Dev Academy
Yamunanagar (Haryana)

Unit - 3

Classification of Elements and Periodicity in Properties

Class
11

Revision Notes

CHEMISTRY

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Part – 1

1.1 Introduction

- Currently, a total of 118 elements have been discovered or synthesized.
- The initial 94 elements occur naturally, while elements with atomic numbers 95 to 118 have been synthesized in laboratories or nuclear reactors.
- With such large number of elements, it is quite difficult to understand individually the chemistry of all the elements.
- To address this challenge, scientists systematically arrange all known elements based on their properties.
- Elements with similar properties are placed in same vertical groups, results in formation of periodic table.
- The **periodic table** organizes all known elements into systematic groups according to their physical and chemical characteristics.

1.2 Genesis of Periodic Classification

1.2.1 Döbereiner's Triads:

- In 1817, Johann Wolfgang Doebereiner, a German chemist, made the first attempt towards the classification of elements.
- He organized the known elements into sets of three, where the atomic weight of the middle element is approximately the arithmetic mean of the atomic weights of the first and third elements. The properties of the central element are also intermediate between those of the other two.

Triad	Atomic Weight	Arithmetic mean
Lithium	7	$\frac{7+39}{2}$ = 23
Sodium	23	
Potassium	39	
Calcium	40	$\frac{40+137}{2}$ = 88.5
Strontium	88	
Barium	137	
Sulfur	32	$\frac{32+128}{2}$ = 80
Selenium	79	
Tellurium	128	
Chlorine	35.5	$\frac{35.5+127}{2}$ = 81.25
Bromine	80	
Iodine	127	

1.2.2 Newland’s Octaves:

- In 1865, the English chemist John Alexander Newlands proposed the law known as Newland’s Law of Octaves.
- According to this law, when elements are arranged in order of their increasing atomic weights, every eighth element shares similar properties with the first one, akin to the resemblance between musical notes.
- However, this law was applicable solely to lighter elements with atomic weights up to 40 u.

Octaves →	**Sa** **1**	**Re** **2**	**Ga** **3**	**Ma** **4**	**Pa** **5**	**Dha** **6**	**Nee** **7**
Elements → **(Atomic Wt.)**	Li (7)	Be (9)	B (11)	C (12)	N (14)	O (16)	F (19)
Elements → **(Atomic Wt.)**	Na (23)	Mg (24)	Al (27)	Si (28)	P (31)	S (32)	Cl (35.5)
Elements → **(Atomic Wt.)**	K (39)	Ca (40)					

1.2.3 Mendeleev’s Periodic Table:

- In 1869, the Russian chemist Dmitri Mendeleev introduced the periodic table, which was developed based on his concept known as Mendeleev Periodic Law.
- According to Mendeleev's Periodic Law, the physical and chemical properties of elements are the periodic function of their atomic weights.
- This law can be summarized as follows: "When elements are arranged in ascending order of their atomic weights, elements with similar properties are repeated after regular intervals."
- The repetition of elemental properties after specific and consistent intervals is termed the periodicity of properties.

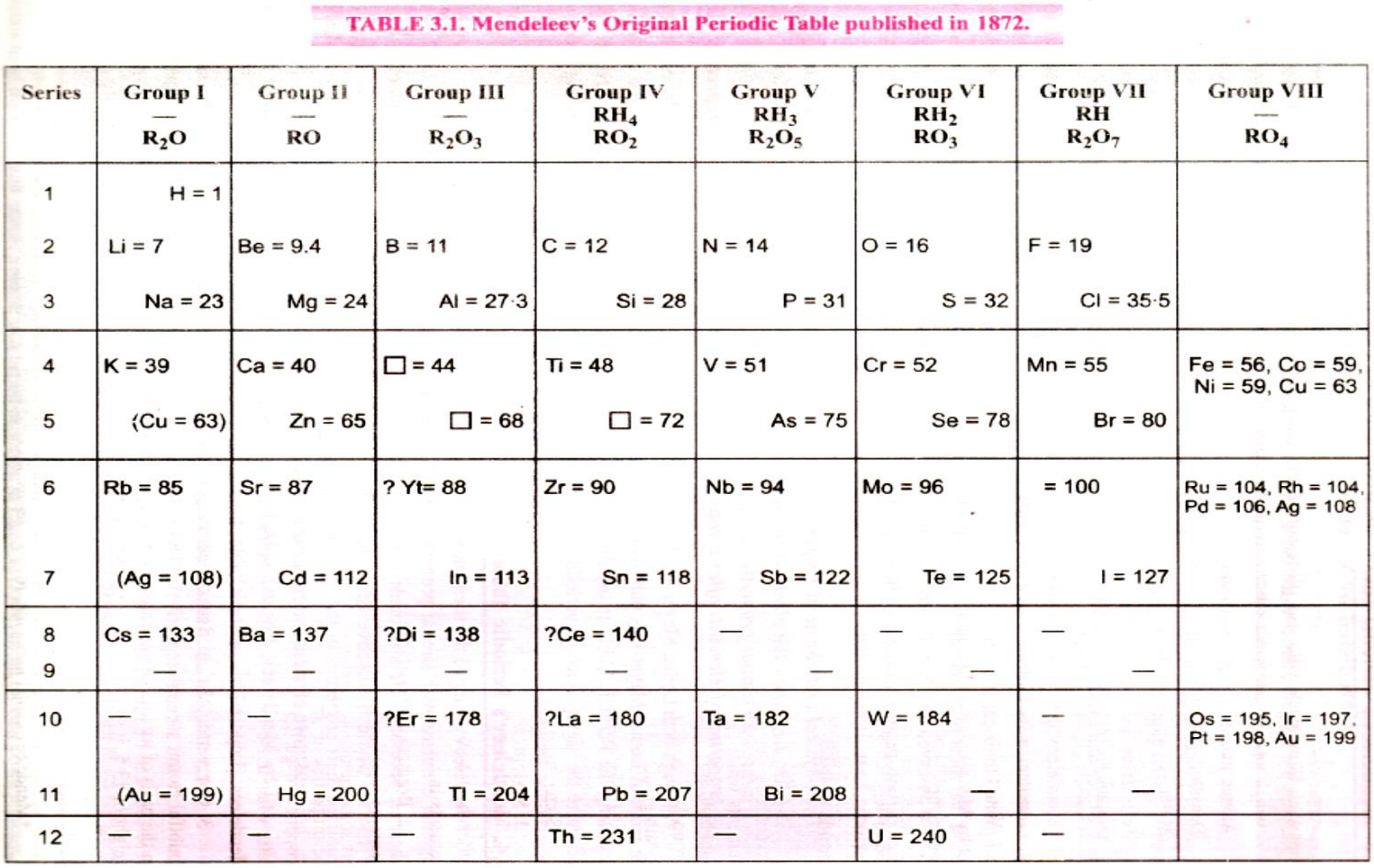


TABLE 3.1. Mendeleev's Original Periodic Table published in 1872.

Series	Group I — R_2O	Group II — RO	Group III — R_2O_3	Group IV RH_4 RO_2	Group V RH_3 R_2O_5	Group VI RH_2 RO_3	Group VII RH R_2O_7	Group VIII — RO_4
1	H = 1							
2	Li = 7	Be = 9.4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27·3	Si = 28	P = 31	S = 32	Cl = 35·5	
4	K = 39	Ca = 40	☐ = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63
5	(Cu = 63)	Zn = 65	☐ = 68	☐ = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	? Yt= 88	Zr = 90	Nb = 94	Mo = 96	= 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	I = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	—	—	—	
9	—	—	—	—	—	—	—	
10	—	—	?Er = 178	?La = 180	Ta = 182	W = 184	—	Os = 195, Ir = 197, Pt = 198, Au = 199
11	(Au = 199)	Hg = 200	Tl = 204	Pb = 207	Bi = 208	—	—	
12	—	—	—	Th = 231	—	U = 240	—	

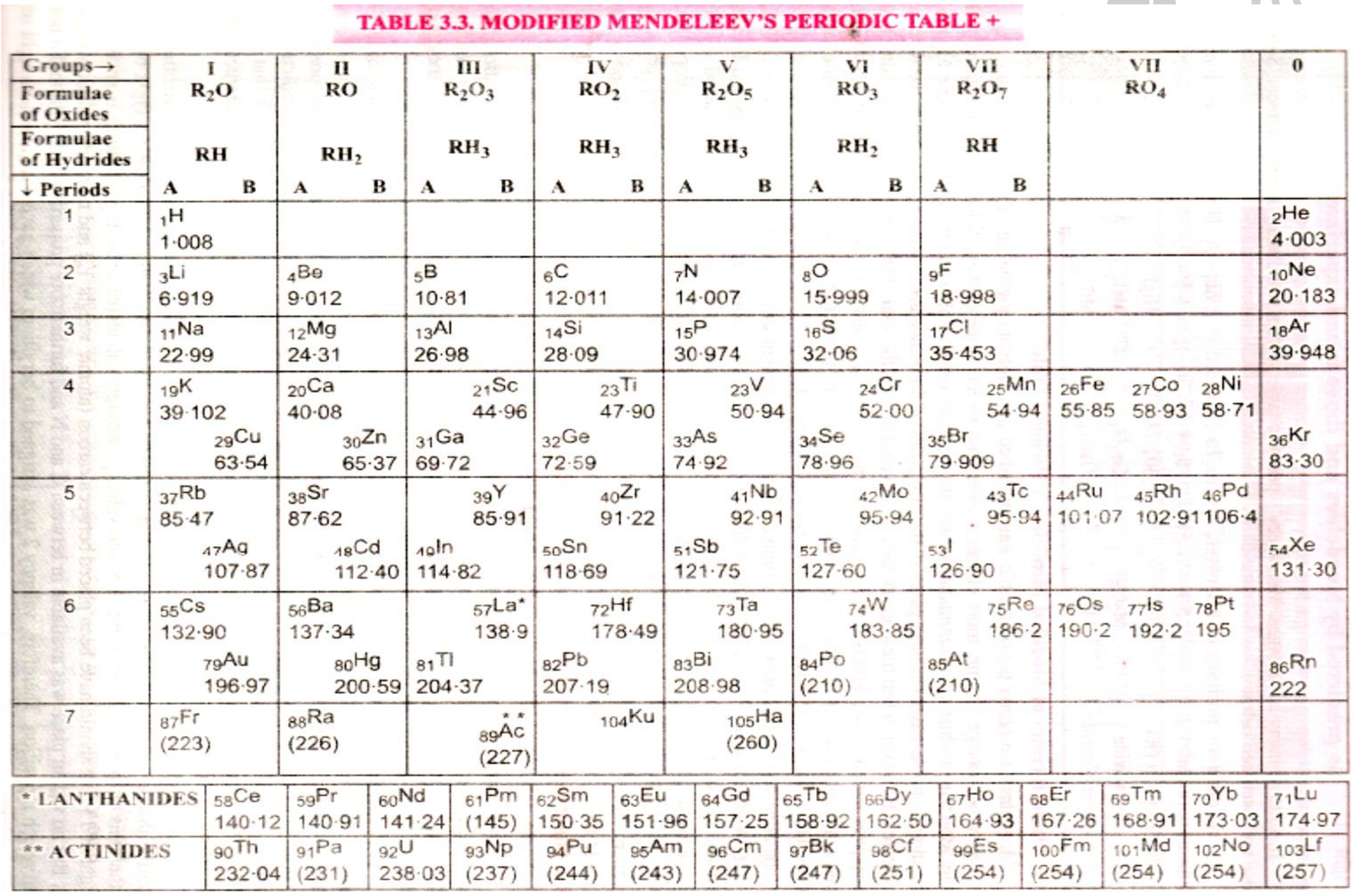


TABLE 3.3. MODIFIED MENDELEEV'S PERIODIC TABLE +

Groups→	I		II		III		IV		V		VI		VII		VII	0
Formulae of Oxides	R_2O		RO		R_2O_3		RO_2		R_2O_5		RO_3		R_2O_7		RO_4	
Formulae of Hydrides	RH		RH_2		RH_3		RH_3		RH_3		RH_2		RH			
↓ Periods	A	B	A	B	A	B	A	B	A	B	A	B	A	B		
1	$_1$H 1·008															$_2$He 4·003
2	$_3$Li 6·919		$_4$Be 9·012		$_5$B 10·81		$_6$C 12·011		$_7$N 14·007		$_8$O 15·999		$_9$F 18·998			$_{10}$Ne 20·183
3	$_{11}$Na 22·99		$_{12}$Mg 24·31		$_{13}$Al 26·98		$_{14}$Si 28·09		$_{15}$P 30·974		$_{16}$S 32·06		$_{17}$Cl 35·453			$_{18}$Ar 39·948
4	$_{19}$K 39·102	$_{29}$Cu 63·54	$_{20}$Ca 40·08	$_{30}$Zn 65·37	$_{31}$Ga 69·72	$_{21}$Sc 44·96	$_{32}$Ge 72·59	$_{23}$Ti 47·90	$_{33}$As 74·92	$_{23}$V 50·94	$_{34}$Se 78·96	$_{24}$Cr 52·00	$_{35}$Br 79·909	$_{25}$Mn 54·94	$_{26}$Fe 55·85 $_{27}$Co 58·93 $_{28}$Ni 58·71	$_{36}$Kr 83·30
5	$_{37}$Rb 85·47	$_{47}$Ag 107·87	$_{38}$Sr 87·62	$_{48}$Cd 112·40	$_{49}$In 114·82	$_{39}$Y 85·91	$_{50}$Sn 118·69	$_{40}$Zr 91·22	$_{51}$Sb 121·75	$_{41}$Nb 92·91	$_{52}$Te 127·60	$_{42}$Mo 95·94	$_{53}$I 126·90	$_{43}$Tc 95·94	$_{44}$Ru 101·07 $_{45}$Rh 102·91 $_{46}$Pd 106·4	$_{54}$Xe 131·30
6	$_{55}$Cs 132·90	$_{79}$Au 196·97	$_{56}$Ba 137·34	$_{80}$Hg 200·59	$_{81}$Tl 204·37	$_{57}$La* 138·9	$_{82}$Pb 207·19	$_{72}$Hf 178·49	$_{83}$Bi 208·98	$_{73}$Ta 180·95	$_{84}$Po (210)	$_{74}$W 183·85	$_{85}$At (210)	$_{75}$Re 186·2	$_{76}$Os 190·2 $_{77}$Is 192·2 $_{78}$Pt 195	$_{86}$Rn 222
7	$_{87}$Fr (223)		$_{88}$Ra (226)			$_{89}$Ac** (227)		$_{104}$Ku		$_{105}$Ha (260)						

* LANTHANIDES	$_{58}$Ce 140·12	$_{59}$Pr 140·91	$_{60}$Nd 141·24	$_{61}$Pm (145)	$_{62}$Sm 150·35	$_{63}$Eu 151·96	$_{64}$Gd 157·25	$_{65}$Tb 158·92	$_{66}$Dy 162·50	$_{67}$Ho 164·93	$_{68}$Er 167·26	$_{69}$Tm 168·91	$_{70}$Yb 173·03	$_{71}$Lu 174·97
** ACTINIDES	$_{90}$Th 232·04	$_{91}$Pa (231)	$_{92}$U 238·03	$_{93}$Np (237)	$_{94}$Pu (244)	$_{95}$Am (243)	$_{96}$Cm (247)	$_{97}$Bk (247)	$_{98}$Cf (251)	$_{99}$Es (254)	$_{100}$Fm (254)	$_{101}$Md (254)	$_{102}$No (254)	$_{103}$Lf (257)

1.2.3.1 Characteristics of Modified Mendeleev’s Periodic Table

- Incorporating New Elements:
    - Mendeleev initially arranged the 63 known elements in his periodic table.
    - With the discovery of more elements, adjustments were needed to accommodate the new entries.
    - The modified form of Mendeleev's Periodic Table emerged as a result of these adjustments.
- Periods and Groups:
    - Periods: The table features seven horizontal rows, known as periods. Mendeleev referred to horizontal rows as "series."
    - Groups: Vertical columns in the table are called groups. Groups I to VIII and a zero group (absent during Mendeleev's time) are present.
- Sub-Groups A and B:
    - Except for the zero group and group VIII, all groups are divided into sub-groups A and B.
    - Elements of sub-group A are positioned on the left, while sub-group B elements are on the right.
- Triads and Noble Gases:
    - Group VIII: Contains nine elements forming three triads.
    - Zero Group: Comprises noble gases or inert gases.

1.2.3.2 Advantages and Significance of Modified Mendeleev’s Periodic Table

- Systematic Study of Elements:
    - All elements systematically arranged by increasing atomic weights.
    - It simplifies the study of elements and their compounds.
    - Predicting properties of a group becomes feasible by understanding constituent elements' nature.
- Prediction of New Elements:
    - Original Mendeleev's Table published with 63 known elements.
    - Mendeleev predicted properties of unknown elements, leaving gaps for their discovery. Example: Gallium (Eka-aluminum) and Germanium (Eka-silicon) named and predicted by Mendeleev.
    - Gallium and Germanium properties matched Mendeleev's predictions upon discovery.

Property	Eka-Al (Predicted)	Gallium (Found)	Eka-Si (Predicted)	Germanium (Found)
Atomic Mass	68	70	72	72.6
Density	5.9	5.94	5.5	5.36
Oxide Formula	M_2O_3	Ga_2O_3	MO_2	GeO_2
Chloride Formula	MCl_3	$GaCl_3$	MCl_4	$GeCl_4$

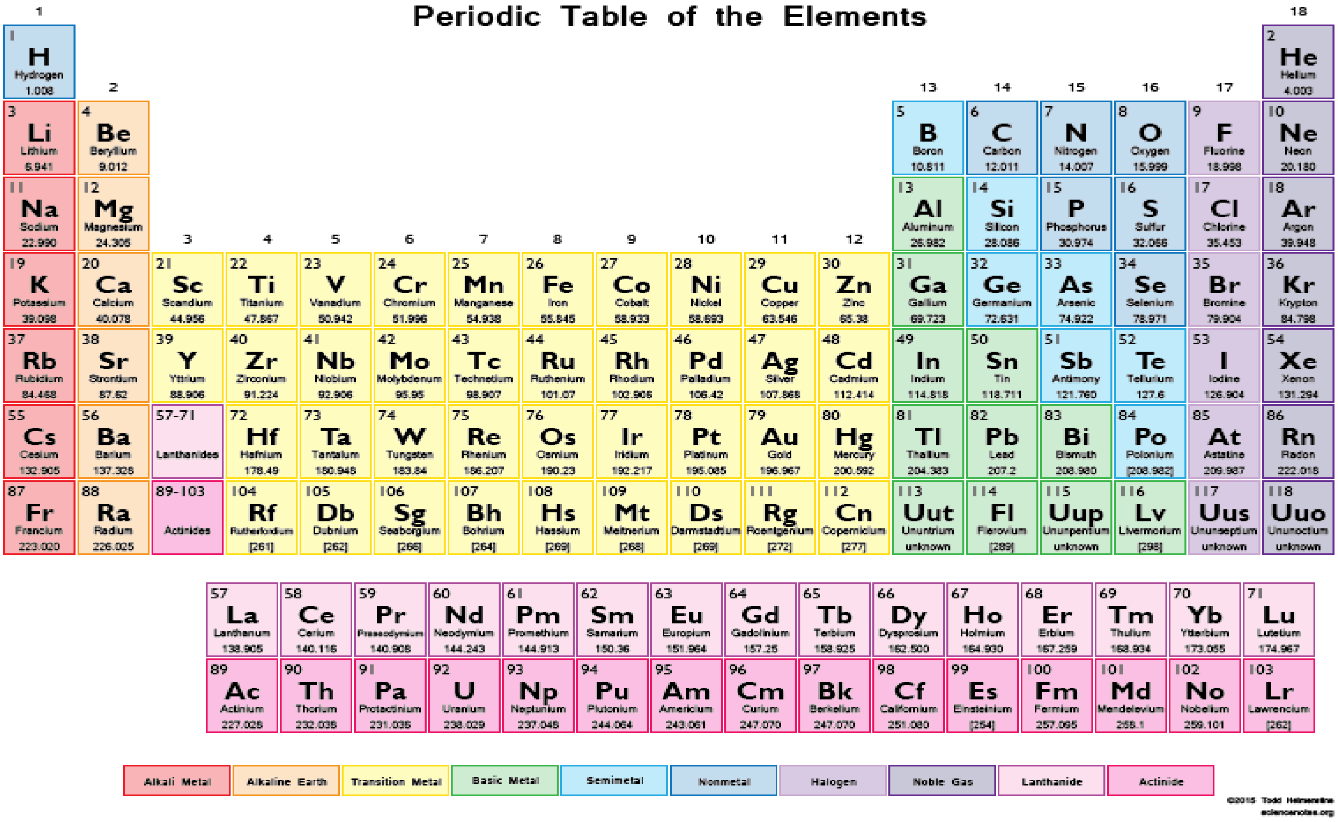
- Correction of Atomic Weights:
  - Mendeleev's table played in correcting the doubtful atomic weights of some elements.
  - Example: Atomic mass of Beryllium (Be) was corrected from 13.5 to 9.
    - The equivalent weight of Be is 4.5.
    - Initially, Beryllium was believed to be trivalent, with an atomic mass of 13.5 (calculated as 4.5 × 3 = 13.5). This conflicted with its proper placement between C and N, and its properties did not match its position.
    - By reassigning a valency of 2 (bivalent) for Beryllium, its atomic weight became 9 (calculated as 4.5 × 2 = 9). This adjustment positioned Beryllium correctly between Li and B.

1.2.3.3 Drawbacks of Modified Mendeleev's Periodic Table

- Anomalous Position of Hydrogen:
  - Hydrogen placed in group IA but shares similarities with both group IA (alkali metals) and group VII A (halogens).
  - Hydrogen's placement is not fixed; it is considered anomalous or controversial.
- Position of Isotopes:
  - Isotopes are atoms of the same element with different atomic weights.
  - Ideally, isotopes should have separate positions in the table.
  - However, no separate places are given to the isotopes.
- Anomalous Pairs of Certain Elements:
  - There are certain pairs of elements which have not been arranged according to their increasing atomic weights.
  - Example: Ar (39.9 u) is placed before of K (39 u).

- Lack of Explanation for Periodicity:
    - Mendeleev did not provide an explanation for the cause of periodicity in elements' physical and chemical properties.
- Mixing of Metals and Non-Metals:
    - Metals and non-metals not segregated in Mendeleev's Periodic Table.

1.2.4 Modern's Periodic Table or Bohr's Table:



Periodic Table of the Elements

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H Hydrogen 1.008																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.066	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.631	33 As Arsenic 74.922	34 Se Selenium 78.971	35 Br Bromine 79.904	36 Kr Krypton 84.798
37 Rb Rubidium 84.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium 98.907	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.414	49 In Indium 114.818	50 Sn Tin 118.711	51 Sb Antimony 121.760	52 Te Tellurium 127.6	53 I Iodine 126.904	54 Xe Xenon 131.294
55 Cs Cesium 132.905	56 Ba Barium 137.328	57-71 Lanthanides	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 [208.982]	85 At Astatine 209.987	86 Rn Radon 222.018
87 Fr Francium 223.020	88 Ra Radium 226.025	89-103 Actinides	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [269]	109 Mt Meitnerium [268]	110 Ds Darmstadtium [269]	111 Rg Roentgenium [272]	112 Cn Copernicium [277]	113 Uut Ununtrium unknown	114 Fl Flerovium [289]	115 Uup Ununpentium unknown	116 Lv Livermorium [298]	117 Uus Ununseptium unknown	118 Uuo Ununoctium unknown

57 La Lanthanum 138.905	58 Ce Cerium 140.116	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 237.048	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]

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1.2.4.1 Atomic Number and Modern Periodic Law

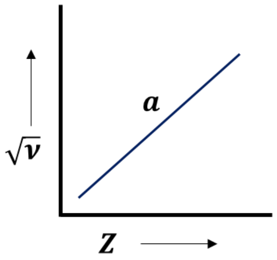
- In 1912, an English physicist named Henry Moseley demonstrated that atomic number is a fundamental property predicting elements' physical and chemical behavior.
- Moseley proposed the Modern Periodic Law based on an experiment involving high-speed electron bombardment of metals.

- He observed X-ray frequencies emitted by 38 elements (from Al to Au) when bombarded by high-speed electrons.
- He found that the square root of X-ray frequency ($\sqrt{\nu}$) was proportional to the element's atomic number (Z), not its atomic weight.

  This relationship was expressed as;

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

  Where, 'a' as the proportionality constant and 'b' as a constant for X-ray lines in a series.



- Graphing $\sqrt{\nu}$ against Z yielded a straight line, leading Moseley to conclude atomic number was a more fundamental property than atomic weight.
- **The Modern Periodic Law:**

  The physical and chemical properties of elements are the periodic function of their atomic numbers.

  This law can be summarized as follows: "When elements are arranged in ascending order of their atomic numbers, elements with similar properties are repeated after regular intervals."

1.2.4.2 Theoretical justification of Modern Periodic Law:

- An element's chemical properties depend on how its atoms interact with other reactants.
- These interactions involve electrons present in valence shell of an atom, not just the nucleus.
- But the number of electrons in atom depends upon the atomic number of an element. So, the physical and chemical properties of an element are the periodic functions of its atomic number.

1.2.4.3 Cause of Periodicity

- Periodicity in properties arises from the recurrence of similar outer electron configurations at regular intervals.
- These intervals are 2, 8, 8, 18, 18, 32, and 32.
- These numbers are occasionally referred to as "magic numbers."
- Illustration with example;

Elements (1st group)	Atomic number	Electronic configuration
Li	3	[He] $2s^1$
Na	11	[Ne] $3s^1$
K	19	[Ar] $4s^1$
Rb	37	[Kr] $5s^1$
Cs	55	[Xe] $6s^1$
Fr	87	[Rn] $7s^1$

Elements (14 group)	Atomic number	Electronic configuration
C	6	[He] $2s^2$ $2p^2$
Si	14	[Ne] $3s^2$ $3p^2$
Ge	32	[Ar] $4s^2$ $4p^2$
Sn	50	[Kr] $5s^2$ $5p^2$
Pb	82	[Xe] $6s^2$ $6p^2$

Elements (16 group)	Atomic number	Electronic configuration
O	8	[He] $2s^2$ $2p^4$
S	16	[Ne] $3s^2$ $3p^4$
Se	34	[Ar] $4s^2$ $4p^4$
Te	52	[Kr] $5s^2$ $5p^4$
Po	84	[Xe] $6s^2$ $6p^4$

Elements (17 group)	Atomic number	Electronic configuration
F	9	[He] $2s^2$ $2p^5$
Cl	17	[Ne] $3s^2$ $3p^5$
Br	35	[Ar] $4s^2$ $4p^5$
I	53	[Kr] $5s^2$ $5p^5$
At	85	[Xe] $6s^2$ $6p^5$

1.2.4.4 Important Features of Long Form of the Periodic Table:

- Groups:
  - Modern Periodic Table consists of 18 vertical columns known as groups.
  - Elements within the same group share a similar outer electronic configuration.
  - Groups are numbered from 1 to 18.
- Periods:
  - Modern Periodic Table consists of 7 horizontal rows known as periods.
  - Periods are numbered from 1 to 7, from top to bottom.
  - The period number corresponds to an element's highest principal quantum number (n).
- Lanthanoids and Actinoids:
  - In the long-form Periodic Table, the 6th and 7th periods include 14 elements each.
  - These elements are placed in separate panels at the bottom of the table.
  - The 6th period's panel is known as Lanthanoids, while the 7th period's panel is referred to as Actinoids.
  - These two panels together are termed f-block elements.

1.2.4.5 IUPAC Nomenclature of Elements with Atomic Number more than 100:

- Names derived directly from atomic numbers using numerical roots for 0 and numbers from 1 to 9, followed by the suffix 'ium'.
- In specific cases, names are shortened: e.g., bi + ium becomes "bium," tri + ium becomes "trium," and enn + nil becomes "ennil."
- The element's symbol is formed from the initial letters of the numerical roots corresponding to the digits composing the atomic number.

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

Atomic Number	Recommended name	Symbol
101	Unnilunium	Unu
102	Unnilbium	Unb
103	Unniltrium	Unt
104	Unnilquadium	Unq
105	Unnilpentium	Unp
106	Unnilhexium	Unh
107	Unnilseptium	Uns
108	Unniloctium	Uno
109	Unnilennium	Une
110	Ununnilium	Uun
111	Unununium	Uuu
112	Ununbium	Uub
113	Ununtrium	Uut
114	Ununquadium	Uuq
115	Ununpentium	Uup
116	Ununhexium	Uuh
117	Ununseptium	Uus
118	Ununoctium	Uuo
119	Ununennium	Uue

1.2.4.6 Number of elements in different Periods:

- 1st Period:
    - Involves the filling of 1s orbital.
    - Contains only **2** elements.
- 2nd Period:
    - Involves the filling of 2s and 2p orbitals.
    - Contains **8** elements.
- 3rd Period:
    - Involves the filling of 3s and 3p orbitals.
    - Contains **8** elements.
- 4th Period:
    - Involves the filling of 4s, 3d, and 4p orbitals.
    - Contains **18** elements.
- 5th Period:
    - Involves the filling of 5s, 4d, and 5p orbitals.
    - Contains **18** elements.
- 6th Period:
    - Involves the filling of 6s, 4f, 5d, and 6p orbitals.
    - Contains **32** elements.
- 7th Period:
    - Involves the filling of 7s, 5f, 6d, and 7p orbitals.
    - Contains **32** elements.

1.2.4.7 Prediction of Period, Group and Block of given Element:

- Period Prediction:
    - The period of an element corresponds to the principal quantum number of its valence shell.
- Block Prediction:
    - The element's block is determined by the type of orbital that receives its last electron.
- Group Prediction:
    - For s-block elements:
        - Group number equals the number of valence electrons.
    - For p-block elements:
        - Group number equals 10 plus the number of electrons in the valence shell.

- For d-block elements:
    - Group number equals the number of electrons in the (n-1)d subshell plus the number of electrons in the valence shell.

1.2.4.8 Advantages of the Long Form of Periodic Table

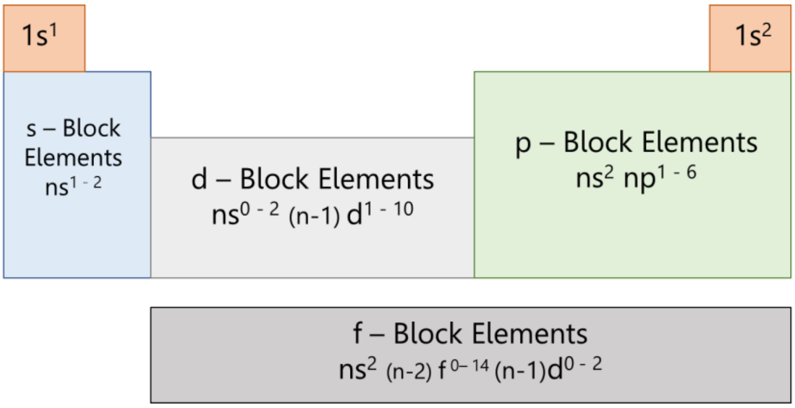
- Simplicity and Memorability:
    - The long-form table is easy to understand and remember.
- Based on Atomic Number:
    - It is structured according to atomic number, not atomic mass.
- Isotope Resolution:
    - Isotope placement issue resolved as the all isotopes have the same atomic number.
- Effective Grouping:
    - Elements organized into s-, p-, d-, and f-blocks, enhancing comprehension of electronic configurations and properties.
- Electronic Configuration Relation:
    - The element's position correlates with its electronic configuration.
- Cause of Periodicity:
    - Elements in the same group show noticeable similarities due to similar outer electron configurations.
- Comprehensive Classification:
    - Distinct classifications for metals, non-metals, metalloids, transition elements, lanthanoids, and actinoids.
- Predictive Potential:
    - Enables predicting elements' properties even before their actual discovery.

1.2.4.9 Defects in the Long Form of the Periodic Table:

- Position of Hydrogen:
    - Hydrogen shows similarities to both group 1 (alkali metals) and group 17 (halogens).
    - Due to this dual resemblance, the positioning of hydrogen in the periodic table is anomalous.
- Lack of Lanthanoid and Actinoid Placement:
    - Lanthanoids and actinoids are not suitably accommodated within the periodic table.

- Helium's Placement:
  - Helium, with an electronic configuration of $1s^2$, is positioned among p-block elements in group 18.

1.3 Division of Elements into s-, p-, d-, and f-Blocks



1.3.1 s-Block Elements:

- Elements in which the last electron enters the s-orbital of the valence shell are termed s-Block elements.
- Since, s-subshell has only one orbital which can accommodate only two electrons. Therefore, there are only two groups fall under s-Block:
  - Group 1 (Hydrogen and Alkali Metals)
  - Group 2 (Alkaline Earth Metals)
- General Outer Shell Electronic Configuration of s-Block Elements:

[Noble gas] ns^{1-2}, where n ranges from 2 to 7.

1.3.1.1 General Characteristics of s-Block Elements:

- They are soft metals.
- They have low ionization enthalpies due to their large atomic radii.
- They have low melting and boiling points.
- They have low densities.
- They are highly electropositive elements.
- They exhibit oxidation states of +1 (alkali metals) and +2 (alkaline earth metals).
- They act as strong reducing agents.
- They serve as good conductors of heat and electricity.
- Their compounds are primarily ionic in nature (except for beryllium).
- They give rise to characteristic flame colors (except for Be and Mg).

1.3.2 p-Block Elements:

- These are the elements where the last electron enters the p – orbitals of the valence shell.
- The p-subshell consists of three degenerate p – orbitals, each capable of accommodating two electrons.
- This configuration allows for a total of six electrons to be accommodated, resulting in six groups of p – block elements.
- The groups are numbered 13, 14, 15, 16, 17, and 18 (excluding He).
    - Group 17 elements are known as halogens (salt producing).
    - Group 16 elements are termed chalcogens (ore forming).
    - Group 18 elements are recognized as Noble gases or Inert gases.
- General Outer Shell Electronic Configuration of p-Block Elements:

$$\textbf{[Noble gas] } ns^2\, np^{1-6}$$

1.3.2.1 General Characteristics of p-Block Elements:

- p-Block elements include metals, non-metals, and metalloids.
- They have high ionization enthalpies due to their small atomic size.
- These elements predominantly form covalent compounds.
- Some elements within this group display variable oxidation states in their compounds.
- Their oxidizing character tends to increase from left to right within a period.
- Conversely, their reducing character tends to increase from top to bottom within a group.
- Many p-block elements exhibit semiconducting properties.

1.3.3 d-Block Elements:

- d-Block elements are those in which the last electron enters the d-orbital of the penultimate (n-1) energy level.
- The d-subshell has five d-orbitals, each capable of accommodating two electrons, allowing for a maximum of ten electrons.
- Consequently, there are 10 groups of d-block elements spanning from 3 to 12 on the Periodic Table.
- These elements are situated between the s- and p-blocks. Hence, d-block elements are also known as transition elements.

- General Outer Shell Electronic Configuration of d-Block Elements:

$$\textbf{[Noble Gas] } (n-1)d^{1-10}\, ns^{0-2}$$

1.3.3.1 Transition Series

- The transition elements encompass four transition series: 3d, 4d, 5d, and 6d.
    - First Transition Series (**3d series**): Includes 10 elements from scandium (21=Sc) to zinc (30=Zn), positioned in the fourth period.
    - Second Transition Series (**4d series**): Includes 10 elements from yttrium (39=Y) to cadmium (48=Cd), found in the fifth period.
    - Third Transition Series (**5d series**): Consists of 10 elements, beginning with lanthanum (57=La) and followed by hafnium (72=Hf) to mercury (80=Hg), situated in the sixth period.
    - Fourth Transition Series (**6d series**): Comprises 10 elements, starting with actinium (89=Ac) and extending from rutherfordium (104=Rf) to copernicium (Z=112), positioned in the seventh period.

1.3.3.2 General Characteristics of d-Block Elements:

- They are hard metals (except for Hg, which is a liquid at room temperature).
- They have high melting and boiling points.
- They are good conductors of heat and electricity.
- They show variable oxidation states.
- Their Ionization enthalpies lie between s- and p-block elements.
- They can form both ionic and covalent compounds.
- Their compounds are often colored and paramagnetic.
- They have ability to form colored complexes.
- Many elements, such as Ni, Pt, Pd, Co, Mn, Fe, Cr, are utilized as catalysts.
- Many transition metals readily form alloys.
- Notable exception: Zn, Cd, and Hg do not exhibit several characteristic properties of transition elements.

1.3.4 f-Block Elements:

- f-Block elements are those in which the last electron enters the f-orbital of the ante-penultimate (n-2) energy level.
- General Outer Shell Electronic Configuration of f-Block Elements:

$$\textbf{[Noble gas] } ns^2\, (n-1)d^{0-1}\, (n-2)f^{1-14}$$

1.3.4.1 Inner-transition Series:

- The f-block elements are organized into two series, each comprising 14 elements. In total, there are 28 f-block elements on the periodic table.
- These elements are situated at the lower part of the periodic table and are referred to as inner-transition elements.
    - First Series (**4f series**): Following lanthanum (57La), this series consists of elements known as lanthanoids. This series spans from cerium (58=Ce) to lutetium (71=Lu) and is also referred to as the rare earth elements.
    - Second Series (**5f series**): This series follows actinium (89Ac) and includes 14 elements from thorium (90=Th) to lawrencium (103=Lr). These elements are radioactive.

1.3.4.2 General Characteristics of f-Block Elements:

- They are characterized as hard and heavy metals.
- Majority of them have high melting and boiling points.
- These elements exhibit variable oxidation states. However, the most common oxidation state of lanthanoids and actinoids is +3. But, in certain compounds of actinoids, +4 oxidation state is also preferred.
- Compounds of f-block elements are often coloured and paramagnetic due to the presence of unpaired electrons in their valence shells.
- They possess a strong tendency to form complexes.
- Most of the elements in actinoid series are radioactive in nature.

1.4 Metals, Non-metals, and Metalloids

1.4.1 Metals:

- About 78 % of all known elements are metals and appear on the left-hand side of the Periodic Table. All s-block, d-block and f-block elements are metals.

1.4.1.1 General Characteristics of Metals:

- Electrical Conductance: Metals are excellent conductor of electricity. Hence, they are used in making electric wires.
- Heat Conductance: Metals are good conductor of Heat. Hence, they are used in making cooking utensils.

- Sound Conductance: Metals are good conductors of sound; hence, they are sonorous. When struck by a hard material, they produce a deep ringing sound.
- Malleability and Ductility:
    - Metals can be beaten into thin sheets. This property is called malleability.
    - Metals can be drawn into thin wires. This property is called ductility.
    - Gold (Au) is the most ductile metal. A wire of about 2 km length can be drawn from one gram of gold.
- Hardness: All metals are generally hard (Except sodium (Na) and potassium (K), which are soft and can be cut with knife). However, hardness varies from metal to metal.
- Metallic lustre: In pure state, metals have a shining surface and can be polished. This property is called metallic lustre.
- Physical State: Metals are solid at room temperature (except mercury (Hg), which is liquid at room temperature).
- Basic oxide: Metals forms basic oxides such as CaO, MgO, Na_2O, MnO_2 etc.
- Melting and boiling points: All metals have high melting and boiling points.
    - Tungsten (W) has the highest melting and boiling points whereas the mercury (Hg) has the lowest.
- High density:
    - Iridium (Ir : 22.56 g/cc) and Osmium (Os : 22.57 g/cc) have the highest densities.
    - Lithium (Li : 0.54 g/cc) has the lowest density.
    - Iron (Fe) has density = 7.9 g/cc
- Electro-positivity: Due to large atomic size, metals have low ionization enthalpies, they are electropositive in nature. Hence, they exist as cation in their compounds.
- Examples of metals: Gold (Au), Silver (Ag), Copper (Cu), Zinc (Zn), Mercury (Hg), Lead (Pb), Tungsten (W), Sodium (Na), Potassium (K), Iron (Fe) etc.

1.4.2 Non-Metals:

- They are located at top right-hand side of the Periodic Table. The number of non-metals is very few (i.e., 18).

1.4.2.1 General Characteristics of Non-Metals:

- Non-metals generally act as insulators of heat and electricity, except for graphite, which is an allotrope of carbon.
- Non-metals lack metallic luster and appear dull (except iodine).

- Non-metals are brittle in nature.
- Non-metals tend to have lower densities compared to metals.
- Due to their small atomic size, non-metals possess high ionization enthalpies, making them highly electronegative and tend to form anions in compounds.
- Most non-metals exist as gases at room temperature, including hydrogen, oxygen, nitrogen, fluorine, chlorine, and noble gases.
  - Few non-metals like Carbon, Sulphur, Phosphorus, and Iodine are solid at room temperature.
  - Bromine stands out as the sole liquid non-metal at room temperature.
- Non-metals do not produce sonorous sounds when struck against other materials.
- Many non-metals produce acidic oxides, like Sulphur dioxide, when dissolved in water.
- Some non-metals exhibit allotropy, adopting different structural arrangements under varying conditions.
  - Carbon exists in two crystalline allotropic forms: diamond and graphite.
  - Phosphorus displays three allotropic forms: white phosphorus, red phosphorus, and black phosphorus.
- Notable non-metals include fluorine, chlorine, bromine, iodine, hydrogen, oxygen, nitrogen, and noble gases.

1.4.3 Metalloids or Semi-metals:

- The elements silicon, germanium, arsenic, antimony, and tellurium show the properties of both metals and non-metals. These elements are called as metalloids and semi-metals.

1.4.4Trends in metallic and non-metallic behaviour:

- The metallic character increases as we go down a group.
- The non-metallic character increases as we go from left to right in a period.

Questions

1. Write the names and atomic number of the following elements:
    1. The fourth alkali metal
    2. The third alkaline earth metal
    3. The fifth element of the first transition series
    4. The first inner transition element
    5. The sixth noble gas
    6. The fourth halogen
    7. The third chalcogen
2. Predict the block of elements, if outer electronic configuration of elements is:
    1. $3s^2\,3p^3$
    2. $3d^5\,4s^1$
    3. $3s^2\,3p^6\,4s^2$
    4. $5d^1\,6s^2$
    5. $4f^1\,5d^1\,6s^2$
3. Arrange the following elements in the increasing order of metallic character: B, Al, Mg, K
4. Predict the period, group number and block of the following elements having atomic number:

    (i) 9 (ii) 12 (iii) 29 (iv) 54 (v) 84 (vi) 38 (vii) 58

    Also classify them as representative elements (noble gas, transition, and inner transition)

Part – 2

2.1 Periodic Properties of Elements

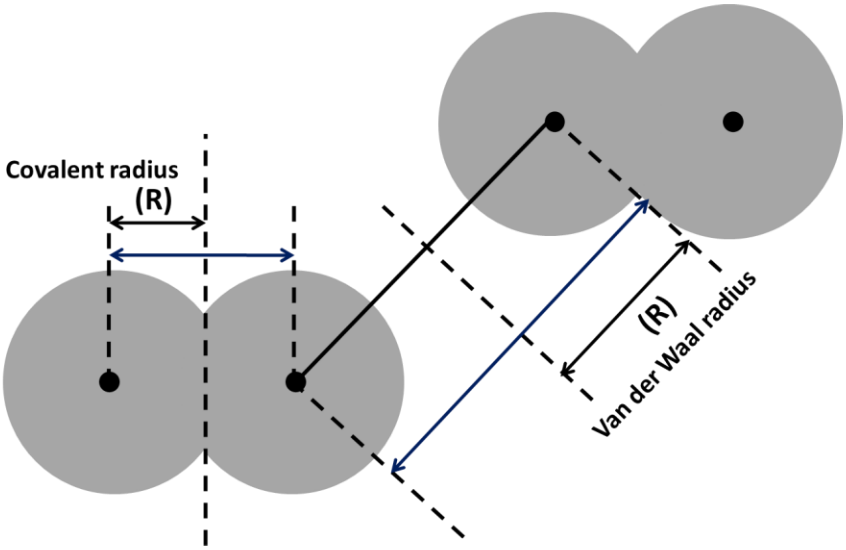
- Properties of elements directly or indirectly linked to their electronic configuration are termed periodic properties.
- These properties can be categorized into the following two groups:
  - **Properties of Individual Atoms:**
    Valency, atomic and ionic radii, ionization enthalpy, and other similar traits are properties attributed to individual atoms and are directly associated with their electronic configurations.
  - **Properties of Groups of Atoms:**
    Bulk properties such as melting point, boiling point, heat of fusion, density, and atomic volume are collective attributes and indirectly influenced by the electronic configurations of various atoms.

2.2 Atomic Radius and Atomic Size

- The atomic radius refers to the distance from the nucleus's center to the outermost electron-containing shell in an isolated atom.

2.2.1 Types of Atomic Radii:

- Depending on whether an element is a non-metal or a metal, three distinct atomic radius types are used.



- **Covalent Radius:**

  It can be defined as half of the distance between nuclei of two covalently bonded atoms of the same element within a molecule. Typically employed for non-metals.

  $$R_{covalent} = \frac{1}{2} \text{ [internuclear distance between two bonded atoms]}$$

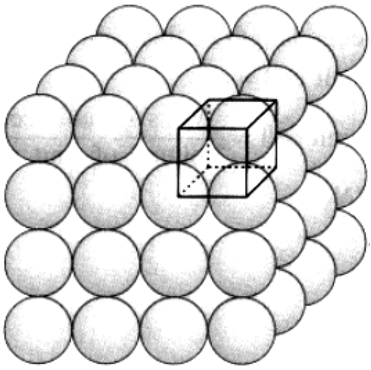
  Alternatively, $R_{covalent} = \frac{1}{2}$ [bond length]

- **Van der Waals Radius:**

  It can be defined as half the internuclear distance between similar non-bonded isolated atoms belonging to neighboring molecules of an element.

- **Metallic Radius:**

  It can be defined as half the internuclear distance between two adjacent metal ions in the metallic crystal lattice.



- Comparison of different atomic radii types:

  Vander Waal radius > Metallic radius > Covalent radius.

2.2.2 Factors Affecting Atomic Size:

- **Principal Quantum Number Value:**

  A higher principal quantum number (n) corresponds to a greater distance between the nucleus and the outermost shell, leading to an increase in atomic size.
- **Nuclear Charge:**

  As atomic number (number of protons) increases, nuclear charge also rises, resulting in an increased force of attraction between electrons and the nucleus. Consequently, atomic size decreases.

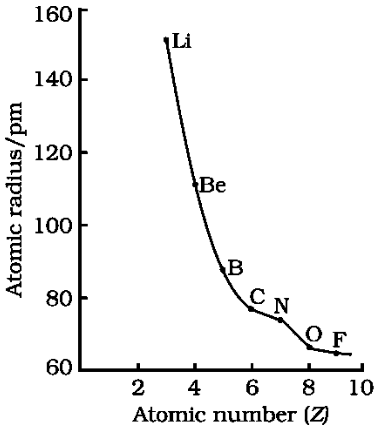
- **Shielding effect**
  It is a phenomenon where inner electrons in an atom create a protective barrier around the nucleus. This barrier helps outer electrons feel less attraction from the positively charged nucleus, resulting in an increase in atomic size.

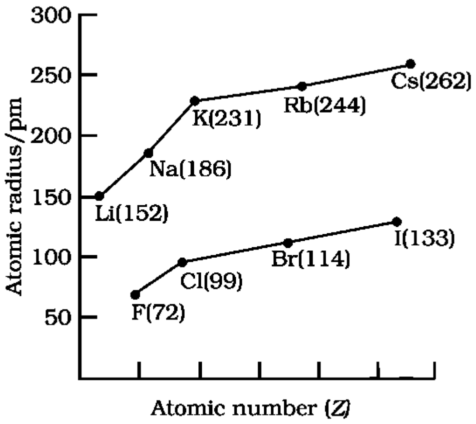
2.2.3 Variation of Atomic Radii in the Periodic Table:

- **Along the Group:**
  Atomic radius increases as we move down a group because of the addition of new energy shells or higher principal quantum number.
- **Along the Period**:
  - Atomic radius decreases as we move from left to right across a period.
  - This occurs because the nuclear charge increases by one unit in each successive element, while the main shell's number remains constant.
  - This increased nuclear charge pulls electrons closer to the nucleus, leading to a reduction in atomic radius.

Atom (Period II)	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**
Atomic radius (pm)	152	111	88	77	74	66	64
Atom (Period III)	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**
Atomic radius (pm)	186	160	143	117	110	104	99

Atom (Group 1)	**Atomic Radius**	**Atom (Group 17)**	**Atomic Radius**
Li	152	**F**	64
Na	186	**Cl**	99
K	231	**Br**	114
Rb	244	**I**	133
Cs	262	**At**	140





2.2.4 Difficulties in Determining Atomic Radius:

- **Isolation Challenge:**
  Isolating a single atom for determining its radius is not feasible.
- **Indistinct Boundary:**
  The atom's electron cloud lacks a well-defined boundary, making precise size determination difficult.
- **Uncertainty Principle:**
  The precise location of an electron around the nucleus cannot be pinpointed due to the Heisenberg uncertainty principle.
- **Variable Size:**
  The size of an atom changes from one bonded state to another.

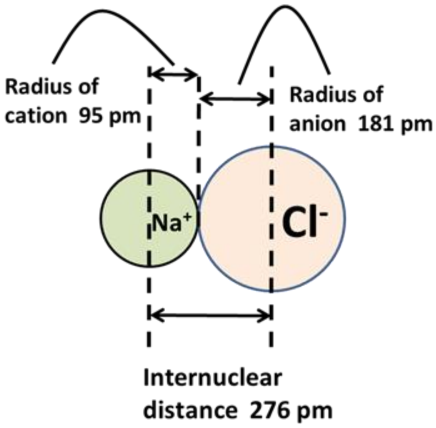
2.3 Ionic Radii

- Ionic radius is the effective distance from the nucleus of an ion up to which it influences its electronic cloud.

2.3.1 Pauling's Method for Determination of ionic radii:

- The internuclear distance of an ionic compound is the sum of the radii of the two oppositely charged ions involved (cation and anion).
- By knowing the radius of one ion, the radius of another ion can be calculated by subtracting the radius value of one ion from the measured internuclear distance in its compound.
  - Internuclear distance = Sum of radii of cation and anion.
  - Ionic radius of anion = Internuclear distance - Ionic radius of cation.

    Or
  - Ionic radius of cation = Internuclear distance - Ionic radius of anion.



- **Example:** Internuclear distance between Na^+ and Cl^- ions in NaCl crystal has been determined to be 276 pm. The absolute value of the ionic radius of Na^+ ion has determined to be 95 pm. Therefore, ionic radius of Cl^- ion is = 276 – 95 = 181 pm

2.3.2 Formation of Ions:

- Positive ions (cations) form when atoms lose electrons.

  $$A \rightarrow A^{+n} + ne-$$

- Negative ions (anions) form when atoms gain electrons.

  $$B + ne- \rightarrow B^{-n}$$

2.3.3 Size Comparison:

- Cations have smaller radii than their parent atoms. This is due to:
  - Decrease in the number of shells. Example:

    $$\underset{1s^2\ 2s^2\ 2p^6\ 3s^1}{Na} \longrightarrow \underset{1s^2\ 2s^2\ 2p^6}{Na^+} + e^-$$

  - Increase in effective nuclear charge leading to stronger attraction by the nucleus.
- Anions have larger radii than their parent atoms. This is due to:
  - Decrease in effective nuclear charge leading to weaker attraction by the nucleus.
  - Addition of electrons causing electron repulsions and expansion of electron cloud. Example:

    $$\underset{1s^2\ 2s^2\ 2p^6\ 3s^3\ 3p^5}{Cl + e^-} \longrightarrow \underset{1s^2\ 2s^2\ 2p^6\ 3s^3\ 3p^6}{Cl^-}$$

2.3.4 Variation of Ionic Radii in the Periodic Table:

- **Down a Group:**
  - As you move down a group (vertical column) in the periodic table, the ionic radii of both cations and anions increase.
  - This is because each successive element in the group adds a new energy shell, making the electron cloud farther from the nucleus.

- **Across a Period:**
  - As you move across a period (horizontal row) from left to right, the ionic radii generally decrease.
  - This is due to the increase in effective nuclear charge as the atomic number increases.
  - The increased nuclear charge pulls the electrons closer to the nucleus, resulting in smaller ionic radii.

- **Isoelectronic Series:**
  - Ions of different elements which have the same number of electrons but have different magnitude of nuclear charge are called isoelectronic ions.
    - Example: Al^{+3}, Mg^{+2}, Na^{+}, F^{-1}, O^{-2}, N^{-3} all have 10 electrons.
  - In isoelectronic series (ions with the same number of electrons), the ionic radii decrease as the nuclear charge increases.
  - As you move from cations to anions in the same isoelectronic series, the number of protons in the nucleus increases, causing stronger attraction and reducing the size of the electron cloud.
    - For example, among isoelectronic species like Al^{+3}, Mg^{+2}, Na^{+}, F^{-1}, O^{-2}, N^{-3}, the ionic radii follow the order: $Al^{+3} < Mg^{+2} < Na^{+} < F^{-1} < O^{-2} < N^{-3}$.

2.4 Ionisation Enthalpy:

- It is the minimum energy required to remove the outermost electron or least tightly bound electron from a gaseous neutral atom, converting it into a gaseous cation.
- It is represented by a symbol $\Delta_i H$.
- The process may be represented as;

$$M_{(g)} + \Delta_i H \rightarrow M^{+}_{(g)} + e\text{-}_{(g)}.$$

- Units of Ionisation Enthalpy;
  - It is measured in units of eV per atom, kcal/mol, or kJ/mol.
    - 1 eV = 1.602×10^{-19} J/atom = 23.06 kcal/mol = 96.49 kJ/mol.
  - **Note:**
    - 1 eV = 3.827×10^{-20} cal/atom
    - 1 cal = 4.184 J

2.4.1 Successive Ionisation Enthalpies:

- Energies needed to remove successive electrons from a gaseous atom.
- Example;
  - $M_{(g)} + \Delta_iH_1 \rightarrow M^+_{(g)} + e^-_{(g)}$
  - $M^+_{(g)} + \Delta_iH_2 \rightarrow M^{+2}_{(g)} + e^-_{(g)}$
  - $M^{+2}_{(g)} + \Delta_iH_3 \rightarrow M^{+3}_{(g)} + e^-_{(g)}$
- Order of Successive ionisation enthalpies: $\Delta_iH_3 > \Delta_iH_2 > \Delta_iH_1$.

2.4.2 Factors Affecting Ionisation Enthalpy:

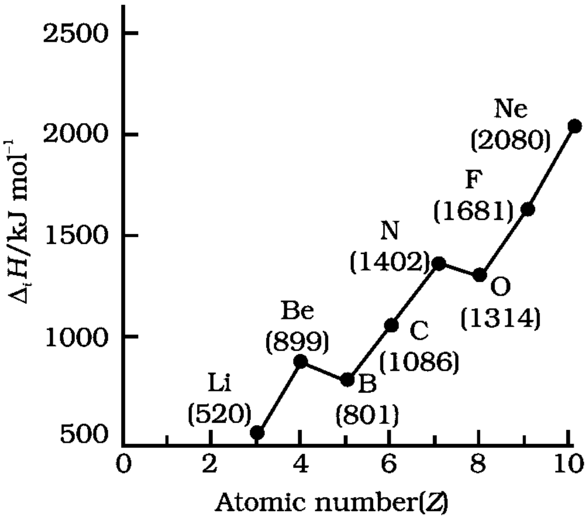
- **Atomic Radius:**
  Larger the atomic radius, weaker will be the force of attraction between the electron and nucleus. As a result, lesser amount of energy is required to remove the valence electron.
- **Nuclear Charge:**
  Higher the nuclear charge, stronger the attraction between nucleus and the valence electron, hence, greater is the amount of energy required to remove valence electron.
- **Screening Effect of inner shell electrons:**
  The inner shell electrons shield valence electrons as a result force of attraction between the valence electron and nucleus decreases. Hence, ionisation enthalpy decreases with increase in shielding effect.
- **Half-Filled or Completely Filled Orbitals:**
  If an atom has exactly half-filled or completely-filled orbitals, its ionisation enthalpy is higher than expected normally from its position in Periodic Table.

2.4.3 Variation of Ionisation Enthalpy in the Periodic Table:

- **Variation within a Group:**
  - Top to Bottom: Moving down a group → Ionisation enthalpy decreases.
  - Reason: Atomic size increases due to adding new main shells.
  - Effect: Valence electrons move farther from the nucleus, weakening attraction, and causing lower ionisation enthalpy.

- **Variation along a Period:**
  - Left to Right: Across a period → Ionisation enthalpy generally increases.
  - Cause: Atomic number rises, leading to increased nuclear charge.
  - Result: Valence electrons experience stronger attraction, raising ionisation enthalpy.

- **Variation within the 2nd Period:**



Here are some fluctuations in the first ionization enthalpies.

- **Li to Be:**
  Ionization enthalpy Increase due to higher nuclear charge and smaller atomic radius of Be.
- **Be to B:**
  Ionization enthalpy decrease because B's electronic configuration ($1s^2$ $2s^2$ $2p^1$) is less stable than Be's ($1s^2$ $2s^2$).
- **B to N:**
  Ionization enthalpy increase due to greater nuclear charge and reducing atomic radii.
- **N to O:**
  Ionization enthalpy decrease due to O's less stable configuration ($1s^2$ $2s^2$ $2p_x^2$ $2p_y^1$ $2p_z^1$) compared to N's ($1s^2$ $2s^2$ $2p_x^1$ $2p_y^1$ $2p_z^1$).
- **O to Ne:**
  Ionization enthalpy increase due to a rise in nuclear charge.

2.5 Electron Gain Enthalpy:

- It can be defined as the minimum energy released when a neutral gaseous atom accepts an extra electron to forms a gaseous negative ion (anion).
- It is also called **electron affinity** (Δ_{eg} H).
- The process can be represented by an equation given below:
  - $X_{(g)}$ + e- → $X^{-1}{}_{(g)}$; $\Delta_{eg}H$ = – ve
  - Example;
    - $Cl_{(g)}$ + e- → $Cl^{-1}{}_{(g)}$; $\Delta_{eg}H$ = –349 kJ/mol.

- **Halogen vs. Noble Gases:**
  - The value of electron gain enthalpy for halogens is highly negative, but noble gases have a large positive electron gain enthalpy (e.g., +48 kJ/mol in the case of He).
  - This is because in halogens, electron acceptance is strongly favored as they attain noble gas configuration by accepting an electron.
  - However, in the case of noble gases, the electron enters the next higher principal quantum number, requiring additional energy to be supplied upon adding an extra electron.

2.5.1 Successive Electron Gain Enthalpies:

- Successive electron gain enthalpy refers to the energy required to add a second or third electron to an anion (negatively charged ion) that already has one or more electrons added to it.
- The value of successive electron gain enthalpies is generally positive, indicating that energy is absorbed during these processes due to the electrostatic repulsion between negatively charged electrons already present in the anion.
- Example:
  - $O^{-1}{}_{(g)}$ + e- → $O^{-2}{}_{(g)}$; $\Delta_{eg}H_2$ = +780 kJ/mol
  - $S^{-1}{}_{(g)}$ + e- → $S^{-2}{}_{(g)}$; $\Delta_{eg}H_2$ = +590 kJ/mol

2.5.2 Factors Affecting Electron Gain Enthalpy:

- **Atomic Size:**
  Smaller the size of the atom, greater is the tendency to add an electron and hence, more negative electron gain enthalpy.

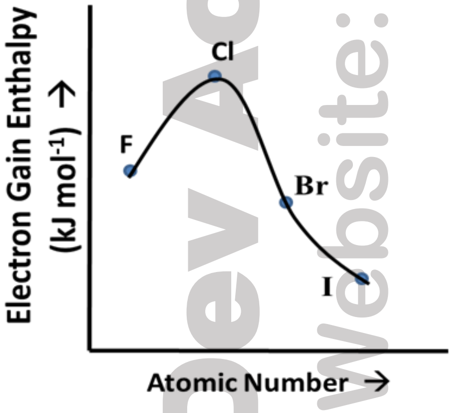
- **Effective Nuclear Charge:**
  Larger the effective nuclear charge, stronger the pull on the incoming electron, which results in more negative electron gain enthalpy.
- **Electronic Configuration:**
  Elements having exactly half-filled or completely-filled orbitals are very stable. As a result, energy must be supplied to add an electron. Hence, their electron gain enthalpies are positive. Because of this reason, the electron gain enthalpies of all the noble gases are positive.

2.5.3 Variation of Electron Gain Enthalpy in the Periodic Table:

- **Along the Period:**
  - Generally, electron gain enthalpy becomes more negative as we move from left to right in a period. This is because the nuclear charge increases while the atomic size decreases.
  - Exception:
    Halogens have the most negative electron gain enthalpies, but noble gases have positive electron gain enthalpies.

- **Down the Group:**
  - Generally, electron gain enthalpy becomes less negative as we move down a group due to the increasing atomic size, leading to a decrease in effective nuclear charge for the electron.

Group 1	$\Delta_{eg}H$	Group 16	$\Delta_{eg}H$
H	– 73		
Li	– 60	O	– 141
Na	– 53	S	– 200
K	– 48	Se	– 195
Rb	– 47	Te	– 190
Cs	– 46	Po	– 174



Group 17	$\Delta_{eg}H$	Group 0	$\Delta_{eg}H$
		He	+ 48
F	– 328	Ne	+ 116
Cl	– 349	Ar	+ 96
Br	– 325	Kr	+ 96
I	– 295	Xe	+ 77
At	– 270	Rn	+ 68

- Exception:
    - Nitrogen, Oxygen, and Fluorine (elements of the second period) have lower electron gain enthalpies compared to Phosphorus, Sulphur, and Chlorine, respectively (elements of the third period).
- Reason:
    - This is because the elements of the second period have the smallest atomic size among the elements in their groups. Due to this small size, there is considerable electron-electron repulsion within the atom, making it less receptive to incoming electrons compared to other elements in the groups.

2.6 Electronegativity

- Electronegativity is the tendency of an atom to attract shared pair of electrons towards itself in a covalent bond. It is denoted by x (kii).
- Unlike electron gain enthalpy, it is not directly measurable, but various scales like Pauling, Mulliken-Jaffe, and Allred-Rochow have been developed.

2.6.1 Factors Affecting Electronegativity:

- **State of Hybridization:**
  The electronegativity of an atom depends on its bonding state. For instance, sp-hybridized carbon is more electronegative than sp^2 or sp^3-hybridized carbon. This because electronegativity increases with increase in s-character.
- **Oxidation State:**
  Electronegativity increases with increase in the oxidation state of the element.
- **Nature of Substituent:**
  Electronegativity varies with the substituent attached to the atom. For instance, CF_3I has a more electronegative carbon atom than CH_3I.

2.6.2 Trends in the Periodic Table:

- Unlike electron affinity, electronegativity follows regular trend in Periodic table
- **Along a period:**
  Electronegativity of elements increase along a period from left to right.
- **Down a group:**
  Electronegativity of elements decrease down the group due to increased atomic size and decreased nuclear charge.

Electronegativity Values (on Pauling scale) Across the Periods

Atom (Period II)	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**
Electronegativity	1.0	1.5	2.0	2.5	3.0	3.5	4.0
Atom (Period III)	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0

Electronegativity Values (on Pauling scale) Down a Family

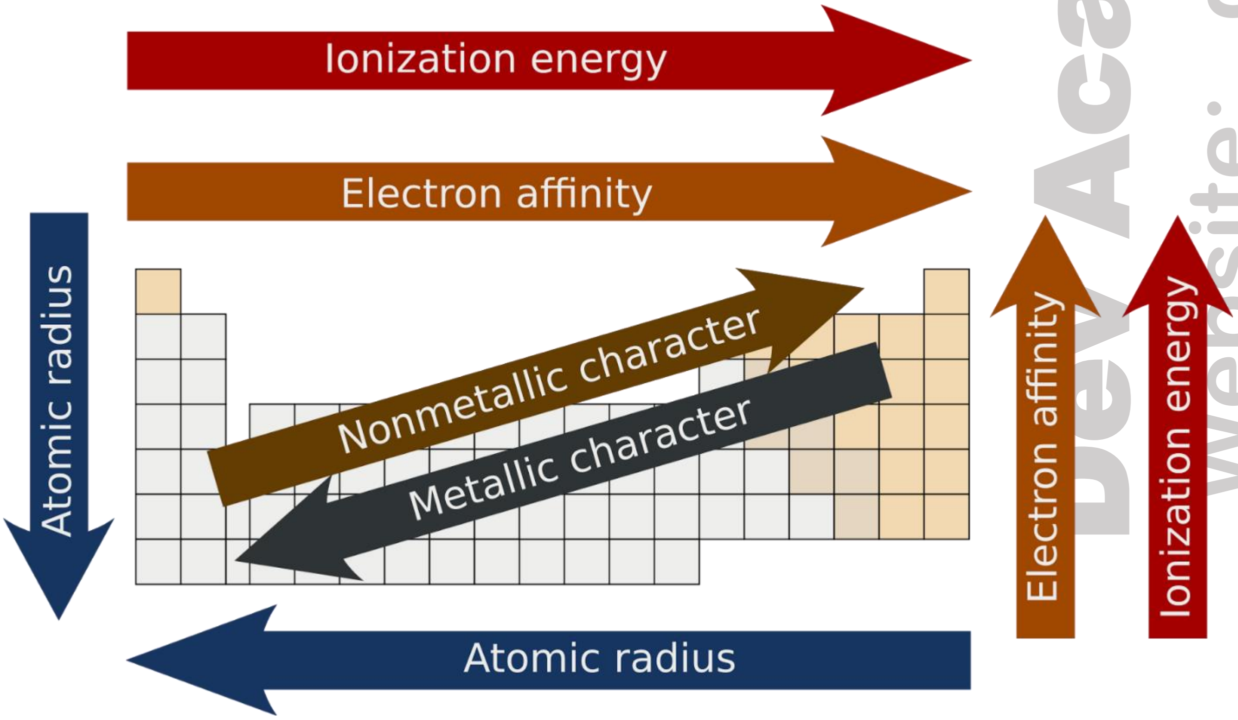
Atom (Group I)	**Electronegativity Value**	**Atom (Group 17)**	**Electronegativity Value**
Li	1.0	**F**	4.0
Na	0.9	**Cl**	3.0
K	0.8	**Br**	2.8
Rb	0.8	**I**	2.5
Cs	0.7	**At**	2.2

2.6.3 Applications of Electronegativity:

- **Metallic and Non-metallic Character:**
    - Higher electronegativity corresponds to greater non-metallic character, while lower electronegativity relates to more metallic character.
    - For example, Fluorine (4.0) is highly non-metallic, and Caesium (0.7) is very metallic.
- **Polar and Non-polar bond:**
    - Non-polar bonds are formed between two similar atoms with equal electronegativity, resulting in shared electrons being equally attracted.
      Example: H_2, Cl_2, O_2, N_2
    - Polar bonds are formed between two atoms with different electronegativity, causing the more electronegative atom to acquire a partial negative charge, and the less electronegative atom a partial positive charge.
      Example: HF, HCl, HBr, HI

2.6.3 Differences between Electronegativity and Electron affinity:

Sr. No.	Electronegativity	Electron Gain Enthalpy
1.	It is the tendency of an atom in a molecule to attract shared pair of electron towards itself.	It is the amount of energy released when isolated gaseous atom takes up an electron.
2.	It is the property of the bonded atoms.	It is the property of an Isolated atom.
3.	It has no unit as it is a number.	It is measured in Kcal mol^{-1} or Kj mol^{-1}.
4.	Different scales give different values of electronegativity of the same element.	Its value is absolute for an element.
5.	It is regular property. It increases along the period and decreases down a group.	It does not show regular trend in periods and in groups as a number of exceptions are known.
6.	The value of electronegativity is not fixed and changes with the hybrid state of the atom.	Every element has definite value of electron gain enthalpy.



Part – 3

3.1 Periodic Trends in Chemical Properties

3.1.1 Valence and Oxidation States:

- Valence refers to the number of electrons in outermost shell of an atom.
- Oxidation state of an element represents the charge acquired by an atom in its compound.
    - In general, the oxidation state of an atom is generally equal to either the number of valence electrons (for s- and p-block elements) or equal to eight minus the number of valence electron.

Groups	**1**	**2**	**13**	**14**	**15**	**16**	**17**	**18**
Number of valence electrons	1	2	3	4	5	6	7	8
Oxidation state	+1	+2	+3	–4, +4	–3, +3, +5	–2, +2, +6	–1, +1, +7	0, +8

- **Understanding Oxidation States:**
    - Consider compounds like OF_2 and Na_2O.
    - In OF_2, fluorine shares electrons with oxygen, resulting in an oxidation state of -1 for fluorine and +2 for oxygen.
    - In Na_2O, oxygen's higher electronegativity gives it oxidation state of -2, while sodium has an oxidation state of +1.

3.1.2 Acidic, Basic, and Amphoteric character of oxides:

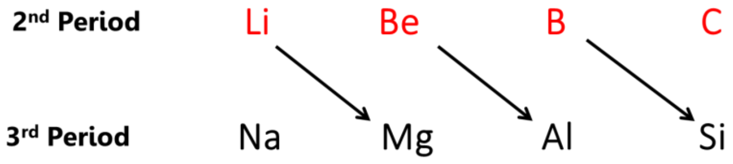
- It depends on the combined effect of several factors. For example:
- **Oxidation state:**
    - Among the oxides of the same element, higher the oxidation state of the element, stronger is the acid.
    - Example: P_4O_{10} is stronger acid in which P exhibit +5 oxidation state, while P_4O_6 is weaker acidic in which P exhibit +3 oxidation state.
- **Ionization enthalpy:**
    - If the ionization enthalpy of an element is high, its oxide will be acidic.
    - If the ionization enthalpy of an element is low, its oxide will be basic.
    - If the ionization enthalpy is intermediate, its oxide will be amphoteric.
    - On moving down the group, the acidic character first decreases, then become amphoteric and finally changed to basic.

B_2O_3 (weak acid) Al_2O_3 (Amphoteric) Ga_2O_3 (Amphoteric)

- **Electronegativity:**
    - Higher the electronegativity of the element, more acidic is its oxide.
- **Metals and Non-metals:**
    - Oxides formed by metals are basic, while those formed by non-metals are acidic.
        - CaO, Na_2O, and MgO are basic in nature.
        - SO_2, NO_2, P_4O_6, and CO_2 are acidic in nature.
    - However, oxide of hydrogen (i.e., H_2O), oxide of Zinc (ZnO), and oxide of Aluminium (Al_2O_3) are amphoteric.

3.1.3 Anomalous Properties of Second Period Elements:

- **Diagonal Relationship:**
    - The similarities in behavior between the first three elements in second period with the three elements of third period of next group is called diagonal relationship
        - Lithium shows diagonal relationship with Magnesium.
        - Beryllium shows diagonal relationship with Aluminium.
        - Boron shows diagonal relationship with Silicon.



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    - Elements like lithium and beryllium differ from other alkali and alkaline earth metals by forming compounds with covalent character. However, this behavior is more similar to the second element in the next group.
    - Causes of Diagonal Relationship:
        - Small atomic size.
        - High electronegativity.
        - High charge/radius ratio.
        - Absence of d-orbital in valence shell.
- **Distinctive Behavior of N, O, and F within their Group:**
    - **Nitrogen (N):**
        - Nitrogen primarily forms trihalides, exemplified by NCl_3.
        - In contrast, other elements in its group have the capacity to create both trihalides (e.g., PCl_3) and pentahalides (e.g., PCl_5).

  - **Oxygen (O):**
    - Oxygen predominantly forms dihalides, as seen in OF_2.
    - Conversely, fellow group members have the ability to produce tetrahalides (e.g., SF_4), and hexahalides (e.g., SF_6).
  - **Fluorine (F):**
    - Fluorine exhibits an oxidation state of only -1 in its compounds, such as HF.
    - In contrast, other members of its group are capable of displaying various oxidation states, including +1 (IF), +3 (IF_3), +5 (IF_5), and +7 (IF_7).

3.1.4 Atomic Volume:

- It is defined as the volume occupied by one mole of the atoms.
- Generally, lower atomic volume results in higher density, increased hardness and brittleness, less malleability and ductility.
- Boron is known to have lowest atomic volume while Francium has the highest atomic volume.
- **Trends in Periodic Table:**
  - As we move down the group, the atomic volume generally increases due to increase in number of shells.
  - However, as we move along the period, the atomic volume first decrease to minimum and then increase.

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