PERIODIC CLASSIFICATION OF ELEMENTS

Lavoisier Classification

Lavoisier classified the elements as metals and non-metals depending upon the physical properties such as malleability, ductility, lustre, hardness etc.
This method of classification was not found to be satisfactory because in this type of classification, there was no place for "the elements having properties intermediate between metals and non metals. In addition to this, all the elements classified as metals do not have similar properties. For example, sodium and lead were classified as metals. But their chemical properties were entirely different from one another. Thus, this classification was found to be inadequate.

Dobereiner's Law of Triads

In 1829, German chemist Johann Dobereiner made an attempt to arrange elements on the basis of atomic masses. He prepared a law known as "Dobereiner's Law".

It states that **"When elements are arranged in increasing order of atomic masses in a triad, the atomic mass of the middle element is equal to the arithmetic mean of the other two elements."**
Thus he classified all the elements in groups of three elements. These groups of three elements were called as triads. The examples for triads are **Li, Na, K** and **Cl, Br, I**.
But in some triods, all the three elements possessed nearly the same atomic mass, hence the law was rejected.

Newland's law of octaves

In 1864, John Newland, an English chemist and musician, arranged all the known elements in order of increasing atomic masses. He observed that "the properties of every eighth element following it". In other words it was just like the repetition of the eighth mode in an octave in music as shown below.

Sa	Re	Ga	Ma	Pa	Dha	Ni	Sa
1	2	3	4	5	6	7	8

On this basis, Newland proposed a law for classification of elements which is known as Newland's law of octaves.

According to this law "When elements are arranged in the order of increasing atomic masses, the properties of the eighth element are similar to the properties of the first element."
Newland's arranged the elements in horizontal rows of seven elements as shown ahead.

Elements arranged according to Newland's Law of Octaves.

Element	Li	Be	B	C	N	O	F
Element	Na	Mg	Al	Si	P	S	Cl
Musical	sa	Re	ga	ma	pa	dha	ni
Notes	sa						

If we start with lithium, then according to Newland's law, properties of the eighth element sodium should be similar to that of lithium. Practically, it has been found that lithium and sodium have similar chemical properties.

Again, if we take fluorine as the first element, then according to Newland's law, the properties of the eighth element chlorine should be similar to the properties of fluorine. It has been actually found that chlorine and fluorine have similar chemical properties.

Merits: (i) This is system of classification worked well for lighter metals (only upto atomic number 20).
(ii) Newland's classification gave an important idea to other scientists that there is periodicity in the properties of the elements. Thus, he need for arranging elements in vertical columns and horizontal rows became evident.

Limitations: (i) Newland's law of octaves could classify the elements up to calcium only.
(ii) It could not be applied to heavier elements.
(iii) When the noble gases were discovered, the idea of octaves could not be held.

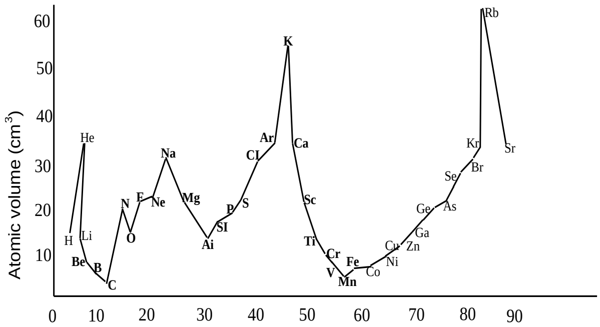
Lothar meyer's atomic volume curve

In 1869, a German chemist, Julius Lothar Meyer calculated atomic volumes of elements by dividing atomic mass of the element by its density.

$$\text{Atomic Volume} = \frac{\text{Atomic mass}}{\text{density}}$$

He plotted the atomic volumes of elements against their atomic masses and obtained a curve. The curve was called as Lothar Meyer's atomic volume curve.

From the curve of atomic volumes, Lothar Meyer observed that the elements having similar chemical properties occupy the similar positions in each part of the curve.



Lother Meyer's atomic volume curve

For example

(i) Alkali metals-Sodium (Na), Potassium (K) and rubidium (Rb) having similar chemical properties occupy the similar positions in the curve. They all occur at peaks of the curve.
(ii) Halogen-fluorine (F), chlorine (Cl) and bromine (Br) having similar properties are present at similar positions (troughs) in the curve.

Mendeleev's Periodic Law

In 1869, a Russian Scientist D'mitri I. Mandeleef studied the properties of the elements and concluded that atomic mass is the fundamental property of elements.
The law states that, **"The properties of the elements are a periodic function of their atomic masses"**.

Mendeleev's periodic table

Mendeleev prepared a tabular chart and arranged all the known 63 elements in the table in the order of increasing atomic masse. His periodic table comprised of eight vertical columns and seven horizontal rows.

Mendeleev's Periodic Table

Period	Series	Group I R_2O A B	Group II (RO) A B	Group III R_2O_2 A B	Group IV RH_4 R_2O_5 A B	Group V RH_3 R_2O_3 A B	Group VI RH_2 RO_3 A B	Group VII RH R_2O_7 A B	Group VIII - RO_4
1	1	H = 1							
2	2	Li = 7	Be = 9.4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	3	Na=23	Mg=24	Al=27.5	Si = 28	P = 31	S = 32		Fe = 56
	4	K = 39	Ca=40	? = 40	Ti = 48	V = 51	Cr =	Cl=35.5	Co= 59
4							52	Mn = 55	Ni =
	5	Cu=63.5	Zn=65	? = 68	? = 72				58.7
	6	Rb = 85	Sr=87	?Yi=88	Zr = 90	As=75	Se	Br	
5						Nb=94	= 78	=80	Ru=104
							Mo =	? = 100	Rh=104
	7	Ag=108		IN =113	Sn =118		96		Pd=106
6	8	Cs=133	Cd=112	$?D_1$=138	?Ce=140				
			Ba=137	-	-	Sb=122		I = 127	
	9	(-)		$?E_r$=178	?La=180	-	Te =	-	
7	10	-				-	27.6	-	Os=195
			-			Ta =182	-	-	Ir=197
	11				Pb=207				Pt=198
	12	Au=199	Hg=200	TL=204	Th=231		-	-	
		-	-	-		Bi	W =		- -
						=208	184	-	
						-			
							-		
							U =		
							240		

While preparing the periodic table, Mandeleef laid more stress on the similarity in the properties of the elements, i.e., grouping of similar elements together. The order of increasing atomic masses was not rigidly followed. He had to place some elements with higher atomic mass before the elements with lower atomic masses. For example, cobalt (at. Mass 59) and nickel (58.71), tellurium (127.6) and iodine (126.9) were placed in inverted order of their atomic masses. This was done due to similarities in chemical properties.

One of the striking applications of Mendeleev's classification was that he **left certain gaps** in his periodic table. He thought that the elements which fit for these gaps would be discovered later on. He even predicted the properties for those unknown elements. It is interesting to known that the missing elements were discovered later on and their properties were found very close to those predicted by Mendeleev.

For example, Mendeleev found that Titanium (Ti) has properties similar to those of Silicon (Si), So he placed Titanium below Silicon. This left a gap below Aluminium. He predicted that the element which would have properties similar of boron and aluminum is yet to be discovered. He named the element as eka- aluminum i.e. an element with properties similar to that of aluminum. This element was discovered five years later named Gallium (Ga). Since it has properties similar to aluminum, it was placed below aluminum. In addition to this eka-boron (Scandium), eka-silicon (Germanium) and eka-manganese (Technetium) were predicted by Mendeleev.

Merits of Mendeleev Periodic Table

1. Formation of groups and sub groups
2. Scope for discovery of new elements
3. Correction of the atomic masses of many elements

For example, beryllium was assigned an atomic mass of 13.5 but was arranged before carbon (atomic mass 12) in the table. It was suggested that the reported mass of beryllium is incorrect. Later on it was found to be so. The correct mass of beryllium is 9.

Discrepancies in Mendeleev's Periodic table. It is not able to explain

1. The position of the hydrogen and the position of the isotopes
2. Incorrect grouping of some elements
3. Position of some dissimilar elements in the same group
   For example, alkali metals (Li, Na, K, etc) have properties entirely different from those of coinage metals (Cu, Ag, Au). But they are placed in the same group.
4. No distinction between metals and non-metals

Modern Periodic Law

In 1913, Henry Crwyn Jefferys Moseley proved, on the basis of ingenious experiments that atomic number(s) is the fundamental property of an element. Since chemical properties of an elements depend on electronic configuration, therefore atomic number should be the basis of classification. This led to the idea of modern periodic law. It states that

" The properties of the elements are periodic function of their atomic numbers". In other words if the elements are arranged in an increasing order of atomic number, then after a regular interval, the properties of the elements are repeated. Thus a modern periodic table was formulated based on the atomic number.

Periodicity and Its Cause

The repetition of properties of elements after regular intervals when the elements are arranged in the increasing atomic numbers in called periodicity.Now the question arises, why there is periodicity in the properties of elements when they are arranged in increasing order of atomic number. This can be explained as follows:

Cause of Periodicity

We know that the chemical properties of an element depend on the number of electrons in the outermost shell (valence electrons). Now, since electronic configuration of outermost shell is repeated after a regular interval, chemical properties of elements are also repeated. Since electronic configuration depends on the atomic number. We can say that properties of elements are periodically repeated when elements are arranged in increasing order of atomic number.

Modern Periodic Table or Long Form of Periodic Table

The modern periodic table was presented by Bohr. It is based on the modern periodic law, In this table, elements are arranged in order of increasing atomic number in such a way that the elements having the same number of valence electrons come in the same group. Though many tables are in existence, the most satisfactory form, the separate "long form" is shown. The table is known as modern periodic table has 7 horizontal rows called periods and 18 vertical columns called groups.

Periodic Table of the Elements (Long Form)

s-Block Elements · d-Block Elements · p-Block Elements

Group: New Notation (1–18); CAS Version (IA–VIIA, 0 (Zero))

STATE: Gas G · Liquid L · Solid S · Not found in nature X

Key to Chart: G 8 O Oxygen 15.9994 — Atomic number, Symbol, Name, Atomic mass

Period	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIII	9 VIII	10 VIII	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 0(Zero)
1	1 G **H** Hydrogen 1.008																1 G **H** Hydrogen 1.008	2 G **He** Helium 4.003
2	3 S **Li** Lithium 6.941	4 S **Be** Beryllium 9.0121											5 S **B** Boron 10.811	6 S **C** Carbon 12.011	7 G **N** Nitrogen 14.007	8 G **O** Oxygen 15.999	9 G **F** Fluorine 18.998	10 G **Ne** Neon 20.180
3	11 S **Na** Sodium 22.990	12 S **Mg** Magnesium 24.305											13 S **Al** Aluminium 26.982	14 S **Si** Silicon 28.086	15 S **P** Phosphorus 30.914	16 S **S** Sulphur 32.066	17 G **Cl** Chlorine 35.453	18 G **Ar** Argon 39.948
4	19 S **K** Potassium 39.098	20 S **Ca** Calcium 40.078	21 S **Sc** Scandium 44.956	22 S **Ti** Titanium 47.867	23 S **V** Vanadium 50.942	24 S **Cr** Chromium 51.996	25 S **Mn** Manganese 54.938	26 S **Fe** Iron 55.845	27 S **Co** Cobalt 58.933	28 S **Ni** Nickel 58.693	29 S **Cu** Copper 63.546	30 S **Zn** Zinc 65.39	31 S **Ga** Gallium 69.723	32 S **Ge** Germanium 72.61	33 S **As** Arsenic 74.922	34 S **Se** Selenium 78.96	35 L **Br** Bromine 79.904	36 G **Kr** Krypton 83.30
5	37 S **Rb** Rubidium 85.468	38 S **Sr** Strontium 87.62	39 S **Y** Yttrium 88.906	40 S **Zr** Zirconium 91.224	41 S **Nb** Niobium 92.906	42 S **Mo** Molybdenum 95.94	43 X **Tc** Technetium (98)	44 S **Ru** Ruthenium 101.07	45 S **Rh** Rhodium 102.906	46 S **Pd** Palladium 106.42	47 S **Ag** Silver 107.868	48 S **Cd** Cadmium 112.411	49 S **In** Indium 114.818	50 S **Sn** Tin 118.710	51 S **Sb** Antimony 121.760	52 S **Te** Tellurium 177.60	53 S **I** Iodine 126.904	54 G **Xe** Xenon 131.29
6	55 S **Cs** Cesium 132.905	56 S **Ba** Barium 137.327	57 S **La*** Lanthanum 138.906	72 S **Hf** Hafnium 178.49	73 S **Ta** Tantalum 180.948	74 S **W** Tungsten 183.84	75 S **Re** Rhenium 186.207	76 S **Os** Osmium 190.23	77 S **Ir** Iridium 192.217	78 S **Pt** Platinum 195.078	79 S **Au** Gold 196.967	80 L **Hg** Mercury 200.59	81 S **Tl** Thallium 204.383	82 S **Pb** Lead 207.2	83 S **Bi** Bismuth 208.980	84 S **Po** Polonium (209)	85 S **At** Astatine (210)	86 G **Rn** Radon (222)
7	87 S **Fr** Francium (223)	88 S **Ra** Radium (226)	89 S **Ac**** Actinium (227)	104 X **Rf** Rutherfordium (261)	105 X **Db** Dubnium (262)	106 X **Sg** Seaborgium (263)	107 X **Bh** Bohrium (262)	108 X **Hs** Hassium (265)	109 X **Mt** Meitnerium (266)	110 X **Uun** Ununnilium (269)	111 X **Uuu** Unununium (272)	112 X **Uub** Ununbium (277)	113	114 X **Uuq** Ununquadium	115	116 X **Uuh** Ununhexium		

Metals · Metalloids · Nonmetals

f-Block Elements

*Lanthanides	58 S **Ce** Cerium 140.116	59 S **Pr** Praseodymium 140.908	60 S **Nd** Neodymium 144.908	61 X **Pm** Promethium (145)	62 S **Sm** Samarium 150.36	63 S **Eu** Europium 151.964	64 S **Gd** Gadolinium 157.25	65 S **Tb** Terbium 158.925	66 S **Dy** Dysprosium 162.50	67 S **Ho** Holmium 164.930	68 S **Er** Erbium 167.26	69 S **Tm** Thulium 168.934	70 S **Yb** Ytterbium 173.04	71 S **Lu** Lutetium 174.967
**Actinides	90 S **Th** Thorium 232.038	91 S **Pa** Protactinium 231.036	92 S **U** Uranium 238.029	93 X **Np** Neptunium (237)	94 S **Pu** Plutonium (244)	95 X **Am** Americium (243)	96 X **Cm** Curium (247)	97 X **Bk** Berkelium (247)	98 X **Cf** Californium (251)	99 X **Es** Einsteinium (252)	100 X **Fm** Fermium (257)	101 X **Md** Mendelevium (258)	102 X **No** Nobelium (259)	103 X **Lr** Lawrencium (262)

Note : (1) The new IUPAC format numbers the group from 1 to 18. The previous IUPAC numbering system and the system used by Chemical Abstracts Service (CAS) are also shown. For radioactive elements that do not occur in nature, the mass number of the most stable isotope is given in brackets with the longest half-life.

(2) The symbols for elements 104-109 used in this table are those proposed by the American Chemical Society and 110-116 proposed by IUPAC.

Main Features of the Modern Periodic Table

Periods

The long form of periodic table has 7 horizontal rows called periods. In each period, the elements have consecutive atomic numbers i.e., atomic number of elements increases by one of unit from left to right. First element of every period is an alkali metal and the last element is an inert gas. The number of elements in different periods is different. First period contains 2 elements (Very short period), second and third period contains 8 elements (Short period), fourth and fifth contains 18 elements(Long period), sixth period contains 32 elements (Very long period) and seventh is incomplete period.

It should be noted that the elements in a period have

(i) Different number of valence electrons. The first element of period has 1 valence electrons whereas the last element (except He) has 8 valence electron.

(ii) Each element has different valence electrons i.e., they have different electronic configuration and hence the elements in a period have different chemical properties.

(iii) The elements of first period have only one shell, the K shell that can accommodate maximum of two electrons. The elements of second and third periods have respectively 2 (K and L) and 3 (K, L and M) shells and so on. The electronic configuration of first two periods given.

(iv) As we move along a period from left to right (Li to Ne), the L shell gets progressively filled with electrons. Similarly, from 3 rd period (Na to Ar) the M shell, for 4th period the N shell is filled and so on.

Groups

The long form of periodic table has 18 vertical columns known as groups. These groups are named as I A to VII A, I B to VIII B and zero groups. Group IA is in the extreme left whereas zero group is on the extreme right side of the periodic table. According to the latest periodic table the groups are given running numbers from 1 to 18.

(i) Group I A to VII A or group 1, 2, 13 to 17 have normal elements. It means that their inner shells are complete and only outermost shell is incomplete. The group number of each element is the same as the number of valence electrons in its atom. For example, Na has 1 valence electron, in its atom. For example, Na has 1 valence electron, it is placed in group I A or 1. All the alkali metals (Li, Na, K, Rb, Cs, Fr) are placed in group I since they have one valence electron. Similarly, halogens (F, Cl, Br, I) have seen valence electrons and hence are placed in group VII A or group 17.

(ii) The elements of zero group are inert gases: The outermost shell of all these elements is complete. All of them (except He) have 8 electrons in the outermost shell.

(iii) The elements of group I B to VIII B or group 3 to 12 are known as transition elements: In these elements the penultimate shell (the shell inner to outermost shell) is also incomplete. Therefore, these elements show variable valency.

(iv) Lanthanides and Actinides are placed in two separate horizontal rows below seventh period.

Merits of the long form of periodic table

1. In long form of the periodic table, elements are arranged in increasing order of atomic number which is the most fundamental property of elements.
2. Long form of periodic table explains why elements in the same group show similar chemical properties. This is because elements in the same group have similar electronic configuration. Therefore, they have similar chemical properties. Whereas elements with different electronic configuration are placed in different groups, hence they differ in chemical properties.
3. No separate position of isotopes is required in long form of periodic table. This because long form of periodic table is based on the atomic number of elements. Since isotopes have the same atomic numbers, they should be placed together. Hence no separate place is required for isotopes.

4. There is a clear separation of metals, non-metals and metalloids in long form of periodic table.
5. Normal elements transition elements and noble gases find separated positions in long form of periodic table.

Defects of Long Form of Periodic Table

1. Position of Hydrogen : Position of hydrogen in controversial in the periodic table also. Hydrogen is a non-metal but is placed with metals in the periodic table.
2. Position of Helium : On the basis of electronic configuration, helium should be placed in group II, but it is placed in zero with noble gases on the basis of its properties.

Advantages of Periodic Table

1. It has become easier to remember the properties of elements : This is because the elements with similar properties are placed in the same group in periodic table. Therefore, by learning the properties of groups only, one can remember the properties of elements.
2. The nature of bonds formed by an element can be known by knowing its position the periodic table. For example, if an element is placed among the metals, it will form ionic compounds. Whereas, a non-metal element with form covalent compounds.
3. Many elements have been discovered with the help of periodic table. Mandeleef predicted the existence and expected properties of some unknown elements on the basis of gaps in his periodic table. These elements were discovered later on.
4. Atomic masses of many elements are corrected on the basis of periodic table. Mandeleef arranged the elements on the basis of their atomic weights. In doing so anomalies were found atomic weights of elements. Later on, atomic weights of these elements were corrected.
5. A periodic table is useful in teaching and studying chemistry of elements and their compounds.

Prediction of Properties of Elements Using Periodic Table

If the position of an element in the periodic table is known, we can predict many of the properties. This is because modern periodic table is based on electronic configuration of elements. Therefore, by knowing position of elements in periodic table, we can know its electronic configuration and can predict its properites, for example :

- **Valence electrons:** The group number of an element tells the number of valency electrons.
- **Number of shells:** The periodic number of an element is equal to the number of electronic shells in its atom.
- **Valency of the element:** By knowing the number of valence electrons, the valency of element can be calculated.

Illustration 1: ***An element X belongs to third period and group II or 2 of the periodic table. State (i) No. of valence electrons (ii) valency (iii) metallic or non-metallic character (iv) electronic configuration.***

Solution: (i) We know that number of valence electrons in an atom of an element is equal to its group number. Since the element X belongs to group II, it has 2 valence electrons.

(ii) Since valencies of elements belonging to group I, II and III are equal to the number of valence electrons, the valency of element X will be 2.

(iii) Elements on the left hand side on the periodic table are metals. Since group II is on the left side of the periodic table, the element X is a metal.

(iv) Group number of an element is equal to the number of valence electrons. Hence the number of electrons in the outermost shell is 2.

The period number of elements is assigned on the basis of number of electron shell filled. Therefore, as the element X has 3 electron shells. Thus electronic configuration of element X will be : 2, 8, 2.

Illustration 2: Electronic configuration of an element M is 2, 8, 7. State giving reasons (i) its position in the periodic table (ii) whether it is a metal or a non-metal.

Solution : (i) Group number of an element is equal to the number of valency electrons. Since the element M has 7 electrons in the outermost shell, i.e., number of valency electrons is 7, It should be placed in group VII of periodic table.
(ii) We know that period number of an element is equal to the number of shells. Since group VII elements are non-metals, element M is a non-metal.

Periodicity in Properties

Properties of elements depend upon the electronic configuration.
As the elements of the same group have same electronic configuration of the valencies shell, they have similar chemical properties. It is observed that there is a regular variation in the properties of elements when we move from left to right in a period or we move down in a group. This regular variation in the properties is called as periodicity. Due to this periodicity in the properties of elements, this table is known as periodic table. Now we will discuss the variation in properties of elements in periods and groups.

1. Atomic Size (Atomic radii):

The atomic size means the radius of an atom. It is defined as the distance between the centre of nucleus and the outermost shell of an isolated atom.

(i) For metal atoms, atomic radius is taken as half of the inter nuclear distance between the two metal ions in a metallic crystal.
(ii) For non metals the atomic radii is taken as half of inter nuclear distance between two atoms bound by a single covalent bond.

Variation of a Period

Atomic size of elements decreases when we move from left to right in a period. Atomic radii of inert gases are abnormally large in their respective periods.
The units of atomic radii is expressed in Angstroms. (1A = 10^{-10} m)
we know that atomic size decreases from left to right in a period. This is because the number of electrons and protons increase as we move from left to right. Due to increased number of of protons in the nucleus, the nuclear charge increases on electrons. This increased nuclear charge pulls the electrons more closer towards the nucleus and hence the radii of the atoms decrease.

Variation of a Group

Atomic size of elements increases on moving from top to bottom in a group. The first element of a group is the smallest in size and that last element is the largest one. For example, in group I, lithium is the first element, it has the smallest atom. As we move down, the atomic size increases and the last element, francium (Fr) has the largest atomic radius. The increase in size of atoms on moving down a group is due to the increase in number of shells. As we move down in a group, the number of electrons shells also increases and hence the size increases.

In group VII or group 17, fluorine atom is the smallest and iodine atom is the largest.

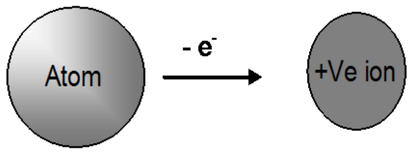
2. Ionic Radius

Cation

A cation (positive ion) is formed by the electron(s) from an atom, usally in order to attain a noble gas configuration.

Na → Na^+ + e^-
(2, 8, 1) (2, 8)

Mg → Mg^{2+} + $2e^-$
(2, 8, 2) (2, 8)
Al → Al^{3+} + $3e^-$
(2, 8,3) (2, 8)

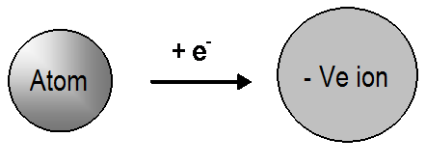


After the formation of a cation, the nuclear charge exceeds the electronic charge. Therefore, the force of attraction between the electrons and the nucleus increases. So the cation is smaller than the atom.

Anions
An anion (negative ion) is formed by the addition of electron(s) to an atom with a view to attaining a noble gas configuration. As the electrons increase in number, the repulsion among them also increases. This results in an increases in size. For example, the radius of the sulphur atom is 102 pm, while that of S^{2-} is 170 pm.

F + e^- → F^-
(2, 7) (2, 8)

Cl + e^{--} → Cl^-
(2, 8, 7) (2, 8, 8)
O + $2e^-$ → O^{2-}
(2, 6) (2, 8)
S + $2e^-$ → S^{2-}
(2, 8, 6) (2, 8, 8)



Periodic Trend
The radii of ions of similar charge follow the same periodic trend as the atomic radii of the parent elements. In other words, ionic radius (for ions of similar charge) increases down a group and decreases down a group and decreases from left to right in a period.

Isoelectronic species: are those which have same number of electrons but different nuclear charge.

Atom or Ion	Atomic No. Z	No. of electrons(e)	z/e ratio	Size in Å
C^{4-}	6	10	0.6	2.60
N^{3-}	7	10	0.7	1.71
O^{2-}	8	10	0.8	1.40
F^-	9	10	0.9	1.30
Ne	10	10	1.0	1.12
Na^+	11	10	1.1	0.95
Mg^{++}	12	10	1.2	0.65
Al^{+++}	13	10	1.3	0.50

H^+ and Cs^+ are the smallest and largest cation respectively H^- and I^- are the smallest and largest anion respectively.

Illustration 3: ***Compare the size of Cl, Cl^-, and Cl^+ ion***

Solution: $\frac{Z}{e}$ ratio for Cl = $\frac{17}{17}$ = 1.00

$Cl^- = \frac{17}{18} = 0.944$

$Cl^+ = \frac{17}{16} = 1.0625$

$Cl^- > Cl > Cl^+$

3. Ionisation Energy (IE)

It is defined as the amount of energy (in kcal or kJ) which is required to remove an electron completely from an isolated gaseous atom or ion of an element.

$$M(g) + I.E. \longrightarrow M^+(g) + e^-$$

An atom has as many ionization energies as it has electrons.

- Energy required to remove the first electron is called first ionisation energy ($I.E_1$). This changes the neutral atom to a unipositive ion.
- Now when the second electron is to be removed it has to be done from a unipositive charged ion. It is called second ionisation energy or successive ionisation energy ($I.E_2$).

The atom of energy required for this process ($I.E_2$) is greater than the first ionisation (1st I.E.). It is because with the removal of first electron, the nucleus force of attraction on the remaining electrons increase. Thus,

3rd I.E. > 2nd I.E. > 1st I.E. or $IE_3 > IE_2 > IE_1$

Factors effecting ionization energy

1. **<u>Atomic size:</u>** Smaller the size of atom more strong attraction exist between the nucleus and valence electrons. Thus removal of electron needs greater energy. Hence ionization energy will be more. Larger the size of atom ionization energy will be less. Ionization energy is inversely proportional to atomic size.
2. **<u>Effective nuclear charge :</u>** Ionization energy is directly proportional to effective nuclear charge(Zeff). Greater the nuclear charge more will be the ionization energy.
3. **<u>Shielding effect :</u>** Ionization energy is inversely proportional to shielding effect (screening effect) ,thus greater the shielding effect lesser the ionization energy.
4. **<u>Penetration power of orbital :</u>** The order of penetration of orbital is s>p>d>f. Thus removal of electron from 's' orbital is difficult than 'p' orbital, removal of electron from 'p' orbital is difficult compared to 'd' orbital.
5. **<u>Electronic configuration :</u>** Exactly half filled and completely filled orbitals are highly stable due to symmetry and exchange energy. Thus np^3 , np^6 , nd^5 , nd^{10} are highly stable . Thus ionization energy will be high in those atoms.

Variation in a Period

The ionization energy increases from left to right in a period. It is because with the increase in atomic number the nuclear force of attraction increases. The ionization energy of the elements of second period are given in the table below.

First I.E. of Elements of Period 2.

Elements of 2nd period	Li	Be	B	C	N	O	F
I.E. (kJ Mol^{-1})	500	900	801	1085	1400	1314	1680

Li < Be > B < C < N > O < F (expected – Li < Be < B < C < N < O < F)

Variation in a Group

The ionisation energy decreases from top to bottom. It is because the atomic size increases in a group. As a result of this the force of attraction between the nucleus and valence electronic decreases and hence the ionization energy. This idea is illustrated in the table given below.

First Ionisation Energy (kJ Mol^{-1}) of Elements of Group 1

Element of Group 1	Na	K	Rb	Cs
Ionisation Energy (kJ Mol^{-1})	496	420	403	376

Na < Mg > Al < Si < P > S < Cl (expected- Na < Mg < Al < Si < P < S < Cl)

4. Electron Affinity (EA)

It is defined as '' the energy released when an electron is added to an outer most orbit of an isolated neutral gaseous atom. It is a measure of the attraction or affinity of the gaseous atom for an extra electron.

$$Cl\,(g) + e^- \longrightarrow Cl^-\,(g) + \text{Energy}$$

Factors effecting electron affinity

1. **Atomic size:** Smaller the size of atom more strong attraction exist between the nucleus and valence electrons. Thus addition of electron releases greater energy. Electron affinity (Electron gain enthalpy will be more negative) will be more. Larger the size of atom electron affinity will be less. Electron affinity is inversely proportional to atomic size.
2. **Effective nuclear charge :** Electron affinity is directly proportional to effective nuclear charge(Zeff). Greater the nuclear charge more will be the electron affinity.
3. **Shielding effect :** Electron affinity is inversely proportional to shielding effect (screening effect) ,thus greater the shielding effect lesser the electron affinity.
4. **Penetration power of orbital :** The order of penetration power of orbital is s>p>d>f. Thus addition of electron to 's' orbital is easier than 'p' orbital, addition of electron to 'p' orbital is easier compared to 'd' orbital.
5. **Electronic configuration :** Exactly half filled and completely filled orbitals are highly stable due to symmetry and exchange energy. Thus np^3 , np^6 , nd^5 , nd^{10} are highly stable . Thus the atoms with half filled and complete filled orbital have less electron affinity.

Variation in a Period

The electron affinity increases from left to right along a period. It is because as the atomic size decreases, the force of attraction of nucleus with electrons increases and therefore electron affinity. In second and third period electron affinity order is

Li > Be < B < C > N < O < F (expected order – Li < Be < B < C < N < O < F)

Na > Mg < Al < Si > P < S < Cl (expected oerder- Na < Mg < Al < Si < P < S < Cl)

Variation in a Group

The electronic affinity decreases from top to bottom in a group. It is because as the size increases the nuclear force of attraction decreases.

For halogens Cl > F > Br > I (Expected – F > Cl > Br > I)

For chalcogens S > Se > O > Te (Expected- O > S > Se > Te)

It is important to note that some elements like Beryllium (Be) of group 2, Nitrogen, Phosphorous, Arsenic etc. of group 15 and noble gases have zero electron affinity values.

5. Electronegativity

The relative tendency of an atom to attract the shared pair electron towards itself is called electro-negativity. There is no unit of electronegativity. It depends upon its ionisation potential and electron affinity values. Higher ionisation potential and higher electron affinity values imply higher electronegativity value.

Electronegativity values are important to make certain prediction.

i) ***Nature of the bond between two atoms***

a) If $X_A – X_B = 0$, bond is purely covalent

b) If $X_A – X_B$ difference is lesser than1.7 the bond is polar covalent.

c) If $X_A – X_B$ is 1.7 the bond is 50% covalent and 50% ionic.

d) If $X_A – X_B$ is higher than 1.7 the bond is more ionic and less covalent.

ii) ***Stability of the bond***, greater the difference, of the electron negativity, more stable is the bond

Electro negativities of Representative Elements

S – block							
I	**II**	**III**	**IV**	**V**	**VI**	**VII**	**0**
H 2.1							**He 0**
Li 1.0	**Be 1.5**	**B 2.0**	**C 2.5**	**N 3.0**	**O 3.5**	**F 4.0**	**Ne 0**
Na 0.9	**Mg 1.2**	**Al 1.5**	**Si 1.8**	**P 2.1**	**S 2.5**	**Cl 3.0**	**Ar 0**
K 0.8	**Ca 1.0**	**Ga 1.6**	**Ge 1.8**	**As 2.0**	**Se 2.4**	**Br 2.8**	**Kr 0**
Rb 0.8	**Sr 1.0**	**In 1.7**	**Sn 1.8**	**Sb 1.9**	**Te 2.1**	**I 2.5**	**Xe 0**
Cs 0.7	**Ba 0.9**	**Tl 1.8**	**Pb 1.8**	**Bi 1.9**	**Po 2.0**	**At 2.2**	**Rn 0**

Periodicity in Electronegativity

1. In a period moving from left to right, electro negativity value increases.

2. In a period the electronegativity value of 1A alkali metal is minimum and that of VII A halogen is maximum.

3. In a group moving from top to bottom, the electro negativity decreases because atomic radius increases.

4. The electronegativity values of F is maximum and that of Cs is minimum in the periodic table.

5. The electronegativity of Cs (55) should be more than Fr (87) but in practice it is less. This is due to the fact that there in increase of 32 units in nuclear change of Fr which makes the value of effective nuclear charge quite high.

6. The electronegativity of inert gas elements of zero group is zero.

6. Metallic and Non-Metallic Properties

In the long form of the periodic table, metals are present on the left hand side whereas non metals on the right hand side. There is a line of demarcation between metals and non metals. The bordering elements boron, silicon, germanium, arsenic, antimony, tellurium and polonium are intermediate in properties (metals and non metals) and are called Semimetals or metalloids. The metallic character of an element is expressed in terms of its tendency to lose electronic (electronic character) whereas the non-metallic character in terms of gaining of electrons (electronegative character).

Variation in a Period

On moving from left to right in a period, the metallic character of elements decreases and non-metallic character increases.

Metals have a tendency to form positive ions by losing electrons, i.e., they are electropositive, whereas non-metals are electronegative, i.e., they form negative ions by accepting electrons. Since metallic character decreases from left to right in a period, we can say that electropositive character decreases and electronegative character increases from left to right in a period.

Elements of 3rd Period Na Mg Al Si P S Cl

——— electropositive character decreases ⟶

——— electronegative character increases ⟶

Variation in a Group

In a group when we move from top to bottom we observe that the size increases. As a result of this the force of attraction between the nucleus and outermost electron decreases. In other words, element with large size will lose electrons easily, hence, it will be more electropositive or more metallic.

In group I, lithium is least metallic (least electropositive) and francium is most metallic (most electropositive), that means metallic character increases and non metallic character decreases from top to bottom in a group.
It should also be noted that on the other hand, the non-metallic character or the electronegative character decreases on moving down in a group.

In group, VII or 17, fluorine is the most electronegative element whereas iodine is least electronegative.
The decrease in electronegative character also can be explained on the basis of atomic size. Since electronegative character depends on the tendency of accepting electrons, the element with the smallest size will be most electronegative because it will accept electrons easily. This is because there is more attraction between nucleus and electrons and hence electrons are accepted easily. Thus, an element with smallest atomic size will be most electronegative.

On the other hand, as the atomic size increases (due to addition of new shells) the attraction between nucleus and electrons decreases and hence the elements will not accept electrons easily to form negative ion. In other words, it will be less electronegative. Therefore, element with largest atomic size will be least electronegative.

7. Nature of Oxides

On moving from left to right in a period, basic nature of oxides decreases and acidic nature of oxides increases.

Oxides of Elements	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_2	Cl_2O_7
Nature of Oxides	Strongly Acidic	Basic	Amphoteric	Neutral	Acidic	Highly acidic	Strongly acidic

Basic nature of oxides decreases ⟶   Acidic nature of oxides increases ⟶

We know that oxides of metal are basic whereas oxides of non-metals are generally acidic in nature. Now, since metallic character decreases and non-metallic character increases from left to right in a period, the basic nature of oxides decreases and acidic nature increases.

Illustration 4: The atomic radii of element A, B and C belonging to the same period are given below :

Elements	***A***	***B***	***C***
Atomic Radii (pm)	***125***	***162***	***190***

Answer the following questions : (i) Which element will have the highest atomic number and which element will have the lowest atomic number ? (ii) Arrange these elements in decreasing order of metallic character.

Solution : We know that atomic radii of elements decreases from left to right in a period. Since these elements belong to the same period, the correct order of these elements from left to right on the basis of atomic radii should be. C B A

⟵ 190 162 125 ⟵

(i) We know that atomic number increase when we move from left to right in a period. Hence C will have the lowest atomic number and A will have the highest atomic number.

(ii) We know that metallic character of elements decreasing from left to right in a period. Hence C is most metallic and A is least metallic.
Thus the order of decreasing metallic character is, C > B > A.

Illustration 5. The atomic number of 2 elements A and B are 13 and 20 respectively. Write their electronic configuration and identify the group/groups to which these elements belong.
Solution: The electronic configurations of A and B are

A : 2, 8, 3 B : 2, 8, 8, 2

We find that A and B have the different number of valence electrons, so they will be placed in the difference groups. A belongs to group 13 whereas B belongs to group 2.
ASSIGNMENTS

SUBJECTIVE

LEVEL - I

1. State Dobereiner's law of triads and Newland's law of octaves.
2. What's Modem periodic law ?
3. State Mendeleef's periodic law
4. Give two examples of elements discovered after Mandeleef gave the periodic table.
5. How does the nature of oxides vary in going from left to right in a periodic table ?
6. How the value of ionisation energy increases in a period ?
7. Why electron affinity of non metals is more than metals ?
8. Why atomic size decreases along a period ?
9. The atom of an elements X has 5 valence electrons. What is the group number of element X. What is the formula of its oxide ?
10. What do you understand by the term periodicity ? Are the properties of the elements placed in a group same ? Illustrate

LEVEL - II

1. (a)Discuss the position of hydrogen in the long form of periodic table.
   (b) An element X is in the second group of the periodic table. Give the formula of its oxide.
2. Define ionisation energy. First ionisation energy of two elements A and B are 500 kJ mol^{-1} and 375 kJ mol^{-1} respectively. Comment on their relative positions in a group as well as in a period.
3. An atom has the electronic configuration of 2, 8, 7
   (a) What is the atomic number of the periodic table ?
   (b) To which of the following element would it be chemically similar ? (atoms numbers are given). N(7), F(9), P(15), Ar(18)
4. Given below are the melting points and the atomic radii of three elements X, Y and Z of the periodic table, each having n electrons in the outermost shell of their atoms :

Elements	X	Y	Z
Melting Points (^{o}C)	180.5	97.5	63.4
Atomic radii (A^{o})	1.23	1.57	202

   Answer the following :
   (a) Do these elements X, Y and Z belong to the same group or to the same period?
   (b) Which element will be least metallic ?
5. The atomic numbers of some elements in a period are : x = 17, y = 11, z =15
   (i) Calculate their valencies
   (ii) Explain with reasons which of them will be most metallic and which element will be most metallic and which element will be least metallic in nature.

OBJECTIVE

LEVEL - I

1. Which of the following is the most active metal?
(A) Al (B) K (C) Na (D) Mg
2. Among the following oxides, which is the most acidic?
(A) SiO_2 (B) P_4O_{10} (C) SO_3 (D) Cl_2O_7
3. Which of the following elements has the smallest atomic size?
(A) H (B) Li (C) Na (D) K
4. Which of the following species should have largest size?
(A) F (B) Cl (C) Br^- (D) Br^-
5. The biggest ion among the following is (Ba At n° 56)
(A) Al^{3+} (B) Ba^{2+} (C) Mg^{2+} (D) Na^+
6. Which of the following elements, which has the highest ionisation energy?
(A) N (B) O (C) F (D) Cl
7. Which of the following bonds is nonpolar?
(A) H-Cl (B) F-F (C) O-H (D) H-F
8. Which of the following elements is the most electropositve?
(A) Na (B) K (C) Cs (D) C
9. Which of the following oxides is expected to be the most basic?
(A) H_2O (B) Na_2O (C) MgO (D) K_2O
10. Which of the following elements has highest electron affinity
(A) F (B) O (C) Cl (D) N

LEVEL - II

1. Which of the series of elements listed below would have nearly the same atomic radii?
(A) F,Cl,Br,I (B) Na,K,Rb,Cs (C) Li,Be,B,C (D) Fe,Co,Ni,Cu
2. In which of the following pairs is the second atom larger the first?
(A) Br,Cl (B) Na, Mg (C) Sr,Ca (D) N,P
3. The electrons affinities of halogens F,Cl,Br,I decrease in the order
(A) F > Cl > Br > I (B) Cl > F > Br > I (C) I > Cl > F > Br (D) F > Br > Cl > I
4. Which of the following electronic configuration have more atomic size?
(A) $1s^22s^22p^63s^23p^5$ (B) $1s^22s^22p^3$ (C) $1s^22s^22p^5$ (D) $1s^22s^22p^63s^1$
5. The correct order of increasing electron affinity of the following elements is
(A) C < Be < N < O < F (B) Be < C< O < N < F
(C) Be < N < C < O < F (D) Be < C < N < F < O
6. Which of the following elements do not being to Group 2?
(A) Ra (B) Ca (C) Sr (D) Ga
7. Which of the following elements are inner transition elements?
(A) Cesium (B) Tungsten (C) Thorium (D) Copper
8. Arrange the elements with the following electronic configurations in increasing order of atomic sizes.
(i) $1s^22s^22p^5$(F) (ii) $1s^22s^22p^4$(O) (iii) $1s^22s^22p^3$(N) (iv) $1s^22s^22p^63s^23p^4$(S)
(A) ii < i < iii < iv (B) iv < iii < ii < i (C) iii < ii < iv < i (D) i < ii < iii < iv
9. In which of the following are the orders of electron affinity correct?
(A) S > O (B) O > S (C) C < Si (D) N > P
10. Among the following, the most electronegative element will have the outermost electronic configuration
(A) ns^2np^3 (B) ns^2np^6 (C) ns^2np^4 (D) ns^2np^5

LEVEL - III

1. The pair of atomic numbers which belong to the same group
(A) 9, 14 (B) 17, 51 (C) 6, 53 (D) 12, 56
2. Mendeleef's periodic table is based on the
(A) Atomic weight (B) atomic number (C) atomic radius (D) atomic volume
3. Largest ion among the following is
(A) Na^+ (B) O^{-2} (C) S^{-2} (D) Cl^-
4. Which one of the following is diagonally related pair?
(A) Be, Al (B) Li, Mg (C) B, Si (D) All the above
5. The law of octaves applies to
(A) B, C, N (B) As, K, Ca (C) Be, Mg, Ca (D) None
6. Among the following metals, which one has the weakest metallic character?
(A) Li (B) Na (C) K (D) Cs
7. In the long term of periodic table, all the non-metals are placed in
(A) s–block (B) p–block (C) d–block (D) f–block
8. The cause of periodicity of properties is
(A) Increasing atomic radius
(B) Increasing atomic weight
(C) Number of electrons in the valency orbit
(D) The recurrence of similar outer electronic configuration
9. Highest ionization potential in a period is shown by
(A) alkali metals (B) transition elements (C) halogens (D) noble gases
10. The properties of the elements, as well as the formulae and properties of their compounds depend in a periodic manner on the atomic weights of the elements. This periodic law was given by
(A) J.L. Meyer (B) John A.R. Newlands (C) Dobernier (D) D.I. Mendeleev
11. Ionization potential in a period is lowest for
(A) halogens (B) inert gases
(C) alkali metals (D) alkaline earth elements
12. The law of triads is applicable to
(A) C, N, O (B) Li, Na, K (C) Cl, Br, I (D) Both (B) and (C)
13. The atoms of elements belonging to the same group of periodic table have the same number of
(A) protons (B) electrons
(C) neutrons (D) electrons in the outermost shell
14. The ionization potential in a group from top to bottom
(A) decreases (B) increases
(C) remains the same (D) increases and decreases
15. The correct order of relative sizes is
(A) $I^- > I^+ > I$ (B) $I^- > I > I^+$ (C) $I > I^+ > I^-$ (D) $I^+ > I^- > I$

ANSWERS

OBJECTIVE

LEVEL - I

1.	**B**	2.	**D**	3.	**A**	4.	**D**
5.	**B**	6.	**C**	7.	**D**	8.	**C**
9.	**D**	10.	**C**				

LEVEL - II

1.	**D**	2.	**D**	3.	**B**	4.	**D**
5.	**C**	6.	**D**	7.	**C**	8.	**D**
9.	**A**	10.	**D**				