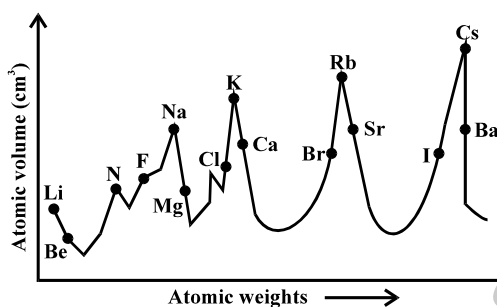


CLASSIFICATION OF ELEMENTS
AND PERIODICITY IN PROPERTIES

www.einsteinclasses.com
Free Quality Education

- C1(a) Dabereiner :** grouped elements in Triads. He pointed out that there were sets of three elements (Triads) which showed similar chemical properties. He also noted that the atomic weight of the central element of the Triad was approximately the mean of the atomic weights of the other two members. The properties of the middle element were in between of those end members. e.g. Li, Na, K, Ca, Sr, Ba or Cl, Br, I etc.
- (b) John Newland :** He developed a law of octaves.
He observed that similar elements are repeated at 8th place like the 8th note of music. The elements are arranged in the order of increase of atomic weights. Similar element means that physical & chemical properties of element will be same. e.g. Li has the same property as of Na.
- (c) Lothar Meyer arrangement :** He studied the physical properties such as atomic volume, melting point and boiling points of various elements. On this basis he plotted a graph between Atomic volume (cm^3) Vs. Atomic weights.



Observation :

- The most electropositive alkali metals (Li, Na, K, Rb and Cs) occupy the peaks on the curve.
- The less strongly electropositive elements i.e. alkaline earth metals (Be, Mg, Ca, Sr and Ba) occupy the descending positions on the curve.
- The most electronegative elements i.e., halogens (F, Cl, Br and I) occupy the ascending positions on the curve.

On the basis of these observations, Lothar Meyer proposed that the physical properties of the elements are a periodic functions of their atomic weights.

- (d) Mendeleev's Periodic law and table :** Mendeleev arranged all the elements in the order of increase of atomic weight.
- A table formed with the help of classification of elements is called periodic table.
 - The method of arranging similar elements in one group and separating them from dissimilar elements is called classification of elements.
 - He prepared the periodic table on the basis of periodic law i.e.,
"The properties of elements are periodic function of their atomic weight".
 - Mendeleev's periodic table consists of seven horizontal rows known as periods and nine vertical column known as groups.
 - Periods :** Out of seven periods, first three periods are short periods while the fourth, fifth and sixth periods are called long periods.
 - There are nine groups in all including 8th group of transition elements and zero group of inert gases.
 - All the group from I to VII (except zero and VIII) were divided into sub-groups.
 - The group number represents the valency.
 - The elements of same sub group resemble to each other more closely and differ from other sub groups.
 - He predicted the properties of the missing elements from the known properties of the other elements in the same group. e.g. gallium and germanium were not discovered at that time. He named these elements as Eka-Aluminium and Eka-silicon.

Defects in Mendeleev's Periodic Table :

1. **Position of Hydrogen :** Hydrogen resembles both the alkali metal and halogen. Hence its position in periodic table is undecided.
 2. **Position of Isotopes :** According to Mendeleev's periodic table, isotopes should occupy different positions in the periodic table, but this is not so.
 3. **Position of VIII group elements :** Nine elements in the VIII group do not fit into the system.
 4. **Positions of Lanthanides and Actinides :** Their position was not justified according to the periodic law and cannot be arranged in the order of their increasing atomic weight.
 5. **Dissimilar elements placed in the same group :** Alkali metals (Li, Na, K etc.) are placed with coinage metals (Cu, Ag, Au).
 6. **Similar elements placed Apart :** Chemically similar elements like Cu and Hg, Ag and Ti, Au and Pt have been placed in different groups.
 7. **Anomalous pair of elements :** Some elements of higher atomic weight have been placed before the elements of lower atomic weight. e.g. Argon (At. wt. = 39.9) has been placed before potassium (At. wt. = 39.1); cobalt (At. wt. = 58.94) is placed before nickel (At. wt. = 58.69); Tellurium (127.5) has been placed before iodine (126.9).
- (e) **Modern Periodic law and the present form of periodic table :** As a result of modern researches it is established that atomic number is a fundamental property not the atomic weight.

Thus this led Moseley to change the basis of classification of elements from atomic weight to atomic number. The modern periodic law given by Moseley is :

“The properties of elements are periodic functions of their atomic numbers, i.e., if elements are arranged in the order of their atomic numbers. Similar elements are repeated after regular intervals”. He also gave the following formulae

$$\text{i.e., } \sqrt{\nu} = a(z - b) \quad ; \quad \begin{array}{l} \nu = \text{frequency of X-rays} \\ a, b = \text{constant for all lines in a given series of X-rays} \\ z = \text{atomic number} \end{array}$$

Now we will discuss why elements with similar properties reoccurs after certain regular intervals.

Cause of Periodicity : The cause of periodicity in properties is the repetition of similar outer electronic configuration at certain regular intervals which indeed determines the physical and chemical properties of the elements and their compounds.

- (i) A modern version of table contains horizontal rows known as periods (which Mendeleev called series). Elements having similar outer electronic configuration in their atoms are grouped in vertical columns; these are referred as Groups or Families.
- (ii) According to IUPAC (International Union of pure and Applied Chemistry), the groups are numbered from 1 to 18 replacing the older notation of groups O, IA, II A VIII B.
- (iii) There are seven periods (three short periods and four long ones).
- (iv) The first period contains 2 elements. The subsequent periods consists of 8, 8, 18, 18 and 32 elements respectively.
- (v) The 7th period is incomplete.
- (vi) Till now elements upto 112 and 114 have been discovered.
- (vii) Elements with $z = 113, 115$ and beyond are not known.

C2 Arrangement of Elements in Periodic Table is according to Electronic Configuration of the Elements :**Electronic configuration in Periods :**

- (i) Each successive period in the periodic table is associated with the filling of next higher principal energy shell i.e. $n = 1, n = 2$ etc.
- (ii) Number of atomic orbitals in each period is twice the number of atomic orbitals available in the energy level is being filled.
- (iii) The first period starts with the filling of the lowest energy level (1s) and has the two elements – hydrogen ($1s^1$) and helium ($1s^2$).

- (iv) Second period starts with lithium.
- (v) The third period begins with sodium and added electrons enters a 3s orbital. This shell has nine orbitals (one 3s, three 3p and five 3d) but 3d orbitals are of higher energy than 4s according to $(n + l)$ rule. Therefore 3d orbitals are only filled after filling the 4s-orbitals. Thus it contains only eight elements.
- (vi) In the fourth period, the filling of electrons is in the fourth energy level i.e., $n = 4$. It starts with 4s. But in this filling of 4d and 4f orbital does not take place. After filling 4s orbital, filling of 3d orbital and then 4p orbitals takes place. 4d and 4f are of higher energy than 5s. Thus it contains 18 elements.
- (vii) In the fifth period the filling of electrons starts with 5s orbital and then 4d orbitals are filled and then three 5p orbitals are filled.

Sixth Period : Corresponds to filling of sixth energy level i.e. $n = 6$. In this period filling takes place in (one 6s, seven 4f, five 5d and three 6p) orbitals. There are 16 orbitals are available, thus 32 elements can be there.

- (viii) Filling up of the 4f orbital begins with cerium ($z = 58$) and ends at lutetium ($z = 71$) to give the 4f-inner transition series called the Lanthanoid Series. They are placed at the bottom of the periodic table.
- (ix) Similar to the sixth period, the seventh period corresponds to the filling of seventh energy shell i.e. $n = 7$. It is also expected to contain thirty two elements corresponding to filling of sixteen orbitals i.e. one 7s, seven 5f, five 6d and three 7p.
- (x) Filling up of 5f orbitals after actinium ($z = 89$) gives the 5f inner transition series known as Actinoid Series. Thus 4f, 5f transition series of elements are placed separately in the periodic table to maintain its structure.

Group wise electronic configuration

- (xi) Elements in the same vertical column or group have similar electronic configuration, have same number of electrons in the outer orbitals and have similar properties.

C3 Types of elements : s-, p-, d-, f- blocks

-s Block elements : These elements contain 1 or 2 electrons in the s-orbital of their respective outer most shell.

- (i) The elements of group 1 having outer-most electronic configuration ns^1 are called as alkali metals.
- (ii) The elements of group 2 having outer-most electronic configuration ns^2 are called alkaline earth metals.
- (iii) **Properties of s- block elements :** They are reactive metals with low ionization energy.
- (iv) They lose the outermost electron readily to form +1 oxidation state in case of alkali metals.
- (v) Alkaline earth metals can lose two electrons to acquire +2 oxidation state very easily.
- (vi) Metallic character and reactivity increases as we move down the group.
- (vii) The compounds of s-block are predominantly ionic with exception of Be (beryllium)

P-block Elements : These elements contain 1-6 electrons in the P-orbital of their respective outermost shells.

- (i) General electronic configuration of outermost shell is ns^2np^{1-6} , where $n \rightarrow 2-7$.
- (ii) These include elements belonging to group of 13, 14, 15, 16, 17 and 18 excluding helium.

Properties of P-Block Elements :

- (i) All the orbitals in the valence shell of the noble gases are completely filled by electrons. Thus it is difficult to alter this stable arrangement by addition or removal of electrons.
- (ii) Thus noble gases exhibit low chemical reactivity.
- (iii) Group - 17 elements are known as halogens and then have high negative electron gain enthalpy.
- (iv) Similarly group - 16 elements i.e., oxygen family are also known as chalcogens, have high tendency to gain electrons.
- (v) The non-metallic character increases as we move from left to right in a period.
- (vi) Metallic character increases as we move down the group.
- (vii) Their ionization energies are higher than s-block elements.
- (viii) They mostly form covalent compounds.
- (ix) Some of them show more than one oxidation states in their compounds.

d-Block Elements : These elements characterise by filling of inner d-orbital by electrons and therefore referred to d-Block elements.

- (i) General electronic configuration of d-Block elements is $(n - 1) d^{1-10} ns^{0-2}$.
- (ii) They are the elements belonging to 3 to 12 groups.

Properties of d-Block Elements :

- (i) Their ionization energies are between 5 and p-block elements.
- (ii) They show variable oxidation states.
- (iii) They form both ionic and covalent compounds
- (iv) Their compounds are generally coloured and paramagnetic.
- (v) Most of transition metals form alloys.
- (vi) Most of transition elements are used as catalyst i.e., V, Cr, Mn, Fe, Co, Ni, Cu etc.

Exception : Zn, Hg, Cd have $(n - 1)d^{10}ns^2$ electronic configuration, do not show most of properties of transition elements as they have complete d subshell and in them last electron enters the s-subshell not the d-subshell.

Similarly of Zn, Cd, Hg with transition elements are :

- (i) they form complexes like d-block elements.
- (ii) they form covalent compounds.
- (iii) first ionization energies are much higher

f-Block Elements : The two rows at the bottom of periodic table called Lanthanoids and Actinoids have outer most electronic configuration as : $(n - 2)f^{1-14}(n - 1)d^{0-1}ns^2$

- (i) Last electron enters the f-subshell.
- (ii) Thus two series elements are called f-Block elements or inner-transition elements.

Properties of f-block elements :

- (i) They are all metals.
- (ii) They show variable oxidation states.
- (iii) Their compounds are generally coloured.
- (iv) Most of the elements of actinide series are radioactive. In this series after uranium elements are called as transuranium elements.
- (v) Classification of periodic table can be the basis of metals and non-metals.

C4 Properties of Metals

- (i) They are good conductors of heat and electricity
- (ii) They are usually solids at room temperature
- (iii) They are malleable (can be flattened into thin sheets by hammering) and ductile (can be drawn into wires)

Properties of Non-Metals

- (i) They are bad conductor of heat and electricity
- (ii) They are usually solids or gases at room temperature
- (iii) They cannot be drawn into thin sheets or wires.
- * Metallic character increases down the group.
- * Metallic character decreases along the period.
- * The change from metals to non-metals in periodic table is not abrupt. The elements on the border line are semi-metals or metalloids e.g. (Si, As, Sb, Te).

C5 Prediction of period, group and block of given elements :

- (a) Period of an element corresponds to principal quantum number of the valence shell.
- (b) The block of an element is type of orbital which receives the last electron.
- (c) Group is predicted as follows :
 - (i) for s-block elements : group number is equal to the number of valence electrons.
 - (ii) for p-block elements : group number is equal to 10 + number of valence electrons.
 - (iii) for d-block elements : group number is equal to number of electrons in $(n - 1) d$ subshell + number of electrons in valence shell (nth shell)

C6 Periodic trends in properties of elements :

1. **Atomic radii** : is the distance from the centre of nuclei to the point upto which the density of electron cloud is maximum.

It is of four types :

- (i) covalent radii
- (ii) vander wall radii
- (iii) metallic radii
- (iv) Ionic radii

Covalent radii : r_{covalent} = $\frac{1}{2}$ [Internuclear distance between two covalently bonded atoms of same molecule]
= $\frac{1}{2}$ [bond length]

Vander wall radii : $r_{\text{vander wall}}$ = $\frac{1}{2}$ [Internuclear distance between two non-bonded atoms] of different or neighbouring molecules

Metallic radii : = $\frac{1}{2}$ [internuclear distance between two adjacent atoms in the metallic lattice]

Variation of Atomic radii in periodic table :

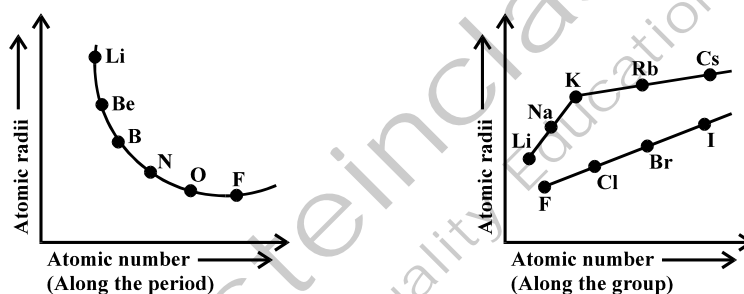
1. Along the period it decreases as nuclear charge increases.

Exception :

The size of atoms of inert gases are however larger than the halogen elements.

Reason :

As we move along the period charge increases while the number of shells remain the same thus along the period size decreases.



2. Along the group the atomic radii of elements increases with increase in atomic number as we move from top to bottom of group.

Reason : As we move down the number of shell increases thus distance between the outer electrons and nucleus increases. Also with increase of atomic number the nuclear charge down the group increases. Thus atomic radii should decrease, but effect of increased nuclear charge is reduced due to screening or shielding effect on the valence electrons by the electrons present in the inner shells.

- Comparison of the ionic radii with atomic radii of same atom – $\text{Mg} > \text{Mg}^{2+}$

Reason : In Mg $n_p = 12, n_{e^-} = 12$

Mg^{2+} $n_p = 12, n_{e^-} = 10$

Thus in Mg 12 protons are attracting by the 12 electrons outer nuclear part. Whereas in Mg^{2+} the 12 protons are attracting 10 e^- outside the nucleus thus 12 protons will attract 10 e^- electrons more strongly thus size of Mg^{2+} decreases as compared to Mg.

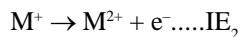
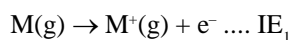
- * Atoms or ions with same electronic configuration are called isoelectronic. If we consider a series of isoelectronic species (atoms or ions) then size decreases with increase of atomic number.

e.g.	O^{2-}	I,	Ne,	Na^+	Mg^{2+}
	(1.40)	(1.36)	(1.31)	(0.95)	(0.65)
Atomic number	8	9	10	11	12

- * The size of same cation decreases with increase of magnitude of positive charge i.e. Fe^{3+} is smaller than Fe^{2+} , Cu^{2+} is smaller than Cu^+ so on.....

Ionisation Energy

First I.E. : Amount of energy required to remove a single electron from the outer shell of a neutral gaseous atom.



(IE₂) Second Ionisation Energy : Amount of energy required to remove the second electron from an atom who has already lost one electron.

$$\text{IE}_3 > \text{IE}_2 > \text{IE}_1$$

As the number of electrons in outer shell decreases so attraction of nucleus for remaining electrons increases thus I.E. increases.

Trends of first I.E. along the period and down the group is as follows :

- (a) First I.E. decreases as we move down the group. Reason is that as we move down the group no. of shell increases thus nuclear attraction decreases. Although nuclear charge also increases but its effect is weakened by shielding supplied by the inner shells to the outer most shell.
- (b) As we move along the period ionisation energy increases. The reason for this is because of decrease of size of atom along the period. As the number of electrons are added to same shell along the period thus nuclear attraction for the outer electrons increases with increase of nuclear charge.

Exception :

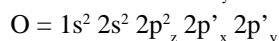
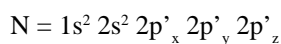
- (1) I.E. (B) < I.E. (Be)

Electronic Configuration : B = $1s^2 2s^2 2p^1$, Electronic configuration Be = $1s^2 2s^2$

As in Be e^- is to be removed from completely filled shell thus its I.E. is more than B. Here is extra stability of subshell is the cause of this irregularity

- (2) I.E. of O < I.E. N

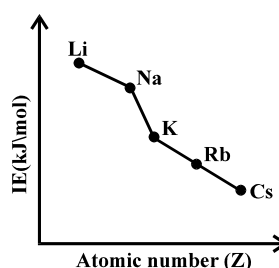
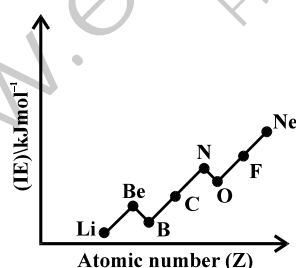
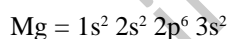
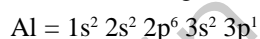
Electronic configuration



The extra stability of half-shell is cause of irregularity in IE pattern.

- (3) I.E. of Al < I.E. of Mg

Electronic configuration

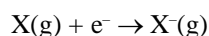


- * The electrons in different orbitals (s, p, d, f) belonging to the same energy level experience different pull of the nucleus.

The I.E. for pulling out an s-electron is maximum and it decreases in pulling out p-electron.

Hence we can say that I.E. for pulling out an electron from a given energy level decreases in the order $s > p > d > f$ orbitals.

Electron Affinity : It is defined as energy given out when an extra electron is taken up by a neutral gaseous atom.



E. A. measures the tightness with which an atom binds an extra electron to it.

E.A. decreases down the group, because an atom gets larger, the attractions of positive nucleus for an outside electron decreases.

Exception : E.A. of F < E.A. of Cl

Reason : It is due to extremely small size of F atom as compare to Cl atom.

Extra electron create strong electron-electron repulsion among all the electrons.

Along the period : Along the period electron affinity increases as we move from left to right.

- (i) E.A. of noble gases are zero.
- (ii) E.A. of N and Be atoms are quite low due to extra stability of half filled orbitals (p^3 in N & s^2 in Be)
- (iii) After taking up an extra electron an atom becomes negatively charged (anion) and now second electron is to be added to it. The anion will repel the incoming of an electron and an additional energy will be required to add it to the anion.

First E.A. \rightarrow negative (energy released)

Second E.A. \rightarrow positive (energy released)

Electronegativity : It is the measure of the ability of an atom in a combined state (i.e. in a molecule) to attract itself to the electrons within a chemical bond or tendency to attract bond pair towards itself.

- * Non-metals have high value of electronegativity than metals.
F, O, N & Cl are highly electronegative than [K, Rb, Cs] (metals) which are electropositive in nature.
- * Electronegativity along the period increases & down the group decreases.
- * It is measured with the help of Pauling scale & Mulliken scale.

Screening (shielding) effect :

In d-block elements (transition element) while writing electronic configuration of elements, it is seen that new electrons are added to inner shells i.e. penultimate shells. Thus nuclear attraction for out electrons gets affected. As the new electrons enter the inner shells they tend to shield or screen the outer shell electrons from nucleus and thus decreases the nuclear attractive force. This is called as screening effect. Due to this effect, the atomic size of transition elements remain same as we move along the period which should decrease as we move from left to right. Thus ionisation energy, electron affinity and other properties remains nearly same as we move along the period.

C7 Hydration and Hydration Energy :

Hydration energy is the enthalpy change that accompanies the dissolving of one mole of gaseous ions in water.



- (i) Size of the ion and its charge determines extent of hydration.
- (ii) Greater the charge, smaller the size of the ion, greater the attraction for the lone pair of O of H_2O , hence greater the extent of hydration and hence greater the hydration energy.
 - size of the hydrated ions increases,
 - ionic mobility decreases [heavier (hydrated) ions moves slower]

C8 Acid-Base character of oxides :

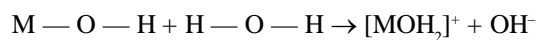
On moving across a period, the basic character of the oxides gradually changes first into amphoteric and finally into acidic character.

- (i) On moving down the group, reverse behaviour is observed, i.e., from more acidic to more basic.
Oxides of the element M in H_2O produce MOH

- If electronegativities difference of M and O is greater than that of H and O in H_2O then MOH is acidic due to formation of H_3O^+

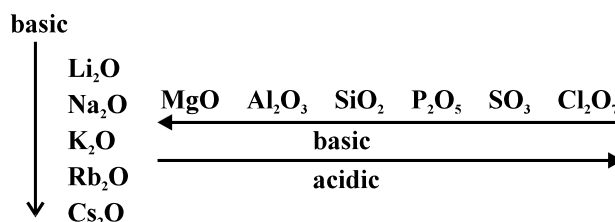


- If electronegativities difference of M and O is less than that of H and O in H_2O then MOH is basic due to formation of OH^-



- (ii) Stability of oxides decreases across a period.

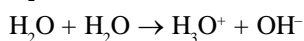
(Min. \rightarrow Max)



- (iii) Oxides of the following elements are amphoteric

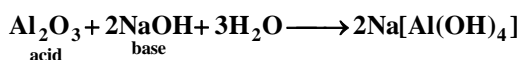
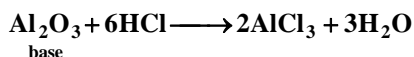
H, Be, Al, Ga, In, Tl, Sn, Pb, Sb, Bi, Po

- H_2O is amphoteric (also called amphiprotic)



Since it is H^+ acceptor (base) as well as H^+ donor (acid).

- BeO , Al_2O_3 , SnO_2 , PbO_2 ,.... are amphoteric since they form salts with acid as well as with base



- oxide is acidic if it reacts with a base.
- oxide is basic if it reacts with an acid.