1. Details of Module and its structure

Module Detail	
Subject Name	Chemistry
Course Name	Chemistry 01 (Class XI, Semester 01)
Module Name/Title	Classification of Elements and Periodicity In Properties: Part 2
Module Id	kech_10302
Pre-requisites	Atomic number, electronic configuration & periodic classification
Objectives	After going through this module, the learner will be able to: 1. Understand the significance of electronic configuration as a basis for classification. 2. Classify elements into *s, p, d, f* blocks and understand the main characteristic properties of these elements. 3. Classify elements into metals and non-metals and understand their characteristic properties.
Keywords	Electronic configuration, *s,p,d,f* blocks, alkali metals, transition elements halogens, alkaline-earth metals, main-group elements

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1. Introduction: Classification of elements based on their properties

The Group 18 elements of the periodic table (the noble gases) undergo few chemical reactions. This stability results from the gases' special electron configurations. Helium's highest occupied level, the 1s orbital, is completely filled with electrons. And the highest occupied levels of the other noble gases contain stable octets. Generally, the electron configuration of an atom's highest occupied energy level governs the atom's chemical properties.

2. Electronic Configurations of Elements and the Periodic Table:

An atom is characterized by a set of four quantum numbers, and the principal quantum number (n) defines the main energy level known as shell. The filling of electrons into different subshells, also referred to as orbitals (s, p, d, f) in an atom is governed by Aufbau principle. The distribution of electrons into orbitals of an atom is called its electronic configuration. An element's location in the Periodic Table reflects the quantum numbers of the valence shell (last orbital filled). This section explains the direct connection between the electronic configurations of the elements and the position of the element in the long form of the Periodic Table. The period of an element is determined by its electron configuration.

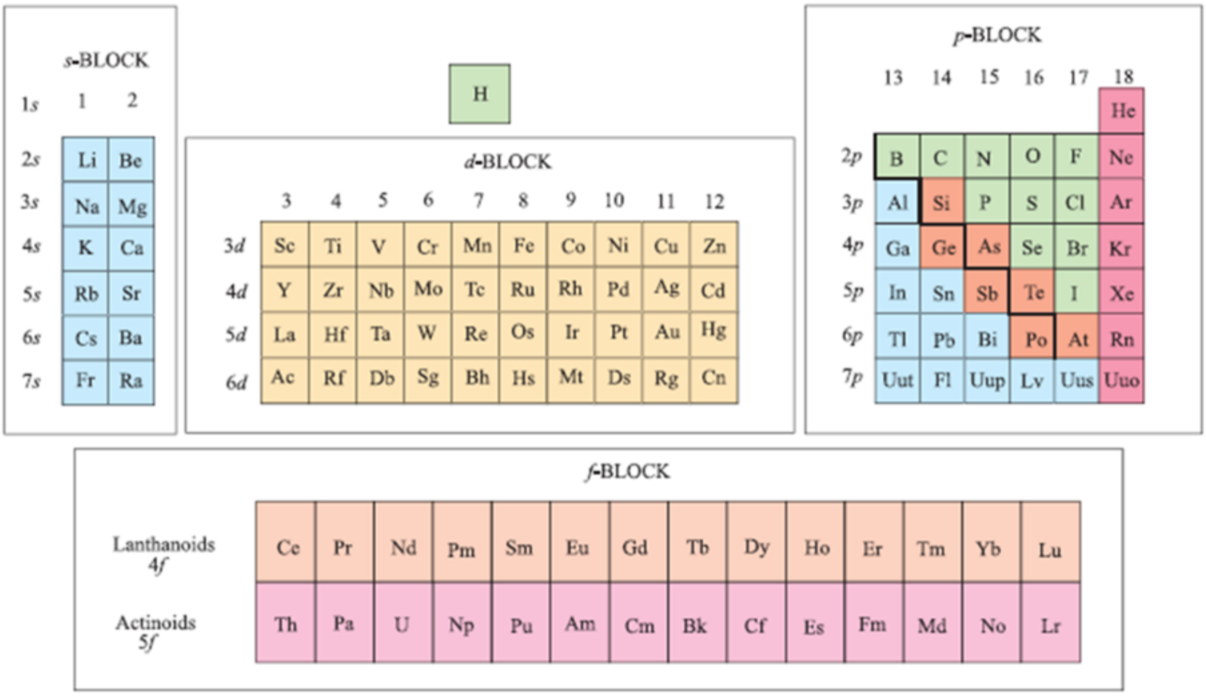

Fig. 3 The types of elements in the Periodic Table based on the orbitals that are being filled. Also shown is the broad division of elements into METALS ( ) , NON-METALS ( ) and METALLOIDS ( ).

3. Electronic Configurations in Periods:

The elements are arranged in the periodic table, vertically in groups and horizontally in rows, or periods. These are arranged based on the similarity in their chemical properties. There are a total of seven periods of elements in the modern periodic table. The length of each period is known from the number of electrons that fill the orbitals (sublevels) in that period. Refer to table below.

Relationship between period length and the orbitals being filled in the periodic table

Period number	Number of elements in the period	Order of filling of electrons in the valence shell orbitals
1	2	1s
2	8	2s 2p
3	8	3s 3p
4	18	4s 3d 4p
5	18	5s 4d 5p
6	32	6s 4f 5d 6p
7	32	7s 5f 6d 7p

The period indicates the value of n for the outermost or valence shell. In other words, successive period in the Periodic Table is associated with the filling of the next higher principal energy level (n = 1, n = 2, etc.). It can be seen that the number of elements in each period is twice the number of atomic orbitals available in the energy level that is being filled.

The first period (n = 1) starts with the filling of the lowest level (1s) and therefore has two elements — hydrogen (ls1) and helium (ls2) when the first shell (K) is completed.

The second period (n = 2) starts with lithium and the third electron enters the 2s orbital. The next element, beryllium has four electrons and has the electronic configuration 1s2, 2s2. Starting from the next element boron, the 2p orbitals are filled with electrons when the L shell is completed at neon (2s2 2p6). Thus there are 8 elements in the second period.

The third period (n = 3) begins at sodium, and the added electron enters a 3s orbital. Successive filling of 3s and 3p orbitals gives rise to the third period of 8 elements from sodium to argon.

The fourth period (n = 4) starts at potassium, and the added electrons fill up the 4s orbital. Now you may note that before the 4p orbital is filled, filling up of 3d orbitals becomes energetically favorable and we come across the so called 3d transition series of elements. This starts from scandium (Z = 21) which has the electronic configuration 3d14s2. The 3d orbitals are filled at zinc (Z=30) with electronic configuration 3d10 4s2. The fourth period ends at krypton with the filling up of the 4p orbitals. Altogether we have 18 elements in this fourth period.

The fifth period (n = 5) beginning with rubidium is similar to the fourth period and contains the 4d transition series starting at yttrium (Z = 39). This period ends at xenon with the filling up of the 5p orbitals.

The sixth period (n = 6) contains 32 elements and successive electrons enter 6s, 4f, 5d and 6p orbitals, in the order — filling up of the 4f orbitals begins with cerium (Z = 58) and ends at lutetium (Z = 71) to give the 4f-inner transition series which is called the Lanthanide series.

The seventh period (n = 7) is similar to the sixth period with the successive filling up of the 7s, 5f, 6d and 7p orbitals and includes most of the man-made radioactive elements. This period will end at the element with atomic number 118 which would belong to the noble gas family.

Filling up of the 5f orbitals after actinium (Z = 89) gives the 5f-inner transition series known as the Actinide series. The 4f and 5f-inner transition series of elements are placed separately in the Periodic Table to maintain its structure and to preserve the principle of classification by keeping elements with similar properties in a single column.

Q1: How would you justify the presence of 18 elements in the 5th period of the Periodic Table?

Solution: For the value of n = 5, l can have values 0, 1, 2 and 3. The increasing order of the energy of the available orbitals is, 5s < 4d < 5p. The total number of orbitals available in valence shell of the fifth period is 9 namely 1 of 5s, 5 of 4d and 3 of 5p. And the maximum number of electrons that can be accommodated is 18; and therefore 18 elements can be present in the 5th period.

Q2: Two elements X and Y have atomic numbers 12 and 16 respectively. Write the electronic configuration for these two elements. And mention the period of the modern periodic table to which these two elements belong?

Solution: Electronic configuration of X with atomic number 12 is 2, 8, 2.

Electronic configuration of Y with atomic number 16 is 2, 8, 6.

Both these elements belong to the third period as the valence shell for both the elements as evident from the electronic configuration is the third shell.

4. Groupwise Electronic Configurations:

Elements in the same vertical column or group have similar valence shell electronic configurations, the same number of electrons in the outer orbitals, and similar properties. For example, the Group 1 elements (alkali metals) all have ns1 valence shell electronic configuration as shown below. Thus it can be seen that the properties of an element have periodic dependence upon its atomic number and not on relative atomic mass.

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	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^2 6p^6 7s^1$

5. Electronic Configurations and Types of Elements: s-, p-, d-, f- BLOCKS

The Aufbau (build up) principle and the electronic configuration of atoms provide a theoretical foundation for the periodic classification.

The elements in a vertical column of the Periodic Table constitute a group or family and exhibit similar chemical behaviour. This similarity arises because these elements have the same number and same distribution of electrons in their outermost orbitals. Thus, elements can be classified into four blocks viz., s-block, p-block, d-block and f-block depending on the type of atomic orbitals that are being filled with electrons. This is illustrated in Fig. 3. Two exceptions are noticed due to this categorization.

i) Strictly, helium belongs to the s-block but its positioning in the p-block along with other group 18 elements is justified because it has a completely filled valence shell (1s2) and as a result, exhibits properties characteristic of other noble gases.

ii) The other exception is hydrogen. It has only one s-electron and hence can be placed in group 1 (alkali metals). It can also gain an electron to achieve a noble gas arrangement and hence it can behave similar to a group 17 (halogen family) elements. Because it is a special case, it is placed separately at the top of the Periodic Table as shown in Fig. 2 and Fig. 3.

The salient features of the four types of elements marked in the Periodic Table

i). **s-Block Elements:**

- The elements of Group 1 (alkali metals) and Group 2 (alkaline earth metals) which have ns1 and ns2 outermost electronic configuration belong to the s-Block Elements.
- They are all reactive metals with low ionization enthalpies.
- They lose the outermost electron(s) readily to form 1+ ion (in the case of alkali metals) or 2+ ion (in the case of alkaline earth metals).
- The metallic character and the reactivity increases as we go down the group. Because of high reactivity they are never found pure in nature.
- The compounds of the s-block elements, with the exception of those of lithium and beryllium are predominantly ionic.

ii) **The p-Block Elements:**

- The p-Block Elements comprise of the elements that belong to Group 13 to 18 and these together with the s-Block Elements are called the Representative Elements or Main Group Elements.
- The outermost electronic configuration varies from ns2 np1 to ns2 np6 in each period.
- At the end of each period is a noble gas element with a closed valence shell ns2np6 configuration.
- All the orbitals in the valence shell of the noble gases are completely filled by electrons and it is very difficult to alter this stable arrangement by the addition or removal of electrons.
- The noble gases thus exhibit very low chemical reactivity.
- Preceding the noble gas family are two chemically important groups of non-metals.
- They are the halogens of Group 17 and the chalcogens of Group 16.

- These two groups of elements have highly negative electron gain enthalpies and readily add one or two electrons respectively to attain the stable noble gas configuration.
- The non-metallic character increases as we move from left to right across a period and metallic character increases as we go down the group.

iii) **The d-Block Elements also known as the Transition Elements:**

- These are the elements of Group 3 to group 12 in the centre of the Periodic Table.
- These are characterized by the filling of inner d orbitals by electrons and are therefore referred to as d-Block Elements.
- These elements have the general outer electronic configuration (n-1)d1-10 ns0-2.
- They are all metals.
- They mostly exhibit variable valence (oxidation states) and form coloured ions in solution.
- These show paramagnetism and are often used as catalysts. However, Zn, Cd and Hg which have the electronic configuration, (n-1)d10ns2 do not show most of the properties of transition elements.
- In a way, transition metals form a bridge between the chemically active metals of s-block elements and the less active elements of Groups 13 and 14 and thus take their familiar name "Transition Elements".

iv) **The f-Block Elements - Inner-Transition Elements**

- The two rows of elements at the bottom of the periodic table, called Lanthanides, Ce(Z = 58) – Lu(Z = 71) and Actinides, Th(Z = 90) – Lr (Z = 103).
- These are characterized by the outer electronic configuration $(n-2)f^{1-14}$ $(n-1)d^{0-1}$ ns^2.
- The last electron added to each element is filled in f- orbital.
- These two series of elements are hence called the Inner-Transition Elements (f-Block Elements) and these are all metals.
- Within each series, the properties of the elements are quite similar.
- The chemistry of the early actinides is more complicated than the corresponding lanthanides, due to the large number of oxidation states possible for the actinide elements.
- Actinide elements are radioactive and many of the actinide elements have been made only in nanogram quantities or even less by nuclear reactions and their chemistry is not fully studied. The elements after Uranium are called Transuranium Elements.

Let us now solve certain problems to enrich the concepts studied in this module.

Example: In which family / group would you place these elements with atomic numbers 117 & 120, though 120 element is not yet discovered? Also give the electronic configuration of each of the elements.

Solution: From the long form of periodic table, it can be deduced that the element with atomic number 117, would belong to the halogen family which is Group 17 and the electronic configuration would be like that of noble gas radon plus $5f^{14}\ 6d^{10}\ 7s^{2}\ 7p^{5}$.

The element with atomic number 120, will be placed in Group 2 which is alkaline earth metals, and will have the electronic configuration of $[Uuo]8s^{2}$.

6. Metals and Non-Metals:

In addition to classifying elements into s, p, d and f blocks, another broad classification of elements is based on their properties and the elements can be divided into Metals and Non-Metals.

A few important observations of metals are as below:

- Metals comprise of more than 78% of all known elements and appear on the left side of the Periodic Table.
- Metals are usually solids at room temperature. Few exceptions being mercury a semi solid and two metals gallium and caesium also have very low melting points namely 303 degree kelvin for gallium and 302 degree kelvin for caesium.
- Metals have high melting and boiling points.
- They are good conductors of heat and electricity.
- They are malleable, that means these can be flattened into thin sheets by hammering and ductile which means that these can be drawn into wires.
- In contrast to metals, non-metals are located at the top right hand side of the Periodic Table.
- In a horizontal row, the property of elements changes from metallic on the left to non-metallic on the right.
- Non-metals are usually solids or gases at room temperature with low melting and boiling points (boron and carbon are exceptions).
- They are poor conductors of heat and electricity.
- Most non-metallic solids are brittle and are neither malleable nor ductile.
- The elements become more metallic as we go down a group in the Periodic Table.

The change from metallic to non-metallic character is not abrupt as shown by the thick zig-zag line in Fig. but the elements silicon, germanium, arsenic, antimony and tellurium bordering this line and running diagonally across the Periodic Table show properties that are characteristic of both metals and non-metals. These elements are called Semi-metals or Metalloids.

Problems based on the electronic configuration and position of elements in the periodic table. Example: Considering the atomic number and position in the periodic table, arrange the following elements in the increasing order of metallic character: Si, Be, Mg, Na, P.

Solution: We know that the Metallic character increases down a group and decreases along a period as we move from left to right. Hence the order of increasing metallic character is:

P < Si < Be < Mg < Na.

7. Summary:

In this Module, we have discussed the application of the **Periodic Law.** Mendeleev's **Periodic Table** was based on atomic masses. Modern **Periodic Table** arranges the elements in the order of their atomic numbers in seven horizontal rows **(periods)** and eighteen vertical columns **(groups** or **families).** Atomic numbers in a period are consecutive, whereas in a group they increase in a pattern. Elements of the same group have similar **valence shell** electronic configuration and, therefore, exhibit similar chemical properties. However, the elements of the same period have incrementally increasing number of electrons from left to right, and therefore, have different valencies. Four types of elements can be recognized in the periodic table on the basis of their electronic configurations. These are **s-block, p-block, d-block** and **f-block** elements. **Hydrogen** with one electron in the 1s orbital occupies a unique position in the periodic table. **Metals** comprise more than seventy eight per cent of the known elements. **Non-metals** which are located in the top of the periodic table are less than twenty in number. Elements which lie at the border line between metals and non-metals (eg. Si, Ge, As) are called **metalloids** or **semi-metals.** Metallic character increases with increasing atomic number in a group whereas decreases from left to right in a period. The physical and chemical properties of elements vary periodically with their atomic numbers.