The Periodic Table

Doberiener's Triads and Newlands Law of Octaves

There are 118 elements that are known at present. Some elements have similar properties whereas some others have completely contrasting properties.

You must have observed that in a grocery store, things are kept in an orderly manner. For example, soaps are stacked at one place while biscuits are kept separately at another place. Scientists too tried to arrange elements based on their properties. However, as more and more elements were discovered, it became increasingly difficult to arrange these elements.

Hence, scientists began to look for some pattern in the properties of these elements. Let us study in this part how famous scientists such as Johann Wolfgang Dobereiner and John Newlands arranged the elements discovered at that time.

In 1829, **Johann Wolfgang Dobereiner,** a German chemist, classified elements into groups based on their properties. He kept all elements having similar properties in one group. Most of his groups had three elements each. Thus, he called these groups as **triads.** He was the first person to illustrate the relationship between the atomic masses of elements and their properties.

Mass number is the sum of the number of protons and neutrons in an element.

He also gave a law known as the **Law of Triads**. It states that when three elements in a triad are listed in the increasing order of their atomic masses, the atomic mass of the middle element will roughly be the average of the atomic masses of the other two elements. This is demonstrated in the following animation.

The above example illustrates Dobereiner's Law of Triads.

Similarly, this law can also be proved for the triads of other elements as shown below:

Other metals		**Non-metals**	
Element	**Atomic mass**	**Element**	**Atomic mass**

Calcium	40	Chlorine	35.5
Strontium	88	Bromine	80
Barium	137	Iodine	127

Average mass of calcium and barium = $\frac{40+137}{2}$

= 88.5

Average mass of chlorine and iodine = $\frac{35.5+127}{2}$

= 81.2

In both cases, the average atomic mass of the middle element is approximately equal to the average atomic mass of the other two elements.

Hence, Dobereiner was able to identify only three triads from the elements known at that time as shown in **table 1**.

Li	Ca	Cl
Na	Sr	Br
K	Ba	I

Table 1: Dobereiner's triads

Limitations of Dobereiner's classification of elements:

- All known elements could not be classified into groups of triads on the basis of their properties.
- Not all groups obeyed the Law of Triads. For example, nitrogen family does not obey the Law of Triads.

Nitrogen family

<table>
<tr><th>Element</th><th>Atomic mass</th></tr>
<tr><td>Nitrogen</td><td>14</td></tr>
<tr><td>Phosphorus</td><td>31</td></tr>
<tr><td>Arsenic</td><td>74.9</td></tr>
<tr><td colspan="2">Average mass of nitrogen and arsenic = $\frac{14+74.9}{2}$

= 44.45

In this case, the mass of phosphorus is not equal to the average mass of the other two elements i.e., nitrogen and arsenic. Hence, the elements of this group do not obey Dobereiner’s Law of Triads.</td></tr>
</table>

Although Dobereiner tried to classify elements into groups based on their properties, his classification was not very useful.

In 1866, **John Newlands,** an English scientist, arranged the known elements in an increasing order of their atomic masses.

He began with hydrogen, which has the lowest atomic mass. He observed that if the elements are arranged in the increasing order of their atomic masses, then every eighth element (starting from a given element) had properties similar to those of the first element.

Therefore, he arranged 56 elements in seven groups such that elements having same properties were present below each other in the form of a group. A part of the original Newlands' Octaves is shown in **table 2**.

Notes of music	sa	re	ga	ma	pa	da	ni
Arrangement of elements	H	Li	Be	B	C	N	O
	F	Na	Mg	Al	Si	P	S
	Cl	K	Ca	Cr	Ti	Mn	Fe
	Co and Ni	Cu	Zn	Y	In	As	Se
	Br	Rb	Sr	Ce and La	Zr	—	—

Table 2: Newlands' Octaves

He compared his table to the **octaves of music**. It is for this reason that he called his model the **Law of Octaves** or **Newlands' Law of Octaves**.

DO YOU KNOW?

There are seven musical notes in the Indian system of music. They are - *sa, re, ga, ma, pa, da, ni*. Similarly, in the west, they use the notations - *do, re, mi, fa, so, la, ti*. Every eighth note is similar to the first one and it is the first note of the next scale.

According to **Newlands' Law of Octaves**, the properties of fluorine (eighth element starting from hydrogen) are similar to those of hydrogen. Similarly, the properties of sodium are similar to those of lithium; the properties of magnesium are similar to those of beryllium; and so on.

Newlands' arrangement of atoms showed for the first time that elements could be arranged and grouped based on some fundamental property such as the atomic mass.

Do you know that Newlands' Law of Octaves was found applicable only for the elements having low atomic masses?

Newlands' Law of Octaves has many limitations, which are discussed below.

- This law was not applicable throughout the arrangement. It was applicable only till calcium.
- Newlands' assumed that only 56 elements would exist in nature and believed that no more elements would be discovered. However, several elements were discovered in the following years. These elements did not follow the Law of Octaves.
- The positions of cobalt and nickel could not be explained according to Newlands' Law of Octaves. He kept cobalt and nickel in the same slot. They were also placed in the same column as fluorine and chlorine, which have completely different properties.

- The properties of iron are similar to those of cobalt and nickel. However, iron was placed away from them in a different column.

Mendeleev's Periodic Table

In 1866, John Newlands, an English scientist, arranged the known elements in an increasing order of their atomic masses, and his model is known as **Newlands' Law of Octaves**. Newlands began with hydrogen, which has the lowest atomic mass.

He observed that if elements were arranged in the increasing order of their atomic masses, then every eighth element (starting from a given element) had properties similar to those of the first element.

Therefore, he arranged the elements in seven groups such that elements having the same properties were present below each other in the form of a group. However, Newlands' Law of Octaves has many limitations. Hence, it was not highly accepted.

Limitations of Newlands' Law of Octaves:

It was not applicable throughout the arrangement. It was applicable only till calcium.

Newlands assumed that only 56 elements would exist in nature and believed that no more elements would be discovered. However, several elements were discovered in the following years.

The positions of cobalt and nickel could not be explained according to Newlands Law of Octaves. They were placed in the same column below fluorine and chlorine, which have completely different properties.

The properties of iron are similar to those of cobalt and nickel. However, iron was placed away from them in a different group.

Lother Meyer's arrangement of elements

In 1869, a German chemist, Lother Meyer plotted the different physical properties of elements, like atomic volume, density, melting point, boiling point, thermal conductivity etc. against their atomic weights. He found that these properties varied in a periodic fashion. Based on this, he arranged elements in a way which later nearly resembled Mendeleev's arrangement of elements.

After the failure of Newlands Law of Octaves, scientists continued to correlate the properties of elements with their atomic masses. **Dmitri Ivanovich Mendeleev** (1834-1907), a Russian scientist, also tried to relate the properties of elements with their atomic masses.

Mendeleev was the first person to introduce the concept of a periodic table. His periodic table was based on the atomic mass as well as the properties of elements. He arranged elements in the increasing order of their atomic masses and grouped the elements having similar properties together. Mendeleev formulated the periodic law in 1869.

Mendeleev's Periodic Law:

This law states that the properties of elements are a periodic function of their atomic masses. This means that arrangement of elements in the increasing order of their atomic masses, results in repetition of their properties after regular intervals.

Only 63 elements were known when Mendeleev first started his work. He arranged these elements in the form of a table in the increasing order of their atomic masses. This table is known as **Mendeleev's periodic table** as shown in the table. His table contains vertical columns called '**groups**' and horizontal rows called '**periods**'.

Groups	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
Periods ↓	A B	A B	A B	A B	A B	A B	A B	Transition series
1	H 1.008							
2	Li 6.939	Be 9.012	B 10.81	C 12.011	N 14.007	O 15.999	F 18.998	
3	Na 22.99	Mg 24.31	Al 29.98	Si 28.09	P 30.974	S 32.06	Cl 35.453	
4 First series:	K 39.102	Ca 40.08	Sc 44.96	Ti 47.90	V 50.94	Cr 50.20	Mn 54.94	Fe 55.85 Co 58.93 Ni 58.71
Second series:	Cu 63.54	Zn 65.37	Ga 69.72	Ge 72.59	As 74.92	Se 78.96	Br 79.909	
5 First series:	Rb 85.47	Sr 87.62	Y 88.91	Zr 91.22	Nb 92.91	Mo 95.94	Tc 99	Ru 101.07 Rh 102.91 Pd 106.4
Second series:	Ag 107.87	Cd 112.40	In 114.82	Sn 118.69	Sb 121.75	Te 127.60	I 126.90	
6 First series:	Cs 132.90	Ba 137.34	La 138.91	Hf 178.49	Ta 180.95	W 183.85		Os 190.2 Ir 192.2 Pt 195.09
Second series:	Au 196.97	Hg 200.59	Tl 204.37	Pb 207.19	Bi 208.98			

Mendeleev's periodic table

Do you know that Mendeleev's periodic table was a huge success? Let us now study the achievements of Mendeleev's periodic table.

Achievements of Mendeleev's periodic table:

- Mendeleev left some gaps in his periodic table. He had predicted that some elements were yet to be discovered. He left these gaps deliberately so that these undiscovered elements could get a place in his periodic table.

- Mendeleev named the undiscovered elements using the Sanskrit word *Eka* (meaning one) as a prefix, with the name of the preceding element in the same group. For example, gallium was not discovered in Mendeleev's time.

  Therefore, he left a gap for it in his periodic table and named it *Eka*-aluminium. He also predicted the properties of these undiscovered elements based on their positions in the periodic table. A comparison of the properties of gallium as predicted by Mendeleev and its actual properties is given in the following table.

Property	***Eka*-aluminium**	**Gallium**
Atomic mass	68	69.7
Density (g/cm^3)	5.9	5.93
Melting point (°C)	Low	30.15
Oxides formula	Ea_2O_3	Ga_2O_3
Chlorides formula	$EaCl_3$	$GaCl_3$

Comparison of the properties of *Eka*-aluminium and gallium

In the periodic table, Ea is used as the symbol for *Eka*-aluminium. It can be clearly observed from the table that Mendeleev's predictions were almost exact. This extraordinary achievement made Mendeleev's periodic table very popular.

- Noble gases were not discovered at the time when Mendeleev gave the periodic table. These were discovered in recent times as they are very inert and are present in very low concentrations in the atmosphere.

  When these gases were finally discovered, they got a place in his periodic table as a separate column. The accommodation of these gases in the periodic table did not disturb the positions of other elements. This underlined the strength of Mendeleev's periodic table.

- Although Mendeleev arranged the elements in the increasing order of their atomic masses, there were instances where he placed an element with a slightly higher atomic mass before an element with a slightly lower atomic mass.

  For example, cobalt, whose atomic mass is 58.9, was placed before nickel whose atomic mass is 58.7. This was done to maintain consistency in the properties of the elements present in a group i.e., to group the elements with similar properties together.

  Though, Mendeleev's periodic table was a huge success at that time, it still had many limitations. Now, let us discuss the limitations of Mendeleev's periodic table.

Limitations of Mendeleev's periodic table:

1. Hydrogen's position was not justified in Mendeleev's periodic table. He positioned hydrogen in the first column above alkali metals. He did so because hydrogen and alkali metals have similar properties. For example, hydrogen reacts with halogens, oxygen, and sulphur to form compounds whose formulae are similar to those of alkali metals.

Compounds of H	Compounds of Na
HCl	$NaCl$
H_2O	Na_2O
H_2S	Na_2S

Hydrogen and alkali metals reacting with halogens

However, hydrogen also resembles halogens in many ways. Like halogens, hydrogen is a gas, and exists as a diatomic molecule (H_2). It forms covalent compounds like halogens unlike alkali metals. Hence, it can also be placed above the halogen group.

Therefore, Mendeleev was not able to explain the position of hydrogen. In other words, the position of hydrogen in Mendeleev's periodic table was not justified. This was the first limitation of Mendeleev's periodic table.

2. The discovery of isotopes revealed another limitation of Mendeleev's periodic table. Since Mendeleev's periodic table was based on atomic masses of elements, isotopes should be placed in different columns despite the fact that they represent the same element.

Atoms of the same elements having different number of neutrons are called isotopes. Isotopes have the same number of protons and electrons, but different number of neutrons. For example, the isotopes of chlorine are Cl-35 and Cl-37. They have the same atomic number, but different atomic masses.

Modern Periodic Table

Mendeleev made a successful effort in grouping elements in the form of his periodic table. He had many achievements, but there were many limitations in his Periodic Table as well.

Some limitations of Mendeleev's periodic table are listed below:

The position of hydrogen was not justified in Mendeleev's periodic table.

The discovery of isotopes revealed another limitation of Mendeleev's periodic table.

Although Mendeleev arranged the elements in the increasing order of their atomic masses, there were instances where he had placed an element with a slightly higher atomic mass before an element with a slightly lower atomic mass.

The limitations of Mendeleev's periodic table forced scientists to believe that atomic mass could not be the basis for the classification of elements.

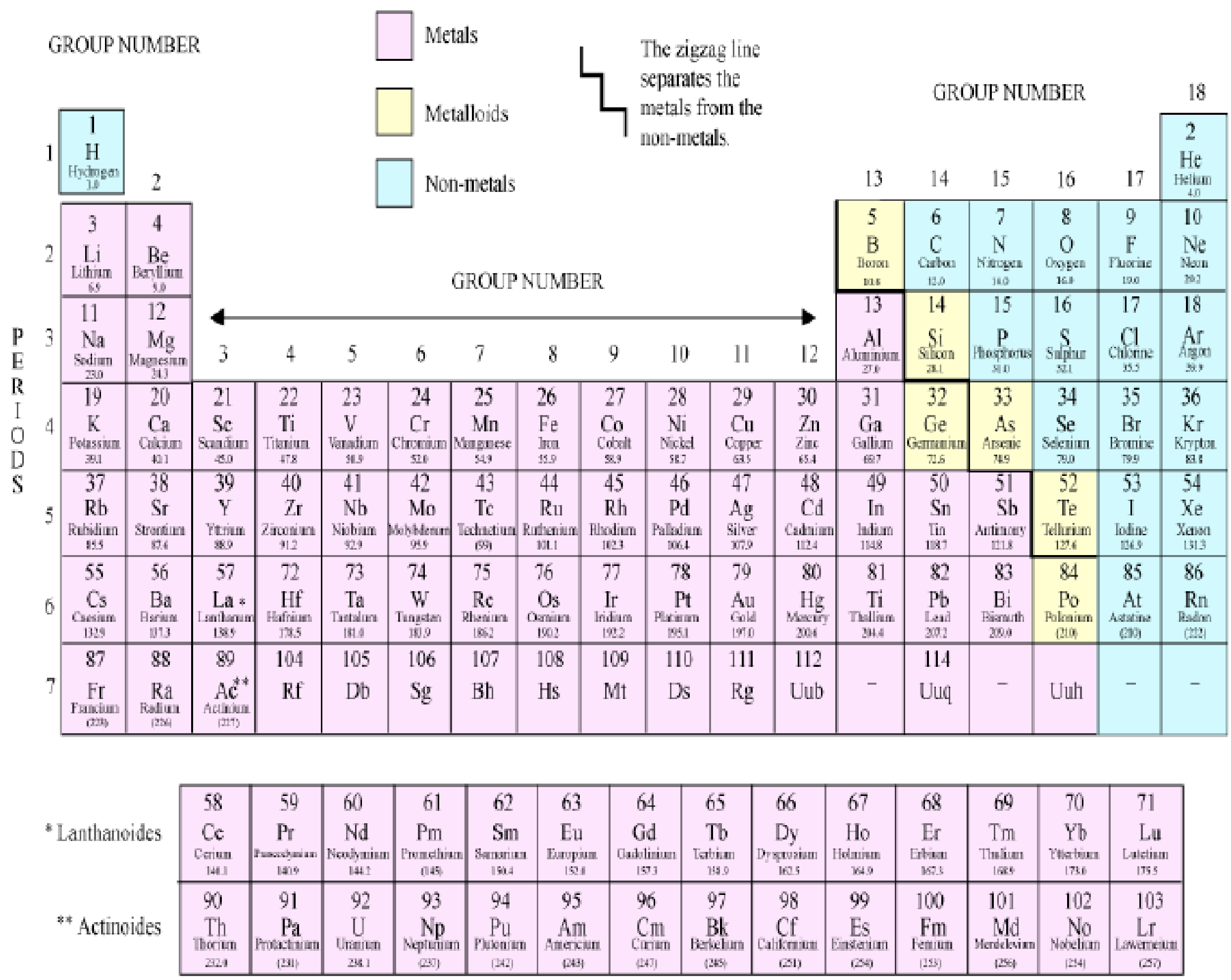
In 1913, **Henry Moseley** demonstrated that atomic number (instead of atomic mass) is a more fundamental property for classifying elements. The atomic number of an element is equal to the number of protons present in an atom of that element. Since the number of protons and electrons in an atom of an element is equal, the atomic number of an element is equal to the number of electrons present in a neutral atom.

Atomic number = Number of protons = Number of electrons

The number of protons or electrons in an element is fixed. No two elements can have the same atomic number. Hence, elements can be easily classified in the increasing order of their atomic numbers. In the light of this fact, Mendeleev's Periodic Law was done away with. As a result, the modern periodic law came into the picture.

The modern periodic law states that the properties of elements are a periodic function of their atomic numbers, not their atomic masses.

The table that is obtained when elements are arranged in the increasing order of their atomic numbers is called the **Modern Periodic Table** or **Long Form of the Periodic Table** as shown in the figure.



The Modern periodic table

In the modern periodic table, the elements are arranged in rows and columns. These rows and columns are known as **periods** and **groups** respectively. The table consists of 7 periods and 18 groups.

Do You Know:

In the modern periodic table, hydrogen is placed above alkali metals because of resemblance with their electronic configurations. However, it is never regarded as an alkali metal. This makes hydrogen a unique element.

If you look at the modern periodic table, you will find that all elements in the same group contain the same number of valence electrons. Let us see the following activity to understand better.

Activity 1: Look at group two of the modern periodic table. Write the name of the first three elements followed by their electronic configurations.

What similarity do you observe in their electronic configurations? How many valence electrons are present in these elements?

The first three elements of group two are beryllium, magnesium, and calcium. All these elements contain the same number of valence electrons. The number of valence electrons present in these elements is 2. On the other hand, the number of shells increases as we go down the group.

Again, if you look at periods in the modern periodic table, you will find that all elements in the same period contain the same valence shell. Let us see the following activity to understand better.

Activity 2: Look at the elements of the third period of the modern periodic table. Write the electronic configuration of each element and calculate the number of valence electrons present in these elements.

What do you observe from the given activity? Do these elements contain the same number of shells? How many valence electrons are present in these elements?

You will find that elements such as sodium, magnesium, aluminium, silicon, phosphorus, sulphur, chlorine, and argon are present in that period.

The valence shell in all these elements is the same, but they do not have the same number of valence electrons.

Name of the element	**Electronic configuration (K, L, M)**
Sodium	2, 8, 1
Magnesium	2, 8, 2
Aluminium	2, 8, 3

Silicon	2, 8, 4
Phosphorus	2, 8, 5
Sulphur	2, 8, 6
Chlorine	2, 8, 7
Argon	2, 8, 8

Thus, the number of electrons in the valence shell increases by one unit as the atomic number increases by one unit on moving from left to right in a period.

Let us calculate the number of elements that are present in the first, second, third, and fourth periods.

The maximum number of electrons that a shell can hold can be calculated using the formula $2n^2$. Here, *n* represents the number of shells from the nucleus. For example, *n* is equal to 1, 2, and 3 for K, L, and M shells respectively. Hence, the maximum number of electrons that each of these shells can hold can be calculated by substituting the value of *n* in the given formula.

Number of electrons that K shell can accommodate = $2n^2$

= 2×1^2

= 2

Hence, K shell can accommodate only 2 electrons and only two elements are present in the first period.

Similarly, the second and third shell (L and M respectively) can accommodate 8 and 18 electrons respectively. Since the outermost shell can contain only 8 electrons, there are only 8 elements in both the periods.

Important Note:

The position of an element in the Modern Periodic Table tells us about its chemical reactivity. The valence electrons determine the kind and the number of bonds formed by an element.

IUPAC Nomenclature for Elements with Atomic Number $>$ 100

- Latin word roots for various digits are listed in the given table.

Notation for IUPAC Nomenclature of Elements

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

- Latin words for various digits of the atomic number are written together in the order of digits, which make up the atomic number, and at the end, 'ium' is added.

- Nomenclature of elements with the atomic number above 100 is listed below.

Nomenclature of Elements with Atomic Number Above 100

Atomic number	Name	Symbol	IUPAC Official Name	IUPAC Symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg

107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassnium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnilium	Uun	Darmstadtium	Ds
111	Unununnium	Uuu	Rontgenium	Rg
112	Ununbium	Uub		
113	Ununtrium	Uut		
114	Ununquadium	Uuq		
115	Ununpentium	Uup		
116	Ununhexium	Uuh		
117	Ununseptium	Uus		
118	Ununoctium	Uuo		

Characteristics of the Modern Periodic Table

Characteristic of Periods:

● All elements of the same period have similar number of shells, i.e. their electrons are filled in the same valence shell.
● Electronic configuration and chemical properties of the elements change along the period.
● Atomic size of the elements decreases along the period from left to right.
● Metallic character decreases while non-metallic character increases along the period.

Characteristic of Groups:
● In Groups 2 and 18, differences in atomic numbers of the elements are 8, 8, 18, 18 and 32.
●In Groups 13 and 17, differences in atomic numbers of the elements are 8,18,18 and 32.
● The elements of the same group have the same number of valence electrons and valency.
● The elements of the same group have identical chemical properties. Physical properties, i.e. melting point, boiling point, density, etc., of the elements of the same group vary gradually.
● Atomic radii of an element of the same group increases down the group.

Advantages of the Modern Periodic Table:
● The modern periodic table is based on atomic number, which is a fundamental property.
● It easily and clearly correlates the position of elements with its electronic configuration.
● It has more logical completion of each period. At the end of every period, there is an element with noble gas configuration.
● This table is easy to remember.
● It does not have any subgroup and every group is independent.
● The isotopes of different elements are placed in the same group as they all have the same atomic number.
● All the transition elements are present in the middle as the properties of transition metals are intermediate to the left and the right portions of the periodic table.
● It completely separates metals from non-metals.
● It justifies the positions of some elements, which earlier were a misfit.
● It provides justification for the placement of lanthanides and actinides at the bottom of the periodic table.

Defects in the Modern Periodic Table:
● It does not denote a fix position for hydrogen. Lanthanides and actinides have not been given position inside the main body of the periodic table.
● It does not justify the exact electronic configuration of some transition and inner transition elements.

Electronic Configuration and the Periodic Table

Electronic Configuration in Periods

- Period indicates the value of '*n*' (principal quantum number) for the outermost or valence shell.

- Successive periods in the periodic table are associated with the filling of the next higher principal energy level (n = 2, n = 3, etc).

- First period (n = 1) → hydrogen ($1s^1$) and helium ($1s^2$) [2 elements]

- Second period (n = 2) → Li ($1s^2\ 2s^1$), Be ($1s^2\ 2s^2$), B ($1s^2\ 2s^2\ 2p^1$) to Ne ($2s^2\ 2p^6$) [8 elements]

- Third period (n = 3) → filling to 3*s* and 3*p* orbitals gives rise to 8 elements (Na to Ar)

- Fourth period (n = 4) → 18 elements (K to Kr) **−** filling of the 4*s* and 4*p* orbitals

  3*d* orbital is filled up before 4*p* orbitals (3*d* orbitals → energetically favourable)

- 3*d*-transition series → Sc ($3d^1\ 4s^2$) to Zn ($3d^{10}\ 4s^2$)

- Fifth period (n = 5) → 18 elements (Rb to Xe)

- 4*d*-transition series starts at Ytterbium and ends at Cadmium.

- Sixth period (n = 6) → 32 elements; electrons enter 6*s*, 4*f*, 5*d*, and 6*p* orbitals successively. Elements from Z = 58 to Z = 71 are called 4*f*-inner transition series or lanthanoid series (filling up of the 4*f* orbitals).

- Seventh period (n = 7) → electrons enter at 7*s*, 5*f*, 6*d*, and 7*p* orbitals successively. Filling up of 5*f* orbitals after Ac (Z = 89) gives 5*f*-inner transition series or the actinoid series.

Electronic Configuration in Groups

- Same number of electrons is present in the outer orbitals (that is, similar valence shell electronic configuration).

- Electronic configuration of group 1 elements is given in the following table.

Atomic number	**Symbol**	**Electronic configuration**
3	Li	$1s^2\ 2s^1$ **(or)** [He]$2s^1$
11	Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$ **(or)** [Ne]$3s^1$

19	K	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$ **(or)** $[Ar]4s^1$
37	Rb	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 5s^1$ **(or)** $[Kr]5s^1$
55	Cs	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 5s^2\ 5p^6\ 6s^1$ **(or)** $[Xe]6s^1$
87	Fr	$[Rn]7s^1$

Electronic Configurations and Types of Elements

- ***s*- Block Elements**

- Group 1 (alkali metals) − ns^1 (outermost electronic configuration)
- Group 2 (alkaline earth metals) − ns^2 (outermost electronic configuration)
- Alkali metals form +1 ion and alkaline earth metals form +2 ion.
- Reactivity increases as we move down the group.
- They are never found in the pure state in nature. (Reason − they are highly reactive)

- ***p* - Block Elements**

- Elements belonging to Groups 13 to 18
- Outermost electronic configuration varies from ns^2np^1 to ns^2np^6 in each period.
- Group 18 (ns^2np^6) − noble gases
- Group 17 − halogen
- Group 16 − chalcogens
- Non-metallic character increases from left to right across a period.

- ***d*- Block Elements (Transition Elements)**

- Elements of group 3 to group 12
- General electronic configuration is $(n-1)\,d^{1-10}\,ns^{0\text{-}2}$.
- Called transition elements
- Zn, Cd, and Hg with $(n-1)\,d^{10}\,ns^2$ configuration do not show properties of transition elements.
- All are metals. They form coloured ions, exhibit variable oxidation states, paramagnetism, and are used as catalysts.

- ***f*- Block Elements**

- Lanthanoids → Ce (Z = 58) to Lu (Z = 71)

- Actinoids → Th (Z = 90) to Lr (Z = 103)
- Outer electronic configuration → $(n - 2)\, f^{1-14}\, (n - 1)\, d^{0-1}\, ns^2$
- They are called inner-transition elements.
- All are metals.
- Actinoid elements are radioactive.
- Elements after uranium are called **Transuranium** elements.

Metals, Non-metals, and Metalloids

- Metals → Appear on the left side of the periodic table
- Non-metals → Located at the top right-hand side of the periodic table
- Elements change from metallic to non-metallic from left to right.
- Elements such as Si, Ge, As, Sb, Te show the characteristic properties of both metals and non-metals. They are called semi-metals or metalloids.

Trends in the Modern Periodic Table

If you observe the long form of the periodic table or the modern periodic table, then you will find that there are certain trends followed. These trends help us in determining various properties of the elements.

For example, elements like sodium, magnesium, aluminium, silicon, phosphorus, sulphur, chlorine, and argon are present in the third period. **Do you know why they are placed like this?**

It is because all these elements have one thing in common.

You must have observed that these elements contain the same number of shells, and the last electron enters shell M as shown in the given table:

Elements of the third period

Element	Electronic configuration K L M
Sodium	2, 8, 1
Magnesium	2, 8, 2
Aluminium	2, 8, 3
Silicon	2, 8, 4
Phosphorus	2, 8, 5

Sulphur	2, 8, 6
Chlorine	2, 8, 7
Argon	2, 8, 8

Do you know the valency of the elements present in the third period?

Valency is defined as the number of electrons an atom requires to lose, gain, or share in order to complete its valence shell to attain the stable noble gas configuration. Valencies of the elements can also be determined by the number of electrons present in the outermost shell known as the valence shell.

Therefore, on moving across a period, from left to right, the valency first increases from 1 to 4, and then decreases from 4 to 0.

Element	Valence electrons	Valency
Sodium, Na	1	1
Magnesium, Mg	2	2
Aluminium, Al	3	3
Silicon, Si	4	4
Phosphorus, P	5	3, 5
Sulphur, S	6	2
Chlorine, Cl	7	1
Argon, Ar	8	0

Do you know how the valency of an element, present in the same group, changes on moving down the group?

The number of valence electrons present in atoms of elements belonging to a particular group remains the same. As a result, valency also remains the same. Hence, the valency of group 1 elements is one, group 2 elements is two, and so on.

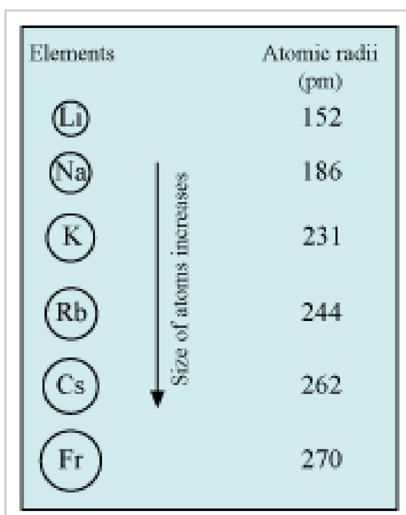
For example, the valency of the first four elements of group 1 of the periodic table is given as follows. The valency of all these four elements is 1.

Valency of the first four elements of group 1

Element	**Electronic configuration K L M**	**Valency**
Hydrogen	1	1
Lithium	2, 1	1
Sodium	2, 8, 1	1
Potassium	2, 8, 8, 1	1

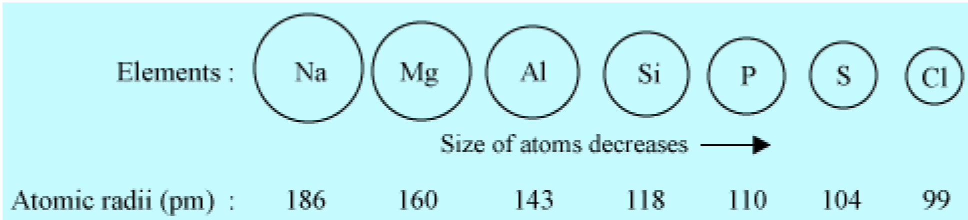
Atomic size is the radius of an atom in its neutral state i.e. the distance of the nucleus from its valence shell in an isolated gaseous atom.

Do you know how atomic size changes as we move from top to bottom or across the periodic table. Let us see.



Changes in atomic size on moving down a group

The atomic radii of the elements present in the third period are given in **Pico meters.**



Changes in atomic size on moving across a period

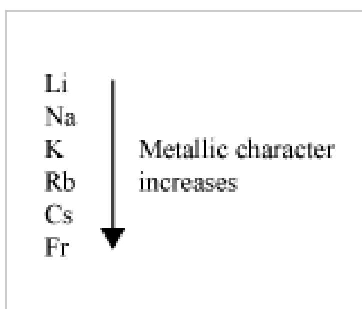
Do you know that the metallic character of elements can also be determined with the help of the periodic table?

Metals are electropositive in nature as they lose electrons to form positive ions called cations. For example, Na loses one electron to form Na^+ ion.

We have learnt that there is a decrease in the atomic size, as we move across a period. As a result, it becomes difficult for an atom to lose electrons as we move from left to right across the periodic table. Hence, metallic character decreases on moving across a period from left to right.

On the other hand, atomic size increases when we move down a group. This makes it easier for an element to lose electrons. Hence, metallic character increases on moving down a group.

Therefore, the elements present in the 1st group are the most metallic in their respective periods as they contain only one electron, and have the largest atomic size in their respective periods.



Changes in metallic character on moving down a group

Do you know that with the help of a periodic table, we can also predict the nature of the oxides formed by the various elements? The oxides formed by metals are basic in nature, whereas the oxides of non-metals are acidic in nature.

Hence, the elements present on the left hand side of the periodic table will form basic oxides, whereas the elements present on the right hand side of the periodic table will form acidic oxides.

Elements show periodicity because of their **valence shell configuration**. All elements showing periodicity in properties have the same number of electrons in the last or valence shell. The properties that will be discussed here are:

- Atomic radius
- Ionisation potential
- Electron affinity
- Electronegativity
- Metallic character

Atomic Radius: The atomic radius is usually considered as the distance from the centre of the nucleus to the outermost shell i.e., to a point where the electron density is effectively zero.

Across the period i.e., from left to right: Atomic radius decreases

Down the group i.e., from top to bottom: Atomic radius increases

Reason: Across the period, the effective nuclear charge increases. This is due to the fact that the number of electrons increases (in the same subshell), and so does the number of protons in the nucleus. This pulls the valence shell of electrons in an atom towards itself, thus decreasing the atomic radius. But as we move down the group, the number of shells keeps on increasing along with the number of protons. The space required to accommodate the extra shells becomes the dominating factor and therefore the atomic size increases.

Ionisation Potential: It is the energy required to remove one mole of electrons from the valence shell of one mole of isolated gaseous atoms.

Across the period i.e., from left to right: Ionisation potential increases

Down the group i.e., from top to bottom: Ionisation potential decreases

Reason: Across the period, the effective nuclear charge increases. This causes the atomic radius to decrease, thus getting the valence shell closer to the nucleus. This makes it difficult to remove electrons.

But as we move down the group, the number of shells keeps on increasing along with the number of electrons. The distance from the nucleus coupled with the interference of the electron between the nucleus and the valence shell renders the valence electrons weakly bound to the nucleus.

Electronegativity: The tendency of an atom of an element to attract a shared pair of electrons towards itself when bonded with another element in a compound is called electronegativity.

Across the period i.e., from left to right: Electronegativity increases

Down the group i.e., from top to bottom: Electronegativity decreases

Reason: Across the period, the effective nuclear charge increases, thus decreasing the atomic radius. This favours the increase in electronegativity of elements across the period. But as we move down the group, the number of shells keeps on increasing and therefore the atomic size increases and the electronegativity decreases.

Metallic character: It is defined as the tendency of an atom to lose electrons.

Across the period i.e., from left to right: Metallic character decreases

Down the group i.e., from top to bottom: Metallic character increases

Reason: Across the period, the effective nuclear charge increases, thus decreasing its atomic radius. This favours the increase of electronegativity and therefore the tendency to lose electrons is low. This accounts for the decrease in the metallic character along a period. But as we move down the group, the number of orbits keeps on increasing and therefore the atomic size increases. This means that the electronegativity decreases. This enhances the loss of electrons and therefore the metallic character increases down a group.

Alkali Metals

- Group 1 elements: lithium, sodium, potassium, rubidium, caesium, and francium
- Called alkali metals because their hydroxides are alkaline in nature

Electronic Configuration

- General outer electronic configuration is ns^1.
- The electronic configurations of the alkali metals are given in the following table.

Element	Electronic configuration
Lithium (Li)	$1s^2\ 2s^1$
Sodium (Na)	$1s^2\ 2s^2\ 2p^6\ 3s^1$

Potassium (K)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Rubidium (Rb)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 5s^1$
Caesium (Cs)	[Xe] $6s^1$
Francium (Fr)	[Rn] $7s^1$

Atomic and Ionic Radii

- Increase with increase in atomic number (i.e. on moving down the group)

Ionization Enthalpy

- Low
- Decreases on moving down the group, which is because of the following reasons:
- Increase in size
- Increase in screening of outermost electrons from the nuclear charge

Hydration Enthalpy

- Decreases on moving down the group

Physical Properties

- Silvery white, soft, and light metals
- Density increases down the group. (exception: Na is denser than K)
- Low melting and boiling points − Reason: weak metallic bond due to the presence of only one electron in the outermost shell
- They and their salts impart characteristic colour to an oxidising flame. Reason: after getting heat from the flame, the valence electrons get excited to higher energy level. Then they come back to ground state by emitting radiation in the visible region.
- Crimson red − Lithium (Li)
- Golden yellow − Sodium (Na)

- Pale violet − Potassium (K)

Chemical Properties

- **Highly reactive because of**

- large size
- low ionization enthalpy
- Reactivity increases down the group

- **Reaction with air:**

- Tarnish in dry air
- Reason: formation of oxides and then hydroxides with moisture
- Burns vigorously in air

$$4Li + O_2 \longrightarrow 2Li_2O \text{ (monoxide)}$$
$$2Na + O_2 \longrightarrow Na_2O_2 \text{ (Peroxide)}$$
$$M + O_2 \longrightarrow MO_2 \quad \text{(superoxide)}$$
$$(M = K, Rb, Cs)$$

  - Kept in kerosene oil − Reason: highly reactive towards air and water
  - Li reacts directly with nitrogen of air to form lithium nitride (Li_3N).

- **Reaction with water:**

$$2M + 2H_2O \longrightarrow 2M^+ + 2OH^- + H_2$$

(M = alkali metal)

- **Reaction with acids:**
- Reacts vigrously with even dil. acids to give out hydrogen gas.
- 2M + 2HCl → 2MOH

- **Reaction with dihydrogen:**

$$2M + H_2 \longrightarrow 2MH \quad \text{(hydride)}$$

  - Alkali metal hydrides are ionic solids with high melting point.
  - Reaction with acids: React voilently with dilute acids

  $2\ M + 2\ HCl \rightarrow 2\ MCl + H_2$

- **Reaction with halogen:**

- Reacts vigorously to form halides which are ionic in nature (exception: LiX is some what covalent − Reason: High polarizing capability of Li^+ ion)
- LiI is the most covalent.

- **Reducing nature**

- Strong reducing agent
- Form unipositive ions easily, $M(g) \rightarrow M^+ (g) + e^-$
- Li is the strongest and Na is the weakest.

- **Solutions in liquid ammonia**

- Dissolve in liquid ammonia to give conducting deep blue solution
- $M + (x + y)NH_3 \longrightarrow [M(NH_3)_x][e(NH_3)_y]^-$

- Reason for blue colour − The ammoniated electron absorbs energy in the visible region of light.

- The solution is paramagnetic in nature and liberates hydrogen to form amide.
- $M^+_{(am)} + e^- + NH_{3(l)} \longrightarrow MNH_{2(am)} + \frac{1}{2}H_{2(g)}$ ('am' stands for solution in ammonia)

- Its concentrated solution is bronze in colour and diamagnetic in nature.

- The metals also react with alcohol, gaseous ammonia, and alkynes

Uses

- In making useful alloys. For example, it is used with lead in making 'white metal' bearings for motor engines

- Li is used in making electrochemical cells.

- Na is used in making Na/Pb alloy, which are used to make organolead compounds − $PbEt_4$ and $PbMe_4$. These compounds were earlier used as anti-knock additives to petrol.

- Liquid Na is used as a coolant in fast breeder nuclear reactors.

- K plays an important role in biological systems

- KCl − used as a fertilizer

- KOH − in the manufacture of soft soap and as an absorbent of CO_2
- Cs is used in devising photoelectric cells.

Alkaline Earth Metals

- Group 2 elements: Beryllium, magnesium, calcium, strontium, barium and radium.
- They are called alkaline earth metals because their oxides and hydroxides are alkaline in nature, and their oxides are found in earth's crust.

Electronic Configuration

- General outer electronic configuration is ns^2
- The electronic configurations of the alkaline earth metals are given in the following table:

Element	Electronic Configuration
Beryllium (Be)	$1s^2 2s^2$
Magnesium (Mg)	$1s^2 2s^2 2p^6 3s^2$
Calcium (Ca)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Strontium (Sr)	[Kr] $5s^2$
Barium (Ba)	[Xe] $6s^2$
Radium (Ra)	[Rn] $7s^2$

Atomic and ionic radii

- Smaller than the corresponding alkali metals

Reason: Greater nuclear charge

- Increase with the increase in atomic number (i.e., on moving down the group)

Ionisation Enthalpy

- Low

- Decreases on moving down the group. This is because of:

- Increase in size
- Increase in the screening of valence electrons from the nuclear charge

- Higher first ionisation enthalpy than that of the corresponding alkali metals. This is due to smaller size and higher nuclear charge as compared to the corresponding group 1 alkali metals.

- Second ionisation enthalpy is lower than that of the corresponding alkali metals

Hydration enthalpy

- Decreases on moving down the group (i.e. with increase in ionic size)

- Larger hydration enthalpy than that of the corresponding alkali metal ions. This is why compounds of alkaline earth metals are more extensively hydrated than those of alkali metals.

- Example**:** NaCl is not hydrated whereas $MgCl_2$ exists as $MgCl_2.6H_2O$. KCl does not form hydrate whereas $CaCl_2$ exists as $CaCl_2 \cdot 6H_2O$.

Physical properties

- Silvery white, lustrous and soft (harder than alkali metals) (Exception − Beryllium and Magnesium are greyish)

- Melting and boiling points of these metals are higher than those of the corresponding alkali metals.

  Reason: Smaller size than alkali metals

- Strongly electropositive in nature. Reason: Low ionisation enthalpies

- Flame test:

- Calcium − Brick red

- Strontium − Crimson
- Barium − Apple green
- Mg and Be do not impart colour because the electrons are too strongly bound to be excited.

- Good conductors of heat and electricity

Chemical Properties

- Less reactive than alkali metals

- Reactivity increases down the group

- Reaction with air and water:

- Be and Mg are kinetically inert to oxygen and water. Reason − formation of oxide film on their surface

- However, when Mg and powdered Be burn, they give MgO and Mg_3N_2, and BeO and Be_3N_2 respectively.

- Ca, Sr and Ba react readily with air to form oxides and nitrides; also react with water (even cold), to form hydroxides

- Reaction with halogens:

- At high temperature

$$M + X_2 \rightarrow MX_2 \text{ (Halides)}$$

$$(X = F, Cl, Br, I)$$

  - BeF_2 is prepared by the thermal decomposition of $(NH_4)_2BeF_4$

  - $BeCl_2$ is prepared from BeO by heating it with carbon and chlorine to about 600-900 K

$$BeO + C + Cl_2 \rightleftharpoons BeCl_2 + CO$$

- Reaction with hydrogen:

- React with hydrogen to form hydrides (Exception − Beryllium)
- Preparation of BeH_2 from $BeCl_2$ and $LiAlH_4$ −

$$2BeCl_2 + LiAlH_4 \rightarrow 2BeH_2 + LiCl + AlCl_3$$

- Reaction with acids:

$$M + 2HCl \rightarrow MCl_2 + H_2$$

(M = Alkali earth metal)

- Reducing nature:

- Strong reducing agent (weaker than corresponding alkali metals)
- Be is a weaker reducing agent than other alkaline earth metals.

- Solutions in liquid ammonia:

Dissolve in liquid ammonia to give deep, blue-black solution

$$M + (x + y)NH_3 \rightarrow [M(NH_3)_x]^{2+} + 2[e(NH_3)_y]^{-}$$

Uses

- **Beryllium:**

- For making alloys such as Cu−Be alloys which are used for making high strength springs
- Metallic Be, for making windows of X-ray tubes

- **Magnesium:**

- For making alloys with Al, Zn, Mn and Sn
- Mg−Al alloys are used in the construction of aircraft
- Magnesium is used in flash powders and bulbs, incendiary bombs and signals
- Milk of magnesia (suspension of $(MgOH)_2$ in water) is used as an antacid
- $MgCO_3$ is a component in toothpaste

- **Calcium:**

- In the process of extraction of metals from oxides which are difficult to reduce with carbon
- For removing air from vacuum tubes. Reason − reacts with oxygen and nitrogen at high temperature

- **Barium:**

- For removing air from vacuum tubes. Reason − reacts with oxygen and nitrogen at high temperature

- **Radium:**

- In radiotherapy (e.g., in the treatment of cancer)

General Characteristics of Group VIIA or 17 Elements

Group 17 Elements

- Fluorine (F), Chlorine (Cl), Bromine (Br), Iodine (I) and Astatine (At)
- Collectively known as halogens
- Astatine, a radioactive element

Occurrence

- Fluorine is present as insoluble fluorides − fluorspar (CaF_2), cryolite, and fluoroapatite.

- Fluorine is also present in soils, rivers, bones and teeth of animals.

- Sea water contains chlorides, bromides and iodides of Na, K, Mg and Ca.

- Chlorine is also present in sodium chloride, carnallite, $KCl.MgCl_2.6H_2O$

- Seaweeds contain up to 0.5% of iodine and chile saltpetre contains up to 0.2% of sodium iodate.

Atomic Properties

- Outermost shell contains 7 electrons.

- Atomic and ionic radii increase from F to I due to increasing number of quantum shells.

- Ionisation enthalpy decreases down the group due to increase in atomic size.

- Halogens have the maximum electron gain enthalpy in the corresponding periods.

- Very high electronegativity; electronegativity decreases down the group

Physical Properties

- F and Cl are gases, Br is liquid and I is solid.

- Melting and boiling points steadily increase down the group.

- Halogens are coloured. F is pale yellow, Cl is greenish yellow, Br is reddish brown and I is violet.

- Br and I are only sparingly soluble in water, but soluble in various organic solvents such as $CHCl_3$, CCl_4, CS_2 and hydrocarbons.
- They are all poisonous and have strong pungent odour.

Chemical Properties

- Valency is −1.
- Halogens are highly reactive.
- They all form diatomic molecules.
- They react with metals and non-metals to form halides.
- Reactivity decreases down the group. Fluorine is the most reactive while iodine the least reactive haloge.
- F_2 is the strongest oxidising agent.
- A halogen oxidises halide ions of higher atomic number.
- Fluorine oxides water to oxygen.

$$2F_{2(g)} + 2H_2O_{(l)} \longrightarrow 4H^+_{(aq)} + 4F^-_{(aq)} + O_{2(g)}$$

- Cl_2 and Br_2 react with water to form corresponding hydrohalic and hypohalous acids.

$$X_{2(g)} + H_2O_{(l)} \longrightarrow HX_{(aq)} + HOX_{(aq)} \quad (X = Cl \text{ or } Br)$$

- I^- can be oxidised by oxygen in acidic medium.

$$4I^-_{(aq)} + 4H^+_{(aq)} + O_{2(g)} \longrightarrow 2I_{2(s)} + 2H_2O_{(l)}$$

- All react with hydrogen to form hydrogen halides (HX).
- Acidic strength order: HF < HCl < HBr < HI
- They form many oxides with oxygen, but most of them are unstable.
- Fluorine forms two oxides: OF_2 and O_2F_2
- They react with metals to form metal halides.

General Characteristics of Group Zero or 18 Elements

Group 18 Elements

- Helium (He), Neon (Ne), Argon (Ar), Krypton (Kr), Xenon (Xe), and Radon (Rn)
- All are gases.
- Chemically uncreative − Hence, they are termed as noble gases.

Occurrence

- Occur in atmosphere (except Rn)
- Atmospheric abundance in dry air is ~ 1%. (Argon is the major constituent)
- Xe and Rn are the rarest elements of the group.
- Rn is the decay product of ^{226}Ra.

$$^{226}_{88}Ra \longrightarrow ^{222}_{86}Ra + ^{4}_{2}He$$

- He and Ne are found in minerals of radioactive origin such as pitchblende, monazite, cleveite, etc.

Atomic Properties

- Outer shell remains completely filled.
- High ionisation enthalpy
- Reason − Stable electronic configuration
- However, ionisation enthalpy decreases down the group (that is, with the increase in atomic size).
- Atomic radii increase down the group.

Physical Properties

- Monoatomic
- Colourless, odourless, and tasteless

- Sparingly soluble in water
- Low melting and boiling points.

  He has the lowest boiling point of 4.2 K.

- Unusual property of diffusing through most commonly used laboratory materials such as rubber, glass, or plastics

  Chemical Properties

- Less reactive

  Reason:

- Completely filled valence shell electronic configuration
- Only xenon-fluorine compounds have been synthesised.
- XeF_2, XeF_4, and XeF_6

  Uses of Noble Gases

- **Helium**
- In filling balloons for meteorological observations as it is a non-inflammable and light gas
- In gas-cooled nuclear reactors
- **Liquid Helium**
- As a cryogenic agent
- As a diluent for oxygen in modern diving apparatus

  Reason − Very low solubility in blood

- To produce and sustain powerful superconducting magnets, which are essential parts of modern NMR spectrometers and MRI instruments
- **Neon**
- In discharge tubes and fluorescent bulbs

- Neon bulbs − Used in botanical gardens and in green houses
- **Argon**
- To provide an inert atmosphere in high temperature metallurgical processes
- For filling electric bulbs
- In the laboratory for handling substances that are air-sensitive
- **Xenon and Krypton**
- In light bulbs designed for special purposes