SEMESTER–I (Period–III)

C10CH4

4
TOPIC

The Periodic Table/Periodic Chemistry

Learning Objectives

Upon completion of this topic, learners will be able to:

- Discuss the history and development of the periodic table
- Identify that elements are placed on the periodic table due to similar properties
- Identify the main blocks, groups and the periods of the periodic table
- Discuss the chemical and physical properties of the groups
- Discuss the periodic trends.

Introduction

Before the beginning of eighteenth century, only about 30 elements were known and it was quite easy to study and remember their individual properties. However, the situation became difficult with the discovery of large number of elements in the later years. At this stage, the scientists felt the need of some simple method to facilitate the study of the properties of various elements and their compounds. After numerous attempts the scientists were ultimately successful in arranging the elements in such a way so that similar elements were grouped together and different elements were separated. This arrangement of elements is known as *classification of elements* and it led to the formulation of *periodic table*. Thus, ***periodic table*** *may be defined as the table giving the arrangement of all the known elements according to their properties so that similar elements fall within the same vertical column and dissimilar elements are separated.*

4.1. HISTORY AND ORIGIN OF THE PERIODIC LAW

Earlier attempts on classification of elements were based on atomic masses. The formulation of a satisfactory periodic law took place only after 1860.

4.1.1. Dobereiner's Triads

The German chemist, Johann Dobereiner (1829) made the first significant attempt to show a relationship between atomic masses and the chemical properties of the elements. He observed that certain similar elements exist in groups of three elements which he named **triads**. An interesting feature of these triads was that the atomic mass of middle member was the arithmetic mean of the atomic masses of the other two members of the triad. For example, lithium, sodium and potassium constituted one triad. Atomic masses of lithium, sodium and potassium are 7, 23 and 39. We can observe that atomic mass of sodium is equal to the average of atomic masses of lithium and potassium.

$$\text{Atomic mass of sodium} = \frac{\text{Atomic mass of lithium} + \text{Atomic mass of potassium}}{2}$$

$$= \frac{7+39}{2} = 23$$

Sulphur, selenium, tellurium and chlorine, bromine, iodine are two more examples of triads. This classification of elements into triads was not satisfactory as it could be applied only to a limited number of elements.

4.1.2. Newland's Law of Octaves

The distinction of correlating the chemical properties of the elements with their atomic masses goes to J.A.R. Newlands. In 1866, he arranged the elements in the order of increasing atomic masses and noted a striking similarity between every eighth element. Newlands named this generalization, the *law of octaves*, due to its similarity to the musical scale. He stated that:

If the elements are arranged in the order of increasing atomic masses, the eighth element, starting from a given one is a kind of repetition of the first—like the eighth note in an octave of music.

Table 4.1 shows the first three rows of Newlands' table.

Table 4.1. Newlands' Arrangement of Elements

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

Newlands' arrangement worked well for the first 17 elements but failed beyond calcium. Moreover, with the discovery of noble gases, the eighth element no longer remained a similar element.

4.1.3. Lothar Meyer's Curves

In 1870, the German chemist Julius Lothar Meyer plotted the atomic volumes (i.e., the atomic mass divided by density) of the elements against their atomic masses. From his graph, Lothar Meyer was able to produce a table showing periodic arrangement of elements. A graph of atomic volume against atomic number has been shown in Fig. 4.1. This graph is similar to the plot of atomic volume versus atomic mass.

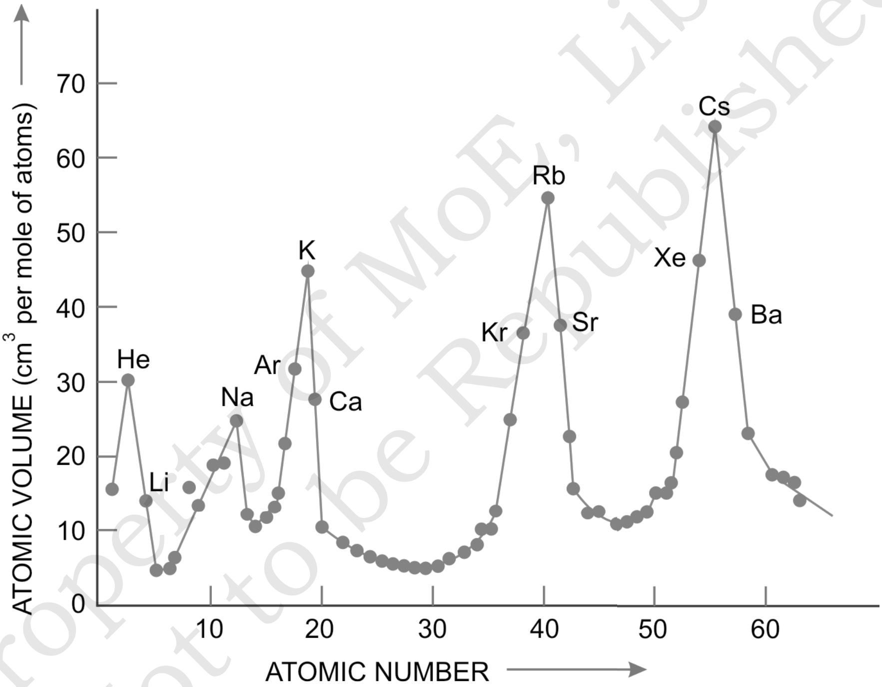

Fig. 4.1. Change of atomic volume with atomic number

It may be noted that the elements belonging to same chemical family occur at similar points on the curves. For example, alkali metals appear on the peaks whereas alkaline earth metals are on the descending portions of the curve while noble gases are on the ascending portions of the curve.

Although Lothar Meyer's curves showed a periodic repetition of properties with atomic masses, most of the credit for arranging the elements in a periodic table is given to Mendeleev.

4.1.4. Mendeleev's Periodic Table

Dmitri Ivanovich Mendeleev (1834–1907)

Dmitri Ivanovich Mendeleev produced a form of periodic table from which the modern periodic table was developed. When Mendeleev started his work 63 elements were known. He arranged all the elements known to him in the order of increasing atomic mass (then known as atomic weight) and showed that elements with similar properties recurred at regular intervals. The elements were arranged in such a way that the elements with similar properties fall in the same vertical column. **These vertical columns of similar elements are called groups and the horizontal rows of elements are called periods**. Mendeleev's periodic table has been shown in Table 4.2.

Table 4.2. Mendeleev's Periodic Table

Group	**I**		**II**		**III**		**IV**		**V**		**VI**		**VII**		**VIII**
Oxide: **Hydride:**	R_2O RH		RO RH_2		R_2O_3 RH_3		RO_2 RH_4		R_2O_5 RH_3		RO_3 RH_2		R_2O_7 RH		RO_4
	A	**B**	**A**	**B**	**A**	**B**	**A**	**B**	**A**	**B**	**A**	**B**	**A**	**B**	
Period 1	H														
Period 2	Li		Be		B		C		N		O		F		
Period 3	Na		Mg		Al		Si		P		S		Cl		
Period 4	K	Cu	Ca	Zn	*	*	*	Ti	As	V	Se	Cr	Br	Mn	Fe Co Ni
Period 5	Rb	Ag	Sr	Cd	In	Y	Sn	Zr	Sb	Nb	Te	Mo	I	Tc	Ru Rh Pd
Period 6	Cs	Au	Ba	Hg	Ti	La	Pb	Hf	Bi	Ta		W			Os Ir Pt

This arrangement of elements was based on the physical and chemical properties of the elements and also the formulae of the compounds they formed with oxygen and hydrogen. Mendeleev selected oxygen and hydrogen as the elements because they are very reactive and formed compounds with most of the elements. The formulae of the hydrides and oxides formed by an element were treated as one of the basic properties for its classification. For example, all the metals of group IA form hydrides with molecular formula MH.

4.1.5. Modern Periodic Law

Mendeleev's periodic table had a number of drawbacks. A large number of scientists made attempts to remove these. In 1914, Henry Moseley, the English physicist, observed that physical and chemical properties of the elements are determined by their atomic numbers instead of their atomic masses. This observation led to the development of **modern periodic law**. The modern periodic law states that:

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

It means that if the elements are arranged in order of increasing atomic numbers, the elements with similar properties recur after regular intervals. Many new forms of periodic table have been proposed in recent times with modern periodic law as the guiding principle, but the general plan of the table remained the same as proposed by Mendeleev.

Based on this modern periodic law, the **modern periodic table** (Fig. 4.2 on page 94) was prepared. In this periodic table the elements are arranged into different groups and periods based on the electronic configuration.

4.2. STRUCTURE OF THE PERIODIC TABLE

Let us now, study structure of the periodic table.

4.2.1. Groups and Periods

4.2.1.1. Groups

A vertical column of elements in the periodic table is called a ***group***. A group consists of a series of elements having similar valence shell configuration. There are 18 vertical columns in the long form of periodic table. Thus, there are 18 groups in the long form of periodic table which are numbered from 1 to 18 according to recommendations of IUPAC.

Earlier, the designation of these groups was the same as in the Mendeleev's periodic table. The relationship between the two ways of numbering the groups is given below:

1	2	3	4	5	6	7	8	9	10
IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII		
11	12	13	14	15	16	17	18		
IB	IIB	IIIA	IVA	VA	VIA	VIIA	0		

PERIODS ↓ GROUPS →

This numbering system is used by the international Union of Pure and Applied Chemistry (IUPAC)

12 — Atomic number
2 8 2 — Electrons in each energy level
Mg — Element symbol
Magnesium
24.305 — Average atomic mass

TRANSITION METALS

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	1 **H** Hydrogen 1.0079 1																	2 **He** Helium 4.0026 2
2	3 **Li** Lithium 6.941 2 1	4 **Be** Beryllium 9.0122 2 2											5 **B** Boron 10.81 2 3	6 **C** Carbon 12.011 2 4	7 **N** Nitrogen 14.007 2 5	8 **O** Oxygen 15.999 2 6	9 **F** Fluorine 18.998 2 7	10 **Ne** Neon 20.179 2 8
3	11 **Na** Sodium 22.990 2 8 1	12 **Mg** Magnesium 24.305 2 8 2											13 **Al** Aluminium 26.982 2 8 3	14 **Si** Silicon 28.086 2 8 4	15 **P** Phosphorus 30.974 2 8 5	16 **S** Sulphur 32.06 2 8 6	17 **Cl** Chlorine 35.453 2 8 7	18 **Ar** Argon 39.948 2 8 8
4	19 **K** Potassium 39.098 2 8 8 1	20 **Ca** Calcium 40.08 2 8 8 2	21 **Sc** Scandium 44.958 2 8 9 2	22 **Ti** Titanium 47.90 2 8 10 2	23 **V** Vanadium 50.941 2 8 11 2	24 **Cr** Chromium 51.996 2 8 13 1	25 **Mn** Manganese 54.938 2 8 13 2	26 **Fe** Iron 55.847 2 8 14 2	27 **Co** Cobalt 58.933 2 8 15 2	28 **Ni** Nickel 58.71 2 8 16 2	29 **Cu** Copper 63.546 2 8 18 1	30 **Zn** Zinc 65.38 2 8 18 2	31 **Ga** Gallium 69.72 2 8 18 3	32 **Ge** Germanium 72.59 2 8 18 4	33 **As** Arsenic 74.922 2 8 18 5	34 **Se** Selenium 78.96 2 8 18 6	35 **Br** Bromine 79.904 2 8 18 7	36 **Kr** Krypton 83.80 2 8 18 8
5	37 **Rb** Rubidium 85.468 2 8 18 8 1	38 **Sr** Strontium 87.62 2 8 18 8 2	39 **Y** Yttrium 88.906 2 8 18 9 2	40 **Zr** Zirconium 91.22 2 8 18 10 2	41 **Nb** Niobium 92.906 2 8 18 12 1	42 **Mo** Molybdenum 95.94 2 8 18 13 1	43 **Tc** Technetium (97) 2 8 18 14 1	44 **Ru** Ruthenium 101.07 2 8 18 15 1	45 **Rh** Rhodium 102.91 2 8 18 16 1	46 **Pd** Palladium 106.4 2 8 18 18	47 **Ag** Silver 107.87 2 8 18 18 1	48 **Cd** Cadmium 112.41 2 8 18 18 2	49 **In** Indium 114.82 2 8 18 18 3	50 **Sn** Tin 118.69 2 8 18 18 4	51 **Sb** Antimony 121.75 2 8 18 18 5	52 **Te** Tellurium 127.60 2 8 18 18 6	53 **I** Iodine 126.90 2 8 18 18 7	54 **Xe** Xenon 131.30 2 8 18 18 8
6	55 **Cs** Cesium 132.91 2 8 18 18 8 1	56 **Ba** Barium 137.33 2 8 18 18 8 2	57 **La** Lanthanum 138.91 2 8 18 18 9 2	72 **Hf** Hafnium 178.49 2 8 18 32 10 2	73 **Ta** Tantalum 180.95 2 8 18 32 11 2	74 **W** Tungsten 183.85 2 8 18 32 12 2	75 **Re** Rhenium 186.21 2 8 18 32 13 2	76 **Os** Osmium 190.2 2 8 18 32 14 2	77 **Ir** Iridium 192.22 2 8 18 32 15 2	78 **Pt** Platinum 196.09 2 8 18 32 17 1	79 **Au** Gold 196.97 2 8 18 32 18 1	80 **Hg** Mercury 200.59 2 8 18 32 18 2	81 **Tl** Thallium 204.37 2 8 18 32 18 3	82 **Pb** Lead 207.2 2 8 18 32 18 4	83 **Bi** Bismuth 208.98 2 8 18 32 18 5	84 **Po** Polonium (209) 2 8 18 32 18 6	85 **At** Astatine (210) 2 8 18 32 18 7	86 **Rn** Radon (222) 2 8 18 32 18 8
7	87 **Fr** Francium (223) 2 8 18 32 18 8 1	88 **Ra** Radium 226.03 2 8 18 32 18 8 2	89 **Ac** Actinium (227) 2 8 18 32 18 9 2	104 **Rf** Rutherfordium (261) 2 8 18 32 32 10 2	105 **Db** Dubnium (262) 2 8 18 32 32 11 2	106 **Sg** Seaborgium (263) 2 8 18 32 32 12 2	107 **Bh** Bohrium (264) 2 8 18 32 32 13 2	108 **Hs** Hassium (265) 2 8 18 32 32 14 2	109 **Mt** Meitnerium (266) 2 8 18 32 32 15 2	110 **Ds** Darmstadtium (269) 2 8 18 32 32 16 2	111 **Rg** Roentgenium (272)	112 **Cn** Copernicium (285)	113 **Nh** Nihonium (286)	114 **Fl** Flerovium (289)	115 **Mc** Moscovium (290)	116 **Lv** Livermorium (293)	117 **Ts** Tennessine (294)	118 **Og** Oganesson (294)

Lanthanide Series	58 **Ce** Cerium 140.12 2 8 18 20 8 2	59 **Pr** Praseodymium 140.91 2 8 18 21 8 2	60 **Nd** Neodymium 144.24 2 8 18 22 8 2	61 **Pm** Promethium (145) 2 8 18 23 8 2	62 **Sm** Samarium 150.4 2 8 18 24 8 2	63 **Eu** Europium 151.96 2 8 18 25 8 2	64 **Gd** Gadolinium 157.25 2 8 18 25 9 2	65 **Tb** Terbium 158.93 2 8 18 27 8 2	66 **Dy** Dysprosium 162.50 2 8 18 28 8 2	67 **Ho** Holmium 164.93 2 8 18 29 8 2	68 **Er** Erbium 167.26 2 8 18 30 8 2	69 **Tm** Thulium 168.93 2 8 18 31 8 2	70 **Yb** Ytterbium 173.04 2 8 18 32 8 2	71 **Lu** Lutetium 174.97 2 8 18 32 9 2
Actinide Series	90 **Th** Thorium 232.04 2 8 18 32 18 10 2	91 **Pa** Protactinium 231.04 2 8 18 32 20 9 2	92 **U** Uranium 238.03 2 8 18 32 21 9 2	93 **Np** Neptunium 237.06 2 8 18 32 22 9 2	94 **Pu** Plutonium (244) 2 8 18 32 24 8 2	95 **Am** Americium (245) 2 8 18 32 25 8 2	96 **Cm** Curium (247) 2 8 18 32 25 9 2	97 **Bk** Berkelium (247) 2 8 18 32 27 8 2	98 **Cf** Californium (251) 2 8 18 32 28 8 2	99 **Es** Einsteinium (254) 2 8 18 32 29 8 2	100 **Fm** Fermium (257) 2 8 18 32 30 8 2	101 **Md** Mendelevium (258) 2 8 18 32 31 8 2	102 **No** Nobelium (259) 2 8 18 32 32 8 2	103 **Lr** Lawrencium (260) 2 8 18 32 32 9 2

Fig. 4.2. Modern Periodic Table

It may be mentioned here that **all the elements belonging to a particular group have same number of valence electrons and hence exhibit similar properties**. All the elements belonging to the same group constitute a family. For example, elements of group 1 are known as **alkali metals**. Similarly, elements of group-2 are known as **alkaline earth metals**, elements of group 16 are known as **chalcogens**, elements of group 17 are known as **halogens** and elements of group 18 are **noble gases**.

4.2.1.2. Periods

A horizontal row of a periodic table is called a **period**. A period constitutes *a series of elements having same valence shell.* There are **seven periods** in all, which are numbered as 1, 2, 3, 4, 5, 6 and 7.

Each period begins with the filling of new energy shell. In fact, *the number of period also represents the number of the valence shell of elements present in it.*

Different periods can accommodate different number of elements. It may be noted that periods 1, 2 and 3 contain 2, 8, 8 elements respectively and are called **short periods**. There are 18 elements in 4th and 5th periods and they are called the **long periods**. Sixth period containing 32 elements is the **longest period**. Seventh period also contains 32 elements.

Occurrence of 2, 8, 8, 18, 18 and 32 elements in the first, second, third, fourth, fifth, sixth and seventh period can be explained by keeping the following points in mind:

(a) All the elements in the periodic table are arranged in the order of their increasing atomic number.

(b) Each period in the periodic table starts with the beginning of a new energy level. For example, in case of first period the K-shell starts filling while in case of second period L-shell starts filling.

(c) Filling of electron shells takes place according to a well defined set of rules.

The first period begins with the first energy shell. First energy shell can accommodate only 2 electrons in it. **Hence, there are only 2 elements in the first period**.

The second period begins with the filling of second energy level. This shell can accommodate only eight electrons. **Hence, there are only eight elements in the second period**.

After the addition of eight electrons to the second shell, the second shell becomes fully filled. Now, the next electron enters the third shell and the elements now go to third period. The third shell has capacity to accommodate 18 electrons. However, after the third shell receives 8 electrons, the filling of fourth shell begins (because the outermost shell cannot have more than eight electrons). **Therefore, third period has only eight elements**.

With the beginning of the filling of fourth energy shell, the elements start entering the fourth period. Similarly, we can explain why there are 18, 18 and 32 elements in fourth, fifth and sixth period respectively.

There are two rows of elements at the bottom of the periodic table. The first row contains elements with atomic numbers 58 to 71. These 14 elements which follow lanthanum are called **lanthanoids**. The second row contains elements with atomic numbers 90 to 103. These 14 elements which follow actinium are known as **actinoids**.

Different elements belonging to a particular period have different electronic configurations and have different number of valence electrons. That is why, elements belonging to a particular period have different properties.

Example 4.1: *State one reason for keeping fluorine and chlorine in the same group of periodic table.*

Solution: Fluorine and chlorine have similar valence shell electronic configurations. Hence, they have similar properties and are placed in the same group of the periodic table.

$$_{9}\text{F} \ : \ 2, 7$$

$$_{17}\text{Cl} \ : \ 2, 8, 7$$

Both fluorine and chlorine have 7 electrons in the valence shell.

Example 4.2: *Name two other elements which are in the same group as:*

(i) *Carbon* (ii) *Fluorine* (iii) *Sodium, respectively.*

Solution:

(i) The two other elements which are in the same group as carbon are: *silicon (Si)* and *germanium (Ge).*

(ii) The two other elements which are in the same group as fluorine are: *chlorine (Cl)* and *bromine (Br).*

(iii) The two other elements which are in the same group as sodium are: *lithium (Li)* and *potassium (K).*

4.2.2. Metals, Non-metals and Metalloids

In addition to the classification of elements into *s*-, *p*-, *d*- and *f*-blocks, it is possible to divide them into metals, non-metals and metalloids. More than 78% of the elements are metals. **Metals** are present on the left side and the centre of the periodic table. Metals are the elements which are malleable and ductile, possess lustre, are good conductors of heat and electricity and have high densities. Metals usually have high melting and boiling points, and are generally solids at room temperature. Mercury is the only metal which is liquid at room temperature. Gallium (303 K) and caesium (302 K) also have very low melting points.

Non-metals are much less in number than metals. There are only about 20 non-metals. Non-metals are located at the top right hand side of the Periodic Table. Non-metals have low melting and boiling points. They usually are solids or gases at room temperature. Non-metals are neither malleable nor ductile. They are poor conductors of heat and electricity. In a period, the non-metallic character increases as we move from left to right. In a group, the non-metallic character decreases and metallic character increases on going down a group. There is no sharp line dividing metals from non-metals. A zig-zag line separates metals from non-metals as shown in Fig. 4.3. The borderline elements such as silicon, germanium, arsenic, antimony and tellurium exhibit characteristic properties of metals as well as non-metals. These elements are called **semi-metals** or **metalloids**.

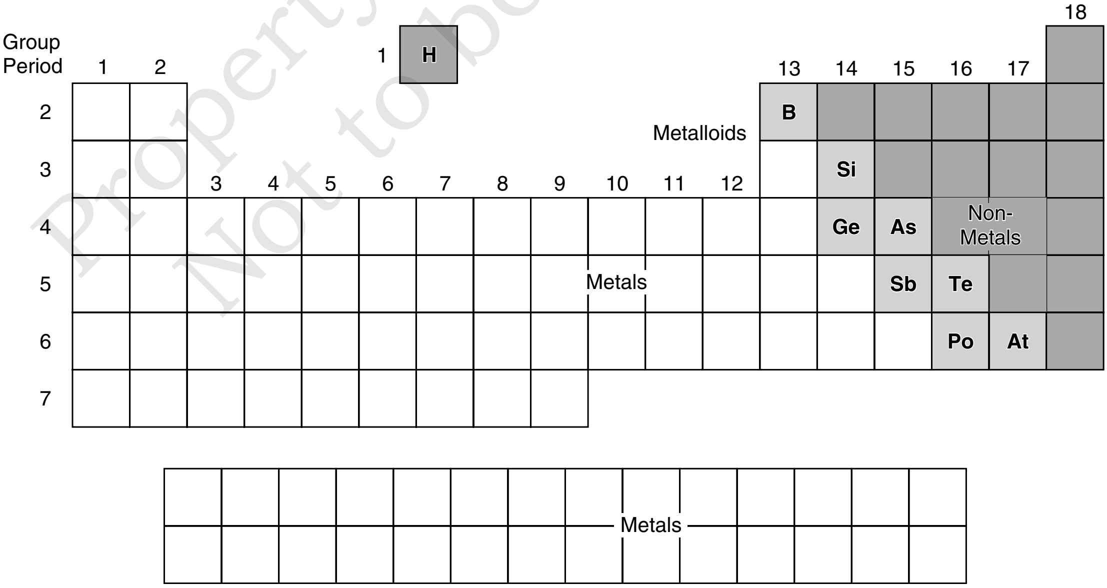

Fig. 4.3. Position of metals, non-metals and metalloids in the periodic table

4.3. TRENDS IN PERIODIC PROPERTIES

Most of the properties of the elements such as electronegativity, ionization energy, electron affinity, atomic radius, metallic character, etc., are directly related to the electronic configuration of the atoms. These properties undergo periodic variation with the change in the atomic number within a period or a group. These properties indirectly control the physical properties such as melting point, boiling point, density, etc. Let us now study the variation of some of the atomic properties in the periodic table.

4.3.1. Electronegativity

Electronegativity *may be defined as the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.*

The main factors on which the electronegativity depends are nuclear charge and atomic radius.

- *Greater the nuclear charge, greater is the electronegativity.*
- *Smaller the atomic radius, greater is the electronegativity.*

4.3.1.1. Variation across a Period

In a period, electronegativity increases in moving form left to right. This is due to the reason that nuclear charge increases whereas atomic radius decreases as we move from left to right in a period.

4.3.1.2. Variation down a Group

In a group, electronegativity decreases on moving down the group. This is due to the effect of increased atomic radius.

4.3.2. Ionization Energy

Ionization energy *may be defined as the amount of energy required to remove the most loosely bound electron from the isolated gaseous atom in its ground state.*

$$\mathbf{A(g) + Energy\ (I.E.) \longrightarrow A^{+}\ (g) + e^{-}}$$

Ionization energy is expressed in terms of *kilo joules per mole of atoms (kJ mol^{-1}).*

4.3.2.1. Variation in a Period

In general, *the ionization energy increases with the increase in atomic number across the period.* This can be attributed to the fact that moving across the period from left to right,

(i) nuclear charge increases regularly;
(ii) atomic size decreases.

Thus, due to the gradual increase in nuclear charge and simultaneous decrease in atomic size, the attractive force between the nucleus and the electron cloud increases.

4.3.2.2. Variation in a Group

The ionization energies of elements decrease regularly with the increase in atomic number within a group.

The decrease in the value of ionization energy within the group can be explained on the basis of net effect of the following factors:

As we move down the group there is:

(i) a gradual increase in the atomic size.
(ii) increase in the shielding effect on the outermost electron due to increase in the number of inner electrons.

Example 4.3: *From each set, choose the atom which has the largest ionization energy and explain your answer*

(i) *F, O, N* (ii) *Mg, P, Ar.*

Solution:

(i) F has the highest ionization energy among F, O and N because it has smallest size and highest nuclear charge. In general, ionization energy increases as we go from left to right in a period.
(ii) Ar (a noble gas) has the highest ionization energy among the elements Mg, P and Ar because it has stable electronic configuration and maximum nuclear charge.

4.3.3. Electron Affinity

The tendency of a gaseous atom to form anion is expressed in terms of *electron affinity*.

Electron affinity *may be defined as the energy change taking place when an isolated gaseous atom accepts an electron to form a monovalent gaseous anion.*

The values of electron affinity are expressed in *kilo joules per mole of atoms*. For example, electron affinity of chlorine is – 348 kJ mol^{-1}.

$$\mathbf{Cl(g) + e^- \longrightarrow Cl^-(g);\ E_{ea} = -\,348\ kJ\ mol^{-1}}$$

Depending on the element, the process of adding an electron can be either exothermic or endothermic.

The magnitude of electron affinity measures the tightness with which the atom can hold the additional electron. ***The large negative value of electron affinity reflects the greater tendency of an atom to accept the electron***.

4.3.3.1. Factors Affecting Electron Affinity

Following factors affect the electron affinity of an atom:

1. **Nuclear Charge.** Greater the magnitude of nuclear charge greater will be the attraction for the incoming electron and as a result, larger will be the negative value of electron affinity.
2. **Atomic Size.** Larger the size of an atom more will be the distance between the nucleus and the additional electron and smaller will be the negative value of electron affinity.
3. **Electronic Configuration.** Elements having the stable electronic configuration have tendency to accept the electron and larger will be the positive value of its electron affinity.

4.3.3.2. Variation across a Period

On moving across the period, *the atomic size decreases* and *nuclear charge increases*. Both these factors result into greater attraction for the incoming electron, therefore, *electron affinities tend to become more negative as we go from left to right across a period.*

4.3.3.3. Variation down a Group

On moving down a group, *the atomic size* as well as nuclear charge increases. But the effect of increase in atomic size is much more pronounced than that of nuclear charge. Consequently, *electron affinity becomes less negative on going down the group.*

4.3.4. Atomic Radius

The atomic size is very important property of the atoms because it is related to many other chemical and physical properties. In dealing with atomic size, the atom is assumed to be a sphere and its radius determines the size. In general, **atomic radius** is defined as ***the distance of closest approach, to another identical atom***.

4.3.4.1. Variation of Atomic Radii in the Periodic Table

Atomic radii usually depend upon *nuclear charge* and *number of main energy levels* of an atom. The periodic trends in atomic radii have been described as follows:

Variation in a Period

In general, *the atomic radii decrease with the increase in the atomic number in a period.* For example, atomic radii decrease from lithium to fluorine in second period.

The decrease of atomic radii along a period can be explained on the basis of nuclear charge. In moving from left to right across the period, the nuclear charge increases progressively by one unit but the additional electron goes to the same principal shell. As a result, the electron cloud is pulled closer to the nucleus by the increased effective nuclear charge. This causes the decrease of atomic size.

Variation in a Group

In general, *the atomic radii increase from top to bottom within a group of the periodic table.*

In moving down a group, the nuclear charge increases with increase in atomic number but at the same time, there is a progressive increase in the principal energy, shells. The number of electrons in the outermost shell, however, remains the same. Since, the effect of additional energy level is more pronounced than the effect of increased nuclear charge, therefore, *atomic size goes on increasing as we move down a group.*

Atomic raddi increase down the group. Atomic radii decrease across the period.

4.3.5. Electropositivity and Metallic Character

Tendency of atoms of an element to lose electrons and form positive ion is known as ***electropositivity****.*

A more electropositive element has more metallic character.

Whether an element behaves as a metal or a non-metal is directly related to its ionization energy. The elements having low values of ionization energies are metals whereas elements having high values of ionization energies are non-metals. The border line elements behave as metalloids.

4.3.5.1. Variation across a Period

Metallic character decreases across a period from left to right. On the other hand, non-metallic character increases with increase in atomic number across a period.

4.3.5.2. Variation down a Group

On going down a group from top to bottom, the metallic character of elements increases.

In each group, the first element is least metallic while the last element is most metallic.

Metals are located on the left hand side and the centre of the periodic table. Non-metals are located at the top right hand side of the periodic table. A zig-zag line separates metals from non-metals. The borderline elements such silicon, germanium, arsenic, antimony, etc., behave as metalloids. (Fig. 4.4)

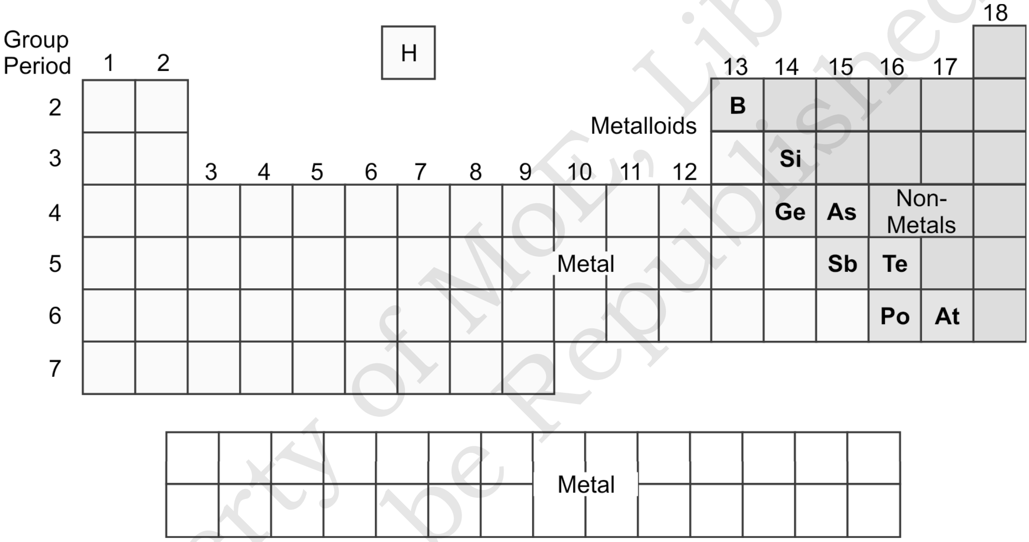

Fig. 4.4. Position of metals, non-metals and metalloids in the periodic table

4.3.6. Ionic to Covalent Bonding in Compounds

Because of the nature of ionic and covalent bonds, the materials produced by those bonds tend to have quite different macroscopic properties. The atoms of covalent materials are bound tightly to each other in stable molecules, but those molecules are generally not very strongly attracted to other molecules in the material. On the other hand, the atoms (ions) in ionic materials show strong attractions to other ions in their vicinity. This generally leads to low melting points for covalent solids, and high melting points for ionic solids. For example, the molecule carbon tetrachloride is a non-polar covalent molecule, CCl_4. It's melting point is –23°C. By contrast, the ionic solid NaCl has a melting point of 800°C.

You can anticipate some things about bonds from the positions of the constituents in the periodic table. Elements from opposite ends of the periodic table will generally form ionic bonds. They will have large differences in electronegativity and will usually form positive and negative ions. The elements with the largest electronegativities are in the upper right of the periodic table, and the elements with the smallest electronegativities are on the bottom left. If these extremes are combined, such as in RbF, the dissociation energy is large. As can be seen from the illustration below, hydrogen is the exception to that rule, forming covalent bonds.

Elements which are close together in electronegativity tend to form covalent bonds and can exist as stable free molecules. Carbon dioxide is a common example.

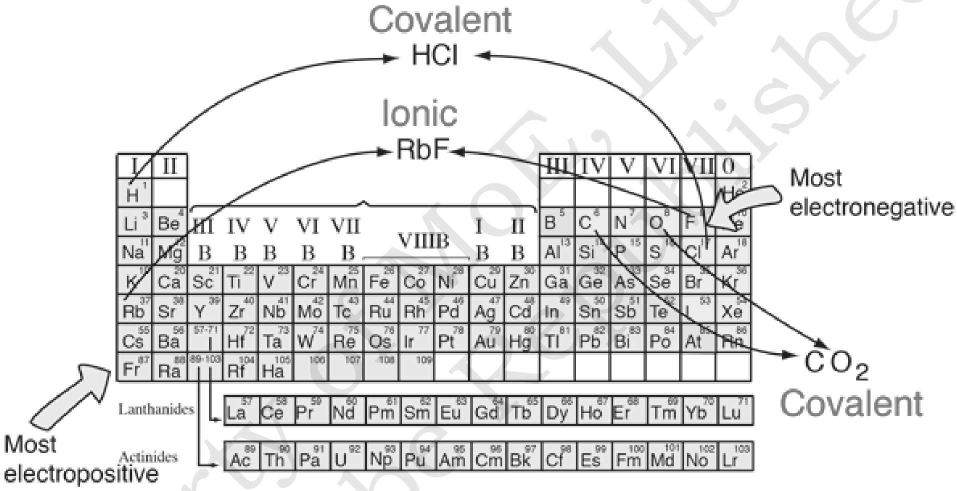

Fig. 4.5.

4.3.7. Non-Metallic Character

Non-metallic elements have strong tendency to gain electrons. Therefore, electronegativity is directly related to the non-metallic character of elements. We can also say that the electronegativity is inversely related to the metallic character of elements.

4.3.7.1. Variation across a Period

In a period, non-metallic character increases from left to right. This is due to the increase in electronegativities across a period.

4.3.7.2. Variation down a Period

In a group, non-metallic character decreases on moving down a group. This is due to the decrease in electronegativities down a group.

4.3.8. Lattice Energy

The lattice energy is the energy change occurring when one mole of a solid ionic compound forms in its gaseous state. It also refers to the energy required to disassociate one mole of a solid compound into its component gaseous ions. Lattice energy can be released (exothermic) or absorbed (endothermic) depending on whether the compound forms or disassociates.

General Reaction during Formation of Solid Ionic Compound:

$$A^+ + B^- \rightarrow AB + \text{Lattice energy}$$

General Reaction during Dissociation of Solid Ionic Compound:

$$AB\ (s) \rightarrow A^+(g) + B^-(g)$$

The following factors affect lattice energy:

1. **Ionic radius:** As the ionic radius increases, the lattice energy decreases. In other words, the bond between opposite ions is strongest when the ions are small. The following table shows the lattice energy values (in kJ/mol) for the ionic bond formed between alkali metals and halogens. It is clear that the bond between Li^+ and F^- (LiF) has the highest lattice energy and that between Cs^+ and I^- (CsI) has the lowest.

	F^-	Cl^-	Br^-	I^-
Li^+	1036	853	807	757
Na^+	923	787	747	704
K^+	821	715	682	649
Rb^+	785	689	660	630
Cs^+	740	659	631	604

2. **Ionic charge:** As the ionic charge increases, the lattice energy increases. In other words, the ionic bond becomes stronger as the charge on the ions becomes large. The lattice energy is proportional to the product of the two ionic charges. The following table shows the lattice energies for salts of OH^- and O^{2-}. It is clear that the bond between Na^+ and OH^- (NaOH) has the smallest lattice energy, and that between Al^{3+} and O^{2-} (Al_2O_3) has the greatest.

	OH^-	O^{2-}
Na^+	900	2481
Mg^{2+}	3006	3791
Al^{3+}	5627	15916

From the above tables, one can observe that the lattice energy increases with atomic charge and decreases with ionic radius. Across a period, the atomic charge increases, and down a group, the ionic radius increases. Hence, the lattice energy increases from left to right across a period and decreases from top to bottom down a group.

4.4. MAIN GROUP ELEMENTS

The main group elements are any of the chemical elements belonging to the *s* and *p* blocks of the periodic table. The *s*-block elements are group 1 (alkali metals) and group 2 (alkaline earth metals). The *p*-block elements are groups 13-18 (basic metals, metalloids, non-metals, halogens, and noble gases). The *s*-block elements usually have one oxidation state (+1 for group 1 and +2 for group 2). The *p*-block elements may have more than one oxidation state, but when this happens, the most common oxidation states are separated by two units. Specific examples of main group elements include helium, lithium, beryllium, boron, carbon, nitrogen, oxygen, and fluorine.

Let us discuss properties of some main group elements.

4.4.1. Helium (He)

Helium is a member of the noble gas family. The noble gases are the elements in Group 18 (VIIIA) of the periodic table. Helium was first discovered in the Sun. In 1868 Pierre Janssen (1824-1907), a French astronomer, studied light from the Sun during a solar eclipse. He found proof that a new element existed in the Sun. He called the element helium.

2
He
Helium
4.0026

4.4.1.1. Physical Properties of Helium

Following are some physical properties of helium:

- It is a colourless, odourless, tasteless and inert gas.
- It is completely inert (i.e.,it doesn't react with any other element).
- It has a very low boiling point (4.2 K).
- It is less dense than any other known gas except hydrogen.

4.4.1.2. Chemical Properties of Helium

Following are some chemical properties of helium:

- It is non-toxic.
- It is less soluble in water than any other gas.
- It is not flammable.

4.4.2. Lithium (Li)

Lithium is part of the Group 2 Alkaline Earth Metals. It is found in beryl and emerald, minerals that were known to the ancient Egyptians. It is widely distributed in earth's crust and is estimated to occur in Earth's igneous rocks to the extent of 0.0002 percent.

3
Li
Lithium
6.941

4.4.2.1. Physical Properties of Lithium

Following are some physical properties of lithium:

- Lithium is a silvery white metal.
- It is the lightest of all solid elements.
- It is a good conductor of heat and electricity.
- Its vapours impart calamine red colour to the flame.
- It gives alloys with number of metals and forms amalgam.

4.4.2.2. Chemical Properties of Lithium

Following are some chemical properties of lithium:

1. **Reaction with air:** Lithium is not affected by dry air but in moist air it is readily oxidized. When heated in air above 450K, it burns to give lithium monoxide and lithium nitride.

$$4\ Li + O_2 \rightarrow 2\ Li_2O$$
$$6\ Li + N_2 \rightarrow 2\ Li_3N$$

2. **Reactions with Water:** It decomposes cold water forming lithium hydroxide and hydrogen.

$$2\ Li + 2\ H_2O \rightarrow 2\ LiOH + H_2$$

4.4.3. Beryllium (Be)

Beryllium is part of the Group 2 Alkaline Earth Metals. It is found in nature and is combined with other elements in minerals, including beryl and chrysoberyl. In its purest form, beryllium is a steel-gray and lightweight alkaline earth metal.

4
Be
Beryllium
9.01218

4.4.3.1. Physical Properties of Beryllium

Following are some physical properties of beryllium:

- Beryllium is a silvery-white metal.
- It has low density.
- It is a non-magnetic, hard and brittle metal.
- It is a toxic element.

4.4.3.2. Chemical Properties of Beryllium

Following are some chemical properties of beryllium:

1. **Reaction with water:** Beryllium metal does not react with water or steam, even if the metal is heated to red heat.
2. **Reaction with acids:** The surface of beryllium metal is covered with a thin layer of oxide that helps protect the metal from attack by acids, but powdered beryllium metal dissolves readily in dilute acids such as sulphuric acid (H_2SO_4), hydrochloric acid (HCl), or nitric acid (NO_3), to form solutions containing the aquated Be(II) ion together with hydrogen gas (H_2).

$$Be(s) + H_2SO_4(aq) \rightarrow Be_2+(aq) + SO_4^{2-}(aq) + H_2(g)$$

4.4.4. Boron (B)

Boron is part of the Group 13 of the periodic table. Boron is not present in nature in elemental form. It is found combined in borax, boric acid, kernite, ulexite, colemanite and borates. Vulcanic spring waters sometime contains boric acids.

5
B
Boron
10.811

4.4.4.1. Physical Properties of Boron

Following are some physical properties of boron:

- It is a hard and black-coloured non-metallic solid.
- It has an unusually high melting point.
- At room temperature, it is a poor electrical conductor, but it is a good conductor at high temperatures.

4.4.4.2. Chemical Properties of Boron

Following are some chemical properties of boron:

1. **Reaction with air:** Boron do not react with oxygen at room temperature but at a higher temperature it reacts to form boron trioxide (B_2O_3).

$$4\ B(s) + 3\ O_2(g) \rightarrow 2\ B_2O_3(s)$$

2. **Reaction with water:** Boron does not react with water under normal conditions.
3. **Reaction with halogens:** When boron undergoes halogenation, the product formed is boron trihalides. The reaction with bromine (Br) is given below.

$$2B\ (s) + 3Cl_2(g) \rightarrow 2BCl_3(l)$$

4.4.5. Carbon (C)

Carbon is part of the Group 14 of the periodic table. It is the seventeenth most abundant element found on earth. It is found in the minerals of most metals in the form of carbonates.

6
C
Carbon
12.011

4.4.5.1. Physical Properties of Carbon

Following are some physical properties of carbon:

- Carbon is a unique element. It occurs in many forms. Some examples of the pure form of carbon are coal and soot.
- It is soft, dull grey or black non-metal.
- It occurs in a number of allotropic forms such as diamond and graphite.
- It is available in various shapes.

4.4.5.2. Chemical Properties of Carbon

The chemical properties of carbon are observed during the chemical reactions. Carbon takes part in four main reactions:

1. **By heating carbon in limited supply of oxygen:** Carbon monoxide is formed by incomplete combustion of carbon or carbon containing compounds (such as hydrocarbons) in the limited supply of oxygen.

$$C(s) + \tfrac{1}{2}\ O_2(g) \rightarrow \underset{\text{Carbon monoxide}}{CO(g)}$$

2. **Reaction with carbon dioxide:** When carbon reacts with carbon dioxide it produces carbon monoxide.

$$C(s) + CO_2(g) \rightarrow \underset{\text{Carbon monoxide}}{2\ CO\ (g)}$$

3. **Reaction with iron oxide:** Carbon is more reactive than iron, so it can displace iron from iron oxide.

$$2Fe_2O_3 + 3C \rightarrow 4Fe(l) + 3CO_2(g)$$

4. **Reaction with conc. H_2SO_4:** When carbon react with conc. sulphuric acid it produces carbon dioxide, sulphur dioxide and water.

$$C(s) + 2H_2SO_4\ (aq) \rightarrow CO_2\ (g) + 2SO_2\ (g) + 2H_2O(l)$$

4.4.6. Nitrogen (N)

7
N
Nitrogen
14.007

Nitrogen is part of the Group 15 of the periodic table. It is essential to life on Earth. It is a component of all proteins, and it can be found in all living systems. Nitrogen is crucial to life, but in excess it can also be harmful to the environment.

4.4.6.1. Physical Properties of Nitrogen

Following are some physical properties of nitrogen:

- It is colourless, tasteless and odourless gas.
- It is non-toxic in nature.
- It is almost insoluble in water (23.2° cm^3 per litre of water at 273 K and 1 bar pressure).
- Its freezing point and boiling point are 63 K and 77.2 K respectively.

4.4.6.2. Chemical Properties of Nitrogen

Following are some chemical properties of nitrogen:

1. **Reaction with highly electropositive metals like lithium, calcium magnesium etc.** These metals burn in the atmosphere of dinitrogen of form their respective nitrides

$$6Li + N_2 \rightarrow 2Li_3N$$
$$3Ca + N_2 \rightarrow Ca_3N_2$$
$$3Mg + N_2 \rightarrow Mg_3N_2$$

2. **Reaction with non-metal like dihydrogen and dioxygen:**

$$N_2 + 3H_2 \xrightarrow[\text{750 K pressure}]{\text{Fe/MO}} \underset{\text{Ammonia}}{2NH_3}$$

$$N_2 + O_2 \xrightarrow[\text{elertric arc}]{\text{3000 K}} \underset{\text{Nitric oxide}}{2NO}$$

3. **Reaction with compounds like Al_2O_3 and calcium carbide (CaC_2):**

$$\underset{\text{Alumina}}{Al_2O_3} + N_2 + 3C \xrightarrow{2100\,K} \underset{\text{Aluminium nitride}}{2AlN} + 3CO$$

$$\underset{\text{Calcium carbide}}{CaC_2 + N_2} \xrightarrow{1300\,K} \underset{\text{Calcium cyanamide}}{CaNCN} + \underset{\text{Coke}}{C}$$

Both AlN and $CaCN_2$ hydrolyse with boiling water to give ammonia.

4.4.7. Oxygen (O)

Oxygen is part of the Group 16 of the periodic table. It is the third most abundant element on the earth. About 50% of the earth's crust consists of oxygen. In free state it constitutes 21% of atmosphere by volume. In the combined state oxygen is present in water, minerals and the bodies of plants and animals. The human body has about 65% oxygen by weight. It is present in the bodies of animals and plants. Sand contains nearly 56% oxygen.

8
O
Oxygen
15.999

4.4.7.1. Physical Properties of Oxygen

Following are some physical properties of oxygen:

- Oxygen is a colourless, odourless and tasteless gas.
- Its density is higher than air.
- Oxygen is a very poor conductor of heat and electricity.
- Oxygen is soluble in some liquids such as water, alcohol, etc.

4.4.7.2. Chemical Properties of Oxygen

Oxygen is a highly reactive element. It is easily capable of combining with other elements. The most important chemical property of oxygen is that it supports combustion. It also combines with elements at room temperature, for example, the formation of rust. Decaying is an example of oxygen reacting with compounds. Carbon dioxide and water are the main products of decay.

1. **Reaction of Oxygen with Metals:** Metals react with oxygen to form basic oxides which dissolve in water to form basic oxide (alkalies) and turn red litmus blue. For example,

   Sodium burns brightly in oxygen with a golden yellow flame forming sodium oxide.

$$4Na + O_2 \rightarrow 2Na_2O$$

Sodium oxide thus formed dissolves in water forming NaOH (Sodium hydroxide) which turns red litmus blue.

2. **Reaction of Oxygen with Non-Metals:** Non-metals burn in oxygen to form acidic oxides. The acidic oxides dissolve in water and turn blue litmus red. For example,

   Burning sulphur combined with oxygen forms SO_2 gas which has a pungent and suffocating smell.

   $$S + O_2 \rightarrow SO_2$$

   SO_2 dissolves in water to form sulphurous acid which turns blue litmus red.

   $$SO_2 + H_2O \rightarrow H_2SO_3 \text{ (sulphurous acid)}$$

4.4.8. Fluorine (F)

Fluorine is part of the Group 17 of the periodic table. Fluorine occurs naturally in the crust of the earth where it is present in rocks, coal, and clay. Through wind-blown soil, fluorides are released into the air. Fluorine is the 13th most abundant element in the crust of the Earth.

9
F
Fluorine
18.998

4.4.8.1. Physical Properties of Fluorine

Following are some physical properties of fluorine:

- Fluorine can be found in nature as a gas.
- It is a light gas with a pale yellow colour and a faint smell.
- It is a flammable gas.

4.4.8.2. Chemical Properties of Fluorine

Following are some chemical properties of fluorine:

1. **Reaction of fluorine with hydrogen:** Fluorine reacts quickly with hydrogen to form hydrogen fluoride. The reaction can be explosive under the right conditions.

   $$H_2(g) + F_2(g) \rightarrow 2\ HF(g)$$

2. **Reaction of fluorine with metals/metal ions:** Fluorine reacts with sodium to form sodium fluoride.

   $$2\ Na(s) + F_2(g) \rightarrow 2\ NaF(s)$$

   The reaction with metals is a general reaction for most metals.

   Similarly Tellurium reacts with excess fluorine to form tellurium(VI) fluoride.

   $$Te(s) + 3\ F_2(g) \rightarrow TeF_6(s)$$

3. **Reaction of fluorine with noble gasses:** Krypton will react with fluorine, F_2, when cooled to –196 °C (liquid nitrogen) and zapped with an electric discharge or X-rays, forming krypton(II) fluoride (KrF_2).

$$Kr(s) + F_2(s) \rightarrow KrF_2(s)$$

This compound decomposes when heating to room temperature.

4.5. PERIOD THREE COMPOUNDS

4.5.1. Properties of Period three Compounds

Properties of some period 3 compounds are given below.

1. Hydrides

The hydrides of period 3 are given below:

Hydrides	NaH	MgH_2	AlH_3	SiH_4	PH_3	H_2S	HCl
Bonding	Ionic	Ionic/ Covalent	Covalent/ Polymer	Covalent	Covalent	Covalent	Covalent
Physical State	Solid	Solid	Solid	Gas	Gas	Gas	Gas

The hydrides of Na and Mg are ionic while that of other elements of period 3 are covalent. Aluminium hydride is polymeric $(AlH_3)_n$. Hydrides of Na, Mg and Al are solids while that of other elements of period 3 are gases. Argon does not form hydride.

NaH, MgH_2 and AlH_3 react vigorously with water and yield basic solutions.

$$NaH + H_2O \longrightarrow NaOH + H_2$$
$$MgH_2 + 2H_2O \longrightarrow Mg(OH)_2 + 2H_2$$
$$2AlH_3 + 3H_2O \longrightarrow 2Al(OH)_3 + 3H_2$$

SiH_4, PH_3 and H_2S are not much soluble in water and their aqueous solutions are almost neutral. HCl is highly soluble in water and its aqueous solution is strongly acidic. It neutralizes bases to form salts.

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

2. Oxides

The oxides of third period elements are given below in tabular form:

Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10} P_4O_6	SO_3 SO_2	Cl_2O_7 Cl_2O
Bonding	Ionic	Ionic	Ionic	Covalent	Covalent	Covalent	Covalent
Nature of Oxide	Alkaline	Alkaline	Alkaline	Acidic	Acidic	Acidic	Acidic

The oxides of metallic elements are ionic solids. SiO_2 is a giant covalent network solid. The remaining non-metallic oxides are covalent molecular compounds.

Acid-Base Behaviour of Some Period 3 Oxides

On moving from left to right among the oxides we observe that metal oxides form strongly basic oxides on the left-hand side to strongly acidic ones on the right, via an amphoteric oxide (aluminium oxide) in the middle. An *amphoteric oxide* is one which shows both acidic and basic properties.

Sodium oxide is a simple strongly basic oxide. Sodium oxide reacts exothermically with cold water to produce sodium hydroxide solution.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

As a strong base, sodium oxide also reacts with acids. For example, it would react with dilute hydrochloric acid to produce sodium chloride solution.

$$Na_2O + 2HCl \longrightarrow 2NaCl + H_2O$$

Magnesium oxide is not as strongly basic as sodium oxide.

Magnesium oxide has a slight reaction with the water to produce hydroxide ions in solution.

$$MgO + H_2O \longrightarrow Mg(OH)_2$$

Magnesium oxide reacts with warm dilute hydrochloric acid to give magnesium chloride solution.

$$MgO + 2HCl \longrightarrow MgCl_2 + H_2O$$

3. Hydroxides

The acid-base character of hydroxides of elements of period 3 changes from basic to amphoteric to acidic across the period from left to right.

Sodium hydroxide and magnesium hydroxide are basic.

Both react with acids to form salts. For example, with dilute hydrochloric acid, you get colourless solutions of sodium chloride or magenesium chloride.

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

$$Mg(OH)_2 + 2HCl \longrightarrow MgCl_2 + 2H_2O$$

Aluminium hydroxide is amphoteric.

Like sodium or magnesium hydroxides, it reacts with acids. It shows the basic nature of aluminium hydroxide.

$$Al(OH)_3 + 3HCl \longrightarrow AlCl_3 + 3H_2O$$

But aluminium hydroxide also reacts with sodium hydroxide solution

$$Al(OH)_3 + NaOH \longrightarrow NaAl(OH)_4$$

4. Chlorides

The chlorides of period 3 elements are:

Chlorides	NaCl	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_5 PCl_3	S_2Cl_2
Bonding	Ionic	Ionic	Ionic covalent	Covalent	Covalent	Covalent

Sodium chloride and magnesium chloride are ionic and consist of giant ionic lattices at room temperature. Aluminium chloride and phosphorus (V) chloride change their structure from ionic to covalent when the solid turns to a liquid or vapour.

The others are simple covalent molecules.

Sodium and magnesium chlorides are solids with high melting and boiling points because of the large amount of heat which is needed to break the strong ionic attractions.

The rest are low melting point solids or liquids.

Reactions with Water

The simple ionic chlorides (sodium and magnesium chloride) dissolve in water. The other chlorides all react with water in different ways to form a variety of products. The reaction with water is known as **hydrolysis**.

Solid aluminium chloride reacts with the water rather than just dissolving in it. In the first instance, hexaaquaaluminium ions are formed together with chloride ions.

$$AlCl_3(s) + 6H_2O(l) \longrightarrow [Al(H_2O)_6]^{3+}(aq) + 3Cl^-(aq)$$

Silicon tetrachloride is a colourless liquid at room temperature which fumes in moist air.

It fumes in moist air because it reacts with water in the air to produce hydrogen chloride.

$$SiCl_4 + 2H_2O \longrightarrow SiO_2 + 4HCl$$

4.5.2. Thermal stability of CO_3^{2-}, NO_{3-} of Li, Na, K, Mg and Ca

The trioxocarbonate(IV) of Group-1 metal are known to be relatively more stable to heat as compared to those of group-2 metals which are less stable.

Carbonates and Hydrogen Carbonates

The trioxocarbonate(IV) of alkali metals except lithium trioxocarbonate(IV) are stable to heat. The trioxocarbonate(IV) of group-2 metals *i.e.*, magnesium and calcium and that of lithium decompose on heating, forming an oxide along with the evolution of carbon dioxide. For example,

$$Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2$$

$$MgCO_3 \xrightarrow{\Delta} MgO + CO_2$$

$$Na_2CO_3 \xrightarrow{\Delta} \text{no effect.}$$

Reason. *The stability of carbonate towards heat depends upon the relative stability of the resulting metal oxide. More is the stability of the resulting metal oxide lesser is the stability of the carbonate towards heat and vice versa. Now the stability of resulting metal oxides decreases down the group due to decrease in lattice enthalpy, (because of bigger size) therefore the stability of carbonates towards heat increases.*

Li_2CO_3 decomposes on heating because it gives lithium oxide. The small O^{2-} anion is very strongly attracted to Li^+ ion resulting in an ionic compound with high lattice energy and therefore it is more stable than Li_2CO_3. Greater is the stability of resulting oxide more is its tendency of formation and hence lower is the thermal stability of the carbonate.

In period 3, magnesium trioxocarbonate(IV) decomposes more readily than sodium trioxocarbonate(IV). It is because the resulting MgO is more stable than $MgCO_3$ because of strong ionic bond between Mg^{2+} and O^{2-} ions. Thus its lattice energy is high and hence the stability. The stabilities of carbonates of group-2 metals increase on moving down the group. For example, $BeCO_3$ decomposes at 373 K, $MgCO_3$ at 813 K, $CaCO_3$ at 1173 K, $SrCO_3$ at 1563 K and $BeCO_3$ at 1633 K.

Lithium and group-2 metals do not form solid hydrogencarbonates, although they exist in solution. On heating these solutions, the hydrogencarbonates decompose to form carbonates and CO_2 gas is liberated. The solid hydrogencarbonates of alkali metals decompose between 375 and 575 K.

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

Nitrates

Trioxonitrate(V) of alkali metals (Na, K, etc.), except $LiNO_3$, decompose on strong heating forming nitrites and oxygen. For example,

$$2KNO_3(s) \xrightarrow{\text{heat}} 2KNO_2(s) + O_2(g)$$

Trioxonitrate(V) of Mg and Ca metals and $LiNO_3$ decompose on heating to form oxides, nitrogen dioxide and oxygen.

$$2LiNO_3(s) \xrightarrow{\Delta} Li_2O(s) + 2NO_2(g) + O_2(g)$$

$$2Ca(NO_3)_2(s) \xrightarrow{\Delta} 2CaO(s) + 4NO_2(g) + O_2(g)$$

Thermal stabilities of trioxonitrate(V) of group-1 *i.e.*, Li, Na, K, Rb, Cs, etc., and group-2 metals Be, Mg, Ca, Sr, Ba etc., increase on moving down the group from top to bottom.

EXPERIMENT 4.1

Objectives:

To demonstrate thermal stabilities of some trioxocarbonated(IV) in the laboratory.

Requirements:

The experiment requires, three test tubes, gas holder, lime water.

Procedure:

1. Take the three test tubes and to each of it add about 1g each of Li_2CO_3, Na_2CO_3 and $CuCO_3$.
2. To the mouth of the test tubes attach gas detector containing lime water. The simple experimental set up is shown in Fig. 4.6.

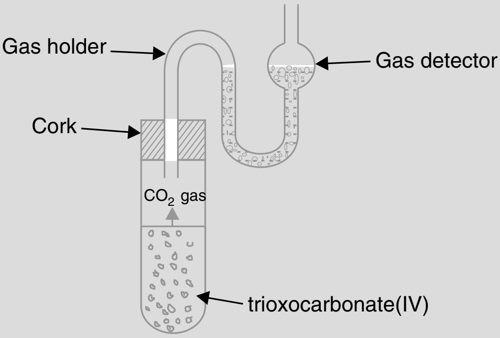

Fig. 4.6.

Observations:

The trioxocarbonate(IV)s of lithium and copper will decompose on heating and will turn lime water milky.

$$Li_2CO_3(s) \xrightarrow{\Delta} Li_2O(s) + CO_2(g)$$

$$CuCO_3(s) \xrightarrow{\Delta} CuO(s) + CO_2(g)$$

The trioxocarbonate of copper decomposes at much slower rate than that of lithium.

Na_2CO_3 will not decompose to produce carbon dioxide gas.

4.6. PERIOD FOUR METALS

Metals in the period 4 are K, Ca, Sc, Ti, etc. In this section, we will discuss about K and Ca.

4.6.1. Potassium

Potassium is the first element in the periodic table's fourth period. The name potassium is derived from the mineral Potash. For hundreds of years, the element has been used. It, along with lithium, rubidium, sodium, caesium, and francium, is an alkali metal. Potassium has an atomic mass of 39.098 atomic mass units. It is represented by the letter 'K.'

Physical Properties of Potassium

Following are some physical properties of potassium:

- Potassium is an alkali metal.
- It is a highly reactive element and does not occur in a free state.
- It is a soft, silvery-white metal.

- Potassium has a density less than that of water (0.89 g/cm^3). Hence, it can float on the water surface.
- It is malleable in nature.
- Potassium has a melting point of 63.5 °C and a boiling point of 759 °C.

Chemical Properties of Potassium

Following are some chemical properties of potassium:

- It gives out hydrogen gas when reacts with water. The reaction is volatile and can cause an explosion.

$$2K + 2H_2O \rightarrow 2KOH + H_2\uparrow$$

- It is highly reactive with nitrogen, phosphorous, sulphur, and fluorine.
- It rapidly gets dissolved when reacted with dilute sulphuric acid and gives up potassium ions along with hydrogen gas.

$$2K + H_2SO_4 \rightarrow 2K^+ + SO_4 + H_2\uparrow$$

- Potassium forms potassium halides when gets reacted with halogens.

$$2K + Cl_2 \rightarrow 2KCl$$

- Potassium forms potassium halides when gets reacted with halogens.

$$2K + Cl_2 \rightarrow 2KCl$$

Uses of Potassium

Potassium is widely used in our day to day life and some of which are mentioned below:

- Industries use potassium for making soaps, detergents, dyes, gunpowder, etc.
- Potassium is used for muscle contraction.
- Excess potassium diet helps to reduce blood pressure and prevents heart strokes.
- Potassium carbonate is used for the production of glass.
- It has a high demand for fertilizers.
- It can also be used as a medium of heat exchange and is used in nuclear power plants.

4.6.2. Calcium

Calcium is the second element in the fourth period of the periodic table. Calcium (Ca) is an essential mineral that helps our bones stay strong and capable of bearing our weight. Calcium is also employed by our nervous system to aid in the transmission of impulses throughout our bodies.

Physical Properties of Calcium

Following are some physical properties of calcium:

- Calcium doesn't occur naturally in the free-state.
- It is used as an alloying agent for aluminum, lead, copper, and other base metals.
- It is a form of soft metal.
- Calcium is a good conductor of electricity.
- It's malleable and ductile in nature.

Chemical Properties of Calcium

Following are some chemical properties of calcium:

- The dissolved form of calcium bicarbonate is found in hard water.

$$CaCO_3 + CO_2 \rightarrow Ca(HCO_3)_2 + H_2O$$

- When calcium comes in contact with air, it forms a coating of nitride and oxide.

$$2Ca + O_2 \rightarrow 2CaO$$

$$3Ca + N_2 \rightarrow Ca_3N_2$$

- Compounds of calcium are highly reactive to acids.

$$CaCO_3 + HCl \rightarrow CaCl_2 + HO + CO_2\uparrow$$

Uses of Calcium

Calcium can be used for many purposes. Some of them are mentioned below:

- Calcium helps to maintain strong bones to perform many necessary functions.
- It is needed for nerves to carry messages between the brain and every body part.
- It can be used as a reducing agent in the metal extraction process.
- It is also used as an alloying agent for the production of some metals.
- Calcium concatenate is used as a food additive.
- Calcium carbide is used for the production of plastics and acetylene gas.

4.7. GROUP 7 ELEMENTS

The Group 7 elements are placed in the vertical column, second from the right-hand side of the periodic table. All Group 7 elements have 7 electrons in their outer shell. Fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At), all belong to Group 7. All these elements are called halogens. Let us discuss about these in detail.

4.7.1. Physical Properties of Group 7 Elements

Following are some physical properties of group 7 elements:

1. **Physical State.** The halogens are all diatomic and exist as F_2, Cl_2, Br_2 and I_2. The intermolecular forces are very weak in halogens. The nature of forces is van der Waals' and their magnitude increases down the group.

   Thus, F_2 and Cl_2 are gases, bromine is a volatile liquid and iodine is a volatile solid.

2. **Colour.** Halogens are coloured.

   The colour of different halogens are as follows:

Halogen	*Fluorine*	*Chlorine*	*Bromine*	*Iodine*
Colour	Pale yellow	Greenish yellow	Reddish orange	Dark violet

   Thus, the colour deepens down the group.

3. **Melting and Boiling Points.** Melting and boiling points increase with increase in atomic number. This indicates that the strength of intermolecular forces of attraction between the molecules increases with the increase in atomic number.

   The intermolecular forces in halogens are van der Waal's forces which increase with the size of the molecule.

4. **Ionization Energies.** Ionization energies of all the halogens are very high. Therefore, they have a less tendency to lose electron. However, this tendency increases down the group.

5. **Electronegativity.** The halogens have very high electronegativity. Halogens are the most electronegative elements in their respective periods. Electronegativity decreases on descending the group. Fluorine is the most electronegative element in the Periodic Table.

6. **Non-metallic Character.** All the halogens have very high values of ionization energies and exhibit non-metallic character. The non-metallic character, however, decreases down the group.

Iodine shows some distinct metallic properties *e.g.*, it possesses metallic lustre and forms positive ions like I^+, I^{3+}, etc.

4.7.2. Chemical Properties of Group 7 Elements

Following are some properties of Group 7 elements:

1. **Oxidizing Power:** Elements of this group are great oxidizing agents. *Fluorine* can oxidize all halide particles to halogen in a solution. However, oxidizing power decreases as we move down the group. *Chlorine* can oxidize bromide to bromine and iodide to iodine.

$$Cl_2 + 2Br^- \rightarrow Br_2 + 2Cl^-$$

$$Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-$$

Bromine can oxidize iodide to iodine.

$$Br_2 + 2I^- \rightarrow I + 2Br^-$$

2. **Reaction with Hydrogen:** Acidic hydrogen halides are formed when halides react with hydrogen. The reactivity of halogen towards halogen decreases as we move down group 17. Therefore, their acidity also decreases as we move down the group.

In dark: $H_2 + F_2 \rightarrow 2HF$

In sunlight: $H_2 + Cl_2 \rightarrow 2HCl$

3. **Reaction with Metals:** Halogens react with metals instantly due to their high reactivity to form metal halides.

Sodium reacts with chlorine to form sodium chloride which releases a large amount of heat energy and yellow light as it is an exothermic reaction.

$$2Na(s) + Cl_2\,(g) \rightarrow 2NaCl(s)$$

Metal halides are ionic in nature due to the high electronegativity of halogen and electro positivity of metals. The ionic character decreases down the group.

4.7.3. Uses of Group 7 Elements

Uses of some group 7 elements are discussed below:

Fluorine

It is mainly used for the manufacture of UF_6 for nuclear power generation. It is also used for the preparation of many fluorinating agents. The important organic chemicals derived from fluorine are the *chlorofluorocarbons and polytetrafluoroethylene* (teflon). Chlorofluorocarbons known as *freons* are used as refrigerants and in aerosols. Teflon is a plastic which is not attacked by chemical reagents and is heat-resistant. It is used for many special applications. An important inorganic chemical made from HF is cryolite (Na_3AlF_6) which is used for the production of aluminium. Some of the uses of HF are in the glass industry as an *etching agent* and in the manufacture of fluoride salts. Prominent among the fluorides is NaF used for the fluorination of water; one part per million level fluoride in drinking water prevents tooth decay.

Chlorine

The chief uses of chlorine are:

(i) In the production of organic compounds like polyvinyl chloride, chlorinated hydrocarbons, pharmaceuticals, herbicides, pesticides, etc.

(ii) It is used to bleach paper pulp and textiles and as disinfectant for sterilizing drinking water.

(iii) It is used in the preparation of solvents such as CCl_4, $CHCl_3$, CH_2Cl_2 and trichloroethene.

(iv) In the production of inorganic compounds like HCl, PCl_3, PCl_5 sodium hypochlorite (NaOCl), bleaching powder ($CaOCl_2$) etc.

Bromine

Some important uses of bromine are:

(i) Bromine is used in the preparation of ethylene bromide which is used as an additive to *leaded* petrol.

(ii) Bromine is used to make AgBr for photography.

(iii) It is also used in the preparation of dyes, fire retardants disinfectants, fumigants and medicines.

Iodine

Some important uses of iodine are:

(i) Iodine is used as an anticeptic in the form of an alcoholic solution, which is known as *tincture of iodine*.

(ii) It is also used for the preparation of iodoform and potassium iodide.

(iii) Iodide ion is necessary for the normal functioning of the thyroid gland. Insufficient iodide in the diet leads to *goitre* (enlargement of thyroid gland). Hence, sodium or potassium iodide is added to table salt and this type of salt is known as "*iodized*" salt.

4.8. ELEMENTS OF THE FIRST TRANSITION SERIES

In general, any element which corresponds to the *d-block* of the modern periodic table (which consists of groups 3-12) is considered to be a transition element. The elements of the first transition series are scandium (Sc), titanium (Ti), vanadium (V), chromium (Cr), manganese (Mn), iron (Fe), cobalt (Co), nickel (Ni), copper (Cu), and zinc (Zn).

4.8.1. Properties of the Elements of the First Transition Series

Following are some common characteristics of elements of the first transition series:

- **Physical States:** Except for mercury, all transition elements exist in solid state.
- **Metallic Property:** All the transition elements are metals; this is because the number of electrons in outermost shell is only 2.
- **Magnetic Property:** Most of the transition metals are paramagnetic in nature. This is due to the presence of unpaired electrons in the transition elements.
- **Variable Oxidation State:** Transition metals can form compounds with a wide range of oxidation states. Some of the observed oxidation states of the elements of the first transition series are shown in Table. 4.3.

Table 4.3. Oxidation States of Elements of the First Transition Series

^{21}SC	^{22}Ti	^{23}V	^{24}Cr	^{25}Mn	^{26}Fe	^{27}Co	^{28}Ni	^{29}Cu	^{30}Zn
								1+	
		2+	2+	2+	2+	2+	2+	2+	2+
3+	3+	3+	3+	3+	3+	3+	3+	3+	
	4+	4+	4+	4+					
		5+							
			6+	6+	6+				
				7+					

As we move from left to right across the first transition series, we see that the number of common oxidation states increases at first to a maximum towards the middle of the table, then decreases.

- **Formation of Coloured Compounds:** Many compounds of transition elements are coloured in contrasts to those of *s* and *p* block elements. In compound state due to the surrounding groups (ligands), the *d*-orbitals of transition elements are not degenerate but split into two groups of different energy.
- **Catalytic Property:** Many transition metals and their compounds have catalytic properties. For example, V_2O_5, Fe, $FeCl_3$, Ni, Pd, etc. This is due to following reasons:
  - (i) *Variable oxidation state*: Due to variable oxidation state they form unstable intermediate compounds and provide a new path with lower activation energy for the reaction (Intermediate compound formation theory).
  - (ii) *Large Surface area*: Finely divided transition metals or their compounds provide a large surface area for adsorption and the adsorbed reactants react faster due to the closer contact.
- **Formation of Alloys:** Alloys are homogenous solid solutions of two or more metals obtained by melting the components and then cooling the melt. These are formed by metals whose atomic radii differ by not more than 15% so that the atoms of one metal can easily take up the positions in the crystal lattice of the other. Since transition metals have similar atomic radii, they form alloys very readily.

GLOSSARY

- **Covalent Radius:** Half of the inter-nuclear distance between two atoms of the element held by a single covalent bond.
- **Electron Affinity:** The energy change taking place when an electron is added to an isolated gaseous atom of the element.
- **Electronegativity:** It is the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.
- **Group:** A vertical column of elements in the periodic table.
- **Ionization Energy:** The energy required to remove the outermost electron from an isolated gaseous atom of the element.
- **Metallic Radius:** Half of the inter-nuclear distance between two nearest atoms in the metallic lattice.
- **Period:** A horizontal row of elements in the periodic table.

- **Periodic Table:** Arrangement of elements in the increasing order of atomic number such that elements with similar properties fall under same vertical column.
- **Triads:** The elements arranged according to their increasing atomic mass, in a group of three.

SUMMARY

- The periodic table is the arrangement of all the known elements according to their properties so that similar elements fall within the same vertical column and dissimilar elements are separated.
- According to the law of octaves, *"If the elements are arranged in the order of increasing atomic masses, the eighth element, starting from a given one is a kind of repetition of the first—like the eighth note in an octave of music."*
- The modern periodic law states that: "The physical and chemical properties of the elements are the periodic function of their atomic numbers."
- Modern periodic table arranges the elements in the order of their atomic numbers. There are 118 elements in the Modern periodic table.
- A horizontal row of a periodic table is called a period. There are seven periods in all, which are numbered as 1, 2, 3, 4, 5, 6 and 7.
- A vertical column of elements in the periodic table is called a group. There are 18 groups in the modern periodic table.
- Electronegativity is the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.
- Electron affinity is the energy change taking place when an isolated gaseous atom accepts an electron to form a monovalent gaseous anion.
- The lattice energy is the energy change occurring when one mole of a solid ionic compound forms in its gaseous state.
- The main group elements are any of the chemical elements belonging to the *s* and *p* blocks of the periodic table. Specific examples of main group elements include helium, lithium, beryllium, boron, carbon, nitrogen, oxygen, and fluorine.
- Group 7 elements are called halogens.
- Any element which corresponds to the *d*-block of the modern periodic table (which consists of groups 3-12) is considered to be a transition element.

EVALUATION

I. Multiple Choice Questions

1. In Mendeleev's periodic table, the elements are arranged according to their

(a) Atomic size
(b) Atomic number
(c) Atomic mass
(d) None of these

2. In the modern periodic table the elements are arranged according to their

(a) Atomic size
(b) Atomic number
(c) Atomic mass
(d) None of these

3. Elements of Group-2 are called

(a) Alkali metals
(b) Alkaline earth metals
(c) Atomic mass
(d) None of these

4. Elements of which group are called chalcogens?

(a) Group-1
(b) Group-2
(c) Group-16
(d) Group-17

5. The tendency of gaseous atom to form anion is expressed in terms of

(a) Ionization energy
(b) Electron affinity
(c) Electronegativity
(d) None of these

6. On moving down a group, the atomic size

(a) increases
(b) decreases
(c) becomes stable
(d) None of these

7. On moving across a period from left to right, the metallic character

(a) increases
(b) decreases
(c) becomes stable
(d) None of these

8. The energy change occurring when one mole of a solid ionic compound forms in its gaseous state.

(a) Ionization energy
(b) Lattice energy
(c) Electronegativity
(d) None of these

9. Which of the following is not a Group-7 element?

(a) Fluorine
(b) Aluminium
(c) Bromine
(d) Iodine

II. Fill in the Blanks

1. A horizontal row of periodic table is known as _______ .
2. The elements in a _______ have same valence shell.
3. A vertical column of elements in periodic table is known as _______ .
4. The elements in a _______ have same outer shell configuration.
5. Group 1 elements are known as _______ .
6. Group 17 elements are known as _______ .
7. Noble gases are the elements belonging to group _______ .
8. The elements of group 16 are known as _______ .
9. Alkali metals occur at the _______ in Lothar Meyer's atomic volume curve.
10. Transition metals belong to _______ block of the periodic table.

III. Answer the Following Questions

1. What is a group and a period of periodic table?
2. What is the number of groups and number of periods in the long form of periodic table?
3. What property did Mendeleev use to arrange elements in the periodic table?
4. State the periodic law on which the modern periodic table based. Who proposed this law?
5. What are the common features of the electron structures of elements in: (i) Group 2 (ii) Period 3.
6. The atomic number of carbon is 6 and that of silicon is 14. Write the electronic configurations of carbon and silicon atoms. In which group or groups do these elements occur?
7. On the basis of electronic configuration how will you identify:
   - (i) chemically similar elements
   - (ii) the first element of a period
   - (iii) the last element of a period.
8. Consider the following elements:

   Ca, Na, Mg, Al, Be, Si
   - (i) Which of these elements belong to the same period?
   - (ii) Which of these elements belong to the same group?
9. Define electronegativity. How does it vary along a period and along a group?
10. What are the elements of the first transition series? Write their common properties.