

Classification of Elements and Periodicity in Properties Class 11 Notes Chemistry

Chapter 3

• Genesis of Periodic Classification 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.

Triad	:	Lithium	Sodium	Potassium	Atomic mass of
Atomic mass	:	7	23	39	$\text{Na} = \frac{39 + 7}{2} = 23$
Triad	:	Chlorine	Bromine	Iodine	Atomic mass of
Atomic mass	:	35.5	80	127	$\text{Br} = \frac{127 + 35.5}{2} = 81.25$
Triad	:	Calcium	Strontium	Barium	Atomic mass of
Atomic mass	:	40	88	137	$\text{Sr} = \frac{40 + 137}{2} = 88.5$

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:

Li	Be	B	C	N	O	F
(7)	(9)	(11)	(12)	(14)	(16)	(19)
Na	Mg	Al	Si	P	S	Cl
(23)	(24)	(27)	(28)	(31)	(32)	(35.5)
K	Ca					
(39)	(40)					

Limitations of Newlands' Law of Octaves

(i) 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.

(ii) 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 of Mendeleev's Periodic Table

(i) In the periodic table, the elements are arranged in vertical rows called groups and horizontal rows known as periods.

(ii) 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.

(iii) There are seven periods (numbered from 1 to 7) or, horizontal rows in the Mendeleev's periodic table.

Importance of Mendeleev's Periodic Table

(i) This made the study of the elements quite systematic in the sense that if the properties of one element in a particular group are known, those of others can be predicted.

(ii) This helped to a great extent in the discovery of these elements at a later stage.

(iii) Mendeleev corrected the atomic masses of certain elements with the help of their expected positions and properties.

Defects in Mendeleev's Periodic Table

(i) Hydrogen has been placed in group IA along with alkali metals. But it also resembles halogens of group VII A in many properties. Thus, its position in the Mendeleev's periodic table is controversial.

(ii) Although the elements in the Mendeleev's periodic table have been arranged in order of their atomic masses, but in some cases the element with higher atomic mass precedes the element with lower atomic mass.

(iii) We know that the isotopes of an element have different atomic masses but same atomic number. Since, periodic table has been framed on the basis of increasing atomic masses of the elements, different positions must have been allotted to all the isotopes of a particular element.

(iv) According to Mendeleev, the elements placed in the same group must resemble in their properties. But there is no similarity among the elements in the two sub-groups of a particular group.

PERIODIC SYSTEM OF THE ELEMENTS IN GROUPS AND SERIES

SERIES	GROUP OF ELEMENTS											
	0	I	II	III	IV	V	VI	VII	VIII			
1		Hydrogen H 1.008										
2	Helium He 4.0	Lithium Li 7.03	Beryllium Be 9.1	Boron B 11.0	Carbon C 12.0	Nitrogen N 14.04	Oxygen O 16.00	Fluorine F 19.0				
3	Neon Ne 19.9	Sodium Na 23.5	Magnesium Mg 24.3	Aluminium Al 27.0	Silicon Si 28.4	Phosphorus P 31.0	Sulphur S 32.06	Chlorine Cl 35.45				
4	Argon Ar 38	Potassium K 39.1	Calcium Ca 40.1	Scandium Sc 44.1	Titanium Ti 48.1	Vanadium V 51.4	Chromium Cr 52.1	Manganese Mn 55.0	Iron Fe 55.9	cobalt Co 59	Nickel Ni 59	(Cu)
5		Copper Cu 63.6	Zinc Zn 65.4	Gallium Ga 70.0	Germanium Ge 72.3	Arsenic As 75	Selenium Se 79	Bromine Br 79.95				
6	Krypton Kr 81.8	Rubidium Rb 85.4	Strontium Sr 87.6	Yttrium Y 89.0	Zirconium Zr 90.6	Niobium Nb 94.0	Molybdenum Mo 96.0		Ruthenium Ru 101.7	Rhodium Rh 103.0	Palladium Pd (Ag) 106.5	
7		Silver Ag 107.9	Cadmium Cd 112.4	Indium In 114.0	Tin Sn 119.0	Antimony Sb 120.0	Tellurium Te 127.6	Iodine I 126.9				
8	Xenon Xe 128	Caesium Cs 132.9	Barium Ba 137.4	Lanthanum La 139	Cerium Ce 140							
9												
10				Ytterbium Yb 173		Tantalum Ta 183	Tungsten W 184		Osmium Os 191	Iridium Ir 193	Platinum Pt (Au) 194.9	
11		Gold Au 197.2	Mercury Hg 200.0	Thallium Tl 204.1	Lead Pb 206.9	Bismuth Bi 208						
12			Radium Ra 224		Thorium Th 232		Uranium U 239					
	R	R ₂ O	RO	R ₂ O ₃	HIGHER SALINE OXIDES RO ₂ R ₂ O ₅ RO ₃			R ₂ O ₇	RO ₄			
					HIGHER CASEOUS HYDROGEN COMPOUNDS RH ₄ RH ₃ RH ₂			RH				

Fig. 3.1 Mendeleev's Periodic Table published earlier.

- (v) In some cases, elements with similar properties have been placed in different groups.
- (vi) Lanthanoids and actinoids were placed in two separate rows at the bottom of the periodic table without assigning a proper reason.
- (vii) No proper explanation has been offered for the fact that why the elements placed in group show resemblance in their properties.

• Modern Periodic Law

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

• Present Form of the Periodic Table (Long form of Periodic Table)

The long form of periodic table, also called Modern Periodic Table, is based on Modern periodic law. In this table, the elements have been arranged in order of increasing atomic numbers.

• Nomenclature of Elements with Atomic No. more than 100

TABLE 3.1. Notation for IUPAC Nomenclature of Elements

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

TABLE 3.2. Nomenclature of Elements with Atomic Number above 100

<i>Atomic Number</i>	<i>Name</i>	<i>Symbol</i>	<i>IUPAC Official Name</i>	<i>IUPAC Symbol</i>
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	Röntgenium*	Rg*
112	Ununbium	Uub	*	*
113	Ununtrium	Uut	+	
114	Ununquadium	Uuq	*	*
115	Ununpentium	Uup	+	
116	Ununhexium	Uuh	*	*
117	Ununseptium	Uus	+	
118	Ununoctium	Uuo	+	

• Structural Features of the Periodic Table
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:

- (i) All the elements present in a group have same general electronic configuration of the atoms.
- (ii) The elements in a group are separated by definite gaps of atomic numbers (2, 8, 8,18, 18,32).
- (iii) The atomic sizes of the elements in group increase down the group due to increase the number of shells.
- (iv) The physical properties of the elements such as m.p., b.p. density, solubility etc., follow a systematic pattern.
- (v) 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.

s-Block Elements

General electronic configuration: ns^{1-2} Characteristics of s-block elements:

- (i) All the elements are soft metals.
- (ii) They have low melting and boiling points.
- (iii) They are highly reactive.
- (iv) Most of them impart colours to the flame.
- (v) They generally form ionic compounds.

p-Block Elements

General electronic configuration: $ns^2 np^{1-6}$

Characteristics of p-block elements:

- (i) The compounds of these elements are mostly covalent in nature.
- (ii) They show variable oxidation states.
- (iii) In moving from left to right in a period, the non-metallic character of the elements increases.
- (iv) The reactivity of elements in a group generally decreases downwards.
- (v) At the end of each period is a noble gas element with a closed valence shell $ns^2 np^6$ configuration.
- (vi) Metallic character increases as we go down the group.

d-Block Elements

General electronic configuration: $(n-1)d^{1-10} ns^0-2$

The d-block elements are known as transition elements because they have

incompletely filled d-orbitals in their ground state or in any of the oxidation states.

IUPAC (1934)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
Bohr (1920)	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIII			IB	IIB	IIIB	IVB	VB	VIB	VIIA	VIIIA	0
← s-Block Elements → $[ns^1-2]$ ← p-Block Elements → $[ns^2 np^{1-6}]$																			
Period ↓	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
1	1 H Hydrogen 1.0079																	2 He Helium 4.00260	
2	3 Li Lithium 6.941	4 Be Beryllium 9.01218											5 B Boron 10.81	6 C Carbon 12.011	7 N Nitrogen 14.0067	8 O Oxygen 15.9994	9 F Fluorine 18.998403	10 Ne Neon 20.179	
3	11 Na Sodium 22.98977	12 Mg Magnesium 24.305											13 Al Aluminum 26.98154	14 Si Silicon 28.0855	15 P Phosphorus 30.97376	16 S Sulphur 32.06	17 Cl Chlorine 35.453	18 Ar Argon 39.948	
4	19 K Potassium 39.0983	20 Ca Calcium 40.08	21 Sc Scandium 44.9559	22 Ti Titanium 47.88	23 V Vanadium 50.9415	24 Cr Chromium 51.996	25 Mn Manganese 54.9380	26 Fe Iron 55.847	27 Co Cobalt 58.9332	28 Ni Nickel 58.69	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.72	32 Ge Germanium 72.58	33 As Arsenic 74.9216	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.80	
5	37 Rb Rubidium 85.4678	38 Sr Strontium 87.62	39 Y Yttrium 88.9059	40 Zr Zirconium 91.22	41 Nb Niobium 92.9064	42 Mo Molybdenum 95.93	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.905	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.69	51 Sb Antimony 121.75	52 Te Tellurium 127.60	53 I Iodine 126.9045	54 Xe Xenon 131.29	
6	55 Cs Caesium 132.9054	56 Ba Barium 137.33	57 La Lanthanum (138.905)	72 Hf Hafnium 178.49	73 Ta Tantalum 180.9479	74 W Tungsten 183.85	75 Re Rhenium 186.207	76 Os Osmium 190.2	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.9665	80 Hg Mercury 200.59	81 Tl Thallium 204.383	82 Pb Lead 207.2	83 Bi Bismuth 208.9804	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)	
7	87 Fr Francium (223)	88 Ra Radium 226.0254	89 Ac Actinium 227.0278 (180-103)	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (263)	107 Bh Bohrium (264)	108 Hs Hassium (265)	109 Mt Meitnerium (266)	110 Uuh Ununium (267)	111 Uuq Ununium (268)	112 Uub Unbinium (269)	113 Uut Untrium (270)	114 Uuq Unquadrium (271)	115 Uup Unpentium (272)	116 Uuh Unhexium (273)	117 Uus Unseptium (274)	118 Uuo Unoctium (276)	
← f-Block Elements → $[(n-2)f^{1-14} (n-1)d^{0-1} ns^2]$																			
* LANTHANIDE SERIES																			
	58 Ce Cerium 140.12	59 Pr Praseodymium 140.9077	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tm Terbium 158.9254	66 Dy Dysprosium 162.50	67 Ho Holmium 164.9304	68 Er Erbium 167.257	69 Tm Thulium 168.9342	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.967					
** ACTINIDE SERIES																			
	90 Th Thorium 232.0377	91 Pa Protactinium 231.0362	92 U Uranium 238.0289	93 Np Neptunium 237.0482	94 Pu Plutonium 244	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (260)					

The masses in parentheses () are the mass numbers of the most stable or best-known isotopes.
The elements 113, 115, 117 are not known but are included to show their expected positions.
The elements 114, 116 and 118 have only been reported recently.

Characteristics of d-block elements:

- (i) They are all metals with high melting and boiling points.

- (ii) The compounds of the elements are generally paramagnetic in nature.
- (iii) They mostly form coloured ions, exhibit variable valence (oxidation states).
- (iv) They are of tenly used as catalysts.

f-Block Elements

General electronic configuration: $(n - 2) f^{1-14} (n - 1) d^{0-1} ns^2$

They are known as inner transition elements because in the transition elements of d-block, the electrons are filled in $(n - 1) d$ sub-shell while in the inner transition elements of f-block the filling of electrons takes place in $(n - 2) f$ subshell, which happens to be one inner subshell. Characteristics of f-Block elements:

- (i) The two rows of elements at the bottom of the Periodic Table, called the Lanthanoids Ce ($Z = 58$) – Lu ($Z = 71$) and Actinoids Th ($Z = 90$) – Lr ($Z = 103$).
- (ii) These two series of elements are called Inner Transition Elements (f-Block Elements).
- (iii) They are all metals. Within each series, the properties of the elements are quite similar.
- (iv) Most of the elements of the actinoid series are radio-active in nature.

• Metals

- (i) Metals comprise more than 78% of all known elements and appear on the left side of the Periodic Table.
- (ii) Metals are solids at room temperature.
- (iii) Metals usually have high melting and boiling points.
- (iv) They are good conductors of heat and electricity.
- (v) They are malleable and ductile.

• Non-metals

- (i) Non-metals are located at the top right hand side of the Periodic Table.
- (ii) Non-metals are usually solids or gases at low temperature with low melting and boiling points.
- (iii) They are poor conductors of heat and electricity.
- (iv) The non-metallic character increases as one goes from left to right across the Periodic Table.
- (v) Most non-metallic solids are brittle and are neither malleable nor ductile.

• Metalloids

The elements (e.g., silicon, germanium, arsenic, antimony and tellurium) show the characteristic, of both metals and non-metals. These elements are also called semimetal.

• Noble Gases

- These are the elements present in group 18.
- Each period ends with noble gas element.
- All the members are of gaseous nature and because of the presence of all the occupied filled orbitals, they have very little tendency to take part in chemical combination.
- These are also called inert gases.

• Representative Elements

The elements of group 1 (alkali metals), group 2 (alkaline earth metals) and group 13 to 17 constitute the representative elements. They are elements of s-block and p-block.

• Transition Elements

The transition elements include, all the d-block elements and they are present in the centre of the periodic table between s and p-block elements.

• **Inner Transition Elements**

Lanthanoids (the fourteen elements after Lanthanum) and actinides (the fourteen elements after actinium) are called inner transition elements. They are also called f-block elements.

The elements after uranium are also called transuranic elements.

• **Periodic Trends in Properties of Elements**

Trends in Physical Properties

Atomic Radii: It is defined as the distance from the centre of the nucleus to the outermost shell containing the electrons. Depending upon whether an element is a non-metal or a metal, three different types of atomic radii are used. These are:

(a) Covalent radius (b) Ionic Radius (c) van der Waals radius (d) Metallic radius.

(a) Covalent Radius: It is equal to half of the distance between the centres of the nuclei of two atoms held together by a purely covalent single bond.

(b) Ionic Radius: It may be defined as the effective distance from the nucleus of an ion upto which it has an influence in the ionic bond.

(c) van der Waals Radius: Atoms of Noble gases are held together by weak van der Waals forces of attraction. The van der Waals radius is half of the distance between the centre of nuclei of atoms of noble gases.

(d) Metallic Radius: It is defined as half of the internuclear distance between the two adjacent metal ions in the metallic lattice.

• **Variation of Atomic Radius in the Periodic Table**

Variation in a Period: Along a period, the atomic radii of the elements generally decreases from left to right.

TABLE 3.3. Atomic Radii/pm Across the Periods

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

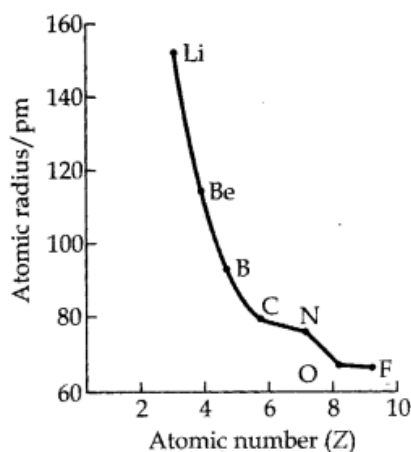


Fig. 3.2 Variation of atomic radius with atomic number across the second period.

Variation in a group: The atomic radii of the elements in every group of the periodic table increases as we move downwards.

TABLE 3.4. Atomic Radii/pm Down a family

Atom (Group I)	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

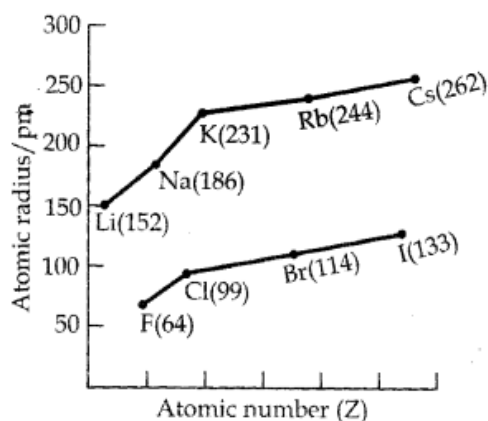


Fig. 3.3 Variation of atomic radius with atomic number for alkali metals and halogens.

• Ionic Radius

The ionic radii can be estimated by measuring the distances between cations and anions in ionic crystals.

In general, the ionic radii of elements exhibit the same trend as the atomic radii.

Cation: The removal of an electron from an atom results in the formation of a cation. The radius of cation is always smaller than that of the atom.

Anion: Gain of an electron leads to an anion. The radius of the anion is always larger than that of the atom.

Isoelectronic Species: Some atoms and ions which contain the same number of electrons, we call them isoelectronic species. For example, O^{2-} , F^{-} , Na^{+} and Mg^{2+} have the same number of electrons (10). Their radii would be different because of their different nuclear charges.

• Ionization Enthalpy

It is the energy required to remove an electron from an isolated gaseous atom in its ground state.

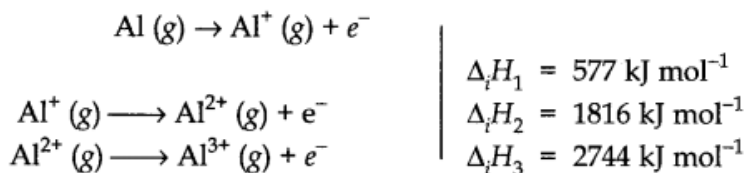


The unit of ionization enthalpy is kJ mol^{-1} and the unit of ionization potential is electron volt per atom.

Successive Ionization Enthalpies

If a gaseous atom is to lose more than one electron, they can be removed one after the other i.e., in succession and not simultaneously. This is known as successive ionization enthalpy (or potential).

The ionization enthalpies required to remove first, second and third electron from a gaseous atom are known as first ($\Delta_i H_1$), second ($\Delta_i H_2$) and third ($\Delta_i H_3$) ionization enthalpies respectively. The successive ionization enthalpies for Al (g) atom are as follows:



Thus $\Delta_i H_3 > \Delta_i H_2 > \Delta_i H_1$

• Variation of Ionization Enthalpies in the Periodic Table:

Variation of Ionization Enthalpy Along a Period

Along a period ionization enthalpies are expected to increase in moving across from left to the right, because the nuclear charge increases and the atomic size decreases.

Variation of Ionization Enthalpy in a Group

The ionization enthalpies of the elements decrease on moving from top to the bottom in any group.

The decrease in ionization enthalpies down any group is because of the following factors.

- (i) There is an increase in the number of the main energy shells (n) in moving from one element to the other.
- (ii) There is also an increase in the magnitude of the screening effect due to the gradual increase in the number of inner electrons.

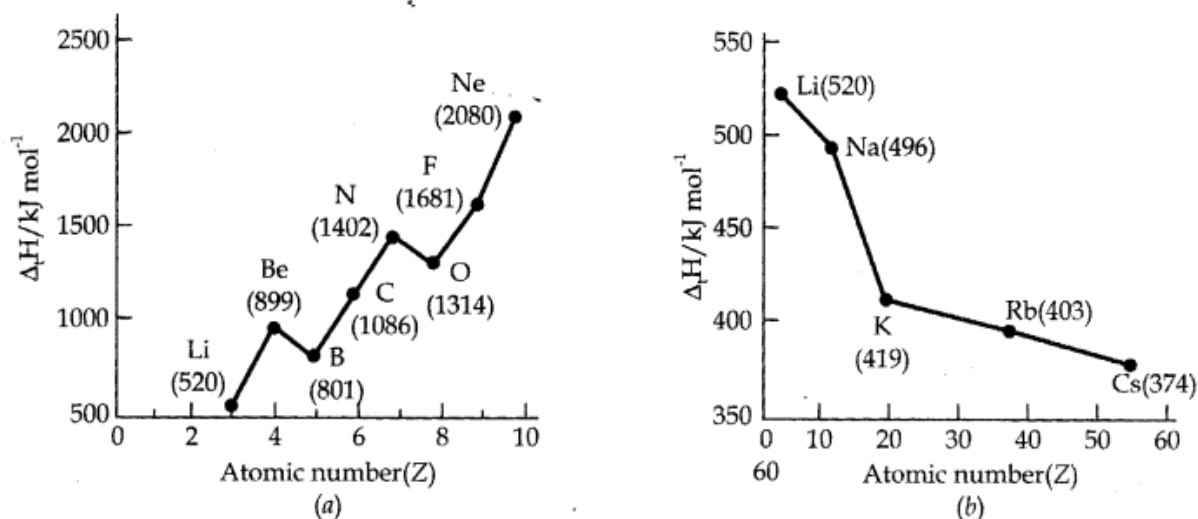
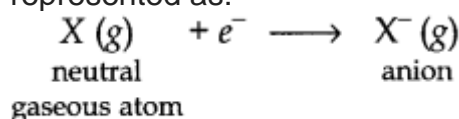


Fig. 3.4 (a) First ionization enthalpies ($\Delta_i H$) of elements of the second period as a function of atomic number (Z), (b) $\Delta_i H$ of alkali metals as a function of Z.

• Electron Gain Enthalpy

Electron Gain Enthalpy is the energy released when an electron is added to an isolated gaseous atom so as to convert it into a negative ion. The process is represented as:



$$\Delta H = \Delta_{eg} H$$

For majority of the elements the electron gain enthalpy is negative. For example, the

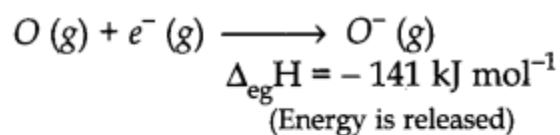
electron gain enthalpy for halogens is highly negative because they can acquire the nearest noble gas configuration by accepting an extra electron.

In contrast, noble gases have large positive electron gain enthalpies because the extra electron has to be placed in the next higher principal quantum energy level thereby producing highly unstable electronic configuration.

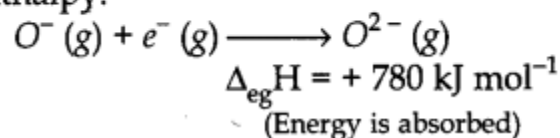
Successive Electron Gain Enthalpies

We have studied that electrons from a gaseous atoms are lost in succession (i.e., one after the other). Similarly, these are also accepted one after the other, i.e., in succession. After the addition of one electron, the atom becomes negatively charged and the second electron is to be added to a negatively charged ion. But the addition of second electron is opposed by electrostatic repulsion and hence the energy has to be supplied for the addition of second electron. Thus the second electron gain enthalpy of an element is positive.

For example, when an electron is added to oxygen atom to form O^- ion, energy is released. But when another electron is added to O^- ion to form O^{2-} ion, energy is absorbed to overcome the strong electrostatic repulsion between the negatively charged O^- ion and the second electron being added. Thus, first electron gain enthalpy:



Second Electron Gain Enthalpy:



Factors on which Electron Gain Enthalpy Depends

(i) Atomic size: As the size of an atom increases, the distance between its nucleus and the incoming electron also increases and electron gain enthalpy becomes less negative,

(ii) Nuclear charge: With the increase in nuclear charge, force of attraction between the nucleus and the incoming electron increases and thus electron gain enthalpy becomes more negative.

(iii) Symmetry of the Electronic Configuration: The atoms with symmetrical configuration (having fully filled or half filled orbitals in the same sub-shell) do not have any urge to take up extra electrons because their configuration will become unstable.

In that case the energy will be needed and electron gain enthalpy ($\Delta_{eg}H$) will be positive. For example, noble gas elements have positive electron gain enthalpies.

Variation of Electron Gain Enthalpy Across a Period

Electron gain enthalpy becomes more negative with increase in the atomic number across a period.

Variation of Electron Gain Enthalpy in a Group

Electron gain enthalpy becomes less negative as we go down a group.

TABLE 3.5. Electron Gain Enthalpies/(kJ mol⁻¹) of Some Main Group Elements

Group 1	$\Delta_{eg}H$	Group 16	$\Delta_{eg}H$	Group 17	$\Delta_{eg}H$	Group 0	$\Delta_{eg}H$
H	-73					He	+48
Li	-60	O	-141	F	-328	Ne	+116
Na	-53	S	-200	Cl	-349	Ar	+96
K	-48	Se	-195	Br	-325	Kr	+96
Rb	-47	Te	-190	I	-295	Xe	+77
Cs	-46	Po	-174	At	-270	Rn	+68

• **Electronegativity**

A qualitative measure of the ability of an atom in a chemical compound to attract shared electrons to itself is called electronegativity. Unlike ionization enthalpy and electron gain enthalpy, it is not a measurable quantity.

However, a number of numerical scales of electronegativity of elements viz, Pauling scale, Milliken- Jaffe scale, Allred Kochow scale have been developed. The electronegativity of any given element is not constant; it varies depending on the element to which it is bound.

Across a Period

Electronegativity generally increases across a period from left to right.

In a Group

It decreases down a group.

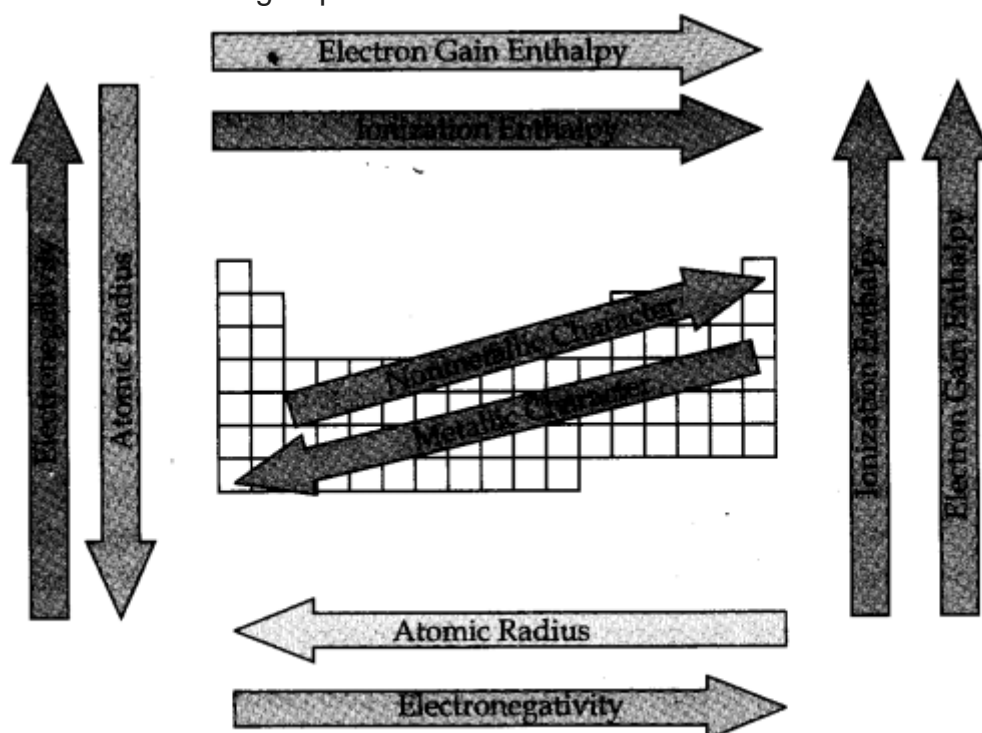


Fig 3.5 The periodic trends of elements in the periodic table.

TABLE 3.6 (a) Electronegativity Values (on Pauling scale) Across the Periods

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

Table 3.6 (b) Electronegativity Values (on Pauling scale) Down a Group

<i>Atom (Group I)</i>	<i>Electronegativity Value</i>	<i>Atom (Group 17)</i>	<i>Electronegativity Value</i>
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

• **Periodic Trends in Chemical Properties along a Period**

(i) Metallic character: Decrease across a period maximum on the extreme left (alkali metals).

(ii) Non-metallic character: Increases along a period. (From left to right).

(iii) Basic nature of oxides: Decreases from left to right in a period.

(iv) Acidic nature of oxides: Increases from left to right in a period.

• **Variation from Top to Bottom on Moving Down a Group**

(i) Metallic character. Generally increases because increase in atomic size and hence decrease in the ionization energy of the elements in a group from top to bottom.

(ii) Non-metallic character. Generally decreases down a group. As electronegativity of elements decreases from top to bottom in a group.

(iii) Basic nature of oxides. Since metallic character or electropositivity of elements increases in going from top to bottom in a group basic nature of oxides naturally increases.

(iv) Acidic character of oxides. Generally decreases as non-metallic character of elements decreases in going from top to bottom in a group.

(v) Reactivity of metals. Generally increases down a group. Since tendency to lose electron increases.

(vi) Reactivity of non-metals. Generally decreases down the group, Higher the electronegativity of non-metals, greater is their reactivity. Since electronegativity of non-metals in a group decreases from top to bottom, their reactivity also decreases.

• **Anomalous Properties of Second Period Elements**

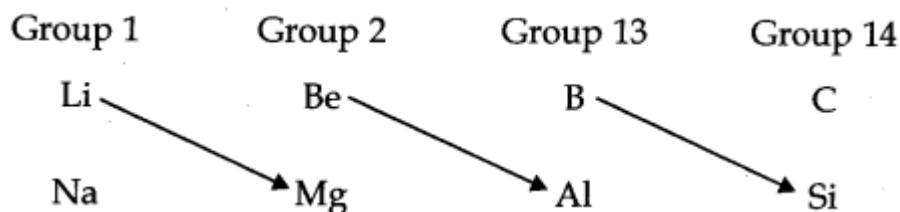
The first element of each of the group 1 (lithium) and 2 (beryllium) and group 13-17 (boron to fluorine) differs in many respects from the other members of their respective groups. For example, lithium unlike other alkali metals, and beryllium unlike other alkaline earth metals

form compounds which have significant covalent character; the other members of these groups, predominantly form ionic compounds.

It has been observed that some elements of the second period show similarities with the elements of the third period placed diagonally to each other, though belonging to

different groups.

For example,



This similarity in properties of elements placed diagonally to each other is called diagonal relationship.

- **Mendeleev's Periodic Law.** Physical and chemical properties of elements are periodic function of their atomic masses.
- **Modern Periodic Law.** Physical and chemical properties of the elements are periodic function of their atomic numbers.
- **Groups.** There are 18 groups. These are vertical rows.
- **Periods.** There are 7 periods. These are horizontal rows.
- **Representative Elements.** The S and P block of elements are known as representative elements.
- **Transition Elements.** They are also called d-block elements. They have general electronic configuration $(n - 1) d^{1-10} ns^{0-2}$.
- **Inner Transition Elements.** Lanthanoids (the fourteen elements after Lanthanum) and actinides (the fourteen elements after actinium) are called inner transition elements. General electronic configuration is $(n - 2) f^{1-14}(n - 1) d^{0-1} ns^2$. They are also called f-block elements.
- **Metals.** Present on the left side of the periodic table. Comprise more than 78% of the known elements.
- **Non-metals.** Mostly located on the right hand side of the periodic table.
- **Metalloids.** Elements which line as the border line between metals and non-metals (e.g., Si, Ge, As) are called metalloids or semimetals.
- **Atomic Radii and Ionic Radii,** increase down the group decrease along the period.
- **Ionization Enthalpy.** Increases along the period and decreases down the group.
- **Noble Gas Elements.** Elements with symmetrical configuration are chemically inert in nature.
- **Electric Nuclear Charge.** $Z = \text{Nuclear charge} - \text{Screening constant}$.
- **Electronegativity.** Increases along a period decreases down the group,
- **Chemical Reactivity.** Chemical reactivity is highest at the two extremess of a period and lowest in the centre.
- **Oxides of Elements.** Oxides formed of the Elements on the left are basic and of elements on the right are acidic in nature. Oxides of elements in the centre are amphoteric or neutral.