5 PERIODIC CLASSIFICATION

5.1 INTRODUCTION

At present 120 elements are known to us. Around the year 1800, only 30 elements were known. All these had seemingly different properties.

As different elements were being discovered, scientists gathered more and more information about the properties of these elements. They found it difficult to organise all that was known about the elements. They started looking for some pattern in their properties. On the basis of which they could study such a large number of elements with ease.

5.2 WHY DO WE NEED TO CLASSIFY ELEMENTS ?

Before the begining of eighteenth century, only a very few elements were known and it was quite easy to study and remember their individual properties. However, the situation became difficult with the discovery of large number of elements in the later years. At this stage the scientists felt the need of some simple method to facilitate the study of the properties of various elements and their compounds. After numerous attempts the scientists were ultimately successful in arranging the elements in such a way so that similar elements were grouped together and different elements were separated.

5.3 MAKING ORDER OUT OF CHAOS-EARLY ATTEMPT AT THE CLASSIFICATION OF ELEMENTS :

5.3.1 Dobereiners triads

(i) He made groups of three elements having similar chemical properties called **TRIAD**.

(ii) In Dobereiner triad, at wt. of middle element is equal to the average atomic weight of first and third element.

Dobereiner's Triads :

Triad **Relative atomic**

	Arithmetic mean of atomic masses of	
	masses respectively	**first and the third elements.**
S, Se, Te	32, 79, 128	$\frac{32+128}{2} = 80$
Cl, Br, I	35.5, 80, 127	$\frac{35.5+127}{2} = 81.25$
Ca, Sr, Ba	40, 88, 137	$\frac{40+137}{2} = 88.5$

5.3.2 Limitations

Doberenimer could identify only three triads from the elements known at that time. hence this system of classification into triads was not found to be useful for classifying many other elements which were not able to form any traids like all three previous triads.

Illustration 1

X, Y, Z are three members of a Dobereiner's triad. The atomic mass of X is 7 μ and that of Y is 23 μ. What is the atomic mass of Z?

Solution

The triads is X, Y, Z. Let the atomic mass of Z be *x*, then according to Dobereiner.

$$\text{Atomic mass of Y} = \frac{\text{Atomic mass of X} + \text{Atomic mass of Z}}{2}$$

$$23\,\mu = \frac{7\mu + x}{2}$$

$x = 2 \times 23\,\mu - 7\mu = 46\,\mu - 7\,\mu = 39\,\mu$

Therefore, the atomic mass of Z is 39 μ.

5.4 NEWLAND LAW OF OCTAVES:

(i) He arranged the elements in the increasing order of their atomic mass and observe that properties of every 8th element was similar to the 1st one. Like in the case of musical vowels notation.

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

(ii) At that time inert gases were not known.

Sa (do)	re (re)	ga (mi)	ma (fa)	pa (so)	da (la)	ni (ti)
H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe
Co and Ni	Cu	Zn	Y	In	As	Se
Br	Rb	Sr	Ce and Na	Zr	–	–

(iii) The properties of Li are similar to 8th element i.e. Na, Be are similar to Mg and so on.

Drawback or Limitation :

- It was found that the **law of octaves** was applicable only upto calcium, as after calcium every eighth element did not possess properties simlar to that of the first.
- It was assumed by **Newlands** that only 56 elements existed in nature and no more elements would be discovered, whose properties did not fit into the Law of Octaves.
- In order to fit elements into his table, Newlands adjusted two elements in the same slot, but also put some unlike elements under the same note. Can you find examples of these from Table? Note that cobalt and nickel are in the same slot and these are placed in the same column as fluorine, chlorine and bromine which have very different properties than these elements, Iron, which resembles cobalt and nickel in properties has been placed far away from these elements.

  Thus, Newland's law of Octaves worked well with lighter elements only

*5.5 LOTHER MEYER'S CURVE

(i) He plotted a curve between atomic weight and atomic volume of different elements.

(ii) The following observation can be made from the curve-

(a) Most electropositive elements i.e. alkali metals (Li, Na, K, Rb, Cs etc.) occupy the peak positions on the curve.

(b) Less electropositive i.e. alkali earth metal (Be, Mg, Ca, Sr, Ba) occupy the descending position on the curve.

(c) Metalloids (B, Si, Te, At etc) and transition metals occupy bottom part of the curve.

(d) Most electronegative i.e. halogens (F, Cl, Br, I) occupy the ascending position on the curve.

Note : Elements having similar properties occupy similar position on the curve.

Conclusion :

On the basis of the curve **Lother Meyer** proposed that the physical properties of the elements are periodic function of their atomic wt. and this become the base of **Mendeleev's periodic table**.

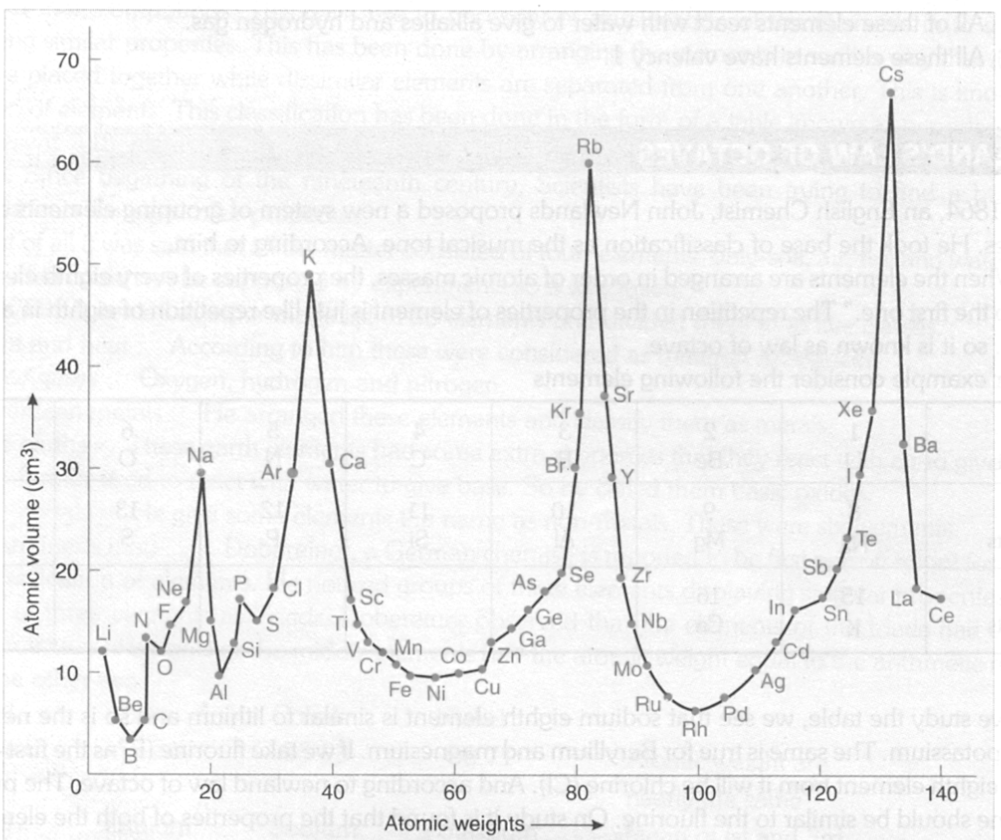


Illustration 2

What is periodic function.

Solution

When the elements are arranged in the increasing order of their atomic weight., elements having similar properties gets repeated after a regular interval.

5.6 MENDELEEV'S PERIODIC TABLE (1869)

Mendeleev's Periodic Table was published in a German journal in 1872. In the formula for oxides and hydrides at the top of the columns, the letter R is used to represent any of the elements in the group. Note the way formulae are written. For example, the hydride of carbon, CH_4 is written as RH_4 and the oxide CO_2 as RO_2.

Group	**I**		**II**		**III**		**IV**		**V**		**VI**		**VII**		**VIII**
Oxide Hydride	R_2O RH		RO RH_2		R_2O_3 RH_3		RO_2 RH_4		R_2O_5 RH_3		RO_3 RH_2		R_2O_7 RH		RO_4
Periods ↓	A	B	A	B	A	B	A	B	A	B	A	B	A	B	Transition Series
1	H 1.008														
2	Li 6.939		Be 9.012		B 10.81		C 12.011		N 14.007		O 15.999		F 18.998		
3	Na 22.99		Mg 24.31		Al 29.98		Si 28.09		P 30.974		S 32.06		Cl 35.453		
4 First series	K 39.102		Ca 40.08			Sc 44.96		Ti 47.90		V 50.94		Cr 50.20		Mn 54.94	Fe Co Ni 55.85 58.93 58.71
Second series		Cu 63.54		Zn 65.37	Ga 69.72		Ge 72.59		As 74.92		Se 78.96		Br 79.909		
5 First series	Rb 85.47		Sr 87.62			Y 88.91		Zr 91.22		Nb 92.91		Mo 95.94		Tc 99	Ru Co Ni 101.07 102.91 106.4
Second series		Ag 107.87		Cd 112.40	In 114.82		Sn 118.69		Sb 121.75		Te 127.60		I 126.90		
6 First series	Cs 132.90		Ba 137.34			La 138.91		Hf 178.49		Ta 180.95		W 183.85			Os Ir Pt 190.2 192.2 195.09
Second series		Au 196.97		Hg 200.59	Tl 204.37		Pb 207.19		Bi 208.98						

(i) **Mendeleev's Periodic Law** : The physical and chemical properties of elements are the periodic function of their atomic weight.

(ii) Characteristic of Mendeleev's periodic table.

(a) It was based on atomic weight.

(b) 63 elements were known at that time.

(c) Noble gases were not discovered.

(d) He was the first scientist to classify the elements in a systematic manner i.e. in horizontal rows and in vertical columns.

(e) Horizontal rows are called periods and there were 7 periods in Mendeleev's periodic table.

(f) Vertical columns are called groups and there were 8 groups in Mendeleev's Periodic table.

(g) Each group upto VII is divided into A & B subgroups, 'A' sub groups elements are called normal elements and 'B' subgroups elements are called transition elements.

(h) The VIII group contains 9 elements in three rows (Transitional metals group)

(i) The elements belonging to same group exhibit similar properties.

5.6.1 Achievements of Mendeleev's Periodic Table

(i) While developing the Periodic Table, there were a few instances where Mendeleev had to place an element with a slightly greater atomic mass before an element with a slightly lower atomic mass. The sequence was inverted so that elements with similar properties could be grouped together. For example, cobalt (atomic mass 58.9) appeared before nickel (atomic mass 58.7).

(ii) Further, Mendeleev left some gaps in his Periodic Table. Instead of looking upon these gaps as defects, Mendeleev predicted the existence of some elements that had not been discovered at that time.

Mendeleev named them by prefixing a Sanskrit numeral, Eka (one) to the name of preceding element in the same group.

For instance, scandium, gallium and germanium, discovered later, have properties similar to B, Al, Si that's why they were termed as Eka-boron, Eka-aluminium and Eka-silicon respectively.

The properties of Eka-aluminium predicted by Mendeleev and those of the element, gallium which was discovered later and replaced Eka-aluminium are listed as follows

Table: Properties of Eka-aluminium and gallium

Property	Preducted properties of Eka-alumnium	Real properties of Gallium
Atomic mass	68	69.7
Formula of Oxide	E_2O_3	Ga_2O_3
Formula of Chloride	ECl_3	$GaCl_3$

(iii) This provided convincing evidence for both the correctness and usefulness of Mendeleev's Periodic Table. Further, it was the extraordinary success of Mendeleev's prediction that led chemists not only to accept his Periodic Table but also recognise him, as the originator of the concept on which it based.

(iv) Noble gases like helium (He), neon (Ne) and argon (Ar) have been discovered later because they are very inert and present in extremely low concentrations in our atmosphere. One of the strengths of Mendeleev's Periodic Table was that, when these gases were discovered, they could be placed in a new group without disturbing the existing order.

5.6.2 Limitations of Mendeleev's Classification

Position of Hydrogen

Electronic configuration of **hydrogen** resembles that of **alkali metals**. Like alkali metals, hydrogen combines with halogens, oxygen and sulphur to form compounds having similar formulae.

On the other hand, just like halogens, hydrogen also exist as diatomic molecules and it combines with metals and non-metals to form covalent compounds.

Illustration 3

Many scientists before Mendeleev also used atomic mass as the basis of classification, but why did only Mendeleev succeed.

Solution

The secret of Mendeleev's success was that although the classification was based on atomic mass, but at many places he did not follow this rule rigidly. He laid major stress on the similarity in the chemical and physical properties.

Illustration 4

Why did Mendeleev leave many gaps in his periodic table?

Solution

Mendeleev predicted that there were many elements get to be discovered. So, when none of the elements known at that time fit into a particular position, he left a gap there. Later, when more elements were discovered these were found to fit into these gaps.

5.7 MAKING ORDER OUT OF CHOAS-THE MODERN PERIODIC TABLE

In 1913, **Henry Moseley** showed that the atomic number of an element is a more fundamental property that its atomic mass as described below. Accordingly, Mendeleev's Periodic Law was modified and ... can

... this

number increases by one in going from one element to the next. Elements, when arranged in order of increasing atomic number Z, lead us to the classification known as the **modern periodic table**. Prediction of properties of elements could be made with more precision when elements were arranged on the basis of increasing atomic number.

*5.7.1 Modern Periodic Table

Long form of periodic table :-

Block : The periodic table is divided into four main blocks (s, p, d and f) depending upon the subshell to which the valence electron enters into.

(i) Elements of group 1 and 2 constitute s-block.

(ii) Elements of group 13,14,15,16,17,18 constitute p-block.

(iii) Elements of groups 3,4,5,6,7,8,9,10,11,12 constitute d-block.

(iv) The f-block element comprise two horizontal rows placed at the bottom of the periodic table to avoid its unnecessary expansion.

Elements of s and p-block are called normal or representative elements, those of d-block are called transition elements while the f-block elements are called inner transition elements.

Groups:- The 18 vertical columns are called groups. The elements belonging to a particular group is known as a family and is usually named after the first number. A part from this some of the groups are given typical name as examplified beneath.

(i) Elements of group 1 are called **ALKALI METALS .**

(ii) Elements of group 2 are called **ALKALINE EARTH METALS.**

(iii) Elements of group 15 are called **PNICOGENS.** Means choking to system due to toxicity.

(iv) Elements of group 16 are called **CHALCOGENS.**

(v) Elements of group 17 are called **HALOGENS.**

(vi) Elements of group 18 are called **NOBLE GASES OR AEROGENS.**

All the other groups are named after the first member of each group.

Periods : The horizontal rows are called periods. There are Seven Periods in the long form of the periodic Table.

(i) 1^{st} Period ($_1H \rightarrow {}_2He$) contains 2 elements. It is the **shortest period**.

(ii) 2^{nd} Period ($_3Li \rightarrow {}_{10}Ne$) and 3rd Period ($_{11}Na \rightarrow {}_{18}Ar$) contains 8 elements each. These are **short period.**

(iii) 4^{th} Period ($_{19}K \rightarrow {}_{36}Kr$) and 5th period ($_{37}Rb \rightarrow {}_{54}Xe$) contains 18 elements each. These are **long periods.**

(iv) 6^{th} Period ($_{55}Cs \rightarrow {}_{86}Rn$) consists of 32 elements and is the **longest period.** It constitutes Lanthenide series.

(v) 7^{th} period starting with $_{87}Fr$ is incomplete period and consists of 19 elements.

5.7.2 What is Periodicity

The recurrence of properties of the elements, after a certain regular intervals, when they are arranged in the increasing order of their atomic numbers, is called **periodicity.**

5.7.3 Cause of Periodicity

We know that properties of elements depend upon the number of valence electrons, i.e. the number of electrons in the outermost shell. When the elements are arranged in the increasing order of their atomic numbers, then the elements having the same numbers of valence electrons are repeated at regular intervals of 2, 8, 8, 18, 18, 32....... Since the chemical properties depend upon the number of valence electrons, therefore, chemical properties are repeated at regular intervals, i.e. there is periodicity in the chemical properties of the elements.

5.7.4 Advantages of Long Form over Mendeleev's Peridic Table

There are several advantages of long form of periodic table over Mendeleev's periodic table. Some of these are as follows:

(i) It is based upon atomic number which is a fundamental property instead of atomic mass.

(ii) The elements have been grouped as s,p,d and f-block elements. which helps us to understand the electronic configuration in a better way.

(iii) In the long form of periodic table, the elements are arranged in the increasing order of their atomic numbers, therefore, no separate place is required for isotopes.

(iv) The position of some of the elements which were a misfit on the basis of atomic mass is now explained on the basis of increase in atomic number. For example, argon proceeds potassium because argon has atomic number 18 which is less than that of potassium which is 19.

(v) Metals, non-metals, metalloids, transition elements, lanthanoids and actinoids and actinoids are now better classified.

5.8 DIVISION OF THE PERIODIC TABLE IN *s, p, d* and *f* blocks :

1. **s-Blocks Elements :**

The elements in which the last electron enters the s-sub-shell of their outermost energy level and electronic configuration is ns^1 or ns^2 (I or II group) are called **s-block elements are :**

(i) They are soft metals.

(ii) They have low ionisation energies.

(ii) They are very reactive and form ionic compounds.

(iv) They show oxidation states of +1 group and +2 group.

(v) They are good reducing agents.

2. *p*- Block Elements :

The elements in which the last electron enters the p-sub-shell of their outermost energy level are called **p-block elements.** The exception is helium ($1s^2$).

The general configuration of their outermost shell is $\mathbf{ns^2\ np^{1-6}}$. These elements are kept in group 13 to 18. Some of the general characteristics of p-block elements are:

(i) They show variable oxidation states.

(ii) They form ionic as well as covalent compounds.

(iii) Most of them are non-metals.

(iv) Most of them form acidic oxides.

3. *d*- Block Elements :

(i) They are hard and having high melting point.

(ii) They show variable oxidation states.

(iii) They form coloured compexes.

(iv) They form ionic as well as covalent compounds.

(v) Most of them exhibit paramagnetism

(vi) Most of them possess catalytic properties.

4. *f*- Block Elements :

The elements in which the last electron enters the ***f*- block elements.**

Their general configuration is **(n-2) f^{1-14} (n-1)d^{0-1}, ns^2.** They consist of two series of 28 elements (14 in each) placed at the bottom of the periodic table.

The elements of first series followed by lanthenum ($_{57}La$) are called **Lanthenides**.

The elements of second series followed by actinium ($_{89}Ac$) are called **actinides**.

The general **Characterstics of *f*-block elements are :**

(i) They show variable oxidation states.

(ii) They have high melting points.

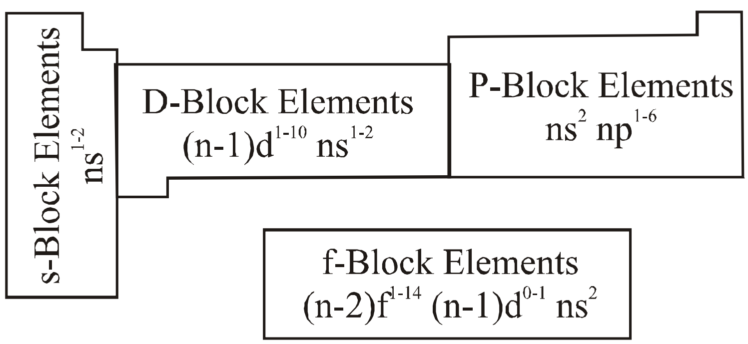
(iii) They have high densities.

(iv) They form coloured compounds.

(v) Most of the elements of actinide series are radioactive.

It may be noted that :

1. The elements of group zero are called ***inert gases, noble gases, rare gases*** **or** ***aerogens.***
2. The elements of p-block (except noble gases) are called *representative or main group elements*. The members of this group of elements have all their occupied subshells filled except their outermost electron shell.



5.9 PERIODIC PROPERTIES

The properties which are directly or indirectly related to their electronic configuration and show gradual change when we move from left to right in a period or from top to bottom in a group are called **periodic properties.**

(a) Atomic volume, atomic size, melting point, boiling point and density are important physical properties which show periodicity.

(b) Some important chemical properties that exhibit periodicity are electronic configuration, ionisation energy, electron affinity, electronegativity, metallic character, nature of oxides, oxidation state and reducting character etc.

(c) Specific heat, refractive index etc. are not periodic properties.

(A) Atomic size

It refers to the distance between the centre of nucleus of atom to its outer most shell of electrons.

Since absolute value of atomic size cannot be determined, it is expressed in terms of the operational definitions such as **ionic radius, covalent radius, van der waal's radius and metallic radius.**

The absolute value of atomic radius cannot be determined because :

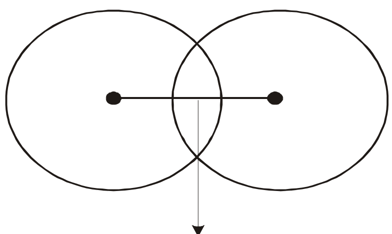
(i) It is not possible to locate the exact position of electrons in an atom as an orbital has no sharp boundaries.

(ii) It is not possible to isolate an individual atom for its size determination.

(iii) In a group of atoms, the probability distribution of electrons is influenced by the presence of neighbouring atoms. Thus, size of an atom may change in going from one environment to other.

However, the value of atomic radii are derived from bond length measured by various techniques such as **X-ray diffraction, electron diffraction, infra-red spectroscopy, nuclear magnetic resonance spectroscopy etc.**

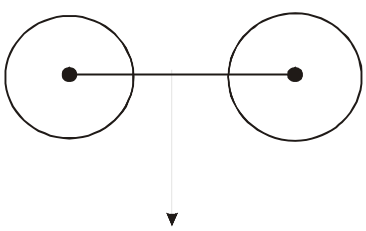
(i) Covalent radius : It is defined as half of the distance between two successive nuclei of two covalent bonded like atoms in a molecule. If the bond length in between the two atoms is say A-A = d, then



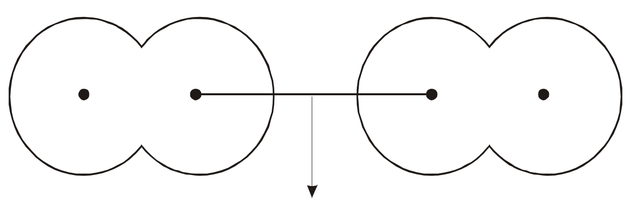
Internuclear distance between
two covalently bonded atoms

Covalent radius ($r_{cov.}$) = ½. [Internuclear distance between two covalently bonded like atoms] = ½d

(ii) Vander Wall's radius ($r_{v,waat}$) . It is defined as one half of the distance between the nuclei of two non bonded isolated atoms or two adjacent atoms belonging to two neighbouting molecules of an element in the solid state.



Distance between two
isolated atoms



Distance between adjacent atoms belonging
to two neighbouting molecules in solid state.

In general, vander waals radius > Covalent radius of an atom.

(iii) Metallic radius : It is defined as half the distance between two successive nuclei of two adjacent metal atoms in the metallic closed packed crystal lattice.

Factors influencing atomic radius :

(a) Multiplicity of bond . Covalent radii depends on the multiplicity of bonds. e.g.

	Bond length	Radius of carbon atom
$H_3C—CH_3$	154 pm	$\frac{154}{2} = 77$ pm,
$H_2C = CH_2$	134 pm	$\frac{134}{2} = 67$ pm
$HC \equiv CH$	120 pm	$\frac{120}{2} = 60$ pm

(b) Percentage of ionic character : Covalent radius of H in HCl, HBr, and HI are different.

(iv) IONIC RADIUS :

Ionic radius may be defined as the effective distance from the nucleus upto which it has an influence in the ionic bond.

When an atom changes into ion, its size changes appreciably.

(a) **The radius of cation is smaller than that of the parent atom.**

(b) **The radius of anion is larger than that of parent atom.**

(c) ***The ions having same number of electrons but different magnitude of nuclear charge are called iso-electronic ions.***

For example, each one of sulphids (S^{2-}), chloride (Cl^-), Potassium (K^+), and Calcium (Ca^{2+}) ion has eighteen electrons but they have nuclear charge, +16, +17, +19 and +20 respectively.

Within the series of iso-electronic ions, size of the ions decreases with the increase in the magnitude of nuclear charge. For example, N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+}, Al^{3+}, are iso-electronic and have 10 electrons each. The size of these ions are in the order.

$Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$

The ionic radii of these ions are given in Table below :

Ionic Radii of Some Iso-electronic Ions

Atoms	N^{3-}	O^{2-}	F^-	Na^+	Mg^{2+}	Al^{3+}
Nuclear Charge	+7	+8	+9	+11	+12	+13
Size (pm)	1.71	1.40	1.26	0.95	0.65	0.50

Trends in Atomic size

Variation of atomic size in a period. On moving from left to right in a period of the periodic table, the size of atoms goes on decreasing.

Variation of atomic radii of Element of Third Period

Element	Na	Mg	Al	Si	P	S	Cl
Atomic Radii (pm = 10^{-12} m)	186	160	143	117	110	104	99

Variation of atomic Size in a Group. On moving down a group of the periodic table, the size of the atom goes on increasing. For Example. In case of group I elements, the size increases from 152 pm in case in lithium to 262 pm in case of cesium (Table)

Group 1 Elements	**Atomic Radii (pm)**
Lithium (Li)	152
Sodium (Na)	186
Potassium (K)	231
Rubidium (Rb)	244
Cesium (Cs)	262
Francium (Fr)	-

Illustration 8

Give examples of three cations and three anions which are isoelectronic with Argon.

Solution **Cations : K^+ Ca^{2+} Sc^{3+}**

Anions : Cl^- S^{2-} P^{3-}

(B) IONISATION ENERGY

The amount of energy required to remove the most loosely bound electron from an isolated gaseous atom is called ionisation energy.

The process may be represented as :

$A(g) + I.E. \rightarrow A^+ + e^-$

Ionisation energy is expressed either in terms of electron volts per atom (eV/atom) or kilo jules per mole of atoms (kj/mol).

1 ev per atom = 96.3 kj mol^{-1}

The ionisation energy defined above is more precisely the first ionisation energy (IE_1).

Successive Ionisation energies :

Once the first electron has been removed from the gaseous atom, it is possible to remove second electron from the monopositive ion.

The amount of energy required to remove the most loosely bound electron from the isolated monopositive ion of the element is called second ionisation energy (IE_2). Similarly the energy required to remove the outermost electron from isolated dipositive ion is called third ionisation energy ($I.E._3$).

$A^+ (g) + I.E._2 \rightarrow A^{2+}(g) + e^-$

$A^{2+} (g) + I.E._3 \rightarrow A^{3+}(g) + e^-$

The second (I.E.2) third (I.E.3), fourth (I.E.4) etc. ionisation energies are collectively known as successive ionisation energies. It may be noted that :

$\mathbf{IE_3 > IE_2 > IE_1}$

Factor on which Ionisation Energy Depends :

1. **Atomic Size :** I.E. decreases with increases in atomic size.
2. **Nuclear Charge :** I.E. increases with increase in nuclear charge.
3. **Screening Effect of the Inner Electrons. :** Larger the number of electrons in the inner shells, greater is the screening effect on the outermost electron and hence lower is the ionisation energy.

4. **Penetration Effect :** Penetration power of various sub-shells of a particular energy level is in the order s> p> d> f. Therefore, for the same shell it is easier to remove an electron from p-sub-shell than from s-sub-shell. Greater the penetration power, higher is the I.E.

5. **Electronic Configuration. :** The elements having stable electronic configurations have relatively higher values of I.E. for example,

(i) The noble gases have stable configuration (ns^2np^6). They have highest ionisation energies within their respectively periods.

(ii) The elements like N ($1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$) and P ($1s^2, 2s^2, 2p^6, 3s^2, 3p_x^1, 3p_y^1, 3p_z^1$) have configurations in which orbitals belonging to same sub-shell are exactly half-filled. Such configurations are quite stable and consequently, their ionisation energies are relatively high.

(iii) The elements like Be ($1s^2, 2s^2$) and Mg ($1s^2, 2s^2, 2p^6, 3s^2$) have all electrons paired. Such configurations being stable also result in the higher values of ionisation energy.

Variation across the period : In general, the value of ionisation energy increases with the increases in atomic number across a period. This can be attributed to the fact that in moving across the period, nuclear charge increases and atomic size decreases regularly and hence, the electrons are more tightly bound to the nucleus.

Variation in a group : The value of ionisation energies decreases regularly with the increases in atomic number of shell and distance from the nucleus along a group.

Illustration 9

Out of Na^+ and Ne which has higher ionisation energy? Explain why.

Solution

Na^+ has higher ionisation energy than Ne. Na^+ and Ne are isoelectronic species. However, the nuclear charge in Na^+ is more than in Ne. Hence, the electons are more tightly held in Na^+ and it has higher ionisation energy.

Illustration 10

Out of Al^+ and Mg^+ which has higher I.E. ?

OR

Out of Al and Mg which has higher second I.E. ?

Solution

In both Al^+ and Mg^+ the outermost electron is removed from 3s-orbital.

Al^+ : $1s^2, 2s^2, 2p^6, 3s^2$

Mg^+ : $1s^2, 2s^2, 2p^6, 2s^1$

Al^+ has higher I.E. than Mg^+ because nuclear charge in Al^+ (13 units) in higher than in Mg^+ (12 units) Moreever. Al^+ has stable configuration (Fully filled 2s sub-shell).

(C) ELECTRON GAIN ENTHALPY

Electron Gain Enthalpy may be defined as the amount of energy released when an electron is added to an isolated gaseous atom of the element.

The process may be expressed as :

$A(g) + e^- \rightarrow A^-(g) + E.A.$

The large value of electron reflects the greater tendency of an atom to accept the electron.

Units : The values of electron affinity are expressed either in electron volt per atom or kilo Jules per mole of atoms.

Factors Affecting Electron gain Enthalpy :

1. **Nuclear charge :** Greater the magnitude of nuclear charge, larger will be the value of electron affinity.
2. **Atomic size :** Larger the size of an atom, smaller will be the value of electron affinity.
3. **Electronic configuration. :** Stable the electronic configuration of an atom, lesser will be its tendency to accept the electron and lower will be the value of its electron affinity.

Variation in a period : On moving across the period the atomic size decreases and nuclear charge increases. Both these factors result into greater attraction for the incoming electron, therefore, **electron affinity, in general, increases in a period from left to right.** However, some irregularities are observed in a general trends. These are mainly due to the stable electronic configurations of certain elements. For example, electron affinities of noble gases, nitrogen, beryllium and magnesium are negative (regarded as zero) due to their stable electronic configurations.

Variation along a group : On moving down a group, the atomic size and nuclear charge increase but the effect of increase in atomic size is much more pronounced that of nuclear charge and thus the additional electron feels less attraction by the nucleus.

Consequently, electron affinity decreases from top to bottom in a group. However, electron affinities of elements of the second period are lower than those for the elements of the third period. This is possibly due to the strong inter electron repulsion forces operating within the relatively compact 2p-sub-shell.

Illustration 11

Arrange F, Cl, Br and I in the decreasing order of their electron gain enthalpies and account for the order assigned.

Solution

The decreasing order of electon affinities is Cl > F > Br > I.

In case of fluorine electron goes to second energy level which is very small in size and has already seven electrons present in it. Therefore, the added electron feels some repulsion from the electrons already present in the valence shell. This neutralises to some extent the attractive force of nucleus on the electron being added.

On the other hand in chlorine added electron goes to third energy level which is of larger size. Hence electron affinity of fluorine is less than that of chlorine.

$F + e^- \rightarrow F^-$

$1s^2, 2s^2, 2p^5$

$Cl + e^- \rightarrow Cl^-$

On the other hand, if we compare the electron affinities of Cl and Br, then electron affinity of Cl is more than that of Br because Cl atom is smaller in size and hence effective nuclear charge in Cl is more than in Br. Therefore, added electron is more strongly attracted in Cl. Similarly, electron affinity of Br is higher than that of I. So, the overall of electron affinities of halogens is Cl > Br > I.

Illustration 12

In each of the following sets arrange the elements in the increasing order of their electron affinities:

(i) C, N, O ***(ii) O, N, S***

(iii) S, Cl, Ar ***(iv) F, Cl, Br***

Solution

(i) N < C < O (ii) N < O < S (iii) Ar < S < Cl

(iv) Br < F < Cl

(D) ELECTRONEGATIVITY

Electronegativity may be defined as the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.

the main factors on which the electronegativity depends are effective nuclear charge and atomic radius.

Electronegativity $\propto$ Effective Nuclear Charge

Electronegativity $\propto \dfrac{1}{\text{size}}$

In period : electronegativity increases in moving from left to right. This is due to the reason that nuclear charge increases whereas atomic radius decreases as we move from left to right in a period. Halogens have the highest value of electronegativity in their respective periods.

In a group : electronegativity decreases on moving down the group. This is due to the effect of increased atomic radius.

(E) VALENCY

Valency may be defined as combining capacity of an element.

The Valency of an element is usually determined by the number of electrons in the outermost shell of its atoms. This is because the outer electrons are largely responsible for the chemical behaviour as these electron usually participate in chemical bonding.

(i) Variation of valency in a period. On moving from left to right in each short period, the valency of the elements first increases from 1 to 4 and then decreases and becomes 0 (zero). The trend for the elements of third period is shown below in table.

Variation of Valency of Element of Third Period

Element	Na	Mg	Al	Si	P	S	Cl	Ar
Valency	1	2	3	4	3	2	1	0

(ii) Variation of Valency in a Group. On moving down a group, the number of valence electrons remains the same and due to this all the elements in a particular group have the same valency. For example,

All elements of group 1 shows a **valency** of 1

All elements of group 2 shows a **valency** of 2

All elements of group 13 shows a **valency** of 3

All elements of group 14 shows a **valency** of 4

All elements of group 15 shows a **valency** of 3

All elements of group 16 shows a **valency** of 2

All elements of group 17 shows a **valency** of 1

All elements of group 18 shows a **valency** of 0

Illustration 13

(i) How do you think the tendecy to lose electrons will change in a group?

(ii) How will this tendency change in a period

Solution

As the effective nuclear charge acting on the valence shell electrons increases across a period, the tendency to lose electrons will decrease. Down the group, the effective nuclear charge experienced by valence electrons is decreasing because the outermost electrons are farther away from the nucleus. Therefore, these can be lost easily. Hence metallic character decreases across a period and increases down a group.

Non-metals on the other hand are electronegative. They tend to form bonds by gaining electrons. Let us learn about the variation of this property.

Illustration 14

(i) How would the tendency to gain electrons change as you go from left to right across a period?

(ii) How would the tendency to gain electrons change as you go down a group?

Solution

As the trends in the electronegativity show, non-metals are found on the right hand side of the Periodic Table towards the top.

These trends also help us to predict the nature of oxides formed by the elements because it is known to you that the oxides of metals are basic and that of non-metals are acidic in general.

Illustration 5

The following species are isoelectronic with the noble gas neon, Arrange them in order on increasing size : Na^+, F^-, O^{2-}, Mg^{2+}, Al^{3+}.

Solution

In Na^+, F^-, O^{2-}, Mg^{2+}, Al^{3+}, the nuclear charges are 11, 9, 8, 12 and 13 respectively. Among isoelectronic species, greater the nuclear charge smaller is the size. Therefore, the sizes of the above ionic species are in the order :

$Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-}$.

Illustration 6

Out of Na^+ and Na which has smaller size and why?

Solution

Na^+ has smaller size than Na, Na^+ has one electron less than Na. However, Na and Na^+ have same nuclear charge. Therefore, electrons in Na^+ are more tighty held than in Na. So, removal of one electron from Na leads to complete removal of the third shell so that in Na^+, the outermost shell is second. Hence, Na^+ has smaller size than Na.

Metallic and Non-metallic Character. Generally, **Metals** posses 1, 2 or 3 electrons in their respective valence shells and thus have a strong tendency to lose these electrons to form positive ions. therefore, metals are also called **electropositive elements** and the metallic character is also called **Electropositive Character.**

Non-Metals, on the other hand, generally have 4 to 8 electrons in their respective outermost shells and thus have a tendency to gain electrons to form negative ions. Therefore non-metals are also called **Electronegative elements** and the non metallic character is also called **electronegative character.**

Variation in a period: On moving from left to right in a period, the metallic character decreases while the non-metallic character increases.

Second period elements $\underbrace{\text{Li} \quad \text{Be}}_{\text{Metals}} \quad \underbrace{\text{B} \quad \text{C} \quad \text{N} \quad \text{O} \quad \text{F}}_{\text{Non - Metals}}$

Nature of elements $\xrightarrow[\textit{Non – Metallic character increases}]{\textit{Metallic character decreases}}$

Thus in the second period, lithium is the most metallic element followed by beryllium. The non metallic character starts with boron and keeps on increassing. Thus fluorine is the most non metallic element of the second period.

The change from metallic to non metallic character is more striking among the elements of the third period as shown below :

Third period elements $\underbrace{\text{Na} \quad \text{Mg} \quad \text{Al}}_{\text{Metals}} \quad \underbrace{\text{Si}}_{\text{Metalloid}} \quad \underbrace{\text{P} \quad \text{S} \quad \text{Cl}}_{\text{Non - Metals}}$

Nature of elements $\xrightarrow[\textit{Non – Metallic character increases}]{\textit{Metallic character decreases}}$

In general, the most metallic element lies on the extreme left hand side while most non metallic element lies on the extreme right hand side of any period.

Variation in a group. On moving down in a group, the metallic character or electropositive character increases.

Elements of Group 1

Element	Symbol	Metallic character
Lithium	Li	Least metallic element
Sodium	Na	
Potassium	K	
Rubidium	Rb	
Cesium	Cs	
Francium	Fr	Most metallic element

Conversely, on moving down a group, the non metallic character or electronegative character decreases.

Elements of Group 17

Element	Symbol	Metallic character
Fluorine	F	Most non metallic element
Chlorine	Cl	Non-Metallic character or electronegative character decreases on moving down the group
Bromine	Br	
Iodine	I	
Astatine	At	Least non-metallic element

Division of elements into Metals and Non-metals: In the long form of the periodic table, the elements have been broadly divided into metals and non metals by the tick zig zag line funning diagonally across the periodic table. those elements which lie on the left hand side of this line are Metals and those which lie on the right hand side of this line are non **metalic**. However, the elements, silicon, germanium, arsenic antimony metallic and tellurium which lie along the border of this line show the properties of both metals and non metals. These elements are called semi metals or **metalloids**.

The metallic/non mettalic character also helps us to predict the nature of oxides formed by the elements. in general, the oxides of metals are basic while those of non metals are acidic in nature.

An element behaves as a metal or a non-metal is directly related to its ionization energy. The elements having low values of ionization energies are metals whereas elements having high values of ionization energies are non-metals.

Illustration 7

Elements having atomic number 3 to 18 are shown in the form of a table by using certain letters of the alphabet (These letters are not the usual symbols of these compound)

3 A	4	5	6	7	8	9	10 G
11 B	12 C	13	14 D	15	16	17 F	18

(a) Which of these is

(i) a noble gas ***(ii) a halogen***

(iii) an alkali metal ***(iv) an element with valency 4?***

(b) Write the formula of the compound formed when A react with F?

(c) Write the electronic configuration of element G?

Solution

(a) (i) G (ii) F (iii) A and B (iv) D

(b) $A \rightarrow A^+ + e^-$; $F + e^- \rightarrow F^-$; $A^+ + F^- \rightarrow A^+F^- \rightarrow AF$

So the formula of the compound formed is AF

(c) G has the atomic number 10. So the electronic arrangement in G is 2,8.

*5.9 SOME IMPORTANT NOTES

1. Based on number of incomplete shells the elements are classified into four types. Inert gases, representative elements, transition elements and inner transistion elements.

(a) General electronic configuration of inert gases (except He) is ns^2np^6 (outermost shell completely filled.)

(b) In representative elements, the outermost shell only in incomplete. General electronic configurations varies from ns^1 to ns^2 np^6.

(c) In transition elements, two outer most shells are incomplete (nth and (n-1)th]. General electronic configuration is $(n-1)\ d^{1-10}\ ns^{1-2}$.

(d) In inner transition elements, three outer shells are incomplete. General electronic configuration is $(n-2)f^{1-14}\ (n-1)d^{0-1}\ ns^2$.

Diagonal Relationship :- Certain elements of second period exhibit similarty in properties as shown by the elements diagonally placed to them in the third row, e.g., This is called **diagonal relationship.** Li and Mg; Be and Al; B and Si as shown below.

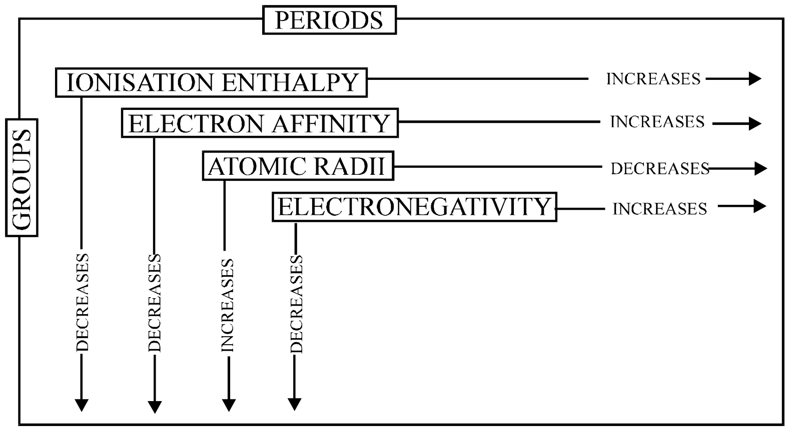
	Group 1	Group 2	Group 3	Group 4
2nd Period	Li	Be	B	C
3rd Period	Na	Mg	Al	Si

This is due to identical ionic radii and polarizing power (i.e., charge/ size ratio of the pairs of these elements). Elements of second period are known as **bridge elements.**

Anomalous behaviour of first element of a group.

The first element of a group exhibits difference in its properties in certain respects from the rest of the elements of its group. This is due to its small size, high electronegativity and non avaiability of d-electrons. This anomalous behaviour is shown by the elements of the second raw (period) i.e. , Li to F.

Some Important General Periodic Trends



Solved Examples

Example 1

Explain why there are only 18 elements in the fifth period.

Solution

The fifth period begins with the filling of 5s orbital and continious fill the filling of sixth energy level (6s) starts. The sub-shells which follow 5s are, 4d, 5p, 6s .... Thus, the elements which involves filling of 5s, 4d and 5p, subshells can accept 18 electronis in all, there are 18 elements in the fifth period.

Example 2

Out of Al and Al^{3+} which has smaller size and why?

Solution

Al^{3+} is formed by removal of 3 electrons from Al. Thus number of electrons in Al^{3+}, is three less than the number of electrons in Al whereas both have same nuclear charge. Therefore, electrons are more tightly held in Al^{3+}, the outer-most shell is second. Hence, Al^{3+} is smaller in size than Al.

$${}_{13}Al : 1s^2, 2s^2, 2p^6, 3s^2, 3p_x^1.$$

Example 3

Arrange Mg, Mg^+ and Mg^{2+} in the increasing order of their sizes. Give explanation for the order assigned.

Solution

Mg, Mg^+ and Mg^{2+} have same nuclear charge (12 units) however, the number of electrons in them is 12, 11 and 10 respectively. Therefore, Mg^{2+}, having least number of electrons, is the smallest whereas Mg, having maximum number of electrons, is the largest. The increasing order of their size is :

$$Mg^{2+} < Mg^+ < Mg$$

Example 4

How do you explain that $_{31}Ga$ has slightly higher I.E. than $_{13}Al$, although it occupies lower position in the group?

Solution

$_{13}Al$: $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$

$_{13}Ga$: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^1$

In Ga, the 10 electrons present in 3d-sub-shell do not shield the outer electrons from the nucleus effectively. As a result effective nuclear charge in Ga increases. This explains why I.E. of Ga is slightly more than that of $_{13}Al$.

Example 5

Out of Ca^{2+}, and Ar which has higher ionisation energy ? Explain briefly.

Solution

Ca^{2+} and Ar are isoelectronic and have 18 electrons each. However, nuclear charge of Ca^{2+}, is 20 whereas that of Ar is 18. Due to higher nuclear charge. electrons in Ca^{2+} are more tightly held and hence it has higher ionisation energy than Ar.

Example 6

Out of aluminium and magnesium which has higher ionisation energy and why?

Solution

Aluminium and magesium belong to same period. Nuclear charge in Al is 13 and in Mg it is 12. So, it is expected that Al should have higher I.E. than Mg. But if we carefully observe the electronic configurations of Al and Mg then we find that Mg has stable configuration (Fully filled s-Sub-shell in valence shell) and hence has higher ionization energy than Al. Moreover in Mg, the electron is removed from 3s-orbital which, due to high penetration and less shielding by the electrons in the lower obritals, experiences stronger force attraction from nucleus as compared with 3p electron of Al. Therefore Mg has higher ionisation energy than Al.

$_{13}Al$: $1s^2, 2s^2, 2p^6, 3s^2, 3p_x^1$

$_{12}Mg$: $1s^2, 2s^2, 2p^6, 3s^2$

Example 7

Out of Al^+ and Mg^+ which, has higher I.E.?

or

Out of Al and Mg which has higher second I.E.?

Solution

Al^+ has higher second I.E. because nuclear charge in Al^+ (13 units) is higher than in Mg^+ (12 units) and Al^+ has stable configuration (Fully filled 3s sub-shell).

Example 8

Arrange F, Cl, Br and I in the decreasing order of their electron affinities and account for the order assigned.

Solution

The decreasing order of electron affinities is Cl > F > Br > I.

In case of fluorine added electron goes to second energy electron feels some repulsion from the electrons already present in the valence shell. On the other hand, in chlorine added electron goes to third level which is of larger size. Hence, **electron affinity of fluorine is less than that of chlorine.**

$$F + e^- \rightarrow F^-$$

$$Cl + e^- \rightarrow Cl^-$$

on the other hand, if we compare the electron affinities of Cl and Br then electron affinity of Cl is more than that of Br because Cl atom is smaller in size and hence effective nuclear charge in Cl is more than in Br. Therefore, added electron is more strongly attracted in Cl. Similarly, electron affinity of Br is higher than that of I. So, the overall order of electron affinities of halogens is Cl > F > Br > I.

Example 9

Explain why electron affinities of noble gases are zero.

Solution

All noble gases have ns^2np^6 valence shell configuration (except helium where it is $1s^2$). In the valence shell of noble gases s and p sub-shells are fully filled, therefore, these configurations are highly stable. Hence, noble gases do not have any tendency to accept more electrons because this would disturb their stable configuration. Hence, electron affinities of noble gases are zero.

Example 10

Given on the side is a part of the periodic table. As we move vertically downward from Li to fr

(a) What happens to the size of the atom from Li to Fr.

(b) What happens to their metallic character?

Solution

(a) When we move vertically downward from Li to Fr the atomic size increases, i.e. Fr is the biggest atom.

(c) The metallic character increases i.e Fr is the most metallic element (Tendency of large electrons)

Example 11

Given below is a part of the periodic table

Li	Be	B	*l*	N	O	F
Na	Mg	Al	Si	P	S	Cl

As we move horizontaly from left to right

(a) What happens to the metallic character of the elements?

(b) What happens to the atomic size?

Solution

(a) When we move horizontally from left to right, the metallic character decreases.

(b) When we move horizontally from left to right, the atomic size decreases.

Example 12

Element (X) forms a chloride with the formula XCl_2, Which is a solid with high melting point. (X) would most likely be in the same group of the periodic table as

(a) Na (b) Mg (c) Al (D) Si (E) P

Solution

From the information given, the compound XCl_2 is an acidic compound.

The element (X) should have last two electrons to form XCl_2. So (X) belongs to group 2. Therefore, the element (X) lies in the same group as magnesium (Mg). So the answer (b) is correct.

EXERCISE-I

1. What was the basis of Dobereiner's classification?
2. What are the limitation of the Newland's law of octaves?
3. An element (X) is in the third group of the periodic table. What is the formula of its oxide?
4. What valency will be shown by an element having atomic number 17?
5. State the periodic law on which Mendeleev's periodic table was based?
6. Hydrogen could be placed both in group I as well in groiup XVII. Give three points to support this statement.
7. In which part of the periodic table are (a) metals (b) non metals located.
8. How does the electronic distribution in atoms change in a period from left to right?
9. What are the horizontal rows and vertical columns in a periodic table known as?
10. State the modern periodic law?
11. By which common name are the elements of group I and group-17 called?
12. Name an element which as
    (a) most metallic (b) most non-metallic
13. How is the position of an element in the periodic table related to the number of valence electrons?
14. What is a period in the periodic table? How do atomic structures change in a period with an increase in atomic number?
15. State one reason for keeping chlorine and bromine in the same group of the periodic table.
16. Which of the following halognes show the least realtivity towards hydrogen?
    Chlorine, Bromine, Iodine
17. From the stand point of atomic structure, what determines which element will be the first and which one will be the last in a periodic table?
18. Li, Na, Ba, Cl, I, Sr, Ca, Br, K
    Arrange the following and make their triads having similar properties.
19. How many elements were known at the time of Mendeleev's classification?
20. Name three elements whose existance was predicted by Mendeleev?
21. Name two elements whose atomic weight was corrected on the basis of their position in Mendeleev's periodic table.
22. How modern periodic law was able to remove the drawbacks of Mendeleev's periodic table.
23. Which out of nitrogen and phosphorus has more size and why?
24. What happens to the metallic character, melting point and boiling points of the elements as we go down in a group of the periodic table.
25. An element has atomic number 16. Identify its group and period?
26. An element is in 4th group and second period. What will be its atomic number.
27. Is Fe, Co, Ni are dobereiner triad?
28. Atomic wt. of an element X is 39, and that of element Z is 132. What is the atomic weight of their intermediate element Y?
29. The law of triads is not applicable on :
    (A) Cl, Br, I (B) Na, K, Rb (C) S, Sc, Te (D) Ca, Sr, Ba

30. Which of the following set of elements obeyes Newland's octave rule-
    (A) Na, K, Rb (B) F, Cl, Br (C) Be, Mg, Ca (D) B, Al, Ga
31. For which of the pair Newland octave rule is not apllicable :-
    (A) Li, Na (B) C, Si (C) Mg, Ca (D) Cl, Br
32. Which property of elements was used by lother Meyer to classify elements?
33. Which of the following element has highest density?
    (A) Alkali Metal (B) Alkaline Earth Metal (C) Halogen (D) Transition Metal
34. An element (X) is in the third group of the periodic table. What is the formula of its oxide?
35. Use Mendeleev periodic table to predict the formulae for the oxides of the following elements
36. Arrange the following ions in the increasing order of their sizes.
    Cl^{-}, P^{3-}, S^{2-}, F^{-}
37. Name
    (a) Three elements that have only single electron in their outermost shell.
    (b) Two elements that have two electrons in their outmost shell.
    (c) Three elements with filled outermost shell.
    (d) One element that has a single electron in its outermost shell.
38. In the modern periodic table, of the first ten elements which are metals?
39. What is the name and the chemical symbol of the alkaline earth metal with the smallest atomic number?
40. Name two elements you would expect to show same kind of chemical reactivity as magnesium?
41. Name two elements you would expect to show same kind of chemical reactivity as magnesium?
42. Out of metallic radius and covalent radius of an element which is larger and why?
43. Why vander waal's radius of an element is always larger than the covalent radius?
44. Which element has the highest ionisation energy?
45. Among s-block elements which element has the highest ionisation energy?
46. Out of oxygen and sulphur, which has higher electron affinity and why?

EXERCISE-II

1. Did Dobereiner's triads also exist in the columns of Newland's Octaves? Compare and find out?
2. What were the limitations of Dobereiner's Classification?
3. What were the limitations of Newlands' Law of Octaves?
4. Use Mendeleev's periodic table to predict the formula for the oxides of the following elements.
   K , C, Al, Si, Ba,
5. Besides gallium, which other elements have since been discovered that were left by Mendeleev in his Periodic Table? (any two)
6. What were the criteria used by Mendeleev in creating his Periodic Table?
7. Why do you think to noble gases are placed in a separate group?
7. How could the Modern Periodic Table remove various anomalies of Mendeleev's Periodic Table?
8. Name two elements you would expect to show chemical reactions similar to magnesium. What is the basis for your choice?
9. Name :

(a) three elements that have a single electron in their outermost shells.

(b) two elements that have two electrons in their outermost shells.

(c) three elements with filled outermost shells.

10.

(a) Lithium, sodium, potassium are all metals that react with water to liberate hydrogen gas. Is there any similarity in the atoms of these elements?

(b) Helium is an unreactive gas and neon is a gas of extremely low reactivity. What, if anything, do their atoms have in common?

11. In the Modern Periodic Table, which are the metals among the first ten elements?

12. By considering their position in the periodic Table, which one of the following elements would you expect to have maximum metallic characteristic?

Ga, Ge, As, Se, Be

13. Which of the following statements is not a correct statement about the trends when going from left to right across the periods of periodic Table.

(A) The elements become less metallic in nature

(B) The number of valence electrons increases.

(C) The atoms lose their electrons more easily.

(D) The oxides become more acidic.

14. Element X from a chloride with the formula XCl_2 . Which is a solid with a high melting point. X would most likely be in the same group of the Periodic Table as :

(A) Na (B) Mg (C) Al (D) Si

15. Which element has :

(A) two shells, both of which are completely filled with electrons?

(B) the electronic configuration 2,8,2?

(C) a total of three shells, with four electrons in its valence shell?

(D) twice a many electrons in its second shell as in its first shell?

16. (A) What property do all elements in the same column of the Periodic Table as boron have in common?

(B) What property do all elements in the same column of the Periodic Table as flurine have in common?

17. An atom has electronic configuration 2,8,7.

(A) What is the atomic number of this element?

(B) To which of the following elements would it be chemically similar? (Atomic numbers are given in parentheses).

N (7) F (9) P (15) Ar (18)

18. The position of three elements A, B and C in the Periodic Table are shown below -

Group 16	Group 17
-	-
-	A
-	-
B	C

(A) State whether A is a metal or non-metal.

(B) State whether C is more reactive or less reactive than A.

(C) Will C be larger or smaller in size than B?

(D) Which type of ion, cation or anion, will be formed by element A?

19. Nitrogen (atomic number 7) and phosphorus (atomic number 15) belong to group 15 of the Period Table. Write the electronic configuration of these two elements. Which of these will be more electronegative? Why?
20. How does the electronic configuration of an atom relate to its position in the Modern periodic Table?
21. In the Modern periodic Table, Calcium (atomic number 20) is surrounded by elements with atomic number 12,19,21 and 38. Which of these have physical and chemical properties resembling with calcium?
22. Compare and contract the arrangement of elements in Mendeleev's Periodic Table and the Modern Periodic Table.

EXERCISE-III

SECTION-A

- **Fill in the blanks**

1. The d-block elements are known as _______________ metals.
2. In the long form of periodic table, there are _________ horizontal rows known as ________.
3. The size of the atom in the second period of periodic table goes on __________ with increase in atomic number.
4. The size of anion is __________ than its parent atom.
5. Ca^{2+} has a smaller ionic radius than K^{+} because it has ___________.
6. Amongst the species Na^{+}, Mg^{+2} and Al^{+3} the smallest size is of _________
7. The element with highest value of IE_1 is ___________.
8. The most electropositive element in first period is _________ Whereas the most electronegative element is _______________.
9. The 16 group elements are also called ______________
10. Electropositivity ____________ on going down a group.

SECTION-B

1. The number of periods and group in the long form of periodic table are respectively

   (A) 7 and 9 (B) 8 and 18 (C) 7 and 18 (D) 6 and 10

2. The elements which are characterised by the outer shell configurations ns^1 to ns^2 and ns^2p^1 to ns^2p^5 are collectively called

   (A) Transition elements (B) Representative elements
   (C) Lanthanides (D) Inner transition elements

3. With reference to concept of ionisation potential , which one of the following sets is correct?

   (A) U > K > Cs (B) Na > K > Cs (C) Cs > U > B (D) Cs < U < K

4. Compared to the first ionisation, the value of second ionisation potential of an element is

   (A) negligible (B) smaller (C) greater (D) double

5. The ionisation energy decreases in moving down a group. This is due to

   (A) increase in nuclear charge (B) increase in atomic size and nuclear charge
   (C) increase in nuclear charge and decrease in shielding effect
   (D) increase in atomic size and also shielding effect

6. Which of the following statements concerning ionisation energy is not correct?

(A) The second ionisation energy is always more than the first

(B) Within a group, there is a gradual increase in ionisation energy because of nuclear charge

(C) Ionisation energies of the Be is more than B

(D) Ionisation energies of noble gases are high.

7. $r_{(\text{van der Waal})}$ is

(A) half the bond length (B) twice the bond length

(C) half the distance between centres of nuclei of two non bonded atoms of adjacent molecules.

(D) none of these

8. The size of the species Cl, Cl^- and Cl^+ decreases as

(A) $Cl > Cl^+ > Cl^-$ (B) $Cl^+ > Cl^- > Cl$ (C) $Cl^- > Cl^+ > Cl$ (D) $Cl^- > Cl > Cl^+$

9. The ions which are arranged in correct order of increasing radii are

(A) K^+, Ca^{2+}, S^{2-} (B) Be^{2+}, Mg^{2+}, Na^+ (C) O^{2-}, F^-, N^{3-} (D) S^{2-}, O^{2-}, As^{3-}

10. The size of following species increases in the order

(A) $Mg^{2+} < Na^+ < F^- < Al$ (B) $F^- < Al < Na^+ < Mg^{2+}$

(C) $Al < Mg^{2+} < F^- < Na^+$ (D) $Na^+ < Al < F^- < Mg^{2+}$

11. The correct order of ionic radii for the ions S^{2-}. Cl^-, P^{3-}, Ca^{2+} is

(A) $Ca^{2+} > Cl^- > S^{2-} > P^{3-}$ (B) $S^{2-} > P^{3-} > Cl^- > Ca^{2+}$

(C) $P^{3-} < S^{2-} < Cl^- < Ca^{2+}$ (D) $Ca^{2+} < Cl^- < S^{2-} < P^{3-}$

12. The anion O^{2-} is iso-electronic with

(A) F^+ (B) N^{2-} (C) F^- (D) N^{3+}

13. For an element with configuration $5s^24d^{10}5p^3$, the chemistry is likely to be similar to that of

(A) boron (B) oxygen (C) chlorine (D) phosphorus

14. Which of the follwoing elements has zero electron affinity?

(A) Na (B) Ne (C) F (D) none of these

15. The correct order of electron affinity of halogens is

(A) F > Cl > Br > I (B) I > Br > Cl > F (C) Cl > F > Br > I (D) Cl > F < Br < I

16. Outer most configuration of most electronegative element of the periodic table is

(A) $3s^23p^6$ (B) $2s^22p^5$ (C) $4s^24p^5$ (D) $2s^22p^4$

17. Diagonal relationship is shown by

(A) all elements with their diagonally opposite elements

(B) all elements of 3rd and 4th periods

(C) some of the elements of 2nd and 3rd periods

(D) elements of d-block

18. Which of the following pairs show diagonal relationship?

(A) N, P (B) N, O (C) N, S (D) None of these

19. The ionisation energy of an atom is the energy

(A) needed to add an electron to an atom

(B) needed to transfer an electron from one orbit to another

(C) needed to convert the atom into a radioactive isotope

(D) just enough to cause the electron to escape from the isolated atom

20. The maximum tendency to form the gaseous unipositive ion is for the element with configuration

(A) $1s^22s^22p^63s^2$ (B) $1s^22p^63s^1$ (C) $1s^22s^22p^63s^23p^2$ (D) $1s^22s^22p^63s^23p^3$

21. In the periodic table, with the increase in the atomic number, the metallic character of the element
(A) decreases in a period and increases in a group
(B) increases in a period and decreases in a group
(C) increases both in a period and a group
(D) decreases both in a period and a group
22. Amongst the following electronic configuration which one will have highest electron affinity?
(A) $1s^1$ (B) $1s^2 2s^1$ (C) $1s^2 2s^2 2p^4$ (D) $1s^2 2s^2 2p^5$
23. With which of the following configuration the lowest value of first IE is associated
(A) $1s^2 2s^2 2p^6 3s^1$ (B) $1s^2 2s^2 2p^5$ (C) $1s^2 2s^2 2p^6$ (D) $1s^2 2s^2 2p^6 3s^2 2p^2$
24. In the long form of the periodic table, the transition metals are placed in
(A) s-block (B) f-block (C) d-block (D) s and p-block
25. Which among the following elements have lowest value of IE_1?
(A) Pb (B) Sn (C) Si (D) C
26. The valency of noble gases, in general, is
(A) Zero (B) One (C) Three (D) Two
27. Which of the following ions are paramagetic in character?
(A) Zn^{2+} (B) Cu^+ (C) Ni^{2+} (D) Ag^+
28. Which block of the periodic table contains the element with configuration
$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
(A) s-block (B) p-block (C) d-block (D) f-block
29. The correct arrangement of increasing order of atomic radius among Na, K, Mg, Rb is
(A) Mg < K < Na < Rb (B) Mg < Na < K < Rb
(C) Mg < Na < Rb < K (D) Na < K < Rb < Mg
30. The correct order of electron affinity of B, C, N, O is
(A) O > C > N > B (B) B > N > C > O (C) O > C > B > N (D) O > B > C > N
31. The radii of the F, F^-, O and O^{2-} are in the order
(A) $O^{2-} > O > F^- > F$ (B) $F^- > O^{2-} > F > O$ (C) $O^{2-} > F^- > F > O$ (D) $O^{2-} > F^- > O > F$
32. Which of the following are isoelectronic species?

I- $\overset{+}{C}H_3$, II- NH_2^-, III- NH_4^+, IV- NH_3

(A) I, II, III (B) II, III, IV (C) I, II, IV (D) II and I
33. The element with highest electron affinity among the halogens is
(A) F (B) Cl (C) Br (D) I
34. An atom with high electronegativity has
(A) Large size (B) High ionization potential
(B) Low electron affinity (D) Low ionization potential
(C) Fluorine cannot be prepared by electrolysis of fused metal fluorides
(D) Fluorine does not form oxoacid
35. The atomic radius increases as we move down a group because
(A) Effective nuclear charge increases
(B) Atomic mass increases
(C) Additive electrons are accommodated in new electron level
(D) Atomic number increases

SECTION-C

- **Assertion & Reason**

Instructions: In the following questions as Assertion (A) is given followed by a Reason (R). Mark your responses from the following options.

(A) Both Assertion and Reason are true and Reason is the correct explanation of 'Assertion'

(B) Both Assertion and Reason are true and Reason is not the correct explanation of 'Assertion'

(C) Assertion is true but Reason is false

(D) Assertion is false but Reason is true

1. **Assertion:** Helium and beryllium have similar outer electronic configuration.

   Reason: Both are chemically inert

2. **Assertion:** First ionisation energy of nitrogen is higher than oxygen.

   Reason: Across a period effective nuclear charge increases.

3. **Assertion:** Each d-block series contains ten elements.

   Reason: The maximum capacity of d-orbitals is of ten electrons as in each series d-orbitals are gradually filled up.

SECTION-D

- **Match the following (one to one)**

Column-I and **column-II** contains **four** entries each. Entries of column-I are to be matched with some entries of column-II. Only One entries of column-I may have the matching with the same entries of column-II and one entry of column-II Only one matching with entries of column-I

1.

Column I	Column II
(A) Law of Octaves	(P) Silicon
(B) Boron diagonally related to	(Q) Ionization energy
(C) $M^{+2} + E \rightarrow M^{+3} + e$	(R) Cl
(D) Halogen	(S) Newland

2.

Column I	Column II
(A) Dobereiner triad	(P) [Xe] $6s^2$
(B) Mendeleev's periodic law	(Q) Li, Na, K
(C) Atomic number 57	(R) Transition element
(D) Alkaline earth Metal	(S) Atomic weight

EXERCISE-IV

SECTION-A

- **Multiple choice question with one correct answers**

1. Which is not a pair of representative element :

   (A) P, Co (B) Sr. Cs (C) S, K (D) P, Cl

2. In which case the energy released is the minimum in the process:

   (A) $B \rightarrow B^-$ (B) $C \rightarrow C^-$ (C) $N \rightarrow N^-$ (D) $O \rightarrow O^-$

3. Elements of the atomic number (58 to 71) are called :
   (A) Alkaline earth metals (B) Lanthanides
   (C) Actinides (D) Inner transition metals
4. Correct set with increasing electro-negativity is :
   (A) C, N, B (B) B, C, N (C) C, B, N (D) N, C, B
5. Highest melting point is shown :-
   (A) Tungsten (B) Carbon (C) Iron (D) Uranium
6. Element X belong to 4th period- it contains 18 and 1 electrons in penultimate and ultimate orbit. This X should be-
   (A) Normal element (B) Inert Gas
   (C) Transition element (D) Inner transition element
7. Diagonal relationship is exhibited because of :
   (A) Almost equal size (B) Identical electronic configuration
   (C) Both the above (D) None of the above
8. Correct order of Ist ionisation potential is
   (A) N < O < F (B) N > O < F (C) N < O > F (D) N > O > F
9. The minimum difference in the electronegativities will be shown by the pair :
   (A) F, N (B) Na, Cs (C) P, H (D) C, H
10. Decreasing order of Size of anions is :
   (A) $Br^{-} > S^{2-} > N^{3-}$ (B) $N^{3-} > S^{-2} > Br^{-}$ (C) $S^{-2}>Br^{-}>N^{3-}$ (D) None

SECTION-B

- **Multiple choice question with one or more than one correct answers**

1. Factor/ Factors affecting the ionisation potential are :
   (A) Atomic size (B) Nuclear charge
   (C) Shielding effect of orbitals (D) Type of bond in the crystal lattice
2. Which is/are a pair of representative element :
   (A) Pr, Co (B) Sn, As (C) S, K (D) P, Cl
3. Select the pair showing diagonal relationship
   (A) Li, Mg (B) Be, Al (C) B,Si (D) Mg, Sc
4. Which of the following statements is/are correct?
   (A) Radium is the first man-made element
   (B) Most of the actinide elements are man-made
   (C) Alkali metals are the lighest metals.
   (D) Alkaline earth metals are good reducing agents.
5. Which of the following is/are a Dobereiner triad?
   (A) Cl, Br, I (B) Ca, Sr, Ba (C) Li, Na, K (D) Fe, Co, Ni
6. Which of the following statements are correct ?
   (A) Greater is the nuclear charge, greater is the electron gain enthalpy.
   (B) Nitrogen has zero electron gain enthalpy.
   (C) Electron gain enthalpy decreases from fluorine to iodine in the group.
   (D) Chlorine has highest electron gain enthalpy.

7. Which of the following statements are incorrrect?
   (A) Atomic size decreases down a group.
   (B) Radius of cation is more than that of an atom.
   (C) Atomic size decreases along a period.
   (D) Radius of anion is less than that of an atom.
8. Which of the following statements are correct?
   (A) An element which high electronegativity always has high electron gain enthalpy.
   (B) Electron gain enthalpy is the property of an isolated atom.
   (C) Electronegativity is the property of a bonded atom.
   (D) Both electronegativity and electron gain enthalpy are usually directly related to neuclear charge and inversly related to atomic size.

SECTION-C

- **Comprehension**

Passage-1

The electron gain enthalpies or electron affinities of some of the elements of second period are however, lower than the corresponding elements of the third period. This is due to the reason that the elements of the second period have the smallest atomic size amongest the elements in their respecting groups. Electron gain enthalpies of elements having completely filled orbitals are zero.

1. The lower electron gain enthalpy of fluorine is lesser than that of chlorine is due to :
   (A) smaller size (B) smaller nuclear charge
   (C) diffierence in their electronic configuration (D) its highest reactivity.
2. The element that has highest electron gain enthalpy is :
   (A) Fluorine (B) Chlorine (C) Oxygen (D) Nitrogen
3. Electron gain enthalpy of inert gases is :
   (A) high (B) low but positive (C) moderate (D) almost zero

Passage-2

Electronegativity is the tendency of an atom in a molecule to attract the shared pair of electrons towards itself. It depends upon :

(i) atomic size (ii) nuclear charge

Within a period, electronegativity increase from left to right due to increasing nuclear charge.

1. An atom of an electronegative element becomes an ion by :
   (A) gain of electrons (B) loss of electrons
   (C) loss of its radius (D) seroup as a reductent
2. Outermost electronic configuration of least electronegative element in the periodic table is :
   (A) $2s^2, 2p^5$ (B) $3s^2, 3p^6$ (C) $2s^2\ 2p^4$ (D) $6s^2 6p^6 7s^1$

3. An atom with high electronegativity has :

(A) tendency to form +ve ions. (B) high ionization enthalpy

(C) large atomic size (D) low electron gain enthalpy

SECTION-D

- **Match the following (one to many)**

Column-I and **column-II** contains **four** entries each. Entries of column-I are to be matched with some entries of column-II. One or more than one entries of column-I may have the matching with the same entries of column-II and one entry of column-II may have one or more than one matching with entries of column-I

1.

Column I	**Column II**
(A) Coinage metals	(P) Tin
(B) Metal showing maximum isotopes	(Q) Cu, Ag, Au
(C) P-block elements	(R) Br, Kr
(D) Transition elements	(S) Sc, Ti, V

2.

Column I	**Column II**
(A) Halogen	(P) Oxygen
(B) Chalcogen	(Q) Chlorine
(C) Nobel gas	(R) Sulphur
(D) Representative elements	(S) Neon

*****

Answers

EXERCISE-III

SECTION-A

1.	Tranisition	2.	7, Period	3.	decreasing
4.	larger	5.	higher nuclear charge	6.	Al^{+3}
7.	Helium	8.	Hydrogen, also hydrogen		
9.	Chalcogens.	10.	Increases		

SECTION-B

1.	(C)	2.	(B)	3.	(B)	4.	(C)	5.	(D)
6.	(B)	7.	(C)	8.	(D)	9.	(B)	10.	(A)
11.	(D)	12.	(C)	13.	(D)	14.	(B)	15.	(C)
16.	(B)	17.	(C)	18.	(D)	19.	(D)	20.	(B)
21.	(A)	22.	(D)	23.	(A)	24.	(C)	25.	(A)
26.	(A)	27.	(C)	28.	(C)	29.	(B)	30.	(A)
31.	(D)	32.	(B)	33.	(B)	34.	(B)	35.	(C)

SECTION-C

1.	(C)	2.	(D)	3.	(A)

SECTION-D

1. (A)-(S),(B)-(P),(C)-(Q),(D)-(R)
2. (A)-(Q), (B)-(S),(C)-(R),(D)-(P)

EXERCISE-IV

SECTION-A

1.	(A)	2.	(C)	3.	(B)	4.	(B)	5.	(B)
6.	(C)	7.	(A)	8.	(B)	9.	(B)	10.	(A)

SECTION-B

1.	(A,B,C)	2.	(B,C,D)	3.	(A,B,C)	4.	(B,C,D)	5.	(A,B,C)
6.	(A,D)	7.	(A,B,D)	8.	(B,D)				

SECTION-C

Passage-1	1.	(A)	2.	(B)	3.	(D)
Passage-2	1.	(A)	2.	(D)	3.	(B)

SECTION-D

1. (A)-(Q), (B)-(P), (C)-(P,R),(D)-(Q,S) 2.(A)-(Q), (B)-(P,R), (C)-(S),(D)-(P,Q,R)