3 Classification of Elements and Periodicity in Properties

Facts that Matter

• 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	Na = $\frac{39+7}{2} = 23$
Triad	:	Chlorine	Bromine	Iodine	Atomic mass of
Atomic mass	:	35.5	80	127	Br = $\frac{127+35.5}{2} = 81.25$
Triad	:	Calcium	Strontium	Barium	Atomic mass of
Atomic mass	:	40	88	137	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 upto 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 pridicted.

(*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 is 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		
					HIGHER SALINE OXIDES				
	R	R_2O	RO	R_2O_3	RO_2	R_2O_5	RO_3	R_2O_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

Digit	*Name*	*Abbreviation*
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

Atomic Number	*Name*	*Symbol*	*IUPAC Official Name*	*IUPAC Symbol*
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassnium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnillium	Uun	Darmstadtium	Ds
111	Unununnium	Uuu	Rontgenium*	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.
(*vi*) They are good conductors of heat and electricity.

p-Block Elements

General electronic configuration: ns^2np^{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	VIIB	0

← *s*-Block Elements → [$ns^{1,2}$]

⟵ *p* - Block Elements ⟶ [$ns^2\ np^{1-6}$]

d-Block Elements [$(n-1)d^{1-10}\ ns^{1,2}$]

Atomic number: 25
Symbol: Mn
Name: Manganese
Atomic mass: 54.9389

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** Aluminium 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 72.59	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 95.95	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.905	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.3868	48 **Cd** Cadmium 112.41	49 **In** Indium 114.82	50 **Sn** Tin 118.69	51 **Sb** Antimony 121.75	52 **Te** Tellurim 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 *(58.71) 138.905	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.9479	74 **W** Tungsten 183.85	75 **Re** Rhenlum 186.206	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 *(90-103)	104 **Rf** Ruther fordium (261)	105 **Db** Dubnium (262)	106 **Sg** Seaborgium (263)	107 **Bh** Bohrium (262)	108 **Hs** Hassium (265)	109 **Mt** Meitnerium (266)	110 **Uuh** Ununnilium	111 **Uuu**	112 **Uub**	113 **Uut**	114 **Uuq**	115 **Uup**	116 **Uuh**	117 **Uus**	118 **Uuo**

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 **Tn** Terbium 158.9254	66 **Dy** Dysprosium 162.50	67 **Ho** Holmium 164.9304	68 **Er** Erbium 167.9604	69 **Tm** Thulium 168.9342	70 **Yb** Ytterbium 173.04	71 **Lu** Lutetium 174.967
** ACTINIDE SERIES	90 **Th** Thorium 232.0381	91 **Pa** Protactinium 231.0359	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** Nobellium (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 oftenly 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*) Metal 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.
- Eash 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 effictive 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

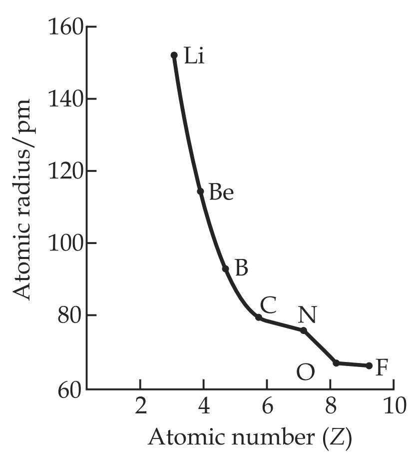

Fig. 3.2 *Variation of atomic radius with atomic number across the second period.*

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

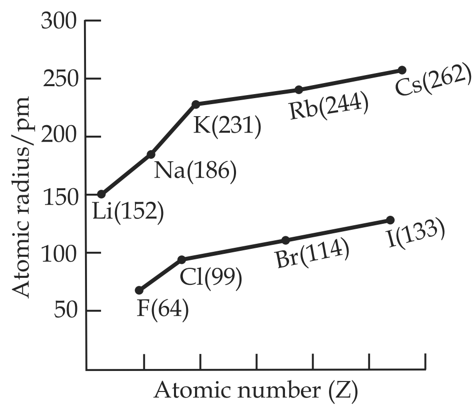

Fig. 3.3 *Variation of atomic radius with atomic number for alkali metals and halogens.*

• 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.

$$M\,(g) + I.E \rightarrow M^{+}\,(g) + e^{-}$$

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 ethalpies respectively. The successive ionization enthalpies for Al (g) atom are as follows:

$$Al\,(g) \rightarrow Al^{+}\,(g) + e^{-} \qquad \Delta_i H_1 = 577 \text{ kJ mol}^{-1}$$
$$Al^{+}\,(g) \longrightarrow Al^{2+}\,(g) + e^{-} \qquad \Delta_i H_2 = 1816 \text{ kJ mol}^{-1}$$
$$Al^{2+}\,(g) \longrightarrow Al^{3+}\,(g) + e^{-} \qquad \Delta_i H_3 = 2744 \text{ kJ mol}^{-1}$$

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 Ethalpy 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.

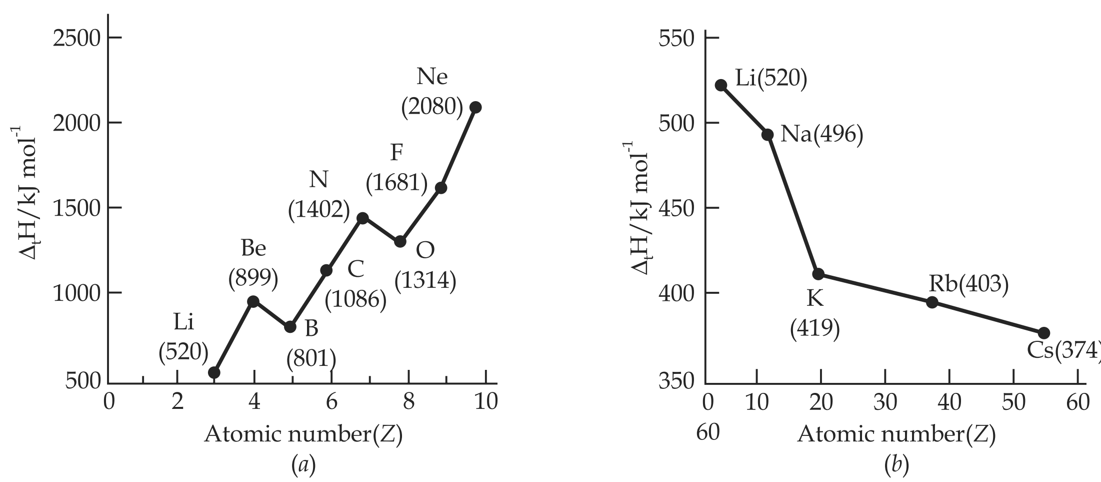

Fig. 3.4 *(a) First ionization enthalpies ($\Delta_i H$) of elements of the second period as a function of atomic number (Z), (b) ΔH of alkali metals as a function of Z.*

- **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:

$$\underset{\substack{\text{neutral}\\\text{gaseous atom}}}{X(g)} + e^- \longrightarrow \underset{\text{anion}}{X^-(g)}$$

$$\Delta H = \Delta egH$$

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:

$$O(g) + e^-(g) \longrightarrow O^-(g)$$

$$\Delta_{eg}H = -141 \text{ kJ mol}^{-1}$$

(Energy is released)

Second Electron Gain Enthalpy:

$$O^-(g) + e^-(g) \longrightarrow O^{2-}(g)$$

$$\Delta_{eg}H = +780 \text{ kJ mol}^{-1}$$

(Energy is absorbed)

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 ethalpy 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.

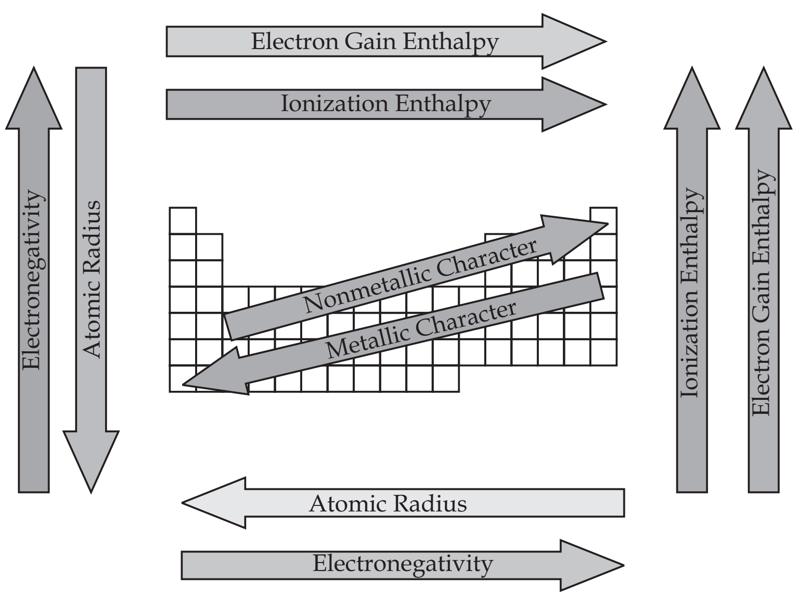

Fig 3.5 *The periodic trends of elements in the periodic table.*

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

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

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

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

- **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:** Increasess 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 oxidise 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 respect 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, pre-dominatly 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,

Group 1	Group 2	Group 13	Group 14
Li	Be	B	C
Na	Mg	Al	Si

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

Words that Matter

- **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 = Nuclear charge – 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.

NCERT TEXTBOOK QUESTIONS SOLVED

Q1. *What is the basic theme of organisation in the periodic table?*

Ans. The basic theme of organisation of elements in the periodic table is to simplify and systematize the study of the properties of all the elements and millions of their compounds. This has made the study simple because the properties of elements are now studied in form of groups rather than individually.

Q2. *Which important property did Mendeleev use to classify the elements in this periodic table and did he stick to that?*

Ans. Mendeleev used atomic weight as the basis of classification of elements in the periodic table. He did stick to it and classify elements into groups and periods.

Q3. *What is the basic difference in approach between Mendeleev's Periodic Law and the Modern Periodic Law?*

Ans. The basic difference in approach between Mendeleev's Periodic Law and Modern Periodic Law is the change in basis of classification of elements from atomic weight to atomic number.

Q4. *On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.*

Ans. The sixth period corresponds to sixth shell. The orbitals present in this shell are 6s, 4f, 5p, and 6d. The maximum number of electrons which can be present in these sub-shell is 2 + 14 + 6 + 10 = 32. Since the number of elements in a period corresponds to the number of electrons in the shells, therefore, sixth period should have a maximum of 32 elements.

Q5. *In terms of period and group where will you locate the element with z = 114?*

Ans. Period - 7 and Group - 14
Block-p.

Q6. *Write the atomic number of the element present in the third period and seventeenth group of the periodic table.*

Ans. The element is chlorine (Cl) with atomic number (Z) = 17.

Q7. *Which element do you think would have been named by*

(i) Lawrence Berkeley Laboratory
(ii) Seaborg's group?

Ans. (*i*) Lawrencium (Lr) with atomic number (z) = 103
(*ii*) Seaborgium (Sg) with atomic number (z) = 106.

Q8. *Why do elements in the same group have similar physical and chemical properties?*

Ans. The elements in a group have same valence shell electronic configuration and hence have similar physical and chemical properties.

Q9. *What does atomic radius and ionic radius really mean to you?*

Ans. **Atomic radius.** The distance from the centre of nucleus to the outermost shell of electrons in the atom of any element is called its atomic radius. It refers to both covalent or metallic radius depending on whether the element is a non-metal or a metal.

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

Q10. *How do atomic radius vary in a period and in a group? How do you explain the variation?*

Ans. Within a group Atomic radius increases down the group.

Reason. This is due to continuous increases in the number of electronic shells or orbit numbers in the structure of atoms of the elements down a group.

Variation across period.

Atomic Radii. From left to right across a period atomic radii generally decreases due to increase in effective nuclear charge from left to right across a period.

Q11. *What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.*
(i) F^- (ii) Ar (iii) Mg^{2+} (iv) Rb^+

Ans. Isoelectronic species are those species (atoms/ions) which have same number of electrons. The isoelectronic species are:

(*i*) Na^+ (*iii*) Na^+
(*ii*) K^+ (*iv*) Sr^{2+}

Q12. *Consider the following species:*

$$N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+}, Al^{3+}$$

(a) What is common in them?
(b) Arrange them in order of increasing ionic radii?

Ans. (*a*) All of them are isoelectronic in nature and have 10 electrons each.
(*b*) In isoelectronic species, greater the nuclear charge, lesser will be the atomic or ionic radius.

$$Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$$

Q13. *Explain why cation are smaller and anions larger in radii than their parent atoms?*

Ans. A cation is smaller than the parent atom because it has fewer electrons while its nuclear charge remains the same. The size of anion will be larger than that of parent atom because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge.

Q14. *What is the significance of the terms – isolated gaseous atom and ground state while defining the ionization enthalpy and electron gain enthalpy?*
[Hint: Requirements for comparison purposes]

Ans. (*a*) **Significance of term 'isolated gaseous atom'.** The atoms in the gaseous state are far separated in the sense that they do not have any mutual attractive and repulsive interactions. These are therefore regarded as isolated atoms. In this state the value of ionization enthalpy and electron gain enthalpy are not influenced by the presence of the other atoms. It is not possible to express these when the atoms are in the liquid or solid state due to the presence of inter atomic forces.

(*b*) **Significance of ground state.** Ground state of the atom represents the normal energy state of an atom. It means electrons in a particular atom are in the lowest energy state and they neither lose nor gain electron. Both ionisation enthalpy and electron gain enthalpy are generally expressed with respect to the ground state of an atom only.

Q15. *Energy of an electron in the ground state of the hydrogen atom is -2.18×10^{-18} J. Calculate the ionization enthalpy of atomic hydrogen in terms of J mol^{-1}.*
[Hint: Apply the idea of mole concept to derive the answer].

Ans. The ionisation enthalpy is for 1 mole atoms. Therefore, ground state energy of the atoms may be expressed as

$$\begin{aligned} E \text{ (ground state)} &= (-2.18 \times 10^{-18} \text{ J}) \times (6.022 \times 10^{23} \text{ mol}^{-1}) \\ &= -1.312 \times 10^{6} \text{ J mol}^{-1} \end{aligned}$$

Ionisation enthalpy $= E_{\infty} - E$ ground state

$= 0 - (-1.312 \times 10^6 \text{ J mol}^{-1})$

$= 1.312 \times 10^6 \text{ J mol}^{-1}$.

Q16. *Among the second period elements, the actual ionization enthalpies are in the order:*
$Li < B < Be < C < O < N < F < Ne$
Explain why

(i) *Be has higher $\Delta_i H_1$ than B?*

(ii) *O has lower $\Delta_i H_1$ than N and F?*

Ans. (*i*) In case of Be ($1s^2\ 2s^2$) the outermost electron is present in 2*s*-orbital while in *B* ($1s^2\ 2s^2\ 2p^1$) it is present in 2*p*-orbital. Since 2*s* - electrons are more strongly attracted by the nucleus than 2*p*-electrons, therefore, lesser amount of energy is required to knock out a 2*p*-electron than a 2*s* - electron. Consequently, $\Delta_i H_1$ of *Be* is higher than that $\Delta_i H_1$ of *B*.

(*ii*) The electronic configuration of

$$N_7 = 1s^2\ 2s^2\ 2p_x^{\ 1}\ 2p_y^{\ 1}\ 2p_z^{\ 1}$$
$$O_8 = 1s^2\ 2s^2\ 2p_x^{\ 2}\ 2p_y^{\ 1}\ 2p_z^{\ 1}$$

We can see that in case of nitrogen 2*p*-orbitals are exactly half filled. Therefore, it is difficult to remove an electron from *N* than from O. As a result $\Delta_i H$ of *N* is higher than that of O.

Q17. *How would you explain the fact that the first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium?*

Ans. Electronic configuration of Na and Mg are

$$\text{Na} = 1s^2\ 2s^2\ 2p^6\ 3s^1$$
$$\text{Mg} = 1s^2\ 2s^2\ 2p^6\ 3s^2$$

First electron in both cases has to be removed from 3*s*-orbital but the nuclear charge of Na (+ 11) is lower than that of Mg (+ 12) therefore first ionization energy of sodium is lower than that of magnesium.

After the loss of first electron, the electronic configuration of

$$\text{Na}^+ = 1s^2\ 2s^2\ 2p^6$$
$$\text{Mg}^+ = 1s^2\ 2s^2\ 2p^6\ 3s^1$$

Here electron is to be removed from inert (neon) gas configuration which is very stable and hence removal of second electron requires more energy in comparison to Mg.

Therefore, second ionization enthalpy of sodium is higher than that of magnesium.

Q18. *What are the various factors due to which the ionization enthalpy of the main group elements tends to decrease down the group?*

Ans. **Atomic size.** With the increase in atomic size, the number of electron shells increase. Therefore, the force that binds the electrons with the nucleus decreases. The ionization enthalpy thus decreases with the increase in atomic size.

Screening or shielding effect of inner shell electron. With the addition of new shells, the number of inner electron shells which shield the valence electrons increases. As a result, the force of attraction of the nucleus for the valence electrons further decreases and hence the ionization enthalpy decreases.

Q19. *The first ionization enthalpy values (in kJ mol^{-1}) of group 13 elements are:*

B	Al	Ga	In	Tl
801	577	579	558	589

How would you explain this deviation from the general trend?

Ans. The decrease in $\Delta_i H_1$ value from B to Al is due to the bigger size of Al.
In Ga there is 10 3d electrons which do not screen as is done by S and P electrons. Therefore, there is an unexpected increase in the magnitude of effective nuclear charge resulting in increased $\Delta_i H_1$ values. The same is with into Tl. The later has fourteen Δf electrons with very poor shielding effect. This also increases, the effective nuclear charge thus the value of $\Delta_i H_1$ increases.

Q20. *Which of the following pairs of elements would have a move negative electron gain enthalpy? (i) O or F (ii) F or Cl.*

Ans. (*i*) ***O* or *F*.** Both O and F lie in 2nd period. As we move from O to F the atomic size decreases.
Due to smaller size of F nuclear charge increases.
Further, gain of one electron by

$$F \rightarrow F^-$$

F^- ion has inert gas configuration, While the gain of one electron by

$$O \rightarrow O^-$$

gives O^- ion which does not have stable inert gas configuration, consequently, the energy released is much higher in going from

$$F \rightarrow F^-$$

than going from $O \rightarrow O^-$. In other words electron gain enthalpy of F is much more negative than that of oxygen.

(*ii*) The negative electron gain enthalpy of Cl (Δ eg H = – 349 kJ mol^{-1}) is more than that of F (Δ eg H = – 328 kJ mol^{-1}).
The reason for the deviation is due to the smaller size of F. Due to its small size, the electron repulsions in the relatively compact 2p-subshell are comparatively large and hence the attraction for incoming electron is less as in the case of Cl.

Q21. *Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first? Justify your answer.*

Ans. For oxygen atom:

$$O\,(g) + e^- \rightarrow O^-\,(g)\ (\Delta \text{ eg } H = -141 \text{ kJ mol}^{-1})$$
$$O^-\,(g) + e^- \rightarrow O^{2-}\,(g)\ (\Delta \text{ eg } H = +780 \text{ kJ mol}^{-1})$$

The first electron gain enthalpy of oxygen is negative because energy is released when a gaseous atom accepts an electron to form monovalent anion. The second electron gain enthalpy is positive because energy is needed to overcome the force of repulsion between monovalent anion and second incoming electron.

Q22. *What is basic difference between the terms electron gain enthalpy and electronegativity?*

Ans. Electron gain enthalpy refers to tendency of an isolated gaseous atom to accept an additional electron to form a negative ion. Whereas electronegativity refers to tendency of the atom of an element to attract shared pair of electrons towards it in a covalent bond.

Q23. *How would you react to the statement that the electronegativity of N on Pauling scale is 3.0 in all the nitrogen compounds?*

Ans. On Pauling scale, the electronegativity of nitrogen, (3.0) indicates that it is sufficiently electronegative. But it is not correct to say that the electronegativity of nitrogen in all the compounds is 3. It depends upon its state of hybridisation in a particular compound, greater the percentage of *s*-character, more will be the electronegativity of the element. Thus, the electronegativity of nitrogen increases in moving from SP^3 hybridised orbitals to SP hybridised orbitals *i.e.,* as $SP^3 < SP^2 < SP$.

Q24. *Describe the theory associated with the radius of an atom as it:*

(a) gains an electron (b) loses an electron?

Ans. (*a*) Gain of an electron leads to the formation of an anion. The size of an anion will be larger than that of the parent atom because the addition of one or more electrons would result in increased repulsion among electrons and decrease in effective nuclear charge.

This the ionic radius of fluoride ion (F^-) is 136 pm whereas atomic radius of Fluorine (F) is only 64 pm.

(*b*) Loss of an electron from an atom results in the formation of a cation. A cation is smaller than its parent atom because it has fomer electrons while its nuclear charge remains the same. For example, The atomic radius of sodium (Na) is 186 pm and atomic radius of sodium ion (Na^+) = 95 pm.

Q25. *Would you expect the first ionization enthalpies of two isotopes of the same element to be the same or different? Justify your answer.*

Ans. Ionization enthalpy, among other things, depends upon the electronic configuration (number of electrons) and nuclear charge (number of protons). Since isotopes of an element have the same electronic configuration and same nuclear charge, they have same ionization enthalpy.

Q26. *What are major differences between metals and non-metals?*

Ans.

Metals	*Non-Metals*
1. Have a strong tendency to lose electrons to form cations. 2. Metals are strong reducing agents. 3. Metals have low ionization enthalpies. 4. Metals form basic oxides and ionic compounds.	1. Non-metals have a strong tendency to accept electrons to form anions. 2. Non-metals are strong oxidising agent. 3. Non-metals have high ionization enthalpies. 4. Non-metals form acidic oxides and covalent compounds.

Q27. *Use periodic table to answer the following questions:*

(a) Identify the element with five electrons in the outer subshell.

(b) Identify the element that would tend to lose two electrons.

(c) Identify the element that would tend to gain two electrons.

Ans. (*a*) Element belonging to nitrogen family (group 15) *e.g.,* nitrogen.

(*b*) Element belonging to alkaline earth family (group 2) *e.g.,* magnesium.

(*c*) Element belonging to oxygen family (group 16) *e.g.,* oxygen.

Q28. *The increasing order of reactivity among group 1 elements is Li < Na < K < Rb < Cs whereas that of group 17 is F > Cl > Br > I. Explain?*

Ans. The elements of Group *I* have only one electron in their respective valence shells and thus have a strong tendency to lose this electron. The tendency to lose electrons in turn, depends upon the ionization enthalpy. Since the ionization enthalpy decreases down the group therefore, the reactivity of group 1 elements increases in the same order Li < Na < K < Rb < Cs. In contrast, the elements of group 17 have seven electrons in their respective valence shells and thus have strong tendency to accept one more electron to make stable configuration. It is linked with electron gain enthalpy and electronegativity. Since both of them decreases down the group, the reactivity therefore decreases.

Q29. *Write the general electronic configuration of s–, p–, d–, and f-block elements?*

Ans. (*i*) *s*-Block elements: ns^{1-2} where $n = 2 - 7$.

(*ii*) *p*-Block elements: $ns^2\ np^{1-6}$ where $n = 2 - 6$.

(*iii*) *d*-Block elements: $(n-1)\ d^{1-10}\ ns^{0-2}$ where $n = 4 - 7$.

(*iv*) *f*-Block elements: $(n-2)\ f^{0-14}\ (n-1)\ d^{0-1}\ ns^2$ where $n = 6 - 7$.

Q30. *Assign the position of the element having outer electronic configuration,* *(i)* $ns^2\ np^4$ *for* $n = 3$ *(ii)* $(n-1)\ d^2\ ns^2$ *for* $n = 4$ *and (iii)* $(n-2)\ f^7\ (n-1)\ d^1\ ns^2$ *for* $n = 6$ *in the periodic table?*

Ans. (*i*) $n = 3$

Thus element belong to 3rd period, *p*-block element.

Since the valence shell contains = 6 electrons.

group No = 10 + 6 = 16

configuration = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

element name is sulphur.

(*ii*) $n = 4$

Means element belongs to 4th period belongs to group 4 as in the valence shell (2 + 2) = 4 electrons.

Electronic configuration.

$= 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$

and the element name is Titanium (T_i).

(*iii*) $n = 6$

Means the element belongs to 6th period. Last electron goes to the *f*-orbital, element is from *f*-block.

group = 3

The element is gadolinium ($z = 64$)

Complete electronic configuration = $[X_e]\ 4f^7\ 5d^1\ 6s^2$.

Q31. *The first ($\Delta_i H_1$) and the second ($\Delta_i H_2$) ionization enthalpies (in kJ mol^{-1}) and the ($\Delta_{eg} H$) electron gain enthalpy (in kJ mol^{-1}) of a few elements are given below:*

Element	$\Delta_i H_1$	$\Delta_i H_2$	$\Delta_{eg} H$
I	*520*	*7300*	*– 60*
II	*419*	*3051*	*– 48*

III	*1681*	*3374*	*– 328*
IV	*1008*	*1846*	*– 295*
V	*2372*	*5251*	*+ 48*
VI	*738*	*1451*	*– 40*

Which of the above elements is likely to be:

(a) the least reactive element *(b) the most reactive metal*

(c) the most reactive non-metal *(d) the least reactive non-metal*

(e) the metal which can form a stable binary halide of the formula MX_2 *(X = halogen)*

(f) the metal which can form a predominantly stable covalent halide of the formula MX (X = halogen)?

Ans. (*a*) The element *V* has highest first ionization enthalpy ($\Delta_i H_1$) and positive electron gain enthalpy ($\Delta_{eg} H$) and hence it is the least reactive element. Since inert gases have positive $\Delta_{eg} H$, therefore, the element-*V* must be an inert gas. The values of $\Delta_i H_1$, $\Delta_i H_2$ and $\Delta_{eg} H$ match that of He.

(*b*) The element II which has the least first ionization enthalpy ($\Delta_i H_1$) and a low negative electron gain enthalpy ($\Delta_{eg} H$) is the most reactive metal. The values of $\Delta_i H_1$, $\Delta_i H_2$ and $\Delta_{eg} H$ match that of *K* (potassium).

(*c*) The element III which has high first ionization enthalpy ($\Delta_i H_1$) and a very high negative electron gain enthalpy ($\Delta_{eg} H$) is the most reactive non-metal. The values of $\Delta_i H_1$, $\Delta_i H_2$ and $\Delta_{eg} H$ match that of *F* (fluorine).

(*d*) The element IV has a high negative electron gain enthalpy ($\Delta_{eg} H$) but not so high first ionization enthalpy ($\Delta_i H_1$). Therefore, it is the least reactive non-metal. The values of $\Delta_i H_1$, $\Delta_i H_2$ and $\Delta_{eg} H$ match that of *I* (Iodine).

(*e*) The element VI has low first ionization enthalpy ($\Delta_i H_1$) but higher than that of alkali metals. Therefore, it appears that the element is an alkaline earth metal and hence will form binary halide of the formula MX_2 (where *X* = halogen). The values of $\Delta_i H_1$, $\Delta_i H_2$ and $\Delta_{eg} H$ match that of *Mg* (magnesium).

(*f*) The element I has low first ionization ($\Delta_i H_1$) but a very high second ionization enthalpy ($\Delta_i H_2$), therefore, it must be an alkali metal. Since the metal forms a predominarly stable covalent halide of the formula *MX* (*X* = halogen), therefore, the alkali metal must be least reactive. The values of $\Delta_i H_1$, $\Delta_i H_2$ and $\Delta_{eg} H$ match that of Li (lithium).

Q32. *Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements:*

(a) Lithium and oxygen *(b) Magnesium and nitrogen*

(c) Aluminium and iodine *(d) Silicon and oxygen*

(e) Phosphorous and fluorine *(f) Element 71 and fluorine.*

Ans. (*a*) LiO_2 (Lithium oxide) (*b*) Mg_3N_2 (Magnesium nitride)

(*c*) AlI_3 (Aluminium iodide) (*d*) SiO_2 (Silicon dioxide)

(*e*) Phosphorous pentafluoride (*f*) *Z* = 71

The element is Lutenium (Lu). Electronic configuration = $[X_e]\ 4f^{14}\ 5d^1\ 6s^2$. With fluorine it will form a binary compound = Lu F_3.

Q33. *In the modern periodic table, the period indicates the value of*

(a) atomic number (b) mass number (c) principal quantum number (d) azimuthal quantum number?

Ans. In the modern periodic table, each period begins with the filling of a new shell. Therefore, the period indicates the value of principal quantum number. Thus, option (*c*) is correct.

Q34. *Which of the following statements related to the modern periodic table is incorrect?*

(a) The p-block has six columns, because a maximum of 6 electrons can occupy all the orbitals in a p-subshell.

(b) The d-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a d-subshell.

(c) Each block contains a number of columns equal to the number of electrons that can occupy that subshell.

(d) The block indicates value of azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration.

Ans. Statement (*b*) is incorrect.

Q35. *Anything that influences the valence electrons will affect the chemistry of the element. Which one of the following factors does not affect the valence shell?*

(a) Valence principal quantum number (n)

(b) Nuclear charge (Z)

(c) Nuclear mass

(d) Number of core electrons.

Ans. (*c*) Nuclear mass.

Q36. *The size of isoelectronic species– F^-, Ne and Na^+ is affected by*

(a) nuclear charge (Z)

(b) valence principal quantum number (n)

(c) electron-electron interaction in the outer orbitals

(d) none of the factors because their size is the same

Ans. (*a*) Nuclear charge (*Z*).

Q37. *Which of the following statements is incorrect in relation to ionization enthalpy?*

(a) ionization enthalpy increases for each successive electron.

(b) The greatest increase in ionization enthalpy is experienced on removal of electrons from core noble gas configuration.

(c) End of valence electrons is marked by a big jump in ionization enthalpy.

(d) Removal of electron from orbitals bearing lower n value is easier than from orbital having higher n value.

Ans. (*d*) is incorrect.

Q38. *Considering the elements B, Al, Mg and K, the correct order of their metallic character is:*

(a) B > Al > Mg > K (b) Al > Mg > B > K (c) Mg > Al > K > B (d) K > Mg > Al > B

Ans. In a period, metallic character decreases as we move from left to right. Therefore, metallic character of K, Mg and Al decreases in the order: K > Mg > Al. However, within a group, the metallic character, increases from top to bottom. Thus, Al is more metallic than *B*. Therefore, the correct sequence of decreasing metallic character is: K > Mg > Al > B, *i.e.,* option (*d*) is correct.

Q39. *Considering the elements B, C, N, F and Si, the correct order of their non-metallic character is:*
(a) B > C > Si > N > F (b) Si > C > B > N > F (c) F > N > C > B > Si (d) F > N > C > Si > B

Ans. In a period, the non-metallic character increases from left to right. Thus, among *B*, *C*, *N* and *F*, non-metallic character decreases in the order: *F* > *N* > *C* > *B*. However, within a group, non-metallic character decreases from top to bottom. Thus, *C* is more non-metallic than Si. Therefore, the correct sequence of decreasing non-metallic character is: *F* > *N* > *C* > *B* > Si, *i.e.,* option (*c*) is correct.

Q40. *Considering the elements F, Cl, O and N, the correct order of their chemical reactivity in terms of oxidising property is:*
(a) F > Cl > O > N (b) F > O > Cl > N (c) Cl > F > O > N (d) O > F > N > Cl

Ans. Within a period, the oxidising character increases from left to right. Therefore, among *F*, *O* and *N*, oxidising power decreases in the order: *F* > *O* > *N*. However, within a group, oxidising power decreases from top to bottom. Thus, *F* is a stronger oxidising agent than Cl. Further because *O* is more electronegative than Cl, therefore, *O* is a stronger oxidising agent than Cl. Thus, overall decreasing order of oxidising power is: *F* > *O* > Cl > *N*, *i.e.,* option (*b*) is correct.

MORE QUESTIONS SOLVED

I. VERY SHORT ANSWER TYPE QUESTIONS

Q1. *State the Modern Periodic Law.*

Ans. Modern Periodic Law states that physical and chemical properties of the elements are a periodic function of their atomic numbers.

Q2. *Why is ionization enthalpy of nitrogen greater than that of oxygen?*

Ans. Nitrogen has exactly half filled *p*-orbitals.

Q3. *Why are electron gain enthalpies of Be and Mg positive?*

Ans. They have fully filled *s*-orbitals and hence have no tendency to accept an additional electron. That's why energy is needed if an extra electron is to be added. Therefore, electron gain enthalpies of Be and Mg are positive.

Q4. *Give four examples of species which are isoelectronic with* ca^{2+}.

Ans. Ar, K^+, Cl^-, S^{2-}, or P^{3-} are isoelectronic with ca^{2+}.

Q5. *Which two elements of the following belong to the same period?*
Al, Si, Ba and O

Ans. Al and Si.

Q6. *Explain why chlorine can be converted into chloride ion more easily as compared to fluoride ion from fluorine?*

Ans. Electron gain enthalpy of Cl is more negative than that of *F*.

Q7. *What are horizontal rows and vertical columns of the periodic table called?*

Ans. Horizontal rows are called periods and vertical columns are called groups.

Q8. *Which has a larger radius?*
(i) Mg or Ca (ii) S or Cl

Ans. (*i*) Ca (*ii*) S.

Q9. *What are representative elements?*

Ans. 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.

Q10. *Give general electronic configuration of f-block elements?*

Ans. General electronic configuration of *f*-block elements = $(n-2) f^{1-14} (n-1) d^{0-1} ns^2$.

Q11. *What are inner transition metals? Why are they called rare earth metals?*

Ans. Lanthanoids (the fourteen elements after Lanthanum) and actinides (the fourteen elements after actinium) are called inner transition elements.

Q12. *Define ionisation enthalpy.*

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

$$M(g) + I.E. \longrightarrow M^+(g) + e^-$$

Q13. *The electronic configuration of an element is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$. Locate the element in the periodic table.*

Ans. (*i*) As the principal quantum number for the valence shell is 4, the element is present in the 4th period.

(*ii*) Since the last electron has been filled in 4s sub-shell (or orbital), the element belongs to *s*-block.

(*iii*) As there is only one electron in the valence *s*-sub-shell, the element is present in group *I*.

II. SHORT ANSWER TYPE QUESTIONS

Q1. *What is the cause of periodicity in properties of the elements? Explain with two examples.*

Ans. The cause of periodicity in properties is the repetition of similar outer electronic configuration after certain regular intervals.

For example, all the elements of group *IA i.e.,* alkali metals, have similar outer electronic configuration as ns^1.

Where *n* refer to the number of outermost principal shell.

In a similar manner all the halogens *i.e.,* elements of group VIIA have similar other electronic configuration *i.e.,* $ns^2\ np^5$ and hence possess similar properties.

Q2. *Show by a chemical reaction with water that Na_2O is a basic oxide and Cl_2O_7 is an acidic oxide.*

Ans. Na_2O reacts with water to form sodium oxide which turns red litmus blue.

$$\underset{\text{Sod.oxide}}{Na_2O} + H_2O \longrightarrow \underset{\text{Sod.hydroxide}}{2NaOH}$$

Therefore, Na_2O is a basic oxide

In contrast, Cl_2O_7 reacts with water to form perchloric acid which turns blue litmus red.

$$Cl_2O_7 + H_2O \longrightarrow \underset{\text{perchloric acid}}{2HClO_4}$$

Therefore, Cl_2O_7 is an acidic oxide.

Q3. *What do you understand by 'Representative elements'? Name the groups whose elements are called representative elements.*

Ans. The elements of *s* and *p*-block are collectively called representative or main group elements. These include elements of group I (alkali metals), group 2 (alkaline earth metals).

Q4. *Name different blocks of elements in the periodic table. Give general electronic configuration of each block.*

Ans. Elements in the long form of the periodic table have been divided into four blocks *i.e.,* *s, p, d* and *f*. This division is based upon the name of the orbital which receives the last electron. General electronic configuration of

s-block elements: ns^{1-2} where $n = 2 - 7$

p-block elements: $ns^2\ np^{1-6}$ where $n = 2 - 6$

d-block elements: $(n-1)\ d^{1-10}\ ns^{0-2}$ where $n = 4 - 7$

f-block elements: $(n-2)\ f^{0-14}\ (n-1)\ d^{0-1}\ ns^2$ where $n = 6 - 7$

Q5. *Elements A, B, C and D have atomic numbers 12, 19, 29, and 36 respectively. On the basis of electronic configuration, write to which group of the periodic table each element belongs.*

Ans. Electronic configuration of *A* (Z = 12)

$= 1s^2\ 2s^2\ 2p^6\ 3s^2$

period = 3, Element's name = *Mg*

block = *s*, Group = II

Electronic configuration of *B* (Z = 19)

Element's name = *K* (potassium)

$= 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

n = 4, period = 4

Block = *s*, Group = *I*

Electronic configuration of *C* (Z = 29)

$= 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

n = 4, period = 4

Block = *d*

Electronic configuration of *D* (Z = 36)

$= 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

period = 4

Block = *p*-Block

group = 18

Q6. *Define the term ionization enthalpy? How does it vary along a period and along a group?*

Ans. **Ionization Enthalpy.** The minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom so as to convert it into a gaseous cation is called its ionization enthalpy or energy. It is represented by $\Delta_i H$.

This process may be represented as

$$M\ (g) + \Delta_i H \rightarrow M^+\ (g) + e^-\ (g)$$

where M (*g*) is isolated gaseous atom.

M^+ (*g*) is the resultant cation (a position ion)

Variation along a period. Moving from left to right in a period, the ionization enthalpy increases with atomic number.

Variation within a group. The ionization enthalpies keep on decreasing regularly as we move down a group from one element to the other.

Q7. *Discuss briefly the various factors on which ionization enthalpy depends.*

Ans. (*i*) **Atomic size.** With the increase in the atomic size, the number of electron shells increases. Therefore, the force that binds the electrons with the nucleus decreases. Thus, the ionization enthalpy decreases with increase in atomic size.

(*ii*) **Nuclear charge.** As the magnitude of the positive charge on the nucleus of an atom increases, the attraction with the electrons also increases. Therefore, the ionization enthalpy increases with the increase in the magnitude of the nuclear charge.

(*iii*) **Screening or shielding effect.** Greater the magnitude of the screening effect, less will be the value of ionization enthalpy or potential.

Q8. *What are Dobereiner's triads? Name two such triads.*

Ans. 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 third elements.

For example:

Triad:	lithium	sodium	potassium
Atomic mass:	7	23	39

$$\text{Atomic mass of Na} = \frac{39+7}{2} = 23$$

Triad:	Chlorine	Bromine	Iodine
Atomic mass:	35.5	80	127

$$\text{Atomic mass of Br} = \frac{127+35.5}{2} = 81.25$$

Q9. *Give the electronic configuration of the transition elements. Write their four important characteristics.*

Ans. The *d*-block elements are known as transition elements.

Electronic configuration = $(n-1)\,d^{1-10}\,ns^{1-2}$

Characteristics of *d*-block elements:

(*i*) They show variable oxidation states.

(*ii*) Their compounds are generally paramagnetic in nature.

(*iii*) Most of the transition elements form coloured compounds.

(*iv*) They are all metals with high melting and boiling points.

Q10. *What is screening or shielding effect? How does it influence the ionization enthalpy ?*

Ans. In a multielectron atom, the electrons present in the inner shells shield the electrons in the valence shell from the attraction of the nucleus or they act as a screen between the nucleus and these electrons. This is known as sheilding effect or screening effect.

As the screening effect increases, the effective nuclear charge decreases. Consequently, the force of attraction by the nucleus for the valence shell electrons decreases and hence the ionization enthalpy decreases.

Q11. *Define electron gain enthalpy. What are its units?*

Ans. The energy which is released by an atom in gaining an electron from outside atom or ion to form negative ion (or anion) is called electron gain enthalpy ($\Delta_{eg}H$).

Unit of electron gain enthalpy is kJ/mol.

In some cases, like in noble gas, atoms do not have any attraction to gain an electron. In that case energy has to be supplied.

For example,

$$Ne\,(g) + e^- \rightarrow Ne^-\,(g)$$
$$\Delta_{eg}\,H = +\,116\ \text{kJ mol}^{-1}$$

III. LONG ANSWER TYPE QUESTIONS

Q1. *Discuss the main features of long form of the periodic table. What are the advantages of long form of periodic table?*

Ans. Main features of long form of periodic table:

(*i*) **Groups.** The vertical columns in the periodic table are known as groups. There are 18 groups in the long form of periodic table.
Each group having the same electronic configuration in the outermost shell.

(*ii*) **Periods.** There are 7 periods in the long form of periodic table.
It is denoted by *n* which means highest principal quantum number.

(*iii*) **Lanthanoids.** Group of 14 elements in the sixth period. They are placed after Lanthanum.

(*iv*) **Actinides.** Group of 14 elements in the seventh period after actinium. Both Lanthanoids and actinoids are placed in separate panel at the bottom of the periodic table.

Advantages of long form of periodic table:

(*i*) It gives a suitable link between the position of element and its electronic configuration.

(*ii*) On the basis of atomic numbers it easier to remember all the elements.

(*iii*) The elements in the same group have similar properties due to their outer-most (valence shell) configuration. Thus it gives is a logical classification.

(*iv*) Justified positions are provided to transition and inner transition elements.

(*v*) It makes the study of elements systematic and simple.

Q2. *Discuss the main characteristics of four blocks of elements in the periodic table? Give their general electronic configuration.*

Ans. *s*-block elements:

(*i*) They are highly reactive elements and thus occurs in combined state. On moving down the group their reactivity increases.

(*ii*) They have good reducing characters.

(*iii*) They generally form electropositive ion by losing 1 or 2 electrons, that's why they are electropositive in nature.

(*iv*) They are good conductors of heat and electricity.

***p*-block elements:**

(*i*) Most of the *p*-block elements show variable oxidation states.

(*ii*) They include both metals and non-metals.

(*iii*) They are generally covalent in nature.

(*iv*) As move from left to right the non-metallic character of the element increases. On moving down the group metallic character increases.

***d*-block elements:**

(*i*) *d*-block elements show variable oxidation states.

(*ii*) They are generally paramagnetic in nature.

(*iii*) Their compounds are generally coloured. Those which form complex compounds.

(*iv*) Most of the elements and their compounds acts as catalyst.

f-block elements:

(*i*) They are generally heavy metals having high melting and boiling points.
(*ii*) Their compounds are generally coloured.
(*iii*) Variable oxidation states are generally shown by these elements.
(*iv*) Most of Activities are radioactive.

General electronic configuration:

s-block — ns^{1-2}
p-block — $ns^2\, np^{1-6}$
d-block — $(n-1)\, d^{1-10}\, ns^{0-2}$
f-block — $(n-2)\, f^{0-14}\, (n-1)\, d^{0-1}\, ns^2$

Q3. *Define electron gain enthalpy. What are its units? Discuss the factors which influence the electron gain enthalpy.*

Ans. Electron gain enthalpy is the energy released when an isolated gaseous atom is converted into a negative ion by adding an extra electron.
Electron gain enthalpy is denoted by the sign Δ_{eg} H.
The process may be represented by

$$\underset{\text{neutral gaseous atom}}{M(g) + e^-} \longrightarrow \underset{\text{anion}}{M^-(g)}$$

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

electron gain enthalpy is negative or positive it depends upon the nature of the element. *For example.* For halogens it is highly negative, because they can acquire the noble gas configuration by accepting an extra electron.
In contrast. For noble gases have positive electron gain enthalpy because energy has to be supplied to the element.

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. Therefore, the force of attraction between the nucleus and the incoming electron decreases and hence the electron gain enthalpy becomes less negative.
(*ii*) **Nuclear charge.** As the nuclear charge increases force of attraction for the incoming electron increases and thus electron gain enthalpy becomes more negative.
(*iii*) **Symmetry of electronic configuration.** Elements having symmetrical configuration (Either half filled or fully filled orbitals in the same sub shell) having no attraction for electron because by accepting electron their configuration becomes less stable. In that case energy has to be supplied to accept electron. Thus electron gain enthalpy will be positive.

Q4. *Discuss the factors that influence the magnitude of ionization enthalpy. What are the general trends of variation of ionization enthalpy in the periodic table? Explain.*

Ans. **Factors affecting Ionization enthalpy.**

(*i*) **Atomic size.** With the increase in atomic size, the number of electron shells increases and thus the force of attraction between the electrons and the nucleus decreases. Therefore the ionization enthalpy decreases.

(*ii*) **Nuclear charge.** As the nuclear charge increases the attraction for the electron also increases that's why ionization enthalpy increases.

(*iii*) **Screening or shielding effect.** In a multi-electron atom, the electron present in the inner shells shield the electrons in the valence shell as a result these electrons experience less attraction from the nucleus. This leads to lesser ionization enthalpy.

Variation along a period. On moving from left to right in a period the nuclear charge increases and the atomic size decreases as a result ionization enthalpies are expected to increase.

Variation within a group. On moving down the group as the atomic size of the elements increases that's why ionization enthalpy decreases down the group.

Q5. *(a) How does atomic radius vary in group in the periodic table?*

(b) Explain

(i) Radius of cation is less than that of the atom.

(ii) Radius of anion is more than that of the atom.

(iii) In iso-electronic ion, the ionic radii decreases with increase in atomic number.

Ans. (*a*) **Variation of atomic radius in a group:**

On moving down the group there is an increase in the principal quantum number and therefore no. of electron shells increases and thus the atomic size increases. Thus the atomic radii of the element increases.

(*b*) (*i*) **Radius of cation is less than that of the atom:**

Since the cation is formed by losing of one or more electrons.

For example,

$$Na \rightarrow Na^+ + e^-$$

Thus the radius of Na^+ will be less than the Na.

(*ii*) **Radius of anion is more than that of the atom.**

Since the anion is formed by gaining one or more electron. Therefore, the atomic radius is larger than the corresponding atom.

(*iii*) In iso-electronic ions, atoms have same number of electrons but different magnitude of nuclear charges. As the nuclear charge increases ionic radius decreases.

For example.

N^{3-}, O^{2-}, F^- have same No. of electrons = 10 but different ionic radii = 171, 140, 136 respectively.

IV. MULTIPLE CHOICE QUESTIONS

1. The highest ionization energy is exhibited by

(*a*) halogens (*b*) alkaline earth metals

(*c*) transition metals (*d*) noble gases

2. Which of the following oxides is neutral?

(*a*) SnO_2 (*b*) CO (*c*) Al_2O_3 (*d*) Na_2O

3. Which of the following is arranged in order of increasing radius?

(*a*) K^+ (*aq*) < Na^+ (*aq*) < Li^+ (*aq*) (*b*) K^+ (*aq*) > Na^+ (*aq*) > Zn^{2+} (*aq*)

(*c*) K^+ (*aq*) > Li^+ (*aq*) > Na^+ (*aq*) (*d*) Li^+ (*aq*) < Na^+ (*aq*) < K^+ (*aq*)

4. What is the electronic configuration of the elements of group 14?

(*a*) $ns^2 np^4$ (*b*) $ns^2 np^6$ (*c*) $ns^2 np^2$ (*d*) ns^2

5. Among the following elements, which has the least electron affinity?

(*a*) Phosphorous (*b*) Oxygen (*c*) Sulphur (*d*) Nitrogen

6. In halogens, which of the following, increases from iodine to fluorine?

(*a*) Bond length

(*b*) Electronegativity

(*c*) The ionization energy of the element

(*d*) Oxidizing power

7. Diagonal relationships are shown by

(*a*) Be and Al (*b*) Mg and Al (*c*) Li and Mg (*d*) B and P

8. Which of the following species are not known?

(*a*) AgOH (*b*) PbI_4 (*c*) PI_5 (*d*) SH_6

(*e*) All of the above

9. Which one of the following is isoelectronic with Ne?

(*a*) N^{3-} (*b*) Mg^{2+} (*c*) Al^{3+} (*d*) all of the above

10. Which element has smallest size?

(*a*) B (*b*) N (*c*) Al (*d*) P

Ans. 1. (*b*) 2. (*b*) 3. (*d*) 4. (*c*) 5. (*d*)
6. (*b*) (*c*) and (*d*) 7. (*a*) and (*c*) 8. (*e*) 9. (*d*) 10. (*b*)

V. HOTS QUESTIONS

Q1. *Arrange the following as stated: (i) N_2, O_2, F_2, Cl_2 (Increasing order of bond dissociation energy) (ii) F, Cl, Br, I (Increasing order of electron gain enthalpy) (iii) F_2, N_2, Cl_2, O_2 (Increasing order of bond length).*

Ans. (*i*) $F_2 < Cl_2 < O_2 < N_2$

(*ii*) I < Br < F < Cl

(*iii*) $N_2 < O_2 < F_2 < Cl_2$

Q2. *The first ionisation enthalpy of magnesium is higher than that of sodium. On the other hand, the second ionisation enthalpy of sodium is very much higher than that of magnesium. Explain.*

Ans. The 1st ionisation enthalpy of magnesium is higher than that of Na due to higher nuclear charge and slightly smaller atomic radius of Mg than Na. After the loss of first electron, Na^+ formed has the electronic configuration of neon (2, 8). The higher stability of the completely filled noble gas configuration leads to very high second ionisation enthalpy for sodium. On the other hand, Mg^+ formed after losing first electron still has one more electron in its outermost (3s) orbital. As a result, the second ionisation enthalpy of magnesium is much smaller than that of sodium.

Q3. *Give reasons:*

(*i*) IE_1 of sodium is lower than that of magnesium whereas IE_2 of sodium is higher than that of magnesium.

(*ii*) Noble gases have positive value of electron gain enthalpy.

Ans. (*i*) The effective nuclear charge of magnesium is higher than that of sodium. For these reasons, the energy required to remove an electron from magnesium is more than the energy required in sodium. Hence, the first ionization enthalpy of sodium is lower than that of magnesium.

However, the second ionization enthalpy of sodium is higher than that of magnesium.

This is because after losing an electron, sodium attains the stable noble gas configuration. On the other hand, magnesium, after losing an electron still has one electron.

(*ii*) Because of stable configuration.

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