
UNIT 1 PERIODIC TABLE AND PERIODIC PROPERTIES

Structure

- 1.1 Introduction
 - Objectives
- 1.2 Development of Periodic Table
- 1.3 Periodic Table and Electronic Configuration of Elements
- 1.4 Periodic Properties
 - 1.4.1 Valence
 - 1.4.2 Atomic Radii
 - 1.4.3 Ionic Radii
 - 1.4.4 Ionisation Energy
 - 1.4.5 Metallic Character
- 1.5 Summary
- 1.6 Answers to SAQs

1.1 INTRODUCTION

Today, more than 110 elements are known. These elements form a large number of compounds. Thus, a huge amount of information and data exists for these elements and compounds. An understanding and analysis of these data is very tedious work if one has to study each element and its compounds. The chemistry of elements and their compounds could be conveniently understood if the vast information available can be organized in a systematic way. In eighteenth and nineteenth centuries, many such attempts were made by chemists to classify the elements about which you will study in this unit. After describing these, we would also explain how all the elements known today are arranged in the form of a table, called **modern periodic table**. This tabular arrangements helps us to study various properties of elements such as *valence*, *atomic sizes*, *ionization energies*, *metallic character*, *melting* and *boiling points* as well as their trends in the periodic table.

Objectives

After reading this unit, you should be able to

- explain the term triad,
- state the law of octaves,
- state periodic law and modern periodic law,
- discuss the main features of the periodic table,
- define periodic properties, and
- describe the variation of periodic properties such as valence, atomic radii, ionic radii, ionisation energy and metallic character down a group and across a period in the periodic table.

1.2 DEVELOPMENT OF PERIODIC TABLE

The first attempt to classify the elements on the basis of their similarities was done by a German chemist Johann Dobereiner (1780-1849). In 1817, Dobereiner found that there are several groups of three elements which have similar properties. Such a set of three elements was called a **triad**. For example, calcium, strontium and barium had similar properties and they formed a *triad*. Dobereiner also noticed that the atomic weight of the middle element of a triad was nearly equal to the arithmetic mean of the atomic weights of the other two elements. Thus, the atomic weight of strontium is 88 which is nearly equal to the arithmetic mean (88.5) of atomic weights of calcium and barium which are 40 and 137, respectively.

Johann Dobereiner

Arithmetic mean of atomic masses of calcium and barium = $\frac{40 + 137}{2} = 88.5$

Similarly, lithium (Li), sodium (Na) and potassium (K) constituted another set of triad.

Element	Li	Na	K
Atomic Weight	7	23	39

Arithmetic mean of atomic weights of Li and K is

$$= \frac{7 + 39}{2} = \frac{46}{2} = 23$$

which is equal to the atomic weight of sodium (Na).

The halogens chlorine (Cl), bromine (Br) and iodine (I) also formed a triad as is shown below :

Element	Cl	Br	I
Atomic Weight	35.5	80	127

Arithmetic mean of atomic masses of chlorine and iodine is

$$\frac{35.5 + 127}{2} = 81.5$$

This value is nearly the same as the atomic weight of bromine which is 80.

Such a grouping in the form of triads could not be done for other known elements. Therefore, this classification had a very limited acceptance.

The next notable attempt was that of an English Chemist John Newlands (1837-1898) in 1864. He arranged the 62 elements known at that time in the

increasing order of their atomic weights and noted that every eighth element had properties similar to the first one. This was called as the **Law of Octaves** because of its similarity with musical notes in which the eighth note is similar to the first one in the octave. However, his idea was not accepted by many at that time.

John Newlands

In 1869, two chemists – the German Julius Lothar Meyer (1830-1895) and the Russian Dmitri Ivanovich Mendeleev (1834-1907) – independently proposed nearly similar schemes of classification. Mendeleev arranged the known elements according to their increasing order of atomic weights in the form of a table, known as **Mendeleev's periodic table**. It is shown in Figure 1.1.

Figure 1.1 : Mendeleev's Early Periodic Table which was Published in 1872

He stated the **Periodic Law** as follows :

The properties of the elements are a periodic function of their atomic weights.

Usually, Mendeleev is given most of the credit for this design of periodic table because he left some spaces for those elements which were not known at that time. Mendeleev, however, was able to predict the existence and properties of these elements. For example, elements gallium and germanium were not known at that time. Mendeleev called them **ekaaluminium** and **ekasilicon** (below silicon), respectively and predicted their properties. Later, when these elements were

Dmitri Mendeleev

discovered, their actual properties were found to be in agreement to those predicted by Mendeleev. Such a comparison is shown for ekasilicon and germanium in Table 1.1.

Table 1.1 : Properties of Ekasilicon and Germanium

Property	Predicted Property for Ekasilicon (E)	Observed Property for Germanium
Atomic weight	72	72 – 59
Density (g cm^{-3})	5.5	5.35
Specific heat ($\text{J g}^{-1} \text{K}^{-1}$)	0.305	0.309
Colour	dark grey	greyish-white
Melting point ($^{\circ}\text{C}$)	high	937
Oxide	EO_2 – White solid – amphoteric	GeO_2 – White solid – amphoteric
Chloride	ECl_4 – density 1.9 g cm^{-3} – b. pt. below 100°C	GeCl_4 – density 1.84 g cm^{-3} – b. pt. below 84°C

With the help of the periodic table, Mendeleev was also able to arrive at the correct atomic weights of indium (In), beryllium (Be) and uranium (U).

In spite of these positive features of Mendeleev's periodic table, there were certain anomalies. For example, an element having higher atomic weight was placed before the element having lower atomic weight. One such pair was potassium and argon where argon having higher atomic weight was placed before potassium which had lower atomic weight.

In 1930, Henry Moosley (1887-1915) gave the concept of atomic numbers. He stated that the atomic number is a more fundamental property than the atomic weight. When Moosley arranged the elements according to the increasing atomic numbers instead of atomic weights, some of the inconsistencies associated with

Henry Moosley

the periodic table were eliminated. In the light of this, the periodic law was modified as follows :

The properties of elements are a periodic function of their atomic numbers.

This is called the **modern periodic law**. The periodic table based on modern periodic law is known as the **modern periodic table**. The original form of periodic table has been modified because of discovery of noble gases and other elements from time to time. Many versions of periodic table are available. One such form is depicted in Figure 1.2.

Figure 1.2 : Periodic Table

However, the form of the Periodic Table which has been recommended by IUPAC has been given in Figure 1.3.

Note that in the Periodic Table, there are 18 vertical columns called **Groups** and 7 rows called **Periods**. The groups are numbered 1 to 18 from left to right whereas the periods are numbered 1 to 7 from top to bottom.

The elements in Groups 1 and 2 and 13 to 17 are called **normal elements, main group elements** or **representative elements**. The elements in Groups 3 to 12 are called **transition elements**. However, Group 18 constitutes the group of **noble gases**.

Figure 1.3 : Periodic Table as Recommended by IUPAC

The elements having atomic numbers 58 to 71 are called **lanthanides** and those with atomic numbers 90 to 103 are **actinides**. These are respectively placed alongwith lanthanum (Ln, atomic number 57) and actinium (Ac, atomic number 89), respectively.

The lanthanides and actinides are together known as **inner transition elements** and these are separately shown below the periodic table.

Note that the first period is the **shortest** period and it contains only two elements. The second and third periods are called **short** periods and both of them contain eight elements each. The next two periods (*viz.* fourth and fifth) are the **long periods** containing 18 elements each. The sixth and seventh periods are called **very long periods** and can have 32 elements each.

1.3 PERIODIC TABLE AND ELECTRONIC CONFIGURATION OF ELEMENTS

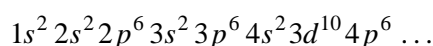
The elements are placed in the periodic table in the increasing order of their atomic number. Thus, as the atomic number (number of protons) increases, the number of electrons also increases. Hence, each element contains one more electron than the preceding element. These electrons occupy various orbitals constituting the shells in the increasing order of their energy.

There are four types of orbitals, namely, *s*, *p*, *d* and *f* and each orbital can accommodate two electrons. There is *one s* orbital, *three p* orbitals, *five d* orbitals and *seven f* orbitals which can be filled by the electrons according to the definite order of energy. Further, each shell or principal energy level can have a definite number of orbitals as shown in Table 1.2.

Table 1.2 : Maximum Numbers of Electrons in Different Energy Shells

No. of Energy Shell	Maximum No. of Electrons
First (Energy Level) Shell <i>1s</i>	2
Second (Energy Level) Shell <i>2s, 2p</i>	$2 + 6 = 8$
Third Shell <i>3s, 3p, 3d</i>	$2 + 6 + 10 = 18$
Fourth Shell <i>4s, 4p, 4d, 4f</i>	$1 + 6 + 10 + 14 = 32$

Thus, the first shell can accommodate only *two* electrons whereas the second, third and fourth shells can have eight, eighteen and thirty-two electrons, respectively. The electrons are filled in various elements according to the following sequence of orbitals.



Let us now understand how electrons are filled in various orbitals in different elements. This representation of electrons in various orbitals is called the **electronic configuration**. Hydrogen, which has only one electron, has the electronic configuration $1s^1$ as the electron occupies the *s* - orbital of first shell

which is the lowest energy orbital available. The symbol $1s^1$ can be interpreted as follows :

Principal Energy Level
(Shell)



Number of Electrons in
the particular Orbital



The next element helium has two electrons in its atom and the second electron also goes to $1s$ orbital. Thus, the electronic configuration of helium (He) is $1s^2$.

Similarly, we can write the electronic configurations of other elements as in Table 1.3.

Table 1.3 : Electronic Configurations of some Elements

Element	Atomic No.	Electronic Configuration
Li	3	$1s^2 2s^1$
Be	4	$1s^2 2s^2$
B	5	$1s^2 2s^2 2p^1$
C	6	$1s^2 2s^2 2p^2$
N	7	$1s^2 2s^2 2p^3$
O	8	$1s^2 2s^2 2p^4$
F	9	$1s^2 2s^2 2p^5$
Ne	10	$1s^2 2s^2 2p^6$

You can see from above that in case of both **hydrogen** and **helium**, the outermost electron goes to the s orbital. Therefore, they are called **s -block elements**.

Similarly, in case of elements from Li to Ne, the outermost electrons go to p -orbitals and hence these elements are known as **p -block elements**.

Similarly, we can write the electronic configurations of elements having atomic numbers 11 (Na) to 18 (Ar), 19 (K) and 20 (Ca) and classify them as **s -block** or **p -block elements**.

In case of Scandium (atomic number 21), the outermost electron goes to the $3d$ orbital. The filling of $3d$ orbital continues up to Zinc (Zn, atomic number 30). Thus, these elements form part of **d -block elements**. Similarly, when $4f$ orbital starts filling from Cerium (Ce, atomic number 58), there is a beginning of **f -block elements**. These elements correspond to atomic numbers 58 to 71 and 90 (Thorium) to 103 (Lr, Lawrencium). We can represent these blocks in the periodic table as given in Figure 1.4.

Figure 1.4 : Periodic Table Showing Various Blocks and Electronic Configurations of Elements

Why do not you try answering the following SAQs?

SAQ 1

- (a) How many elements are present in the third period?
- (b) Write the electronic configuration of the following elements :

Elements	O	P	Cl	Ca	Fe
At. Nos.	8	15	17	20	26

- (c) Write the blocks of the periodic table to which the following elements belong :
- Na
 - S
 - Co
 - Br
 - A
 - Hg
 - U

1.4 PERIODIC PROPERTIES

According to the modern periodic law, **the properties of elements are a periodic function of their atomic numbers**. These atomic numbers in turn give us the **electronic configuration** as you have studied in the previous section. The elements in a particular **group** in periodic table show **similar** electronic configurations. For example, the elements of Group 1 have the following electronic configurations as given in Table 1.4.

Table 1.4 : Electronic Configurations of Group 1 Elements

Elements	Atomic Nos.	Electronic Configuration
Li	3	$1s^2 2s^1$
Na	11	$1s^2 2s^2 2p^6 3s^1$
K	19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Rb	37	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$
Cs	55	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$
Fr	87	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^1$

Note the similarity in these electronic configurations. All the elements have their outermost electron filling the s orbital leading to the s^1 configuration. While we move down, every time a new shell is added and the number prefixed to s^1 denotes that number.

This similarity in electronic configurations is also visible in other groups of elements and is the basis of the similarity in their properties. This is so because physical and chemical properties are governed by the electrons in the outermost shell. The electrons present in the inner shell do not participate in the chemical reactions.

In each period, a new shell starts filling. As we move from left to right, one additional electron enters into the same shell. The number of the period gives the number of the shell into which the electrons get accommodated one by one. This is illustrated in Table 1.5 for third period.

Table 1.5 : Electronic Configurations of the Elements of the Third Period

Element	Symbol	Electronic Configuration
Sodium	Na	$1s^2 2s^2 2p^6 3s^1$
Magnesium	Mg	$1s^2 2s^2 2p^6 3s^2$
Aluminium	Al	$1s^2 2s^2 2p^6 3s^2 3p^1$
Silicon	Si	$1s^2 2s^2 2p^6 3s^2 3p^2$
Phosphorus	P	$1s^2 2s^2 2p^6 3s^2 3p^3$
Sulphur	S	$1s^2 2s^2 2p^6 3s^2 3p^4$
Chlorine	Cl	$1s^2 2s^2 2p^6 3s^2 3p^5$
Argon	Ar	$1s^2 2s^2 2p^6 3s^2 3p^6$

The electrons present in the outermost shell are called **valence electrons**. Since the number of protons as well as that of electrons increases as we move from left to right in a period in the periodic table, the force of attraction between them gradually increases. Thus, there is a gradual variation of properties in a period.

Thus, the properties of elements show a regular variation or trend along a group or a period in the periodic table. We would now discuss the variations of some periodic properties such as valence, atomic sizes, ionization energy, melting and boiling points, metallic character etc. while we move down a group or left to right across a period on the periodic table.

1.4.1 Valence

Valence of an element is its combining capacity. It is equal to the number of electrons present in the outermost or valence shell of an atom of the element. While moving down a group, you can see that the number of valence electrons remains the same. Hence, *the valence in a group remains the same for all those elements which constitute a group*. The valence of representative elements is usually equal to the number of valence electrons or eight minus the number of valence electrons. This is illustrated in Table 1.6 for second period elements. Hence, in a period, the valence first increases from 1 to 4 and then decreases to 1.

Table 1.6 : Valence of Second Period Elements

Elements	Li	Be	B	C	N	O	F
No. of valence electrons	1	2	3	4	5	6	7

Valence (with respect to hydrogen)	1	2	3	4	3	2	1
Compound formed with hydrogen	LiH	BeH ₂	BH ₃	CH ₄	NH ₃	OH ₂ (or H ₂ O)	FH (or HF)

1.4.2 Atomic Radii

Atoms do not have sharp boundaries as the electron cloud gradually drops off with the increasing distance from the nucleus. Then what do we mean by the term **atomic radius** which gives the size of an atom. When atoms are close together in solids and molecules, their centers have definite distance between them. Thus, the **atomic radius of an element is half the distance between the nuclei of neighbouring atoms in a molecule**. For example, in case of bromine molecule, the distance between two nuclei is 228 pm. Hence, its atomic radius is 114 pm. This radius is also known as **covalent radius** of an element. Similarly, the atomic radius of chlorine atom is 99 pm.

Figure 1.5 shows the atomic radii of some of the atoms.

Figure 1.5 : Atomic Radii for Selected Atoms

Let us now analyse the trends in atomic radii along the groups and periods in the Periodic Table. You can see in Figure 1.5 that **the atomic radii increase down a group while they decrease in a period from left to right**.

The increase in atomic radii on moving down a group can be explained on the basis of the fact that in each period, a new shell gets added which lies farther and farther from the nucleus.

On the other hand while we move along a period from left to right, the effective nuclear charge increases with the increase in the numbers of protons. The new electrons occupy the same shell and hence are not able to give much shielding. Therefore, the valence electrons get closer to the nucleus because of increasing effective nuclear charge and the atomic radii decrease. These trends are shown below in Figure 1.6.

Figure 1.6 : Trends in Atomic Radii

The following figure would help in memorizing these trends. Thus, cesium is the largest atom in Group 1 elements. While neon has smallest atom in second period. After understanding the trends in atomic radii, let us now see what happens when the atoms form ions.

Figure 1.7 : Trends of Atomic Radii in the Periodic Table

1.4.3 Ionic Radii

Ionic radius of an ion of an element is its share of distance in an ionic solid. This is shown below in Figure 1.8.

Figure 1.8 : Ionic Radius

When an atom forms a cation, it loses one or more electrons. Since on ion formation, the number of protons remains the same while the number of electrons decreases, hence the effective nuclear charge increases which leads to the decrease in size. Also, at the same time, a shell is completely lost which also leads to the decrease in size. For example, in case of Li atom, the electronic configuration is $1s^2 2s^1$ and its atomic radius is 157 pm. When it forms Li^+ ion by loss of one electron ($2s^1$), the electronic configuration changes to $1s^2$ which means that it has lost its second shell. Simultaneously now there are **three** protons which are attracting only **two** electrons. Remember that in Li atom, **three** protons were attracting **three** electrons. Hence, the effective nuclear

change has increased in case of Li^+ in as compared to Li atom. Thus, we can say that **a cation is always smaller than its parent atom.**

What is the situation if an atom forms an anion. You can easily predict that since the anion contains more electrons than the parent atom, the following factors would lead to an increase in the size.

- The effective nuclear charge is less in an anion as compared to the atom.
- Since the electrons are accommodated in the same shell, there is a repulsion between them.

Thus, **an anion is always larger than its parent atom.**

The trends in ionic radii are, however, similar to those of atomic radii, as discussed above.

Let us now also see the case of **isoelectronic** species. They contain the same number of electrons. Thus, Na^+ , F^- and Mg^{2+} are isoelectronic because all of them have 10 electrons. What about their sizes? Do they have equal radii? Do not forget that each one of them has different number of protons in their nucleus. Hence, the effective nuclear charge in each one of them would be different. The number of protons in them is 11, 9 and 12, respectively. Since, Mg^{2+} has the largest nuclear charge, the electrons are pulled more strongly in it and hence it has the smallest radius. On the other hand, in case of F^- ion, the number of protons is least and hence the effective nuclear charge would be greatest leading to its greater size. Hence, it will be largest amongst the three. The Na^+ ion would be having the radius in between the two.

SAQ 2

Which one of them is larger?

- Na or Na^+
- O or O^{2-}
- Cl^- or Cl
- Mg^{2+} or Mg

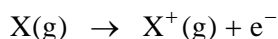
SAQ 3

Arrange the following in the increasing order of their size :

- Na, Na^+ , Mg^{2+}
- Na, Cl, F, Li

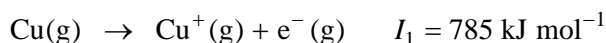
1.4.4 Ionisation Energy

Ionisation energy is the energy required to remove an electron from a gaseous atom or ion. Thus, it is the energy required for the following change

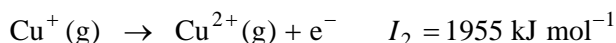


where the atom or ion is assumed to be in its ground state.

The energy required for removing one electron from an atom is called its **first ionisation energy**. For example, when Cu(g) changes to $\text{Cu}^+(\text{g})$, the **first ionisation energy** is 785 kJ mol^{-1} .

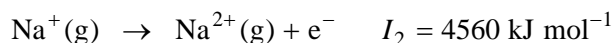
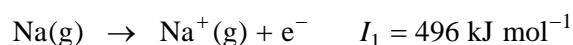


The **second ionisation energy** is the energy required to remove an electron from the singly charged cation in the gas phase. Thus, it corresponds to the following change and has the value 1955 kJ mol^{-1} for copper.



Note that the energy required to remove a second electron is higher than that required to remove the first electron. This makes a sense because in the second case, the electron is to be pulled out from a positively charged cation which has greater effective nuclear charge as compared to the neutral atom.

Let us compare the first and second ionisation energies of Na which are given below :



Note that I_2 is much much greater than I_1 . The reason being that in case of second ionisation, the electron is to be removed from an inner shell or the core of the atom. The core electrons are bound more tightly than the valence electrons. The successive ionisation energies for some elements are given in Table 1.7.

Table 1.7 : Successive Energies for the Elements of the Third Period

Let us now analyse the trends in the ionisation energies of main group elements in the periodic table. We will consider the first ionisation energies. The energy values are depicted in Figure 1.9.

Figure 1.9 : The First Ionisation Energies of Main Group Elements

You can see that **generally the ionisation energy (I_1) decreases down a group while it increases from left to right in a period**. The ionisation energy decreases in a group because in successive periods, each shell is farther away from the nucleus and the electrons are less tightly held. Hence, lesser energy is required to remove them. On moving left to right in a period, the effective nuclear charge increases and hence greater energy is required to remove the electron. These trends of ionisation energies are depicted in Figure 1.10.

Figure 1.10 : Ionisation Energies of Some Representative Elements

Thus, cesium has the lowest ionisation energy while helium and neon rank very high in ionisation energies. This is also shown in Figure 1.11.

Figure 1.11 : Visual Depiction of Trends in Ionisation Energy

Thus, elements at the left bottom of the periodic table easily form cations.

1.4.5 Metallic Character

The most basic division of elements is into **metals** and **non-metals**. Metals have low ionisation energies and have a tendency to form cations. The metallic elements are found on the left hand side of the periodic table while the non-metallic ones are located on the right side of the periodic table. The division between metals and non-metals is not a very sharp one. Many elements show metallic and non-metallic properties under different conditions. They are called **metalloids** or **semi-metals**. They are placed in between the metals and non-metals. Figure 1.12 shows this division of elements.

Figure 1.12 : Location of Metals, Non-metals and Metalloids in Periodic Table

Let us now see how metallic character varies down a group and across a period.

In a group, the metallic character increases from top to bottom in case of main group elements. Across a period, the metallic character decreases on moving from left to right. The non-metallic character, however, shows the opposite trends.

SAQ 4

- (a) Which element has the lowest ionisation energy?
- (b) How does the ionisation energy vary along a Group in the periodic table?

1.5 SUMMARY

In this unit, you have learnt that

- A set of three elements having similar properties was called a *triad* in which the properties of middle element were in between those of the other two elements.
- Modern periodic law states that the properties of elements are periodic function of their atomic numbers.
- Periodic table is the arrangement of elements on the basis of their similarities in properties.
- Elements in a group have similar electronic configurations.
- In the Periodic Table, atomic radii increase on moving down a group while they decrease on going from left to right in a period.
- Cations are smaller than the parent atom while the anions are larger than the parent atom.
- The second ionisation energy is always greater than the first ionisation energy.
- In the periodic table, the ionisation energy decreases down a group while it increases across a period.
- The metallic character increases on moving down a group while it decreases across a period.
- Metals are located on the left hand side in the periodic table while non-metals are present on the right side.

1.6 ANSWERS TO SAQs

SAQ 1

- (a) 8
(b) O – $1s^2 2s^2 2p^4$; (ii) P – $1s^2 2s^2 2p^6 3s^2 3p^3$; (iii) Cl – $1s^2 2s^2 2p^6 3s^2 3p^5$;
(iv) Ca – $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$; (v) Fe – $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
(c) (i) *s*; (ii) *p*; (iii) *d*; (iv) *p*; (v) *s*; (vi) *d*; (vii) *f*

SAQ 2

- (i) Na; (ii) O^{2-} ; (iii) Cl^- ; (iv) Mg

SAQ 3

- (i) $Mg^{2+} < Na^+ < Na$; (ii) $Li < F < Na < Cl$

SAQ 4

- (a) Cesium.
(b) It decreases along a Group.