



Extended or Long Form of the Periodic Table

s-Block Elements		d-Block Elements										p-Block Elements							
Group	1A (1)	2A (2)											3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	0 (18)	
Period 1	1 H 1.0079 Hydrogen	IIA (2)											IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	2 He 4.0026 Helium	
2	3 Li 6.940 Lithium	4 Be 9.0122 Beryllium																9 F 18.998 Fluorine	10 Ne 20.180 Argon
3	11 Na 22.990 Sodium	12 Mg 24.305 Magnesium																17 Cl 35.453 Chlorine	18 Ar 39.948 Argon
4	19 K 39.098 Potassium	20 Ca 40.078 Calcium																35 Br 79.904 Bromine	36 Kr 83.80 Krypton
5	37 Rb 85.468 rubidium	38 Sr 87.62 Strontium																53 I 126.90 Iodine	54 Xe 131.29 Xenon
6	55 Cs 132.91 Cesium	56 Ba 137.33 Barium																85 At 210 Astatine	86 Rn 222 Radon
7	87 Fr 223 Francium	88 Ra 226 Radium																	

Metals

Non Metals

Metalloids

The symbols for elements 104-109 used in this table are those proposed by the American Chemical Society and 110-112 proposed by IUPAC

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f-Block Elements

6	**Lanthanide series										65 Tb 158.92 terbium	66 Dy 162.50 dysprosium	67 Ho 164.93 Holmium	68 Er 167.26 Erbium	69 Tm 168.934 Thulium	70 Yb 173.04 Ytterbium	71 Lu 174.97 Lutetium
7	**Actinide Series										97 Bk 247 berkelium	98 Cf 251 Californium	99 Es 252 Einsteinium	100 Fm 257 Fermium	101 Md 258 Mendelevium	102 No 259 Nobelium	103 Lr 262 Lawrencium

IUPAC designations of groups of elements are given in brackets



IMPORTANT FACTS TO REMEMBER

1. Lowest electronegativity	Cs
2. Highest electronegativity	F
3. Highest ionisation potential	He
4. Lowest ionisation potential	Cs
5. Lowest electron affinity	Noble gases
6. Highest electron affinity	Chlorine
7. Least electropositive element	F
8. Lowest <i>m</i> pt. metal	Hg
9. Highest <i>m</i> . pt. and b. pt. metal	W (Tungsten)
10. Lowest <i>m</i> . pt. and b. pt. non metal	He
11. Notorious element	Hydrogen
12. Lightest element	Hydrogen
13. Smallest atomic radius	H
14. Largest atomic size	Cs
15. Largest monovalent anionic size	I
16. Smallest cation	H ⁺
17. Most electropositive element	Cs
18. Element with electronegativity after <i>F</i>	Oxygen next to Fluorine
19. Group containing maximum no. of gases	Zero group of gaseous elements in periodic table
20. Liquid element of radioactive nature	Fr
21. Volatile d-block elements	Zn, Cd, Hg, Uub
22. Element containing no neutron	H
23. Most abundant element on earth	Oxygen
24. Rarest element on earth	At (Astatine)
25. Most abundant metal on earth	Al
26. Element having maximum tendency	Carbon for catenation
27. Non metal having highest <i>m</i> pt., b.pt.	Carbon (Diamond)

28. Metal showing highest oxidation number	Os (+8)
29. Most electrovalent compound	CsF
30. Most stable carbonate	Cs ₂ CO ₃
31. Strongest alkali (Hydroxide)	CsOH
32. Strongest basic oxide	Cs ₂ O
33. Best electricity conductor among metals	Ag
34. Best electricity conductor among non metals	graphite
35. Most poisonous element	Pu (Plutonium)
36. Liquid non metal	Br
37. Element kept in water	Phosphorus
38. Elements kept in kerosene	IA group element (except Li)
39. Element sublime on heating	I ₂
40. Noble metals	Au, Pt etc.
41. Amphoteric metal	Be, Zn, Al, Sn, Pb
42. Amphoteric non metal	Si
43. Metalloids elements	Si, As, Te, At, Ge, Sb, Po
44. Non metals having metallic lusture	Graphite, Iodine
45. Heaviest naturally occurring element	Uranium
46. Poorest conductor of electricity	Diamond
47. Hardest naturally occurring element	Diamond
48. Lightest solid metal	Li
49. 90% of Sun mass	Hydrogen
50. Amphoteric oxides	BeO, Al ₂ O ₃ , ZnO, PbO ₂ SnO, SnO ₂
	Sb ₂ O ₃ , As ₂ O ₃ etc..
51. Neutral oxides of non metals	NO, CO, H ₂ O, N ₂ O
52. Dry bleacher	O ₃ (ozone)
53. Dry ice	Solid CO ₂

54. Artificial explosive	TNT, RDX (Research Developed Explosive etc.)
55. Oldest known organic acid	CH_3COOH
56. First noble prize of chemistry was given to	Vant Hoff
57. Some isomorphism substances	$\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ (Green Vitriol), $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ (Epsom Salt), $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ (White Vitriol)
58. Some efflorescent substances	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ etc.
59. First man made element	Tc (Technicium)
60. Smallest period	1 st (2 elements)
61. Largest period in periodic table	6 th (32 elements)
62. Largest group in periodic table	3 rd /III B (32 elements)
63. Most abundant d-block metal	Fe
64. Most abundant s-block metal	Ca
65. Element having maximum isotopes	Sn (10)
66. Highest density (Metals)	$\text{Ir} \geq \text{Os}$
67. Highest density (Non Metals)	Boron

Common and IUPAC names of some recently discovered elements

Atomic No.	Common names	IUPAC names
104	Rutherfordium (Rf)/ Kurchatovium (Ku)	Unnilquadium (Unq)
105	Dubnium (Db)/Hahnium (Ha)	Unnilpentium (Unp)
106	Seaborgium (Sg)	Unnilhexium (Unh)
107	Bohrium (Bh)	Unnilseptium (Uns)
108	Hassium (Hs)	Unniloctium (Uno)
109	Meitnerium (Mt)	Unnilennium (Une)

INTRODUCTION:

The arrangement of all the known elements according to their properties in such a way that the elements of similar properties are grouped together in a tabular form is called periodic table. Scientists from the very beginning have attempted to systematize the knowledge they gained through their observations and experiments. Development of the periodic law and the periodic table of the elements is one such attempt. This has brought order in the study of the vast chemistry of more than a hundred elements known now. Therefore, it is quite natural to begin your study of inorganic chemistry with the study of the periodic table in this unit.

By the mid-nineteenth century, more than 60 elements were known and many more were being discovered. During this period, it was also realized that certain groups of elements exhibited similar physical and chemical properties. Was it a mere coincidence or did a relationship exist among the properties of the elements? Attempts to reply such probing questions ultimately resulted in the development of the periodic law.

Periodic table helps us to undergo a systematic study of the various elements found in nature, Without which it would have been impossible for us to study all the elements. By classifying the elements into various groups and periods a comparative study of the elements and their compounds can be done. Several attempts were made to classify the elements which are as follows.

DEVELOPMENT OF PERIODIC TABLE :**(A) LAVOISIER CLASSIFICATION :**

- (i) Lavoisier classified the elements simply in metals and non metals. Metals are the one which have the tendency of losing the electrons.



Non-metals are the one which have the tendency of gaining the electrons.



- (ii) Drawbacks or Limitations:

- (a) As the number of elements increased, this classification became insufficient for the study of elements.
- (b) There are few elements which have the properties of both metals as well as nonmetals and they are called metalloids. Lavoisier could not decide where to place the metalloids.

(B) PROUT'S HYPOTHESIS (Unitary theory):

He simply assumed that all the elements are made up of hydrogen, so we can say that Atomic

weight of element = $n \times$ (Atomic weight of one hydrogen atom)

Atomic weight of H = 1

where n = number of hydrogen atom = 1,2,3, ...

Drawbacks or Limitations :

- (i) Every element cannot be formed by Hydrogen.
- (ii) Atomic weight of all elements were not found as the whole numbers.

Ex. Chlorine (atomic weight 35.5) and Strontium (atomic weight 87.6)

(C) DOBEREINER TRIAD RULE [1817]:

- (i) He made groups of three elements having similar chemical properties called TRIAD.
- (ii) In Dobereiner triad, atomic weight of middle element is nearly equal to the average atomic weight of first and third element.

Ex.	Cl	Br	I	Ca	Sr	Ba	Li	Na	K
	35.5	80.0	127	40	87.6	137	7	23	39
	$\left[x = \frac{35.5 + 127}{2} = 81.2 \right]$			$\left[x = \frac{40 + 137}{2} = 88.5 \right]$			$\left[x = \frac{7 + 39}{2} = 23 \right]$		

Where x = average atomic weight

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

Drawbacks or Limitations : All the known elements could not be arranged as triads. It is not applicable for d and f-block elements.

Illustrations (Optional)

Illustration 1. Atomic weight of an element X is 39 , and that of element Z is 132 . atomic weight of their intermediate element Y, as per dobereiner triad, will be

- (A) 88.5 (B) 93.0 (C) 171 (D) 85.5

Solution. (D)

$$\frac{39 + 132}{2} = 85.5$$

Illustration 2. Which of the following is not a dobereiner triad

- (A) Li, Na, K (B) Mg, Ca, Sr (C) Cl, Br, I (D) S, Se, Te

Solution. (B)

The difference in atomic number are different

Illustration 3. The law of triads is applicable to

- (A) C, N, O (B) H, O, N (C) Na, K, Rb (D) Cl, Br, I

Solution. (D)

The difference in atomic number are same.

Illustration 4. The law of triads is not applicable on

- (A) Cl, Br, I (B) Na, K, Rb (C) S, Se, Te (D) Ca, Sr, Ba

Solution. (B)

The difference in atomic number are different

(D) NEWLAND'S OCTAVE' RULE [1865]

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

Sa	Re	Ga	Ma	Pa	Dha	Ni	Sa
Li	Be	B	C	N	O	F	Na
Na	Mg	Al	Si	P	S	Cl	K

- (ii) At that time inert gases were not known.
 (iii) The properties of Li are similar to 8th element i.e. Na and properties of Be are similar to Mg and so on.

• **Drawbacks or Limitations :**

- (a) This rule is valid only up to Ca, because after Ca due to presence of d-block elements there is difference of 18 elements instead of 8 element.
 (b) After the discovery of Inert gases and including them into the periodic table it becomes the 8th element from Alkali metal so this law had to be dropped out.

Illustrations (Optional)

Illustration 1. Which of the following set of elements obey Newland's octave rule-

- (A) Na, K, Rb (B) F, Cl, Br (C) Be, Mg, Ca (D) B, Al, Ga

Ans. (C)

Illustration 2. For which of the pair Newland's octave rule is not applicable-

- (A) Li, Na (B) C, Si (C) Mg, Ca (D) Cl, Br

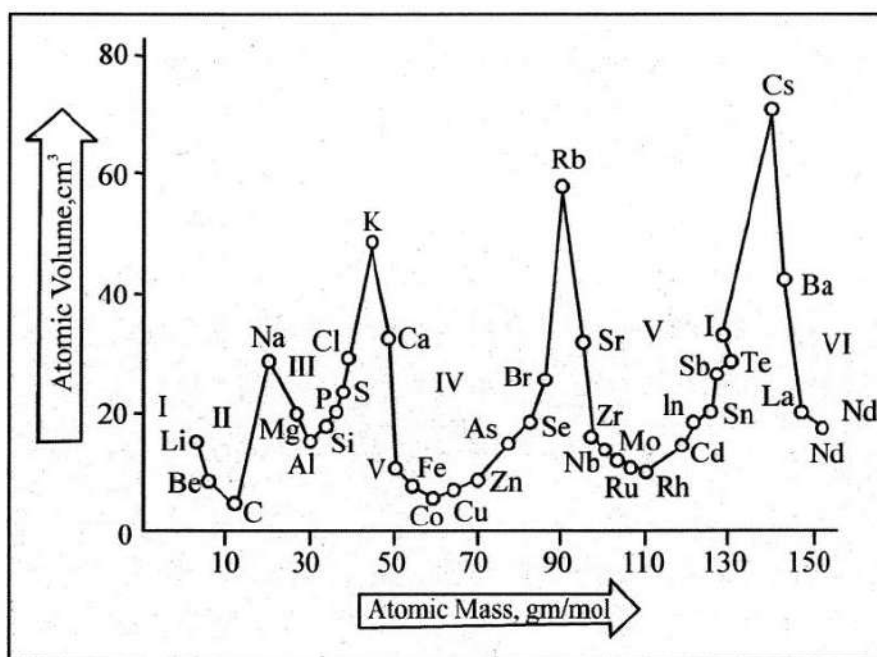
Ans. (D)

(E) LOTHER MEYER'S CURVE [1869] :

- (i) He plotted a curve between atomic weight and atomic volume of different elements.
 (ii) The following observations can be made from the curve –
 (a) Most electropositive elements i.e. alkali metals (Li, Na, K Rb, Cs etc.) occupy the peak positions on the curve.
 (b) Less electropositive i.e. alkaline earth metal (Be, Mg, Ca, Sr, Ba) occupy the descending positions on the curve.
 (c) Metalloids (Si, As, Te, etc.) and transition metals occupy bottom part of the curve.
 (d) Most electronegative i.e. halogens (F, Cl, Br, I) occupy the ascending positions on the curve.

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

Periodic function : When the elements are arranged in the increasing order of their atomic weight, elements having similar properties gets repeated after a regular interval.



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

(a) It was based on atomic weight.

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

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

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

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

(f) Each group upto VII was divided into A & B sub groups 'A' sub group element were called normal elements and 'B' sub group elements were called transition elements.

(g) The VIII group was consisted of 9 elements in three rows (Transition metal group).

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

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

- (a) Study of elements: First time all known elements were classified in groups according to their similar properties. So study of the properties of elements become easier.
- (b) Prediction of new elements : It gave encouragement to the discovery of new elements as some gaps were left in it.

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

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

Ex. Blank space at atomic weight 72 in silicon group was called Eka silicon (means properties like silicon) and element discovered later was named Germanium. Similarly other elements discovered after mendeleev periodic table were.

Eka aluminium	-	Gallium(Ga)	Eka Boron	-	Scandium (Sc)
Eka Silicon	-	Germanium(Ge)	Eka Manganese	-	Technitium (Tc)

Illustrations (Optional)

Illustration 1. Which of the following element was present in Mendeleev's periodic table?

- (A) Sc (B) Tc (C) Ge (D) None of these

Solution. (D)

These elements were added after Mendeleev's periodic table.

- (c) Correction of doubtful atomic weights : Correction were done in atomic weight of some elements.

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

Initially, it was found that equivalent weight of Be is 4.5 and it is trivalent ($V = 3$), so the weight of Be was 13.5 and there is no space in Mendeleev's periodic table for this element.

So, after correction, it was found that Be is actually bivalent ($V = 2$). So, the weight of Be became $2 \times 4.5 = 9$ and there was a space between Li and B for this element in Mendeleev's table.

Note: Corrections done in atomic weight of elements were - U, Be, In, Au, Pt.

Demerits of Mendeleev's periodic table:

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

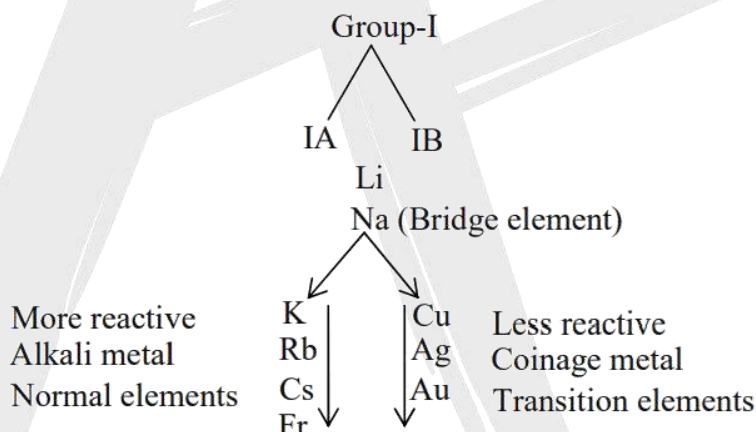
- (b) Position of isotopes: As atomic weight of isotopes differs, they should have been placed at different position in Mendeleev's periodic table, But there were no such places for isotopes in Mendeleev's periodic table.
- (c) Anomalous pairs of elements : There were some pairs of element which did not follow the increasing order of atomic weights.
- Ex. Ar and Co were placed before K and Ni respectively in the periodic table, both having higher atomic weights.

(Ar	K)	(Te	I)
(39.9	39.1)	(127.5	127)
(Co	Ni)	(Th	Pa)
(58.9	58.6)	(232	231)

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

Pt	Au
VIE	IB

- (e) Unlike elements were placed in same group :



Cu, Ag and Au placed in 1st group along with Na, K etc. while they differ in their properties (Only similar in having ns¹ electronic configuration)

- (f) It was not clear that 'lanthanides and Actinides' are related with IIIA group or IIIB group.
- (g) Cause of periodicity: Why physical and chemical properties repeated in a group.

Groups of Elements											
0	I	II	III	IV	V	VI	VII	VIII			
	H 1.008										
He 4.0	Li 7.03	Be 9.1	B 11.0	C 12.0	N 14.04	O 16.00	F 19.0				
Ne 19.9	Na 23.5	Mg 24.5	Al 27.0	Si 28.4	P 31.0	S 32.03	Cl 35.45				
Ar 38	K 39.1	Ca 40.1	Sc 44.1	Ti 48.1	V 51.4	Cr 52.1	Mn 55.0	Fe 55.9	Co 59	Ni 59	Cu
	Cu 63.6	Zn 65.4	Ga 70.0	Ge 72.3	As 75	Se 79	Br 79.95				
Kr 81.8	Rb 85.4	Sr 87.6	Y 89.0	Zr 90.6	Nb 94.0	Mo 96.0	----	Ru 101.7	Rh 103.0	Pd 106.5	Ag
	Ag	Cd	In	Sn	Sb	Te	I				
	107.9	112.4	114.0	119.0	120.0	127.6	126.9				
Xe 128	Cs 132.9	Ba 137.4	La 139	Ce 140	---	---	---				
			Yb 173	---	Ta 183	W 184		Os 191	Ir 193	Pt 194.9	Au
	Au 197.2	Hg 200.0	Tl 204.1	Pb 206.9	Bi 208						
		Ra 224		Th 232		U 239	---				
Rn	R ₂ O	RO	R ₂ O ₃	HIGHER SALINE OXIDES							
				RO ₂	R ₂ O ₅	RO ₃	R ₂ O ₇				
				HIGHER GASEOUS HYDROGEN COMPOUNDS							
				RH ₄	RH ₃	RH ₂	RH	RO ₄			

DO YOURSELF - 1

- Mendeleev's periodic law is based on -
 - Atomic number
 - Atomic weight
 - Number of neutrons
 - None of the above
- The first attempt to classify elements systematically was made by -
 - Mendeleev
 - Newland
 - Lothar Meyer
 - Dobereiner

(G) MODERN PERIODIC TABLE (MODIFIED MENDELEEV PERIODIC TABLE) :**Moseley's Discovery-The modern Concept of Atomic Number.****Moseley's Experiment**

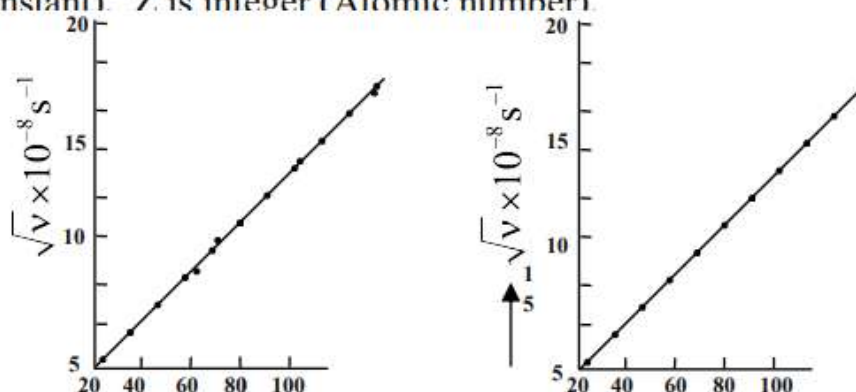
Moseley in his discharge tube experiment observed that nature of x-ray produced by striking of cathode ray on anode metal depends on the nature of metal taken as anode. He observed that wave length of x-ray decreases (or frequency increases) in the regular way on moving from lighter to heavier element. Property of x-ray depends on charge in nucleus (positive charge, proton). From this observation he established a mathematical expression.

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

Where, ν is the frequency of x-ray a is the proportionality constant b is constant (series constant). Z is integer (Atomic number).



Henry G.J. Moseley 1887-1915.
A brilliant English physicist discovered the concept of atomic number in 1913. He had to go to battlefield during World War I and got Killed in 1915



- (i) It was proposed by Moseley.
- (ii) Modern periodic table is based on atomic number.
- (iii) Moseley did an experiment in which he bombarded high speed electron on different metal surfaces and obtained X-rays.

He found out that $\sqrt{\nu} \propto Z$ where

ν = frequency of X-rays

Z = atomic number.

From this experiment, Moseley concluded that the physical and chemical properties of the elements are periodic function of their atomic number. It means that when the elements are arranged in the increasing order of their atomic number elements having similar properties gets repeated after a regular interval. This is also known as 'Modern Periodic Law'.

(iv) Modern periodic law:

The physical & chemical properties of elements are the periodic function of their atomic number.

(v) Characteristics of modern periodic table:

- (a) 9 vertical columns called groups.
- (b) I to VIII group +0 group of inert gases.
- (c) Inert gases were introduced in periodic table by Ramsay.
- (d) 7 horizontal rows called periods.

(H) LONG FORM / PRESENT FORM OF MODERN PERIODIC TABLE :

(It is also called as 'Bohr, Bury, Rang and Werner Periodic Table')

- (i) It is based on the Bohr-Bury electronic configuration concept and atomic number.
- (ii) This model is proposed by Rang & Werner
- (iii) 7 periods and 18 groups
- (iv) According to I. U. P. A. C. 18 vertical columns are named as 1st to 18th group.
- (iv) Modern periodic law: The physical & chemical properties of elements are the periodic function of their atomic number.
- (v) The co-relation between the groups in long form of periodic table and in modern form of periodic table are given below -

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

- (vi) Elements belonging to same group have same number of electrons in the outermost shell so their properties are similar.

Description of periods

Period	No.	Period Sub shell	No. of elements	Element	Name of Period
1.	1	1s	2	${}_1\text{H} - {}_2\text{He}$	Shortest
2.	2	2s, 2p	8	${}_3\text{Li} - {}_{10}\text{Ne}$	Short
3.	3	3s, 3p	8	${}_{11}\text{Na} - {}_{18}\text{Ar}$	Short
4.	4	4s, 3d, 4p	18	${}_{19}\text{K} - {}_{36}\text{Kr}$	Long
5.	5	5s, 4d, 5p	18	${}_{37}\text{Rb} - {}_{54}\text{Xe}$	Long
6.	6	6s, 4f, 5d, 6p	32	${}_{55}\text{Cs} - {}_{86}\text{Rn}$	Longest
7.	7	7s, 5f, 6d, 7p	26	${}_{87}\text{Fr} - {}_{112}\text{Uub}$	Longest*

* All elements of periodic table has been discovered

Conclusion:

1. Period number = outermost shell
2. Number of element in a period = Number of electrons in a period subshell

1st/IA/ Alkali metals $H = 1s^1$ $Li = 1s^2, 2s^1$ $Na = 1s^2, 2s^2, 2p^6, 3s^1$ $K = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$ General electronic configuration = ns^1 Number of valence shell electron = 1	2nd/IIA/Alkaline earth metals $Be = 1s^2, 2s^2$ $Mg = 1s^2, 2s^2, 2p^6, 3s^2$ $Ca = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$ General electronic configuration = ns^2 (n = Outermost shell) Number of valence shell electron = 2
13th/ IIIA/Boron Family $B = 1s^2, 2s^2, 2p^1$ $Al = 1s^2, 2s^2, 2p^6, 3s^2, 3p^1$ $Ga = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^1$ General electronic configuration = $ns^2 np^1$ Number of valence shell electron = 3	14th/IVA/Carbon Family $C = 1s^2, 2s^2, 2p^2$ $Si = 1s^2, 2s^2, 2p^6, 3s^2, 3p^2$ $Ge = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^2$ General electronic configuration = $ns^2 np^2$ Number of valence shell electron = 4
15th/VA/Nitrogen family/Pnicogen (Used in fertilizer as urea) $N = 1s^2, 2s^2, 2p^3$ $P = 1s^2, 2s^2, 2p^6, 3s^2, 3p^3$ $As = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^3$ General electronic configuration = $ns^2 np^3$ Number of valence shell electron = 5	16th/VIA/Oxygen family/Chalcogen (Ore forming) $O = 1s^2, 2s^2, 2p^4$ $S = 1s^2, 2s^2, 2p^6, 3s^2, 3p^4$ $Se = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^4$ General electronic configuration : $ns^2 np^4$ Number of valence shell electron = 6
17th/VIIA/Fluorine family/Halogens (Salt forming) $F = 1s^2, 2s^2, 2p^5$ $Cl = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$ $Br = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^5$ General electronic configuration = $ns^2 np^5$ Number of valence shell electron = 7	18th /Zero group/Inert gases / Noble gases (Less reactive) $Ne = 1s^2, 2s^2, 2p^6$ $Kr = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6$ General electronic configuration = $ns^2 np^6$ (except He) Number of valence shell electron = 8

Ex. Which of the following elements belong to alkali metals?

- (A) $1s^2, 2s^2, 2p^2$
 (B) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, 5s^1$
 (C) $1s^2, 2s^2, 2p^5$
 (D) None of these

Ans. (B)

(vii) Nomenclature of elements :

(a) IUPAC gave names to elements above atomic number 100 as follows -

0	1	2	3	4	5	6	7	8	9
Nil	Un	Bi	Tri	Quad	Pent	Hex	Sept	Oct	enn

(b) In all the elements suffix is - ium.

Ex.

Atomic No.	IUPAC Name	Symbol	Elemental Name	Symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnilium	Uun	Darmstadtium	Ds

(viii) Identification of group, period and block:

(A) When atomic number is given :

$$71 \geq Z \geq 58 \Rightarrow \text{Lanthanides (6}^{\text{th}} \text{ Period)}$$

$$103 \geq Z \geq 90 \Rightarrow \text{Actinoids (7}^{\text{th}} \text{ Period)}$$

Step I:

Group number = IIIB (largest group of periodic table)

Step II:

$Z = 104$ to 118 (Period number = 7)

Group number = last two digits in atomic number of element

Example:

$Z = 104$

Group no. = 4

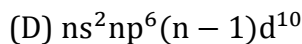
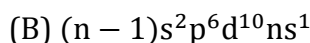
Step III : Group number = 18 + given atomic number - atomic number of next noble gas
 If the value of this formula is negative then use 32 instead of 18 in formula.
 When electronic configuration is given
 Period number (n) = number of outermost shell/highest shell number.

Block identification :

- If np electron present then p - block (ns^2np^{1-6}) group number = 12 + np electrons
- If np electron absent then s/f/d block
 If $(n-2)f^0(n-1)d^0 ns^{1-2} = s$ block
 group number = ns electrons
 If $(n-2)f^{1-14}(n-1)d^{0-1} ns^2 = f$ block
 group number = IIIB
- If any other configuration or $(n-1)d^{1-10} ns^{0-2} = d$ - block group number = (n - 1)d electron + ns electron

DO YOUR SELF - 2

- Which of the following set of atomic numbers represent representative element :-
 (A) 5,13,30,53 (B) 11,33,58,84
 (C) 5,17,31,54 (D) 9,31,53,83
- Which of the following electronic configuration does not belong to same block as others
 (A) $[Xe]4f^{14}5d^{10}6s^2$ (B) $[Kr]4d^{10}5s^2$
 (C) $[Kr]5s^2$ (D) $[Ar]3d^64s^2$
- An element with atomic number 106 has been discovered recently. Which of the following electronic configuration will it possess :-
 (A) $[Rn]5f^{14}6d^57s^1$ (B) $[Rn]5f^{14}6d^57s^2$
 (C) $[Rn]5f^{14}6d^67s^0$ (D) $[Rn]5f^{14}6d^17s^27p^3$
- An element which is recently discovered is placed in 7th period and 10th group. IUPAC name of the element will be :-
 (A) Unnilseptium (B) Ununnilium
 (C) Ununbium (D) None
- The electronic configuration of an element is $1s^22s^22p^63s^23p^63d^{10}4s^1$. What is the atomic number of last element of the same group which is recently discovered :-
 (A) 20 (B) 119
 (C) 111 (D) None
- Which of the following electronic configuration of the outermost shell is characteristic of alkali metals :-

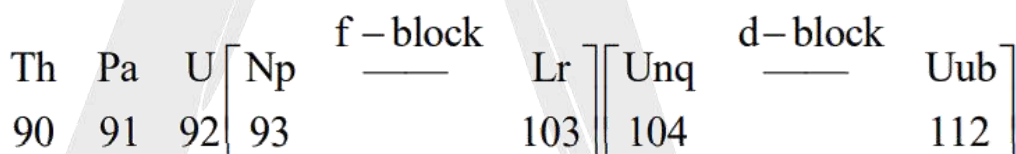


(a) 2nd period elements (Li, Be, B) shows diagonal relationship with 3rd period elements (Mg, Al, Si). Because of same ionic potential value they show similarity in properties.

2nd period Li Be B
3rd period Na Mg Al Si

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

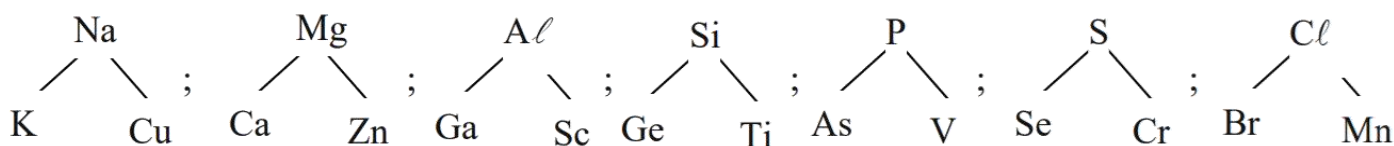
Elements having atomic number more than 92 are known as transuranic element. All transuranic elements are radioactive & artificial.



First man made lanthanoid is - Pm

All actinoids are radioactive but all lanthanoids are not artificial / man made (except Pm)

- (i)** Bridge elements have similar electronic configuration and group nomenclature e.g.



(j) In the periodic table

Number of Gaseous elements - 11 (H, N, O, F, Cl + Noble gases) Number of Liquid elements upto 30°C – 6 (Cs, Fr, Ga, Hg, Br, Uub) Bromine is the only non-metal which exists in liquid form. Number of Solid elements - 95 (if discovered elements are 112)

(k) 0/18th group have all the elements in gaseous form.

(l) 2nd period contains maximum number of gaseous elements - N, O, F, Ne

(m) IIIB/3rd group is called longest group having 32 elements including 14 Lanthanides and Actinides

Sc

Y

La.....Lanthanides (14)

Ac.....Actinides (14)

Defects:

(a) Position of hydrogen is not settled.

(b) Position of helium cannot be justified. It is the only element which belong to s-block but placed in p-block.

(c) Lanthanides & Actinides have not been given space in the main body of periodic table.

Illustrations (Optional)

Illustration 1. Sixth typical element is

(A) Al (B) Mg (C) S (D) O

Ans. (C)

Illustration 2. Elements whose atoms have three outermost shells incomplete are called

(A) s-block (B) p-block (C) d-block (D) f-block

Ans. (D)

Illustration 3. Which of the following statement is wrong

(A) All the actinides are synthetic (man made) elements
(B) In the Lanthanides last electron enters in 4f orbitals
(C) 3rd period elements are typical elements
(D) Lanthanum is d-block element

Ans. (A)

Illustration 4. Which of the following statement is wrong

(A) Total no. of liquid elements in the periodic table.....Six
(B) First metal element in the periodic table is....Li

(C) All type of elements are present in 6th period

(D) Iodine is a gaseous element.

Ans. (D)

ELECTRONIC CLASSIFICATION OF ELEMENTS

The electronic configuration plays great roles in the properties of the elements, On this basis the elements can be divided into four groups.

s-BLOCK ELEMENTS

- (a) The elements of the periodic table in which the last electron enters in s-orbital, are called s-block elements.
- (b) s-orbital can accommodate a maximum of two electrons.
- (c) Their general electronic configuration is ns^{1-2} , where $n = (1 \text{ to } 7)$
- (d) IA group elements are known as alkali metals because they react with water to form alkali. IIA group elements are known as alkaline earth metals because their oxides react with water to form alkali and these are found in the soil or earth crust.
- (e) Fr_{87} and Ra_{88} are radioactive elements while H and He are gaseous elements.

p-BLOCK ELEMENTS

- (a) The elements of the periodic table in which the last electron gets filled up in the p-orbital, called p-block elements.
- (b) The general electronic configuration of p-block elements is $n ns^2, np^{1-6}$ (where $n = 2 \text{ to } 6$)
- (c) p-subshell can accommodate a maximum of six electrons. Therefore, p-block elements are divided into six groups which are IIIA, IVA, VA, VIA, VIIA and zero group.
- (d) The zero group elements having general electronic configuration $ns^2 np^6$ are inert, because their octets are complete.

d-BLOCK ELEMENTS

- (a) The elements, in which the last electron enters into $(n - 1)d$ -orbital are called d-block elements.
- (b) The general electronic configuration of these elements is $(n - 1)s^2, p^6, d^{1-10}, ns^{1-2}$ [except $\text{Pd} \rightarrow 4d^{10}, 5s^0$] where $n = 4 \text{ to } 7$. [except $\text{Th} = (\text{Rn})5f^0, 6d^2, 7s^2$]
- (c) The d-block elements are placed in the groups named IIIB, IVB, VB, VIB, VIIB, VIII, IB and IIB.
- (d) All of these elements are metals.
- (e) d-block elements lies between s & p block elements.

f-BLOCK ELEMENTS

- (a) The elements in which the last electron enters into $(n-2)f$ -orbital are called f-block elements.

- (b) The f-block elements are from atomic number 58 to 71 and from 90 to 103. The general electronic configuration of these elements is $(n-2)s^2, p^6, d^{10}(n-1)s^2, p^6, d^{0-1}, ns^2$ where $n = 6$ to 7 .
- (c) The elements from atomic number 58 to 71 are called lanthanides because they come after lanthanum (57). The elements from 90 to 103 are called actinides because they come after actinium (89).
- (d) The lanthanides occur in nature in low abundance and therefore, these are called rare earth elements.
- (e) All the actinide elements are radioactive.
- (f) All the elements after atomic number **92 (i.e. U^{92})** are transuranic elements.

METALS, NON-METALS & METALLOIDS

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

- (a) **Metals:** Majority of the elements in periodic table are metals and appears on the left side of the periodic table.
- Properties:
- (i) These are usually solid at room temperature [exception - mercury]
 - (ii) They have high melting and boiling point [exception Gallium & Cesium have very low melting point (303 K and 302 K respectively)]
 - (iii) They are good conductor of heat and electricity.
 - (iv) They are malleable (can be flattened into thin sheets by hammering) and ductile (can be drawn into wires)
- (b) **Non-Metals:** These are placed at the top right hand side of periodic table. As we move horizontally along a period, the property of elements changes from metallic (on left) to non-metallic (on the right).
- Properties:
- (i) These are usually solids or gases at room temperature.
 - (ii) They have low melting point and boiling point (exception : Boron, Carbon).
 - (iii) Most Non-metallic solids are brittle and are neither malleable nor ductile.
- (c) **Metalloids (Semi-metals):** Properties of these elements shows the characteristics of both metals and non-metals. Silicon (Si), Germanium(Ge), Arsenic(As), Antimony(Sb) and Tellurium(Te) are metalloids. Si shows predominantly non - metallic behavior.

PERIODICITY

- (A)** The regular gradation in properties from top to bottom in a group and from left to right in a period is called periodicity in properties.
- (a)** In a period, the ultimate orbit remains same, but the number of electron gradually increases.
- (b)** In a group, the number of electron in the ultimate orbit remains same, but the values of n increases. (n = principal quantum number)
- (B) Causes of periodicity:**
- (a)** The cause of periodicity in properties is due to the same outermost shell electronic configuration coming at regular intervals.
- (b)** In the periodic table, elements with similar properties occur at intervals of 2,8,8,18,18 and 32 . These numbers are called as magic numbers.

SCREENING EFFECT (σ) AND EFFECTIVE NUCLEAR CHARGE (Z_{eff}) :

- (i)** Valence shell electron suffer force of attraction due to nucleus and force of repulsion due to inner shell electrons.
- (ii)** The decrease in force of attraction on valence electron due to inner shell is called screening effect or shielding effect, i.e. total repulsive force is called shielding effect.
- (iii)** Due to screening effect valence shell electron experiences less force of attraction exerted by nucleus. (i.e. total attraction force experienced by valence electron is called Z_{eff}).
- (iv)** There is a reduction in nuclear charge due to screening effect. Reduced nuclear charge is called effective nuclear charge.
- All important calculations of Z_{eff} already done in previous chapter.

From top to bottom in a group Z_{eff} remain constant

Group	Element	Li	Na	K	Rb	Cs	
	z_{eff}	1.30	2.20	2.20	2.20	2.20	
Period	Element	Be	B	C	N	O	F
	z_{eff}	1.95	2.6	3.25	3.90	4.55	5.20

For same shell,

The order of shielding effect $s > p > d > f$ (**due to penetration effect**)

Z_{eff} for different ions of an element:-

$$Z_{\text{eff}} \propto \frac{\text{positive charge}}{\text{negative charge}}$$

(i) Z_{eff} f

or different ions of an element

(ii) Z_{eff} for isoelectronic species(i) Z_{eff} for different ions of an elementOrder of Z_{eff} Ex. $\text{N}^+ > \text{N} > \text{N}^-$ (ii) Z_{eff} of isoelectronic speciesEx. $\text{H}^- < \text{Li}^+ < \text{Be}^{+2} < \text{Be}^{+3}$ (2 electrons species) $\text{N}^{-3} < \text{O}^{-2} < \text{F}^- < \text{Na}^+ < \text{Mg}^{+2}$ (10 electron species)

(iii) Element B Al Ga In

Valence e- Z_{eff} 2.6 3.5 5.0 5.0(iv) Element Sc Ti V Cr Mn Fe Co Ni Cu Zn
Valence e- Z_{eff} 3.0 3.15 3.30 2.90 3.60 3.75 3.90 4.05 3.70 4.35(v) For 4 s and 3 d electron of Sc, Y, La, Z_{eff} is same, i.e. 3.0 .**ATOMIC RADIUS :** The average distance of valence shell electron from nucleus is called atomic radius.

It is very difficult to measure the atomic radius because –

(i) The isolation of single atom is very difficult.

(ii) There is no well defined boundary for the atom. (The probability of finding the electron is 0 only at infinity).

So, the more accurate definition of atomic radius is -

- Half the inter-nuclear distance(d) between two atoms in a homoatomic molecule is known as atomic radius.
- This inter-nuclear distance is also known as bond length. Inter-nuclear distance depends upon the type of bond by which two atoms has combined.
- Based on the chemical bonds, atomic radius is divided into four categories -

(A) Covalent radius

(B) Ionic radius

(C) Metallic radius

(D) Vander Waal's radius

(A) Covalent Radius: (SBCR -Single Bonded Covalent Radius)

(a) Covalent bonds are formed by overlapping of atomic orbitals.

(b) Internuclear distance is minimum in this case.

(c) Covalent radius is the half of the internuclear distance between two singly bonded homo atoms.

Ex. If internuclear distance of A_2 molecule is (A-A) and covalent radius is r then

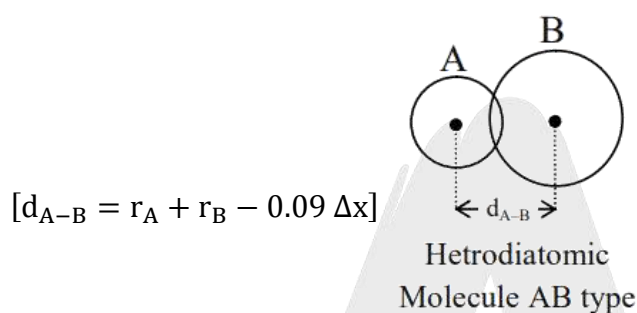
$$d_{A-A} = r_A + r_A \text{ or } 2r_A$$

$$r_A = \frac{d_{A-A}}{2}$$

Ex. In Cl_2 molecule, internuclear distance is 1.98\AA so $r_{\text{Cl}} = \frac{1.98}{2} = 0.99\text{\AA}$.

Note: In a heterodiatomic molecule AB where the electronegativity of atoms A and B are different, the experimental values of internuclear distance d_{A-B} is less than the theoretical values $(r_A + r_B)$.

According to Schomaker and Stevenson -



Where Δx is the difference of electronegativities of the atoms A and B · r_A and r_B are in \AA .

Example: X – X bond length is 1.00\AA and C – C bond length is 1.54\AA . If electronegativities of X and C are 3.0 and 2.0 respectively, then C – X bond length is likely to be? (using Stevenson & schomaker formula).

Sol.

$$r_{C-X} = r_C + r_X - 0.09\Delta x$$

$$= \frac{1.00}{2} + \frac{1.54}{2} - 0.09 [\Delta x = 1]$$

$$= 1.27 - 0.09;$$

Ans. C – X bond length = 1.18\AA

(B) Ionic Radius:

(i) Cationic Radius:

- (a) When a neutral atom loses electron it converts into a cation (positive charged ion)
- (b) Cationic radius is always smaller than atomic radius because after losing electron number of electron reduces, but number of protons remains same, due to this Z_{eff} increases, hence electrons get pulled towards nucleus and atomic radius decreases, moreover after losing all the electrons from the outer most shell, penultimate shell becomes ultimate shell which is nearer to nucleus so size decreases.

(c) Size of cation $\propto \frac{1}{\text{Magnitude of the charge of } Z_{\text{eff}}}$

Ex. (i) $\text{Fe} > \text{Fe}^{+2} > \text{Fe}^{+3}$ (ii) $\text{Pb}^{+2} > \text{Pb}^{+4}$

(iii) $\text{Mn} > \text{Mn}^{+2} > \text{Mn}^{+3} > \text{Mn}^{+4} > \text{Mn}^{+5} > \text{Mn}^{+6} > \text{Mn}^{+7}$

(ii) Anionic Radius :

(a) When a neutral atom gains electron it converts into anion (negative charged ion)

(b) Anionic radius is always greater than atomic radius because in an anion electrons are more than protons so inter electronic repulsion increases, which also increases screening effect (σ) hence effective nuclear charge decreases (Z_{eff}) therefore distance between electron and nucleus increases and size of anion also increases.

Ex. Flourine ($Z = 9$)

	F	F ⁻	
	9	9	
electron	9	10	
so,	$\frac{Z}{e} = \frac{9}{9} = 1$	$\frac{9}{10} = 0.9$	As Z_{eff} of F ⁻ is less than F so F ⁻ > F

(iii) Size of isoelectronic species:

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

Species	K ⁺	Ca ⁺²	S ⁻²	Cl ⁻
Z	19	20	16	17
e	18	18	18	18
$\frac{Z}{e}$	$\frac{19}{18} = 1.05$	$\frac{20}{18} = 1.11$	$\frac{16}{18} = 0.88$	$\frac{17}{18} = 0.94$

Order of radius : $(\text{S}^{-2} > \text{Cl}^{-} > \text{K}^{+} > \text{Ca}^{+2}), (\text{N}^{3-} > \text{O}^{2-} > \text{F}^{-} > \text{Na}^{+} > \text{Mg}^{+2} > \text{Al}^{+3})$

$\underbrace{\hspace{10em}}_{18 e^{-}}$
 $\underbrace{\hspace{10em}}_{10 e^{-}}$

Metallic Radius:

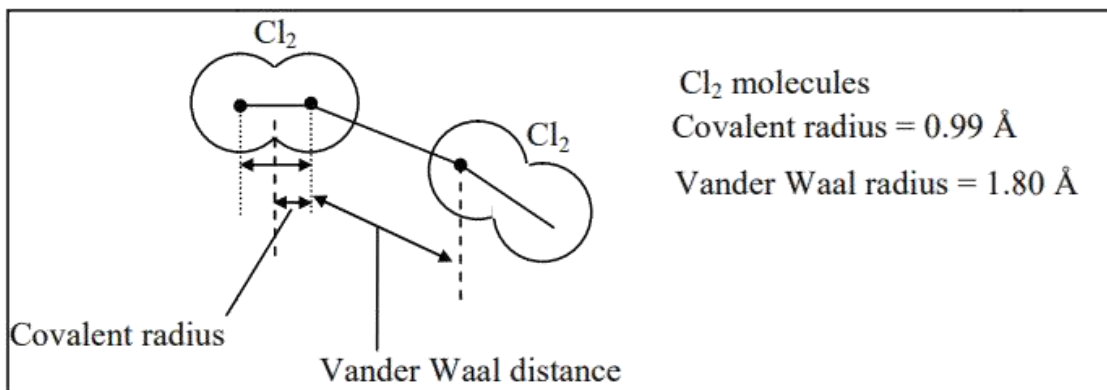
- (a) Half of the inter nuclear distance between two adjacent metallic atoms in crystalline lattice structure.
- (b) There is no overlapping of atomic orbitals, so Metallic radius > Covalent radius

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

(d) Vander Waal's Radius :

- (i) Those atoms (like noble gases) which are not bonded with each other, experiences a weak attractive force to come nearer.
- (ii) Half of the distance between the nuclei of adjacently placed atoms in solid state of a noble gas is Vander Waal's radius.

- (iii) Inert gases have only Vander Waal's radius.
- (iv) In molecules solid nonmetals both covalent and Vander Waal's radius exists.



$$\text{van der Waal's radius} = 2 \times \text{covalent radius}$$

$$\text{van der Waal's radius} > \text{Metallic radius} > \text{Covalent radius}$$

Factors affecting atomic size are:

- (a) Z_{eff} increases, atomic radius decreases
Ex. $\text{Li} > \text{Be} > \text{B} > \text{C} > \text{N} > \text{O} > \text{F}$
- (b) Number of shell(n) increases, atomic radius increases Ex. $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$
- (c) Screening effect increases, atomic radius increases.
- (d) Magnitude of negative charge increases, atomic radius increases
Ex. $\text{O} < \text{O}^- < \text{O}^{2-}$
- (e) Magnitude of positive charge increases, atomic radius decreases
Ex. $\text{Mn} > \text{Mn}^{+2} > \text{Mn}^{+3} > \text{Mn}^{+4}$
- (f) Bond order increases, atomic radius decreases
Ex. $> \text{N} - \text{N} < -\text{N} = \text{N} - < \text{N} \equiv \text{N}$

Periodic variation of atomic size:

- (i) Across a period : It decreases from left to right in a period as effective nuclear charge (Z_{eff}) increases

Ex. $\text{Li} > \text{Be} > \text{B} > \text{C} > \text{N} > \text{O} > \text{F}$

- (ii) In a group : It increases from top to bottom in a group as number of shell increases Ex. $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$

Exceptions: Transition elements

From left to right in a period:

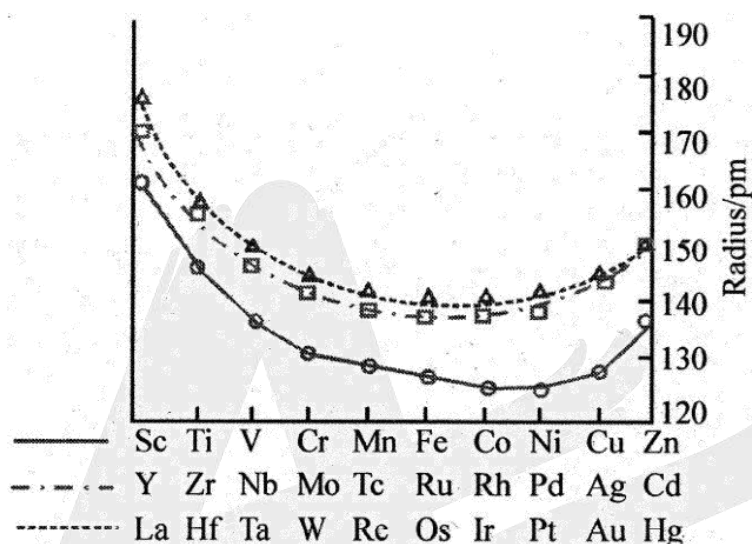
In the first transition series the atomic radii slightly decreases from Sc to Mn because effect of effective nuclear charge is more prominent than the shielding effect. After that it remains almost the same because both the effects balance each other. The atomic size from Cu to Zn slightly increases

because shielding effect is more prominent than effective nuclear charge due to d^{10} configuration of Cu and Zn.

The atomic radii of the elements of 3 d transition series are as under:

Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Atomic radius (pm)	162	147	134	127	126	126	124	124	124	138

Order of covalent radius : $Sc > Ti > V > Cr \approx Mn \approx Fe \approx Co \approx Ni < Cu < Zn$



Trends in atomic radii of transition elements

- **In a group:**
 - (i) The atomic radius of elements increases on moving down the first transition series (3d) to second transition series (4d). This is due to the increases in number of shells with the increase in atomic number.
 - (ii) The atomic radii of second (4d) and third (5d) transition series in a group is almost same except Y(39) and La(57)
 - **In third transition series,** there are fourteen lanthanides in between La(57) of IIIB and Hf (72) of IVB groups, so the atomic radius of Hf (72) decreases much due to lanthanide contraction* in lanthanides. The difference in the nuclear charge in the elements of a group in first and second transition series is +18 units while this difference in second and third transition series is +32 units except Y(39)
 - La (57). Due to the increase of +32 units in the nuclear charge there is a sizeable decrease in the atomic radius which balances the increase in size due to the increase in number of shells.
- So, in a group moving from 5 th period to 6 th period in transition series, the atomic radii of the elements remain almost the same except IIIB. The difference is about 0.02\AA .

(iii) For inner transition elements:

As we move along the lanthanide series, there is a decrease in atomic as well as ionic radius. The decrease in size is regular in ions but not so regular in atoms. This is called lanthanide contraction*.

Lanthanide Contraction :

(a) Outermost electronic configuration of inner transition elements is

$$(n-2)f^{1-14}, (n-1)s^2p^6d^{0-1}, ns^2 (n = 6 \text{ or } 7)$$

(b) Electron enters in $(n-2)f$ orbitals

Because of complicated structure of f-orbital and due to poor shielding of f& d electrons, the outermost shell electrons get highly attracted towards nucleus.

(c) But in lanthanides and actinides, with each electron, nuclear charge increase by +1 so size decreases slowly, this decrement in size is known as lanthanide contraction.

$$\boxed{\text{Nuclear charge} > \text{Screening effect}}$$

So order of radius for transition series

(d) In 1st, 2nd and 3rd transition series

$$3d < 4d \approx 5d \text{ (except III}^{\text{rd}}\text{B)}$$

	IIIB	IVB	Size increases on moving from 3d to 4d series but the size of 4d and 5d series element almost equal due to lanthanide contraction
size increases ↓	Sc	Ti ↓	
	Y	Zr	
	La	Hf	

- Scandide/Transition contraction:**

Note: While atomic size should increase down the group.

IIIB - Atomic size of Ga < Al. size of Al is greater than Ga due to scandide contraction.

(Due to poor shielding of d electrons)

- Different orders of atomic and ionic radii:**

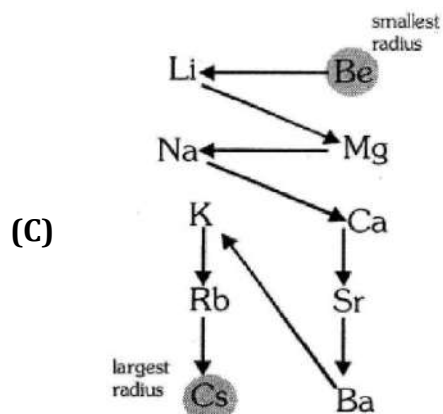
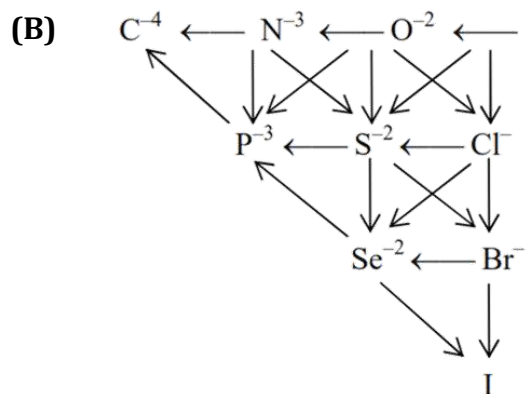
Ni < Cu < Zn	H ⁺ < H < H ⁻	Ti < Zr ≈ Hf	Be < Li < Na
Cr < V < Ti < Sc	I ⁺ < I ₂ I ⁻	Sc < Y < La ≈ Ac	Cl < Na < Rb
Cu < Ag ≈ Au	B < Be < Al < Mg	F < O < Cl < S	

- s-block size variation**

$$H^- > F^-$$

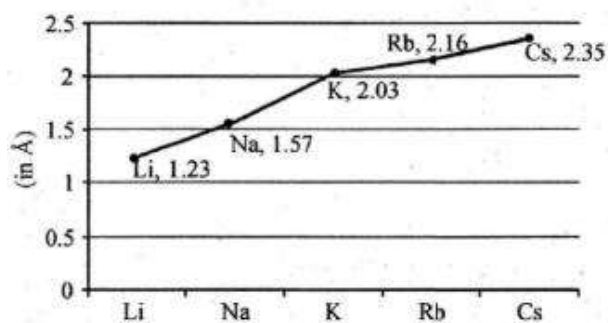
(A) $H^- > Cl^-$

$$H^- > Br^-$$

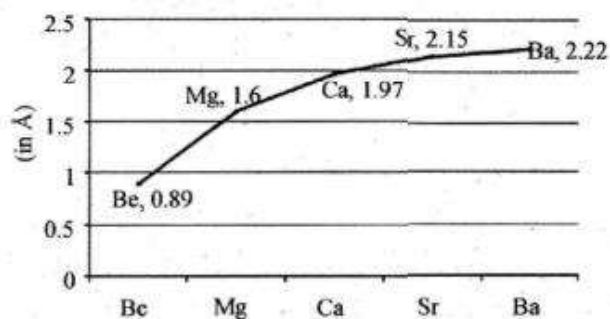


ATOMIC RADIUS (ONLY FOR REFERENCE)

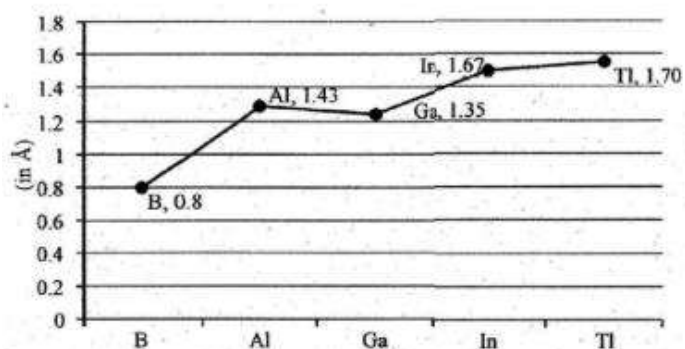
ALKALI METALS



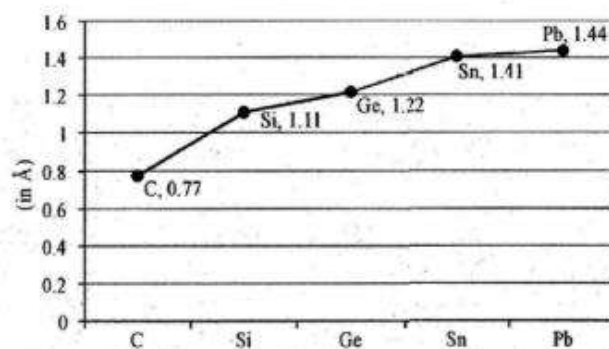
ALKALINE EARTH METALS



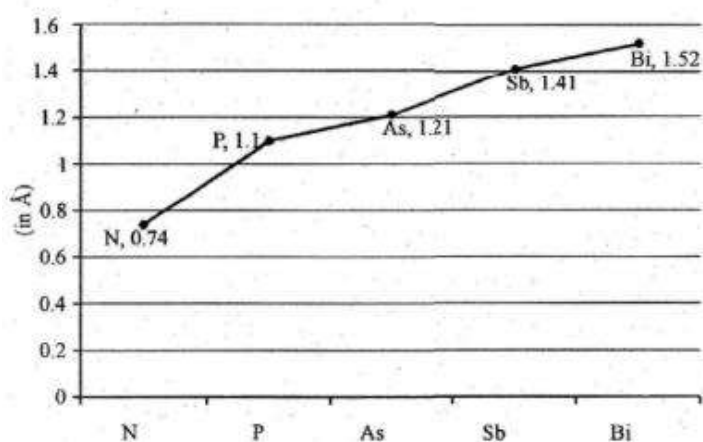
BORON FAMILY



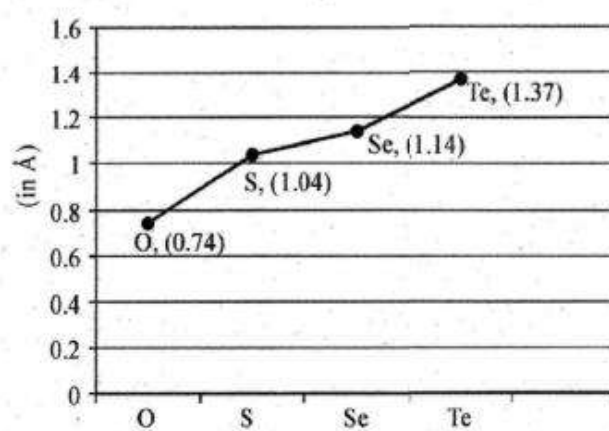
CARBON FAMILY



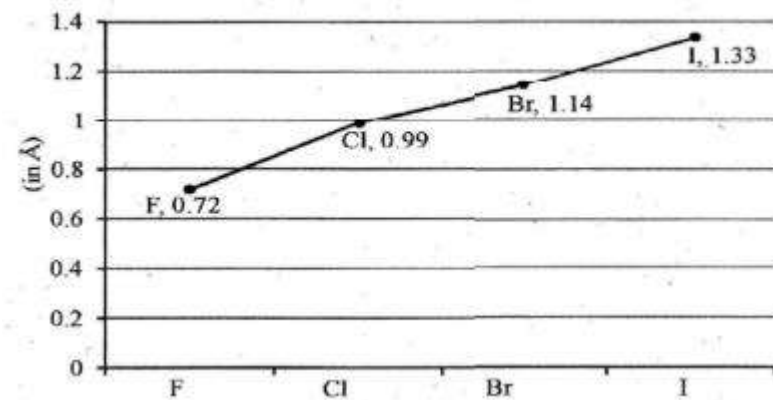
NITROGEN FAMILY (PNICOGENS)



OXYGEN FAMILY (CHALCOGENS)



HALOGENS



ONLY FOR REFERENCE

Group 1		Group 2		Group 13		Group 16		Group 17	
<div> <div>H⁺</div> <div>H</div> <div>35</div> <div>37</div> </div>								<div> <div>H</div> <div>H</div> <div>37</div> <div>208</div> </div>	
Li ⁺	Li	Be ²⁺	Be	B ³⁺	B	O	O ²⁻	F	F ⁻
76	134	45	90	41	82	73	140	71	133
Na ⁺	Na	Mg ²⁺	Mg	Al ³⁺	Al	S	S ²⁻	Cl	Cl ⁻
102	154	72	130	68	118	102	184	99	181
K ⁺	K	Ca ²⁺	Ca	Ga ³⁺	Ga	Se	Se ²⁻	Br	Br ⁻
138	196	100	174	76	126	116	198	114	196
Rb ⁺	Rb	Sr ²⁺	Sr	In ³⁺	In	Te	Te ²⁻	I	I ⁻
152	211	118	192	94	144	135	221	133	220

Sizes of atoms and their ions in pm

Note : In the direction of arrow (→) ionic size increases.

Li ⁺	Be ²⁺	
Na ⁺	Mg ²⁺	Al ³⁺
K ⁺	Ca ²⁺	
Rb ⁺	Sr ²⁺	

O ²⁻	F ⁻
S ²⁻	Cl ⁻
Se ²⁻	Br ⁻
Te ²⁻	I ⁻

Example : Select from each group the species which has the smallest radius stating appropriate reason.

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

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

(b) The ionic radius decreases as more electrons are removed successively. So the correct order is $p^{5+} < p^{4+} < p^{3+}$

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

Solution: Mg^{2+} and O^{2-} both are isoelectronic i.e., have same number of electrons. But Z_{eff} in Mg^{2+} is greater than that in O^{2-} ion.

DO YOURSELF - 3

1. From the given set of species, point out the species from each set having least atomic radius:-

(A) O^{2-}, F^-, Na^+ (B) Ni, Cu, Zn (C) Li, Be, Mg (D) He, Li^+, H^-

Correct answer is

(A) O^{2-}, Cu, Li, H^- (B) Na^+, Ni, Be, Li^+ (C) F^-, Zn, Mg, He (D) Na^+, Cu, Be, He

2. Which has the lowest anion to cation size ratio-

(A) LiF (B) NaF (C) CsI (D) CsF

3. Arrange the elements in increasing order of atomic radius Na, Rb, K, Mg

(A) Na, K, Mg, Rb (B) K, Na, Mg, Rb (C) Mg, Na, K, Rb (D) Rb, K, Mg, Na

4. Which of the following pairs of elements have almost similar atomic radii

(A) Zr, Hf (B) Mo, W (C) Co, Ni (D) All

5. If the ionic radii of K^+ and F^- are nearly the same (i.e. 1.34 Å) then the atomic radii of K and F respectively are

(A) 1.34 Å, 1.34 Å (B) 0.72 Å, 1.96 Å (C) 1.96 Å, 0.72 Å (D) 1.96 Å, 1.34 Å

6. For the element X, student Mansi measured its radius as 102 nm, student Rohit as 203 nm. and Ankur as 100 nm. using same apparatus. Their teacher explained that measurements were correct by saying that recorded values by three students were:-

(A) Crystal, Vander Waal and Covalent radii

(B) Covalent, crystal and Vander Waal radii

(C) Vander Waal, ionic and covalent radii

(D) None is correct.

7. Screening effect is not observed in

- (A) He^+
- (B) Li^{+2}
- (C) H
- (D) All of these

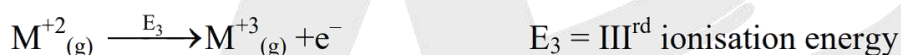
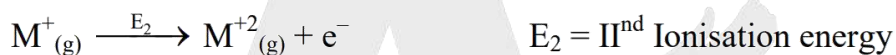
IONISATION POTENTIAL/IONISATION ENERGY/IONISATION ENTHALPY(IP/IE)

(i) Minimum energy required to remove most loosely bonded outer most shell electron in ground state from an isolated gaseous atom is known as ionization energy.

(Isolated \rightarrow Without any bonding with other atom)

(ii) Successive Ionisation Energy

(a) For an atom $M_{(g)}$ successive ionisation energies are as follows -



$$E_1 < E_2 < E_3 \dots \dots \dots \text{ (Always for an element)}$$

(b) Electron cannot be removed from solid state of an atom, it has to be convert into gaseous form.

Energy required for conversion from solid state to gaseous state is called Sublimation energy.

(c) It is always an endothermic process ($\Delta H = +ve$)

(d) It is measured in eV/atom (electron volt/atom) or Kcal/mole or KJ/mole

NOTE: I.P, I.E, ΔH_E are different definition wise. I.P is expressed in volts while I.E. in eV.

FACTORS AFFECTING IONISATION ENERGY

In a period :

(i) **Atomic size:** with increasing size ionisation potential decreases.

(ii) **Effective nuclear charge (Z_{eff})**

$$\text{Ionisation Energy} \propto Z_{\text{eff}} \propto \frac{\text{positive charge}}{\text{negative charge}}$$

Ion with higher oxidation state will have more ionisation energy.

Ex. $\text{Fe}^{+3} > \text{Fe}^{+2} > \text{Fe}$

(iii) Screening effect: Higher is the screening effect on the outermost electrons causes less attraction from the nucleus and can be easily removed, which leads to the lower value of ionisation potential.

Note: Ionisation energy $\propto \frac{1}{\text{atomic size}}$

(v) Penetration power of sub shells

(b) 's' subshell is more closer to nucleus, so more energy will be required to remove e from s-subshell as comparison to p, d & f-subshells.

$$\text{IE}_1 \quad \text{Be} \quad > \quad \text{B}$$
$$\text{IE}_2 \text{ Be} < \text{B}$$

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

Half filled p^3 , d^5 , f^7 or fully filled s^2 , p^6 , d^{10} , f^{14} are more stable than others so it requires more energy.

After losing one e^- , O attains electronic configuration of N, so 1st ionisation energy of O is more than N.

$$\mathbf{IE}_1 \quad \mathbf{N} \quad > \quad \mathbf{0}$$

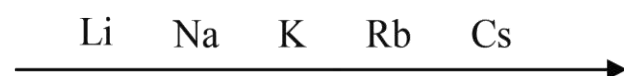
$$\text{Is}^2, 2\text{s}^2, 2\text{p}^2 \quad \text{Is}^2, 2\text{s}^2, 2\text{p}^3$$

(i) In a period : Z_{eff} increases so removal of electron becomes difficult and ionisation energy increases.

Order of IE of IInd period elements

$\text{Li} < \text{B} < \text{Be} < \text{C} < \text{O} < \text{N} < \text{F} < \text{Ne}$

(ii) In a group : Size increases so ionisation energy decreases.



- Ionisation Energy Ga > Al (due to Transition contraction)
- Ionisation Energy of 5 d > 4 d (due to lanthanide contraction)

Ex. Hf > Zr

Illustrations (Optional)

Illustration 1. What is the correct order of ionisation energy :

- (A) $B > Al < Ga > In < Tl$ (B) $B < Al > Ga > In > Tl$
 (C) $B > Al > Ga > In > Tl$ (D) $B > Al < Ga < In < Tl$

Ans. (A)**Illustration 2.** The correct order of decreasing second ionization energy of Li, Be, Ne, C, B

- (A) $Ne > B > Li > C > Be$ (B) $Li > Ne > C > B > Be$
 (C) $Ne > C > B > Be > Li$ (D) $Li > Ne > B > C > Be$

Ans. (D)**Illustration 3.** Which of the following element has highest value of ionisation energy-

- (A) Ti (B) Zr (C) Hf (D) None of these

Ans. (C)**Illustration 4.** What is the correct order of ionisation energy :

- (A) $K < Cu < Cu^+ < K^+$ (B) $K < Cu^+ < Cu < K^+$
 (C) $Cu^+ < K < Cu < K^+$ (D) $K^+ < Cu^+ < Cu < K$

Ans. (A)**(iii) Application of ionization energy :****(A) Metallic and non metallic character :**

Generally for metals Ionization Energy is low.

For Non-metals Ionization Energy is high.

$$\text{Metallic character} \propto \frac{1}{IE}$$

(B) Reactivity of metals :

$$\text{Reactivity of metals} \propto \frac{1}{IE}$$

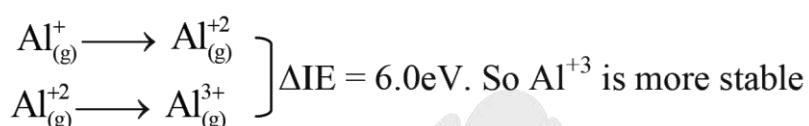
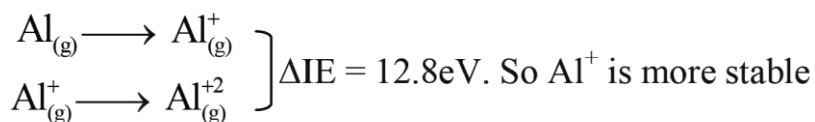
Reactivity of metals increases down the group as ionisation energy decreases.

(C) Stability of oxidation states of an element:**(i)** If the difference between two successive ionisation energies of an element $\geq 16\text{eV}$, then its lower oxidation state is stable.In above example difference between ionisation energy $> 16\text{eV}$, hence Na^+ is more stable.

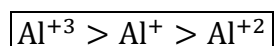
- (ii) If the difference between two successive ionisation energies of an element $\leq 11\text{eV}$, then its higher oxidation state is stable.



In above example difference of ionisation energy $< 11\text{eV}$, So Mg^{+2} is more stable.



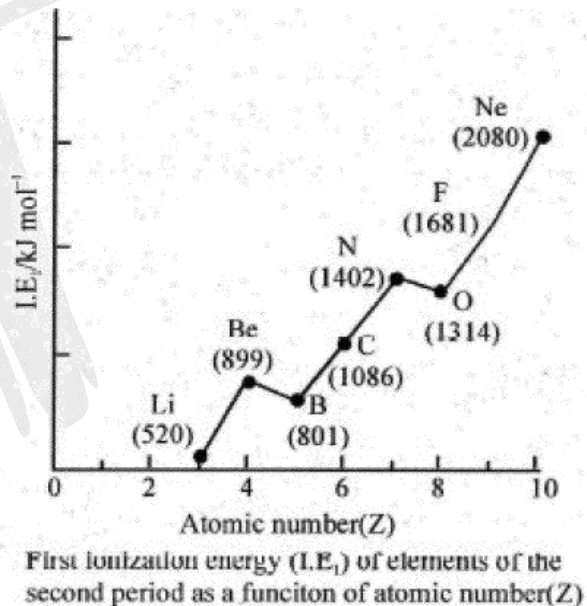
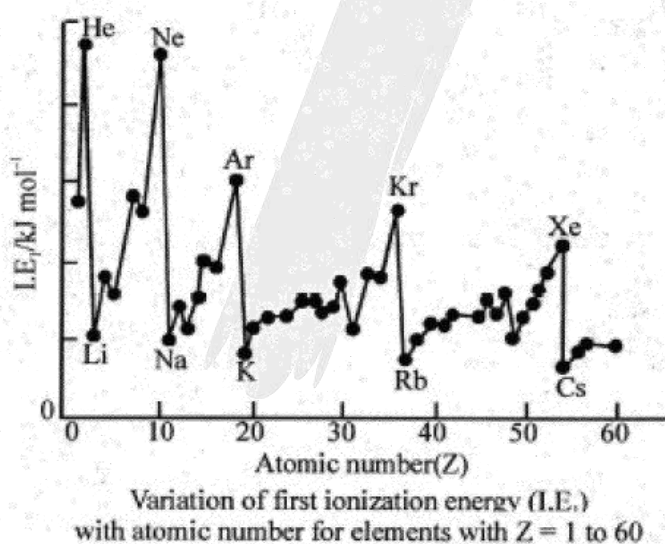
overall order of stability is



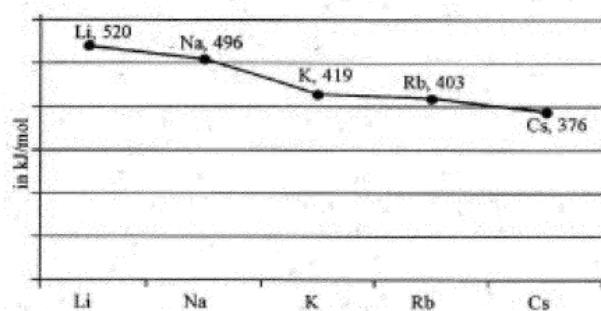
- (D) To determine the number of valence electrons of an element:

Number of valence electrons = number of lower values of ionization energy before 1st highest jump.

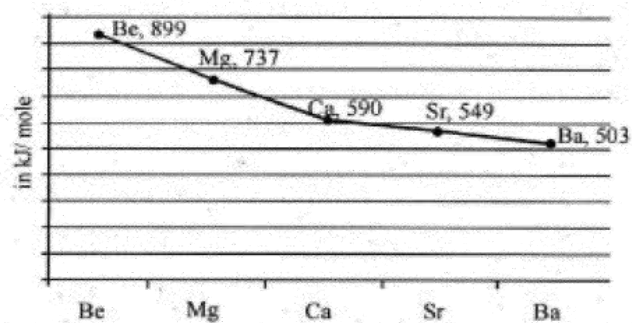
ONLY FOR REFERENCE



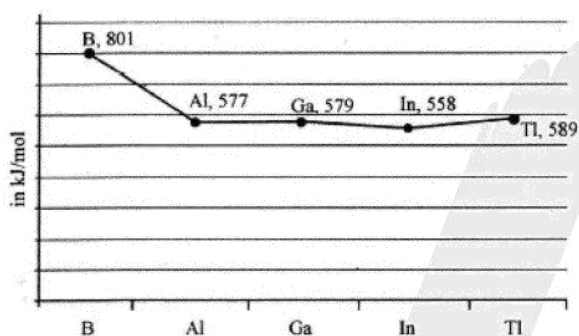
ALKALI METALS



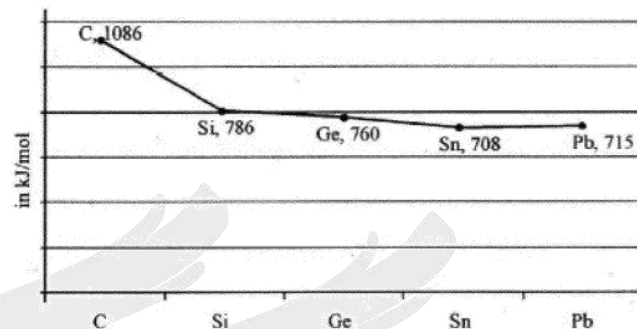
ALKALINE EARTH METALS



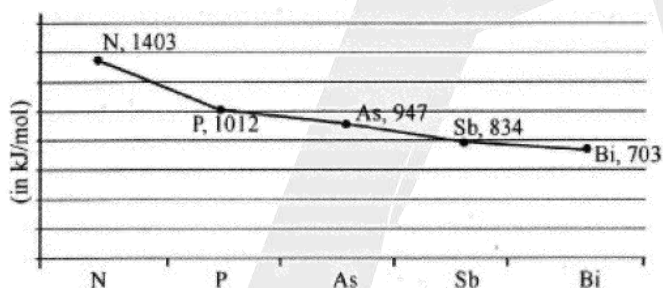
BORON FAMILY



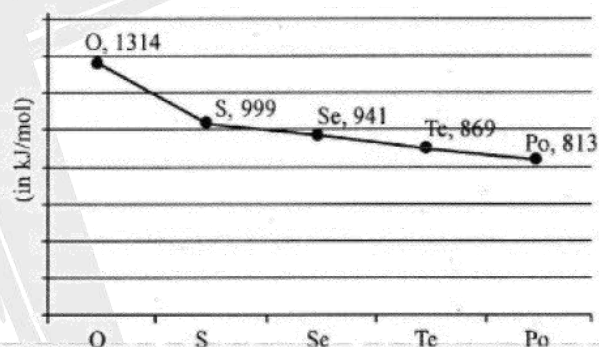
CARBON FAMILY



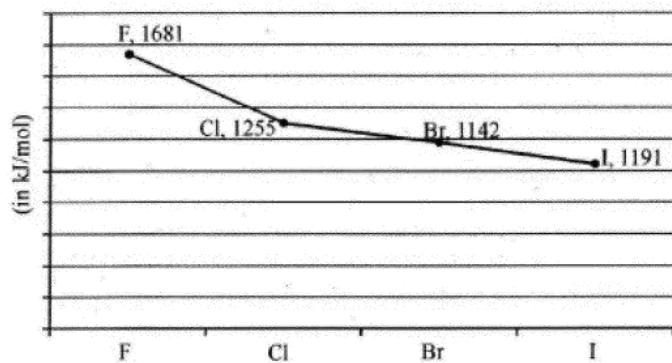
NITROGEN FAMILY (PNICOGENS)



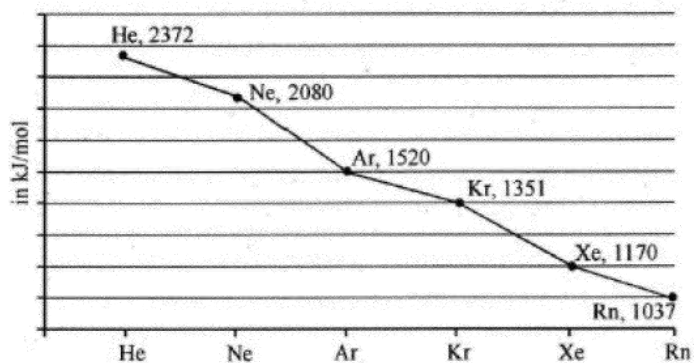
CHALCOGENS



HALOGENS

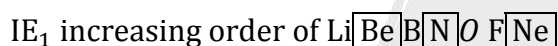
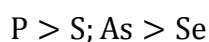
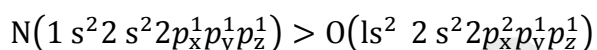


NOBLE GASES



Exception:

- (a) In a period, the ionization energy of IIA group elements is more than the element of IIIA because the penetration power of s-orbital is more. The value of ionization energy of $\text{Be}(1s^2 2s^2)$ is more than $\text{B}(1s^2 2s^2 2p^1)$ because the penetration power of 2s-electrons of Be is more than the 2p electron of B
- (b) In a period, the ionization energy of VA group elements is more than the elements of VIA because the half filled p^3 configuration of VA elements is comparatively of higher stability. VIA group elements (p^4) have the tendency to acquire comparatively more stable (p^3) configuration by the loss of one electron.

Ionization energy order

- (c) **In group IIIA** the ionization potential of Al(13) is nearly equal to the ionization potential of Ga(31) due to transition contraction.
- (d) In group IVA the values of I.P. of Tl(81) and Pb(82) of sixth period is more than that I.P. values of 'In' (49) and Sn (50) of same groups in period fifth. This is because of the electrons are filled in 4 -orbitals for Tl(81) and Pb(82) which do not completely shield the outer electrons. Thus increase in +32 units in nuclear charge results in the increase of ionization potential values.
- (e) On moving from second to third transition series the value of I.P. some what increases except IIB group [Y(39) \rightarrow La(57)]. This is because of 14 electrons are filled in 4f-orbitals of lanthanides which do not shield the 5d electrons effectively. Thus the increase in +32 units in nuclear charge results in the increase of I.P., on moving from left to right this effect decreases and becomes negligible in the later part.

Note:

- (a) Helium (He) has the highest ionisation energy (IE) among all the elements and Cesium (Cs) has the least (IE) value.
- (b) Ionisation potential of inert gases is very high due to most stable $ns^2 np^6$ electronic configuration. ($1s^2$ in He)
- | Element | He | Ne | Ar | Kr | Xe | Rn |
|--------------------------|------|------|------|------|------|------|
| $\text{IE}_1(\text{eV})$ | 24.5 | 21.6 | 15.8 | 14.0 | 12.1 | 10.7 |
- (c) For isoelectronic species I.P. increases with positive charge and decreases with negative charge, e.g. $\text{Al}^{+3} > \text{Mg}^{+2} > \text{Na}^+ > \text{F}^- > \text{O}^{-2} > \text{N}^{-3}$

Application of ionisation potential:**(a) Metallic / Electropositive nature :**

Elements which have tendency to lose electron(s) are known as metals. In general in a group moving from top to bottom metallic or electropositive character increases because I.P. value decreases. In a period moving from left to right the value of I.P. increases so metallic or electropositive character decreases.

(b) Reactivity:

The relative reactivity of the metals increases with the decrease in I.P. values. The I.P. values of IA and IIA metals are comparatively low. So they are comparatively more reactive. The I.P. values of inert gases are very high. So they are almost unreactive.

In a group moving from top to bottom the reactivity of metal atoms increases because their I.P. value decreases.

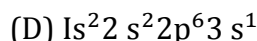
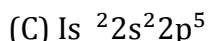
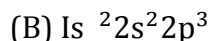
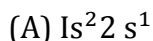
(c) Reducing power:

It is the measure of the tendency to lose electrons. More is the tendency to get itself oxidised and thus more is the reducing nature and therefore reducing power decreases along the period but increases down the group, (with exception lithium which is a powerful reductant due to its high hydration energy).

DO YOURSELF - 4

- IP₁ and IP₂ of Mg are 178 and 348 K. cal mol⁻¹. The enthalpy required for the reaction $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$ is:-
 (A) +170 K. cal (B) + 526 K. cal (C) -170K.cal (D) - 526 K.cal
- The IP₁, IP₂, IP₃, IP₄ and IP₅ of an element are 7.1, 14.3, 34.5, 46.8, 162.2 eV respectively. The element is likely to be:-
 (A) Na (B) Si (C) F (D) Ca
- Which of the following element has 2nd IP < 1st IP
 (A) Mg (B) Ne (C) C (D) None
- In which of the following the energy change corresponds to first ionisation potential only
 (A) $\text{X}_{(\text{g})} \rightarrow \text{X}^+ + \text{e}^-$ (B) $\text{X}_{2(\text{g})} \rightarrow \text{X}_{(\text{g})}^+ + \text{e}^-$
 (C) $\text{X}(\text{s}) \rightarrow \text{X}^+(\text{g}) + \text{e}^-$ (D) $\text{X}_{(\text{aq})} \rightarrow \text{X}_{(\text{aq})}^+ + \text{e}^-$
- In the given process which oxidation state is more stable.
 $\text{M}_{(\text{g})} \rightarrow \text{M}_{(\text{g})}^+ + \text{e}^- \text{IE}_1 = 7.9\text{eV}$
 $\text{M}_{(\text{g})}^+ \rightarrow \text{M}_{(\text{g})}^{+2} + \text{e}^- \text{IE}_2 = 15.5\text{eV}$
 (A) M⁺ (B) M⁺² (C) Both (D) None

6. The electronic configuration of some neutral atoms are given below:-



In which of these electronic configuration would you expect to have highest

(i) IE_1 (ii) IE_2

(A) C, A

(B) B, A

(C) C, B

(D) B, D

7. Match the column.

Column-I**Valence electronic configuration**(a) ns^1 (b) ns^2 (c) $ns^2 np^1$ (d) $ns^2 np^2$ **Column-II****Successive ionisation energies**

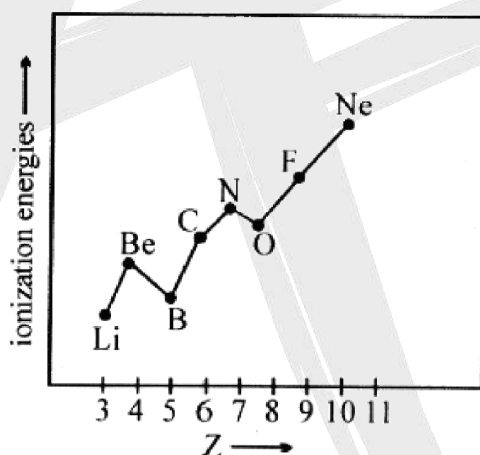
(p) 19, 27, 36, 48, 270

(q) 16, 28, 34, 260

(r) 18, 26, 230, 250

(s) 14, 200, 220, 240

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



(A) above Ne

(B) below Ne but above O

(C) below Li

(D) between N and O

Sol. Na is 3rd period element and is larger than Li. Thus the outer most electron is loosely bounded with nucleus and removal of electron is easier. So option (C) is correct.

Ex. $M(g) \rightarrow M^+(g) + e^-$; $\Delta H = 100\text{eV}$

$M(g) \rightarrow M^{2+}(g) + 2e^-$; $\Delta H = 250\text{eV}$

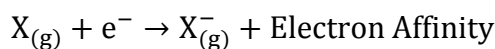
Which is/are correct statement(s)?

(A) IE_1 of $M(g)$ is 100eV(B) IE_1 of $M^+(g)$ is 150eV(C) IE_2 of $M(g)$ is 250eV(D) IE_2 of $M(g)$ is 150eV

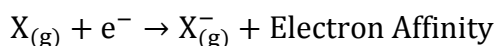
Ans. (A, B, D)

ELECTRON AFFINITY/ELECTRON GAIN ENTHALPY ($EA/\Delta H_{eg}$)

- (A) The amount of energy released when an electron is added to the valence shell of an isolated gaseous atom known as Electron affinity.

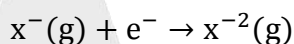


- (B) Generally first electron addition in an isolated gaseous atom is an exothermic process (except stable electronic configuration)



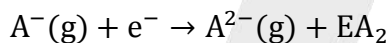
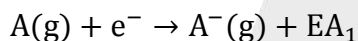
$$\Delta H_{eg} \text{ (first electron gain enthalpy)} = (-) \text{ negative}$$

- (C) Second electron addition in an isolated gaseous atom is always an endothermic process due to inter electronic repulsion.



$$\Delta H_{eg} \text{ (second electron gain enthalpy)} = (+) \text{ positive}$$

- Successive Electron affinities:** Electron affinity just defined is actually first electron affinity since it corresponds to the addition of one electron only. In the process of adding further electron, the second electron will be added to gaseous anion against the electrostatic repulsion between the electron being added and the gaseous anion. Energy instead of being released is supplied for the addition of an electron to an anion. The sum of EA_1 and EA_2 is positive (energy required).



$$EA_1 \text{ is energy released (generally)} \quad \therefore \Delta H_{EG1} = -Ve \text{ (generally)}$$

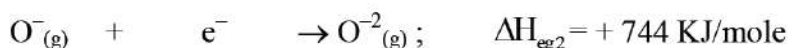
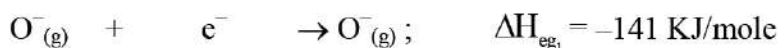
$$EA_2 \text{ is energy required (always)} \quad \therefore \Delta H_{EG2} = +Ve \text{ (always)}$$

$$(EA_1 + EA_2) \text{ is energy required (always)} \quad \therefore \Delta H_{EG1} + \Delta H_{EG2} > 0$$

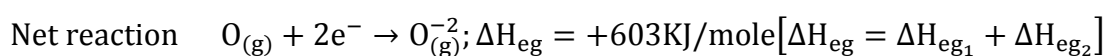
Note: For isolated gaseous atom.

- (i) Except ΔH_{EG1} other electron gain enthalpies i.e. $\Delta H_{EG2}, \Delta H_{EG3} \dots \dots \Delta H_{EGn}$ are positive

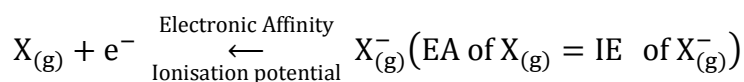
- (ii) $|\Delta H_{EG1}| < |\Delta H_{EG2}|$



Ex.

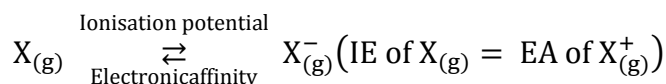


- (D) Formation of polynegative anion like O^{-2}, N^{-3}, C^{-4} etc. is always an endothermic process.
- (5) Electron affinity of neutral atom is equal to ionisation energy of its anion.



Note: [EA of (X) = Zeroth ionisation energy]

(6) IE of neutral atom is equal to electron affinity of its cation



• Factors Affecting the magnitude of electron gain Enthalpy:

- (a) Atomic Size:** In general electron gain enthalpy value becomes less negative with the increasing atomic radius since electrostatic force of attraction decreases between the electron being added and the nucleus due to increase in distance between them.
- (b) Effective Nuclear Charge:** As effective nuclear charge increases, amount of energy released is more. Hence electron gain enthalpy value becomes more negative.
- (c) Screening or Shielding Effect:** Electron gain enthalpy value of the elements becomes less negative with the increasing shielding or screening effect.
- (d) Stability of Half-filled and Completely filled Orbitals:** Half-filled orbitals of a sub shell is comparatively more stable, so it is difficult to add electron in such orbitals and lesser energy is released on addition of electron hence the electron gain enthalpy will decrease.

Ex. (i) $\text{EA} \Rightarrow ns^1 > np^1$

(ii) $\text{EA} \Rightarrow np^2 > np^3$

(7) Variation of electron affinity : (Does not follow any regular order)

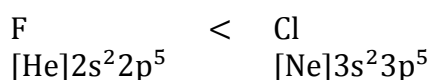
(a) In 2nd period -

Li Be B C N O F Ne

For Be, N, Ne, $\Delta H_{\text{eg}} > 0$.

(b) In Group :

Electron affinity of 3rd period element is greater than electron affinity of 2nd period elements of the respective group.



Due to small size of fluorine, electron density around the nucleus increases. The incoming electron suffers more repulsion. In case of chlorine, electron density decreases due to large size.

Decreasing order of electron affinity

Cl > F > Br > I	S > O > P > N
S > Se > Te > Po > O	Si > C > P > N
Sb > Bi > As > P > N	

Note from NCERT: If energy is released when an electron is added to an atom, the electron affinity is taken as positive, contrary to thermodynamic convention. If energy has to be supplied to add an electron to an atom, then the electron affinity of the atom is assigned a negative sign. However, electron affinity is defined at absolute zero, and therefore, at any other temperature (T), it is given by the equation:

$$\Delta H_{eg} = -E.A. - \frac{5}{2}RT$$

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

Sol.: Order of increasing negative electron gain enthalpy is $N < P < O < S$.

Ex.: Why do halogens have high electron gain enthalpies?

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

Ex.: Which will have the maximum value of electron affinity O^x, O^y, O^z

[x, y and z respectively are 0, -1 and -2]

(A) O^x (B) O^y (C) O^z (D) All have equal

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

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

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

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

$$\frac{5.0 \times 10^{-13} \times 6.023 \times 10^{23}}{10^6} = -30.1 \times 10^4 \text{ J} = -301 \text{ kJ/mol}$$

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

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

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

DO YOUR SELF - 5

- In the formation of a chloride ion, from an isolated gaseous chlorine atom, 3.8eV energy is released, which would be equal to :-
 (A) Electron affinity of Cl^- (B) Ionisation potential of Cl
 (C) Electronegativity of Cl (D) Ionisation potential of Cl^-
- Electron addition would be easier in :-
 (A) O (B) O^+ (C) O^- (D) O^{+2}
- $\text{O}_{(\text{g})} + 2e^- \rightarrow \text{O}^{2-}(\text{g}) \Delta H_{\text{eg}} = 603\text{KJ/mole}$. The positive value of ΔH_{eg} is due to :-
 (A) Energy is released to addition of one electron to O^{-1}
 (B) Energy is required to addition of one electron to O^{-1}
 (C) Energy is needed to addition of one electron to O
 (D) None of the above is correct
- The electron affinity values for the halogens shows the following trend :-
 (A) $\text{F} < \text{Cl} > \text{Br} > \text{I}$ (B) $\text{F} < \text{Cl} < \text{Br} < \text{I}$
 (C) $\text{F} > \text{Cl} > \text{Br} > \text{I}$ (D) $\text{F} < \text{Cl} > \text{Br} < \text{I}$
- The process requiring absorption of energy is.
 (A) $\text{F} \rightarrow \text{F}^-$ (B) $\text{Cl} \rightarrow \text{Cl}^-$ (C) $\text{O} \rightarrow \text{O}^{2-}$ (D) $\text{H} \rightarrow \text{H}^-$
- Choose correct relationship -
 (A) $\Delta H_{\text{eg}}(\text{A}^+) = -\text{IE}_1(\text{A})$ (B) $\Delta H_{\text{eg}}(\text{A}) = -\text{IE}_1(\text{A})$
 (C) $\Delta H_{\text{eg}}(\text{A}^-) = \text{IE}_1(\text{A}^{2-})$ (D) $\Delta H_{\text{eg}}(\text{A}) = \text{IE}_1(\text{A}^+) + \text{IE}_2(\text{A})$

ELECTRONEGATIVITY (EN)

- (i) The tendency of a covalently bonded atom to attract shared pair of electrons towards itself is called electronegativity.
- (ii) A polar covalent or ionic bond of A - B may be broken as
- (a) $A - B \rightarrow A^- + B^+$ (Electronegativity A > Electronegativity B)
- (b) $A - B \rightarrow A^+ + B^-$ (Electronegativity A < Electronegativity B)

Depending on their tendency to attract bonded electron.

Difference between electronegativity and Electron Affinity :

Electronegativity	Electron Affinity
<ul style="list-style-type: none"> - Tendency of an atom in a molecule to attract the bonded electrons - It is not an energetic term. - It regularly increases in a period because it does not depend on stable electronic configuration - It has no unit - It is a relative term 	<ul style="list-style-type: none"> - Energy released when an electron is added to neutral isolated gaseous atom - It is an energetic term. - It does not increase regularly in a period because it depends on stable electronic configuration. - It is measured in eV/atom or KJ/mol - Kcal mol^{-1} - It is an absolute term

- (iii) EN was explained by Pauling for the first time. Electronegativity of some other elements are as follows –

Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	H 2.1
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	F 4.0
K 0.8						Br 2.8
Rb 0.8						I 2.5
Cs 1.0						
Fr 0.7						

In Pauling's scale, elements having
Almost same electronegativity are –

$$C \cong S \cong I = 2.5$$

$$N \cong Cl = 3.0$$

$$P \cong H = 2.1$$

$$Cs \cong Fr = 0.7$$

$$Be \cong Al = 1.5$$

Note : Small atoms normally have more electronegativity than large atoms.

- (iv) Factors Affecting electronegativity:

- (a) **Charge on atom:** The cation will be more electronegative than parent atom. As the positive charge on the cation increases electronegativity increases.

Order of E.N.

$$\text{Ex. } Fe^{+3} > Fe^{+2}$$

The anion will be less electronegative than atom. As the negative charge on the anion

Increases electronegativity decreases.

Ex. $O^{2-} < O^- < O$

(b) Ionization energy and electron affinity: Higher the ionization energy and electron affinity of an atom, greater will be its electronegativity.

(c) Effect of Substitution: The electronegativity of an atom depends upon the nature of substituent attached to that atom. For example - Carbon atom in CF_3I acquires greater positive charge than in CF_3I . Hence, C atom in CF_3I is more electronegative than in CH_3I .

(d) Hybridisation state: s-electrons are more penetrated (on attracted) than p-electron towards the nucleus. So, as the s-character in hybrid orbital increases, electronegativity increases.

(e) **Atomic size**

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

Ex. $F > Cl > Br > I$

(f) **Effective nuclear charge (Z_{eff})**

$$\text{Electronegativity} \propto Z_{\text{eff}} \propto \frac{\text{positive charge}}{\text{negative charge}}$$

Ex. $Mn^{+2} < Mn^{+4} < Mn^{+7}$

$O^{-2} < O^{-1} < O < O^{+1} < O^{+2}$

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

----->

$Z_{\text{eff}} \uparrow \text{EN} \uparrow$

(v) **Variation in Electronegativity:**

(a) Electronegativity decreases down the group.

(b) In period on moving from left to right electronegativity increases.

(c) Electronegativity of Cs and Fr are equal, it is because from ${}_{55}\text{Cs}$ to ${}_{87}\text{Fr}$ only one shell increases but nuclear charge (No. of proton) increases by +32, so effect of nuclear charge balance the effect of increase in number of shell.

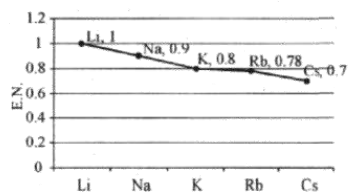
Note:- **Electronegativity of F > Cl but Electron affinity of Cl > F**

(d) In IIIA group, value of electronegativity is irregular when going down the group, because of transition contraction

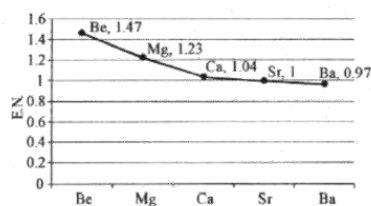
Note:- **Electronegativity of Ga > Electronegativity of Al**

ELECTRONEGATIVITY ONLY FOR REFERENCE

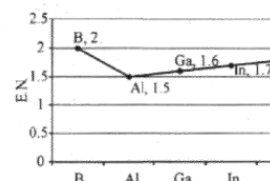
ALKALI METALS



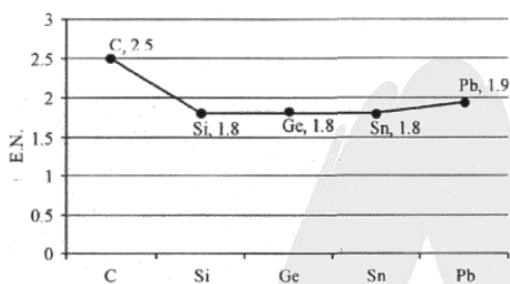
ALKALINE EARTH METALS



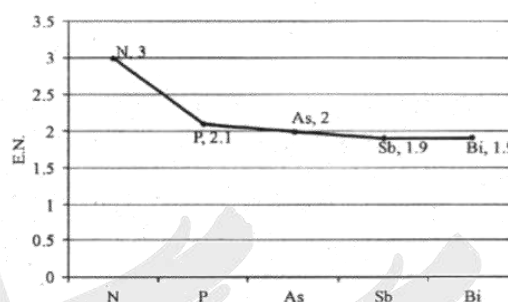
GROUP III A



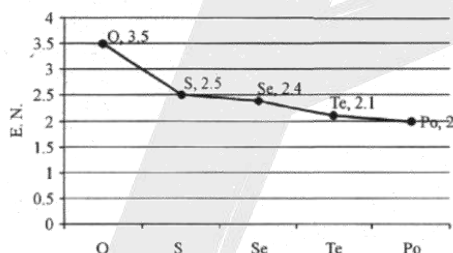
CARBON FAMILY



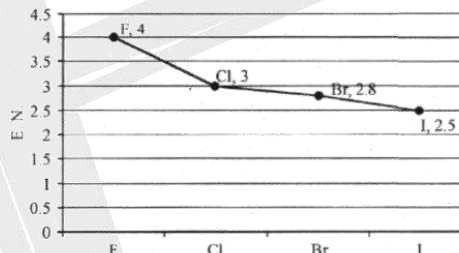
NITROGEN FAMILY (PNICOGENS)



CHALCOGENS



HALOGENS



OTHER SCALES OF ELECTRONEGATIVITY :

Mulliken scale: According to Mulliken electronegativity is average value of ionisation potential and electron affinity of an element,

$$X_m = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{2}$$

$$X_p = \frac{X_m}{2.8}$$

X_p = Pauling's EN where (EN = electronegativity)

X_m = Mulliken's EN

- If ionisation potential and electron affinity are given in eV, then electronegativity by Mulliken on Pauling scale will be

$$X_p = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{5.6}$$

- If ionisation potential and electron affinity are given in K. cal/mole then

$$X_p = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{2 \times 64.5}$$

Allred Rochow's Electronegativity:

The electronegativity of an element is the electrostatic force of attraction between the electron present on the circumference of outermost shell of this atoms and the atomic nucleus. If the distance between the circumference of outermost shell & the nucleus is r and the effective nuclear charge Z_{eff} then -

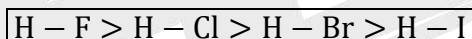
$$\chi_{\text{AR}} = \frac{Z_{\text{eff}} e^2}{r^2} = \frac{0.359 Z_{\text{eff}}}{r^2} \text{ and } Z_{\text{eff}} = Z - \sigma$$

Z = The actual number of charge present in the nucleus i.e. number of protons and σ = shielding constant.

Note:-(I) Z_{eff} to be calculated for $(x + 1)^{\text{th}}$ valence e^- . (B) $X_p = X_{\text{AR}} + 0.744$

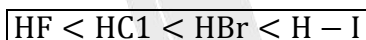
APPLICATIONS OF ELECTRONEGATIVITY

- (a) **Bond strength** : If the electronegativity difference of covalently bonded atoms (ΔX) increases, the bond energy of the covalent bond also increases. For example the order of the $H - X$ bond strength is



As the bond strength decreases, the acidic strength increases.

Order of increasing acidic strength is:-



- (b) **Metallic and non metallic properties of elements :**

- (i) The metallic character decreases as the electronegativity of the element increases.
- (ii) On moving from left to right in a period, the electronegativity of the elements increases.
So the metallic character decreases.
- (iii) On moving down a group, the electronegativity of the elements decreases, so the metallic character increases.

- (c) **Schomaker and Stevenson law**

If in a diatomic molecule electronegativities difference of $A - B$ is more then actual bond length will be reduced. As per Schomaker and Stevenson- The reduction in bond length depends on the difference in electronegativities of atoms in following manner -

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

(Here X_A is E.N. of A & X_B is E.N. of B)

Illustrations (Optional)

Illustration. If bond length of $F_2 = 1.44\text{\AA}$, Bond length of $H_2 = 0.74\text{\AA}$. Find out the bond length of $H - F$? (EN of F is 4.0, EN of H is 2.1)

Solution. $d_{H-F} = r_F + r_H - 0.09(X_F - X_H)$
 $\therefore d_{F-F} = 1.44/2 = 0.72\text{\AA}, d_{H-H} = 0.74/2 = 0.37\text{\AA}$
 $\therefore d_{H-F} = 0.72 + 0.37 - 0.09(4.0 - 2.1)$
 $= 1.09 - (0.09 \times 1.9) = 1.09 - 0.171 = 0.919\text{\AA}$

(d) Acidic & Basic Strength:**(i) Nature of hydrides :**

Stability of molecule \propto Bond energy (strength)

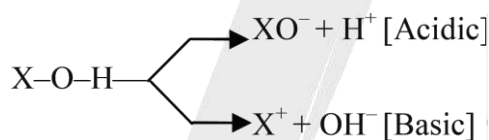
Order of stability of hydrohalides $HF > HCl > HBr > HI$

Order of acidic strength $HF < HCl < HBr < HI$

In VA group $NH_3 < PH_3 < AsH_3 < SbH_3 < BiH_3$

Thermal stability decreases

Acidic character increases

(ii) Nature stability decreases :

As per Gallis, In AOH if electronegativity of A is more than 1.7 (Non metal) then it is acidic in nature.

- If electronegativity of 'A' is less than 1.7 (metal) then AOH will be basic in nature

Ex.	NaOH	ClOH
X_A	0.9	3.0
Nature	Basic	Acidic

- If $X_A - X_O \geq X_O - X_H$ then AO bond will be more polar and will break up as
 $A - O - H \rightarrow A^+ + OH^-$ It shows basic nature
- If $X_A - X_O \leq X_O - X_H$ then OH bond will be more polar and will break up as
 $A - O - H \rightarrow H^+ + AO^-$ It shows basic nature

Ex. In NaOH

$X_O - X_{Na}(2.6) > X_O - X_H (1.4)$ So, NaOH is basic

Ex. In ClOH

$X_O - X_{Cl}(0.5) < X_O - X_H(1.4)$ So, ClOH is acidic

(iii) Nature of oxides :

- Along a period acidic nature increases.
- Down the group basic nature increases

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
←			→			
Basic		Amphoteric		Acidic		

when in periodic table the distance between the element and oxygen increases, basic character increases.

eg. $NO_2 > ZnO > K_2O$

→
acidic character decreases

Note: $BeO, Al_2O_3, ZnO, SnO, PbO, SnO_2, PbO_2, Sb_2O_3$ etc. are amphoteric oxides. CO, H_2O, NO, N_2O etc. are neutral oxides.

Acidic strength of oxide and oxyacid \propto electronegativity

B_2O_3

CO_2

N_2O_5

EN increase, acidic nature increase.

$HNO_3 > H_2CO_3 > H_3BO_3$

$HClO_4 > HBrO_4 > HIO_4$

$SO_3 > SeO_3 > TeO_3$

$H_3PO_4 > H_3AsO_4 > H_3SbO_4$

$N_2O_5 > P_2O_5 > As_2O_5$

$H_2SO_3 > H_2SeO_3 > H_2TeO_3$

$HOI > HOCl > HOBr > HOI$

$N_2O_3 > P_2O_3 > As_2O_3 > Sb_2O_3$

Acidic nature \propto oxidation state

Acidic properties increases with increasing oxidation state of an element

$HClO_4 > HClO_3 > HClO_2 > HClO$

$HNO_3 > HNO_2$

$H_2SO_4 > H_2SO_3$

$N_2O_5 > N_2O_3$

$SO_2 > SO$

$Sb_2O_5 > Sb_2O_3$

Note from NCERT

Basic character oxides increases down the group and decreases in period from left to right.

Where as acidic character decreases down the group and from left to right it increases. With

increase in oxidation number acidic character increases and basic character decreases. The

oxides formed by element on extreme left is most basic (eg: Na_2O) where as on extreme right is the most acidic (eg. Cl_2O) oxides of elements in the centre are amphoteric (eg. Al_2O_3, As_2O_3) or neutral (eg. CO, NO, N_2O).

Increasing Acidic order $-Al_2O_3 < Si_2 < P_2O_5 < SO_3 < Cl_2O_7$. (left to right)

Oxide of Al, Ga, Zn, Sn, As (III), Sb (III), Cr (III), Mn (IV), Pb, Be are amphoteric.

$MnO < Mn_2O_3 < MnO_2 < MnO_3 < Mn_2O_7$

(Basic Basic) (Amphoteric) (Acidic Most acidic)

Non metal form acidic oxide and metal form basic oxides.

Neutral oxide: Oxides having no acidic or basic properties, e.g.: CO, N₂O, NO, H₂O.

Metal oxides in their higher oxidation state like CrO₃, Mn₂O₇ are acidic.

Amphoteric oxides: Oxides which have both acidic and basic nature, e.g.: Al₂O₃, As₂O₃, ZnO, V₂O₅ etc.

(e) Nature of bonds :

- According to Hanny & Smith formula

$$\% \text{ ionic character} = 16(X_A - X_B) + 3.5(X_A - X_B)^2$$

Here X_A = Electronegativity of A

X_B = Electronegativity of B

If $X_A - X_B \geq 2.1$ = Ionic % > 50% i.e. Ionic bond

If $X_A - X_B \leq 2.1$ = Ionic % < 50% i.e. covalent bond

- According to Gallis

$X_A - X_B \geq 1.7$ Ionic

$X_A - X_B \leq 1.7$ Covalent

If $X_A = X_B$; then A – B will be non polar.

Ex. H – H, F – F

If $X_A > X_B$ and difference of electronegativities is small then

$A^{\delta-} - B^{\delta+}$ bond will be polar covalent Ex. H₂O ($H^{\delta+} - O^{\delta-} - H^{\delta+}$)

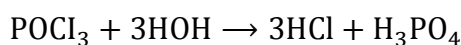
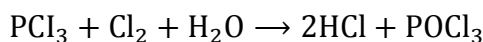
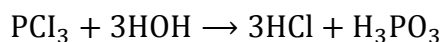
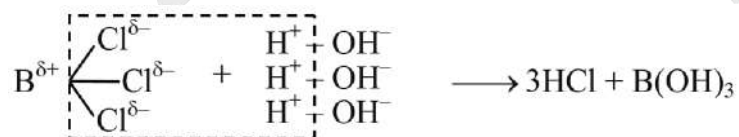
If $X_A \gg X_B$ and $X_A - X_B$ difference of electronegativities is high then $A^- + B^+$ bond will be polar or ionic

(g) Hydrolysis products of AX:

Where A = Other element and X = Halogen

- (i)** If electronegativity of X > Electronegativity of A then on hydrolysis product will be HX

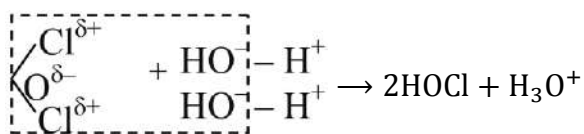
Ex (BCl₃), Electronegativity of Cl > Electronegativity of B



- (ii)** If electronegativity of X < electronegativity of A then on hydrolysis product will be HOX (hypohalous acid)



So On hydrolysis –



DENSITY (OPTIONAL)

(a) In a group volume of an atom increases along with atomic weight but atomic weight increases more than atomic volume, So density increases in a group.

(b) Factors affecting density are

(i) Inner shell configuration

(ii) Atomic weight

(iii) Packing capacity

(iv) Chemical bond

• Periodic variation :

(a) In period density first increases till maximum and then decreases, (s-block to d-block increases, d-block to p-block decreases)

(b) In group from top to bottom density increases regularly

Ex. In VIIA group - F and Cl are gases (Low density)

Br is liquid (density 3.19 g/cm^{-3})

I is solid (density 4.94 g/cm^{-3})

From s-block to p-block packing capacity and strength of bond increases.

• Exceptions

(a) The density of sodium and magnesium is greater than K and Ca respectively. This can be explained by inner shell configuration.

In K, 3d orbital is vacant, 3rd orbit has the capacity to hold 18 electron but it has only 8 electron.

Due to empty orbitals atomic volumes increases and so density decreases

(b) Li is the lightest metal known.

(c) Highest density of solid metal Ir (22.63 gm/cc) and Os (22.6 gm/cc) and liquid metal Hg (13.6 gm/cc)

• Order of density of elements

Li < K < Na < Rb < Cs	Sc < Y < La	Hg < Au < Os < Ir	Cr < Fe < Cu
Ca < Mg < Be < Sr < Ba	Ti < Zr < Hf	Sc < Zn < Fe < Hg	Fe < Cu < Pb
He < Ar < Kr < Rn	Zn < Cd < Hg		Fe < Hg < Au

BOILING POINTS AND MELTING POINTS

(a) It is a property of aggregate of atoms and not of a single atom, i.e. why it is a molecular property.

(b) Melting point of a solid depend on –

- (i) Structure of solid (ii) Chemical bond (iii) Bond energy

(c) Structure of solid made up of (i) atomic solid (ii) molecular solid (iii) metallic solid.

Order of decreasing melting point is: Atomic solid > Metallic solid > Molecular solid

(d) Chemical bond among particles in solid. Order of bond strength is -

Covalent bond > Metallic bond > Vander waal bond

(e) Bond energy - Covalent solids like diamond, SiO_2 etc have only covalent bonds between atoms so their bond energy is higher than molecular or metallic solids.

$$\text{Fe} < \text{Hg} < \text{Au}$$

• **Periodic variation:**

In period from left to right boiling point and melting point first increases then decreases.

Alkali metals - Crystal structure BCC (low boiling point & melting point)

Transition metals - Crystal structure FCC (High boiling point & melting point)

[Inert gases.....Lowest boiling point & melting point (Vander waal force)
[Transition element.....Highest melting point

Metals [W(Tungston) maximum melting point (3410°C)
[Hg (Mercury) lowest melting point (-38°C)

Non metals [Carbon (In the form of diamond) Highest melting point (3727°C)
[Helium Lowest melting point (-270°C)

• **In Group:**

(a) In s-block elements boiling point & melting point decreases down the group.

Li, Na (solid) \rightarrow Cs, Fr (liquid)

It is due to weak cohesive energy of metallic bond in IA.

(b) In d-block elements boiling point & melting point increases down the group (due to lanthanide contraction, z_{eff} increases and hence bond energy increases)

(c) **In p-block elements**

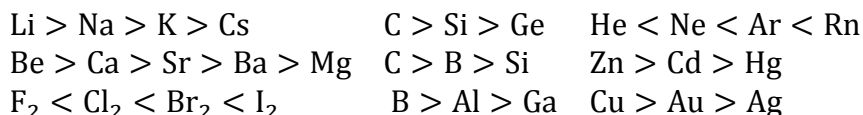
(i) From III A - IVA group boiling point & melting point decreases down the group and from VA to '0' group, boiling point & melting point increases down the group.

(Atomic or molecular weight \propto vander waal force)

(ii) boiling point and melting point of monoatomic molecules are lesser than diatomic molecules. '0' group < Halogens

(iii) Atomic solid non metals like B, C and Si has higher boiling point and melting point due to strong covalent bond.

(iv) Boiling point & melting point of molecular solids are less because of weaker vander waal force among molecules Ex. I_2 .

Order of melting point**(f) Nomenclature of inorganic compounds :**

Prefix-less electronegative element

Suffix-More electronegative element

Ex. Cl₂O (Right) OCl₂, (Wrong)

In Dichloroxide the electronegativity of Cl is less than 'O' So Cl is in prefix position.

Other example:-

OF₂, Oxygen difluoride

ICl Iodine chloride

(g) Bond polarity $\propto \Delta EN$ **Ex.** HF > HCl > HBr > HI**LEARNING BY EXAMPLES :****Example:** The electronegativities of F and H are 4.0 and 2.1 respectively. The percentage ionic character in H and F bond is -

- (A) 43 (B) 34 (C) 94 (D) 39

Solution: (A)

$$\% \text{ Ionic character} = 16(4.0 - 2.1) + 3.5(4.0 - 2.1)^2 = 43$$

Example: Which of the following order is correct for acidic property -

- (A) SiH₄ > PH₃ > H₂S (B) (SiH₄ = PH₃ = H₂S)
 (C) PH₃ > SiH₄ > H₂S (D) SiH₄ < PH₃ < H₂S

Solution: (D)

The acidic character of hydrides increase in a period from left to right and also increases when we move down in a group from top to bottom.

Example: Calculate the electronegativity of carbon from the following data:

$$E_{H-H} = 104.2 \text{ kcal mol}^{-1}, \quad E_{C-C} = 83.1 \text{ kcal mol}^{-1}$$

$$E_{C-H} = 98.8 \text{ kcal mol}^{-1}, \quad X_H = 2.1$$

Solution: Let the electronegativity of carbon be X_C , applying Pauling equation,

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

$$X_C - 2.1 = 0.208 [98.8 - (83.1 \times 104.2)^{1/2}]^{1/2}$$

$$\Rightarrow X_C = 2.5$$

Example: Arrange the following in decreasing basic nature

LiOH, NaOH, KOH, CsOH

Solution: The basic nature of hydroxides and oxides of IA elements increases on descending the group with increase in size of cation as $\text{CsOH} > \text{RbOH} > \text{KOH} > \text{NaOH} > \text{LiOH}$.

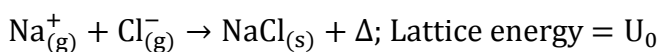
DO YOUR SELF - 6

- Which of the following is affected by stable configuration of an atom :-
 (a) Electronegativity (b) Ionisation potential
 (c) Electron affinity
 Correct answer is :-
 (A) Only electronegativity
 (B) Only ionisation potential
 (C) Electron affinity and ionisation potential
 (D) All of the above
- Which of the following element has the different value of electronegativity :-
 (A) H (B) S (C) C (D) I
- Which is the correct order of electronegativity -
 (A) $\text{Cl} > \text{S} > \text{P} > \text{Si}$ (B) $\text{Si} > \text{Al} > \text{Mg} > \text{Na}$
 (C) $\text{F} > \text{Cl} > \text{Br} > \text{I}$ (D) All
- Which of the following compounds has highest value of bond length -
 (A) CsF (B) CsBr (C) CsI (D) CsCl
- Which of the following will have the most negative electron gain enthalpy and which the least negative ?
 (A) F, Cl (B) Cl, F (C) Cl, S (D) Cl, P
- Which arrangement represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species?
 (A) $\text{S} < \text{O} < \text{Cl} < \text{F}$ (B) $\text{O} < \text{S} < \text{F} < \text{Cl}$
 (C) $\text{Cl} < \text{F} < \text{S} < \text{O}$ (D) $\text{F} < \text{Cl} < \text{O} < \text{S}$

LATTICE ENERGY :

Amount of released energy when one mole of cation and one mole of anion combine to form one mole of ionic compound.

(The anions & cations should be in their gaseous form)



$$\text{LE} = K \frac{q_1 q_2}{r_a + r_c}$$

Factors affecting lattice energy are -

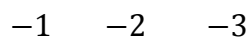
(a) Internuclear distance

(b) Charge on cation

(c) Charge on anion

ORDER OF LATTICE ENERGY:

Based on Internuclear Distance

$$\left\{ \begin{array}{l} \text{LiCl} > \text{NaCl} > \text{KCl} > \text{RbCl} > \text{CsCl} \\ \text{KF} > \text{KCl} > \text{KBr} > \text{KI} \\ \text{NaF} > \text{NaCl} > \text{NaBr} > \text{NaI} \\ \text{Li}_3\text{N} > \text{Na}_3\text{N} > \text{K}_3\text{N} \end{array} \right.$$


Note that in this case, the charge dominates and the size does not affect.

Note: Lattice energy cannot be calculated experimentally, rather using Born Haber cycle.

SOME IMPORTANT DEFINITIONS:

- **Heat of formation** - It is the change in enthalpy associated when the product is formed from its constituent elements present in their standard state.

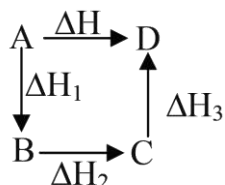
Standard State - It is the physical state of a substance under standard conditions of temperature and pressure.



- **Heat of sublimation** - Energy required to evaporate 1 mole of solid compound to vapour directly.
- **Heat of fusion** - Energy required to convert 1 mole of solid to 1 mole of liquid.
- **Heat of vapourisation** - Energy required to convert 1 mole of liquid to 1 mole of gas.
- **Heat of solvation** - Energy associated with the solvation of 1 mole of an ionic substance in a particular solvent.
- **Bond Dissociation Energy** - Energy required to break the bonds present in 1 mole of substance, provided molecule is covalently bonded.

Hess's Law of constant heat summation

The energy associated with a particular reaction is always constant at a particular temperature and independent of path.

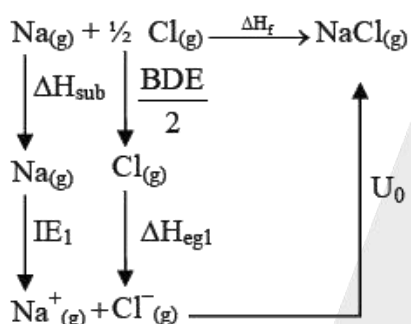


$$\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$$

BORN - HABER CYCLE

This is a cycle which co-relates thermodynamical data with the enthalpy of different physical & chemical routes.

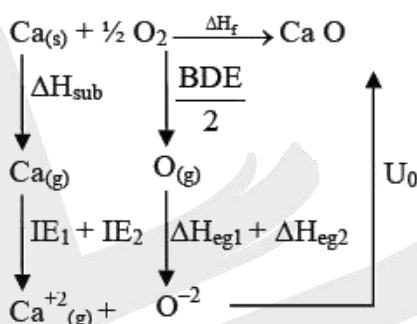
NaCl (For illustration)



Applying Hess's law, we get,

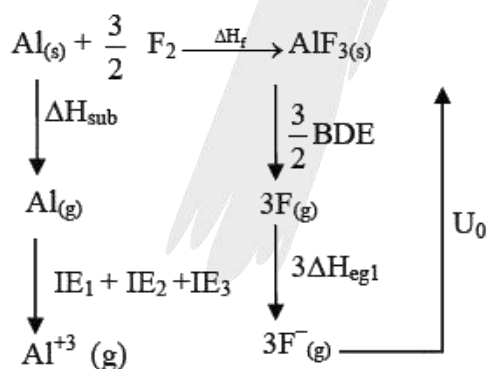
$$\Delta H_f = \Delta H_{\text{sub}} + \text{IE} + \frac{\text{BDE}}{2} + \Delta H_{\text{eg1}} + U_0$$

CaO (For Students)



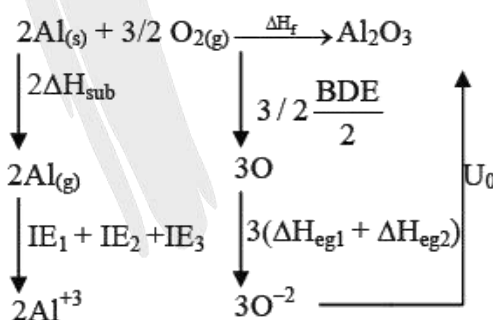
$$\Delta H_f = \Delta H_{\text{sub}} + \Delta \text{IE} + \frac{\text{BDE}}{2} + \Delta H_{\text{eg1}} + \Delta H_{\text{eg2}} + U_0$$

AlF₃ (For Students)



$$\Delta H_f = \Delta H_{\text{sub}} + \Delta \text{IE} + \frac{3}{2} \text{BDE} + 3\Delta H_{\text{eg1}} + U_0$$

Al₂O₃ (For Students)



$$\Delta H_f = \Delta H_{\text{sub}} + (\Delta \text{IE}) + \frac{3}{2} \text{BDE} + 3\Delta H_{\text{eg}} + U_0$$

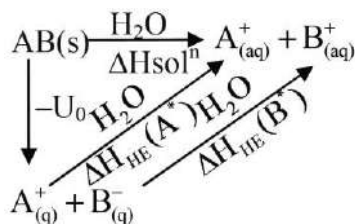
HYDRATION ENERGY

It is the amount of energy released when one mole of gaseous ions undergo hydration, that's special case of solvation. This is the dissolution energy, with the solvent being water.

Hydration energy depends on

(i) Charge (ii) Ionic radius $\text{H. E.} \propto \frac{q_+}{r_+} + \frac{q_-}{r_-}$

(iii) Dielectric constant of solvent $\text{H.E.} \propto \left(1 - \frac{1}{\epsilon_0}\right)$



$$\Delta H_{\text{soln}} = \frac{U_0}{+ve} \frac{\Delta H_{\text{HE}}(\text{A}^+) + \Delta H_{\text{HE}}(\text{B}^-)}{-ve} = \frac{-L.E + H.E -ve}{+ve}$$

If $\Delta H_{\text{soln}} < 0$ when $|\text{HE}| > |\text{LE}|$; Exothermic

If $\Delta H_{\text{soln}} > 0$ when $|\text{HE}| < |\text{LE}|$; Exothermic

If $\Delta H_{\text{soln}} = 0$ when $|\text{HE}| = |\text{LE}|$; Exothermic

Now,

$$F_{\text{air}} = \frac{Kq_1q_2}{r^2} = \frac{1}{4\pi\epsilon_0} \times \frac{q_1q_2}{r^2}$$

$$F_{\text{water}} = \frac{1}{4\pi\epsilon_0 \times 80} \times \frac{q_1q_2}{r^2}$$

$$F_{\text{water}} = \frac{F_{\text{air}}}{80}$$

Hence,

Solubility \Rightarrow $\boxed{\epsilon_0 \text{ value } \uparrow \text{ Solubility } \uparrow, \text{ Force } \downarrow}$

Hydrated Radius:- As the water molecules are attracted by the ions the new radius is called hydrated radius.

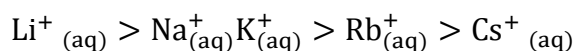
As the charge increases, the no. of water molecules attracted to ion also increases. But as it is just a force, so the hydrated size increases with increase in charge density.

$\boxed{\uparrow \text{ Charge Density } \uparrow \text{ Hydrated size}}$

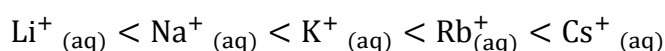
$$\text{Charge Density} = \frac{\text{Charge}}{\text{size}}$$

$\boxed{\uparrow \text{ Hydrated size } \downarrow \text{ Ionic mobility}}$

- Order of Hydrated size:**



- Order of ionic mobility**



We never consider the oxygen anions to exist in water because it is a very strong base and gets converted into OH^- ion.

- Hydration number**

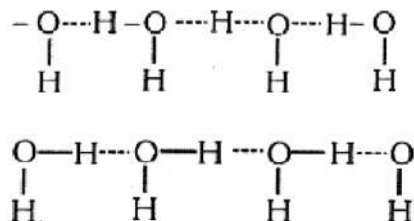
The number of molecules of water with which an ion can combine in an aqueous solution of given concentration.

Ex. $[\text{Na}(\text{H}_2\text{O})_x]^+$. Here, x is the hydration number of sodium.

- Grotthuss conductance**

(This is because of the grotthuss conductance that's why ionic mobility of OH^- is greater than Cl^-)

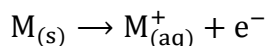
- Grotthuss conductance**



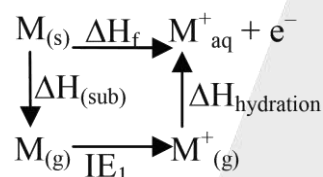
Due to this process, the ionic mobility of OH^- is greater than Cl^- even though the hydrated size is large of OH^- .

REDUCING POWER

It depends on the ease with which the element can remove its electrons and reduce other elements.



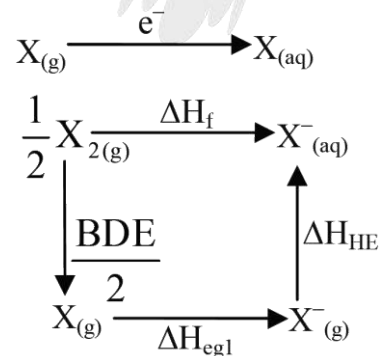
Thermodynamically this chemical would be



$$\Delta H_{\text{f}} = \Delta H_{\text{sub}} + \text{IE}_1 + \Delta H_{\text{hydration}}$$

Li is the best reducing agent because of high enthalpy of hydration, even though its I.E. is higher as compared to Cs.

OXIDISING POWER



$$\Delta H_{\text{f}} = \frac{\text{BDE}}{2} + \Delta H_{\text{eg1}} + \Delta H_{\text{HE}}$$

Here,

$|\Delta H_{\text{HE}}| \uparrow, |\Delta H_{\text{f}}| \uparrow$, Oxidizing power \uparrow

$$\frac{\text{BDE}}{2} \uparrow, |\Delta H_f| \downarrow, \text{Oxidizing power} \downarrow$$

Strongest Oxidising is F_2 even though it has lower E.A. than Cl_2 . This is because of lower B.D.E & high hydration enthalpy.

A

EXERCISE - I

- The element with Atomic number 112 belongs to?
(A) 12th group, 6th period (B) 6th group, 7th period
(C) 12th group, 7th period (D) 6th group, 5th period
- Atomic number of the recently discovered element Nihonium (Nh) is?
(A) 118 (B) 114 (C) 113 (D) 104
- Name of the recently discovered noble gas (atomic number 118) is?
(A) Octauium (Ov) (B) Oganesson (Og)
(C) Moscovium (Me) (D) Tennessine (Ts)
- Unniltrium belongs to which block and group of the periodic table?
(A) d-block, group-10 (B) f-block, group-3
(C) p-block, group-13 (D) d-block, group-11
- In the Lothar Meyer curve, which of the following option is incorrect :
(A) Alkali metals occupied peak position at curve.
(B) Halogens occupied ascending position at curve.
(C) Alkaline earth elements occupy descending position at curve.
(D) Alkali metals are in the lower curve of graph.
- Total number of elements which should be theoretically present in 8th period of the modern long form of periodic table, is
(A) 32 (B) 40 (C) 50 (D) 48
- The period number in the long form of the periodic table is equal to :
(A) magnetic quantum number of any element of the period.
(B) atomic number of any element of the period.
(C) maximum Principal quantum number of any element of the period.
(D) maximum Azimuthal quantum number of any element of the period.

ATOMIC & IONIC RADII

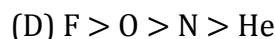
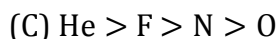
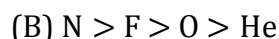
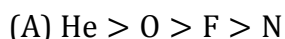
- Which of the following ion has smallest size?
(A) Be^{+2} (B) Ba^{+2} (C) O^{-2} (D) Cl^{-}
- The correct order of atomic radii is :
(A) $\text{F} > \text{Cl} > \text{O} > \text{S}$ (B) $\text{Li} < \text{Na} < \text{K} < \text{Rb}$
(C) $\text{Be} < \text{B} < \text{N} < \text{C}$ (D) Both (A) and (B)

10. When the following four anions are arranged in order of decreasing ionic radius, the correct sequence is:
- (A) Cl^- , F^- , I^- , Br^- (B) Cl^- , Br^- , I^- , F^-
 (C) F^- , I^- , Br^- , Cl^- (D) I^- , Br^- , Cl^- , F^-
11. Choose incorrect option regarding atomic size.
- (A) $\text{Zr} \approx \text{Hf}$ (B) $\text{Fe} \approx \text{Co} \approx \text{Ni}$
 (C) $\text{Y} \approx \text{La}$ (D) All of these
12. Choose the incorrect option
- (A) $\text{Na}^+ > \text{Li}^+ > \text{Mg}^{+2} > \text{Be}^{+2}$ (ionic radius)
 (B) $\text{F}^- < \text{Cl}^- < \text{Br}^- < \text{I}^-$ (ionic radius)
 (C) $\text{Sn}^{4+} > \text{In}^+ > \text{Sn} > \text{In}$ (Z_{eff})
 (D) $\text{Na}^+ > \text{F}^- > \text{O}^{2-}$ (atomic radius)
13. If the ionic radii of K^+ and F^- are nearly the same (i.e., 1.34\AA), then the atomic radii of K and F respectively are:
- (A) 1.34\AA , 1.34\AA (B) 0.72\AA , 1.96\AA
 (C) 1.96\AA , 0.72\AA (D) 1.96\AA , 1.34\AA

IONISATION ENERGY

14. The third ionization energy is least for
- (A) Nitrogen (B) Phosphorus (C) Aluminum (D) Boron
15. Ionisation energy is highest for
- (A) F^- (B) Cl^- (C) Be^- (D) S^-
16. From the given information of element M
- $\text{M} \rightarrow \text{M}^{+3} + 3\text{e}^- \quad H = x\text{kJ/mole}$
 $\text{IE}_1(\text{M}) = y\text{kJ/mole}$
 $\text{IE}_1(\text{M}^{+2}) = z\text{kJ/mole}$
 The value of $\text{IE}_2(\text{M})$ will be
- (A) $(x + y + z)\text{kJ/mole}$ (B) $(x - y - z)\text{kJ/mole}$
 (C) $(x - y + z)\text{kJ/mole}$ (D) $z - (x + y)\text{kJ/mole}$
17. Which of the following species has lowest value of Ionization energy
- (A) F (B) Cl (C) F^- (D) Cl^-
18. Which of the following relation is correct with respect to first (I) and second (II) ionization enthalpies of potassium and calcium ?
- (A) $I_{\text{Ca}} > II_{\text{K}}$ (B) $I_{\text{K}} > I_{\text{Ca}}$ (C) $II_{\text{Ca}} > II_{\text{K}}$ (D) $\Pi_{\text{K}} > \Pi_{\text{Ca}}$

19. The correct order of first IE is-



20. Which of the following statement is incorrect?

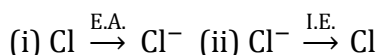
(A) The second ionization energy of sulphur is greater than that of chlorine

(B) The second ionization energy of silicon is greater than that of aluminium

(C) The first ionization energy of aluminium is lower than gallium

(D) The second ionization energy of boron is greater than that of carbon

21. Which of the following statement is correct regarding following process?



(A) | I.E. of process (ii) | = | E.A. of process (i) |

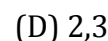
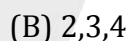
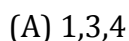
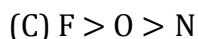
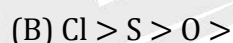
(B) | I.E. of process (iii) | = | I.E. of process (ii) |

(C) | I.E. of process (iv) | = | E.A. of process (i) |

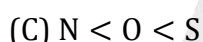
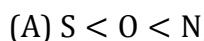
(D) | I.E. of process (iv) | = | I.E. of process (iii) |

ELECTRON AFFINITY

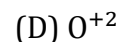
22. Which of the following represent the correct order of electron affinity:



23. Arrange N, O and S in order of increasing electron affinity



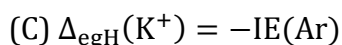
24. Electron addition would be easier in -



25. Which of the following statement is correct ?

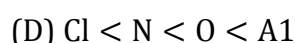
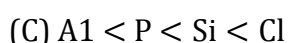
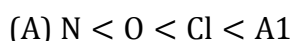
(A) Electron gain enthalpy must be positive for all elements.

(B) Second electron gain enthalpy always remains negative for all the elements.



(D) None of these

26. The increasing order of electron affinity is:



27. The order of electron gain enthalpy (magnitude) of O, S, Se and Te is :
 (A) $O > S > Se > Te$ (B) $S > Se > Te > O$
 (C) $Se > S > Te > O$ (D) $S > O > Se > Te$
28. The element which have highest magnitude of I.E. and electron gain enthalpy both amongst the following (Consider only one electronic change)
 (A) Br (B) O (C) F (D) I
29. Which of following is most unstable species
 (A) Li^- (B) N^- (C) C^- (D) O^-
30. Addition of an electron will be endothermic process for :
 (A) Sr (B) Be (C) N (D) All of these

ELECTRONEGATIVITY

31. Which of the following is most electronegative element among the following?
 (A) H (B) Hg (C) Cr (D) Li
32. According to theory of electronegativity, most suitable name for OF_2 molecule will be
 (A) Oxygen fluoride (B) Difluorine oxide
 (C) Oxygen difluoride (D) Both (B) and (C)
33. Which of the following is correct order of electronegativity:
 (A) $Cs > Rb > Na$ (B) $Li < B < Be$
 (C) $C < N < O$ (D) $Cl > F > Br$
34. Correct electronegativity value of Silicon using Allred-Rochow's scale is
 [where radius of silicon is 1.175\AA]
 (A) 1.82 (B) 2.20 (C) 1.50 (D) 2.00

NATURE OF OXIDES

35. Which of the following represent the correct order of basic strength?
 (A) $Li_2O < Na_2O < K_2O < Rb_2O$ (B) $BeO < MgO < CaO < SrO$
 (C) $K_2O > CaO > Ga_2O_3$ (D) All are correct
36. Which of the following oxide will give the amphoteric
 (A) N_2O (B) CO (C) As_2O_3 (D) BaO
37. Choose the correct order for acidic strength:
 (A) $Na_2O < MgO$ (B) $NO < N_2O$ (C) $SiO_2 > P_2O_5$ (D) $HClO_2 > HClO_3$
38. The weakest base among NaOH, $Ca(OH)_2$, KOH and $Be(OH)_2$ is-
 (A) NaOH (B) $Ca(OH)_2$ (C) KOH (D) $Be(OH)_2$

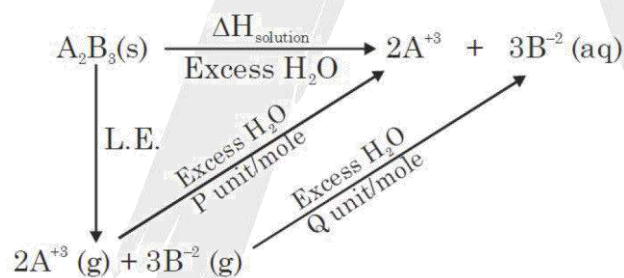
39. Identify the correct order of acidic strength of CO_2 , CuO , CaO
- (A) $\text{CaO} < \text{CuO} < \text{CO}_2$ (B) $\text{CuO} < \text{CaO} < \text{CO}_2$
 (C) $\text{CaO} > \text{CuO} > \text{CO}_2$ (D) $\text{CO}_2 < \text{CaO} < \text{CuO}$

HYDRATION/LATICE ENERGY

40. Order of energy released during hydration of K^+ , Rb^+ and Cs^+ is ?
- (A) $\text{K}^+ < \text{Rb}^+ < \text{Cs}^+$ (B) $\text{Rb}^+ < \text{K}^+ < \text{Cs}^+$
 (C) $\text{Cs}^+ > \text{Rb}^+ > \text{K}^+$ (D) $\text{K}^+ > \text{Rb}^+ > \text{Cs}^+$
41. Which of the following has highest lattice energy?
- (A) LiCl (B) LiF (C) BeO (D) Be_3N_2
42. Hydration energy will be maximum for which of the following ion?
- (A) Be^{+2} (B) Li^+ (C) I^- (D) Mg^{+2}
43. Which of the following compound has highest lattice energy?
- (A) LiF (B) A/N (C) Na_2O_2 (D) Al_2O_3
44. Maximum hydrated radius from the ions given below will be ?
- (A) Mg^{+2} (B) Al^{+3} (C) S^{-2} (D) F^-

BORN HABER'S CYCLE

45. In the given cycle, if P and Q are the magnitude of hydration energies (H.E.) per mole for A^{+3} (g) and B^{-2} (g) ions respectively than the value of $\left(P + \frac{3}{2}Q\right)$ will be
- (Consider only magnitude of H.E.) = (L.E. = Lattice energy of Ionic compound)



- (A) $\frac{\Delta H_{\text{sol}} - \text{L.E.}}{2}$ (B) $\frac{3}{2}(\Delta H_{\text{sol}} - \text{L.E.})$
 (C) $\Delta H_{\text{sol}} - \text{L.E.}$ (D) $\Delta H_{\text{sol}} - \text{L.E.}$

EXERCISE - II

- Which of the following element(s) is/are representative element?
(A) Kr (B) Cs (C) Zn (D) Te
- Which of the following pair of elements shows diagonal relationship:
(A) Li, Mg (B) B, Al (C) B, Si (D) Be, Al
- Which of the following is/are bridging elements
(A) P (B) Ar (C) Al (D) Cl
- Consider the following representation based on long form of periodic table.

		P	
S	X	Q	
	R		

Here P, Q, R and S are up, right, down and left elements with respect to the central element 'X' respectively. According to above representation the correct match is/are:

- | | | | | | |
|-----|----|----|----|----|----|
| | X | P | Q | R | S |
| (A) | Ge | Si | As | Sn | Ga |
| (B) | Te | Se | I | Po | Sb |
| (C) | Sb | As | Te | Bi | Sn |
| (D) | In | Ga | Sn | Tl | Cd |
- In which of the following pairs, the first atom is larger than the second?
(A) Br, Cl (B) Na, Mg (C) Sr, Ca (D) N, P
 - Find the correct order of 2nd ionisation energy?
(A) Al > Mg (B) Cr > Cr⁺ (C) P > S (D) S > Cl
 - The sum of IE₁ and IE₂, IE₃ and IE₄ for element P and Q are given below:

	IE ₁ + IE ₂	IE ₃ + IE ₄
(P)	2.45	8.82
(Q)	2.85	6.11

Then according to the given information the correct statement(s) is/are:

(A) p ²⁺ is more stable than Q ²⁺	(B) p ²⁺ is less stable than Q ²⁺
(C) p ⁴⁺ is more stable than Q ⁴⁺	(D) p ⁴⁺ is less stable than Q ⁴⁺
 - Electron affinity of the elements or ions shown correctly?
(A) S > O⁻ (B) P > N⁻ (C) O⁻ > S⁻ (D) N⁻ > P

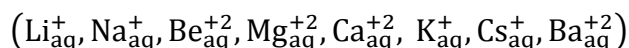
9. Select the endothermic step(s):
 (A) $S^{-}(g) + e^{-} \rightarrow S^{2-}(g)$ (B) $Ne(g) + e^{-} \rightarrow Ne^{-}(g)$
 (C) $N(g) + e^{-} \rightarrow N^{-}(g)$ (D) $Al^{2+}(g) \rightarrow Al^{3+}(g) + e^{-}$
10. Choose the incorrect statement option from given below :
 (A) $Li_{(aq)}^{+} > Na_{(aq)}^{+}$ (Hydrated size) (B) $Li_{(aq)}^{+} < Na_{(aq)}^{+}$ (Ionic mobility)
 (C) $Be^{+2} < Al^{+3}$ (Ionic potential) (D) $KF > KC1$ (lattice energy)
11. Consider the order $O^{2-} < F^{+} < Na^{+} < Mg^{+2}$. Then correct statement(s) is/are :
 (A) Increasing order of Z_{eff} (B) Increasing order of size
 (C) Increasing order of I.E. (D) Increasing order of E.A.
12. Select the correct order of periodic properties of species
 (A) $Fe^{2+} < Fe^{3+}$: Ionic radii (B) $N < O$: Second ionisation energy
 (C) $Cu < Zn$: atomic radius (D) $In < Tl$: First ionisation energy
13. Which of the following process is endothermic ?
 (A) $N_{(g)}^{-} \rightarrow N_{(g)}$ (B) $Mg_{(g)}^{+2} \rightarrow Mg_{(g)}^{+}$
 (C) $P_{(g)}^{-} \rightarrow P_{(g)}$ (D) $Cl_{(g)} \rightarrow Cl_{(g)}^{2-}$
14. Which of the following order is CORRECT for the property indicated against it
 (A) $O < S < Cl < Br$ (Atomic size) (B) $S < N < O < F$ (Electronegativity)
 (C) $Pb < Sn < Si < C(I_1)$ (D) $O < S < Cl(EA_1)$
15. In which of the following statements is the correct order with respect to the given property?
 (A) $Fe^{3+}(aq) > Fe^{2+}(aq)$: Ionic mobility order.
 (B) $Br^{-}(aq) < Cl^{-}(aq) < F^{-}(aq)$: Hydrated radius order.
 (C) $SbH_3 > AsH_3 > PH_3$: $-H$ bond polarity order.
 (D) $Br > C > H > B$: Electronegativity order.
16. If electronegativity of P, Q, R, S elements are respectively 0.8, 1.2, 1.0 and 2.5 then choose the hydroxides which are basic in nature
 (A) POH (B) QOH (C) ROH (D) SOH
17. Which of the following oxides is/ are amphoteric?
 (A) Na_2O (B) CaO (C) Al_2O_3 (D) SnO
18. Which of the following order of acidic strength is /are true ?
 (A) $SO_2 < SO_3$ (B) $P_4O_6 < P_4O_{10}$ (C) $HOCl < HOI$ (D) $Na_2O < MgO$

EXERCISE - III

Integer Answer Type (0 to 9):

- The element with the lowest atomic number that has a ground state electronic configuration of $(n-1)d^5ns^2$ is located in period.
- Period number of S | c = \underline{x}
Modern periodic table group number of Tl = y (according to 1 to 18 convention)
Find the value of $y - x$
- Calculate the covalent radius of 'F' if
 $r_H = 37\text{pm}$, electronegativity of H & F are 2.1 & 4.0 respectively.
[Given : $d_{H-F} = 83.9\text{pm}$]
[Write your answer as sum of digits till you get the single digit answer]
- Among the following total no. of species which have lower ionisation energy than carbon is (Li, Be, B, N, F, Ne, $C^{-}C^{+2}$)
- Element 'X' has maximum electron affinity in periodic table then find the number of elements which have low first ionisation energy as compared to element 'X' ?
F, Na, S, Si, P, Ar
- How many pairs are, in which first species has lower ionisation energy than second species:
(i) N and O (ii) Br and K (iii) Be and B (iv) I and I^{-}
(v) Li and Li^{+} (vi) O and S (vii) Ba and Sr
- According to Haney-Smith formula, if E.N. difference between A & B is 2.1, then A-B molecule is expected to have $x\%$ ionic character, find $\frac{x}{10}$ value.
- How many of the following elements are more electronegative than Boron.
H, Li, Be, C, N, O, F
- $A_{(g)}^{+} \rightarrow A_{(g)} \Delta H = -2\text{eV}$
 $2 A_{(g)} \rightarrow A_{(g)}^{+} \Delta H = x$
Find the value of x .
- How many of the following energy terms are involved in the transformation of
 $Na_{(s)} \rightarrow Na_{(aq)}^{+}$
 $IE, \Delta H_{\text{sub}}, \Delta H_{\text{diss}}, \Delta H_{\text{HE}}, \Delta H_{\text{LE}}$
- How many of the following reactions proceed with the absorption of energy
 $Cl(g) \rightarrow Cl^{-}(g)$ $O^{-2}(g) \rightarrow O^{-}(g)$ $N(g) \rightarrow N^{-}(g)$ $Mg(g) \rightarrow Mg^{+2}(g)$
 $Cl^{-}(g) \rightarrow Cl^{-2}(g)$ $Be(g) \rightarrow Be^{-}(g)$
- In the given molecules total number of oxide & hydroxide which is amphoteric in nature is -
 $Pb(OH)_2, V_2O_5, As_2O_3, N_2O, Al_2O_3, Cr_2O_3, Be(OH)_2, ZnO, Si(OH)_4, P_4O_{10}, Cl_2O_7, BiOH, CO$

13. Total no. of cations among the following which have higher electrical conductance (in aqueous solution) than Al^{+3} is-



14. Determine the lattice energy released due to formation of MgS lattice in kcal/mol in nearest possible integers from the following data.

$$\Delta H_f(\text{MgS}) = -82.2 \text{ kcal/mol.}$$

$$\Delta_{\text{sub}}(\text{Mg}) = 36.5 \text{ kcal/mol.}$$

$$\text{for Mg, } IE_I + IE_{II} = 520.6 \text{ kcal/mol}$$

$$\text{for S, } \Delta_{\text{atom}} H = 1065.6 \text{ kcal/mol.}$$

$$\text{For S, } \Delta_{\text{eg}} H_1 + \Delta_{\text{eg}} H = -72.5 \text{ kcal/mol.}$$

15. Calculate the lattice enthalpy of magnesium bromide from the, data shown in the Table.

Processes	$\Delta H(\text{kJ/mol})$
Sublimation of Mg(s)	+148
Ionization of mg(g) to $\text{Mg}^{+2}(\text{g})$	+2187
Vaporization of $\text{Br}_2(\ell)$	+31
Dissociation of $\text{Br}_2(\text{g})$	+193
Electron gain by Br(g)	-331
Formation of $\text{MgBr}_2(\text{g})$	-524

[Write your answer as sum of digits till you get the single digit answer]

EXERCISE - IV

Ionisation energies of unknown elements are given below :

Element	IE (in kcal/mol)		
	I	II	III
M	209	548	919
N	100	735	1101
O	119	1092	1653
P	1500	2017	2320

- Which pair of elements represent same group?
(A) M, N (B) N, O (C) O, P (D) M, P
- Which amongst them is expected to be a noble gas?
(A) M (B) N (C) O (D) P

Paragraph for question nos. 3 to 4

In 1931, Pauling defined electronegativity of an atom as the tendency of the atom to attract shared electron (bond pair) towards itself when combined in compound. It is a unit less & dimensionsless quantity. If two atoms have similar electronegativities, the bond between them is largely covalent. But when difference between the electronegativity is large, the bond has high degree of polar character. Pauling's approach for obtaining values of electronegativity is explained as Pauling scale and given as:

$$X_A - X_B = 0.208\sqrt{\Delta_{A-B}} \text{ (when energies are expressed in kcal mol}^{-1}\text{)}$$

Where, $X_A = E.N. \text{ of } A$

$X_B = E.N. \text{ of } B$

$\Delta_{AB} = \text{Resonance energy of bond}$

- Select the incorrect statement?
(A) P is more likely to form covalent bond with H.
(B) According to Pauling, oxygen is the second most electronegative element in periodic table
(C) Electronegativity is independent to the electronic configuration
(D) All the statements given above are incorrect
- Bond formed between given elements will be least covalent in:
(A) Li & Cl (B) Cs & Cl (C) Be & I (D) F & I

Paragraph for question no. 5 to 7

The electron affinity is an inherent property of the atom and it depends upon several factors.

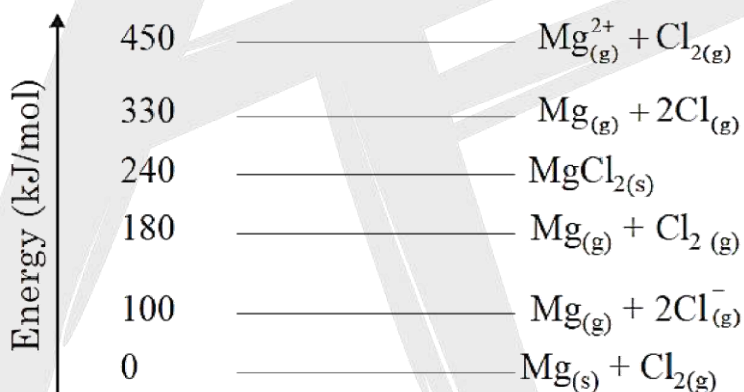
- The correct electron affinity order is
(A) $F > Cl$ (B) $Cl > F$ (C) $S < P$ (D) $N > O$

6. Which of the following is an incorrect statement ?
- (A) The first ionization energy of A^- is equal to the electron affinity of A.
- (B) The second electron affinity is always greater than the first electron affinity for an element.
- (C) The process $O \rightarrow O^{2-}$ is endothermic
- (D) The process $Li \rightarrow Li^+$ is endothermic
7. Which of the following reactions would not proceed in the forward direction spontaneously?
- (A) $Xe + He^+ \rightarrow Xe^+ + He$
- (B) $Si + Cl^+ \rightarrow Si^+ + Cl$
- (C) $F^- + I \rightarrow F + I^-$
- (D) $Be^- + B \rightarrow B^- + Be$

Paragraph for Question 8 to 9

Born Haber cycle is used to calculate enthalpy of formation of ionic compounds. This is based on Hess's law which states that enthalpy of reaction does not depend upon the path followed by the reaction.

Energy of some systems are given below (hypothetical data):



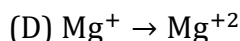
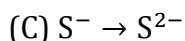
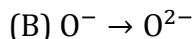
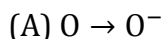
8. The enthalpy of formation (in kJ/mole) of $MgCl_{2(s)}$ is:
- (A) 510 (B) 320 (C) 180 (D) 240
9. Electron affinity (in kJ/ mole) of chlorine is:
- (A) 170 (B) 115 (C) 230 (D) 105

Paragraph for Question 10 to 11

Electron affinity and electronegativity are very closely related, yet they are very different.

10. Identify the only INCORRECT statement
- (A) Be has almost zero electron affinity
- (B) Between F and Cl, formation of Cl^- is more exothermic than formation of F^-
- (C) IE_1 of Cl^- = electron affinity of Cl
- (D) Electronegativity depends only on electron affinity

11. Which of the following processes is exothermic?



12. **Column-I**

(A) 52

(B) 30

(C) 57

(D) 91

Column-II

(P) p-block

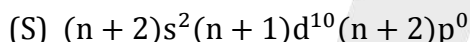
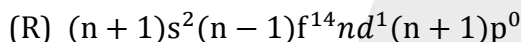
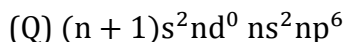
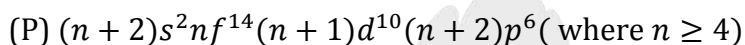
(Q) Transition elements

(R) d-block element

(S) Representative element

(T) Inner transition element

13. **List - I**



List - II

(A) s-block

(B) p-block

(C) d-block

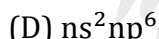
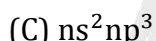
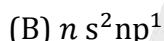
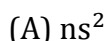
(D) f-block

Codes :

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

14. **Column-I**

(Configuration)



Column-II

(Ionisation energy in kJ/mol)

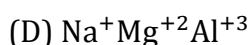
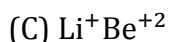
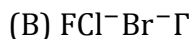
(P) 2100

(Q) 1400

(R) 800

(S) 900

15. **Column-I**



Column-II

(P) Have decreasing order of released hydration energy

(Q) Have decreasing order of hydrated radius

(R) Have decreasing order of electrical conductivity in aqueous Solution

(S) Have decreasing order of released lattice energy while forming lattice with common anion

(T) Have decreasing order of their gaseous ionic radius

EXERCISE - V (JEE MAIN)

- According to the Periodic law of elements, the variation in properties of elements is related to their :- [AIEEE-2003]
 - Nuclear masses
 - Atomic numbers
 - Nuclear neutron-proton number ratio
 - Atomic masses
- The atomic numbers of vanadium, (V), chromium (Cr), manganese (Mn) and iron (Fe) are respectively 23,24,25 and 26 . Which one of these may be expected to have the highest second ionization enthalpy? [AIEEE-2003]
 - Cr
 - Mn
 - Fe
 - V
- Among Al_2O_3 , SiO_2 , P_2O_5 and SO_3 , the correct order of acid strength is [AIEEE-2004]
 - $\text{Al}_2\text{O}_3 < \text{SiO}_2 < \text{SO}_3 < \text{P}_2\text{O}_5$
 - $\text{SiO}_2 < \text{SO}_3 < \text{Al}_2\text{O}_3 < \text{P}_2\text{O}_5$
 - $\text{SO}_3 < \text{P}_2\text{O}_5 < \text{SiO}_2 < \text{Al}_2\text{O}_3$
 - $\text{Al}_2\text{O}_3 < \text{SiO}_2 < \text{P}_2\text{O}_5 < \text{SO}_3$
- The formation of the oxide ion $\text{O}^{2-} - \text{g}$) requires first an exothermic and then an endothermic step as shown below [AIEEE-2004]

$$\text{O}_{(\text{g})} + \text{e}^- = \text{O}_{(\text{g})}^-, \Delta H^\circ = -142 \text{ kJ mol}^{-1}$$

$$\text{O}_{(\text{g})}^- + \text{e}^- = \text{O}_{(\text{g})}^{2-}, \Delta H^\circ = 844 \text{ kJ mol}^{-1}$$

This is because

 - O^- ion will tend to resist the addition of another electron
 - Oxygen has high electron affinity
 - Oxygen is more electronegative
 - O^- ion has comparatively larger size than oxygen atom
- In which of the following arrangements, the order is NOT according to the property indicated below? [AIEEE-2005]
 - $\text{Al}^{3+} < \text{Mg}^{2+} < \text{Na}^+ < \text{F}^-$ - increasing ionic size
 - $\text{B} < \text{C} < \text{N} < \text{O}$ - increasing first ionization enthalpy
 - $\text{I} < \text{Br} < \text{F} < \text{Cl}$ - increasing electron gain enthalpy (with negative sign)
 - $\text{Li} < \text{Na} < \text{K} < \text{Rb}$ - increasing metallic radius
- Which of the following oxides is amphoteric in character? [AIEEE-2005]
 - SnO_2
 - SiO_2
 - CO_2
 - CaO

7. Which of the following factors may be regarded as the main cause of lanthanide contraction? [AIEEE-2005]
- (A) poor shielding of one of 4 electron by another in the subshell
 (B) effective shielding of one of 4f electrons by another in the subshell
 (C) poorer shielding of 5 d electrons by 4f electrons
 (D) greater shielding of 5 d electrons by 4f electrons
8. The increasing order of the first ionization enthalpies of the elements B, P, S and F (lowest first) is:- [AIEEE-2006]
- (A) $F < S < P < B$ (B) $P < S < B < F$
 (C) $B < P < S < F$ (D) $B < S < P < F$
9. The set representing the correct order of ionic radius is [AIEEE-2009]
- (A) $Li^+ > Na^+ > Mg^{2+} > Be^{2+}$ (B) $Mg^{2+} > Be^{2+} > Li^+ > Na^+$
 (C) $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$ (D) $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$
10. The correct order of electron gain enthalpy with negative sign of F, Cl, Br and I, having atomic number 9, 17, 35 and 53 respectively, is [AIEEE-2011]
- (A) $I > Br > Cl > F$ (B) $F > Cl > Br > I$
 (C) $Cl > F > Br > I$ (D) $Br > Cl > I > F$
11. The increasing order of the ionic radii of the given isoelectronic species is [AIEEE-2012]
- (A) $K^+, S^{2-}, Ca^{2+}, Cl^-$ (B) $Cl^-, Ca^{2+}, K^+, S^{2-}$
 (C) $S^{2-}, Cl^-, Ca^{2+}, K^+$ (D) $Ca^{2+}, K^+, Cl^-, S^{2-}$
12. In which of the following arrangements, the sequence is not strictly according to the property written against it? [JEE-MAIN 2012]
- (A) $CO_2 < SiO_2 < SnO_2 < PbO_2$: increasing oxidising power
 (On-Line)
 (B) $B < C < O < N$: increasing first ionisation enthalpy
 (C) $NH_3 < PH_3 < AsH_3 < SbH_3$: increasing basic strength
 (D) $HF < HCl < HBr < HI$: increasing acid strength
13. Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar? [JEE-MAIN-2013]
- (A) $Ca < S < Ba < Se < Ar$ (B) $S < Se < Ca < Ba < Ar$
 (C) $Ba < Ca < Se < S < Ar$ (D) $Ca < Ba < S < Se < Ar$

14. The first ionisation potential of Na is 5.1eV. The value of electron gain enthalpy of Na^+ will be

[JEE-M AIN-2013]

- (A) -2.55eV (B) -5.1eV (C) -10.2eV (D) $+2.55\text{eV}$

15. Electron gain enthalpy with negative sign of fluorine is less than that of chlorine due to :

[JEE-MAIN 2013 (On-Line)]

- (A) Smaller size of chlorine atom
(B) Bigger size of $2p$ orbital of fluorine
(C) High ionization enthalpy of fluorine
(D) Smaller size of fluorine atom

16. Given

[JEE-M AIN-2013 (On-Line)]

Reaction	Energy Change (in kJ/mol)
$\text{Li(s)} \rightarrow \text{Li(g)}$	161
$\text{Li(g)} \rightarrow \text{Li}^+(\text{g})$	520
$\frac{1}{2} \text{F}_2(\text{g}) \rightarrow \text{F(g)}$	77
$\text{F(g)} + \text{e}^- \rightarrow \text{F}^-(\text{g})$	(Electron gain enthalpy)
$\text{Li(g)} + \text{F}^-(\text{g}) \rightarrow \text{LiF(s)}$	-1047
$\text{Li(s)} + \frac{1}{2} \text{F}_2(\text{g}) \rightarrow \text{LiF(s)}$	-617

Based on data provided, the value of electron gain enthalpy of fluorine would be :

- (A) -300 kJ mol^{-1} (B) -328 kJ mol^{-1}
(C) -350 kJ mol^{-1} (D) -228 kJ mol^{-1}

17. The order of increasing sizes of atomic radii among the elements O, S, Se and As is :

[JEE-MAIN 2013 (On-Line)]

- (A) $\text{As} < \text{S} < \text{O} < \text{Se}$ (B) $\text{O} < \text{S} < \text{As} < \text{Se}$
(C) $\text{Se} < \text{S} < \text{As} < \text{O}$ (D) $\text{O} < \text{S} < \text{Se} < \text{As}$

18. Which of the following is the correct order of second ionization potential of C, N, O and F :

[JEE-MAIN 2013 (On-Line)]

- (A) $\text{O} > \text{F} > \text{N} > \text{C}$ (B) $\text{O} > \text{N} > \text{F} > \text{C}$
(C) $\text{C} > \text{N} > \text{O} > \text{F}$ (D) $\text{F} > \text{O} > \text{N} > \text{C}$

19. Which of the following series correctly represents relations between the elements from X to Y?

[JEE-MAIN 2014 (On-Line)]

$\text{X} \rightarrow \text{Y}$

- (A) $_{18}\text{Ar} \rightarrow _{54}\text{Xe}$ Noble character increases
(B) $_3\text{Li} \rightarrow _{19}\text{K}$ Ionization increases
(C) $_6\text{C} \rightarrow _{32}\text{Ge}$ Atomic radii increases
(D) $_{18}\text{F} \rightarrow _{35}\text{Br}$ Electron gain enthalpy with negative sign increases

20. The ionic radii (in Å) of N^{3-} , O^{2-} and F^- are respectively
[JEE-MAIN 2015 (Off-Line)]
- (A) 1.71, 1.40 and 1.36 (B) 1.71, 1.36 and 1.40
(C) 1.36, 1.40 and 1.71 (D) 1.36, 1.71 and 1.40
21. In the long form of the periodic table, the valence shell electronic configuration of $5s^2 5p^4$ corresponds to the element present in:
[JEE-MAIN 2015 (On-Line)]
- (A) Group 16 and period 5 (B) Group 17 and period 6
(C) Group 17 and period 5 (D) Group 16 and period 6
22. Which of the following atoms has the highest first ionization energy?
[JEE-MAIN 2016 (Off-Line)]
- (A) Sc (B) Rb (C) Na (D) K
23. The non-metal that does not exhibit positive oxidation state is : [JEE-MAIN 2016 (On-Line)]
- (A) Oxygen (B) Fluorine (C) Iodine (D) Chlorine
24. The group having isoelectronic species is: [JEE MAIN 2017-Offline]
- (A) O^{2-} , F^- , Na^+ , Mg^{2+} (B) O^- , F^- , Na^+ , Mg^{2+}
(C) O^{2-} , F^- , Na^+ , Mg^{2+} (D) O^- , F^- , Na^+ , Mg^{+2}
25. The electronic configuration with the highest ionization enthalpy is:
[JEE MAIN 2017 - Online]
- (A) $[\text{Ne}]3s^2 3p^1$ (B) $[\text{Ne}]3s^2 3p^2$
(C) $[\text{Ne}]3s^2 3p^3$ (D) $[\text{Ar}]3d^{10} 4s^2 4p^3$
26. Which of the following is an oxide? [JEE MAIN 2017 - Online]
- (A) KO_2 (B) BaO_2 (C) SiO_2 (D) CsO_2
27. For Na^+ , Mg^{2+} , F^- & O^{2-} , the correct order of increasing ionic radii is :
[JEE MAIN 2018-Online]
- (A) $\text{Na}^+ < \text{Mg}^{2+} < \text{F}^- < \text{O}^{2-}$ (B) $\text{Mg}^{2+} < \text{Na}^+ < \text{F}^- < \text{O}^{2-}$
(C) $\text{Mg}^{+2} < \text{O}^{2-} < \text{Na}^+ < \text{F}^-$ (D) $\text{O}^{2-} < \text{F}^- < \text{Na}^+ < \text{Mg}^{2+}$
28. In general, the properties that decrease and increase down a group in the periodic table, respectively, are :
[JEE MAIN 2019-Online]
- (A) electronegativity and electron gain enthalpy.
(B) electronegativity and atomic radius.
(C) atomic radius and electronegativity.
(D) electron gain enthalpy and electronegativity.

29. When the first electron gain enthalpy ($\Delta_{\text{eg}}H$) of oxygen is -141 kJ/mol , its second electron gain enthalpy is : **[JEE MAIN 2019-Online]**
 (A) almost the same as that of the first
 (B) a positive value
 (C) a more negative value than the first
 (D) negative but less negative than the first
30. The electronegativity of aluminium is similar to : **[JEE MAIN 2019-Online]**
 (A) Carbon (B) Beryllium (C) Boron (D) Lithium
31. The effect of lanthanoid contraction in the lanthanoid series of elements by and large means : **[JEE MAIN 2019-Online]**
 (A) decrease in atomic radii and increase in ionic radii
 (B) increase in both atomic and ionic radii
 (C) increase in atomic radii and decrease in ionic radii
 (D) decrease in both atomic and ionic radii
32. The correct order of the atomic radii of C, Cs, Al and S is : **[JEE MAIN 2019-Online]**
 (A) $S < C < Cs < Al$ (B) $C < S < Al < Cs$
 (C) $C < S < Cs < Al$ (D) $S < C < Al < Cs$
33. The correct option with respect to the Pauling electronegativity values of the elements is :- **[JEE MAIN 2019-Online]**
 (A) $Te > Se$ (B) $P > S$ (C) $Si < Al$ (D) $Ga < Ge$
34. The element with $Z = 120$ (not yet discovered) will be an/a : **[JEE MAIN 2019-Online]**
 (A) transition metal (B) inner-transition metal
 (C) alkaline earth metal (D) alkali metal
35. Correct order of electron gain enthalpy (kJ/mole) of F, Cl, Br, I **[JEE MAIN 2020-Online]**
 (A) $-296, -325, -333, -349$ (B) $-333, -349, -325, -296$
 (C) $-349, -333, -325, -296$ (D) $-325, -333, -349, -296$
36. Correct order of first ionization energy of the following metals Na, Mg, Al, Si in KJmol^{-1} respectively are: **[JEE MAIN 2020-Online]**
 (A) 497, 737, 577, 786 (B) 497, 577, 737, 786
 (C) 786, 739, 577, 497 (D) 739, 577, 786, 487
37. Which one of the following amongs each pair will release maximum energy on gaining one electron **[JEE MAIN 2020-Online]**
 (A = F, Cl), (B = S, Se), (C = Li, Na)
 (A) (A) = Cl, (B) = S, (C) = Li (B) (A) = F, (B) = Se, (C) = Li
 (C) (A) = F, (B) = Se, (C) = Na (D) (A) = Cl, (B) = S, (C) = Na

38. In general the property (magnitudes only) that show an opposite trend in comparison to other properties across a period is [JEE MAIN 2020-Online]

- (A) Electron gain enthalpy (B) Electronegativity
(C) Ionization enthalpy (D) Atomic radius

39. Among the statements (I – IV), the correct ones are [JEE MAIN 2020-Online]

- (I) Be has smaller atomic radius compared to Mg.
(II) Be has higher ionization enthalpy than Al.
(III) Charge/radius ratio of Be is greater than that of Al.
(IV) Both Be and Al form mainly covalent compounds.

40. The atomic number of Unnilunium is [JEE MAIN 2020-Online]

41. B has a smaller first ionization enthalpy than Be. Consider the following statements :

[JEE MAIN 2020-Online]

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

The correct statement are :

- (A) (i), (ii), (iii), (iv) (B) (i), (iii), (iv)
(C) (ii), (iii), (iv) (D) (i), (ii), (iii)

42. Identify the elements X and Y using the ionisation energy values given below: [JEE MAIN 2021]

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

- (A) X = Na; Y = Mg (B) X = Mg; Y = F
(C) X = Mg; Y = Na (D) X = F; Y = Mg

43. The characteristics of elements X, Y and Z with atomic numbers, respectively, 33, 53 and 83 are: [JEE MAIN 2021]

- (A) X and Y are metalloids and Z is a metal.
(B) X is a metalloid, Y is a non-metal and Z is a metal.
(C) X, Y and Z are metals.
(D) X and Z are non-metals and Y is a metalloid

44. The first ionization energy of magnesium is smaller as compared to that of elements X and Y , but higher than that of Z . the elements X , Y and Z , respectively, are : [JEE MAIN 2021]
 (A) chlorine, lithium and sodium (B) argon, lithium and sodium
 (C) argon, chlorine and sodium (D) neon, sodium and chlorine
45. Given below are two statements : one is labelled as [JEE MAIN 2021]
 Assertion (A) and the other is labelled as Reason (R).
Assertion (A) : Metallic character decreases and non-metallic character increases on moving from left to right in a period.
Reason (R) : It is due to increase in ionisation enthalpy and decrease in electron gain enthalpy, when one moves from left to right in a period.
 In the light of the above statements, choose the most appropriate answer from the options given below:
 (A) (A) is false but (R) is true.
 (B) (A) is true but (R) is false.
 (C) Both (A) and (R) are correct and (R) is the correct explanation of (A).
 (D) Both (A) and (R) are correct but (R) is not the correct explanation of (A).
46. The IUPAC nomenclature of an element with electronic configuration $[Rn]5f^{14}6d^17s^2$ is : [JEE MAIN 2022]
 (A) Unnilbium (B) Unnilunium
 (C) Unnilquadium (D) Unniltrium
47. In which of the following pairs, electron gain enthalpies of constituent elements are nearly the same or identical ? [JEE MAIN 2022]
 (A) Rb and CS (B) Na and K
 (C) Ar and Kr (D) I and At
 Choose the correct answer from the options given below:
 (A) (A) and (B) only (B) (B) and (C) only
 (C) (A) and (C) only (D) (C) and (D) only
48. Element "E" belongs to the period 4 and group 16 of the periodic table. The valence shell electron configuration of the element, which is just above 'E' in the group is [JEE MAIN 2022]
 (A) $3s^2 \cdot 3p^4$ (B) $3d^{10} \cdot 4s^2, 4p^4$
 (C) $4d^{10} \cdot 5s^2, 5p^4$ (D) $2s^2, 2p^4$

49. Match List-I with List-II.

[JEE MAIN 2022]

List-I (Oxide)		List-II (Nature)	
(A)	Cl_2O_7	(I)	Amphoteric
(B)	Na_2O	(II)	Basic
(C)	Al_2O_3	(III)	Neutral
(D)	N_2O	(IV)	Acidic

Choose the correct answer from the options given below:

(A) (A) – (IV), (B) - (III), (C) - (I), (D) - (II)

(B) (A) – (IV), (B) - (II), (C) - (I), (D) - (III)

(C) (A) – (II), (B) - (IV), (C) - (III), (D) - (I)

(D) (A) – (I), (B) - (II), (C) - (III), (D) - (IV)

50. Inert gases have positive electron gain enthalpy. Its correct order is [JEE MAIN 2023]

(A) $\text{Xe} < \text{Kr} < \text{Ne} < \text{He}$ (B) $\text{He} < \text{Ne} < \text{Kr} < \text{Xe}$ (C) $\text{He} < \text{Xe} < \text{Kr} < \text{Ne}$ (D) $\text{He} < \text{Kr} < \text{Xe} < \text{Ne}$

51. Which of the following represents the correct order of metallic character of the given elements? [JEE MAIN 2023]

(A) $\text{Si} < \text{Be} < \text{Mg} < \text{K}$ (B) $\text{Be} < \text{Si} < \text{Mg} < \text{K}$ (C) $\text{K} < \text{Mg} < \text{Be} < \text{Si}$ (D) $\text{Be} < \text{Si} < \text{K} < \text{Mg}$

52. Match List-I with List-II

[JEE MAIN 2023]

LIST-I (Atomic number)		LIST-II (Block of periodic table)	
(A)	37	I.	p-block
(B)	78	II.	d-block
(C)	52	III.	f-block
(D)	65	IV.	s-block

Choose the correct answer from the options given below:

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

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

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

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

53. Which of the following elements have half-filled f-orbitals in their ground state?

(Given : atomic number Sm = 62; Eu = 63; Tb = 65; Gd = 64, Pm = 61)

[JEE MAIN 2023]

A. Sm

B. Eu

C. Tb

D. Gd

E. Pm Choose the correct answer from the options given below:

(A) B and D only

(B) A and E only

(C) A and B only

(D) C and D only

EXERCISE - VI (JEE ADVANCE)

1. **Statement-1:** F atom has a less negative electron gain enthalpy than Cl atom. [JEE 2000]
Statement-2: Additional electron is repelled more efficiently by 3p electron in Cl atom than by 2p electron in F atom.
(A) Statement-1 is true, statement-2 is true and statement-2 is correct explanation for statement-1.
(B) Statement-1 is true, statement-2 is true and statement-2 is NOT the correct explanation for statement-1.
(C) Statement-1 is true, statement-2 is false.
(D) Statement-1 is false, statement-2 is true.
2. The correct order of radii is: [JEE 2000]
(A) $N < Be < B$ **(B)** $F^- < O^{2-} < N^{3-}$
(C) $Na < Li < K$ **(D)** $Fe^{3+} < Fe^{2+} < Fe^{4+}$
3. The IE_1 of Be is greater than that of B. [T/F] [JEE 2001]
4. The set representing correct order of IP_1 is [JEE 2001]
(A) $K > Na > Li$ **(B)** $Be > Mg > Ca$
(C) $B > C > N$ **(D)** $Fe > Si > C$
5. Identify the least stable ion amongst the following: [JEE 2002]
(A) Li^- **(B)** Be^-
(C) B^- **(D)** C^-
6. The increasing order of atomic radii of the following group 13 elements is [JEE 2016]
(A) $Al < Ga < In < Tl$ **(B)** $Ga < Al < In < Tl$
(C) $Al < In < Ga < Tl$ **(D)** $Al < Ga < Tl < In$
7. The option(s) with only amphoteric oxides is (are): [JEE Advance - 2017]
(A) Cr_2O_3, CrO, SnO, PbO **(B)** NO, B_2O_3, PbO, SnO_2
(C) Cr_2O_3, BeO, SnO, SnO_2 **(D)** ZnO, Al_2O_3, PbO, PbO_2

ANSWER KEY

Do yourself-1

1. (B) 2. (A)

Do yourself-2

1. (D) 2. (C) 3. (A) 4. (B) 5. (C) 6. (C)

Do yourself-3

1. (B) 2. (D) 3. (C) 4. (D) 5. (C) 6. (A) 7. (D)

Do yourself-4

1. (B) 2. (B) 3. (D) 4. (A) 5. (B) 6. (A)
7. ((a) s, (b) r, (c) q, (d) p)

Do yourself-5

1. (D) 2. (D) 3. (B) 4. (A) 5. (C) 6. (A, C)

Do yourself-6

1. (C) 2. (A) 3. (D) 4. (C) 5. (D) 6. (B)

EXERCISE-1

1. (C) 2. (C) 3. (B) 4. (B) 5. (D) 6. (C) 7. (C)
8. (A) 9. (B) 10. (D) 11. (C) 12. (D) 13. (C) 14. (C)
15. (B) 16. (B) 17. (C) 18. (D) 19. (C) 20. (B) 21. (A)
22. (D) 23. (C) 24. (D) 25. (D) 26. (C) 27. (B) 28. (C)
29. (B) 30. (D) 31. (A) 32. (C) 33. (C) 34. (A) 35. (D)
36. (C) 37. (A) 38. (D) 39. (A) 40. (D) 41. (D) 42. (A)
43. (B) 44. (B) 45. (A)

EXERCISE-2

1. (B,D) 2. (A,C,D) 3. (A,C,D) 4. (ABCD) 5. (ABC) 6. (AD) 7. (AD)
8. (AB) 9. (ABCD) 10. (C) 11. (ACD) 12. (BCD) 13. (CD)
14. (BD)
15. (BCD) 16. (ABC) 17. (CD) 18. (ABD)

EXERCISE-3

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

EXERCISE-4

1. (B) 2. (D) 3. (D) 4. (B) 5. (B) 6. (B) 7. (C)
8. (D) 9. (B) 10. (D) 11. (A)
12. (A) \rightarrow P, S; (B) \rightarrow R; (C) \rightarrow Q, R; (D) \rightarrow T 13. C
14. (A) \rightarrow S; (B) \rightarrow R; (C) \rightarrow Q; (D) \rightarrow P
15. (A) \rightarrow P, Q, S; (B) \rightarrow P, Q, S; (C) \rightarrow R, T; (D) \rightarrow R, T

EXRCISE-V (JEE MAIN)

- | | | | | | | |
|---------|---------|---------|---------|-----------|---------|---------|
| 1. (B) | 2. (A) | 3. (D) | 4. (A) | 5. (B) | 6. (A) | 7. (C) |
| 8. (D) | 9. (D) | 10. (C) | 11. (D) | 12. (C) | 13. (C) | 14. (B) |
| 15. (D) | 16. (B) | 17. (D) | 18. (A) | 19. (C) | 20. (A) | 21. (A) |
| 22. (A) | 23. (D) | 24. (C) | 25. (C) | 26. (C) | 27. (B) | 28. (B) |
| 29. (B) | 30. (B) | 31. (D) | 32. (B) | 33. (D) | 34. (C) | 35. (B) |
| 36. (A) | 37. (A) | 38. (D) | 39. (B) | 40. (101) | 41. (D) | 42. (A) |
| 43. (B) | 44. (C) | 45. (B) | 46. (D) | 47. (C) | 48. (A) | 49. (B) |
| 50. (C) | 51. (A) | 52. (D) | 53. (A) | | | |

EXRCISE-V (JEE ADVANCE)

- | | | | | | |
|---------|--------|-----------|--------|--------|--------|
| 1. (C) | 2. (B) | 3. (TRUE) | 4. (B) | 5. (B) | 6. (B) |
| 7. (CD) | | | | | |