

## CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

### 1 NEED FOR CLASSIFICATION

It is very difficult to study individually the chemistry of more than one hundred elements known and their innumerable compounds. This practical problem was felt by scientists and after numerous attempts the scientists were ultimately successful in arranging the elements. This arrangement of elements provided a proper classification of elements, which led to the formation of a periodic table.

Periodic table may be defined as the table, which classifies all the known elements in accordance with their properties in such a way that the elements with similar properties are grouped together in the same vertical column and dissimilar elements are separated from one another.

The periodic table provides a systematic framework for organizing the large available information on the chemical behavior of the elements into simple logical patterns. Using the position of an element in the periodic table, it is possible to predict its chemical behavior.

### 2 HISTORICAL DEVELOPMENT OF CLASSIFICATION OF ELEMENTS

Several scientists for long have tried to classify the elements and to find patterns in their properties. Some of them are as:

#### 2.1 DOBEREINER'S TRIADS

He arranged similar elements in groups of three and showed that atomic weights are either nearly the same or the atomic weight of the middle element is approximately the arithmetic mean of the other two. e.g.

Element	Atomic weight	Element	Atomic weight	Element	Atomic weight
Li	7	Ca	40	Cl	35.5
Na	23	Sr	88	Br	80
K	39	Ba	137	I	127

**Drawback:** This applies only to limited number of elements.

#### 2.2 NEWLAND LAW OF OCTAVE

When lighter elements are arranged in order of their increasing atomic weight, the properties of every eighth element is similar to the first, like eighth note of a musical scale.

Element	Li	Be	B	C	N	O	F
Atomic weight	7	9	11	12	14	16	19
Element	Na	Mg	Al	Si	P	S	Cl
Atomic weight	23	24	27	29	31	32	35.5
Element	K	Ca					
Atomic weight	39	40					

**Drawbacks:** (a) It could not be applied for the element of atomic mass greater than 40.

(b) With the discovery of noble gases, the above sequence was disturbed and the arrangement was no longer valid.

#### 2.3 LOTHER MEYER ARRANGEMENT

By plotting graph between atomic volume and atomic mass of element it was observed that elements having similar properties occupy similar positions on the curve e.g. alkali metals occupy maxima of the curve. Most electronegative elements found ascending portions of the curve and it was found that physical properties are the periodic function of their atomic weights.

By 1868, Lothar Meyer had developed a table of the elements that closely resembles with the Modern Periodic Table. However, his work was not published until after the work of Dimitri Mendeleev, the scientist who is generally credited with the development of the Modern Periodic Table.

## 2.4 MENDELEEV'S PERIODIC TABLE

While Dobereiner initiated the study of periodic relationship, it was Mendeleev who was responsible for publishing the Periodic law for the first time.

Thus, in the development of the periodic table the work of Dmitri Mendeleev was a breakthrough. To explain his scheme of classification, he gave a periodic law which is known as *Mendeleev's periodic law*. This law states as follows, *the physical and chemical properties of elements are periodic function of their atomic masses* i.e. when the elements are arranged in order of their increasing atomic masses, elements with similar properties are repeated after certain regular intervals.

On the basis of this periodic law, he constructed a periodic table in such a way that elements were arranged horizontally in the order of their increasing atomic masses.

Mendeleev's system of classifying elements was more elaborated than that of Lothar Meyer's. He fully recognized the significance of periodicity and used broader range of physical and chemical properties to classify the elements.

### 2.4.1 Characteristics of Mendeleev's Periodic Table

1. It has 9 vertical columns called groups I, II, III, IV, V, VI, VII, VIII and zero group. Except VIII and zero groups, each group is subdivided into two groups designated as A and B.
2. It has seven horizontal rows called periods. The I, II and, III periods are short periods, IV and V periods are long periods and VI period is the longest period.

### 2.4.2 Importance of Periodic table

1. Classification based on atomic mass facilitates systematic study of the elements.
2. It helped in prediction of new element e.g. Ga, Ge. Keeping his primary aim of arranging the elements of similar properties in the same group, he proposed that some of the elements were still undiscovered and, therefore, left several gaps in the table. For example, both gallium and germanium were unknown at the time when Mendeleev published his Periodic Table. He left the gap under aluminium and a gap under silicon, and called these elements Eka-aluminium and Eka-silicon. Mendeleev predicted not only the existence of gallium and germanium, but also described some of their general physical properties.
3. Correction of doubtful atomic mass. The atomic mass of Be was taken as 13.5 but periodic classification helped in arriving at its correct atomic mass i.e. 9.

### 2.4.3 Drawbacks of Mendeleev's Periodic Table

1. **Position of H:** H is placed along with alkali metals in group I(A). It can also be placed along with halogens in group VIIA.
2. **Anomalous positions of some elements:** Although most of the elements have been arranged in the increasing order of their atomic masses, but in some cases, the element having higher atomic mass precedes the element with lower atomic mass. For example, Ar (Atomic mass = 39.9) precedes K (Atomic mass = 39.1) and similarly, Co (Atomic mass = 58.9) precedes Ni (Atomic mass = 58.7).
3. **Position of isotopes:** Different isotopes of the same element are not given separate position in periodic table.

- No co-relation of elements in sub-groups:** Some dissimilar elements are grouped together while some similar elements are placed in different groups. Alkali metals (Li, Na) and coinage metals (Cu, Ag, Au) are placed in the same group although they have different properties. Copper and mercury have been placed in different groups although they have similar chemical properties.
- Position of lanthanides and actinides:** These elements could not be accommodated in the main periodic table. Instead they have been placed in two separate rows at the bottom of the periodic table.
- Cause of periodicity:** No proper explanation has been offered as to why the elements placed in a group show resemblance in their properties.

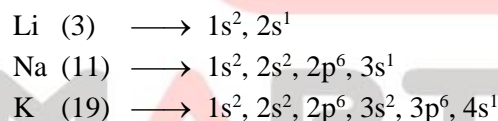
### 3 MODERN PERIODIC LAW AND LONG FORM OF PERIODIC TABLE

When Mendeleev developed his Periodic Table, chemists knew nothing about the internal structure of atom. However, the beginning of the 20<sup>th</sup> century witnessed profound developments in theories about sub-atomic particles. In 1913, the English physicist, Henry Moseley observed regularities in the characteristic X-ray spectra of the elements. A plot of  $\sqrt{\nu}$  (where  $\nu$  is frequency of X-rays emitted) against atomic number ( $Z$ ) gave a straight line and not by the plot of  $\sqrt{\nu}$  vs. atomic mass. Thus, Moseley observed that atomic number is a more fundamental property of an element than its atomic mass, and physical and chemical properties of the elements are determined by their atomic numbers. Based on this, **modern periodic law** was given which states that **physical and chemical properties of the elements are periodic functions of their atomic numbers**. i.e. if the elements are arranged in increasing order of their atomic numbers, the elements with similar properties are repeated after certain regular intervals.

#### 3.1 CAUSE OF PERIODICITY

When the elements are arranged in an order of increasing atomic number, the periodic repetition of elements with similar properties is observed after certain regular intervals. This is called periodicity.

The cause of periodicity in properties is due to repetition of similar outer electronic configuration after certain regular intervals as shown below:



#### 3.2 FEATURES OF LONG FORM OF PERIODIC TABLE

##### 3.2.1 Periods

A horizontal row of a periodic table is known as a period. Entire table is divided into 7 periods. The first period contains 2 elements. The subsequent periods consist of 8, 8, 18, 18 and 32 elements, respectively. The seventh period is incomplete and like the sixth period would have a theoretical maximum (on the basis of quantum numbers) of 32 elements. In this form of the Periodic Table, 14 elements of both sixth and seventh periods (lanthanoids and actinoids, respectively) are placed in separate panels at the bottom. Thus, 1<sup>st</sup>, 2<sup>nd</sup> and 3<sup>rd</sup> period are short periods, 4<sup>th</sup> and 5<sup>th</sup> periods are long and 6<sup>th</sup> period is the longest and 7<sup>th</sup> period is incomplete.

##### 3.2.2 Groups

A vertical column of the periodic table is called a group. It is divided into 18 groups. According to IUPAC, these groups are numbered from 1 to 18, replacing the older notation of groups I A ... VII A, VIII, I B ... VII B and O.

#### 3.3 ADVANTAGES OF LONG FORM OF PERIODIC TABLE

Long form of periodic table has the following advantages:

1. Same group elements have similar electronic configuration, hence they have similar chemical properties.
2. Elements are properly divided into reactive metals, heavy metals and non metals.
3. Since same group elements have similar electronic configuration, the study of elements become very simple.
4. Since, classification is based on the atomic number and not on the atomic mass, the idea of placing isotopes at one place is fully justified.

### 3.4. DRAWBACKS OF THE LONG FORM OF PERIODIC TABLE

The main defects of this table are as follows:

1. **Position of H:** H is still placed along with alkali metals as well as halogens.
2. **Separate position of lanthanides and actinides:** Like Mendeleev's periodic table lanthanides and actinides are placed separately at the bottom of the periodic table.

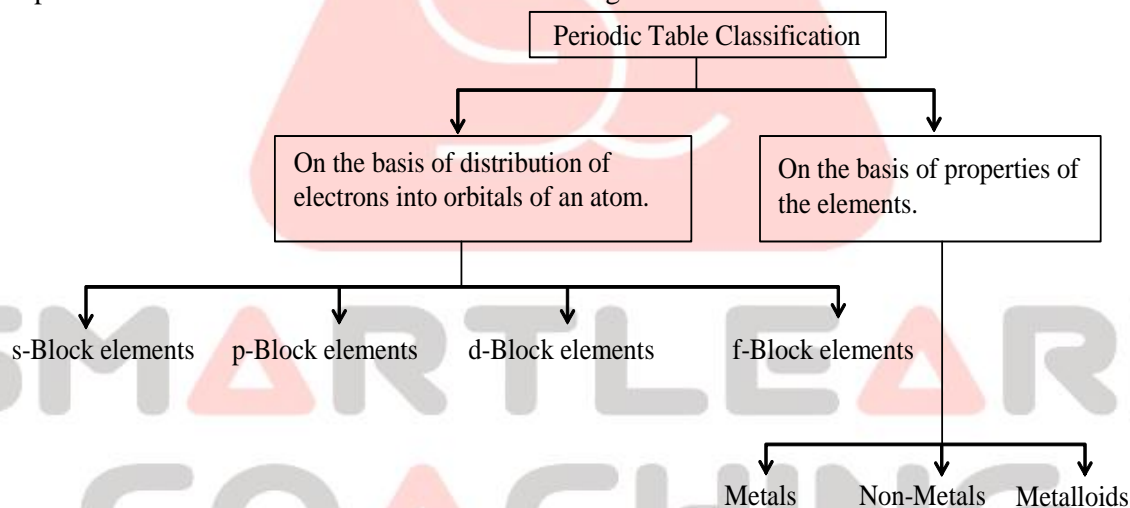
### PRACTICE PROBLEMS

PP1. What is the modern periodic law?

PP2. What is the difference between Mendeleev's and modern periodic table?

## DIVISION OF PERIODIC TABLE

The entire periodic table can be divided into two broad categories as indicated below:



### 4.1 CLASSIFICATION ON THE BASIS OF ELECTRONIC DISTRIBUTION OF ELECTRONS INTO ORBITALS OF AN ATOM

Entire periodic table is divided into four blocks known as *s*, *p*, *d* and *f* based on the orbital in which last electron enters.

**4.1.1 s-BLOCK ELEMENTS:** The elements of group 1 (alkali metals) and Group 2 (alkaline earth metals) which have  $ns^1$  and  $ns^2$  outermost electronic configuration belong to the s-Block Elements. The general features of these elements are:

1. The general outer shell electronic configuration of s-block elements is  $ns^{1-2}$ .
2. They are soft metals with low melting and boiling points.
3. They have low ionization potential and are very reactive.
4. Alkali metals show +1 oxidation state while alkaline earth metals show +2 oxidation state.
5. Most of the elements of this block impart characteristic colour to the flame.



6. They are strong reducing agents.
7. All of them are good conductors of heat and electricity.

**4.1.2 p-BLOCK ELEMENTS:** The p-Block elements comprise those belonging to group 13 to 18 and these together with the s-Block elements are called the Representative Elements or Main Group Elements. The general features of these elements are:

1. The general outer shell electronic configuration of p-block elements is  $ns^2, np^{1-6}$ .
2. This block has metals, non-metals and metalloids.
3. Some of the elements show more than one oxidation states.
4. They form mostly covalent compounds.
5. No member of p-block or the salt of it imparts a characteristic color to the flame because the energy released by excited elements does not appear in the visible region of the spectrum.

**Note:** Group-18 elements are known as noble gases; Group-17 elements are known as halogens and the group-16 elements are known as chalcogens.

### 4.1.3 d-BLOCK ELEMENTS

These are the elements of group 3 to 12 in the centre of the periodic table. These are characterized by the filling of inner d orbitals by electrons and are therefore referred to as d-block elements. The general features of these elements are:

1. The general outer shell electronic configuration is  $(n - 1)d^{1-10} ns^{0-2}$ .
2. They are also known as transition elements because their properties show a gradual transition from the properties of s-block to the p-block elements.
3. They are hard, malleable and ductile metals.
4. They form both ionic and covalent compounds.
5. Their compounds are generally coloured and paramagnetic.
6. Most of the elements or their compounds find their use as catalyst due to their tendency to show variable oxidation states.
7. Most of the transition elements form alloys.

**4.1.4 f-BLOCK ELEMENTS:** The two rows (series) of the elements at the bottom of the periodic table, called the Lanthanides and Actinides are classified into the f-block elements. The general features of these elements are:

1. The general outer shell electronic configuration is  $(n - 2)f^{1-14} (n - 1)d^{0-1} ns^2$ .
2. They are known as inner transition elements since they form transition series within the transition elements of d-block.
3.  ${}_{58}\text{Ce}$  to  ${}_{71}\text{Lu}$  constitutes the 4f inner transition series and are called Lanthanides because they follow lanthanum in the 5d transition series. They closely resemble lanthanum in their properties. They are also called rare earth elements since they occur rarely in earth crust.
4.  ${}_{90}\text{Th}$  –  ${}_{103}\text{Lr}$  constitutes the 5f inner transition series and are known as actinides since they follow actinium in the 6d transition series.
5. All actinides are radioactive elements. Three elements namely Thorium, Protoactinium and Uranium occur in nature but the remaining 11 elements i.e. from neptunium to Lawrencium ( ${}_{93}\text{Np}$  –  ${}_{103}\text{Lr}$ ) have been prepared artificially. These 11 elements are called Transuranium elements since they follow uranium in periodic table and also have been derived from it through nuclear reactions.
6. They show variable oxidation states.
7. Their compounds are generally coloured.
8. They have strong tendency to form complexes.

## 4.2 CLASSIFICATION ON THE BASIS OF PROPERTIES OF THE ELEMENTS.

### I. Metals:

Metals comprise more than 78% of all known elements and appear on the left side of the Periodic Table. The general properties of metals are listed below:

- (i) Metals are usually solids at room temperature. Mercury is an exception.
- (ii) Metals usually have high melting and boiling points.
- (iii) They are good conductors of heat and electricity.
- (iv) They are malleable (can be flattened into thin sheets by hammering) and ductile (can be drawn into wires)

## II. Non-metals:

In contrast to metals, non-metals are located at the top right hand side of the periodic table. In fact, in a horizontal row, the property of elements changes from metallic on the left to non-metallic on the right and the elements become more metallic as we go down a group. The characteristics of the non-metals are as follows:

- (i) Non-metals are usually solids or gases at room temperature with low melting and boiling points (boron and carbon are exceptions).
- (ii) They are poor conductors of heat and electricity.
- (iii) Most non-metallic solids are brittle and are neither malleable nor ductile.

## III. Metalloids

Certain elements (Si, Ge, As, Sb, Te, Po and At) show properties that are characteristic of both metals and non-metals. These elements are called semi-metals or metalloids.

## 5. EFFECTIVE NUCLEAR CHARGE

The effective nuclear charge is the actual charge felt by the valence electrons. Between the outer most valence electrons and the nucleus of an atom, there exists finite number of shells containing electrons. Due to the presence of these intervening electrons, the valence electrons are unable to experience the attractive pull of the actual number of protons in the nucleus. These intervening electrons act as shield between the valence electrons and protons in the nucleus. Thus, the presence of intervening (shielding) electrons reduces the electrostatic attraction between the protons in the nucleus and the valence electrons because intervening electrons repel the valence electrons. The concept of effective nuclear charge allows us to account for the effects of shielding on periodic properties.

## 6 THE PERIODICITY OF ATOMIC PROPERTIES

The electronic configuration of the elements shows a periodic variation with increasing atomic number. Consequently, there are also periodic variations in physical and chemical behaviour of the elements. Let us now study the variation of some of the atomic properties in the periodic table. These properties are classified as follows:

- (i) Atomic radius
- (ii) Ionisation energy
- (iii) Electron gain enthalpy
- (iv) Electronegativity

## 7 TRENDS IN PHYSICAL PROPERTIES

### 7.1 ATOMIC RADIUS

It is defined as the distance from the centre of the nucleus to the outer most shell containing electrons.

#### Difficulties in determining radius

- (a) Exact size of electron cloud can not be determined as the probability of finding an electron even at large distance from the nucleus never becomes zero.
- (b) It is not possible to isolate a single atom.

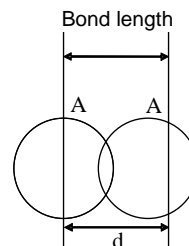
- (c) The probability of finding an electron in an atom is appreciably affected by the presence of the other atoms within its molecule or its neighbourhood.
- (d) The atomic radius also changes when the atom is present in different bonding state.

## Type of Radius

### 7.1.1 Covalent radius

It is defined as one-half of the distance between the nucleus of two covalently bonded atoms of the same element in a molecule.

$r_{\text{covalent}} = \frac{d}{2}$  where 'd' is the internuclear distance between two bonded atoms.



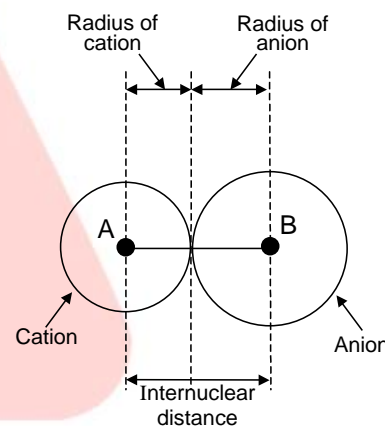
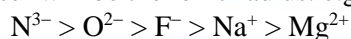
Overlapping of atomic orbitals to form a covalent bond.

### 7.1.3 Metallic radius

It is defined as one-half of the internuclear distance between two adjacent metal atoms in the metallic lattice.

### 7.1.4 Ionic radius

Ionic radii may be defined as the effective distance from the nucleus of the ion to the point upto which it exerts its influence on its electron cloud. Radius of cation is always smaller than that of the parent atom because it has fewer electrons while its nuclear charge remains the same. e.g. the radius of  $\text{Na}^+$  is lesser than Na. Similarly the radius of anion is larger than that of the parent atom because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge. e.g. the radius of  $\text{I}^-$  is greater than I. For isoelectronic species greater the nuclear charge lesser will be the ionic radius. e.g.



$$d_{\text{AB}} = r_{\text{cation}} + r_{\text{anion}}$$

### 7.1.5 Variation of atomic radii in the periodic table

In general, the atomic radii decreases with increase in the atomic number in a period. It is because of increase in the effective nuclear charge along the period.

In general, the atomic radii increases from top to bottom within a group. As we move down the group there is progressive increase in the principal energy level, which increase the distance between nucleus and outermost electron and the size increases.

## 7.2 IONIZATION ENERGY

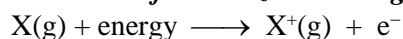
**Ionization energy is defined as the minimum energy (in kJ/mole) required to remove the most loosely bound electron from an isolated gaseous atom in its ground state.** Alternatively, **ionization energy is the amount of energy required in kilojoules needed to knock out one mole of electrons from one mole of isolated gaseous atoms.** In this definition, gaseous atoms are specifically used because an atom in the gas phase is virtually uninfluenced by its neighbours and so, there are no intermolecular forces to take into account while measuring ionization energy.

The magnitude of ionization energy is a measure of how “tightly” the electron is held in the atom. Higher is the ionization energy, more difficult it is to remove the electron.

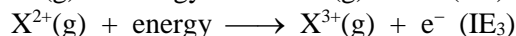
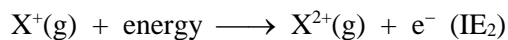
### 7.2.1 Successive Ionization Energies

For a multi-electron atom, the amount of energy required to remove the first electron from the atom in its

ground state is called the **first ionization energy** ( $IE_1$ ).



The second ionization energy ( $IE_2$ ) and the third ionization energy ( $IE_3$ ) are shown in the following equations:



The pattern continues for the removal of subsequent electrons. Theoretically, an atom can have as many ionization energy values as the number of electrons present in that atom.

The ionization enthalpies required to remove first, second, third etc. electrons from an isolated gaseous atom are called successive ionization energies. It may be seen that,  $IE_2$  is always greater than  $IE_1$  and  $IE_3$  is always greater than  $IE_2$  and so on  $IE_3 > IE_2 > IE_1$ . After removing first electron, the resultant force of attraction on the remaining outermost electrons increases and hence, more energy is required to remove the next electron.

### 7.2.2 Factors affecting the ionization energy.

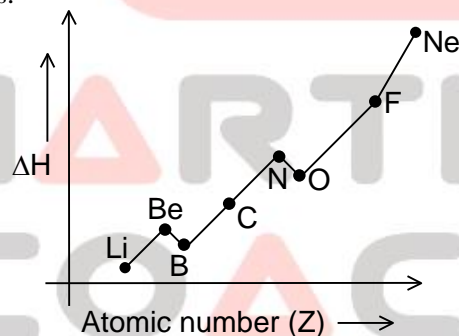
Ionization energy depends upon the following factors:

- Penetration effect of the electrons. i.e. ionization energy of  $ns > np > nd > nf$ , because s-orbital is closest to the nucleus.
- Greater the nuclear charge greater will be the ionization energy.
- Greater the atomic size lesser will be the ionization energy as the distance of the outermost electron from the nucleus increases.
- Greater the screening effect of inner electrons, lesser will be the ionization enthalpy.
- Exactly half filled or completely filled orbitals have higher ionization enthalpy because half filled and fully filled configuration are more stable than partially filled configuration.

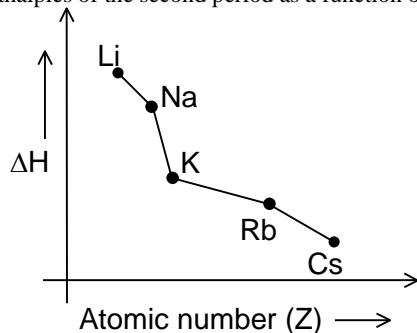
### 7.2.3 Variation of ionisation energy in the periodic table

In general the value of ionisation energy increases with the increase in atomic number across the period. This is due to gradual increase in effective nuclear charge and simultaneous decrease in atomic size. As a result, the attractive force between the nucleus and the electron cloud increases. Consequently, the electron is more and more tightly bound to the nucleus.

As we move down the group, there is a gradual increase in the atomic size due to progressive addition of new energy shells. As a result the nuclear hold on the valence electrons decreases gradually and the ionisation energy also decreases.



First ionization enthalpies of the second period as a function of atomic number (Z).



First ionization enthalpies of the first group as a function of atomic number (Z)

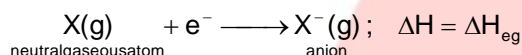


From the above graphs, it is clear that the first ionization enthalpy of boron ( $Z = 5$ ) is slightly less than that of beryllium ( $Z = 4$ ) even though the former has a greater nuclear charge. When we consider the same principal quantum level, an s-electron is attracted to the nucleus more than a p-electron. In beryllium, the electron removed during the ionisation is a s-electron whereas the electron removed during ionization of boron is a p-electron. The penetration of a 2s-electron to the nucleus is more than that of a 2p-electron; hence the 2p-electron of boron is more shielded from the nucleus by the inner core of electrons than the 2s-electrons of beryllium. Therefore, it is easier to remove the 2p-electron from boron compared to the removal of a 2s-electron from beryllium. Thus, boron has a smaller first ionization enthalpy than beryllium.

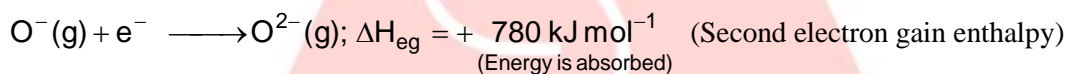
The first ionization enthalpy of oxygen is smaller as compared to nitrogen. This can be attributed due to stability of half filled orbitals of 2p-subshell of nitrogen.  
 $r$  ionization =  $5140 \times 10^{-2} = 51.40$  kJ.

### 7.3 ELECTRON GAIN ENTHALPY

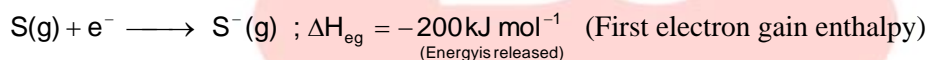
**Electron gain enthalpy of an element may be defined as the energy released when a neutral isolated gaseous atom accepts an extra electron to form the gaseous negative ion, i.e., anion. It is denoted by  $\Delta H_{eg}$ . Higher is the amount of energy released, more is the electron affinity value.** This process may be represented as:



After the addition of one electron, the atom becomes negatively charged and the second electron is to be added to a negatively charged ion. But the addition of second electron is opposed by electrostatic repulsion and hence the energy has to be supplied for the addition of a second electron. Thus, the second electron gain enthalpy of an element is positive.



Similarly, the second electron gain enthalpy of S is also positive as can be seen from the data given below:



The elements which have stable configuration have positive first electron gain enthalpy.

#### Factors on which the electron gain enthalpy depends.

Some important factors on which electron gain enthalpy depends are discussed below:

##### 1. Atomic size

As the size of an atom increases, distance between the nucleus and the last shell which receives the incoming electron increases. As a result, the force of attraction between the nucleus and the incoming electron decreases and hence the electron gain enthalpy becomes less negative.

##### 2. Nuclear charge

As the nuclear charge increases, the force of attraction between the nucleus and the incoming electron increases and hence, the electron gain enthalpy becomes more negative.

##### 3. Electronic configuration

Elements having exactly half-filled or completely filled orbitals are very stable. As a result, energy has to be supplied to add an extra electron. Hence their electron gain enthalpies have large positive values since they do not accept additional electron so easily.

#### 7.3.1 Variation of electron gain enthalpy in the periodic table

On moving across the period, the atomic size decreases and effective nuclear charge increases. Both these factors result into greater attraction for the incoming electron. Therefore, electron gain enthalpies tend to become more negative as we go from left to right across a period. On moving down a group, the atomic size as well as nuclear charge increases. But the effect of increase in atomic size is much more pronounced than that of the nuclear charge and thus, the additional electron feels less attraction. Consequently, electron gain enthalpy becomes less negative on going down the group.

The electron gain enthalpy of second period element is lower than that of third period elements. This is probably due to small size of the atom of second period element. The addition of an extra electron produces high electron charge density in a relatively compact 2p-subshell resulting in strong electron-electron repulsion. The repulsive forces between electrons imply low electron gain enthalpy. Electron gain enthalpy of Cl is maximum in periodic table.

## 7.4 ELECTRONEGATIVITY

**The ability of an atom to attract the shared pair of electrons towards itself is known as electronegativity.** It is a unitless quantity because it is relative with respect to F for which the electronegativity value is fixed as 4. (i) Electronegativity of inert gases is 0. (ii) It is not a property of an isolated atom but rather property of an atom in a molecule. (iii) Because of electronegativity covalent molecule acquire some ionic character. (iv) Fluorine is the most electronegative element of the periodic table.

### 7.4.1 Factors affecting electronegativity

- Size of atom:** As the size increases, electronegativity decreases because the distance of electrons from nucleus increases.
- Oxidation state of element:** Electronegativity increases with higher positive oxidation state because it has more effective nuclear charge. As for example electronegativity of  $\text{Fe}^{3+}$  ion is greater than that of  $\text{Fe}^{2+}$  ion.
- Effective nuclear charge:** Electronegativity increases with increase in nuclear charge because the attraction on shared pair of electrons increases.

### 7.4.2 Application of electronegativity

- To predict the nature of bond.
- To calculate the percentage ionic character.
- To explain the variation in bond angle.
- To explain the bond length variation.

### 7.4.3 Variation of electronegativity in the periodic table

In a period, electronegativity increases on moving from left to right. This is because the effective nuclear charge increases whereas atomic radius decreases as we move from left to right in a period. In a group, electronegativity decreases as we move down the group. This is due to the effect of increased atomic radius.

## 6 PERIODIC TRENDS IN CHEMICAL PROPERTIES

### 8.1 Periodicity of valency or oxidation states

Valency of representative elements is usually equal to the number of electrons in the outermost orbit and or equal to eight minus the number of outermost electrons as shown below:

Group	1	2	13	14	15	16	17	18
Number of valence electrons	1	2	3	4	5	6	7	8
Valency	1	2	3	4	3,5	2,6	1,7	0,8

Now a days the term oxidation state is frequently used for valency. The oxidation state of an element in a particular compound can be defined as the charge acquired by its atom on the basis of electronegative consideration from other atoms in the molecule.

### 8.2 Anomalous Properties Of Second Period Elements-Diagonal Relationship

A diagonal relationship is a similarity in properties between diagonal neighbours in the main groups of the periodic table. Because these properties lie in a diagonal pattern, it is not surprising to find that the elements within a diagonal band show similar chemical properties. Diagonal relationship helps in making predictions about the properties of elements and their compounds.

Diagonal relationship is exhibited by lithium and magnesium; beryllium and aluminium. For example, lithium and magnesium react directly with nitrogen to form nitrides. Beryllium and aluminum, both are amphoteric i.e. reacts with acids as well as bases.

Reasons for different chemical behavior of the first member of a group of elements in the s- and p-blocks compared to that of the subsequent members in the same group:

- (i) The anomalous behaviour is attributed to their small size, large charge / radius ratio and high electronegativity of the elements.
- (ii) The first member of the group don't have d orbitals.
- (iii) The first member of p-block elements display greater ability to form p-p multiple bonds to itself (e.g.  $C=C$ ,  $C\equiv C$ ,  $N=N$ ,  $N\equiv N$ ) and to other second period elements (e.g.  $C=O$ ,  $C=N$ ,  $C\equiv N$ ,  $N=O$ ) compared to subsequent members of the same group.



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