Classification of Elements and Periodicity in Properties

Ionization Enthalpy

A quantitative measure of the tendency of an element to lose electron is given by its **Ionization Enthalpy.** It represents the energy required to remove an electron from an isolated gaseous atom (X) in its ground state.

Energy is always required to remove electrons from an atom and hence ionization enthalpies are always positive.

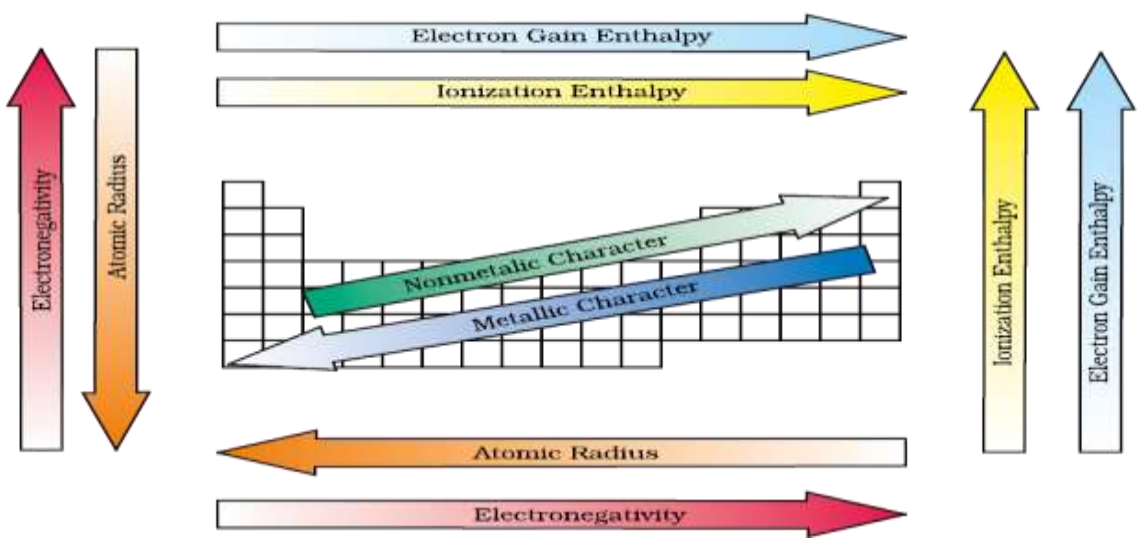
Ionization enthalpy generally increases as we go across a period and decreases as we descend in a group.

Electron Gain Enthalpy

Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion.
Depending on the element, the process of adding an electron to the atom can be either endothermic or exothermic.
As a general rule, electron gain enthalpy becomes more negative with increase in the atomic numbed across a period. The effective nuclear charge increases from left to right across a period and consequently it will be easier to add an electron to a smaller atom since the added electron on an average would be closer to the positively charged nucleus.

Electronegativity

A qualitative measure of the ability of an atom in a **chemical compound** to attract shared electrons to itself is called **electronegativity.**
Electronegativity generally increases across a period from left to right (say from lithium to fluorine) and decrease down a group (say from fluorine to astatine) in the periodic table.



The periodic trends of elements in the periodic table

Electronic configurations of elements and the periodic table

In the preceding unit we have learnt that an electron in an atom is characterized by a set of four quantum numbers, and the principal quantum number (n) defines that main energy level known as **shell.** We have also studied about the filling of electrons into different sub-shells, also referred to as **orbitals** (s, p, d, f) in an atom. 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 last orbital filled. In this section we will observe a direct connection between the electronic configurations of the elements and the long form of the Periodic Table.

(a) Electronic configurations in Periods – 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 filing of the next higher principal energy level (n=1, n=2, etc.). It can be really 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 ($1s^1$) and helium ($1s^2$) 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 configurations $1s^22s^2$. Starting from the next element boron, the 2p orbitals are filled with electrons with the L shell is completed at neon ($2s^22p^6$). 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 favourable and we come across the so called 3d **transition series** of elements. This starts from scandium (Z=21) which has the electronic configuration $3d^14s^2$. The 3d orbitals are filled at zinc (Z=30) with electronic configuration $3d^{10}4s^2$. 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 begins with cerium (Z=58) and ends at lutetium (Z=71) to give the 4f-**inner transition series** which is called the **lanthanoid 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 **actinoid series.** The 4f- and 5f-**innter 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.

(b) 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 ns^1 valence shell electronic configuration as shown below.

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	$[Urn]7s^1$

Thus it can be seen that the properties of an element have periodic dependence upon its atomic number and not on relative atomic mass.

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 behavior. This similarity arises because these elements have the same number and same distribution of electrons in their outermost orbitals. We can classify the elements 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 figure. We notice two exceptions to this categorization. Strictly, helium belongs to the s-blocks but its positioning in the p-block along with other group 18 elements is justified because it has a completely filled valence shell ($1s^2$) and as a result, exhibits properties characteristic of other noble gases. 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, we shall place hydrogen separately at the top of the Periodic Table as shown in figure 1 and fig. 2. We will briefly discuss the salient features of the four types of elements marked in the Periodic Table. More about these elements will be discussed later.

The s-Block Elements

The elements of Group 1 (alkali metals) and Group 2 (alkaline earth metals) which have ns^1 and ns^2 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 from 1+ion (in the case of alkali metals) or 2+ion (in the case of alkaline earth metals). The metallic character and the reactivity increase 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.

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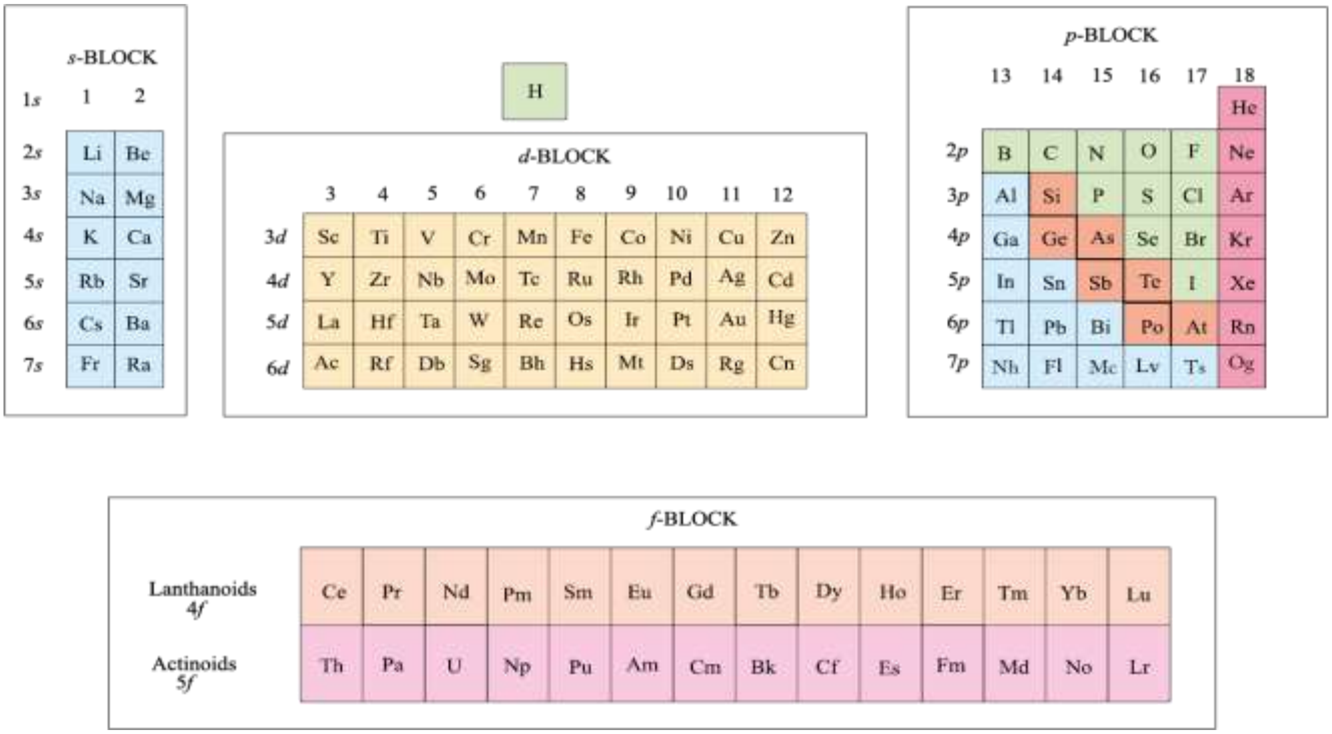

The types of elements in the periodic table based on the orbitals that are being filled. Also shown in the broad division of elements into METALS, NON-METALS and METALLOIDS

The p-Block Elements

The **p-Block Elements** comprise those belonging 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 ns^2np^1 to ns^2np^6 in each period. At the end of each period is a noble gas element with a closed valence shell ns^2np^6 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** (Group 17) and the **chalcogens** (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.

The d-Block Elements (Transition Elements)

These are the elements of Group 3 to 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 $(n-1)\ d^{1-10}ns^{1-2}$

except for Pd where its electronic configuration is $4d^{10}5s^{0}$. They are all metals. They mostly from coloured ions. Exhibit variable valence (oxidation states), paramagnetism and oftenly used as catalysts. However, Zn, Cd and Hg which have the electronic configuration, (n-1) $d^{10}ns^{2}$ 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".**

The f-Block Elements (Inner-Transition Elements)

The two rows of elements at the bottom of the Periodic Table, called the **Lanthanoids,** Ce (Z=58) – Lu (Z=71) and **Actinoids,** the outer electronic configuration (n-2) $f^{1\text{-}14}$ (n-1) $d^{0\text{-}1}ns^{2}$. 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).** They are all metals. Within each series, the properties of the elements are quite similar. The chemistry of the early actinoids is more complicated than the corresponding lanthanoids, due to the large number of oxidation states possible for these actinoid elements. Actinoid elements are radioactive. Many of the actinoid 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.**

Metals, Non-metals and Metalloids

The elements can be divided into **Metals** and **Non-metals.** Metals comprise more than 78% of all known elements and appear on the left side of the **Periodic Table.** Metals are usually solids at room temperature [mercury is an exception; gallium and caesium also have very low melting points (303K and 302K, respectively)]. Metals usually have high melting and boiling points. They are good conductors of heat and electricity. They are malleable (can be flattened into thin sheets by hammering) and ductile (can be drawn into wires). In contrast, non-metals are located at the top right hand side of the **Periodic Table.** In fact, in a horizontal row, the property of elements change 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; the non-metallic character increases as one goes from left to right across the **Periodic Table.** The change from metallic to non-metallic character is not abrupt as shown by the thick zig-zag line in Figure. The elements (e.g., 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.**