groups.

(c) Which of these are representative elements?

Soln.: (a)

Element	At. No.	Electronic configuration	Group
A	12	$1s^2\,2s^2\,2p^6\,3s^2$	2
B	19	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1$	1
C	29	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^1$	11
D	36	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6$	18

[According to old system, these elements belong to groups : IIA (*A*), IA (*B*), IB (*C*), VIIA or zero (*D*)]

(b) Prediction of blocks

(i) *A* receives the last electron in 3*s*-subshell, therefore, it belongs to *s*-block.
Period = 3rd
Group = No. of valence electrons = 2

(ii) *B* receives the last electron in 4*s*-orbital, therefore, it belongs to *s*-block.
Period = 4th
Group = No. of valence electrons = 1.

(iii) *C* receives the last electron in 3*d*-orbital, therefore, it belongs to *d*-block.
Period = 4th
Group = No. of electrons in *ns* and (*n* ñ 1) *d* subshells = 11

(iv) *D* receives the last electron in the 4*p*-orbital, therefore, it belongs to *p*-block.
Period = 4th
Group = 10 + valence electrons = 10 + 8 = 18

(c) The elements *A* and *B* are representative elements.

LONG FORM OF THE PERIODIC TABLE

The table is based on the modern periodic law. It is an improved form of the Mendeleevís periodic table. This table has the following features:

- All the elements have been arranged in the increasing order of atomic numbers.
- Elements with similar electronic configurations have similar properties and hence have been placed together at one place. Elements with different electronic configurations have different properties and hence have been placed at different places in the periodic table. The periodic table consists of
    - Seven horizontal rows called periods.
    - Eighteen vertical columns called groups or families.
    - Four blocks.

Periods

There are seven horizontal rows called periods. The 112 known elements have been accommodated in these periods according to the following scheme:

- **1st period :** The first main energy shell (K-shell) is completed. As the maximum capacity of K-shell is of 2 electrons, it consists of only two elements, hydrogen ($1s^1$) and helium ($1s^2$).
- **2nd period :** The second main energy shell is completed, *i.e.*, 2*s* and 2*p* are completed. It includes eight elements from Li ($2s^1$) to Ne ($2s^2\ 2p^6$).
- In **third period** corresponding to $n = 3$, there are nine orbitals: one 3*s*, three 3*p* and five 3*d*. However, we know from energy level diagram for multielectron atoms that 3*d*-orbitals are higher in energy than 4*s*-orbitals. Consequently, 3*d*-orbitals are filled after filling 4*s*-orbital. Hence, this period involves the filling of only four orbitals (3*s* and 3*p*) and contains **eight elements** from sodium ($Z = 11$) to argon ($Z = 18$).
- The **fourth period** corresponding to $n = 4$ involves the filling of one 4*s*-and three 4*p*-orbitals (4*d* and 4*f* orbitals are higher in energy than 5*s*-orbitals and are filled later). In between 4*s*-and 4*p*-orbitals, five 3*d*-orbitals are filled which have energies in between these orbitals. Thus, in all nine orbitals are to be filled and, therefore, there are **eighteen elements** in fourth period from potassium ($Z = 19$) to krypton ($Z = 36$).
- The **fifth period,** like the fourth period also consists of 18 elements. It begins with rubidium ($Z = 37$) with filling of 5*s*-orbital and ends at xenon ($Z = 54$) with the filling up of the 5*p*-orbitals.
- The **sixth period** contains 32 elements ($Z = 55$ to 86) and successive electrons enter into 6*s*, 4*f*, 5*d*- and 6*p*-orbitals, in that order. It starts with cesium ($Z = 55$) and ends at radon ($Z = 86$).
- The **seventh period**, though expected to have 32 elements is, incomplete and contains only 19 elements at present. This also provides a theoretical justification for periodicity occurring at regular intervals of 2, 8, 8, 18, 18 and 32. These numbers i.e., 2, 8, 18 and 32 are also called **magic numbers**.
- The first three periods containing 2, 8 and 8 elements respectively are called **short periods**, the next three periods containing 18, 18 and 32 elements respectively are called **long periods.**

Groups

- Vertical columns in the periodic table starting from top to bottom are called groups or families. There are 18 groups in the periodic table which are numbered from 1ñ18.
- On the basis of electronic configuration, the elements may be divided into four groups:

(i) *s*-block elements

(a) These are present in the left part of the periodic table.
(b) These are 1 and 2 group elements.
(c) These are metals.
(d) In these elements last electron fills in the *s*-orbital.
(e) Electronic configuration of valence shell is $ns^{1ñ2}$ ($n = 1$ to 7).

(ii) *p*-block elements

(a) These are present in right part of the periodic table.

(b) These constitute 13 to 18 groups of the periodic table.

(c) Most of these elements are metalloids and nonmetals but some of them are metals also.

(d) The last electron fills in *p*-orbital of valency shell.

(e) The electronic configuration of valence shell is $ns^2np^{1\text{ ñ }6}$ (n = 2 to 7).

(f) ns^2np^6 is stable noble gas configuration. The electronic configuration of He is $1s^2$.

(iii) *d*-block elements

(a) These are present in the middle part of the periodic table (between *s* and *p* block element).

(b) These constitute 3 to 12 groups of the periodic table.

(c) All are metals.

(d) The last electron fills in (n ñ 1)d orbital.

(e) The outermost electronic configuration is $(n\text{ ñ }1)d^{1\text{ ñ }10}\,ns^{1\text{ ñ }2}$ (n = 4 to 7).

(f) There are three series of *d*-block elements:
3*d* series - Sc (21) to Zn (30)
4*d* series - Y (39) to Cd (48)
5*d* series - La (57), Hf (72) to Hg (80)

(iv) *f*-block elements

(a) These are placed separately below the main periodic table.

(b) These are mainly related to group 3 of the periodic table.

(c) There are two series of *f*-block elements:
4*f* series - Lanthanides - 14 elements - Ce (58) to Lu (71)
5*f* series - Actinides - 14 elements - Th (90) to Lr (103)

(d) The last electron fills in (n ñ 2) *f*-orbital.

(e) Their outermost electronic configuration is $(n\text{ ñ }2)f^{1\text{ ñ }14}(n\text{ ñ }1)s^2(n\text{ ñ }1)p^6(n\text{ ñ }1)d^{0\text{ ñ }1}ns^2$ (n = 6 and 7).

BOHRíS CLASSIFICATION OF ELEMENTS

On the basis of electronic configuration of the incomplete shells, the elements are classified into five main categories-

1. Inert gases
2. Representative elements
3. Transition elements
4. Inner transition elements
5. Transuranium elements.

Although this classification is convenient for understanding of chemical properties of the elements, it overlooks the specific properties of the individual elements.

1. Inert gases

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

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

2. Representative or normal elements

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

(b) Inner shells are complete.

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

3. Transition elements

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

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

(c) Their outermost electronic configuration is similar to *d*-block elements *i.e.* $(n\ \text{ñ}\ 1)d^{1\ \text{ñ}\ 10}\,ns^{1\ \text{ñ}\ 2}$.

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

4. Inner transition elements

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

(b) The last shell contains two electrons. Penultimate shell may contain eight or nine electrons and prepenultimate shell contains more than 18 upto 32 electrons.

(c) Their outermost electronic configuration is similar to *f*-block elements *i.e.* $(n\ \text{ñ}\ 2)f^{\ 1\ \text{ñ}\ 14}(n\ \text{ñ}\ 1)s^2(n\ \text{ñ}\ 1)p^6(n\ \text{ñ}\ 1)d^{\ 0\ \text{ñ}\ 1}\,ns^2$.

5. Transuranium elements

Elements of the seventh period after atomic number 93 (*i.e.* actinides) are synthetic elements and are called transuranium elements.

In 1934, an Italian Physicist Enrico Fermi had observed that when an element is bombarded with slow neutrons, the element is transformed into a new element having next higher atomic number. First transuranic element, having atomic number 93 was identified by American Physicist Edwin Mcmilan and Philip H. Abelson. In the next year, element number 94 was discovered in uranium fission products by American Chemist Glenn T. Seaborg and coworkers. The elements 93 and 94 were named Neptunium (Np) and Plutonium (Pu) respectively for Neptune and Pluto, the planets discovered after Uranus.

$${}_{92}\mathrm{U}^{238} + {}_{0}n^{1} \rightarrow {}_{92}\mathrm{U}^{239} + \gamma\text{-rays}$$

$${}_{92}\mathrm{U}^{239} \rightarrow {}_{93}\mathrm{Np}^{239} + {}_{\text{ñ}1}\mathrm{e}^{0}$$

$${}_{93}\mathrm{Np}^{239} \rightarrow {}_{94}\mathrm{Pu}^{239} + {}_{\text{ñ}1}\mathrm{e}^{0}.$$

PERIODICITY IN PROPERTIES

- The term periodicity in properties in the classification of elements means that same properties of the elements reappear at definite intervals when the elements are arranged in order of their increasing atomic numbers. In modern periodic table, these intervals are 2, 8, 8, 18, 18 and 32, *i.e.,* similar properties are observed with elements belonging to the same subgroup which have been arranged in subgroups after the difference of either 2 or 8 or 18 or 32 in atomic numbers as similar valency-shell electronic configuration recur after certain regular intervals of atomic number. **This is the cause of periodicity in properties.**

Electronic configurations of alkali metals

Element	At. No.	Electronic configuration
Li	3	$1s^2\ 2s^1$
Na	11	$1s^2\ 2s^2\ 2p^6\ 3s^1$
K	19	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Rb	37	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 5s^1$
Cs	55	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6\,4d^{10}\,5s^2\,5p^6\,6s^1$
Fr	87	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 5s^2\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^1$

Thus, the **cause of periodicity** of the properties of elements is the repetition of similar electronic configuration of their atoms in the outermost energy shell (or valence shell) after certain regular intervals.

Atomic Radii

- It is usually defined **as the distance between the nucleus and outermost shell where electrons are present.** Three types of radii are commonly used, *i.e.*, (a) covalent radii (b) crystal radii (c) van der Waalsí radii.

- **Covalent radius**

- It is defined as half of the distance between the two nuclei of two like atoms bonded together by a single covalent bond.
- Considering a homonuclear diatomic molecule A_2, bonded together by a single covalent bond, it is assumed that electron clouds of each atom touch each other. Let the bond length be $d_{A\ \text{ñ}\ A}$.

  Then $d_{A\ \text{ñ}\ A} = r_A + r_A = 2r_A$

  $$\text{So} \quad r_A = \frac{d_{A-A}}{2}$$

- In a heteronuclear diatomic AB molecule if both atoms are linked by a single covalent bond and have nearly same electronegativity, the bond length d_{A_B} is equal to sum of covalent radii of the two atoms.

  $$d_{A\text{ñ}B} = r_A + r_B$$

- If the covalent bond is formed between two elements of different electronegativity then we use the following relation:

  $$d_{A\text{ñ}B} = r_A + r_B \text{ ñ } 0.09\ (X_A \text{ ñ } X_B)$$

  where X_A and X_B are electronegativity of A and B respectively. This relation was given by **Stevenson** in 1941.

- **Crystal radii :** It is defined as one half of the distance between the nuclei of two adjacent metal atoms in the metallic closed packed crystal lattice in which metal exhibits a coordination number of 12.

- **van der Waalsí radii :** It is half of the distance between the nuclei of two nonbonded neighbouring atoms of two adjacent molecules.

$$r_{\text{covalent}} < r_{\text{crystal}} < r_{\text{van der Waals}}$$

Atomic radius in the nth orbit is given by

$$r_n = \frac{n^2 a_0}{Z*}$$

where n is principal quantum number (*i.e.*, number of shell), a_0, the Bohrís radius of H-atom (= 0.529≈) and Z^*, the effective nuclear charge.

The screening effect or shielding effect

In a multielectron atom, the electrons of the valence shell (outermost shell) are attracted towards the nucleus and repelled by the electrons present in the inner shells. On account of this, the combined effect of the attractive and repulsive force acting on the valence electron is that the valence electron experiences less attraction from the nucleus. This decreased force of attraction exerted by the nucleus on the valence electrons due to the presence of electrons in the inner shells, is called screening effect or shielding effect.

Effective atomic number

Due to screening effect the valence electron experiences less attraction towards nucleus. This results in decrease in the nuclear charge (*Z*) actually present on the nucleus. The reduced nuclear charge is termed **effective nuclear charge** and is represented by Z^*. It is related to actual nuclear charge (*Z*) by the following formula :

Z^* = (*Z* ñ σ) where σ is screening constant.

The magnitude of ëσí is determined by the **Slaterís rules.** The contribution of inner electrons to the magnitude of ëσí is calculated in the following ways:

Slater's rule for Estimating Effective Nuclear Charges, Z^*

(1) Write down the electronic configuration of the element in the following order and grouping:
(1*s*), (2*s*, 2*p*), (3*s*, 3*p*), (3*d*), (4*s*, 4*p*), (4*d*), (4*f*), (5*s*, 5*p*) etc.

(2) Electrons in any group higher in the sequence show that the electron under consideration contribute nothing to the shielding σ.

(3) Then for an electron in an *ns* or *np* orbital
   (a) all other electrons in (*ns*, *np*) group contribute σ = 0.35 each
   (b) all electrons in the (*n* ñ 1) shell contribute σ = 0.85 each
   (c) all electrons in the (*n* ñ 2) or lower shell contribute σ = 1.00 each

(4) For electron in an *nd* or *nf* orbital, all electrons in the same group contribute σ = 0.35 each; those in group lying lower in the sequence than the (*nd*) or (*nf*) group contribute σ = 1.00 each.

- Radius is also dependent on the extent of force of attraction which pulls outer shell inward.

Variation in Period

	Li	Be	B	C	N	O	F	Ne
Z	3	4	5	6	7	8	9	10
σ	1.7	2.05	2.40	2.75	3.10	3.45	3.80	4.15
Z*	1.30	1.95	2.60	3.25	3.90	4.55	5.20	5.85
n	2	2	2	2	2	2	2	2
r_n (pm)	123	90	80	77	75	74	72	160

In a period, left to right :

- *Z* (atomic no.) increases (by one unit)
- Z^* also increases (but by 0.65 unit)
- *n* (number of shells) remains constant

Thus $r_n \propto 1/Z^*$

In case of Noble gases (as in Ne) there is no covalent bond formation, hence only van der Waalsí radius is considered. Thus there is high jump in the value of radius from F (72 pm) to Ne (160 pm).

Variation in a Group

Element	*Z*	σ	*Z**	*n*	Radius (pm)
Li	3	1.7	1.3	2	123
Na	11	8.8	2.2	3	157
K	19	16.8	2.2	4	203
Rb	37	34.8	2.2	5	216
Cs	55	52.8	2.2	6	235
Fr	87	84.8	2.2	7	ó

In a group, top to bottom :

- Z increases
- Z^* almost remains constant
- n increases

  Thus, $r_n \propto n^2$

  Hence atomic radius in a group is dependent on the value of n.

Illustration - 4 : Calculate the screening constant in zinc **(a)** for a 4*s*-electron **(b)** for a 3*d*-electron.

Soln.: (a) The electronic configuration of zinc (30) is, $(1s)^2$ $(2s\ 2p)^8$ $(3s\ 3p)^8$ $(3d)^{10}$ $(4s)^2$

σ = 10 ◊ 1.0 + 18 ◊ 0.85 + 1 ◊ 0.35 = 25.65

(b) For 3*d*-electron, σ = 18 ◊ 1.0 + 9 ◊ 0.35 = 21.15

It is clear that the magnitude of screening constant in the case of *s*- and *p*-block elements increases in a period as well as in a group as the atomic number increases.

Ionic Radii

It is defined as the distance between the nucleus and outermost shell of an ion or it is the distance between the nucleus and the point where the nucleus exerts its influence on the electron cloud.

Metal ions are smaller than the atoms from which they are formed.

When a positive ion is formed, the number of positive charges on the nucleus exceeds the number of orbital electrons, and the effective nuclear charge (which is the number of charges on the nucleus to the number of electrons) is increased. This results in the remaining electrons being more strongly attracted by the nucleus. These electrons are pulled in further reducing the size.

A positive ion is thus always smaller than the corresponding atom, and the more the electrons are removed, the smaller the ion becomes.

Thus $Mg > Mg^+ > Mg^{2+}$

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

The negative ion is always larger than that of the corresponding atom.

- Negative ion is formed by gain of one or more electrons in the neutral atom and thus number of electrons increases but magnitude of nuclear charge remains the same.
- Due to decrease in nuclear charge per electron, there is an expansion of outer shell. Thus size of anion is increased.

  $O^{2ñ}$ > $O^{ñ}$ > O

  $I^{ñ}$ > I > I^+

These can be explained on the basis of Z/e ratio $\left(\frac{\text{Nuclear charge}}{\text{No. of electrons}}\right)\cdot$ When Z/e ratio increases, the size decreases and when Z/e ratio decreases, the size increases.

	Na	Na^+	Cl	$Cl^{ñ}$
Z/e	$\frac{11}{11}=1.0;$	$\frac{11}{10}=1.1;$	$\frac{17}{17}=1.0;$	$\frac{17}{18}=0.95;$

So $Na^+ < Na$ $Cl^{ñ} > Cl$

For isoelectronic species the size decreases with an increase of atomic number. This is illustrated in the following table:

Atom or Ion	Atomic Number (Z)	No. of electrons (e)	Z/e ratio	Size in ≈
$O^{2ñ}$	8	10	0.8	1.40
$F^{ñ}$	9	10	0.9	1.30
Ne	10	10	1.0	1.12
Na^+	11	10	1.1	0.95
Mg^{2+}	12	10	1.2	0.65

H^+ and Cs^+ are the smallest and largest cations respectively.
$H^{ñ}$ and $I^{ñ}$ are the smallest and largest anions respectively.

Illustration - 5 : Arrange the following ions in the order of increasing size:
Be^{2+}, $Cl^{ñ}$, $S^{2ñ}$, Na^+, Mg^{2+}, $Br^{ñ}$

Soln.: $Be^{2+} < Mg^{2+} < Na^+ < Cl^{ñ} < S^{2ñ} < Br^{ñ}$

Illustration - 6 : Which of the following species will have the largest and the smallest size?

$$Mg,\ Mg^{2+},\ Al,\ Al^{3+}$$

Soln.: Atomic radii decrease across a period. So, Mg will have larger size than Al. Cations are smaller than their parent atoms. So size is : $Mg^{2+} < Mg$ and $Al^{3+} < Al$. Now, among isoelectronic ions, the ion with the larger positive nuclear charge will have a smaller radius *i.e.*, size is : $Al^{3+} < Al$. Hence, the largest species is Mg and the smallest species is Al^{3+}.

Illustration - 7 : Which of the following atoms and ions are isoelectronic?
(i) Al^{3+} (ii) F (iii) $Cl^{ñ}$ (iv) $O^{2ñ}$ (v) Na (vi) Mg^{2+}
Arrange the isoelectronic ions in the decreasing order of their size.

Soln.: The number of electrons in these atoms or ions are :

Ion or atom	Al^{3+}	F	$Cl^{ñ}$	$O^{2ñ}$	Na	Mg^{2+}
No. of electrons	10	9	18	10	11	10

Thus Al^{3+}, $O^{2ñ}$ and Mg^{2+} are isoelectronic ions because all the three ions have ten electrons. Now nuclear charge in Al^{3+} is +13, in $O^{2ñ}$ is +8 and in Mg^{2+} is +12. With increase in nuclear charge (electrons remain same), size will decrease. Consequently, the size follows the order:

$$O^{2ñ} > Mg^{2+} > Al^{3+}.$$

Illustration - 8 : Give the formula of a species that will be isoelectronic with the following atoms or ions?
(i) Ar (ii) $F^{ñ}$ (iii) K^+ (iv) Sr^{2+}

Soln.: Isoelectronic species are those which have same number of electrons.
(i) Ar has 18 electrons. Therefore, the species $P^{3ñ}$, $S^{2ñ}$, $Cl^{ñ}$, K^+, Ca^{2+}, etc., are isoelectronic to Ar.
(ii) $F^{ñ}$ has 10 electrons. Therefore, the species $N^{3ñ}$, $O^{2ñ}$, Ne, Na^+, Mg^{2+}, etc. are isoelectronic to $F^{ñ}$.
(iii) K^+ has 18 electrons. Therefore, the species $P^{3ñ}$, $S^{2ñ}$, $Cl^{ñ}$, Ar, Ca^{2+}, etc., are isoelectronic to K^+.
(iv) Sr^{2+} has 36 electrons. Therefore, the species $Br^{ñ}$, Kr, Rb^+, etc., are isoelectronic to Sr^{2+}.

Ionisation Potential or Ionisation Energy

The minimum amount of energy required to remove the most loosely bound electron from an isolated atom in the gaseous state is known as ionisation potential or ionisation energy or first ionisation potential (I_1) of the element.

$$M \xrightarrow{\text{energy}} M^+ + e^-$$

The energy required to remove the second electron from the monovalent cation is called second ionisation potential (I_2).

$$M^+ + I_2 \rightarrow M^{2+} + e^{ñ}$$

Similarly, we have third, fourth . . . ionisation potentials.

$$M^{2+} + I_3 \rightarrow M^{3+}\ e^{ñ}$$

$$M^{3+} + I_4 \rightarrow M^{4+} + e^{ñ}$$

It is observed that I_2 is higher than I_1, I_3 is higher than I_2 and so on, *i.e.*, $I_1 < I_2 < I_3 < I_4$. The increase in the values of successive ionisation potentials can be explained on the basis that effective nuclear charge increases from $M_{(g)}$ to $M^{n+}_{(g)}$, *i.e.*, force of attraction of the outermost electron towards nucleus increases.

Factors affecting the value of ionisation potential

Properties	Effect
1. Atomic Size	Larger the atomic size, smaller is the value of ionisation potential.
2. Screening Effect	Higher the screening effect, the lesser is the value of ionisation potential.
3. Nuclear charge	Ionisation potential increases with the increase in nuclear charge.
4. Penetration effect or shape of orbtial	Values of ionisation potential *s, p, d* and *f* electrons are as : $s > p > d > f$.

If an atom has fully filled or half-filled orbitals, its *IE* is higher than expected normally from its position in the periodic table. *e.g.*

	Li	Be	B	C	N	O	F	Ne
	$2s^1$	$2s^2$	$2s^22p^1$	$2s^22p^2$	$2s^22p^3$	$2s^22p^4$	$2s^22p^5$	$2s^22p^6$
IE (eV)	5.4	9.3	8.3	11.2	14.5	13.6	17.4	21.6

Be (fully filled 2*s* orbital) and N (half filled 2*p* orbital) have higher values than expected due to stable configurations.

- **Variation of (*IE*) in a group**

Force of attraction between electrons and nucleus decreases and tendency to remove the valence electron increases. Hence (*IE*) decreases on moving down the group.

- **Variation of (*IE*) in a period**

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.

- The energies required to remove subsequent electrons from the atom in the gaseous state, are known as **successive ionisation energies.**

- **Successive ionisation energies are higher :** The second ionisation energies are higher than the first ionisation energies. This is mainly due to the fact that after the removal of the first electron, the atom changes into monovalent positive ion. In the ion, the number of electrons decreases but the nuclear charge remains the same. As a result of this, the remaining electrons are held more tightly by the nucleus and it becomes difficult to remove the second electron. Hence, the value of second ionisation energy (IE_2) is greater than that of the first (IE_1).

Illustration - 9 : The ionisation energy of lithium is 520 kJ mol$^{\text{ñ}1}$. Calculate the amount of energy required to convert 210 mg of lithium atoms in gaseous state into Li^+ ion.

Soln.: $Li_{(g)} \rightarrow Li^+_{(g)}$, IE = 520 kJ mol$^{\text{ñ}1}$

Mass of lithium = 210 mg = 210 ◊ $10^{\text{ñ}3}$ g = 21 ◊ $10^{\text{ñ}2}$ g.

∵ Number of moles of lithium = $\frac{21 \times 10^{-2}}{7} = 3 \times 10^{-2}$ mole

∴ Energy required to convert 3 ◊ $10^{\text{ñ}2}$ mole atoms of lithium into Li^+ ions is
= 520 ◊ 3 ◊ $10^{\text{ñ}2}$ = 15.60 kJ.

Illustration - 10 : How much energy in joules must be needed to convert all the atoms of sodium to sodium ions present in 2.3 mg of sodium vapours? Ionisation enthalpy of sodium is 495 kJ mol$^{\text{ñ}1}$ (at. mass of Na = 23).

Soln.: According to the definition of ionisation enthalpy,

$Na_{(g)} + IE \rightarrow Na^{+}_{(g)} + e^{ñ}_{(g)}$; IE = 495 kJ $mol^{ñ1}$

The amount of energy needed to ionise 1 mole of sodium vapours = 495 kJ $mol^{ñ1}$

Moles of sodium vapours present in given sample = $\frac{2.3 \times 10^{-3}}{23} = 1 \times 10^{-4}$ mole

∴ Amount of energy needed to ionise 1 ◊ $10^{ñ4}$ mol of sodium vapours

= 495 ◊ 1 ◊ $10^{ñ4}$ = 0.0495 kJ or 49.5 J.

Electron Affinity or Electron Gain Enthalpy

- The amount of energy released when an electron is added to an isolated gaseous atom to produce a monovalent anion is called electron affinity or first electron affinity.

  $A + e \xrightarrow{\quad} A^{-} + \text{energy}$
- Most elements have a negative electron affinity. This means they do not require energy to gain an electron; instead, they release energy. Atoms more attracted to extra electrons have a more negative electron affinity. Chlorine most strongly attracts extra electrons; radon most weakly attracts an extra electron.
- Electron gain enthalpy is expressed in terms of enthalpy change when an electron is added to an isolated gaseous atom. Larger the energy released in this process, more negative the value of electron gain enthalpy or more positive the value of electron affinity.

Although electron affinities vary in an irregular manner across the table, some patterns emerge. Generally, non-metals have more negative electron affinities than metals. However, the noble gases are an exception: they have positive electron affinities.

Factors affecting the value of electron affinity

Properties	Effects
1. Nuclear charge	Electron affinity increases with the increase in nuclear charge.
2. Atomic size	With the increase in atomic size, electron affinity decreases.
3. Electronic configuration	Electron affinities are low or almost zero in cases of stable configurations i.e. half filled or full-filled valence shell.

	Li	Be	B	C	N	O	F	Ne
	$2s^1$	$2s^2$	$2s^22p^1$	$2s^22p^2$	$2s^22p^3$	$2s^22p^4$	$2s^22p^5$	$2s^22p^6$
eV	ñ0.61	0.0	ñ 0.30	ñ1.25	ñ 0.20	ñ1.48	ñ3.6	0.0

Be, N and Ne have low values due to stable configurations.

- The electron affinities increase across a row (since the radius slightly decreases, because of the increased attraction from the nucleus, and the number of electrons in the top shell increases, helping the atom reach maximum stability) in the periodic table and decrease going down a family (because of a large increase in radius and number of electron that decrease the stability of the atom, repulsing each other).
- Electron affinities are not limited to the elements but also apply to molecules. For instance the electron affinity for benzene is positive, that of naphthalene near zero and that of anthracene positive.
- **Successive electron affinities**

  Like ionisation energies, the second and higher electron affinities are also possible. However, second electron is added to a negatively charged ion and the addition is opposed by coulombic repulsions. The energy has to be supplied to force the second electron into the anion.

Illustration-11 : Which of the following will have the most negative electron gain enthalpy and which the least negative? P, S, Cl, F. Explain your answer.

Soln.: These elements can be arranged in different periods and groups as :

Group No.	15	16	17
2nd period			F
3rd period	P	S	Cl

We know that electron gain enthalpy becomes more negative across a period as we move from left to right. Within the group, electron gain enthalpy becomes less negative as we move down a group. Therefore, P should have the least negative electron gain enthalpy and F should have the most negative electron gain enthalpy. However, adding an electron to smaller 2*p*-orbital leads to greater inter-electronic repulsions than adding an electron to larger 3*p*-orbital. Hence, Cl has more negative electron gain enthalpy than F. Therefore, Cl has the most negative electron gain enthalpy and P has the least negative electron gain enthalpy.

Illustration - 12 : Which element will have the greatest negative electron gain enthalpy? Give reasons?

(a) [Ne] $3s^2$ $3p^3$ (b)[Ne] $3s^2$ $3p^4$ (c) [Ne]$3s^2$ $3p^5$ (d) [Ne] $3s^2$ $3p^6$ $4s^1$ $3d^5$

Soln.: The element corresponding to configuration (c) will have he highest negative electron gain enthalpy. It corresponds to chlorine (halogen) which has very strong tendency to accept an electron to acquire noble gas configuration.

Illustration-13 : Which of the following pairs would have a higher electron gain enthalpy?

(i) N or O (ii) F or Cl (iii) Br or I (iv) B or Al

Soln.: Higher negative electron gain enthalpy.

(i) O (ii) Cl (iii) Br (iv) B

Electronegativity

- Electronegativity is a measure of the tendency of an element to attract electrons to itself. In a molecule, tendency of the atom to attract bonding pair towards itself is its electronegativity. *B* is said to be more electronegative than *A* if it pulls bonding pair towards itself.

  $$(A)\ \overset{\bullet}{\times}\ (B) \longrightarrow \overset{\delta+}{(A)} - \overset{\delta-}{(B)}$$
- An arbitrary value of 4.0 has been assigned to fluorine (most electronegative element) and the electronegativities of other elements have been calculated against this standard by application of following formula:

  $$X_A \text{ ñ } X_B = 0.208[E_{A-B} - (E_{A-A} \times E_{B-B})^{1/2}]^{1/2}$$

  where X_A and X_B are the electronegativities of two atoms *A* and *B* and $E_{A\text{ ñ }B}$, $E_{(A\text{ ñ }A)}$ and $E_{B\text{ ñ }B}$ are bond energies of molecules *A* ñ *B*, A_2 and B_2, respectively
- Mulliken regarded electronegativity as the average value of ionisation potential and electron affinity of an atom.

  $$\text{Electronegativity} = \frac{IP + EA}{2}$$
- On Pauling scale, $\text{electronegativity of an atom} = \dfrac{IP + EA}{5.6}$

  Values of *IP* and *EA* are taken in eV
- Electronegativity of an atom $= \dfrac{IP + EA}{2 \times 62.5}$

  values of *IP* and *EA* are taken in kilo calories per mole

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Illustration-14 : Calculate the electronegativity of fluorine from the following data :

$E_{\text{H ñ H}}$ = 104.2 kcal $\text{mol}^{\text{ñ}1}$, $E_{\text{F ñ F}}$ = 36.6 kcal $\text{mol}^{\text{ñ}1}$, $E_{\text{H ñ F}}$ = 134.6 kcal $\text{mol}^{\text{ñ}1}$, $\chi_H = 2.1$

Soln.: According to the equation, $E_{A\text{ ñ }B}$ ñ $(E_{A\text{ ñ }A}$ ◊ $E_{B\text{ ñ }B})^{1/2} = \Delta'$

$\Delta' = E_{\text{H ñ F}}$ ñ $(E_{\text{H ñ H}}$ ◊ $E_{\text{F ñ F}})^{1/2}$ = 134.6 ñ (104.2 ◊ 36.6)$^{1/2}$ = 72.85 kcal

We have, $0.182\sqrt{\Delta'} = \chi_A - \chi_B$

$\therefore$ χ_F ñ $\chi_H = 0.182(72.85)^{1/2} = 1.55$

$\therefore$ $\chi_F = 1.55 + \chi_H = 1.55 + 2.1 = 3.65$

Illustration-15 : Calculate the electronegativity of carbon from the following data :

$E_{\text{H ñ H}}$ = 104.2 kcal mol$^{\text{ñ}1}$, $E_{\text{C ñ C}}$ = 83.1 kcal mol$^{\text{ñ}1}$, $E_{\text{C ñ H}}$ = 98.8 kcal mol$^{\text{ñ}1}$, $\chi_H = 2.1$

Soln.: According to the equation, $E_{A\,\text{ñ}\,B}$ ñ $(E_{A\,\text{ñ}\,A}$ ◊ $E_{B\,\text{ñ}\,B})^{1/2} = \Delta'$

$\Delta' = E_{\text{C ñ H}}$ ñ $(E_{\text{H ñ H}}$ ◊ $E_{\text{C ñ C}})^{1/2}$ = 98.8 ñ $(104.2$ ◊ $83.1)^{1/2}$ = 5.75 kcal

We have, $0.182\sqrt{\Delta'} = \chi_A - \chi_B$

$\therefore$ χ_C ñ $\chi_H = 0.182(5.75)^{1/2} = 0.44$

$\therefore$ $\chi_C = 0.44 + \chi_H = 0.44 + 2.1 = 2.54$

Valency

The electrons present in the outermost shell are called valence electrons and these electrons determine the valency of the atom.

In case of representative elements, the valency is generally equal to either number of valence electrons or equal to eight minus the number of valence electrons. However, the transition elements, exhibit variable valency.

Variation of Valency in the Periodic Table

(a) Variation in a period : The number of valence electrons increases from 1 to 8 on moving across a period, the valency of the elements with respect to hydrogen and chlorine increases from 1 to 4 and then decreases to zero. This may be illustrated by taking the examples of elements of second and third periods as given below in Table. The number in brackets represents the valency of the elements in the compound.

Table: Variation of valency of elements of second and third periods.

Elements of second period						
Li	Be	B	C	N	O	F
Valency with respect to H						
LiH	BeH_2	BH_3	CH_4	NH_3	H_2O	HF
(1)	(2)	(3)	(4)	(3)	(2)	(1)
Valency with respect to Cl						
LiCl	$BeCl_2$	BCl_3	CCl_4	NCl_3	Cl_2O	ClF
(1)	(2)	(3)	(4)	(3)	(2)	(1)
Elements of third period						
Na	Mg	Al	Si	P	S	Cl
Valency with respect to H						
NaH	MgH_2	AlH_3	SiH_4	PH_3	H_2S	HCl
(1)	(2)	(3)	(4)	(3)	(2)	(1)

(b) Variation in a group : On moving down a group, the number of valence electrons remains same and, therefore, all the elements in a group exhibit same valency. For example, all the elements of group 1 have valency one *i.e.*, are monovalent and those of group 2 have valency two or divalent.

Illustration-16 : Are the oxidation states and covalency of Al in $[AlCl(H_2O)_5]^{2+}$ same?

Soln.: No. The oxidation state of Al is +3 and the covalency is 6.

Atomic volume

Atomic volume increases in a group from top to bottom. The increase is due to the increase in the number of energy shells.

In a period from left to right, atomic volume varies cyclically, *i.e.*, it decreases at first for some elements, becomes minimum in the middle and then increases. The following two factors explain this trend:

(a) Atomic radii decrease due to increase of nuclear charge.

(b) The number of valency electrons increases in a period.

As to accommodate all the valency electrons, the volume increases. These two factors oppose each other. The effect of first factor is more on the left hand side and that of the second factor is more on the right hand side.

Increases ↓								Increases ↓
Li	Be	B	C	N	O	F	Ne	
13	5	5	5	14	11	15	17	
Na	Mg	Al	Si	P	S	Cl	Ar	
24	14	10	12	17	16	19	24	
K	Ca	Ga	Ge	As	Se	Br	Kr	
46	26	12	13	16	16	23	33	

The maximum value of atomic volume (87) is observed in the case of francium (last member of alkali group).

Density

The density of the elements in solid state varies periodically with their atomic numbers. At first, the density increases gradually in a period and becomes maximum somewhere for the central members and then starts decreasing afterwards gradually.

Li	Be	B	C	N	O	F
0.5	1.8	2.3	2.2	ó	ó	ó
Na	Mg	Al	Si	P	S	Cl
1.0	1.7	2.7	2.3	1.8	2.1	ó
K					Se	Br
0.9					4.8	3.1
Rb					Te	I
1.5					6.2	4.9
Cs						
1.9						

In solids, osmium has the highest density (22.6) and in liquids, Hg has the highest density (13.6).

Melting and boiling points

The melting points of the elements exhibit same periodicity with rise of atomic number. It is observed that elements with low values of atomic volumes have high melting points while elements with high values of atomic volumes have low melting points. In general, melting points of elements in any period at first increase and become maximum somewhere in the centre and thereafter begins to decrease.

The melting points of 2nd and 3rd period elements are as shown in the table.

Elements	Li	Be	B	C	N	O	F	Ne
M.Pt. (K)	454	1550	2303	4000	63	54	53	24
Elements	Na	Mg	Al	Si	P	S	Cl	Ar
M.Pt. (K)	370.8	924	933	1693	317	392	172	83.6

Tungsten has the maximum melting point (3410∞C) amongst metals and carbon has the maximum melting point (3727∞C) amongst non-metals. Helium has the minimum melting point (ñ270∞C). The metals Cs, Ga and Hg are known in liquid state at 30∞C.

The boiling points of the elements also show similar trends, however, the regularities are not so striking as noted in the case of melting points.

COMPETITION WINDOW

- **Importance of electron gain enthalpy**

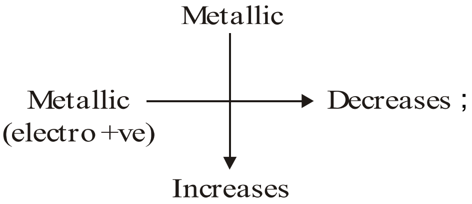
➢ Certain properties of the elements can be predicted on the basis of values of electron gain enthalpies.

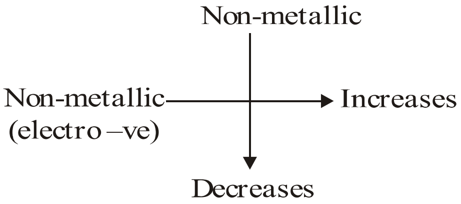
(i) The elements having high negative values of electron gain enthalpy are capable of accepting electron easily. They form anions and thus form ionic (electrovalent) compounds. These elements are electronegative in nature.

(ii) The elements having high negative electron gain enthalpies act as strong oxidising agents, for example, F, Cl, Br, O, S, etc.

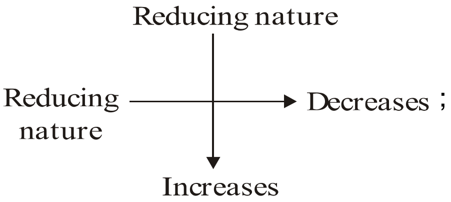
➢ On the basis of the general trend of ionisation enthalpy and electron gain enthalpy, the following properties can be predicted.

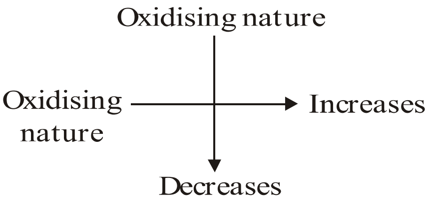
(i) Metallic nature decreases in a period while non-metallic nature increases. Metallic nature increases in a group while non-metallic nature decreases. The arrow (↓) represents a group and (→) represents a period.



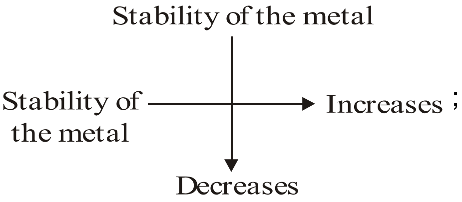


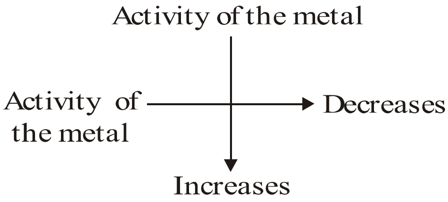
(ii) Reducing nature decreases in a period while oxidising nature increases. The reducing nature increases in a group while oxidising nature decreases.



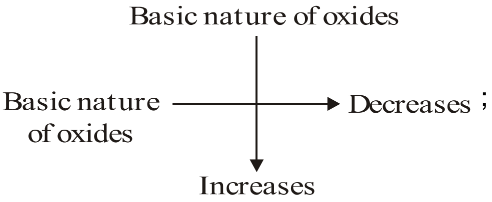


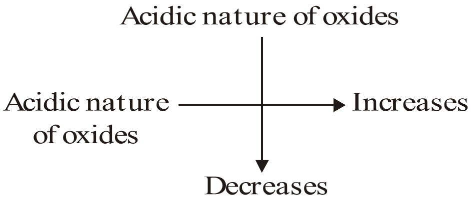
(iii) Stability of metal increases while activity of the metal decreases in a period and in a group stability decreases while activity increases.





(iv) The basic nature of the oxides decreases in a period while acidic nature increases. In a group, basic nature increases while acidic nature decreases.





- **Importance of electronegativity**

  The following predictions can be made from values of electronegativity.

- Nature of the bond between two atoms can be predicted from the electronegativity difference of the two atoms.
  - (a) The difference X_A ñ $X_B = 0$, *i.e.*, $X_A = X_B$, the bond is purely covalent. For example, H_2, Cl_2, O_2 and N_2 molecules are purely covalent and non-polar.
  - (b) The difference X_A ñ X_B is small, *i.e.*, $X_A > X_B$, the bond is polar covalent.
  - (c) The difference X_A ñ X_B is 1.7, the bond is 50% covalent and 50% ionic.
  - (d) The difference X_A ñ X_B is very high, the bond is more ionic and less covalent. The molecule will be represented in such case as BA ($B^{+}A^{ñ}$).
- The elements having low values of electronegativity are metals while the elements having high values of electronegativity are non-metals.
- (X_O ñ X_A) difference predicts the nature of the oxides formed by the elements A. X_O is electronegativity of oxygen.

  X_O ñ X_A is large, the oxide shows basic nature (*e.g.* Na_2O)
  X_O ñ X_A is small, the oxide shows acidic nature (*e.g.* SO_2).

- **Metals, non-metals and metalloids**
- Characteristic properties of metallic and non-metallic elements:

Metallic elements	**Non-metallic elements**
Distinguishing lustre (shine)	Non-lustrous, various colours
Malleable and ductile (flexible) as solids	Brittle, hard or soft
Conduct heat and electricity	Poor conductors
Metallic oxides are basic, ionic	Nonmetallic oxides are acidic, compounds
Cations in aqueous solution	Anions, oxyanions in aqueous solution

- **Metals**
- Most metals are malleable (can be pounded into this sheets; a sugar cube chunk of gold can be pounded into a thin sheet which will cover a football field), and ductile (can be drawn out into a thin wire).

  All are solids at room temperature (except mercury, which is a liquid).
- Metals tend to have low ionisation energies, and typically lose electrons (*i.e.* are oxidised) when they undergo chemical reactions.
  - Alkali metals are always 1 + (lose electron in *s* subshell).
  - Alkaline earth metals are always 2+ (lose both electrons in *s* subshell).
  - Transition metal ions do not follow an obvious pattern, 2 + is common, and 1+ and 3+ are also observed.
- Compounds of metals with non-metals tend to be ionic in nature.
- Most metal oxides are basic oxides; those that dissolve in water react to form metal hydroxides:

  Metal oxide + water → metal hydroxide
  $Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2NaOH_{(aq)}$
  $CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2\,(aq)}$
- Metal oxides exhibit their basic chemical nature by reacting with acids to form salts and water.

  Metal oxide + acid → salt + water
  $MgO_{(s)} + HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_2O_{(l)}$

- **Non-metals**
- Vary greatly in appearance.
- Non-lustrous.
- Poor conductors of heat and electricity.
- The melting points of non-metals are generally lower than metals.
- Seven non-metals exist under standard conditions as diatomic molecules.

  1. $H_{2\,(g)}$ 2. $N_{2\,(g)}$ 3. $O_{2\,(g)}$ 4. $F_{2\,(g)}$ 5. $Cl_{2\,(g)}$ 6. $Br_{2\,(l)}$ 7. $I_{2\,(s)}$

- Non-metals, when reacting with metals, tend to gain electrons (typically attaining noble gas electron configuration) and become anions:

  Non-metal + metal → salt

  $3Br_{2\,(l)} + 2Al_{(s)} \rightarrow 2AlBr_{3\,(s)}$
- Compounds composed entirely of non-metals are molecular substances (not ionic).
- Most non-metal oxides are acidic oxides. Those that dissolve in water react to form acids:

  Non-metal oxide + water → acid

  $CO_{2\,(g)} + H_2O_{\,(l)} \rightarrow H_2CO_{3\,(aq)}$ [carbonic acid]

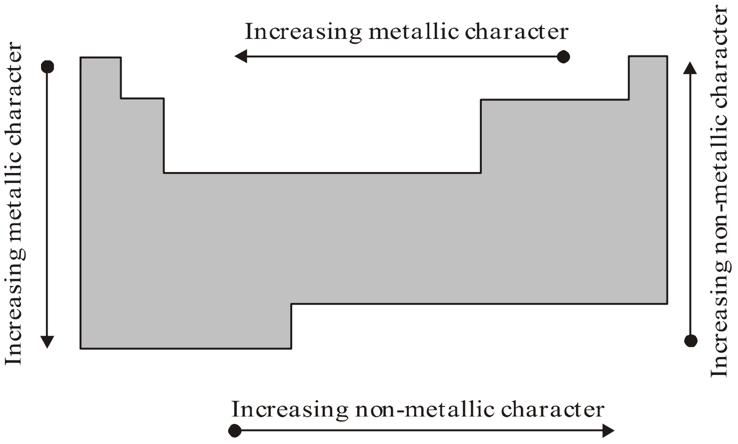
  Carbonated water is slightly acidic)
- Non-metal oxides can combine with bases to form salts.

  Non-metal oxide + base → salt

  $CO_{2\,(g)} + 2NaOH_{\,(aq)} \rightarrow Na_2CO_{3\,(aq)} + H_2O_{\,(l)}$

- **Metalloids**
- Properties intermediate between metals and non-metals.
- Silicon, for example, appears lustrous, but is not malleable or ductile (it is *brittle* - a characteristic of some non-metals). It is a much poorer conductor of heat and electricity than the metals. Metalloids are useful in the semiconductor industry.

- **Trends in metallic and non-metallic character**
- Metallic character is strongest for the elements in the leftmost part of the periodic table, and tends to decrease as we move to the right in any period (non-metallic character increases with increasing ionisation values).
- Within any group of elements (columns), the metallic character increases from top to bottom (the ionization values generally decrease as we move down a group). This general trend is not necessarily observed with the transition metals.



General trends for the metallic character of an element (metallic character is the opposite of non-metallic character)

Miscellaneous Xamples

1. Arrange the following in:

(i) Decreasing ionic size: Mg^{2+}, $O^{2ñ}$, Na^{+}, $F^{ñ}$

(ii) Increasing acidic property: ZnO, Na_2O_2, P_2O_5, MgO

(iii) Increasing first ionisation potential: Mg, Al, Si, Na

(iv) Increasing size: $Cl^{ñ}$, $S^{2ñ}$, Ca^{2+}, Ar

(v) Increasing order of ionic size:$N^{3ñ}$, Na^{+}, $F^{ñ}$, $O^{2ñ}$, Mg^{2+}

(vi) Increasing order of basic character: MgO, SrO, K_2O, NiO, Cs_2O

(vii)Arrange the following ions in order of their increasing radii: Li^{+}, Mg^{2+}, K^{+}, Al^{3+}.

Soln.: (i) $O^{2ñ} > F^{ñ} > Na^{+} > Mg^{2+}$. These are isoelectronic ions (each has 10 electrons). In such case the ionic radii decreases with increase in nuclear charge. The nuclear charge is 8 (for $O^{2ñ}$) 9 (for $F^{ñ}$), 11(for Na^{+}) and 12 (for Mg^{2+}).

(ii) $Na_2O_2 < MgO < ZnO < P_2O_5$
In case of oxides, the acidic strength increases with increase in oxidation state of central atom.

(iii) In case of elements of 3rd period the first *IE* follows the order Na < Mg > Al < Si.

(iv) The correct order is $Ca^{2+} < Ar < Cl^{ñ} < S^{2ñ}$. These are isoelectronic with 18 electrons. Lower the nuclear charge, higher is the ionic radii.

(v) The correct order is $Mg^{2+} < Na^{+} > F^{ñ} < O^{2ñ} < N^{3ñ}$. For isoelectronic ions, ionic radii decreases with increase in nuclear charge.

(vi) The increasing order of basic character is
$NiO < MgO < SrO < K_2O < Cs_2O$
In a group the basic character of oxide increases as we move from top to bottom. Thus we have $K_2O < Cs_2O$ (group 1) and $MgO < SrO$ (group 2)
Also as the group number increases the basic character decreases, so NiO is least basic, Ni is in group 10.

(vii) Out of the given species Al^{3+} and Mg^{2+} are isoelectric (10 electrons) and size of Al^{3+} is smaller, so $Al^{3+} < Mg^{2+}$. Li^{+} and K^{+} belong to group I and Li^{+} is higher up in the group so $Li^{+} < K^{+}$. The correct order is $Al^{3+} < Li^{+} < Mg^{2+} < K^{+}$.

---

2. The first ionisation energy of carbon atom is greater than that of boron atom whereas, the reverse is true for the second ionisation energy. Explain.

Soln.: Along the period from left to right, as atomic size decreases, ionisation energy increases. So, IE_1 of carbon is greater than IE_1 of boron.
IE_2 in case of C^{+} ($1s^22s^22p^1$) will be less than the IE_2 in case of B^{+}($1s^22s^2$). In C^{+} the electron has to be removed from 2*p*-orbital where as in B^{+} it has to be removed from 2*s*-orbital.

---

3. (i) There are only 14 lanthanoids and only 14 actinoids. Explain.

(ii) Why argon (at. mass 39.94) has been placed before potassium (at. mass 39.10) in the periodic table.

Soln.: (i) In lanthanoids and actinoids, the differentiating electron enters to $(n ñ 2)f$-subshell. The maximum capacity of *f*-subshell is of 14 electrons. Thus, there are only 14 lanthanides ($4f^{1\ ñ\ 14}$) and only 14 actinides ($5f^{1\ ñ\ 14}$).

(ii) In modern periodic table, elements have been placed in order of their increasing atomic numbers. The atomic number of argon is 18 and that of potassium is 19. Thus, argon has been placed before potassium.

4. IE_2 for alkali metals show a jump whereas IE_3 for alkaline earth metals show a jump. Why?

Soln.: The jump in *IE* values are noticed when the successive removal of electrons occurs from different shells. In alkali metals (n') removal of I and II electrons occurs from *n*, (*n* ñ 1) shells respectively and thus shows a jump in IE_2 values. This condition occurs during removal of II and III electrons from alkaline earth metals.

---

5. F_2 is more powerful oxidant than Cl_2. Why?

Soln.: Of all the halogens, F has highest positive electrode potential (F_2 = 2.87, Cl_2 = 1.36, Br_2 = 1.09 and I_2 = 0.53 volt). *i.e.* it is most easily reduced hence acts as strongest oxidising agent.

---

6. The radius of a cation is lesser than the atom while that of an anion is greater than the atom. Why?

Soln.: In a cation, the nuclear charge acts on lesser number of electrons and thus, the electron cloud contracts. Hence, the net result is decrease in size. In an anion, the nuclear charge acts on larger number of electrons and thus, the electron cloud expands. Hence, the net result is increase in size. This can also be explained on the basis of *Z/e* ratio. In cation *Z/e* ratio increases, hence size decreases while in an anion *Z/e* ratio decreases and so the size increases.

---

7. Why Be and Mg atoms do not impart colour in flame?

Soln.: Be and Mg atoms are comparatively smaller and their ionisation energies are very high. Hence, their electrons are not excited by the energy of the flame to higher energy state. Therefore, these elements do not give any colour in flame.

---

8. Chlorine has maximum electron affinity. Why?

Soln.: Cl has more negative electron gain enthalpy than F. The less negative electron gain enthalpy of F is due to its very small size which has strong electron-electron repulsions. Therefore F has less tendency to accept an electron and its electron gain enthalpy is less negative than Cl.

---

9. EA_2 of halogens is zero. Give reason.

Soln.: After addition of one electron halogen acquires nearest noble gas configuration. Addition of 2nd electron is not possible and thus EA_2 = 0.

$X : ns^2\ np^5$; $X^{ñ} : ns^2\ np^6$; $X^{2ñ}$: not possible.

---

10. The atomic radius decreases along the period but inert gases have maximum atomic radius in a period. Given reason.

Soln.: Atomic radii of inert gases are Van der Waals radii whereas for rest all atoms covalent radii are determined experimentally. Van der Waals radii is greater in comparison to covalent radii.

---

11. **(a)** Why N has higher 1st ionisation potential than O-atom?

(b) Why Mg has higher 1st ionisation potential than Al-atom?

Soln.: **(a)** The electronic configurations of nitrogen and oxygen are as follows:

N : $1s^2$, $2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$

O : $1s^2$, $2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$

In N, *p*-orbital is half filled which is more stable. It requires more energy to remove an electron. Hence, the *IE* of nitrogen is higher than oxygen atom which has less stable structure.

(b) The electronic configurations of Mg and Al are as follows:

Mg : $1s^2$, $2s^2\ 2p^6\ 3s^2$

Al : $1s^2$, $2s^2\ 2p^6\ 3s^2\ 3p^1$

It is difficult to remove an electron from 3*s* in comparison to 3*p* (3*s* paired (stable) and 3*p* singly occupied). Hence, *IE* of Mg is higher than Al.

---

12. Why there are 2, 8 and 8 elements in first, second and third periods of periodic table respectively?

Soln.: In first period 1*s* is completed. Its capacity is of two electrons. In second period 2*s* 2*p* and in third period 3*s* 3*p* are completed. The capacity of these shells is of 8 electrons each. Thus, 2, 8 and 8 elements are present in first, second and third periods respectively.

13. Why is the electron affinities of halogens so high?

Soln.: The electron is easily accommodated as halogens acquire inert gas configuration after gaining the electron.

---

14. Why is potassium strongly electropositive element?

Soln.: The ionisation potential of potassium is low. Hence, it can lose electron easily. Therefore, it is strongly electropositive in nature.

---

15. **(a)** Between Na^+ and $F^{ñ}$ which has more electronegativity?

(b) Explain the nature of the bond in the following compounds.

($NaCl$, SiO_2, PH_3, HCl)

Soln.: **(a)** Na^+ ion has more electronegativity than $F^{ñ}$ ion, as the former is an electrophile having more tendency to attract electron pair.

(b) $NaCl \longrightarrow$ Ionic bond, as the electronegativity difference is, 3 ñ 0.9 = 2.1
$SiO_2 \longrightarrow$ Polar covalent bond, as the electronegativity difference is, 3.5 ñ 1.8 = 1.7
$PH_3 \longrightarrow$ Pure covalent bond, as the electronegativity difference is, zero.
$HCl \longrightarrow$ Polar covalent bond, as the electronegativity difference is, 3 ñ 2.1 = 0.9.

---

16. Explain, why electron affinity of fluorine is lower than chlorine?

Soln.: Due to smaller size of the valence shell of fluorine atom, seven electrons are much more crowded around the nucleus as compared to chlorine atom with bigger outermost shell (3*s*, 3*p*). So the interelectronic repulsion is greater in fluorine atom than in chlorine atom. Hence fluorine atom feels difficulty in entry of the new electron in its valence shell as compared to chlorine atom. Therefore electron affinity of fluorine is lower than chlorine. Further, as the bond dissociation energy of Cl_2 is higher than F_2, so the former has higher electron affinity.

---

17. **(a)** Explain why Mg has zero electron affinity?

(b) Which property of elements does not vary along the group?

Soln.: **(a)** Mg has fully-filled *s*-orbital arrangement in its valence shell ($3s^2$), which is already quite stable. Therefore, when an electron is added to its atom it goes to the valency shell of higher energy level (3*p*). In this process instead of release, energy is to be supplied. Hence the electron affinity value is ñve and roughly considered as zero.

(b) Along the group, valency of elements remains same, due to similar electronic configuration.

For example, the general outer electronic configuration for the members of group-2 is ns^2, so they have common valency of 2.

---

18. Explain that electron affinity of oxygen is smaller than sulphur.

Soln.: The valence shell of O atom is much smaller than the valence shell of S atom. Therefore 6 electrons in the valence shell of oxygen atom are much more crowded, than in the valence shell of sulphur atom. Hence oxygen atom feels a little bit difficulty in entry of an electron to its valence shell as compared to sulphur atom. So the electron affinity of oxygen is smaller than sulphur atom.

---

19. In a particular period, the noble gases have highest ionisation energy, although they have largest size, explain.

Soln.: Although the noble gases possess largest size for a particular period, but their ionisation energy values are quite high. Because in their atoms not only the valence orbitals are full-filled but also valence shell is saturated in some cases. So high energy is required to disturb their extra stability and hence they have highest ionisation energy.

---

20. **(a)** Arrange the following in increasing order of their electron affinity. (K, Ca, P, S)

(b) Explain, why Al and Ga have same atomic size although both belong to same group?

Soln.: **(a)** Along the period electron affinity value increases, but due to stable fully filled *s*-orbital arrangement in the valence shell of Ca atom, its electron affinity is zero. Therefore the increasing order of electron affinity value is ó

$$Ca < K < P < S$$

(b) $_{13}\text{Al} \longrightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^1$;

$_{31}\text{Ga} \longrightarrow 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^1$

From the electronic configuration of both we can see that Al does not contain any *d*-electron, but Ga contain 10 *d*-electrons in the inner shell. As the inner shell 3*d*-electrons have comparatively poor screening effect and the nuclear charge of Ga is higher than Al, so the effective nuclear charge increases. Therefore atomic radius of Ga decreases and the two radii (of Al and Ga) happens to be same, *i.e.* 1.26 ≈ although both belongs to same group.

---

21. **(a)** Why group-1 metals and their salts impart characteristic colour to flame?

(b) Why the electron affinity of noble gases is zero?

Soln.: Group-1 metals have very low ionisation energy for which their valency electrons are excited to higher energy level by absorbing heat energy from Bunsen flame. The excited electrons then return back to ground state, emit energy in the form of radiation. These radiations appear as different colours in flame. Therefore group-1 metals and their salts impart characteristic colours to flame.

(b) The noble gas elements possess extra stable fully filled orbitals and saturated shells in the valency shell of their atom. So they have no tendency to accept an electron. Even if an electron is added to their atom, then it enters next shell of higher energy level. Therefore in this process instead of release, energy is to be supplied. Hence electron affinity values of noble gases are ñve and roughly considered as zero.

---

22. **(a)** The atomic no. of three elements *A*, *B* and *C* are *Z* ñ 2, *Z* and *Z* + 1 respectively. Of these ë*B*í is inert gas, then predict:

(i) Which element has highest electronegativity?

(ii) Which element has highest ionisation energy?

(iii) What is the formula of the compound between ë*A*í and ë*C*í?

(iv) What is the nature of the bond between ë*A*í and ë*C*í?

(b) The successive ionisation energies for an element are 899, 1757, 14849, 21000 kJ/mole respectively. Predict the number of valence electrons in its atom.

Soln.: *B* (atomic no. = *Z*), an inert gas (group-18)

A (atomic no. = *Z* ñ 2), a member of group-16

C (atomic no. = *Z* + 1), a member of group-1.

(i) ë*A*í has highest electronegativity as it is a non-metal.

(ii) ë*B*í has highest ionisation energy, as it is an inert gas element.

(iii) The valency of ë*A*í is 2 and ë*C*í is 1. So the formula of the compound between ë*A*í and ë*C*í is C_2A.

(iv) The nature of the bond is either ionic or polar covalent, as one is electropositive (*C*) and other is electronegative (*A*).

(b) The number of valence electrons in its atom is 2, as there is sudden jump from 2nd ionisation energy value to 3rd ionisation energy value, which indicates that the removal of 3rd electron needs disturbance of stable noble gas configuration.

---

23. OF_2 is called as oxygen di-fluoride and not written as F_2O, where as Cl_2O is called as chlorine oxide and not written as OCl_2, explain.

Soln.: While writing the formula of a compound less electronegative element is written first, followed by more electronegative element and accordingly the name is written. As electronegativity values of Cl, O, and F are 3, 3.5 and 4 respectively. So the compound of oxygen and fluorine is written as OF_2, while compound of oxygen and chlorine is written as Cl_2O.

24. Ionisation potential and electron affinity of fluorine are 17.42 and 3.45 eV respectively. Calculate the electronegativity of fluorine.

Soln.: According to Mulliken equation

$$X = \frac{IP + EA}{5.6} \text{ when both } IP \text{ and } EA \text{ are taken in eV}$$

$$X_F = \frac{17.42 + 3.45}{5.6} = 3.726$$

---

25. The electron affinity of chlorine is 3.7 eV. How much energy in kcal is released when 2 g of chlorine is completely converted to Cl^- ion in a gaseous state?

Soln.: $Cl + e \longrightarrow Cl^- + 3.7 \text{ eV}$

35.5 3.7 × 23.06 kcal

∴ Energy released for conversion of 2 g gaseous chlorine into Cl^- ions

$$= \frac{3.7 \times 23.06}{35.5} \times 2 = 4.8 \text{ kcal}$$

EXERCISE

Multiple Choice Questions

1. The ionization potential of nitrogen is more than that of oxygen because of

(a) the greater attraction of the electrons by the nucleus

(b) the extra stability of the half filled *p*-orbitals

(c) the smaller size of nitrogen

(d) more penetration effect.

2. Which one of the following represents the electronic configuration of the most electropositive elements ?

(a) [He] $2s^1$ (b) [Xe] $6s^1$

(c) [He] $2s^2$ (d) [Xe] $2s^2$

3. The sizes of A, A^+ and $A^{\tilde{n}}$ follow the order

(a) $A^+ > A^{\tilde{n}} > A$ (b) $A^{\tilde{n}} > A^+ > A$

(c) $A^{\tilde{n}} > A > A^+$ (d) $A > A^{\tilde{n}} > A^+$

4. The maximum valency of halogen group with respect to oxygen is

(a) 1 (b) 5

(c) 6 (d) 7

5. The statement that is not 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) it helps to predict the relative ionicity of the bond between any two elements.

6. The electronic configuration of four elements are given below. Which elements does not belong to the same family as others?

(a) [Xe] $4f^{14}$ $5d^{10}$ $6s^2$ (b) [Kr] $4d^{10}$ $5s^2$

(c) [Ne] $3s^2$ $3p^5$ (d) [Ar] $3d^{10}$ $4s^2$

7. In a period, elements are arranged strictly in sequence of

(a) decreasing charges in the nucleus

(b) increasing charges in the nucleus

(c) constant charges in the nucleus

(d) equal charges in the nucleus.

8. Which of the following structures is associated with the biggest jump between the second and third ionization energies?

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

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

9. Elements *X*, *Y* and *Z* have atomic numbers 19,37 and 55 respectively. Which of the following statements is true about them?

(a) Their ionization potential would increase with increasing atomic number.

(b) *Y* would have an ionization potential between those of *X* and *Z*.

(c) *Z* would have the highest ionization potential.

(d) *Y* would have the highest ionization potential.

10. Al^{3+} has a lower ionic radius than Mg^{2+} ion because

(a) Mg atom has less number of neutrons than Al

(b) Al^{3+} has higher nuclear charge than Mg^{2+}

(c) their electronegativities are different

(d) Al has a lower ionization potential than Mg atom.

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

(a) the properties of the elements are the periodic function of their atomic numbers

(b) non-metallic elements are lesser in number than metallic elements

(c) the first ionization energies of elements along a period do not vary in a regular manner with increase in atomic number

(d) for transition elements the *d*-sub shells are filled with electrons monotonically with increase in atomic numbers.

12. Na^+, Mg^{2+}, Al^{3+} and Si^{4+} are isoelectronic. The order of their ionic size is

(a) $Na^+ > Mg^{2+} < Al^{3+} < Si^{4+}$

(b) $Na^+ < Mg^{2+} > Al^{3+} > Si^{4+}$

(c) $Na^+ > Mg^{2+} > Al^{3+} > Si^{4+}$

(d) $Na^+ < Mg^{2+} > Al^{3+} < Si^{4+}$

13. Mercury is the only metal which is liquid at 0∞C. This is due to its

(a) very high ionization energy and weak metallic bond

(b) low ionization potential

(c) high atomic weight

(d) high vapour pressure.

14. Aqueous solutions of two compounds *M* **ñ** O **ñ** H and *M′* **ñ**O **ñ** H have been prepared in two different beakers. If the electronegativity of *M* = 3.5, *M′* = 1.72, O = 3.0 and H = 2.1, then the solutions respectively are
(a) acidic, acidic (b) acidic, basic
(c) basic, basic (d) basic, acidic.

15. Consider the isoelectronic series : K^+, $S^{2ñ}$, $Cl^ñ$ and Ca^{2+}, the radii of the ions decrease as
(a) $Ca^{2+} > K^+ > Cl^ñ > S^{2ñ}$
(b) $Cl^ñ > S^{2ñ} > K^+ > Ca^{2+}$
(c) $S^{2ñ} > Cl^ñ > K^+ > Ca^{2+}$
(d) $K^+ > Ca^{2+} > S^{2ñ} > Cl^ñ$

16. The incorrect statement among the following is
(a) the first ionization potential of Al is less than the first ionization potential of Mg
(b) the second ionization potential of Mg is greater than the second ionization potential of Na
(c) the first ionization potential of Na is less than the first ionization potential of Mg
(d) the third ionization potential of Mg is greater than the third ionization potential of Al.

17. Set containing isoelectronic species is
(a) C_2^{2-}, NO^+, $CN^ñ$, O_2^{2+}
(b) CO, NO, O_2, CN
(c) CO_2, NO_2, O_2, N_2O_5
(d) CO, CO_2, NO, NO_2

18. Consider the following statements :
I. The radius of an anion is larger than that of the parent atom.
II. The ionization energy generally increases with increasing atomic number in a period.
III. The electronegativity of an element is the tendency of an isolated atom to attract an electron.
Which of the above statements is/are correct?
(a) I alone (b) II alone
(c) I and II (d) II and III

19. The electronic configuration of elements *A, B* and *C* are [He] $2s^1$, [Ne] $3s^1$ and [Ar] $4s^1$ respectively. Which one of the following order is correct for IE_1 (in kJ $mol^{ñ1}$) of *A, B* and *C* ?
(a) $A > B > C$ (b) $C > B > A$
(c) $B > C > A$ (d) $C > A > B$

20. Eka-aluminium and Eka-silicon are known as
(a) gallium and germanium
(b) aluminium and silicon
(c) iron and sulphur
(d) proton and silicon

21. In any period, the valency of an element with respect to oxygen
(a) increases one by one from IA to VII A
(b) decreases one by one from IA to VII A
(c) increases one by one from IA to IV A and then decreases from V A to VII A one by one
(d) decreases one by one from IA to IV A and then increases from V A to VII A one by one.

22. In the ground state, an element has 13 electrons in its M-shell. The element is
(a) manganese (b) chromium
(c) nickel (d) iron.

23. The IUPAC name of the element with atomic number Z = 109 is
(a) Meitnesim (b) Reentgenium
(c) Copernicim (d) Hassim.

24. Pd has exceptional outer electronic configuration as $4d^{10}$ $5s^0$. It belongs to
(a) 4th period, group 11
(b) 5th period, group 10
(c) 6th period, group 9
(d) 3rd period, group 16

25. The ionization enthalpies of Li and sodium are 520 kJ $mol^{ñ1}$ and 495 kJ $mol^{ñ1}$ respectively. The energy required to convert all the atoms present in 7 mg of Li vapours and 23 mg of sodium vapours of their respective gaseous cations respectively are
(a) 52 J, 49.5 J (b) 520 J, 495 J
(c) 49.5 J, 52 J (d) 495 J, 520 J

26. Identify the least stable ion amongst the following.
(a) $Li^ñ$ (b) $Be^ñ$
(c) $B^ñ$ (d) $C^ñ$

27. Element with atomic number 56 belongs to which block?
(a) *s* (b) *p*
(c) *d* (d) *f*

28. Which of the following is not paramagnetic?
(a) $S^{2ñ}$ (b) $N_2^ñ$
(c) $O_2^ñ$ (d) NO

29. The first ionisation potential of Mg, Al, P and S follows the order
(a) Mg < Al < P < S (b) Al < Mg < P < S
(c) Al < Mg < S < P (d) Mg < Al < S < P

30. The atomic radius increases as we move down a group because
(a) effective nuclear charge increases
(b) atomic mass increases
(c) additive electrons are accommodated in new electronic levels
(d) atomic number increases.

31. The correct order of 1st ionisation potential among following elements Be, B, C, N, O is
(a) $B < Be < C < O < N$
(b) $B < Be < C < N < O$
(c) $Be < B < C < N < O$
(d) $Be < B < C < O < N$

32. The set representing the correct order of first ionization potential is
(a) $K > Na > Li$ (b) $Be > Mg > Ca$
(c) $B > C > N$ (d) $Ge > Si > C$

33. The correct order of radii is
(a) $N < Be < B$ (b) $F^{ñ} < O^{2ñ} < N^{3ñ}$
(c) $Na < Li < K$ (d) $Fe^{3+} < Fe^{2+} < Fe^{4+}$

34. The correct statement about *d*-block elements is
(a) they are all metals
(b) they show variable valency
(c) they form coloured ions and complex salts
(d) all the above statements are correct.

35. Which of the following is the smallest cation?
(a) Na^+ (b) Mg^{2+}
(c) Ca^{2+} (d) Al^{3+}

36. What is the maximum number of electrons which can be accommodated in an atom in which the highest principal quantum number is 4?
(a) 10 (b) 18
(c) 36 (d) 54

37. The first ionisation energy of Na, Mg, Al and Si are in the order
(a) $Na < Mg > Al < Si$
(b) $Na > Mg > Al > Si$
(c) $Na > Mg < Al > Si$
(d) $Na < Mg < Al < Si$

38. The ionic radii of $N^{3ñ}$, $O^{2ñ}$, $F^{ñ}$ and Na^+ follows the order
(a) $N^{3ñ} > O^{2ñ} > F^{ñ} > Na^+$
(b) $N^{3ñ} > Na^+ > O^{2ñ} > F^{ñ}$
(c) $Na^+ > O^{2ñ} > N^{3ñ} > F^{ñ}$
(d) $O^{2ñ} > F^{ñ} > Na^+ > N^{3ñ}$

39. The elements in which 4*f*-orbitals are progressively filled in are called
(a) actinides
(b) transition elements
(c) lanthanides (d) halogens.

40. Which of the following pairs is isoelectronic?
(a) Ar and Cl (b) Na^+ and Ne
(c) Na^+ and Mg (d) Mg and Ne

41. Which of the following has least electron affinity?
(a) Oxygen (b) Argon
(c) Nitrogen (d) Boron

42. The first ionization energy will be maximum for
(a) uranium (b) hydrogen
(c) lithium (d) iron.

43. The bond length in LiF will be
(a) less than that of NaF
(b) equal to that of KF
(c) more than that of KF
(d) equal to that of NaF.

44. Electronic configuration of most electronegative elements is
(a) $1s^2\ 2s^2\ 2p^6\ 3s^1$
(b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
(c) $1s^2\ 2s^2\ 2p^5$
(d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

45. Two elements whose electronegativities are 1.2 and 3.0, the bond formed between them would be
(a) ionic (b) covalent
(c) coordinate (d) metallic.

46. The electron affinities of halogens are :
F = 332, Cl = 349, Br = 324,
I = 295 kJ $mol^{ñ1}$.
The higher value for Cl as compared to that of F is due to
(a) higher atomic radius of F
(b) smaller electronegativity of F
(c) weaker electron-electron repulsion in Cl
(d) more vacant *p*-subshell in Cl.

47. Point out the wrong statement, in a given period of the periodic table, the *s*-block elements have, in general, a lower value of
(a) electronegativity (b) atomic radius
(c) ionization energy (d) electron affinity.

48. Correct order of electron affinity of the halogen atoms are
(a) $F < Cl < Br > I$ (b) $F < Cl \sim Br > I$
(c) $F > Cl > Br > I$ (d) $F < Cl > Br > I$

49. Two elements *X* and *Y* contain only one electron in the outer level. Element *X* is reactive and loses electron easily while element *Y* is relatively unreactive and non-corrosive. The elements *X* and *Y* are, respectively
(a) Cs and Li (b) Rb and Na
(c) Li and Cu (d) Ag and Au

50. The IUPAC name of the element, having atomic number 104, is
(a) unnilquadium (b) unnilennium
(c) quadnilunium (d) quadunnilium

51. Mark out the incorrect match: (electronic configuration ñ nature of element)
(a) $3s^2 3p^6$: non-metal
(b) $5s^2 4d^{10} 5p^5$: non-metal
(c) $4s^2 3d^7$: metals
(d) $3s^2 3p^2$: non-metal

52. An atom having electronic configuration as $6s^1\ 4f^{14}\ 5d^{10}$, must be
(a) an alkali metal
(b) an alkaline earth metal
(c) a coinage metal
(d) a lanthanide

53. Elements *A, B, C, D* and *E* have the following electronic configurations
A : [Ar] $3d^{10}4s^1$
B : [Ar] $3d^{10}4s^2 4p^6 5s^1$
C : [Ar] $3d^{10}4s^2 4p^6 4d^{10} 5s^1$
D : [Ar] $3d^{10}4s^2 4p^6 4d^5 5s^1$
Which two elements fall into the same group?
(a) *B* and *C* (b) *A* and *C*
(c) *B* and *D* (d) *A* and *B*

54. Following the Pauling equation, the extra bond energy (in kcal molñ1) for HñCl bond is
[Given : electronegativity : H = 2.1, Cl = 3.0]
(a) 18.72 (b) 4.326
(c) 21.52 (d) 11.21

55. Which of the following is the correct decreasing order of radii?
(a) Al^{3+} Na^+ Ne $F^{ñ}$ $N^{3ñ}$
(b) Ne Al^{3+} Na^+ $F^{ñ}$ $N^{3ñ}$
(c) $N^{3ñ}$ $F^{ñ}$ Ne Na^+ Al^{3+}
(d) Ne $N^{3ñ}$ $F^{ñ}$ Na^+ Al^{3+}

56. For one of the element various successive ionisation energies are given below :

Ionisation energy (kJ mol^{-1})	1st	2nd	3rd	4th	5th
	577.5	1810	2750	11580	14820

The element is
(a) magnesium (b) aluminium
(c) silicon (d) phosphorus

57. The correct decreasing order of ionisation energy is
(a) F B Be H (b) Be B H F
(c) H F B Be (d) F H Be B

58. Which of the following involves maximum amount of energy?
(a) $N + e^{ñ} \rightarrow N^{ñ}$ (b) $O + e^{ñ} \rightarrow O^{ñ}$
(c) $O^{ñ} + e^{ñ} \rightarrow O^{2ñ}$ (d) $F + e^{ñ} \rightarrow F^{ñ}$

59. The correct increasing order of electron affinity is
(a) Li Be B C (b) Si P S Cl
(c) Cl F Br I (d) Mg Al Na Si

60. In general the first ionisation energy decreases in a regular way on descending the main groups. A departure from this trend is observed in Group 13 as follows :

Element	Ionisation energy in kJ mol^{-1}		
	1st	2nd	3rd
B	801	2427	3659
Al	577	1816	2744
Ga	579	1979	2962
In	558	1820	2704

Gallium has higher ionisation energy in comparison to Al because
(a) gallium is harder than aluminium
(b) gallium is preceded by ten transition elements, where 3*d* shell is being filled which makes Ga larger than it would be
(c) gallium is preceded by ten transition elements, where 3*d* shell is being filled which makes Ga smaller than it would be
(d) effective nuclear charge for Ga is much higher than Al

61. Alkali metals are soluble in liquid NH_3. As the concentration of metal increases, solution turns blue to bronze. It reflects the change in magnetic property of the solution
(a) diamagnetic to paramagnetic
(b) paramagnetic to diamagnetic
(c) weak to intense paramagnetic
(d) no change in magnetic property

62. On combustion in excess of air, lithium forms mainly the oxide, sodium forms peroxide while potassium or rubidium or cesium form superoxides. It suggests that as cationic size increases the
(a) anionic polarizability increases, thus stability decreases

(b) anionic size increases, polarizability increases, covalency increases, thus, stability increases
(c) anionic size increases, coordination number increases, lattice energy increases, thus stability increases
(d) anionic size decreases, stability increases

63. The second ionisation energy (IE_{II}) for fluorine is more than nitrogen because of
(a) higher effective nuclear charge
(b) stable electronic configuration
(c) higher oxidation state
(d) low screening effect constant

64. Which of the following nitrates does not give an oxide on heating?
(a) $Cu(NO_3)_2$ (b) $Ca(NO_3)_2$
(c) $LiNO_3$ (d) KNO_3

65. Alkaline earth metal carbonates are sparingly soluble. On moving down the group solubility of alkaline earth metal carbonates
(a) increases (b) decreases
(c) remains constant
(d) none of these

66. On increasing temperature, the ionic character of $AlCl_3$
(a) increases (b) decreases
(c) is not affected
(d) first increases then decreases

67. Mark out the correct increasing order of radius.
(a) As^{3ñ} < Br^{ñ} < K^+ < Mg^{2+}
(b) Mg^{2+} < K^+ < Br^{ñ} < As^{3ñ}
(c) Mg^{2+} < K^+ < As^{3ñ} < Br^{ñ}
(d) K^+ < Mg^{2+} < Br^{ñ} < As^{3ñ}

68. How many elements would be in the second period of the periodic table if the spin quantum numbers had the value $-\frac{1}{2}, 0 \text{ or } +\frac{1}{2}$?
(a) 8 only (b) 10 only
(c) 12 only (d) 13 only

69. The bond dissociation energies of H—H, C—C and C—H bonds respectively are 104.2, 83.1 and 98.8 kcal mol^{ñ1}. The electronegativity of carbon is
(a) 2.50 (b) 2.51
(c) 2.53 (d) 2.52

70. The second electron affinities (in kJ mol^{ñ1}) of oxygen and sulphur respectively are
(a) +590, +780 (b) ñ590, + 780
(c) ñ780, +590 (d) ñ780, ñ590

71. An element has successive ionization energies as 940 (first), 2080, 3090, 4140, 7030, 7870 16000 and 19500 kJ mol^{ñ1}. To which group of the periodic table does this element belong?
(a) 16 (b) 15
(c) 14 (d) 17

72. If the ionic radii of K^+ and F^{ñ} are nearly the same (*i.e.* 1.34 ≈), then the atomic radii of K and F respectively are
(a) 1.96 ≈, 0.72 ≈ (b) 0.72 ≈, 1.96 ≈
(c) 1.34 ≈, 1.34 ≈ (d) 1.96 ≈, 1.34 ≈

73. Which of the following elements with atomic numbers 25, 30, 48 and 80 has the highest vapour pressure at room temperature?
(a) $Z = 80$ (b) $Z = 48$
(c) $Z = 30$ (d) $Z = 25$

74. In which of the following pairs, both the species have nearly the same size?
(a) Fe, Co (b) Li^+, Mg^{2+}
(c) Rb^+, O^{2ñ} (d) All the three

75. Which of the following have same number of electrons in the outermost orbit?
(a) As, Bi (b) N, O
(c) Pb, Sb (d) P, Ca

76. Which one of the following pairs of ions have the same electronic configuration?
(a) Cr^{3+}, Fe^{3+} (b) Fe^{3+}, Mn^{2+}
(c) Fe^{3+}, Co^{3+} (d) Sc^{3+}, Cr^{3+}

77. The radii of F, F^{ñ}, O and O^{2ñ} are in the order of
(a) O^{2ñ} > O > F^{ñ} > F
(b) F^{ñ} > O^{2ñ} > F > O
(c) O^{2ñ} > F^{ñ} > F > O
(d) O^{2ñ} > F^{ñ} > O > F

78. Which of the following represents noble gas configuration?
(a) $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^6$
(b) $1s^22s^22p^63s^23p^63d^{10}4s^24f^{14}5s^25p^65d^1$
(c) $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^65d^16s^2$
(d) $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^65d^66s^2$

79. Which of the following gaseous atoms has the highest value of *IE*?
(a) Si (b) P
(c) Mg (d) Al

80. Which of the following isoelectronic species has the smallest atomic radius:
(a) $F^{ñ}$ (b) $O^{2ñ}$
(c) $N^{3ñ}$ (d) Ne

81. The first (IE_1) and second (IE_2) ionization energies (kJ/mol) of a few elements designated by Roman numerals are given below. Which of these would be an alkali metal?

	IE_1	IE_2
(a) I	2372	5251
(b) II	520	7300
(c) III	900	1760
(d) IV	1680	3380

82. Which of the following is a stable ion?
(a) $Se^{2ñ}$ (b) Ba^{3+}
(c) Rb^{2+} (d) Sn^{4+}

83. Which of the atoms having the following electronic configuration will have the highest first ionization energy?
(a) $[Ne]3s^2\ 3p^4$ (b) $[Ne]3s^2\ 3p^3$
(c) $[Ne]3s^2\ 3p^2$ (d) $[He]2s^2\ 2p^3$

84. The second ionization potential of an element M is the energy required to
(a) remove one mole of electrons from one mole of monovalent gaseous cations of the element
(b) remove one mole of electrons from one mole of gaseous anions
(c) remove one mole of electrons from one mole of gaseous cations of the element
(d) remove 2 moles of electrons from one mole of gaseous atoms.

85. Electronic configuration of four neutral atoms is given below. Which of the corresponding elements would not be paramagnetic?
(a) $1s^2\ 2s^2\ 2p^5$ (b) $1s^2\ 2s^2\ 2p^1$
(c) $1s^2\ 2s^2\ 2p^6$ (d) $1s^2\ 2s^2\ 2p^4$

86. Electronic configuration $1s^2\ 2p^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$ corresponds to
(a) O (b) Cr
(c) N (d) F

87. The element with highest electron affinity is
(a) Br (b) Cl
(c) F (d) I

88. The hydride ion is isoelectronic with
(a) He (b) He^+
(c) Li (d) Be

89. The most common lanthanide is
(a) samarium (b) cerium
(c) lanthanum (d) plutonium

90. Arrange in the increasing order of atomic radii of the following elements O, C, F, Cl, Br
(a) F, Cl, Br, O, C
(b) F, C, O, Cl, Br
(c) F, O, C, Cl, Br
(d) C, O, F, Cl, Br

91. Which of the following transitions involves maximum amount of energy?
(a) $M^+_{(g)} \longrightarrow M^{2+}_{(g)}$
(b) $M^{ñ}_{(g)} \longrightarrow M^+_{(g)}$
(c) $M^{ñ}_{(g)} \longrightarrow M_{(g)}$
(d) $M^{2+}_{(g)} \longrightarrow M^{3+}_{(g)}$

92. Which one of the following is a metalloid?
(a) Ge (b) S
(c) C (d) Pb

93. The telluric helix was given by
(a) Lothar Meyer (b) Mendeleev
(c) Newlands (d) De Chancourtois

94. Which of the following structures is associated with the biggest jump between the second and third ionization energies?
(a) $1s^2\ 2s^2\ 2p^6\ 3s^2$ (b) $1s^2\ 2s^2\ 2p^6\ 3s^1$
(c) $1s^2\ 2s^2\ 2p^2$ (d) $1s^2\ 2s^2\ 2p^1$

95. The lightest metal in the periodic table is
(a) Ca (b) Hg
(c) Na (d) Li

96. The element with the highest first ionization potential is?
(a) Nitrogen (b) Carbon
(c) Boron (d) Oxygen

97. The higher oxide of an element (E) has the formula EO_3. Its hydride contains 2.47% hydrogen, the element is
(a) Te (b) Se
(c) S (d) Si

98. The incorrect statement among the following is
(a) the first ionization potential of Na is less than the first ionization potential of Mg
(b) the second ionization potential of Mg is greater than the second ionization potential of Na
(c) the first ionization potential of Al is less than the first ionization potential of Mg
(d) the third ionization potential of Mg is greater than the third ionization potential of Al.

99. Which one of the following has largest size?
(a) $O^{2ñ}$ (b) O^{+}
(c) O (d) None of these

100. Which one of the following is smallest in size?
(a) $F^{ñ}$ (b) $O^{2ñ}$
(c) $N^{3ñ}$ (d) Na^{+}

101. The electronegativity of the following elements increases in the order
(a) Si, P, C, N (b) N, Si, C, P
(c) C, N, Si, P (d) P, Si, N, C

102. The ionization energy and electron affinity of an element are 13.0 eV and 3.8 eV respectively. Its electronegativity on the Pauling scale is
(a) 3.0 (b) 3.5
(c) 4.0 (d) 2.8

103. The chemical elements are arranged in the order of increasing electronegativities in the sequence:
(a) P, Si, Br, Se, Cl, O
(b) Si, P, Br, Se, Cl, O
(c) Si, P, Se, Br, Cl, O
(d) Se, Si, P, Br, Cl, O

104. In which of the following species, the size of the first species is not more than the second?
(a) Na^{+}, $F^{ñ}$ (b) Fe^{2+}, Fe^{3+}
(c) Li, F (d) S, O

105. An element with atomic number 106 has been discovered recently. Which of the following electronic configurations will it possess?
(a) [Rn] $5f^{14}\ 6d^{6}\ 7s^{0}$
(b) [Rn] $5f^{14}\ 6d^{5}\ 7s^{1}$
(c) [Rn] $5f^{14}\ 6d^{4}\ 7s^{2}$
(d) [Rn] $5f^{14}\ 6d^{1}\ 7s^{2}\ 7p^{3}$

106. Without consulting the periodic table, select from each of the following sets, the elements belonging to same group of the periodic table.
(a) Z = 31, 13, 81, 50
(b) Z = 12, 38, 4, 88
(c) Z = 11, 3, 20, 37
(d) Z = 21, 39, 90, 57

107. The atomic numbers of other elements which lie in the same group as the tenth element in the periodic table are
(a) 18, 32, 54, 86 (b) 8, 18, 36, 84
(c) 2, 18, 30, 36 (d) 2, 18, 36, 54

108. *f*-Block elements are called inner transition elements because
(a) The last electron enters into the *f*-orbital of penultimate shell
(b) They have been taken out of the transition elements
(c) They have properties similar to those of transition elements
(d) The last electron enters into the *f*-orbital of the anti-penultimate shell.

109. Electronic configuration of most electronegative element is
(a) $1s^{2}\ 2s^{2}\ 2p^{4}\ 3s^{1}$ (b) $1s^{2}\ 2s^{2}\ 2p^{4}\ 3s^{2}\ 3p^{5}$
(c) $1s^{2}\ 2s^{2}\ 2p^{5}$ (d) $1s^{2}\ 2s^{2}\ 2p^{6}\ 3s^{1}\ 3p^{6}$

110. The number of *d*-electrons in Ni (at. no. = 28) is equal to that of the
(a) *s* and *p* electrons in $F^{ñ}$
(b) *p* electrons in Ar (at. no. = 18)
(c) *d* electrons in Ni^{2+}
(d) total electrons in N (at. no. = 7)

Assertion and Reason AIIMS

Directions : *In the following questions, a statement of assertion (A) is followed by a statement of reason (R). Mark the correct choice as :*

(a) *If both assertion and reason are true and reason is the correct explanation of assertion*
(b) *If both assertion and reason are true but reason is not the correct explanation of assertion*
(c) *If assertion is true but reason is false*
(d) *If both assertion and reason are false.*

1. **Assertion :** Isoelectronic species have the same number of electrons but different radii.
Reason : Higher the charge, smaller is the ion.

2. **Assertion :** Successive ionisation energies are always higher.
Reason : With every loss of electron, the effective nuclear charge of the species increases.

3. **Assertion :** Sulphur atom has higher electron affinity than oxygen.
Reason : Oxygen is more electronegative than sulphur, thatís why it can hold electrons better.

4. **Assertion :** Fe_2O_3 is more acidic than FeO.
Reason : Higher the oxidation state, higher the electronegativity, thus non-metallic characteristic is higher.

5. **Assertion :** Atomic size of silver is almost equal to that of gold.
Reason : *d*-subshell has low penetration power, and produce poor shielding.

6. **Assertion :** All transition elements are metallic.
Reason : Transition metals have only two electrons in their outer energy level.

7. **Assertion :** Lanthanoid elements are arranged in increasing order of atomic number in the same period, but exhibit close similarities in chemical properties.
Reason : The two outer energy levels for each lanthanoid element have the same electronic configuration.

8. **Assertion :** Known elements may contain as many as 32 electrons in an energy level but only *s* and *p* sub-level electrons are considered for the octet rule.
Reason : For any atom, electrons present in *s* and *p* subshell assume greater stability.

9. **Assertion :** Alkaline earth metal carbonates ($BeCO_3$) are soluble in water.
Reason : Ionic compounds are fairly soluble in water.

10. **Assertion :** Increasing order of ionic radii is : $F^{ñ} < O^{2ñ} < N^{3ñ}$.
Reason : All these species are isoelectronic species.

11. **Assertion :** If the five successive ionisation energies of an element are 700, 2145, 3478, 30450 and 38748 kJ $mol^{ñ1}$ respectively the number of valence electrons is four.
Reason : Ionisation energy decreases abruptly at fourth ionisation energy.

12. **Assertion :** The order of atomic radii is :
$Na > Mg > Al > Si$.
Reason : Atomic radii decreases along a period.

13. **Assertion :** Out of the following compounds, CsF, CsI, LiF and LiI, LiI has the lowest value of ratio of anionic size to cationic size.
Reason : For lowest value of ratio of anionic size to cationic size, anion and cation should be biggest.

14. **Assertion :** For an element, relative stabilities of various oxidation states can be predicted on the basis of their successive ionisation enthalpies.
Reason : Ionisation enthalpy depends on nuclear charge, atomic size etc.

15. **Assertion :** The element with electronic configuration [Xe] $4f^1\ 5d^1\ 6s^2$ is a *f* block element.
Reason : The last electron enters the *f* orbital.

16. **Assertion :** If first and second ionisation potentials of helium atoms are 24.58 eV and 54.4 eV respectively, 762 kJ energy is required to convert 0.1 mole of He atoms into He^{2+} ions.
Reason : Total energy needed to convert one mole of He atoms into He^{2+} ions is 7620.78 kJ.

17. **Assertion :** Increasing order of metallic character is :
$P < Si < Be < Mg < Na$
Reason : Metallic character increases along a period and decreases down a group.

18. **Assertion :** If the electron affinity of bromine is 3.48 eV, 8.02 kcal energy is released when 8 g of bromine is completely converted to $Br^{ñ}$ ions in the gaseous state.
Reason : 8 g of bromine is equal to 0.1 mole of bromine.

19. **Assertion :** Decreasing order of van der Waals radii is :
$Cl > N > O > H$.
Reason : van der Waals radii increases as the number of energy levels increases and decreases as nuclear charge increases.

20. **Assertion :** First ionisation enthalpy of Mg is higher than that of sodium whereas the second ionisation enthalpy of Na is much higher than Mg.
Reason : Both are 2nd period elements.

MCQs NEET / AIPMT AIIMS

Multiple Choice Questions

1. Which one of the following arrangements represents the correct order of least negative to most negative electron gain enthalpy for C, Ca, Al, F and O?

'`( @k; B` ; N ; B ; E
'a(@k; N ; B ; B` ; E
'b(B ; E ; N ; @k; B`
'c(B` ; @k; B ; N ; E

(NEET Karnataka 2013)

2. The first ionisation enthalpy of Na, Mg and Si are 496, 737, 776 kJ/mol respectively. What will be the first ionisation enthalpy potential of Al in kJ/mol?

(a) > 766 kJ/mol (b) > 496 and < 737 kJ
(c) > 737 and < 766 kJ/mol
(d) > 496 kJ/mol **(AIIMS 2013)**

3. Which of the following orders of ionic radii is correctly represented?

(a) $H^{ñ} > H^+ > H$ (b) $Na^+ > F^{ñ} > O^{2ñ}$
(c) $F^{ñ} > O^{2ñ} > Na^+$ (d) $Al^{3+} > Mg^{2+} > N^{3ñ}$

(AIPMT 2014)

4. Which is correct regarding size of atom?

(a) N < O (b) B < Ne
(c) V > Ti (d) Na > K **(AIIMS 2014)**

5. The species Ar, K^+ and Ca^{2+} contain the same number of electrons. In which order do their radii increase?

(a) $Ca^{2+} < K^+ < Ar$ (b) $K^+ < Ar < Ca^{2+}$
(c) $Ar < K^+ < Ca^{2+}$ (d) $Ca^{2+} < Ar < K^+$

(AIPMT 2015, Cancelled)

6. Which of the following elements will have the highest electron affinity ?

(a) Chlorine (b) Nitrogen
(c) Phosphorus (d) Fluorine

(AIIMS 2015)

7. In which of the following options the order of arrangement does not agree with the variation of property indicated against it?

(a) I < Br < Cl < F (increasing electron gain enthalpy)
(b) Li < Na < K < Rb (increasing metallic radius)
(c) $Al^{3+} < Mg^{2+} < Na^+ < F^{ñ}$ (increasing ionic size)
(d) B < C < N < O (increasing first ionisation enthalpy) **(NEET-I 2016)**

8. An element (*X*) belongs to fourth period and fifteenth group of the periodic table. Which one of the following is true regarding the outer electronic configuration of (*X*)? It has

(a) partially filled *d* orbitals and completely filled *s* orbital
(b) completely filled *s* orbital and completely filled *p* orbitals
(c) completely filled *s* orbital and half-filled *p* orbitals
(d) half-filled *d* orbitals and completely filled *s* orbital. **(AIIMS 2016)**

9. The element $Z = 114$ has been discovered recently. It will belong to which of the following family/group and electronic configuration?

(a) Carbon family, [Rn] $5f^{14}\ 6d^{10}\ 7s^2\ 7p^2$
(b) Oxygen family, [Rn] $5f^{14}\ 6d^{10}\ 7s^2\ 7p^4$
(c) Nitrogen family, [Rn] $5f^{14}\ 6d^{10}\ 7s^2\ 7p^6$
(d) Halogen family, [Rn] $5f^{14}\ 6d^{10}\ 7s^2\ 7p^5$

(NEET 2017)

Assertion and Reason

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(c) *If assertion is true but reason is false*
(d) *If both assertion and reason are false.*

10. **Assertion :** For hydrogen like species, energy of an electron in a particular orbit increases with increase in value of *Z*.
Reason : Electronegativity decreases across a period.

(AIIMS 2016)

CHAPTER 3

Classification of Elements and Periodicity in Properties

Multiple Choice Questions

1. **(b) :** Extra stability of the exactly half-filled *p*-orbitals.

2. **(b) :** Elements (a) and (b) are alkali metals whereas (c) and (d) are alkaline earth metals. Amongst alkali and alkaline earth metals, alkali metal with the biggest size is the most electropositive element.

3. **(c) :** $A^{ñ} > A > A^{+}$.

4. **(d) :** Seven as in Cl_2O_7, I_2O_7, etc.

5. **(b) :** It does not help to predict the stable valency states of the elements.

6. **(c) :** Elements (a), (b) and (d) belong to the same group since each one of them has two electrons in the valence shell. In contrast, element (c) has seven electrons in the valence shell and hence does not lie in the same group in which elements (a), (b) and (d) lie.

7. **(b) :** Increasing charges in the nucleus (*i.e.* atomic number).

8. **(c) :** Biggest jump between IE_2 and IE_3 means that the element has two valence electrons, *i.e.* $1s^2\ 2s^2\ 2p^6\ 3s^2$.

9. **(b) :** Elements *X, Y* and *Z* with atomic numbers 19, 37, 55 lie in group 1 (alkali metals). Within a group, *IE* decreases from top to bottom. Therefore, *IE* of *Y* could be between those of *X* and Z.

10. **(b) :** Al^{3+} has a higher nuclear charge (+13) than Mg^{2+} (+12).

11. **(d) :** For transition elements, *d*-sub shells are not filled with electrons monotonically (one at a time) because of extra stability of d^5 and d^{10} configurations. Besides these, there are many other exceptions, *e.g.* Nb, Mo, Ru, Rh, Pd etc.

12. **(c) :** Amongst isoelectronic ions, the size of the cation decreases as the magnitude of the charge increases.

13. **(a) :** Due to very high ionization energy and weak metallic bonding, mercury is a liquid at room temperature.

14. **(b) :** In the compound *M* **ñ** O **ñ** H, if *IE* or E.N. of *M* is low, the compound will act as a base and if the *IE* or E.N. of *M* is high, the compound will behave as an acid. Therefore, *M* **ñ** O **ñ** H will act as an acid as E.N. of *M* is high (3.5) and *M′* **ñ** O **ñ** H will act as a base as E.N. of *M′* is low (1.72).

15. **(c) :** Amongst isoelectronic ions, ionic, radii of anions is more than that of cations. Further size of the anion increases with increase in ñve charge and size of cation decreases with increase in +ve charge.

16. **(b) :** IE_2 of Mg is lower than that of Na because in case of Mg^{2+}, a 3*s*-electron has to be removed, while in case of Na^{+}, an electron from the stable inert gas configuration (Neon) has to be removed.

17. **(a) :** C_2^{2-} (12 + 2), NO^{+} (7 + 8 ñ 1), $CN^{ñ}$ (6 + 7 + 1) and O_2^{2+} (16 ñ 2) are all isoelectronic species.

18. **(c) :** Size of an anion is larger than the corresponding atom and the ionization energy generally increases as the size decreases with increasing atomic number in a period.

19. **(a) :** Ionization energy decreases on moving from top to bottom in a group.

20. **(a) :** Eka-aluminium is known as gallium and eka-silicon is known as germanium because gallium is similar in properties to aluminium and germanium is similar in properties to silicon.

21. **(a) :** In any period, the valency of an element with respect to oxygen increases one by one from group I A (1) to VII A (17) of the periodic table.

22. **(b) :** 13 electrons in M (3rd) shell means $3s^2\ 3p^6\ 3d^5$. Hence complete configuration will be $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$ (at. no. = 24 *i.e.* Cr.

23. **(a)**

24. **(b) :** The exceptional outer electronic configuration of Pd suggests that, in principle its configuration should have been $4d^8\ 5s^2$. As such it lies in 5th period and group 10.

25. **(b) :** No. of moles of Li = $\frac{7}{1000 \times 7} = 10^{-3}$

    No. of moles of Na = $\frac{23}{1000 \times 23} = 10^{-3}$

∴ The amounts of energies required for $10^{ñ3}$ mole each of Li and Na are 520 kJ ◊ $10^{ñ3}$ and 495 kJ ◊ $10^{ñ3}$ or 520 J and 495 J respectively.

26. **(b) :** Be has fully filled 2*s*-subshell ($2s^2$) and therefore shows little tendency to accept an extra electron to form $Be^{ñ}$.

27. **(a)**

28. **(a) :** $S^{2ñ}$ (16 + 2 = 18 electrons) has 2, 8, 8 electronic configuration where all the electrons are paired up and shows diamagnetism.

29. **(c) :** Because of extra stability of exactly half-filled electronic configuration of P its IE_1 is higher than that of S. Further the IE_1 of Mg is higher than that of Al because in case of Mg, electron is to be removed from a 3*s*-orbital while in case of Al, the electron is to be removed from a 3*p*-orbital. Thus, option (c), *i.e.* Al < Mg < S < P is correct.

30. **(c) :** As the number of shells increases, atomic radius increases accordingly.

31. **(a) :** Because of extra stability of exactly half-filled electronic configuration of N, its IE_1 is higher than that of O. Further because of higher nuclear charge, IE_1 of C is higher than that of Be and B. Amongst Be and B, the IE_1 of Be is higher than that of *B* because in case of Be a 2*s*-electron is to be removed while in case of B a 2*p*-electron is to be removed. Thus, the overall order is (a) *i.e.* B < Be < C < O < N.

32. **(b) :** Within a group IE_1 decreases from top to bottom.

33. **(b)** **34.** **(d)** **35.** **(d)**

36. **(c) :** $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6 = 36$.

37. **(a)** **38.** **(a)** **39.** **(c)**

40. **(b) :** Both Na^+ and Ne have 10 electrons each.

41. **(b) :** Inert gases have zero electron-affinity.

42. **(b) :** Being smallest, hydrogen has the maximum IE_1.

43. **(a) :** Due to smaller size of Li than Na or K, LiF has smaller bond length than NaF and KF.

44. **(c) :** F ($1s^2\,2s^2\,2p^5$) is the most electronegative element.

45. **(a) :** When the electronegativity difference is more than 1.7, the bond formed will be ionic in nature.
(3.0 ñ 1.2) = 1.8 which is greater than 1.7

46. **(c)**

47. **(b):** In any period, *s*-block elements have the largest size.

48. **(d) :** Fluorine has unexpectedly less negative electron gain enthalpy than chlorine. Therefore, chlorine has the highest electron gain enthalpy in this group.
It is due to very small size of fluorine atom. As a result, there are strong interelectronic repulsions in the relatively small 2*p* subshell of fluorine and thus, the incoming electron does not feel much attraction. Thus, electron affinity among halogens varies as

F < Cl > Br > I

Chlorine has the highest electron affinity in the periodic table.

49. **(c) :** Alkali metals and coinage elements, both have one valence electron, elements of the same group I (A and B). Alkali metals are highly reactive while coinage elements are very less reactive.

50. **(a) :** un nil quad ium
1 0 4

51. **(d) :** $3s^2\,3p^2$, 3rd period, Group 14.
Thatís silicon ñ Metalloid.

52. **(c) :** $6s^1 4f^{14} 5d^{10}$ ñ Au (Gold) ñ Coinage metal

53. **(b) :** *A* : [Ar] $3d^{10}4s^1$: Cu
C : [Ar] $3d^{10}4s^24p^64d^{10}5s^1$: [Kr] $4d^{10}5s^1$: Ag

54. **(a) :** $\Delta x = 0.208\sqrt{\Delta}$
Δ = Extra bond energy

$$\Delta = \left(\frac{\Delta x}{0.208}\right)^2 = \left(\frac{0.9}{0.208}\right)^2 = (4.326)^2 = 18.72$$

55. **(c) :** Higher the nuclear charge, higher the attraction, smaller the size. (For isoelectronic species)

56. **(b) :** $\left.\begin{array}{l}\text{3rd ionisation energy}\\ \text{4th ionisation energy}\end{array}\right\} = \begin{array}{c}2750\\ 11580\end{array}\ \text{kJ/mol}$

4th ionisation energy is much higher, it means removal of 4th electron is from a stable configuration, hence aluminium is the element.

57. **(d) :** Hydrogen is placed with alkali metals, but its ionisation energy is closer to halogens

H	F	Cl	Br	I
1312.0	1681.0	1255.5	1139.9	1008.4

Beryllium ($2s^2$) has stable electronic configuration than boron ($2s^22p^1$), hence a higher ionisation energy.

58. **(c) :** $O^{ñ} + e^{ñ} \rightarrow O^{2ñ}$ ΔH = + 702 kJ $mol^{ñ1}$
electron is added to uninegative $O^{ñ}$. Energy is given to overcome the repulsive interaction.
$N + e^{ñ} \rightarrow N^{ñ}$ ΔH = 0
$O + e^{ñ} \rightarrow O^{ñ}$ ΔH = ñ 141 kJ/mol
$F + e^{ñ} \rightarrow F^{ñ}$ ΔH = ñ 328 kJ/mol

59. **(d) :** ñve electron affinity is in the order of Mg < Al < Na < Si

+67 ñ50 ñ53 ñ119

60. **(c) :** Ineffective shielding causes decrease in size.

61. **(b) :** Increase in concentration compels the pairing of electron in solution cavity, thus, it becomes diamagnetic from paramagnetic.

62. **(c) :** The stability of the peroxides and superoxides increases as the metal ions become larger. Basically, large cations are stabilised by large anions, since, if both ions are similar in size, the coordination number will be high, and this gives a high lattice energy.

63. **(a) :** Size decreases, effective nuclear charge increases resultantly the ionisation energy increases.

64. **(d) :** $2KNO_3 \rightarrow 2KNO_2 + O_2$

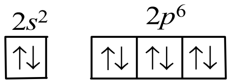
65. **(b)**

66. **(b) :** At low temperature,

$AlCl_3$ exists as close packed lattice of $Cl^{ñ}$ with Al^{3+} occupying octahedral voids. On increasing temperature volume increases, $AlCl_3$ converts into dimeric Al_2Cl_6.

67. **(b) :** $As^{3ñ}$ and $Br^{ñ}$ are isoelectronic species.

68. **(c) :** If *s* could have the value $-\frac{1}{2}, 0, +\frac{1}{2}$ then an orbital can have 3 electrons. The second period starts with $2s^1$ to $2s^2 2p^6 (8e^{ñ})$



4 orbitals are involved, thus, can have 12 electrons and 12 elements can be there.

69. **(c):** $\Delta E = E_{CóH}$ ñ $\sqrt{E_{H-H} \times E_{C-C}}$

$= 98.8-(104.2\times 83.1)^{\frac{1}{2}}$ = 5.75 kcal

x_C ñ $x_H = 0.18\sqrt{\Delta E} = 0.18\,(5.75)^{\frac{1}{2}} = 0.43$

$\therefore\ x_C = 0.43 + x_H = 0.43 + 2.1 = 2.53$

70. **(d):** Since second electron affinities are ñve and O has higher numerical value than S, therefore, option (d) is correct.

71. **(a):** There is an abnormal jump in the 7th ionization energy, therefore, the element has 6 valence electrons and hence it belongs to group 16. Nevertheless, there is a big jump in the 5th ionization energy. This should not be confused with the inert gas configuration but is rather due to the reason that after the removal of four electrons from the *p*-subshell, the 5th electron has to be removed from a *ns*-orbital.

72. **(a):** The size of K will be higher than that of K^+ while size of F will be lower than that of $F^{ñ}$.

73. **(a):** Due to high ionization energy and weak metallic bonding, element with atomic number 80 *i.e.,* Hg is a liquid at room temperature and hence has the maximum vapour pressure at room temperature.

74. **(d):** Rb (2.11 ≈) is much larger than O (0.73 ≈). Since the size of cation is smaller while that of the anion is larger than the parent atom, therefore, the sizes of Rb^+ (1.48 ≈) and $O^{2ñ}$ (1.45 ≈) coincidently match each other. The sizes of Li^+ (0.68 ≈) and Mg^{2+} (0.65 ≈) being diagonally related are nearly the same. The sizes of Fe (1.17 ≈) and Co (1.16 ≈) are nearly the same because the contractive effect of the increased nuclear charge is counter balanced by increase in size due to increased shielding effect of the inner electrons.

75. **(a):** Both As and Bi belong to group 15 of the periodic table and therefore, possess five electrons in the outermost orbit.

76. **(b) :** ${}_{26}Fe^{3+} = [Ar]^{18}\, 3d^5$, ${}_{25}Mn^{2+} = [Ar]^{18}\, 3d^5$.

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

78. **(a) :** Noble gases have configuration $ns^2\, np^6$.

79. **(b):** Mg, Al, Si, P all belong to 3rd period with outer electronic configuration as $3s^2$, $3s^2\, 3p^1$, $3s^2\, 3p^2$ and $3s^2\, 3p^3$ respectively. As the IE_1 increases from left to right in a period, P having exactly half filled configuration will have the highest value of IE_1.

80. **(d):** In case of isoelectronic species, the size decreases as the +ve charge increases. Therefore, Ne (neutral) has the smallest size than those of ñvely charged species.

81. **(b):** As alkali metals have the lowest IE_1 values in their respective periods and also there is a large jump between IE_1 and IE_2 due to the removal of second electron from the noble gas core, therefore, out of the given IE_1 and IE_2 values, (b) belongs to alkali metals.

82. **(a):** In case of Rb^{2+} and Ba^{3+}, the stable noble gas core of Kr and Xe is disturbed and hence are unstable ions. Between Sn^{4+} and $Se^{2ñ}$, the stable noble gas configuration of Kr is achieved only in case of $Se^{2ñ}$ (EAN = 34 + 2 = 36) and hence, is the most stable ion.

83. **(d):** As the element mentioned in (d) belongs to 2nd period (whereas other mentioned electronic configurations are those of elements belonging to 3rd period) and contain exactly half filled electronic configuration, therefore, it would have the highest first *IE*.

84. **(a):** Second *I.P.* of an element will be the energy required to remove one mole of electrons from one mole of monovalent gaseous cations of the element.

85. **(c):** Paramagnetism is associated with number of unpaired electrons in the valence shell of an atom. As configuration (c) contains zero unpaired electrons it would, therefore, not be paramagnetic.

86. **(c):** Element with $Z = 7$ is N.

87. **(b):** Cl has the highest *E.A.* value.

88. **(a):** $H^{ñ}$ ($1s^2$) ion is isoelectronic with He ($1s^2$).

89. **(b):** Cerium is the most common lanthanide.

90. **(c):** The atomic radii of C, O, F which lie in the second period increase in the order: F, O, C. Further, the atomic radii of F, Cl, Br which lie in the same group (group 17) increase in the order: F, Cl, Br. Thus, the overall increasing order is: F, O, C, Cl, Br.

91. **(d):** $IE_3 > IE_2 > IE_1$

92. **(a):** Germanium is a metalloid.

93. **(d):** De Chancourtois.

94. **(a):** Biggest jump between IE_2 and IE_3 means that the element has two valence electrons, *i.e.* $1s^2\ 2s^2\ 2p^6\ 3s^2$.

95. **(d):** Li is the lightest metal in the periodic table.

96. **(a):** Amongst B, C, N and O; N has the highest first ionization energy.

97. **(d):** Since the element (*E*) forms a higher oxide with the formula, EO_3, therefore, it must belong to group 16.

As such Si which lies in group 14 and would from SiO_2 as the higher oxide stands rejected.

98. **(b):** IE_2 of Mg is lower than that of Na because in case of Mg^{3+} 3*s*-electron has to be removed while in case of Na^+, an electron from the stable inert gas configuration (Neon) has to be removed.

99. **(a):** $O^{2ñ} > O > O^+$

100. **(d):** Amongst isoelectronic ions, cations are smaller than anions.

101. **(a):** Si and P are in the 3rd period while C and N are in the 2nd period. Elements in 2nd period have higher electronegativities than those in the 3rd period. Since N has smaller size and higher nuclear charge than C, its electronegativity is higher than that of C. Similarly, the electronegativity of P is higher than that of Si. Thus, the overall order is: Si, P, C, N.

102. **(a):** Electronegativity on the Mullikenís scale = 1/2 (13.0 + 3.8) = 8.4. But since Mullikenís values are nearly 2.8 times as large as Pauling values, therefore, electronegativity of the element on the more commonly used Pauling scale is 8.4/2.8 = 3.0.

103. **(c):** Oxygen belongs to 2nd period; Si, P and Cl belong to 3rd period while Se and Br belong to 4th period. O is the most electronegative followed by Cl and Br. Amongst Cl and Br, Cl is more electronegative than Br. Se being a chalcogen is more electronegative than Si and P. Amongst Si and P, P is more electronegative. Thus, the overall sequence is, Si, P, Se, Br, Cl, O.

104. **(a):** Na^+ and $F^{ñ}$ are isoelectronic ions. Since amongst isoelectronic ions, the size of the anion is larger than that of the cation, therefore, $Na^+ < F^{ñ}$.

105. **(b):** In accordance with Aufbau principle, configuration (b) is correct.

106. **(b):** Elements with Z = 4 (Be), 12 (Mg), 20 (Ca), 38 (Sr), 56 (Ba), 88 (Ra) belong to group 2 (alkaline earths) of the periodic table.

107. **(d):** Tenth element is a member of noble gases (Group 18 or zero). The atomic numbers of other elements of this group are 2 (He), 18 (Ar), 36 (Kr), 54 (Xe) and 86 (Rn).

108. **(d):** In *f*-block elements, the last electron enters into the *f*-orbital of the anti-penultimate shell.

109. **(c) :** Most electronegative element is fluorine with $Z = 9$ *i.e.* $1s^2\ 2s^2\ 2p^5$.

110. **(c) :** Electronic configuration of
$_{28}Ni = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$
$Ni^{2+} = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8$
Thus *d*-electrons in Ni and Ni^{2+} are same.

Assertion and Reason

1. **(c) :** Charge is not defined as positive or negative. [Isoelectronic species, having higher the negative charge, larger the size, higher the positive charge, smaller the size.]

2. **(a)**

3. **(b) :** In case of Sulphur valence shell is less dense than oxygen.

4. **(a)**

5. **(b) :** Atomic size of silver is almost equal to that of gold due to lanthanide contraction.

6. **(a)** 7. **(a)**

8. **(b) :** Electrons in *d*-and *f*-sublevels can never be in the outer level of a neutral atom. The *s*-and *p*-electrons are in the highest energy level in the atom and are the electrons involved in chemical reactions.

9. **(d)**

10. **(b) :** $F^{ñ}$, $O^{2ñ}$ and $N^{3ñ}$ ions are isoelectronic ions [each have 10 electrons] but their nuclear charges are different. Nuclear charges of $F^{ñ}$, $O^{2ñ}$ and $N^{3ñ}$ are +9, +8, +7. The ionic radii of isoelectronic ions decrease with increase in magnitude of nuclear charge. Thus increasing order of ionic radii is : $F^{ñ} < O^{2ñ} < N^{3ñ}$.

11. **(d) :** Ionisation energy increases abruptly at fourth ionisation energy *i.e.* $IE_4 >>> IE_3$ and as the 4th electron requires very-very high energy for its removal as this electron is to be knocked out from the noble gas core. Hence, the number of valence electron is three.

12. **(a)**

13. **(d) :** Since $F^{ñ}$ is the smallest anion and Cs^+ is the largest cation and so for CsF, ratio of anionic size to cationic size is lowest.

14. **(b) :** If the difference in ionisation enthalpies of two successive states is approximately equal to 965-1450 kJ mol^{ñ1} or less, lower oxidation state is not stable while if it is greater than 1450 kJ mol^{ñ1}, the lower oxidation state is more stable *e.g.*, for Al, IE_1 = 579 kJ mol^{ñ1}, IE_2 = 1795 kJ mol^{ñ1}, IE_3 = 2758 kJ mol^{ñ1} and IE_4 = 11580 kJ mol^{ñ1}.

15. **(a) :** The element with electronic configuration (Xe) $4f^{1}5d^{1}6s^{2}$ is a *f* block element because the last electron enters the *f* orbital.

16. **(b) :** Total energy required to convert one mole of He atoms into He^{2+} ions is
(24.58 + 54.4) eV = 78.98 eV
= 78.98 ◊ 96.49 kJ = 7620.780 kJ/mole
$\therefore$ Energy required to convert 0.1 mole of He atoms into He^{2+} ions is 762.078 kJ.

17. **(c) :** The elements can be arranged in different periods and groups as follows :

Period/Group	1	2	13	14	15
2nd	ñ	Be	ñ	ñ	ñ
3rd	Na	Mg	ñ	Si	P

As metallic character decreases along a period and increases down a group, so Na is the most metallic element than Mg and Si and P is the least metallic element. Among Be and Mg, Mg is more metallic. Hence, increasing order of metallic character is :
P < Si < Be < Mg < Na.

18. **(b) :** $\text{No. of moles of bromine} = \frac{8}{80} = 0.1$
$\therefore$ Energy required = 0.1 ◊ 3.48 ◊ 23.06
= 8.02 kcal

19. **(a) :** As number of energy shells for H, O, N and Cl is 1, 2, 2 and 3 respectively so van der Waalís radius of H is the smallest and Cl is the biggest. Both O and N have two energy levels but the nuclear charge on O (+8) is higher than that of N (+7). Thus, the Van der Waalís radius of N is bigger than that of O.
So, the decreasing order is : Cl > N > O > H.

20. **(b) :** Electronic configuration of Na is $1s^22s^22p^63s^1$ & Electronic configuration of Mg is $1s^22s^22p^63s^2$.
In both cases, first electron is to be removed from the 3s orbital but nuclear charge of Na is less than that of Mg, so, first ionisation enthalpy of Mg is higher than that of Na. After the removal of first electron, Na^+ ($1s^22s^22p^6$) has stable inert gas configuration and so, more energy is required to remove second electron while in case of Mg^+ ($1s^22s^22p^63s^1$), removal of second electron is easy as after that Mg^{2+} attains noble gas configuration.

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1. **(d) :** Electron gain enthalpy becomes less negative from top to bottom in a group while it becomes more negative from left to right within a period.

2. **(c) :** Na, Mg, Al and Si are in period 3 and as we move across the period, the atomic size decreases and hence ionisation enthalpy increases. So, the order is Na < Mg < Al < Si.
Hence, *IE* of Al is greater than that of Mg (737 kJ/mol) and lower than that of Si (776 kJ/mol).

3. **(None) :** Cations lose electrons and are smaller in size than the parent atom, whereas anions add electrons and are larger in size than the parent atom. Hence, the order is
$H^{ñ} > H > H^{+}$.
For isoelectronic species, the ionic radii decreases with increase in atomic number *i.e.* nuclear charge.
Hence, the correct orders are
$O^{2ñ} > F^{ñ} > Na^{+}$ and $N^{3ñ} > Mg^{2+} > Al^{3+}$.

4. **(b):** The atomic radii of noble gases are by far the largest in their respective periods. This is due to the reason that noble gases have only van der Waals radii.

5. **(a) :** In case of isoelectronic species, radius decreases with increase in nuclear charge.

6. **(a) :** In general, electron affinity increases from left to right in a period and decreases down the group. But due to interelectronic repulsions in small sized F-atom, Cl has got larger electron affinity than that of F.

7. **(a, d) :** The correct order of increasing negative electron gain enthalpy is : I < Br < F < Cl and the correct order of increasing first ionisation enthalpy is B < C < O < N.

8. **(c) :** The electronic configuration of (X) can be written as

   $X = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

   So, element (X) has completely filled s and d orbitals and half filled p orbitals.

9. **(a) :** The electronic configuration of the element with Z = 114 (flerovium) is [Rn]$5f^{14}\ 6d^{10} 7s^2 7p^2$.

   Hence, it belongs to carbon family which has the same outer electronic configuration.

10. **(c) :** For hydrogen like species, energy of an electron in n^{th} orbit is given by

$$E_n = -\frac{2\pi^2 m Z^2 e^4 k^2}{n^2 h^2}$$

   i.e., $E_n \propto Z^2$.

   Electronegativity increases across a period with decrease in size of atoms.

☺☺☺