**Periodic Classification of Elements**

**Need for the Periodic Classification of Elements**

All existing matter in our surroundings is made up of basic units known as elements. Initially, in 1800, only 31 chemical elements were discovered. After some advancement in technology in 1865, about 63 more elements were discovered. This created the need for the periodic classification of elements.

Presently, there are ***118 elements known to us.*** Out of these 118 chemical elements, some elements are man-made.

* When there were only 31 elements it was relatively easy to study the properties of these chemical elements individually.
* Now the number of elements has raised to 118, it would be very cumbersome to study the properties of every element individual.
* In order to ease the work, scientists started thinking about a method so that the study of elements can be simplified.
* They decided to organize the elements in a periodic table according to the information available about the elements and various characteristics shown by them.

It was observed that elements show periodicity in their properties. Many tables were made to arrange the elements in an ordered manner based on their characteristics to study the properties of elements in a fixed pattern.

**Periodic Classification of Elements Characteristics**

In the long form periodic table the elements are arranged in the order of their atomic numbers. Atomic number of an element is equal to the number of protons inside the nucleus of its atom.

The general features of the long form periodic table are:

* There are in all, 18 vertical columns and 18 groups in the long form periodic table.
* These groups are numbered from 1 to 18 starting from the left.
* There are seven horizontal rows called periods in the long form periodic table. Thus, there are seven periods in the long form periodic table.
* The elements of Groups 1, 2 and 13 to 17 are called the main group elements. These are also called typical or representative or normal elements.
* The elements of Groups 3 to 12 are called transition elements.
* Elements with atomic number 58 to 71 (Ce to Lu) occurring after lanthanum (La) are called lanthanides. Elements with atomic numbers 90 to 103 (Th to Lw) are called actinides. These elements are called f-block elements and also as inner transition elements.

## Modern periodic table

In the year 1913, English physicist **Henry Moseley** studied the wavelength of the characteristic x-rays By using different metals as anti cathode and showed that the square root of the frequency of the line is related to the atomic number. On the basis of the above observations Moseley gave the **modern periodic law** which states that :

“Physical and chemical properties of the elements are the periodic function of their atomic numbers”.

* The modern periodic table consists of **18**vertical columns, called the **groups(1-18)** and **7** Horizontal rows, called **periods.**
* The first period contains **two elements**, Hydrogen and Helium.
* The second period contains **eight elements,**from Lithium to Neon.
* The third period contains **eight elements,**from Sodium to Argon.
* The fourth period contains**eighteen elements**, from Potassium to Krypton.
* The fifth period contains **eighteen elements,**from Rubidium to Xenon.
* The sixth period contains**thirty-two elements.**
* The seventh period is incomplete**.**
* On the basis of electronic configuration, elements are classified into **four Blocks** known as **s, p, d and f- blocks**.
* 1st and 2nd group elements are called**s-block**elements. The general electronic configuration is **ns1-2.**
* 13th to 18th group elements are called **p-block** elements. The general electronic configuration is **ns2 np1-6**.
* 3rd to 12th group elements are called**d-block**elements. The general electronic configuration is (**n-1)d1-10 ns1-2.**
* Lanthanides and actinides elements are called **f-block** elements. The general electronic configuration is **(n-2)f1-14 (n-1)d0-1 ns2**

## Periodic properties of elements

The basic law governing modern periodic table states that the properties of elements are periodic functions of their atomic number. These properties reappear at regular intervals or follow a particular trend at regular intervals. This phenomenon is known as the periodicity of elements.

### 1. Atomic and ionic Radii

#### Atomic Radii:

The atomic radius may be defined as the distance from the centre of the nucleus to the outermost shell containing electrons. Depending upon the nature of bonding in the atoms these are (i) Covalent radii (ii) van der Waals radii  (iii) Metallic radii

**(i) Covalent radii:** One-half of the distance between the centres of the nuclei of two adjacent similar atoms joined to each other by a single covalent bond is known as covalent radii. Eg Cl-Cl bond distance=198 pm covalent radius of Cl= 99 pm.

**(ii)van der Waals radii:** Half of the internuclear distance between two similar adjacent atoms belonging to the two neighbouring molecules of the same substance in the solid state is known as van der Waals radii.

**(iii) Metallic radii:**Half the distance between the centre of the  nuclei of two adjacent atoms  in the metallic crystal is known as metallic radii

As we move from left to right in a period, the atomic radius decreases due to an increase in [effective nuclear charge](https://byjus.com/questions/what-is-nuclear-charge-on-periodic-table/) (Zeff). Along the group, as we move from top to bottom, atomic radius increases  due to increase in principal quantum number which causes an increase in the number of shells and increases in [shielding effect](https://byjus.com/questions/what-is-shielding-effect-and-what-is-screening-effect/).

#### Ionic Radii:

The ions formed by the loss of one or more electrons from the neutral atom are known as cation (positive ion) when the electrons added to the neutral atom form an anion (negative ion). The effective distance from the centre of the nucleus of the ion upto which it exerts its influence on the electron cloud is known as the ionic radii.

The ionic radii change in the same trend as atomic radii. It decreases along the period from left to right and increases down the group from top to bottom. size of cation and anion of any natural atom as: cation< neutral atom < anion

### 2. Ionization enthalpy

The amount of energy required when an electron is removed from the outermost orbit of an isolated gaseous atom is known as Ionisation Enthalpy (IE).

Generally left to right in period IE increases whereas on moving down the group it decreases but half-filled orbital and fully filled orbitals are highly stable and thus have high IE.

**Factors Influencing Ionization Enthalpy**

The ionization enthalpy of an atom depends on the following factors.

**(i)  Size of the atom**

As the distance between the electron and the nucleus increases, i.e., as the size of the atom increases, the outermost electrons are less tightly held by the nucleus. Thus, it becomes easier to remove an outermost electron. Thus ionization enthalpy decreases with increases in atomic size.

**(ii) Charge on the nucleus**

Ionization enthalpy increases with increase in nuclear charge because of the increase in the attractive force between the nucleus and the electron.

**(iii) Screening effect of inner electrons**

Ionization enthalpy decreases when the shielding effect of inner electrons increases. This is because when the inner electron shells increases, the attraction between the nucleus and the outermost electron decreases.

**(iv) Penetration effect of electrons**

The penetration power of the electrons in various orbitals decreases in a given shell (same value of n) in the order: s>p>d>f. Since the penetration power of s-electron towards the nucleus is more, it will be closer to the nucleus and will be held firmly. Thus, for the same shell, the ionization enthalpy would be more to remove the s-electrons in comparison with the p-electron which in turn would be more than that for d-electron and so on.

**(v) Effect of half-filled and completely filled sub-levels**

If an atom has half-filled or completely filled sub-levels, its ionization enthalpy is higher than that expected normally from its position in the periodic table. This is because such atom, have extra stability and hence it is difficult to remove electrons from these stable configurations.

### 3. Electrons gain enthalpy

The electron gain enthalpy is defined as the change in enthalpy which takes place when a gaseous atom gains an extra electron to form a monovalent anion in the gaseous state.

Electron gain enthalpy increases across the periods while it decreases down the group.

Chlorine has the highest electron affinity than fluorine.

### 4. Electronegativity

Electronegativity is the tendency of an atom to attract the shared pair of electrons towards itself in a covalent bond. **Fluorine** is the most electronegative element while **Cesium** is the least.

In the periods left to right electronegativity increases. In the groups while moving down the groups electronegativity decreases.

**Hard and Soft Acids and Bases (HSAB) Principle** is a qualitative concept introduced by Ralph Pearson to explain the stability of metal complexes and the mechanisms of their reactions. However, it is also possible to quantify this concept based on Klopman's FMO analysis using interactions between HOMO and LUMO.

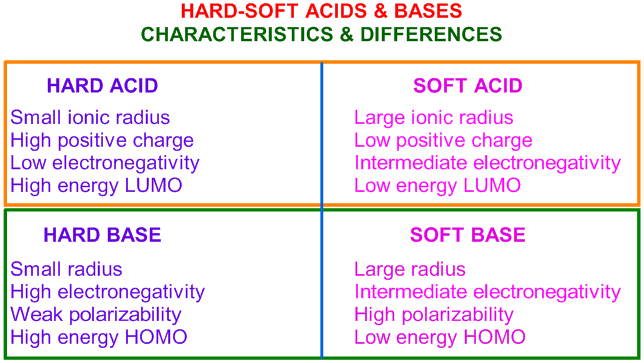
According to HSAB principle, the Lewis acids and bases can be further divided into hard or soft or borderline types.

**Hard Lewis acids** are characterized by small ionic radii, high positive charge, strongly solvated, empty orbitals in the valence shell and with high energy LUMOs.

**Soft Lewis acids** are characterized by large ionic radii, low positive charge, completely filled atomic orbitals and with low energy LUMOs.

**Hard Lewis bases** are characterized by small ionic radii, strongly solvated, highly electronegative, weakly polarizable and with high energy HOMOs.

**Soft Lewis bases** are characterized by large ionic radii, intermediate electronegativity, highly polarizable and with low energy HOMOs.



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| **Type of Acid/Base** | **EXAMPLES** |
| **Hard acids** | H+, Li+, Na+, K+, Be2+, Mg2+, Ca2+, Sr2+, Sn2+  Al3+, Ga3+, In3+, Cr3+, Co3+, Fe3+, Ir3+, La3+, Si4+, Ti4+, Zr4+, Th4+,U4+, VO2+ , UO22+  BeMe2, BF3, BCl3, B(OR)3, AlMe3 |
| **Soft acids** | Cu+, Ag+, Au+, Hg+ , Cs+ , Tl+ , Hg2+ , Pd2+, Cd2+ , Pt2+   Metal atoms in zero oxidation states  BH3 |
| **Borderline acids** | Fe2+ , Co2+ , Ni2+ , Cu2+ , Zn2+ , Pb2+ , B(CH3)3, SO2, NO+ |
| **Hard bases** | H2O, OH-, F-, Cl-, CH3CO2-, PO43-, SO42-, CO32-, NO3-, ClO4-, ROH, RO-, R2O, NH3, RNH2, N2H4 |
| **Soft bases** | S2-, RSH, RS-, R2S, I-, CN-, SCN-, S2O3-, R3P, R3As, (RO)3P, RNC, CO, C2H4, C6H6, R-, H- |
| **Borderline bases** | Aniline, pyridine, N3-, Br-, NO2-, SO32-, N2 |