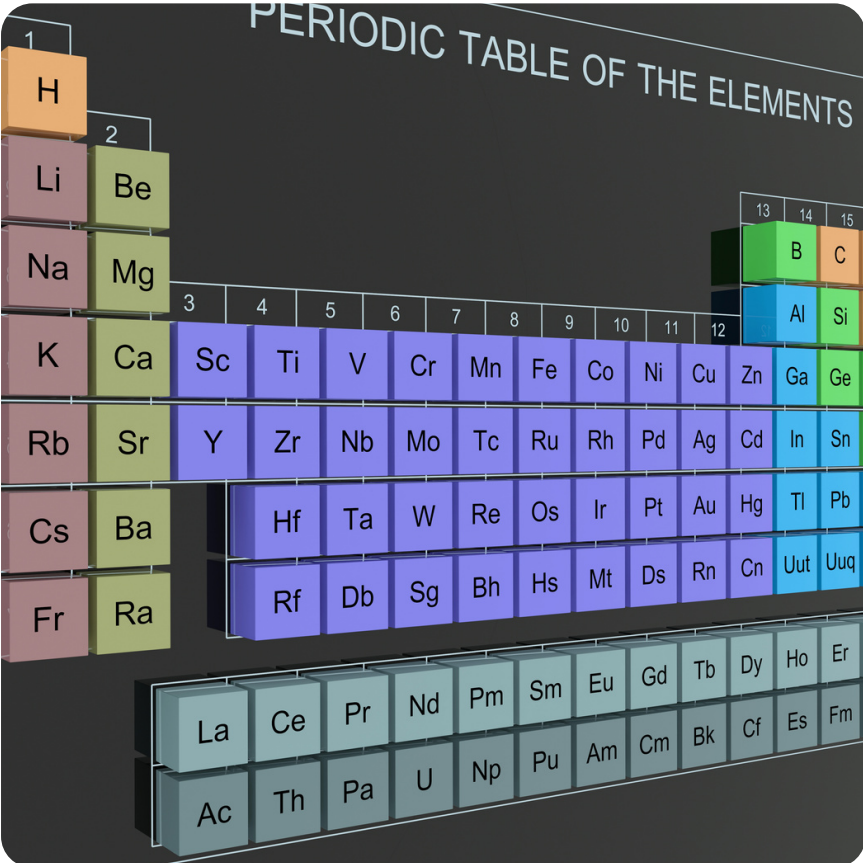


INORGANIC CHEMISTRY

ENTHUSIAST | LEADER | ACHIEVER

PERIODIC TABLE OF THE ELEMENTS



1	2											13	14	15
H												B	C	
Li	Be													
Na	Mg											Al	Si	
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	
Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	
Fr	Ra		Rf	Db	Sg	Bh	Hs	Mt	Ds	Rn	Cn	Uut	Uuq	
		La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	
		Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	

STUDY MATERIAL

Periodic Table

ENGLISH MEDIUM

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Extended or Long Form of the Periodic Table

Extended or Long Form of the Periodic Table

s-Block Elements

IA

Group

Period

1

1

1.0079

H

Hydrogen

Metals

Non metals

Metalloids

0

(18)

2

He

4.0026

Helium

2

Li

6.940

Lithium

3

Na

22.990

Sodium

4

K

39.098

Potassium

5

Rb

85.468

Rubidium

6

Cs

132.91

Cesium

7

Fr

223

Francium

4

Be

9.0122

Beryllium

5

Mg

24.305

Magnesium

6

Ca

40.078

Calcium

7

Sr

87.62

Strontium

8

Ba

137.33

Barium

9

Ra

226

Radium

3

Al

26.982

Aluminium

4

Si

28.086

Silicon

5

P

30.974

Phosphorus

6

S

32.066

Sulphur

7

Cl

35.453

Chlorine

8

Ar

39.948

Argon

11

Na

22.990

Sodium

12

Mg

24.305

Magnesium

13

Al

26.982

Aluminium

14

Si

28.086

Silicon

15

P

30.974

Phosphorus

16

S

32.066

Sulphur

17

Cl

35.453

Chlorine

18

Ar

39.948

Argon

19

K

39.098

Potassium

20

Ca

40.078

Calcium

21

Sc

44.956

Scandium

22

Ti

47.867

Titanium

23

V

50.941

Vanadium

24

Cr

51.996

Chromium

25

Mn

54.938

Manganese

26

Fe

55.847

Iron

27

Co

58.933

Cobalt

28

Ni

58.693

Nickel

29

Cu

63.546

Copper

30

Zn

65.39

Zinc

31

Ga

62.723

Gallium

32

Ge

72.61

Germanium

33

As

74.922

Arsenic

34

Se

78.96

Selenium

35

Br

79.904

Bromine

36

Kr

83.80

Krypton

37

Rb

85.468

Rubidium

38

Sr

87.62

Strontium

39

Y

88.906

Yttrium

40

Zr

91.224

Zirconium

41

Nb

92.906

Niobium

42

Mo

95.94

Molybdenum

43

Tc

98

Technetium

44

Ru

101.07

Ruthenium

45

Rh

102.91

Rhodium

46

Pd

106.42

Palladium

47

Ag

107.87

Silver

48

Cd

112.41

Cadmium

49

In

114.82

Indium

50

Sn

118.71

Tin

51

Sb

121.76

Antimony

52

Te

127.60

Tellurium

53

I

126.90

Iodine

54

Xe

131.29

Xenon

55

Cs

132.91

Cesium

56

Ba

137.33

Barium

57

La*

138.91

Lanthanum

58

Ce

140.12

Cerium

59

Pr

140.91

Praseodymium

60

Nd

144.24

Neodymium

61

Pm

145

Promethium

62

Sm

150.36

Samarium

63

Eu

151.96

Europium

64

Gd

157.25

Gadolinium

65

Tb

158.93

Terbium

66

Dy

162.50

Dysprosium

67

Ho

164.93

Holmium

68

Er

167.26

Erbium

69

Tm

168.93

Thulium

70

Yb

173.07

Ytterbium

71

Lu

174.97

Lutetium

72

Hf

178.49

Hafnium

73

Ta

180.95

Tantalum

74

W

183.84

Tungsten

75

Re

186.21

Rhenium

76

Os

190.23

Osmium

77

Ir

192.22

Iridium

78

Pt

195.08

Platinum

79

Au

196.97

Gold

80

Hg

200.59

Mercury

81

Tl

204.38

Thallium

82

Pb

207.2

Lead

83

Bi

208.98

Bismuth

84

Po

210

Polonium

85

At

210

Astatine

86

Rn

222

Radon

87

Fr

223

Francium

88

Ra

226

Radium

89

Ac**

227

Actinium

90

Th

232.04

Thorium

91

Pa

231.04

Protactinium

92

U

238.03

Uranium

93

Np

237

Neptunium

94

Pu

244

Plutonium

95

Am

243

Americium

96

Cm

247

Curium

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

f-Block Elements

*Lanthanide Series

**Actinide Series

IUPAC designations of groups of elements are given in brackets

PERIODIC TABLE

1.0 INTRODUCTION :

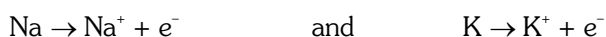
The arrangement of all the known elements according to their properties in such a way that the elements with similar properties are grouped together in a tabular form is called periodic table.

DEVELOPMENT OF PERIODIC TABLE

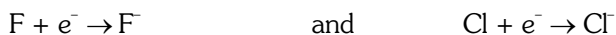
(A) LAVOISIER CLASSIFICATION :

- (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) **Drawback or Limitation :**

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

(B) PROUT'S HYPOTHESIS :

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,....

Drawback or Limitation :

- (i) Every element can not 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)

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

(D) NEWLAND OCTAVE RULE [1865]

- (i) He arranged the elements in the increasing order of their atomic mass and observed that the properties of every 8th element was similar to the 1st element. (like in the case of musical vowels notation)
- (ii) At that time inert gases were not known.

Sa	Re	Ga	Ma	Pa	Dha	Ni	Sa
1	2	3	4	5	6	7	8

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

- (iii) The properties of Li are similar to 8th element i.e. Na and Be are similar to Mg and so on.

Drawback or Limitation :

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

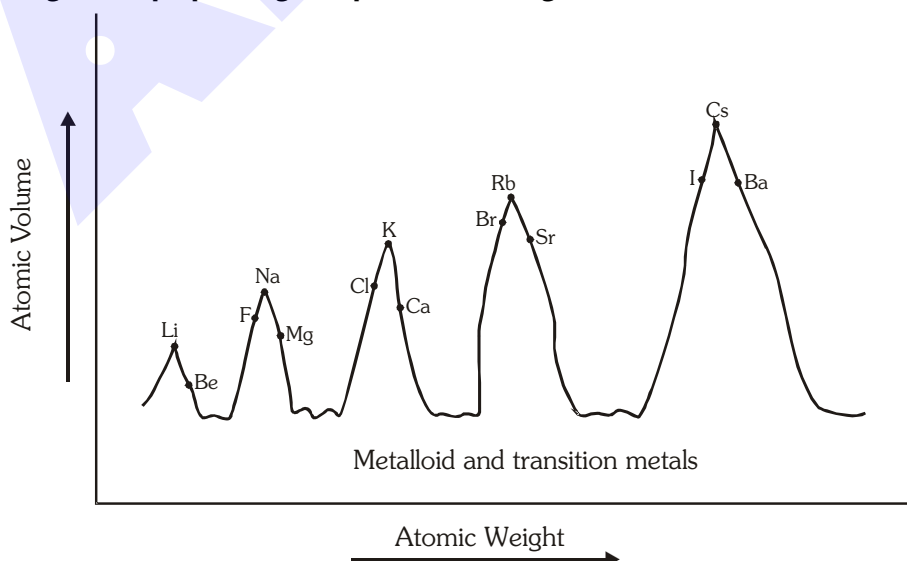
(E) LOTHAR MEYER'S CURVE [1869] :

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

Note : Elements having similar properties occupy similar position 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 has become 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.



(F) MENDELEEV'S PERIODIC TABLE [1869] :

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

(ii) **Characteristics of Mendeleev's periodic table :**

- It is based on atomic weight
- 63 elements were known, noble gases were not discovered.
- He was the first scientist to classify the elements in a systematic manner i.e. in horizontal rows and in vertical columns.
- Horizontal rows are called periods and there were 7 periods in Mendeleev's Periodic table.
- Vertical columns are called groups and there were 8 group in Mendeleev's Periodic table.
- Each group upto VII is divided into A & B subgroups. 'A' sub group element are called normal or representative elements and 'B' sub group elements are called transition elements.
- The VIII group consisted of 9 elements in three rows (Transitional metals group).
- The elements belonging to same group exhibit similar properties.

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

- 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 .
- 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's periodic table were.

Eka Aluminium	– Gallium(Ga)	Eka Boron	– Scandium (Sc)
Eka Silicon	– Germanium (Ge)	Eka Manganese	– Technetium (Tc)

(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 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.

Corrections were done in atomic weight of elements are – **U, Be, In, Au, Pt.**

(iv) **Demerits of Mendeleev's periodic table :**

- 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.
- Position of isotopes** : As atomic wt. of isotopes differs, they should have placed in different position in Mendeleev's periodic table. But there were no such places for isotopes in Mendeleev's table.

- (c) **Anomalous pairs of elements :** There were some pair of elements 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, but having higher atomic weights.

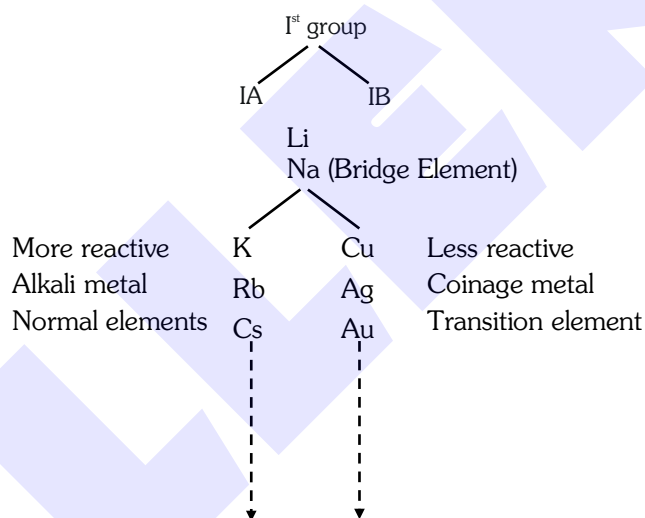
$\begin{pmatrix} \text{Ar} & \text{K} \\ 39.9 & 39.1 \end{pmatrix}$	$\begin{pmatrix} \text{Te} & \text{I} \\ 127.5 & 127 \end{pmatrix}$
$\begin{pmatrix} \text{Co} & \text{Ni} \\ 58.9 & 58.6 \end{pmatrix}$	$\begin{pmatrix} \text{Th} & \text{Pa} \\ 232 & 231 \end{pmatrix}$

- (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
VIII	IB

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



Cu, Ag and Au placed in Ist 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 & chemical properties repeated in a group.

BEGINNER'S BOX-1

- Mendeleev's periodic law is based on -
 (1) Atomic number (2) Atomic weight (3) Number of neutrons (4) None of the above
- The first attempt to classify elements systematically was made by -
 (1) Mendeleev (2) Newland (3) Lothar Meyer (4) Dobereiner
- 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
 (1) 88.5 (2) 93.0 (3) 171 (4) 85.5

Pre-Medical

4. Which of the following is not a Dobereiner triad
 (1) Li, Na, K (2) Mg, Ca, Sr (3) Cl, Br, I (4) S, Se, Te
5. The law of triads is applicable to
 (1) C, N, O (2) H, O, N (3) Na, K, Rb (4) Cl, Br, I
6. The law of triads is not applicable on
 (1) Cl, Br, I (2) Na, K, Rb (3) S, Se, Te (4) Ca, Sr, Ba
7. Which of the following set of elements obeys Newland's octave rule –
 (1) Na, K, Rb (2) F, Cl, Br (3) Be, Mg, Ca (4) B, Al, Ga
8. For which of the pair Newland octave rule is not applicable –
 (1) Li, Na (2) C, Si (3) Mg, Ca (4) Cl, Br
9. Which of the following element was present in Mendeleev's periodic table?
 (1) Sc (2) Tc (3) Ge (4) None of these
10. Is Fe, Co, Ni are dobereiner triad ?

1.1 MODERN PERIODIC TABLE (MODIFIED MENDELEEV PERIODIC TABLE) :

- (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.

LONG FORM / PRESENT FORM OF MODERN PERIODIC TABLE :

(It is also called as 'Bohr, Bury, Rang & 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.

- (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	VII	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	n	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	32	${}_{87}\text{Fr} - {}_{118}\text{Uuo}$	Complete

CONCLUSION

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

DESCRIPTION OF GROUPS :

1st/IA/Alkali metals $\text{H} = 1s^1$ $\text{Li} = 1s^2, 2s^1$ $\text{Na} = 1s^2, 2s^2 2p^6, 3s^1$ $\text{K} = 1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^1$ General electronic configuration = ns^1 Number of valence shell $e^- = 1$	2nd/IIA/Alkaline earth metals $\text{Be} = 1s^2, 2s^2$ $\text{Mg} = 1s^2, 2s^2 2p^6, 3s^2$ $\text{Ca} = 1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^2$ General electronic configuration = ns^2 (n = Number of shell) Number of valence shell $e^- = 2$
13th/IIIA/Boron Family $\text{B} = 1s^2, 2s^2 2p^1$ $\text{Al} = 1s^2, 2s^2 2p^6, 3s^2 3p^1$ $\text{Ga} = 1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^1$ General electronic configuration = $ns^2 np^1$ Number of valence shell $e^- = 3$	14th/IVA/Carbon Family $\text{C} = 1s^2, 2s^2 2p^2$ $\text{Si} = 1s^2, 2s^2 2p^6, 3s^2 3p^2$ $\text{Ge} = 1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^2$ General electronic configuration = $ns^2 np^2$ Number of valence $e^- = 4$
15th/VA/Nitrogen family/Pnicogen (Used in fertilizer as urea) $\text{N} = 1s^2, 2s^2 2p^3$ $\text{P} = 1s^2, 2s^2 2p^6, 3s^2 3p^3$ $\text{As} = 1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^3$ General electronic configuration = $ns^2 np^3$ Number of valence shell $e^- = 5$	16th/VIA/Oxygen family/Chalcogen (Ore forming) $\text{O} = 1s^2, 2s^2 2p^4$ $\text{S} = 1s^2, 2s^2 2p^6, 3s^2 3p^4$ $\text{Se} = 1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^4$ General electronic configuration = $ns^2 np^4$ Number of valence shell $e^- = 6$

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 $e^- = 7$	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 $e^- = 8$
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(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
111	Unununium	Uuu	Rontgenium	Rg
112	Ununbium	Uub	Copernicium	Cn
113	Ununtrium	Uut	Nihonium	Nh
114	Ununquadium	Uuq	Flerovium	Fl
115	Ununpentium	Uup	Moscovium	Mc
116	Ununhexium	Uuh	Livermorium	Lv
117	Ununseptium	Uus	Tennessine	Ts
118	Ununoctium	Uuo	Oganesson	Og

(viii) Identification of group, period and block :
(A) When atomic number is given :

Step I : $71 \geq Z \geq 58 \Rightarrow$ Lanthanoids (6th Period)
 $103 \geq Z \geq 90 \Rightarrow$ Actinoids (7th Period) } f-block

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.

(B) When electronic configuration is given

Period number (n) = number of outermost shell/Highest shell number.

Block identification :

- If np electron present then p - block ($ns^2 np^{1-6}$)
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

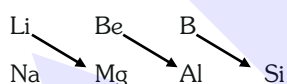
Bohr's Classification

Inert gases	Normal or Representative element	Transition element	Inner transition element
outermost shell complete	outermost shell incomplete	n & n-1 shells incomplete either in atomic or ionic form	n, (n-1), (n-2) shells incomplete
6 element	s & p block element except inert gas upto 118 (Atomic no.)	all d block element except = IIB (Zn, Cd, Hg & Uub) 36 element	f-block elements 28 elements

SOME IMPORTANT POINTS :

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

$$(\text{Ionic potential } (\phi) = \frac{\text{Charge on cation}}{\text{Radius of cation}})$$

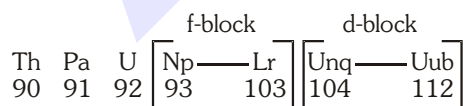


- (b) 3rd period elements (Na, Mg, Al, Si, P, S, Cl) except inert gases are called **typical elements** because they represent the properties of other element of their respective group.

(c) TRANSURANIC ELEMENTS :

Elements having atomic number more than 92 are known as transuranic element.

All transuranic elements are radioactive & artificial.



First man made element is Tc

First man made lanthanoid is Pm

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

- (d) The group containing most electro positive elements – GROUP IA.
(e) The group containing most electro negative elements – GROUP VIIA

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- (f) The group containing maximum number of gaseous elements—GROUP ZERO(18th)
- (g) The group in which elements have generally ZERO valency – GROUP ZERO(18th)
- (h) **In the periodic table**
- | | |
|---|-------------------------------------|
| Number of Gaseous elements | – 11 (H, N, O, F, Cl + Noble gases) |
| Number of Liquid elements | – 6 (Cs, Fr, Ga, Hg, Br, Uub) |
| Number of Liquid elements at room temp. | – 2 |
- Bromine is the only non-metal which exists in liquid form.
- | | |
|--------------------------|---------------------------------------|
| Number of Solid elements | – 95 (if discovered elements are 112) |
|--------------------------|---------------------------------------|
- (i) 0/18 group have all the elements in gaseous form.
- (j) 2nd period contains maximum number of gaseous elements. They are 4 **(N, O, F, Ne)**
- (k) IIIB/3rd group is called longest group having 32 elements including 14 Lanthanides and 14 Actinides
- | | |
|-------------------------|--|
| Sc | |
| Y | |
| La.....Lanthanides (14) | |
| Ac.....Actinides (14) | |

BEGINNER'S BOX-2

- Which of the following set of atomic numbers represents representative element
 (1) 5, 13, 30, 53 (2) 11, 33, 58, 84 (3) 5, 17, 31, 64 (4) 9, 31, 53, 83
- Which of the following electronic configuration does not belongs to same block as others :-
 (1) [Xe] 4f¹⁴ 5d¹⁰ 6s² (2) [Kr] 4d¹⁰ 5s² (3) [Kr] 5s² (4) [Ar] 3d⁶ 4s²
- The electronic configuration of an element is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$. What is the atomic number of next element of the same group which is recently discovered :-
 (1) 20 (2) 119 (3) 111 (4) None
- Select the correct matching of given atomic number with their IUPAC symbol and IUPAC official name :-
 (1) 104 – Unq – Rutherfordium
 (2) 110 – Une – Darmstadtium
 (3) 107 – Uno – Bohrium
 (4) 102 – Unt – Nobelium
- Which of the following statement is wrong :-
 (1) Total no. of liquid elements in the periodic table.....Six
 (2) First metal element in the periodic table is....Li
 (3) All type of elements are present in 6th period
 (4) Iodine is a gaseous element.
- An element which is recently discovered is placed in 7th period and 10th group. IUPAC name of the element will be :-
 (1) Unnilseptium (2) Ununnilium (3) Ununbium (4) None
- Neptunium and Plutonium are found in -
 (1) Pitch blende (2) Malachite (3) Alum (4) None
- According to Soviet scientists name of the element which has atomic number 104-
 (1) Rutherfordium (2) Kurchatovium (3) Unnilquadium (4) Not exist

9. Accordiy to IUPAC nomenclature name of Lawrencium is -
 (1) Unnilbium (2) Unniltrium (3) Unnilquadium (4) Unnilocatium
10. Most of the man-made radioactive element's are present in -
 (1) 6th period (2) 7th period (3) 4th period (4) None

1.2 PERIODICITY

(A) In a period, the ultimate orbit remain same, but the number of e^- gradually increases.

In a group, the number of e^- in the ultimate orbit remains same, but the values of n increases.

(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 e^- 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 e^- due to inner shell e^- is called screening effect or shielding effect.(i.e. total repulsive force is called shielding effect.)
- (iii) Due to screening effect valence shell e^- experiences less force of attraction exerted by nucleus. i.e. total attraction force experieced by valence electrons represented by a number is Z_{eff} .
- (iv) There is a reduction in nuclear charge due to screening effect. Reduced nuclear charge is called effective nuclear charge.
- (v) If nuclear charge = Z , effective nuclear charge = Z_{eff} , σ (Sigma)= Screening constant or shielding constant. So, $Z_{\text{eff}} = (Z - \sigma)$

● Slater's rule to know screening constant (σ)

- (a) For single electron species $\sigma = 0$
- (b) Screening effect (S.E.) for two e^- species 0.30

Ex. In He ($1s^2$)

Screening effect of one $1s e^-$. where $\sigma = 0.30$

$$\therefore Z_{\text{eff}} = Z - \sigma = 2 - 0.30 = 1.7$$

- (c) Screening effect of each ns and np (Outermost orbit) electrons is 0.35
- (d) Screening effect of each $(n - 1)$ penultimate orbit s, p, d electrons is 0.85
- (e) Screening effect of each $(n - 2)$ and below all the e^- present in s, p, d, f is 1.0

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 shielding effect has the order as $s > p > d > f$ (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} for different ions of an element
	(ii) Z_{eff} for isoelectronic species.

(i) Z_{eff} for different ions of an element

Ex. $N^+ > N > N^- = Z_{\text{eff}}$

(ii) Z_{eff} of isoelectronic species

Ex. $H^- < Li^+ < Be^{+2} < B^{+3}$ ($2e^-$ species)

$N^{3-} < O^{2-} < F^- < Na^+ < Mg^{+2}$ ($10e^-$ species)

1.3 ATOMIC RADIUS

The average distance of valence shell e^- from nucleus is called atomic radius. It is very difficult to measure the atomic radius because –

- The isolation of single atom is very difficult.
 - There is no well defined boundary for the atom. (The probability of finding the e^- 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 combine.

Based on the chemical bonds, atomic radius is divided into four categories –

(A) Covalent radius (B) Ionic radius (C) Metallic radius (D) van der Waals radius

(A) Covalent Radius

(SBCR –Single Bonded Covalent Radius)

- Covalent bonds are formed by overlapping of atomic orbitals.
- Internuclear distance is minimum in this case.
- Covalent radius is the half of the internuclear distance between two singly bonded homo atoms.

Ex. If internuclear distance of $A-A(A_2)$ molecule is (d_{A-A}) and covalent radius is r_A then

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

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

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

(B) Ionic Radius
(i) Cationic Radius

- When an neutral atom loses e^- it converts into cation (+ve charged ion)
- Cationic radius is always smaller than atomic radius **because** after loosing e^- number of e^- reduces, but number of protons remains same, due to this Z_{eff} increases, hence electrons are pulled towards nucleus and atomic radius decreases, moreover after loosing 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 $\propto \frac{1}{\text{Magnitude of the charge or } Z_{\text{eff}}}$

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 e^- it converts into anion [Negative charge ion]

(b) Anionic radius is always greater than atomic radius **because** in an anion e^- are more than protons and inter electronic repulsion increases, which also increases screening effect. So effective nuclear charge reduces, so distance between e^- and nucleus increases and size of anion also increases.

Ex. Flourine ($Z=9$)

	F	F ⁻
Proton	9	9
e^-	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 size of F⁻ > F

(c) Size of isoelectronic species :

- Those species having same number of e^- 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
$\frac{Z}{e}$	$\frac{19}{18}$	$\frac{20}{18}$	$\frac{16}{18}$	$\frac{17}{18}$

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

└──────────┘
└──────────┘
18 e⁻
10 e⁻

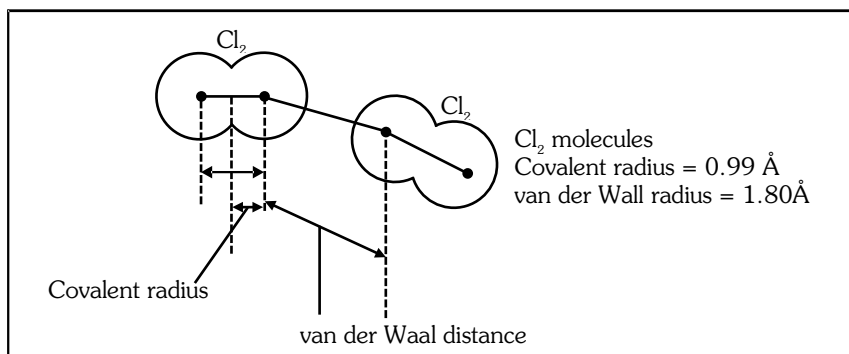
(C) Metallic/Crystal Radius

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

(D) Vander Waals Radius

- (a) Those atoms (like noble gases) which are not bonded with each other, experiences a weak attractive force to come nearer.
- (b) Half of the distance between the nuclei of adjacently placed atoms in solid state of a noble gas is Vander Waals radius.
- (c) Inert gas have only Vander Waal radius.
- (d) In molecules of non metals solid both covalent and Vander Waals radius exists.



$$\text{Vander Waals radius} \cong 2 \times \text{covalent radius}$$

$$\text{Vander Waals radius} > \text{Metallic radius} > \text{Covalent radius}$$

● **Factors affecting atomic size are :**

(a) **In a period**

$$\text{Atomic radius} \propto \frac{1}{Z_{\text{eff}}} \propto \frac{\text{negative charge}}{\text{positive charge}}$$

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

(b) **In a group**

$$\text{Atomic radius} \propto \text{number of shells}$$

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

● **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

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

(ii) **In a group :** It increases from top to bottom in a group as number of shell increases

$$\text{Ex. } \text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$$

Exceptions : Transition elements

$$\begin{array}{ccccc} \text{Sc} & \text{Ti} & \text{V} & \text{Cr} & \text{Mn} \\ \hline Z_{\text{eff}} & > & \text{Screening effect} \end{array}$$

$$\begin{array}{ccc} \text{Fe} & \text{Co} & \text{Ni} \\ \hline Z_{\text{eff}} & \approx & \text{Screening effect} \end{array}$$

$$\begin{array}{cc} \text{Cu} & \text{Zn} \\ \hline Z_{\text{eff}} & < \text{Screening effect} \end{array}$$

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

● **Lanthanide Contraction :**

(a) Outermost electronic configuration of inner transition elements is

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

(b) e^- enters in $(n-2)f$ orbitals

(c) Because of complicated structure of f-orbital and due to poor shielding f electrons, the outermost shell electrons get attracted towards nucleus.

(d) In 1st, 2nd and 3rd transition series, Radii— $3d < 4d \approx 5d$ (**except IIIrd B**)

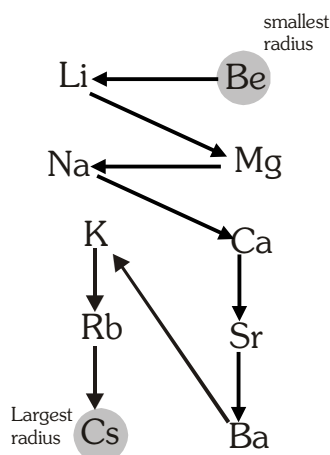
	IIIB	IVB	
size increases ↓	Sc	Ti ↓ size increases	
	Y	Zr	} Almost equal due to lanthanide contraction
	La	Hf	

● **Transition contraction :**

Note : While atomic size should increase down the group.

At. size of Ga \approx At. size of Al, due to transition contraction. (Due to poor shielding of d electrons)

● **s-block size variation**



BEGINNER'S BOX-3

- From the given set of species, point out the species from each set having least atomic radius:-
 (1) O^{2-} , F^- , Na^+ (2) Ni, Cu, Zn (3) Li, Be, Mg (4) He, Li^+ , H^-
 Correct answer is :-
 (1) O^{2-} , Cu, Li, H^- (2) Na^+ , Ni, Be, Li^+ (3) F^- , Zn, Mg, He (4) Na^+ , Cu, Be, He
- Arrange the elements in increasing order of atomic radius Na, Rb, K, Mg :-
 (1) Na, K, Mg, Rb (2) K, Na, Mg, Rb (3) Mg, Na, K, Rb (4) Rb, K, Mg, Na
- Which of the following pairs of elements have almost similar atomic radii :-
 (1) Zr, Hf (2) Mo, W (3) Co, Ni (4) All
- 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 :-
 (1) 1.34 Å, 1.34 Å (2) 0.72 Å, 1.96 Å (3) 1.96 Å, 0.72 Å (4) 1.96 Å, 1.34 Å
- 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 :-
 (1) Crystal, Vander Waal and Covalent radii (2) Covalent, crystal and Vander Waal radii
 (3) Vander Waal, ionic and covalent radii (4) None is correct.
- Screening effect is not observed in :-
 (1) He^+ (2) Li^{+2} (3) H (4) All of these
- Arrange in orders of atomic and ionic radii :
 (a) Ni, Cu, Zn (b) H^+ , H, H^- (c) Ti, Zr, Hf (d) Be, Li, Na
 (e) Cr, V, Ti, Sc (f) I^+ , I, I^- (g) Sc, Y, La, Ac (h) Cl, Na, Rb
 (i) Cu, Ag, Au (j) B, Be, Al, Mg (k) F, O, Cl, S

8. The screening effect of d- electrons is :-
 (1) Equal to the p - electrons (2) Much more than p - electrons
 (3) Same as f - electrons (4) Less than p - electrons
9. Covalent radius of an element represents -
 (1) Half of the distance between two similar atoms when they are bonded together by a single bond.
 (2) Half of the distance between two similar atoms when they are bonded together by a double bond.
 (3) Half of the distance between two similar atoms when they are bonded together by a triple bond.
 (4) covalent radius of an atom can be measured by any method
10. Select the incorrect order of atomic radii.
 (1) $\text{Li} > \text{I} > \text{Br} > \text{Cl} > \text{F}$ (2) $\text{Na} > \text{Li} > \text{F}$ (3) $\text{I} > \text{Br} > \text{Cl} > \text{Li} > \text{F}$ (4) $\text{Cs} > \text{Rb} > \text{K} > \text{Na} > \text{Li}$

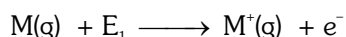
1.4 IONISATION POTENTIAL OR IONISATION ENERGY OR IONISATION ENTHALPY (IP / IE)

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

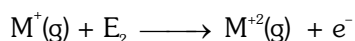
(Isolated \rightarrow Without any bonding with other atom)

(ii) Successive Ionization Energy

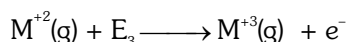
(a) For an atom $M_{(g)}$ successive ionization energies are as follows -



$E_1 = \text{I}^{\text{st}}$ Ionization energy



$E_2 = \text{II}^{\text{nd}}$ Ionization energy



$E_3 = \text{III}^{\text{rd}}$ ionization energy

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

- (b) Electron can not 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) For any neutral atom ionization energy is always an endothermic process ($\Delta H = +ve$)
- (d) It is measured in eV/atom (electron volt/atom) or Kcal/mole or KJ/mole

FACTORS AFFECTING IONISATION ENERGY

In a period

(i) Effective nuclear charge (Z_{eff})

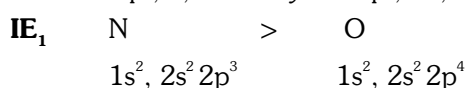
$$\text{Ionisation Energy} \propto Z_{\text{eff}} \propto \frac{\text{positive charge}}{\text{negative charge}}$$

Ion with high positive oxidation state will have high ionisation energy.

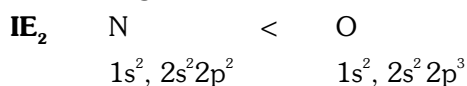
Ex. $\text{Fe}^{+3} > \text{Fe}^{+2} > \text{Fe}$

(ii) Stability of half filled and fully filled orbitals :

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



After loosing one e^- , O attains electronic configuration of N, so II^{nd} ionisation energy of O is more than N.



(iii) Penetration power of sub shells

- (a) Order of attraction of subshells towards nucleus (Penetration power) is $s > p > d > f$
 (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.

Ex. Be B**IE₁** Be > BAfter loosing one e^- , B attains electronic configuration of Be, so IInd ionisation energy of B is more than Be.**IE₂** Be < B**In a group****Atomic size :** $IE \propto \frac{1}{\text{atomic size}}$ **COMPARISON OF IONISATION ENERGY****(i) In a period :** Z_{eff} increases so removal of electron becomes difficult and ionisation energy increases.Order of IE of 2nd period elements $Li < B < Be < C < O < N < F < Ne$ **(ii) In a group :** Size increase so ionisation energy decrease.**Exception :**

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

Ex. Hf > Zr**Application of ionisation energy :****(A) Metallic and non metallic character :**

Generally for metals Ionisation Energy is low.

For Non-metals Ionisation Energy is high.

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

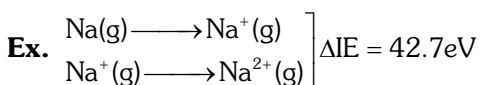
(B) Reactivity of metals :

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

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

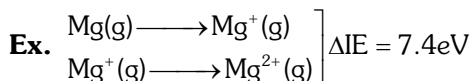
(C) Stability of oxidation states of an element :

- (a) If the difference between two successive ionisation energy of an element $\geq 16\text{eV}$, then its lower oxidation state is stable.

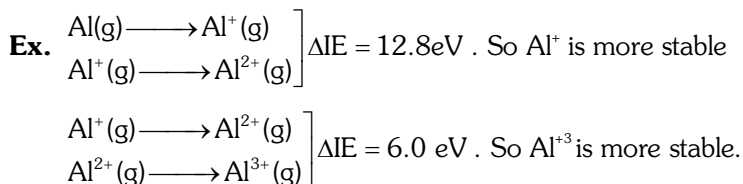
Difference between ionisation energy > 16 eV. So Na^+ is more stable.

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- (b) If the difference between two successive ionisation energy of an element ≤ 11 eV, then its higher oxidation state is stable.



Difference of ionisation energy < 11 eV. So Mg^{+2} is more stable.



Overall order of stability is $\boxed{\text{Al}^{+3} > \text{Al}^+ > \text{Al}^{+2}}$

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

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

BEGINNER'S BOX-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 :-
 (1) + 170 K.cal (2) + 526 K.cal (3) - 170 K.cal (4) - 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:-
 (1) Na (2) Si (3) F (4) Ca
- Which of the following element has 2nd IP < 1st IP
 (1) Mg (2) Ne (3) C (4) None
- In which of the following the energy change corresponds to first ionisation potential only :-
 (1) $\text{X(g)} \rightarrow \text{X}^+(\text{g}) + \text{e}^-$ (2) $\text{X}_2(\text{g}) \rightarrow \text{X}^+(\text{g}) + \text{e}^-$ (3) $\text{X(s)} \rightarrow \text{X}^+(\text{g}) + \text{e}^-$ (4) $\text{X(aq)} \rightarrow \text{X}^+(\text{aq}) + \text{e}^-$
- In the given process which oxidation state is more stable.

$$\begin{array}{l} \text{M(g)} \longrightarrow \text{M}^+(\text{g}) \quad IE_1 = 7.9 \text{ eV} \\ \text{M}^+(\text{g}) \longrightarrow \text{M}^{+2}(\text{g}) \quad IE_2 = 15.5 \text{ eV} \end{array}$$

 (1) M⁺ (2) M⁺² (3) Both (4) None
- The correct order of decreasing second ionization energy of Li, Be, Ne, C, B
 (1) Ne>B>Li>C>Be (2) Li>Ne>C>B>Be (3) Ne>C>B>Be>Li (4) Li>Ne>B>C>Be
- In which of the following element has highest value of ionisation energy-
 (1) V (2) Nb (3) Ta (4) Cs
- The first ionisation enthalpy of Boron is slightly less than Beryllium because :-
 (1) 2p-electron is more shielded than 2s-electron
 (2) effective nuclear charge in Beryllium is more than Boron
 (3) of inert pair effect
 (4) lanthanoid contraction

9. Match the column.

Column-I**Valence electronic configuration**

- (a) ns^1
 (b) ns^2
 (c) $ns^2 np^1$
 (d) $ns^2 np^2$

Select the correct option

- (1) a-p, b-q, c-r, d-s
 (3) a-s, b-r, c-q, d-p

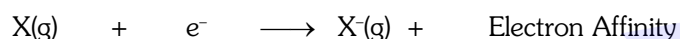
Column-II**Successive ionisation energies**

- (p) 19, 27, 36, 48, 270
 (q) 16, 28, 34, 260
 (r) 18, 26, 230, 250
 (s) 14, 200, 220, 240

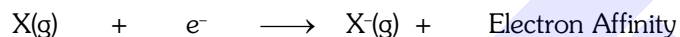
- (2) a-s, b-r, c-p, d-q
 (4) a-r, b-s, c-p, d-q

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

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

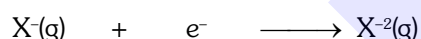


- (2) Generally first electron addition of an isolated gaseous atom is an exothermic process (except stable electronic configuration)

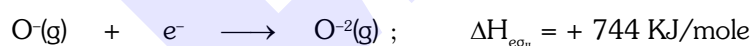


$$\Delta H_{egI} (\text{first electron gain enthalpy}) = -ve$$

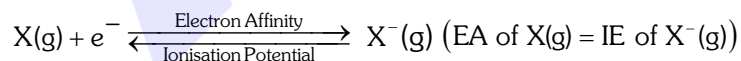
- (3) Second electron addition of an isolated gaseous atom is always an endothermic process due to inter electronic repulsion



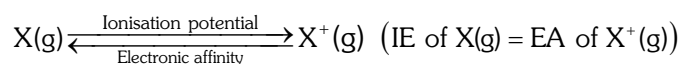
$$\Delta H_{egII} (\text{second electron gain enthalpy}) = \text{positive}$$



- (4) Formation of poly negative 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.



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



- (7) **Factors affecting electron affinity :**

(A) **Atomic size :** Electron Affinity $\propto \frac{1}{\text{Atomic size}}$

(B) **Effective nuclear charge (Z_{eff}) :** Electron Affinity $\propto Z_{eff} \propto \frac{\text{positive charge}}{\text{negative charge}}$

- (C) **Stability of completely filled or half filled orbitals** : Electron affinity of elements having full-filled or half filled configuration is very less or zero so for these elements electron gain enthalpy (ΔH_{eg}) will be positive.

Ex. (i) $EA \Rightarrow ns^1 > ns^2$ (ii) $EA \Rightarrow np^2 > np^3$

(8) **Variation of electron affinity :**

(i) **In 2nd period –**

Li Be B C N O F Ne

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

(ii) **In Group :**

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

F Cl
 $[\text{He}] 2s^2 2p^5$ $[\text{Ne}] 3s^2 3p^5$

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

$\text{Cl} > \text{F} > \text{Br} > \text{I}$	$\text{S} > \text{O} > \text{P} > \text{N}$	$\text{Si} > \text{C} > \text{P} > \text{N}$
---	---	--

Note : N & P have low electron affinity due to stable half filled configuration.

BEGINNER'S BOX-5

- The correct order of electron affinity is :-
 (1) $\text{Be} < \text{B} < \text{C} < \text{N}$ (2) $\text{Be} < \text{N} < \text{B} < \text{C}$
 (3) $\text{N} < \text{Be} < \text{C} < \text{B}$ (4) $\text{N} < \text{C} < \text{B} < \text{Be}$
- In the formation of a chloride ion, from an isolated gaseous chlorine atom, 3.8 eV energy is released, which would be equal to :-
 (1) Electron affinity of Cl^- (2) Ionisation potential of Cl
 (3) Electronegativity of Cl (4) Ionisation potential of Cl^-
- $\text{O(g)} + 2e^- \rightarrow \text{O}^{2-}(\text{g})$ $\Delta H_{\text{eg}} = 603 \text{ KJ/mole}$. The positive value of ΔH_{eg} is due to :-
 (1) Energy is released to add on $1 e^-$ to O^- (2) Energy is required to add on $1 e^-$ to O^-
 (3) Energy is needed to add on $1 e^-$ to O (4) None of the above is correct
- The electron affinity values for the halogens shows the following trend :-
 (1) $\text{F} < \text{Cl} > \text{Br} > \text{I}$ (2) $\text{F} < \text{Cl} < \text{Br} < \text{I}$
 (3) $\text{F} > \text{Cl} > \text{Br} > \text{I}$ (4) $\text{F} < \text{Cl} > \text{Br} < \text{I}$
- The process requiring the absorption of energy is.
 (1) $\text{F} \rightarrow \text{F}^-$ (2) $\text{Cl} \rightarrow \text{Cl}^-$ (3) $\text{O} \rightarrow \text{O}^{2-}$ (4) $\text{H} \rightarrow \text{H}^-$
- Second electron affinity of an element is :-
 (1) Always exothermic (2) Endothermic for few elements
 (3) Exothermic for few elements (4) Always endothermic

7. Process, $\text{Na}^+(\text{g}) \xrightarrow{\text{I}} \text{Na}(\text{g}) \xrightarrow{\text{II}} \text{Na}(\text{s})$
 (1) In (I) energy released, (II) energy absorbed (2) In both (I) and (II) energy is absorbed
 (3) In both (I) and (II) energy is released (4) In (I) energy absorbed, (II) energy released
8. Which of the following configuration will have least electron affinity. Where ($n = 2, 3$)
 (1) ns^2np^5 (2) ns^2np^2 (3) ns^2np^3 (4) ns^2np^4
9. Which of the following will have the most negative electron gain enthalpy and which the least negative ?
 (1) F, Cl (2) Cl, F (3) Cl, S (4) Cl, P
10. Which arrangement represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species ?
 (1) $S < O < Cl < F$ (2) $O < S < F < Cl$ (3) $Cl < F < S < O$ (4) $F < Cl < O < S$

1.6 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 bond of A – B may be broken as
 $\text{A} - \text{B} \longrightarrow \text{A}^{\delta-} - \text{B}^{\delta+}$ (Electronegativity A > Electronegativity B)
 depending on their tendency to attract bonded electron.
- (iii) **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 not depend on stable electronic configuration It has no unit 	<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 depend on stable electronic configuration. It is measured in eV/atom or KJ mol^{-1} or Kcal mole^{-1}.

- (iv) 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						Cl 3.0
Rb 0.8						Br 2.8
Cs 0.7						I 2.5
Fr 0.7						

In Pauling's scale, elements having almost same electronegativity are-

$N = Cl = 3.0$
 $C = S = I = 2.5$
 $P = H = 2.1$
 $Be = Al = 1.5$
 $K = Rb = 0.8$
 $Cs = Fr = 0.7$

Note : Small atoms are normally having more electronegativity than larger atoms.

(v) FACTORS AFFECTING ELECTRONEGATIVITY :
(A) Atomic size

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

Ex. $F > Cl > Br > I$

(B) Effective nuclear charge (Z_{eff})

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

Ex. $Mn^{+2} < Mn^{+4} < Mn^{+7}$
 $O^{-2} < O^{-1} < O < O^{+1} < O^{+2}$
 $Fe < Fe^{+2} < Fe^{+3}$
 $\xrightarrow{Z_{\text{eff}} \uparrow \text{EN} \uparrow}$

(C) % s - character

$$\text{Electronegativity} \propto \% \text{ s - Character Hybridised atom}$$

(vi) PERIODIC TABLE & ELECTRONEGATIVITY :

(A) Electronegativity decreases down the group.

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

(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 balanced the effect of increase in number of shell.

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

Electronegativity of Ga > Electronegativity of Al

(vii) APPLICATION OF ELECTRONEGATIVITY :
(A) Metallic and non metallic nature :

Generally metals have low electronegativity and non metals have high electronegativity, so we can say metallic character increases down the group but decreases along a period.

$$\text{Non Metallic Nature} \propto \text{EN}$$

(B) Bond energy : By increasing difference in electronegativity of bonded atoms, bond length decreases and hence bond energy increases

$$\text{Bond energy} \propto \text{Electronegativity difference}$$

$HF > HCl > HBr > HI$

(C) Schoemaker and Stevenson law

If in a diatomic molecule electronegativities of A – B have more difference. Then actual bond length will be reduced. As per schoemaker & Stevenson– The reduction in bond length depends on the difference in electronegativities of atoms by 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

Ex. 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.

$$d_{H-F} = r_F + r_H - 0.09 (X_F - X_H)$$

$$\therefore r_F = 1.44 / 2 = 0.72 \text{ \AA}, r_H = 0.74 / 2 = 0.37 \text{ \AA}$$

$$\therefore d_{H-F} = 0.72 + 0.37 - 0.09 (4.0 - 2.1)$$

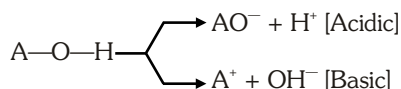
$$= 1.09 - (0.09 \times 1.9) = 1.09 - 0.171 = 0.919 \text{ \AA}$$

(D) Acidic & Basic Strength :**(i) Nature of non-metallic hydrides :**

Stability of molecule \propto Bond energy

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 of hydroxides :**(a) As per Gallis,**

- (i) In AOH if electronegativity of A is more than 1.7 (Non metal) then it is acidic in nature.
 (ii) If electronegativity of 'A' is less than 1.7 (metal) then AOH will be basic in nature
 (b) If $X_A - X_O \geq X_O - X_H$ (X_A = EN of A) then AO bond will be more polar and will break up as
 $\text{A} - \text{OH} \rightarrow \text{A}^+ + \text{OH}^-$ It shows basic nature

Ex. In NaOH
 $X_O - X_{Na} (2.6) > X_O - X_H (1.4)$ So hydroxide is basic

- (c) If $X_A - X_O \leq X_O - X_H$ then OH bond will be more polar and will break up as
 $\text{A} - \text{O} - \text{H} \rightarrow \text{H}^+ + \text{AO}^-$ It shows Acidic nature

In ClOH
 $X_O - X_{Cl} (0.5) < X_O - X_H (1.4)$ So hydroxide is acidic
(iii) Nature of oxides :

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

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

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

NO ₂	>	ZnO	>	K ₂ O
→				
acidic character decreases				

Note : BeO, Al₂O₃, ZnO, SnO, PbO, SnO₂, PbO₂, Sb₂O₃ etc. are amphoteric oxides.

CO, H₂O, NO, N₂O etc. are neutral oxides.

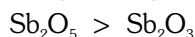
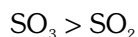
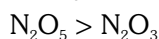
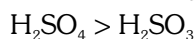
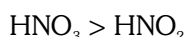
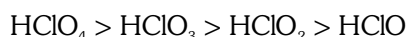
Acidic strength of oxide and oxyacid \propto Electronegativity
--

B ₂ O ₃	CO ₂	N ₂ O ₅
→		
EN increase, acidic nature increase		

HNO ₃ > H ₂ CO ₃ > H ₃ BO ₃	HClO ₄ > HBrO ₄ > HIO ₄	SO ₃ > SeO ₃ > TeO ₃
H ₃ PO ₄ > H ₃ AsO ₄ > H ₃ SbO ₄	N ₂ O ₅ > P ₂ O ₅ > As ₂ O ₅	H ₂ SO ₃ > H ₂ SeO ₃ > H ₂ TeO ₃
HOI > HOCl > HOBr > HOI	N ₂ O ₃ > P ₂ O ₃ > As ₂ O ₃ > Sb ₂ O ₃	

Acidic nature \propto oxidation state

Acidic properties increases with increasing oxidation state of an element



(E) Nature of bonds :

(a) 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

(b) 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

$A^{\delta-} \text{ --- } B^{\delta+}$ bond will be polar covalent

Ex. H_2O ($\text{H}^{\delta+} \text{ --- } \text{O}^{\delta-} \text{ --- } \text{H}^{\delta+}$)

If $X_A \gg X_B$ and $X_A - X_B$ difference of electronegativities is high then

$A^- \text{ --- } B^+$ bond will be polar or ionic

(F) Nomenclature of inorganic compounds :

Prefix — less electronegative element

Suffix — More electronegative element

Ex. Cl_2O (Right) OCl_2 (Wrong)

In Dichloroxide the electronegativity of Cl is less than 'O' i.e. why Cl is in prefix position.

OF_2 Oxygen difluoride

ICl Iodine chloride

(G) Bond polarity $\propto \Delta EN$

Ex. $\text{HF} > \text{HCl} > \text{HBr} > \text{HI}$

(viii) ELECTRONEGATIVITY SCALE :

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

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

where X_p is electronegativity on the basis of Pauling scale.

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

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

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

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

$$X_p = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{2 \times 62.5}$$

BEGINNER'S BOX-6

- Which of the following is affected by stable configuration of an atom :-
 (a) Electronegativity (b) Ionisation potential (c) Electron affinity
 Correct answer is :-
 (1) Only electronegativity (2) Only ionisation potential
 (3) Electron affinity and ionisation potential (4) All of the above
- Which of the following elements have different value of electronegativity :-
 (1) H (2) S (3) Te (4) P
- Which is the correct order of electronegativity -
 (1) $\text{Cl} > \text{S} > \text{P} > \text{Si}$ (2) $\text{Si} > \text{Al} > \text{Mg} > \text{Na}$ (3) $\text{F} > \text{Cl} > \text{Br} > \text{I}$ (4) All
- Electronegativity scale of Pauling is based upon :-
 (1) Bond length (2) Bond energy (3) Atomic radius (4) Covalent radius
- Correct order of electronegativity of N, P, C and Si is :-
 (1) $\text{N} < \text{P} < \text{C} < \text{Si}$ (2) $\text{N} > \text{C} > \text{Si} > \text{P}$ (3) $\text{N} = \text{P} > \text{C} = \text{Si}$ (4) $\text{N} > \text{C} > \text{P} > \text{Si}$
- Outermost electronic configuration of the most electronegative element is :-
 (1) ns^2np^3 (2) ns^2np^6 (3) ns^2 (4) ns^2np^5
- Electronegativity of the following elements increases in the order.
 (1) O, N, S, P (2) P, S, N, O (3) P, N, S, O (4) S, P, N, O
- In periodic table, the basic character of oxides/hydroxides ?
 (1) Increase from left to right and decreases from top to bottom
 (2) decreases from right to left and increases from top to bottom
 (3) decreases from left to right and increases from top to bottom
 (4) Increases from left to right and decreases from top to bottom

ANSWER KEY

BEGINNER'S BOX-1	Que.	1	2	3	4	5	6	7	8	9	10
	Ans.	2	1	4	2	4	2	3	4	4	No
BEGINNER'S BOX-2	Que.	1	2	3	4	5	6	7	8	9	10
	Ans.	4	3	3	1	4	2	1	2	2	2
BEGINNER'S BOX-3	Que.	1	2	3	4	5	6	7	8	9	10
	Ans.	2	3	4	3	1	4		4	1	3
BEGINNER'S BOX-4	Que.	1	2	3	4	5	6	7	8	9	
	Ans.	2	2	4	1	2	4	3	1	3	
BEGINNER'S BOX-5	Que.	1	2	3	4	5	6	7	8	9	10
	Ans.	2	4	2	1	3	4	3	3	4	2
BEGINNER'S BOX-6	Que.	1	2	3	4	5	6	7	8		
	Ans.	3	2	4	2	4	4	2	3		