

I N D E X

IIT JAM INORGANIC – THEORY BOOK

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CHAPTER 1

PERIODICITY

INTRODUCTION:

Dobereiner's Triads

In 1829, Dobereiner arranged certain elements with similar properties in groups of three in such a way that the atomic mass of the middle element was nearly the same as the average atomic masses of the first and the third elements. A few triads proposed by him are listed.

S. No.	Element	Atomic Mass	Mean of I and III
1.	I. Lithium	7	$\frac{7 + 39}{2} = 23$
	II. Sodium	23	
	III. Potassium	39	
2.	I. Calcium	40	$\frac{40 + 137}{2} = 88.5$
	II. Strontium	88	
	III. Barium	137	
3.	I. Chlorine	35.5	$\frac{35.5 + 127}{2} = 81.25$
	II. Bromine	80	
	III. Iodine	127	

Limitations of Dobereiner's Triads

The triads given by Dobereiner were helpful in grouping some elements with similar characteristics together, but he could not arrange all the elements known at that time into triads.

Newlands' Law of Octaves

John Newlands proposed the law of octaves by stating that when elements are arranged in order of increasing atomic masses, every eighth element has properties similar to the first. Newlands called it law of octaves because similar relationship exists in the musical notes also.

This can be illustrated as,

Newlands' Arranged Elements in Octaves:

H	F	Cl	Co/Ni	Br	Pd	I	Pt/Ir
Li	Na	K	Cu	Rb	Ag	Cs	Tl
G	Mg	Ca	Zn	Sr	Cd	Ba/V	Pb
Bo	Al	Cr	Y	Ce/La	U	Ta	Th
C	Si	Ti	In	Zn	Sn	W	Hg
N	P	Mn	As	Di/Mo	Sb	Nb	Bi
O	S	Fe	Se	Ro/Ru	Te	Au	Os

Limitations of Newlands' Law of Octaves

- This classification was successful only up to the element calcium. After that, every eighth element did not possess the same properties as the element lying above it in the same group.
- When noble gas elements were discovered at a later stage, their inclusion in the table disturbed the entire arrangement.

Mendeleev's Periodic Table

Mendeleev's Periodic Law: The physical and chemical properties of the elements are a periodic function of their atomic masses. Mendeleev arranged the elements known at that time in order of increasing atomic masses and this arrangement was called periodic table.

Elements with similar characteristics were present in vertical rows called groups. The horizontal rows were known as periods.

Description Mendeleev's Periodic Table

- In the periodic table, the elements are arranged in vertical rows called groups and horizontal rows known as periods.
- There are nine groups indicated by Roman Numerals as I, II, III, IV, V, VI, VII, VIII and zero. Group VIII consists of nine elements which are arranged in three triads. The zero group contains elements belonging to inert gases or noble gases and elements present have zero valency.
- There are seven periods (numbered from 1 to 7) or, horizontal rows in the Mendeleev's periodic table.

I	II	III	IV	V	VI	VII	VIII		
H 1.01									
Li 6.94	Be 9.01	B 10.8	C 12.0	N 14.0	O 16.0	F 19.0			
Na 23.0	Mg 24.3	Al 27.0	Si 28.1	P 31.0	S 32.1	Cl 35.5			
K 39.1	Ca 40.1		Ti 47.9	V 50.9	Cr 52.0	Mn 54.9	Fe 55.9	Co 58.9	Ni 58.7
Cu 63.5	Zn 65.4			As 74.9	Se 79.0	Br 79.9			
Rb 85.5	Sr 87.6	Y 88.9	Zr 91.2	Nb 92.9	Mo 95.9		Ru 101	Rh 103	Pd 106
Ag 108	Cd 112	In 115	Sn 119	Sb 122	Te 128	I 127			
Ce 133	Ba 137	La 139		Ta 181	W 184		Os 194	Ir 192	Pt 195
Au 197	Hg 201	Ti 204	Pb 207	Bi 209					
		Th 232			U 238				

Mordern Periodic Law

In the periodic table the elements are arranged in order of increasing atomic number that is in order of increased nuclear charge, or increased number of orbital electrons. Thus each element contains one more orbital electron than the preceding element. Instead of listing the 118 elements as one long list, the periodic table arranges

them into several horizontal rows or periods, in such a way that each row begins with an alkali metal and ends with a noble gas. The sequence in which the various energy levels are filled determines the number of elements in each period, and the periodic table can be divided into four main regions according to whether the s, p, d or f levels are being filled.

1

11A

1A

2

IIA

2A

13

IIIA

3A

14

IVA

4A

15

VA

5A

16

VIA

6A

17

VIIA

7A

18

VIIIA

8A

1

H

Hydrogen

1.008

2

He

Helium

4.003

3

Li

Lithium

6.941

4

Be

Beryllium

9.012

11

Na

Sodium

22.99

12

Mg

Magnesium

24.305

19

K

Potassium

39.098

20

Ca

Calcium

40.078

37

Rb

Rubidium

85.468

38

Sr

Strontium

87.62

55

Cs

Cesium

132.905

56

Ba

Barium

137.328

87

Fr

Francium

223.020

10

Ne

Neon

20.180

17

F

Fluorine

18.998

35

Cl

Chlorine

35.453

53

I

Iodine

126.904

85

At

Astatine

208.987

18

Ar

Argon

39.948

36

Kr

Krypton

83.799

54

Xe

Xenon

131.294

86

Rn

Radon

222.018

118

Og

Oganesson

294

13

B

Boron

10.811

14

C

Carbon

12.011

15

N

Nitrogen

14.007

16

O

Oxygen

15.999

17

F

Fluorine

18.998

18

Ne

Neon

20.180

29

Cu

Copper

63.546

30

Zn

Zinc

65.38

47

Ag

Silver

107.868

48

Cd

Cadmium

112.414

63

Eu

Europium

151.964

64

Gd

Gadolinium

157.25

65

Tb

Terbium

158.925

66

Dy

Dysprosium

162.500

67

Ho

Holmium

164.930

68

Er

Erbium

167.259

69

Tm

Thulium

168.934

70

Yb

Ytterbium

173.055

71

Lu

Lutetium

174.967

72

Hf

Hafnium

178.49

73

Ta

Tantalum

180.948

74

W

Tungsten

183.84

75

Re

Rhenium

186.207

76

Os

Osmium

190.23

77

Ir

Iridium

192.217

78

Pt

Platinum

195.085

79

Au

Gold

196.967

80

Hg

Mercury

200.592

81

Tl

Thallium

204.383

82

Pb

Lead

207.2

83

Bi

Bismuth

208.980

84

Po

Polonium

209

85

At

Astatine

208.987

86

Rn

Radon

222.018

87

Fr

Francium

223.020

88

Ra

Radium

226.025

89-103

104

Rf

Rutherfordium

261

105

Db

Dubnium

262

106

Sg

Seaborgium

266

107

Bh

Bohrium

264

108

Hs

Hassium

269

109

Mt

Meitnerium

278

110

Ds

Darmstadtium

281

111

Rg

Roentgenium

280

112

Cn

Copernicium

285

113

Nh

Nihonium

286

114

Fl

Flerovium

289

115

Mc

Moscovium

289

116

Lv

Livermorium

293

117

Ts

Tennessine

294

118

Og

Oganesson

294

57

La

Lanthanum

138.905

58

Ce

Cerium

140.118

59

Pr

Praseodymium

140.908

60

Nd

Neodymium

144.243

61

Pm

Promethium

144.913

62

Sm

Samarium

150.36

63

Eu

Europium

151.964

64

Gd

Gadolinium

157.25

65

Tb

Terbium

158.925

66

Dy

Dysprosium

162.500

67

Ho

Holmium

164.930

68

Er

Erbium

167.259

69

Tm

Thulium

168.934

70

Yb

Ytterbium

173.055

71

Lu

Lutetium

174.967

89

Ac

Actinium

227.028

90

Th

Thorium

232.038

91

Pa

Protactinium

231.036

92

U

Uranium

238.029

93

Np

Neptunium

237.044

94

Pu

Plutonium

244.064

95

Am

Americium

243.061

96

Cm

Curium

247.070

97

Bk

Berkelium

247.070

98

Cf

Californium

251.080

99

Es

Einsteinium

254

100

Fm

Fermium

257.095

101

Md

Mendelevium

258.1

102

No

Nobelium

259.101

103

Lr

Lawrencium

262

Alkali Metal

Alkaline Earth

Transition Metal

Basic Metal

Semimetal

Nonmetal

Halogen

Noble Gas

Lanthanide

Actinide

Periodic Table of the Elements

11A

1A

IIA

IIIA

IVA

VA

VIA

VIIA

VIIIA

8A

1st period	1s	elements in this period 2
2nd period	2s	2p elements in this period 8
3rd period	3s	3p elements in this period 8
4th period	4s 3d	4p elements in this period 18
5th period	5s 4d	5p elements in this period 18
6th period	6s 4f 5d	6p elements in this period 32

Groups

The long form of periodic table also consists of the vertical rows called groups. There are in all 18 groups in the periodic table. Unlike Mendeleev periodic table, each group is an independent group

Characteristics of Groups:

- (1) All the elements present in a group have same general electronic configuration of the atoms.
- (2) The elements in a group are separated by definite gaps of atomic numbers (2, 8, 8, 18, 18, 32) as magic numbers.
- (3) The atomic sizes of the elements in group increase down the group due to increase the number of shells.

- (4) The physical properties of the elements such as m.p., b.p. density, solubility etc., follow a systematic pattern.
- (5) The elements in each group have generally similar chemical properties.

Periods

Horizontal rows in a periodic table are known as periods. There are in all seven periods in the long form of periodic table.

Characteristics of periods:

- (i) In all the elements present in a period, the electrons are filled in the same valence shell.
- (ii) The atomic sizes generally decrease from left to right.

Periodic Classification of Elements Based on Electronic Configuration

The long form of the periodic table is divided into four blocks known as the s, p, d and f-blocks.

S-block Elements

The elements in the s-block of the periodic table have the last electron filled in valence s-sub-shell of the outermost energy shell. Because the s-subshell can only have two electrons, this block comprises only two groups (1,2). The elements included in group 1 have ns^1 electronic configuration and are called alkali metals. Similarly, group 2 contains alkaline earth metals with ns^2 electronic configuration. Thus, the elements of group 1 and 2, including hydrogen and helium in which the s-orbitals are being progressively filled in are called s-block elements.

General Electronic Configuration of s-block Elements is:

$$ns^{1-2}ns^{1-2}$$

where $n=2-7$

p-block Elements

P-block elements are those in which the last electron enters one of the three p-orbitals of their corresponding outermost shells. As the p-subshell can have a maximum of six electrons, there are six groups in this block (13 to 18).

General Electronic Configuration of p-block Elements is:

$$ns^2np^{1-6}ns^2np^{1-6}$$

where $n=2-7$

The elements belong to s and p-blocks in the periodic table are collectively called Representative elements.

d-block Elements

Elements in which the last electron enters any one of the five d-orbitals of their respective penultimate shells are called d-block elements. Because the d-subshell can have a maximum of five orbitals and ten electrons, the d-block is divided into ten vertical columns or groups numbered 3 to 12. The d-block elements are known as the transition elements because they have incompletely filled d-orbitals in their ground state.

General Electronic Configuration of p-block Elements is:

$$(n-1)d^{1-10}ns^{0-2}(n-1)d^{1-10}ns^{0-2}$$

where $n=4-7$

f-block Elements

Elements in which the last electron enters any one of the seven f-orbitals of their respective antepenultimate shells are called d-block elements. All the f-block elements are also called two such series, each with fourteen elements; in the first series, the filling of electrons takes place in the 4f-subshell, known as lanthanoid series (Atomic number 58–71). In the second series, the filling takes place in the 5f-subshell, and it is called the actinide series.

General electronic configuration of p-block elements is:

$$(n-2)f^{0-14}(n-1)d^{0-2}ns^2(n-2)f^{0-14}(n-1)d^{0-2}ns^2$$

where $n=6-7$

Morden Periodic Law

The elements in a group are separated by definite gaps of atomic numbers (2,8,8,18,18,32) as magic numbers and The atomic sizes of the elements in group increase down the group due to increase the number of shells.

Orbitals and Quantum Numbers:

A large number of orbitals are possible in an atom. Qualitatively these orbitals can be distinguished by their size, shape and orientation. An orbital of smaller size means there is more chance of finding the electron near the nucleus. Similarly shape and orientation mean that there is more probability of finding the electron along certain directions than along others. Atomic orbitals are precisely distinguished by what are known as quantum numbers. Each orbital is designated by three quantum numbers labelled as n, l and m_l .

The principal Quantum Number 'n'

It is a positive integer with value of $n = 1, 2, 3, \dots$. The principal quantum number determines the size and to large extent the energy of the orbital. For hydrogen atom and hydrogen like species (He^+ , Li^{2+} , etc.) energy and size of the orbital depends only on 'n'. The principal quantum number also identifies the shell. With the increase in the value of 'n', the number of allowed orbital increases and are given by 'n²'. All the orbitals of a given value of 'n' constitute a single shell of atom and are represented by the following letters

$$n = 1 \ 2 \ 3 \ 4 \ \dots\dots\dots$$

$$\text{Shell} = K \ L \ M \ N \ \dots\dots\dots$$

Size of an orbital increases with increase of principal quantum number 'n'. In other words the electron will be located away from the nucleus. Since energy is required in

shifting away the negatively charged electron from the positively charged nucleus, the energy of the orbital will increase with increase of n .

Azimuthal Quantum Number. 'l'

It is also known as orbital angular momentum or subsidiary quantum number. It defines the three-dimensional shape of the orbital. For a given value of n , l can have n values ranging from 0 to $n - 1$, that is, for a given value of n , the possible value of l are $l = 0, 1, 2, \dots, (n-1)$

For example, when $n = 1$, value of l is only 0. For $n = 2$, the possible value of l can be 0 and 1. For $n = 3$, the possible l values are 0, 1 and 2.

Each shell consists of one or more subshells or sub-levels. The number of sub-shells in a principal shell is equal to the value of n .

For example in the first shell ($n = 1$), there is only one sub-shell which corresponds to $l = 0$. There are two sub-shells ($l = 0, 1$) in the second shell ($n = 2$), three ($l = 0, 1, 2$) in third shell ($n = 3$) and so on. Each sub-shell is assigned an azimuthal quantum number (l). Sub-shells corresponding to different values of l are represented by the following symbols.

Value for l : 0 1 2 3 4 5

Notation for : s p d f g h

Sub-shell

Table shows the permissible values of 'l' for a given principal quantum number and the corresponding sub-shell notation.

Table: Subshell Notations

n	l	Subshell Notation
1	0	1s
2	0	2s
2	1	2p
3	0	3s
3	1	3p
3	2	3d
4	0	4s
4	1	4p
4	2	4d
4	3	4f

Magnetic Orbital Quantum Number. 'ml'

It gives information about the spatial orientation of the orbital with respect to standard set of co-ordinate axis. For any sub-shell (defined by 'l' value) $2l+1$ values of ml are possible and these values are given by:

$$ml = -l, -(l-1), -(l-2) \dots 0, 1 \dots (l-2), (l-1), l$$

Thus for $l = 0$, the only permitted value of $ml = 0$, [$2(0)+1 = 1$, one s orbital]. For $l = 1$, ml can be $-1, 0$ and $+1$ [$2(1)+1 = 3$, three p orbitals]. For $l = 2$, $ml = -2, -1, 0, +1$ and $+2$, [$2(2)+1 = 5$, five d orbitals].

It should be noted that the values of ml are derived from l and that the value of l are derived from n . Each orbital in an atom, therefore, is defined by a set of values for n , l and ml . An orbital described by the quantum numbers $n = 2, l = 1, ml = 0$ is an orbital in the p sub-shell of the second shell. The following chart gives the relation between the subshell and the number of orbitals associated with it.

Value of l	0	1	2	3	4	5
Subshell Notation	s	p	d	f	g	h
Number of Orbitals	1	3	5	7	9	11

Electron spin quantum number (m_s). An electron spins around its own axis, much in a similar way as earth spins around its own axis while revolving around the sun. In other words, an electron has, besides charge and mass, intrinsic spin angular quantum number.

Spin angular momentum of the electron — a vector quantity, can have two orientations relative to the chosen axis. These two orientations are distinguished by the spin quantum numbers m_s which can take the values of $+\frac{1}{2}$ or $-\frac{1}{2}$. These are called the two spin states of the electron and are normally represented by two arrows, $\uparrow$ (spin up) and $\downarrow$ (spin down). Two electrons that have different m_s values (one $+\frac{1}{2}$ and the other $-\frac{1}{2}$) are said to have opposite spins.

An orbital cannot hold more than two electrons and these two electrons should have opposite spins.

To sum up, the four quantum numbers provide the following information:

- n defines the shell, determines the size of the orbital and also to a large extent the energy of the orbital.
- There are n subshells in the n th shell. l identifies the subshell and determines the shape of the orbital (see section 2.6.2). There are $(2l+1)$ orbitals of each type in a subshell, that is, one's orbital ($l = 0$), three p orbitals ($l = 1$) and five d orbitals ($l = 2$) per subshell. To some extent l also determines the energy of the orbital in a multi-electron atom.

- iii) m_l designates the orientation of the orbital. For a given value of l , m_l has $(2l+1)$ values
- iv) the same as the number of orbitals per subshell. It means that the number of orbitals is equal to the number of ways in which they are oriented. iv) m_s refers to orientation of the spin of the electron

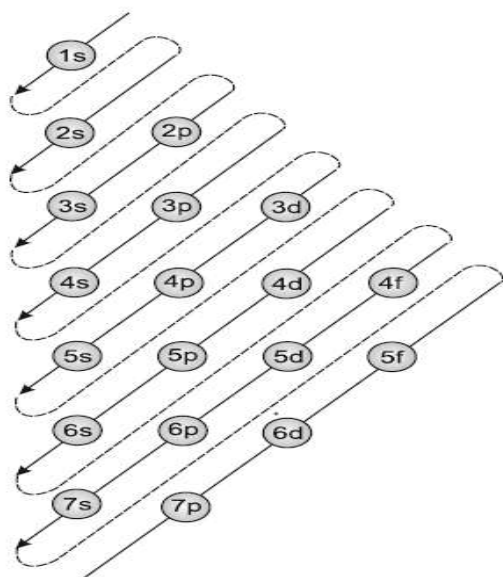
Electronic Configuration of Atoms

The distribution of electrons into orbitals of an atom is called its electronic configuration. If one keeps in mind the basic rules which govern the filling of different atomic orbitals.

Aufbau Principle

The word 'aufbau' in German means 'building up'. The building up of orbitals means the Arrangement of Orbitals with Increasing Energy on the Basis of $(n+l)$ Rule 2022-23 62 CHEMISTRY filling up of orbitals with electrons. The principle states: In the ground state of the atoms, the orbitals are filled in order of their increasing energies. In other words, electrons first occupy the lowest energy orbital available to them and enter into higher energy orbitals only after the lower energy orbitals are filled. As you have learnt above, energy of a given orbital depends upon effective nuclear charge and different type of orbitals are affected to different extent. Thus, there is no single ordering of energies of orbitals which will be universally correct for all atoms. However, following order of energies of the orbitals is extremely useful: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 4f, 5d, 6p, 7s...

The order may be remembered by using the method given in Fig.



Order of filling of orbitals

Pauli Exclusion Principle

The number of electrons to be filled in various orbitals is restricted by the exclusion principle, given by the Austrian scientist Wolfgang Pauli (1926). **According to this principle:** No two electrons in an atom can have the same set of four quantum numbers. Pauli Exclusion Principle can also be stated as: "Only two electrons may exist in the same orbital and these electrons must have opposite spin." This means that the two electrons can have the same value of three quantum numbers n , l and m_l , but must have the opposite spin quantum number. The restriction imposed by Pauli's exclusion principle on the number of electrons in an orbital helps in calculating the capacity of electrons to be present in any subshell. For example, subshell 1s comprises one orbital and thus the maximum number of electrons present in 1s subshell can be two, in p and d subshells, the maximum number of electrons can be 6 and 10 and so on. This can be summed up as:

The maximum number of electrons in the shell with principal quantum number n is equal to $2n^2$.

Hund's Rule of Maximum Multiplicity

This rule deals with the filling of electrons into the orbitals belonging to the same subshell. Order of filling of orbitals (that is, orbitals of equal energy, called degenerate orbitals).

It states: pairing of electrons in the orbitals belonging to the same subshell (p, d or f) does not take place until each orbital belonging to that subshell has got one electron each i.e., it is singly occupied.

Since there are three p, five d and seven f orbitals, therefore, the pairing of electrons will start in the p, d and f orbitals with the entry of 4th, 6th and 8th electron, respectively.

Electronic Configuration

Atomic Number	Symbol	Electronic Configuration
3	Li	$1s^2 2s^1$ (or) $[He] 2s^1$
11	Na	$1s^2 2s^2 2p^6 3s^1$ (or) $[Ne] 3s^1$
19	K	$1s^2 2s^2 2p^6 3s^2 3p^4 s^1$ (or) $[Ar] 4s^1$
37	Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5 5s^1$ (or) $[Kr] 5s^1$
55	Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^1$ (or) $[Xe] 6s^1$
87	Fr	$[Rn] 7s^1$

Elements are classified into - metals – Non-metals - Metalloids

The periodic Table classifies the elements in several ways. You have already learned about the s-, p-, d- and f- block elements. In this article, we will discuss another broad classification of elements – metals and non-metals. The classification is largely based on the chemical properties which are common among the divisions. We will also talk about the third category of elements, called the metalloids.

Metals

- 78% of known elements are metals.
- They take up the left side of the Periodic Table.
- At room temperature, they are usually solids (except mercury).
- They usually have high boiling and melting points.
- They are shiny having a metallic lustre.
- They are good conductors of heat and electricity.
- They are malleable, that is, they can be hammered into thin sheets.
- They are ductile, that is, they can be drawn into wires.
- All the metals tend to lose electrons easily.

Non-Metals

- They occupy the right-hand side of the Periodic Table.
- Most non-metals are brittle and are not malleable or ductile.
- At room temperature, they are usually solids or gases.
- They usually have low boiling and melting points (except for carbon and boron).
- They are poor conductors of heat and electricity.
- They are usually less dense when compared to metals.
- They tend to gain electrons during chemical reactions.

Metalloids.

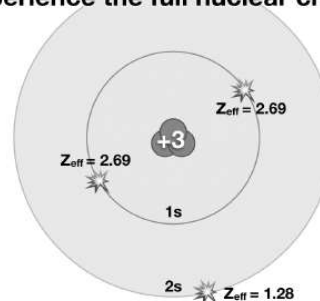
- Metalloids exhibit some properties of metals as well as non-metals. Hence, they are also known as semi-metals.
- Boron, silicon, germanium, arsenic, antimony, and tellurium are the most commonly recognized metalloids.
- These elements run diagonally across the Periodic Table.
- They may have a metallic or dull appearance.

- They are usually brittle and are fair conductors of electricity.
- They may gain or lose electrons during chemical reactions.
- They form alloys with metals.
- They are used as catalysts, biological agents, flame retardants, alloys and semiconductors in industries due to their unique properties

PERIODICITY IN PROPERTIES:

EFFECTIVE NUCLEAR CHARGE (Z_{eff})

Electrons in multi electron atoms **do not** experience the full nuclear charge!



Effective nuclear charge (Z_{eff}) is the charge that each electron experiences.

Effective nuclear charge is the attractive positive charge of nuclear protons acting on valence electrons. The effective nuclear charge is always less than the total number of protons present in a nucleus due to shielding effect.

Effective nuclear charge is behind all other periodic table tendencies

The effective nuclear charge may be approximated by the equation:

$$Z_{\text{eff}} = Z - S$$

Where Z is the atomic number and S is the number of shielding electrons

Trends

The periodic table tendency for effective nuclear charge:

- Increase across a period (due to increasing nuclear charge with no accompanying increase in shielding effect).
- Decrease down a group (although nuclear charge increases down a group, shielding effect more than counters its effect).

For example,

In lithium (Li), none of the three electrons "feel" the full +3 charge from the nucleus. Rather, each electron "feels" a Z_{eff} that is less than the actual Z and that depends on the electron's orbital. The actual nuclear charge in Li is +3;

the 1s electrons experience a $Z_{\text{eff}} = +2.69$, and the 2s electron experiences a $Z_{\text{eff}} = 1.28$. In general, core electrons (or the electrons closest to the nucleus), "feel" a Z_{eff} that is close to, but less than, the actual nuclear charge (Z). On the other hand, outer valence electrons experience a Z_{eff} that is much less than Z .

Atomic Size

Atomic size is the distance between the centre of the nucleus of an atom and its outermost shell.

Factors affecting atomic size

- 1) Principal quantum number increases (n) atomic size increases.
- 2) Z_{eff} increase, size decreases.

PERIODIC TRENDS:

- In a period from left to right atomic size decreases.
- The number of protons are increasing from left to right with increase in atomic number in a period.
- When effective nuclear charge is increasing the force of attraction on outer most electron increasing, hence the result net contraction of atomic size.

For Example,

Atomic order of 1 group: $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs} < \text{Fr}$
 2 group: $\text{Be} < \text{Mg} < \text{Ca} < \text{Sr} < \text{Ba} < \text{Ra}$
 3 group.....etc.

Atomic Radii

Atomic radius is generally stated as being the total distance from an atom's nucleus to the outermost orbital of electron. It can be defined as something similar to the radius of a circle, where the center of the circle is the nucleus and the outer edge of the circle is the outermost orbital of electron.

Types of Atomic Radius

Based on the type of bond, atomic radius is divided into four types as follows:

- Covalent Radius
- Ionic Radius
- Metallic Radius
- Van Der Waals Radius

Covalent Radius

The covalent radius of an atom is the radius of an atom under the covalent bond with another atom(s) of a similar element. The covalent radius of an atom can be determined by measuring bond lengths between pairs of

covalently-bonded atoms. If the two atoms are the same kind, the covalent radius is simply one-half of the bond length. Whilst this is straightforward for some molecules such as Cl_2 and O_2 , in other cases, one has to infer the covalent radius by measuring bond distances to atoms whose radii are already known (e.g., a C–X bond, in which the radius of C is known).

Ionic Radius

The ionic radius is the radius of an atom forming an ionic bond or an ion. The atomic bonds restrict the electrons and nucleus, and due to this, the ions or atoms don't have a specific shape. The measuring unit for the ionic radius is Armstrong ($\text{\AA}$) or picometers (pm). The characteristic radius ranges from 30 to 200 pm. The ionic radius is not static but differs with respect to the spin state of the electrons, coordination number and numerous other parameters. Ionic size grows with increasing coordination number. The ionic size also grows for an ion with a high spin state of an electron than an ion with a low spin state. The ion with a positive charge will have a smaller size than the ion with a negative charge if we consider the charge of the ion.

Metallic Radius

The metallic radius is the radius of an atom joined by the metallic bond. The metallic radius is half of the total distance between the nuclei of two adjacent atoms in a metallic cluster.

Van-der-Waals radii:-

Van-der-Waals radii are determined from the contact distances between unbonded atoms in touching molecules or atoms.

Size of isoelectronic ion:

Isoelectronic ions have the same numbers of electrons Al^{+3} , Mg^{+2} , Na^+ , Ne , F^- , O^{2-} and N^{3-} all have 10 electrons. All have same electronic configuration: $1s^2 2s^2 2p^6$

Ionization Energy

In more technical terms we can describe ionization energy as the minimum energy that an electron in a gaseous atom or ion has to absorb to come out of the influence of the nucleus. It is also sometimes referred to as ionization potential and is usually an endothermic process.

What we can deduce further is that ionization energy gives us an idea of the reactivity of chemical compounds. It can also be used to determine the strength of chemical bonds. It is measured either in units of electronvolts or kJ/mol.

Depending on the ionization of molecules which often leads to changes in molecular geometry, ionization energy can be either adiabatic ionization energy or vertical ionization energy.

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Factors Governing Ionization Energy

Normally, when the ionization energy is high it will be more difficult to remove an electron. There are also several factors that govern the attraction forces.

- If the nucleus is positively charged then the electrons are strongly attracted to it.
- If an electron lies near or close to the nucleus then the attraction will be greater than the one when the electron is further away.
- If there are more electrons between the outer level and the nucleus the attraction forces are less.
- When there are two electrons in the same orbital they experience some form of repulsion. Now, this creates disturbances in the attraction of the nucleus. In essence, ionization energy will be less in paired electrons as they can be removed easily

Ionization Energy and Bohr's Atomic Model

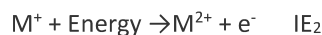
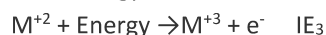
Atomic ionization energy can be further be predicted using Bohr's model of an atom. His model predicts the presence of several paths for the electron to go around the nucleus containing protons and neutrons. Each path or orbit is at a fixed distance from the nucleus. Each orbit also represents fixed energy. Electron is a particle and will have the energy of the orbit present. A particle can absorb energy and jump to the next higher orbits of higher energy. If more energy is available and absorbed, the electron will come out of force of attraction of the nucleus, which means out of the atom.

Ionization Energy

Ionization is a process that involves the removal of electron present in an orbit to outside the atom. As the electron in each orbit has characteristic energy, ionization energy is equal to the difference of energy between the energy of the electron in the initial orbit and the energy of the electron outside the atom (in the infinite orbit from the nucleus).

Amount of energy required to remove outermost electrons from an isolated gaseous atom.

Ionisation is endothermic i.e requires energy hence ΔH is positive



$$IE_3 > IE_2 > IE_1$$

Ionisation energy is also called as ionisation potential.

Unit: Ionization is express in the units of energy, i.e (KJ mol⁻¹) or kcal mol⁻¹.

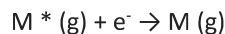
$$1 \text{ eV / atom} = 1.602 \times 10^{-16} \text{ kJ / atom} = 1.602 \times 10^{-16} \times 6.02 \times 10^{23} \text{ kJ / mol}$$

$$= 96.4 \times 10^6 \text{ kJ/mol}^{-1}$$



(ground state)

Ionization energy of M = x kJ mol⁻¹.



Electron affinity of M⁺ = x kJ mol⁻¹.

$$I.E = \frac{13.6 \cdot Z^{*2}}{n^{*2}} \text{ eV / atom}$$

Thus, ionization energy of an atom is equal to the electron affinity of its unipositive cation in magnitude.

$$I.E = 13.6 \times Z^{*2} \cdot 2 \text{ AH} = +x \text{ kJ mol}^{-1} - \text{eV / atom} \quad \text{AH} = -x \text{ kJ mol}^{-1} n$$

Since, I.E. (of hydrogen) = 13.6 eV / atom.

FACTOR AFFECTING I.E:

1. I.E. Varies Inversely with Size of Atom

As the size of atom increases. The outer most electrons are less tightly held by the nucleus, As a result it is easier to remove the e⁻. The ionisation energy decreases with increase in atomic size.

2. I.E. Varies Directly with Nuclear Charge

As nuclear charge increases, it becomes difficult to remove the e⁻, so ionisation energy increases.

3. I.E. Varies Inversely with Screening Effect

The e's of inner shell repel the e's of outer most shells. This effect is known as screening or shielding effect.

The e's of outer most shell are thus shielded or screened from the nucleus by inner e⁻ s. As a result of screening effect the outer most e's do not experience the complete nuclear charge.

So, an increase in no. of inner e's increases screening effect and an increase in screening effect decreases nuclear charge. This results in decrease of I.E.