### Method of prediction of period , group

By knowing the electronic configuration. example: sodium(atomic no-11)-
electronic configuration = 2 , 8 , 1. Sodium have one electron in the last orbit hence it is the member of first group. There are 3 shells occupied by electrons, hence it is the member of 3rd period.

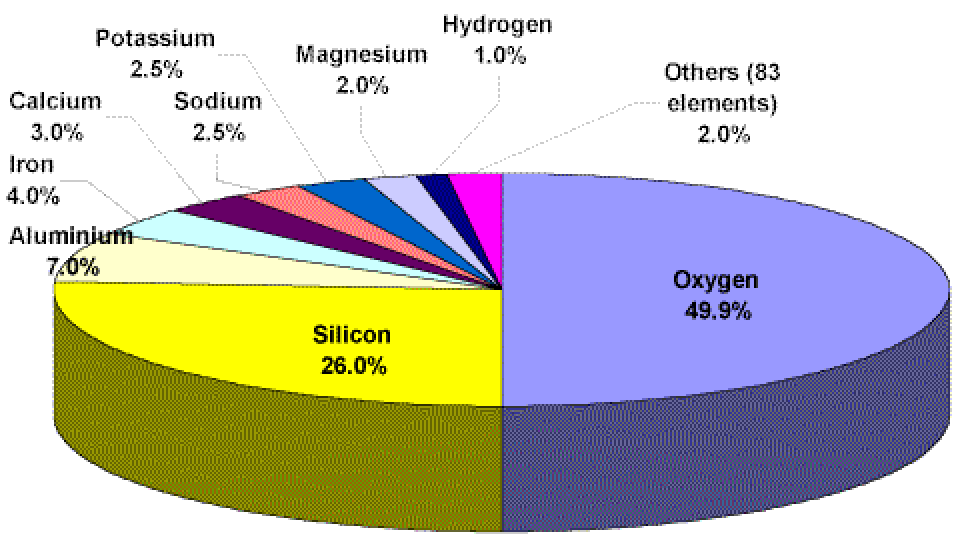
### Define an Element

An **element** is a species of an atom which has the same number of protons in the atomic nuclei.
Examples: Any atom which has 3 protons in its atomic nuclei is Lithium (element).
There are 118 elements that have been identified. First 94 elements occur naturally on Earth while the remaining 24 are synthetic elements.

### History of Classification of Elements

In 1789, Antoine Lavoisier made the earliest attempt to classify the elements. He grouped the elements based on their properties. Several other attempts were made to group elements together over the coming decades. In 1829, Johann Dobereiner recognized triads of elements with chemically similar properties.

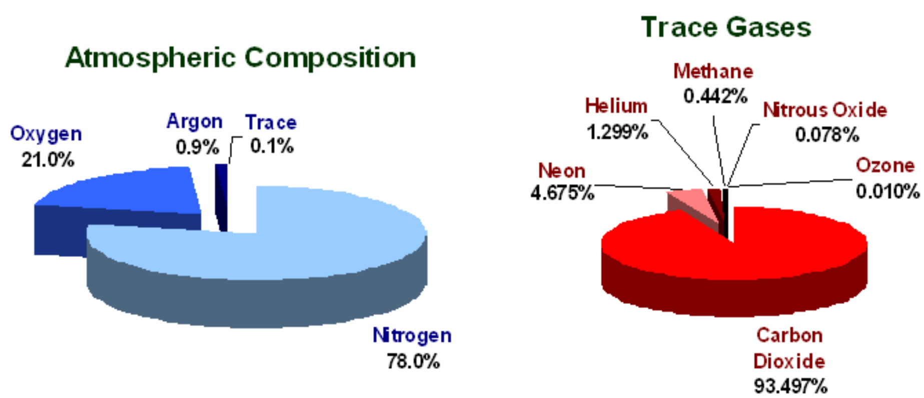
### Occurrence of Elements in Earth's Crust




The most abundant element in the earth's crust is oxygen, making up 46.6% of the earth's mass. Silicon is the second most abundant element (27.7%), followed by aluminum (8.1%), iron

(5.0%), calcium (3.6%), sodium (2.8%), potassium (2.6%). and magnesium (2.1%). These eight elements account for approximately 98.5% of the total mass of the earth's crust.

**Occurrence of Elements in Atmosphere**




The atmosphere is composed of several gases in a certain amount. Nitrogen accounts for 78% of the atmosphere, oxygen 21%, and argon 0.9%. Carbon dioxide accounts for about a 90% of the 0.1% trace gases. Water vapor concentration varies from 0-4% of the atmosphere depending upon location and time.

**Physical Properties of Elements**

Physical Properties of a substance are the properties that are observable. Some of the physical properties are:

Color
Luster
Freezing point
Boiling point
Melting point
Density
Hardness
Odour

**Chemical Properties of Elements**

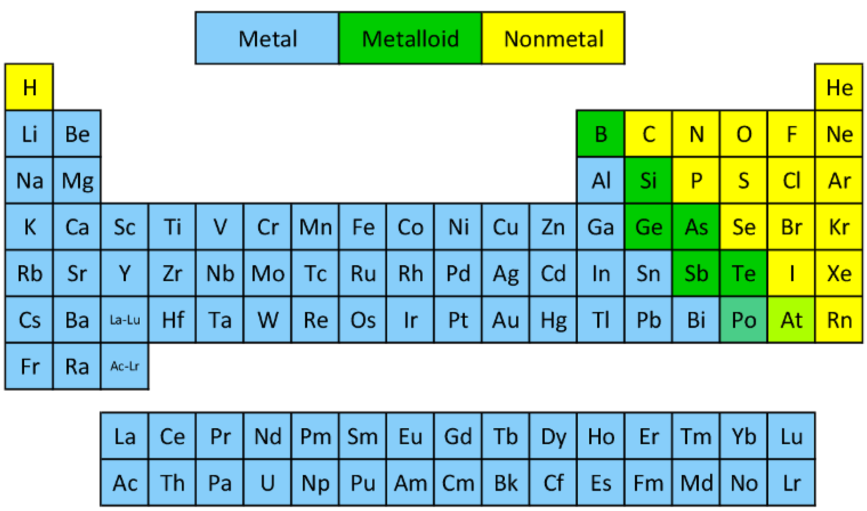
Chemical Properties of a substance determines the chemical behavior of substance. It explains how the substance reacts with other substances. These properties are only observable during a chemical reaction.
Example: Flammability, oxidizing or reducing nature, etc.

**Latest Classification of Elements**

According to the latest theory of periodic table, elements are classified into:

Metallic Elements (Metals)
Non-Metallic Elements (Non-Metals)
Metalloids
Noble Gases

**Metallic Elements and its Properties**




**Metallic Elements (Metals)** are elements which show following properties.

High electrical conductivity
Malleability
Ductility
Luster

Readily loses electrons to form cations
Examples: Sodium, Copper, Magnesium, etc.
Metals occupies the positions as shown in the periodic table. They are further classified into four groups. These are alkali metals, alkali earth metals, rare earth metals and transition metals.

#### Non-Metallic Element and its Properties

**Non-metallic** elements are elements which show following properties.

Good insulator of heat and electricity
Brittle in nature
Non-lustrous
Readily gain electrons to form negative ions
Generally, all non-metallic properties showing elements are non-metal (non-metallic elements).

#### Define Metalloids

**Metalloid** is a chemical element which shows properties of both metal and non-metal. Some of the metalloids are - Silicon, Germanium, Arsenic, Antimony, Tellurium and Polonium.

#### Define Noble Gases

The noble gases (inert gases) are monatomic gases with very low chemical reactivity. They are colourless and odourless in nature. They occupy group 18 position in the periodic table. The six noble gases that occur in nature are helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe). Remaining Elements are all synthetic elements.

**IUPAC Nomenclature of Elements With Atomic Number Greater Than 100**

| Digit | name |
|---|---|
| 0 | Nil |
| 1 | Un |
| 2 | Bi |
| 3 | Tri |
| 4 | Quad |
| 5 | Pent |
| 6 | Hex |
| 7 | Sept |
| 8 | Oct |
| 9 | Enn |

IUPAC name of the element (Atomic Number >100) is dependent on its atomic number. The given table gives the code for the given number. Write the codes in the order of digits and add ium at end to get the name of an element.

Example: Let us take element with atomic number 108.

Code for 1 is un

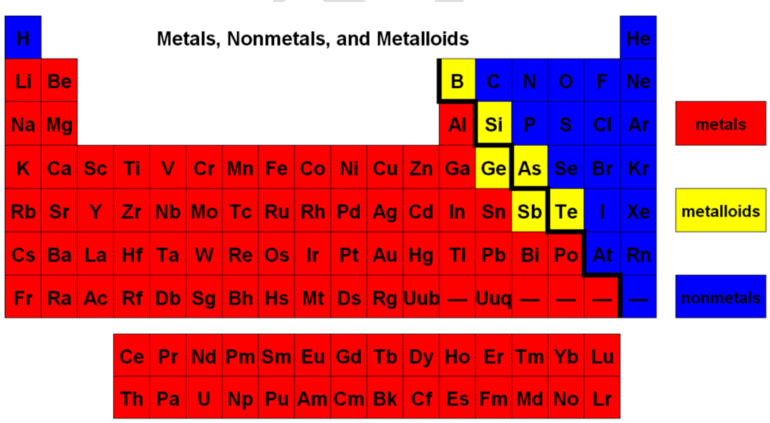
Code for 0 is nil

Code for 8 is oct

Hence, the name of the element is un + nil + oct + ium (ending code) = unniloctium

**Note:** If the last digit code ends with i then add um as ending code instead of ium Example: 102 - Unnilbium

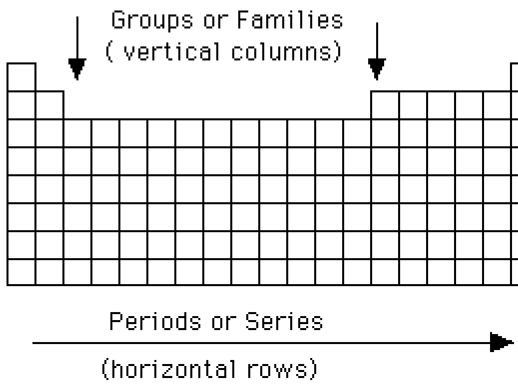
**Position of Elements in the Periodic Table**



Location of Metals, Non-metals, Metalloids and Noble Gases are as follows:

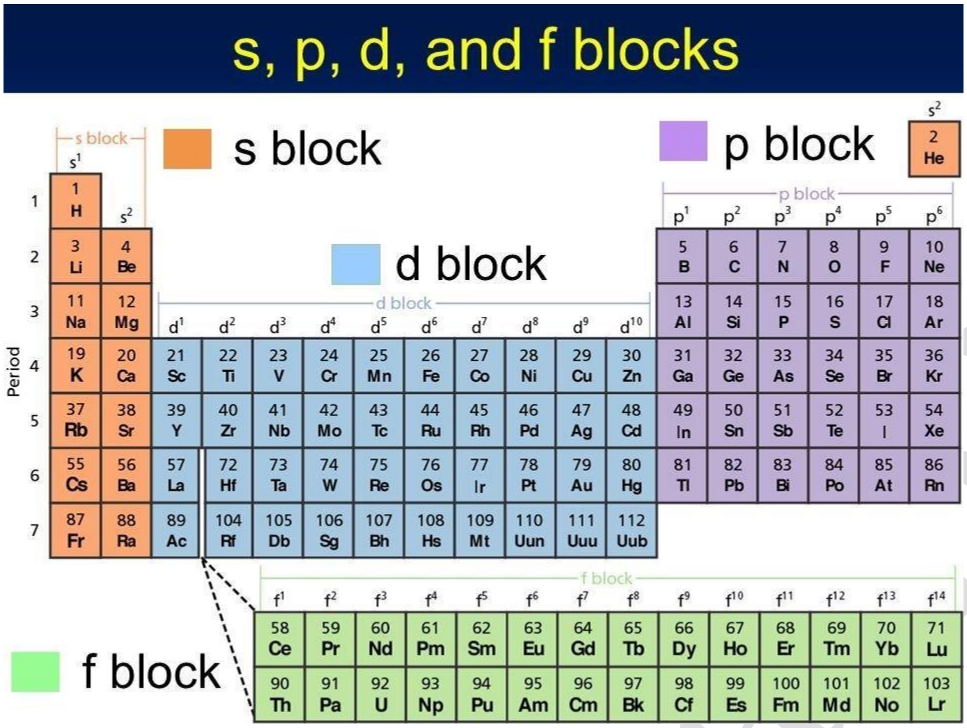
Elements on the left and middle are **Metals**
Elements on the right are **Non-metals**
**Metalloids** form the narrow stair-step area between metals and non-metals
The last group (18) to the right are **Noble gases**

**Periodic Table Chart**




The **Periodic Table** organizes the elements according to their similar chemical and physical properties. The Table has rows and columns.The vertical columns in the periodic table represent **Groups**. The horizontal rows in table represents **Periods**.

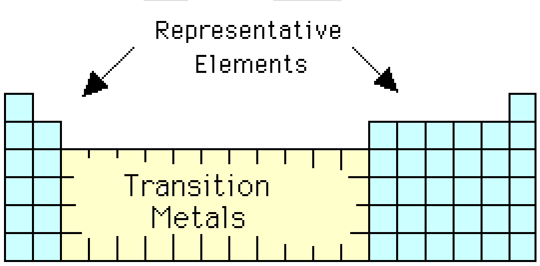
**Types of Elements - s, p ,d and f Blocks**




Elements in the group exhibit similar chemical behaviour. The similarity is because of two factors:

Same number of electrons in outermost orbital
Same distribution of electron in their outermost orbitals.
We can classify the all the elements into 4 blocks: s-block, p-block, d-block and f-block. This classification is based on type of atomic orbitals filled with electrons.

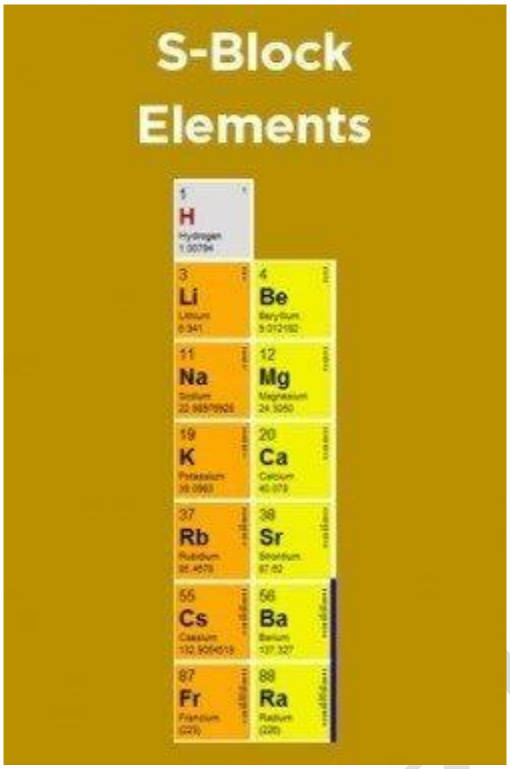
**Define Representative Elements**




The Groups of periodic table can divided into the two categories.

Representative Elements
Transition Metals.
First two and last six groups of periodic table together are representative elements. Transition metals falls in the middle of periodic table.

**Introduction to s-Block Elements**




**s-Block** Elements outer electronic configuration is ns$_1$ns1 and ns$_2$ns2. Group 1 ( Alkali Metals) and Group 2 ( Alkali Earth Metal) belongs to s-block elements. The given figure shows elements present in the s-block of the periodic table.

**Properties of s-Block Elements**

Following are the properties of s-Block elements:

They all are metals
They have low ionization enthalpies
They lose the outermost electron(s) to form cation
Group 1 elements lose 1 electron while the group 2 elements loses 2 electrons

Metallic character increases down the group
Because of high reactivity they are never found pure in nature
Reactivity increases down the group
Compounds of s-Block are ionic in nature
Exception to this is beryllium and lithium

**Properties of p-Block Elements**

Following are the properties of p-block elements.

They are solids/liquids/gases at room temperature (Br is liquid)
They have variable oxidation states
They form acidic oxides
They impart no characteristic colour to the flame
Generally they form covalent compounds
Halogens form salts with alkali metals
They have high ionization potentials
They have very large electron gain enthalpies
The aqueous solutions their oxides are acidic in nature

**Properties of d-Block Elements**

Following are the general properties of d-block elements.

These are metallic in nature
They are hard and have high densities
They have high melting and boiling point
They shows variable oxidation states
They form coloured ions and compounds
The atomic radii decreases with increase in atomic number

**Properties of f-Block Elements**

Following are the general characteristics of f-Block Elements.

They are paramagnetic in nature
Number of radioactive elements is more than the other blocks
They show variable oxidation states
They show shielding effect. Shielding is when an electron becomes less attracted to an atom the further it is away from the nucleus. This is because the forces holding atoms together become weaker as distance increases.

**Properties of individual atoms**

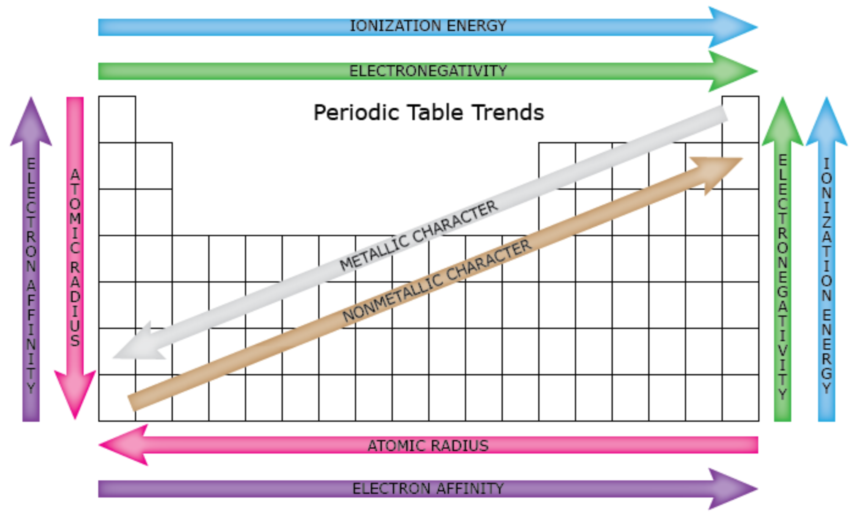
The periodic variation in electron configurations as one moves sequentially through the Periodic Table from H to heavier elements produces a periodic variation in a variety of properties.
The periodic properties are:

Atomic size
Electron affinity
Electronegativity
Ionization enthalpy
Metallic and Non metallic character

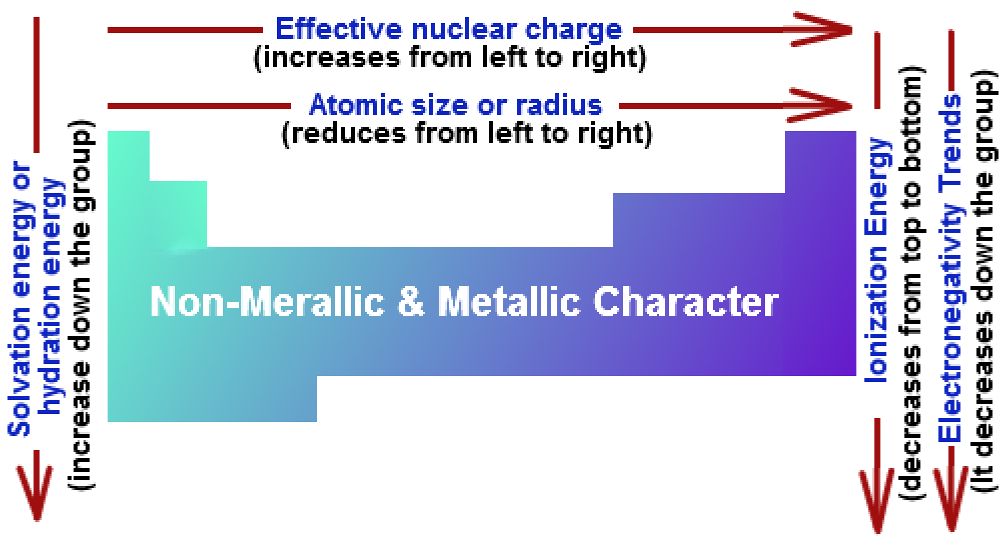
**Properties of groups**

Elements within different groups within the periodic table have different physical and chemical properties. This determines the kinds of reactions these elements have. Different groups also show different trends, in terms of reactivity, as you move down a group. This can also determine how violently a reaction occurs - or whether it happens at all.

**Trends in physical properties**




**Periodic trends in chemical properties**




**Chemical reactivity of the elements**

Reactivity refers to how likely or vigorously an atom is to react with other substances. This is usually determined by how easily electrons can be removed (ionization energy) and how badly they want to take other atom's electrons (electronegativity) because it is the transfer/interaction of electrons that is the basis of chemical reactions.

Metals

Period - reactivity decreases as you go from left to right across a period.

Group - reactivity increases as you go down a group
The farther to the left and down the periodic chart you go, the easier it is for electrons to be given or taken away, resulting in higher reactivity.
Non-metals
Period - reactivity increases as you go from the left to the right across a period. Group - reactivity decreases as you go down the group.
The farther right and up you go on the periodic table, the higher the electronegativity, resulting in a more vigorous exchange of electron

**Explain the periodic trend of other properties**

Across a period, density, boiling and melting point increase gradually. Down the group, density increases gradually. Melting and boiling point of elements decreases gradually. Across a period, basicity of oxides, hydroxides, oxy-acids and hydrides decreases and down the group basicity of oxides, hydrides, hydroxides, oxy-acids increases.

**Group number in a periodic table**

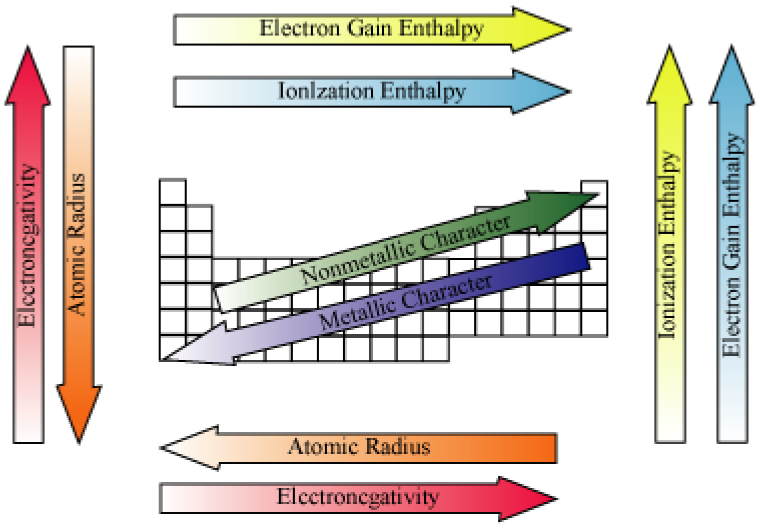
1. In any group, outermost shell electron are known as valance electrons and these electrons are same so main properties of elements of group is similar.
2. Elements are divided into four blocks, which is s, p, d, f according to valance electrons.
3. s-block elements elements of 1 and 2 group.
4. p-block elements elements of 13 to 18 group.
5. d-block elements elements of 3 to 12 group.
6. f-block elements elements of the Lanthanide and Actinide series.
7. Representative Elements elements of s-block and p-block collectively called as Representative elements also known as Normal elements or Typical elements.
8. Transition Elements elements of d-block.
9. Inner Transition Elements elements of f-block, also known as Rare Earth Elements.
10. Alkali Metals elements of 1st group.
11. Alkaline Earth Metas elements of 2nd group.

**Trends in different properties of elements in a periodic table**

In periodic table elements are placed in periods and groups.In periods like atomic number increases but electrons enters in same shell hence atomic size do not increases on moving Left to

Right in periods but in group as atomic number increases number of electrons enters into new shells hence atomic size increases.

**Trends of physical properties of periodic table**



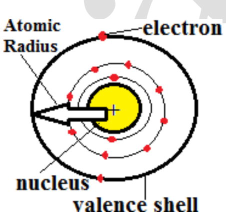

**Reasons for change in trends in periodic properties**

There are change in properties because of the difference in the number of shells and the nuclear charge of various elements.

**Atomic size or atomic radius**

The atomic radius of an element is a measure of the size of its atoms. It represents the mean distance from the nucleus to the outermost boundary of the surrounding cloud of electrons. Atomic radii vary in a predictable manner across the periodic table.

**Define Atomic Size (Atomic Radius)**



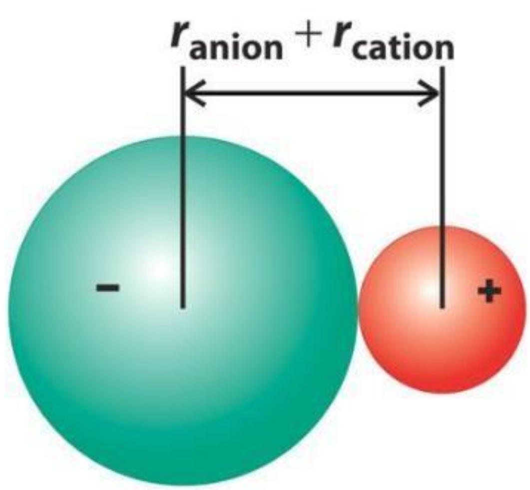
**Atomic Size (Atomic Radius)** is the

distance between the nucleus and valence shell.

**Different Atomic Radius Definitions**

In a practical way, it is difficult to measure the size of an individual atom. But we can estimate the radii of the atom by the bonds it forms. But still, there is no fixed radius of an atom.
We can understand atomic radius by three definitions:
Covalent Radius: If an element is non-metallic
Metallic Radius: If an element is metallic
van der Waals Radius: If an element is a noble gas.

**Define Ionic Radius**



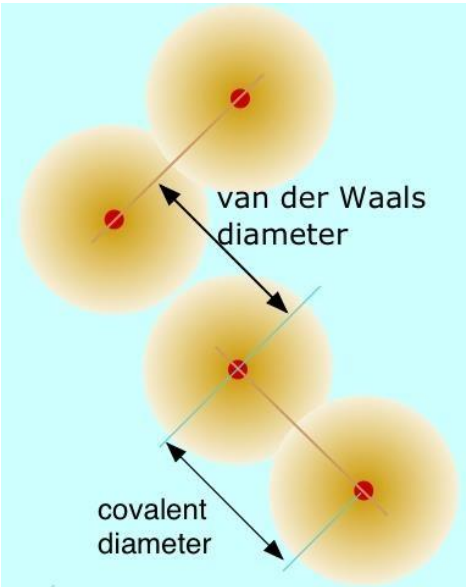

Ionic radius is the radius of an atom's ion in its ionic crystal structure. It is half the distance between two ions that are barely touching each other.
The given diagram shows ionic radii of anion and cation in the molecule.

**Compare Ionic Radius with Atomic Radius**

When electron is removed from an atom cation is formed. The number of electrons in cation is less than its atom. But the number of protons remains same. As there are more protons than electrons, it will pull electrons more closer to balance coulombic force. Hence, the ionic radius of cation will be less than it atomic radius.

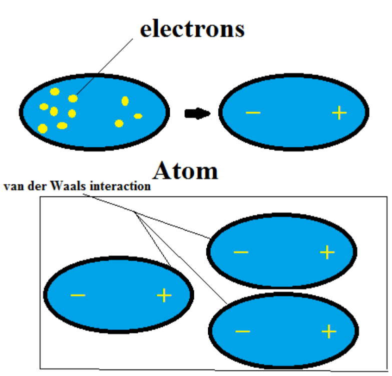
In the case of anion reverse thing will happen. Electrons will go farther away to balance coulombic force. Hence ionic radius of anion will be more than it atomic radius.

**van der Waals Radius**




van der Waals radius is a good approach to estimate the atomic radius of monoatomic gases. It can be also estimated for non-bonded atoms of the same element.
**van der Waals Radius** is half the distance between two non-bonded atoms when the electrostatic forces are balanced. The given diagram shows the difference between covalent and van der Waals radius. Covalent radius is estimated for covalently bonded atoms of the same molecule. Whereas van der Waals radius is estimated for non-bonded atoms of the different molecule.

**Van der Waals Interaction**

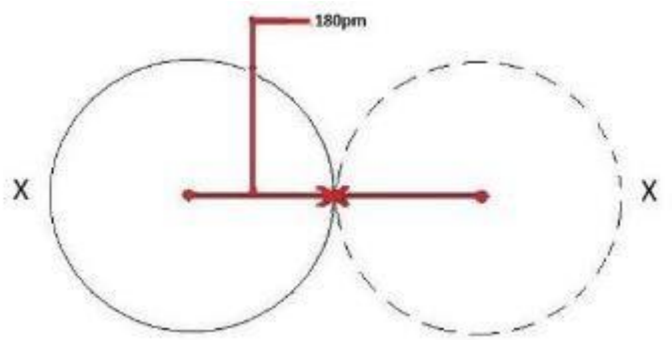


Electrons are constantly moving in an atom. There can be a scenario where there is more number of electrons compared to another. So there is a chance of dipole formation, as seen in the diagram. These dipoles can attract each other cause of electrostatic force. This attractive force is called van der Waals interaction.
*Note: A dipole is a pair of opposite charges separated by small distance.*

**Define Inert Gas Radius**

Inert gases are the most stable element. They don't bond with other elements to gain stability. The only forces that acts on it is van der Waals forces.
Hence, atomic radius is expressed in term of van der Waals radius. This radius is nothing but the **Inert Gas Radius**

**Define Metallic Radius**




The metallic radius is the radius of an atom joined by metallic bond. The metallic radius is half of the distance between the nuclei of two adjacent atoms in the metallic cluster.
Metallic radius for the given diagram is 180 pm.

**Problems Faced to Measure Atomic Radius**

Following are the problems faced to measure atomic radius.
Electron cloud does not have a fixed boundary. Hence, the size of the atom is difficult to find.
None of the atoms (except noble gases) are in the free state. It is difficult to isolate atoms of a molecule as they are unstable. Thus, it is difficult to determine atomic radius.

- The atomic radius changes when the element is in the bonded state.

*Note: Electron cloud is a group of electrons moving around the nucleus*

**78**

**FORMULA**

**Salter rules**

Slater's rules provide numerical values for the effective nuclear charge.Practically electron experience less nuclear pull than the actual one due shielding effect.For each electron in an atom, Slater's rules provide a value for the screening constant, denoted by *s*, *S*, or , which relates the effective and actual nuclear charges is as:

$Z_{eff}=Z-s$.Zeff=Z−s.

**79**

**EXAMPLE**

**Factors which affect atomic size**

Factors affecting atomic size are:

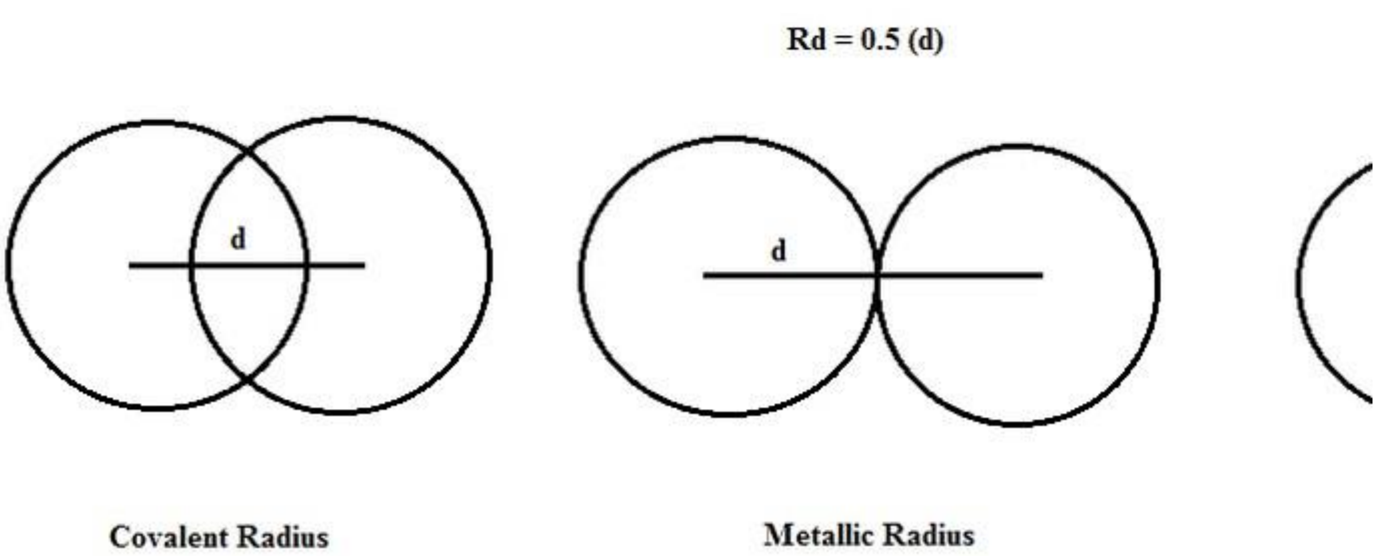
a) Screening effect- caused by mutual repulsion between electrons in the inner shell with those in the outer shell.

b.) Nuclear charge- more the number of protons in the nucleus.more will be the pull on electrons closer to the nucleus which causes the atomic size to decrease

**80**

**DEFINITION**

**Compare Different Radii of an Atom**



Radius of an atom is equal to half of the distance between the nucleus of atoms. Thus, Radius is directly proportional to the distance between the nucleus of atoms. For a covalent bond, the electron cloud of two atoms overlaps with each other. In the case of a metallic bond, the boundaries of atom touch each other. In van der Waals interaction, electron clouds are from each other, such that their boundaries dont intersect. Hence the order of radius is:
van der Waals Radius >> Metallic Radius >> Covalent Radius

**81**

**EXAMPLE**

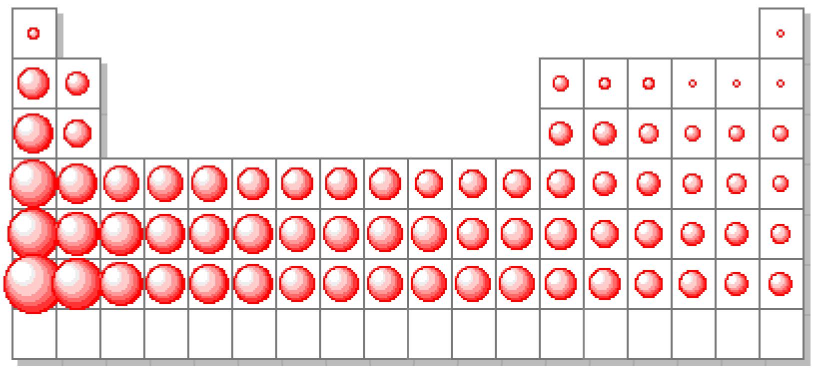
**Trend of atomic size in periodic table**

The atomic radius of an element is a measure of the size of its atoms. It represents the mean distance from the nucleus to the outermost boundary of the surrounding cloud of electrons. Atomic radii vary in a predictable manner across the periodic table.

**82**

**DEFINITION**

**Atomic Size Trend Across the Period**

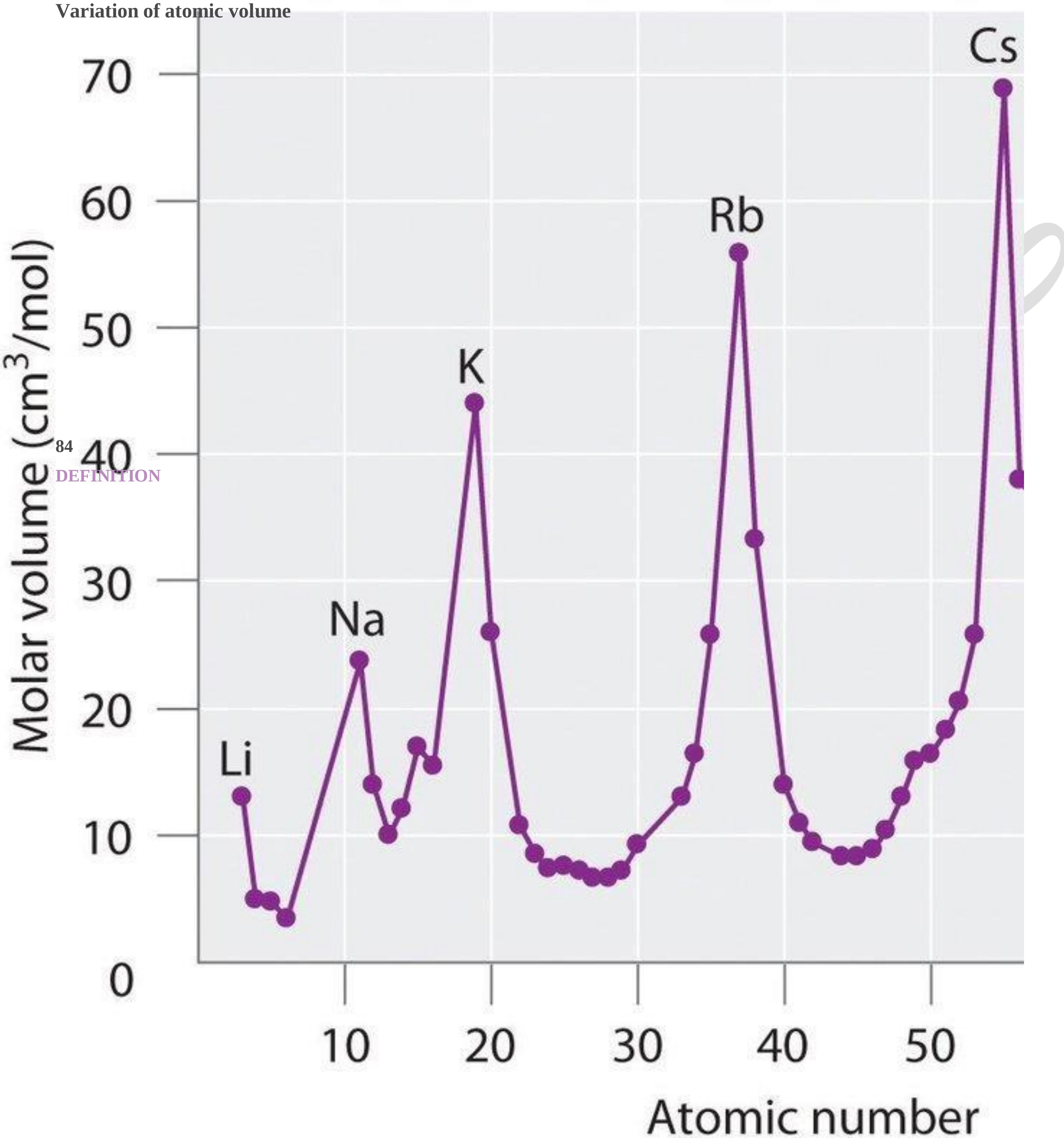


Atomic radius decreases from left to right within a period. This is caused by the increase in the number of protons and electrons across a period. One proton has a greater effect than one electron cause of the coulombic law. Thus, electrons are pulled towards the nucleus, resulting in a smaller radius.

**83**

**DIAGRAM**

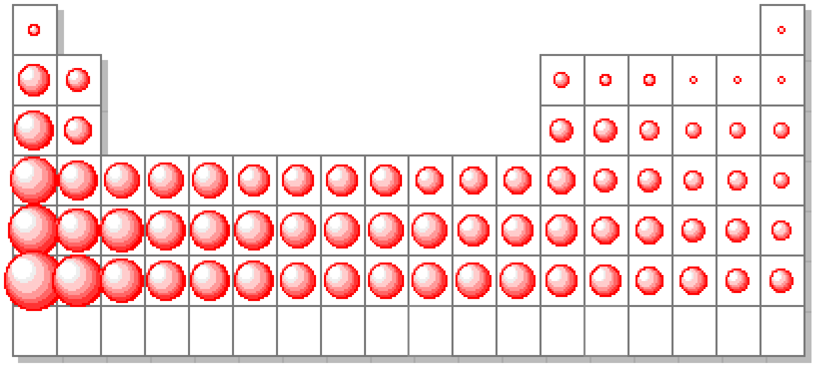
**Variation of atomic volume**






DEFINITION

**Atomic Size Trend Down the Group**



The number of energy levels increases as you move down a group as the number of electrons increases. Each coming energy level is further from the nucleus than the last. Thus, the atomic radius increases as the group and energy levels increase.

**85**

**DEFINITION**

**Define Atomic Volume**

Atomic Volume is the volume occupied by one mole of an element under standard conditions
The atomic volume is a calculated value using the atomic weight and the density.
Formula: Atomic Volume=Atomic weightdensityAtomic Volume=Atomic weightdensity

**86**

**DEFINITION**

**Define Isoelectronic species**

Isoelectronic species are elements or ions that have the equal number of electrons.
Example: $O^{2-}$,$F^{-}$,$Mg^{2+}$O2−,F−,Mg2+ have 10 electrons. Hence, they are isoelectronic species.

**87**

**DEFINITION**

**Atomic size of isoelectronic species**

The ionic radii of isoelectronic ions decrease with the increase in the magnitude of the nuclear charge.

$Al^{3+}$<$Mg^{+}$<$Na^{+}$<$F^{-}$<$O^{2-}$<$N^{3-}$Al3+<Mg+<Na+<F−<O2−<N3−

**88**

**DEFINITION**

**Trend of valency**

In a period, the number of valence electrons increases as we move from left to right side. However, in a group this periodic trend is constant, that is the number of valence electrons remains the same.

**89**

DEFINITION

**Valency across a period**

Valency increases to group 4 and then decreases.

**90**

DEFINITION

**Valency down the group**

The valency down the group remains the same.

**91**

DEFINITION

**Define Ionization Energy.**

Ionization energy is the energy required to remove valence electron from an atom.
Example: Energy required to remove electron from NaNa atom to form $Na^+$Na+ atom is aa kJ. aa kJ is an ionzation energy.

**92**

DEFINITION

**Successive ionization enthalpy**

Ionization energy is the energy required to remove an electron from the outermost shell of an isolated gaseous atom. When the first electron or the most loosely bound electron is removed, the amount of energy required is less than the energy required to remove the electron in the next successive shell. This ionization energy goes on increasing with the number of electrons removed. So the number of electrons removed from the successive no of shells and the energy involved is called successive ionization energy.

**93**

DEFINITION

**ionization enthalpy (ionization potential)**

The ionization energy (IE) is qualitatively defined as the amount of energy required to remove the most loosely bound electron, the valence electron, of an isolated gaseous atom to form a cation.

**94**

DEFINITION

**Shielding effect**

Shielding effect can be defined as a reduction in the effective nuclear charge on the electron cloud, due to a difference in the attraction forces of the electrons on the nucleus. It is also referred to as the screening effect or atomic shielding.

**95**

DEFINITION

**Ionisation energy across the period**

Ionization energy increases across a period due to increase in the nuclear charge.
The following depicts the increase of IE in period.

**96**

**DEFINITION**

**Trend in Ionization energy**

Ionization energies are dependent upon the atomic radius. Since going from right to left on the periodic table, the atomic radius increases, and the ionization energy increases from left to right in the periods and up the groups.
As we move down the group, size increases so the outermost electrons are very far away from the nucleus so the electrons are loosely bounded by the nucleus so it is easy to remove it. Moving across a period, atomic size decreases, so the outermost electrons are nearer to the nucleus. So more force of attraction holds the electrons so more energy is needed to remove the electrons.

**97**

**EXAMPLE**

**Trend in ionization energy down the group**

Ionization energy decreases down the group due increase in the atomic size (addition of new shell).

**98**

**EXAMPLE**

**Electron affinity**

The energy released when an electron is added to a neutral gaseous atom is known as electron affinity. The unit for electron affinity is kilo joules per mole.
It depends mainly on two factors. They are atomic size and nuclear charge.

**99**

**DEFINITION**

**Successive electron gain enthalpies**

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 second electron.

**100**

**EXAMPLE**

**Electron gain enthalpy**

The electron affinity increases across a period while it decreases down a group.
The zero group elements have the lowest electron affinity values. Halogens posses highest

electron affinity in the periodic table. In halogens chlorine posses highest electron affinity in the periodic table.

**101**

**EXAMPLE**

**Trend in electron gain enthalpy**

The electron affinity increases across a period while it decreases down a group.

The zero group elements have the lowest electron affinity values.

**102**

**DEFINITION**

**Electron Affinity**

The amount of energy released when an atom in the gaseous state accepts an electron to form an anion. Factors which affect electron affinity are: Atomic size and Nuclear charge. As atomic radii increases, electron affinity increases. As nuclear charge increases, electron affinity increases. It decreases down a group and increases across a period.

**103**

**DEFINITION**

**Metallic and nonmetallic character**

The tendency of an atom to lose electrons is known as metallic character.

The tendency of an atom to gain electrons is known as non-metallic character.

**104**

**EXAMPLE**

**Factors affecting metallic or nonmetallic character**

Metals are highly electropositive in nature and non-metals are more electronegative in nature.As the electro positively increases metallic character increases and as the electronegativity increases nonmetallic character increases.

**105**

**EXAMPLE**

**Trend of metallic and non-metallic properties in modern periodic table**

In modern periodic table metallic character increases down the group and decreases from left to right across the period.

**106**

**DEFINITION**

**Metallic Character Trend Across the Period**

Metallic character decreases as you move across a period from left to right. Metallic character depends on how quick electron can lose from an atom. As the tendency to lose electrons decreases across the period, metallic character decreases.

**107**

DEFINITION

**Metallic and Non-metallic properties**

The tendency of an element to lose electrons and form positive ions (cations) is called electropositive or metallic character. The tendency of an element to accept electrons to form an anion is called its non-metallic or electronegative character. In each period, metallic character of elements decreases as we move to the right. Elements to the left of the periodic table have a pronounced metallic character while those to the right have a non-metallic character. Conversely, non-metallic character increases from left to right.

As we move down the group the number of shells increases. This causes the effective nuclear charge to decrease due to the outer shells being further away: in effect the atomic size increases. The electrons of the outermost shell experience less nuclear attraction and so can lose electrons easily thus showing increased metallic character.

**108**

DEFINITION

**Electronegativity**

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons.

**109**

EXAMPLE

**Trends in electronegativity across a period from left to right**

Trend-wise, as one moves from left to right across a period in the periodic table, the electronegativity increases.

**110**

DEFINITION

**Electronegativity**

The tendency of an atom to attract electrons to itself when combined in a compound is known as electronegativity. The factors which affect the electronegativity are atomic size and nuclear charge. As atomic radii increases, electronegativity decreases. As nuclear charge increases, electronegativity increases. Electronegativity decreases down a group and increases across a period.

**111**

EXAMPLE

**Factors which affect electronegativity**

## Factors Affecting Electronegativity

- Nuclear charge. The higher the nuclear charge, more will be the electronegativity value of an element, since the nucleus will be able to attract or pull more electrons towards itself.

- Atomic size
- Screening effect or shielding effect.

**112**

**DEFINITION**

**Applications of electronegativity**

The concept of electronegativity is very useful in predicting metallic, non-metallic characters of elements and polarity of bonds.

**113**

**DEFINITION**

**Electrongain enthalpy and electronegativity**

Electronegativity of an element may be defined as the tendency of its atom to attract the shared pair of electrons towards itself in a covalent bond. whereas electrongain enthalpy is the energy released when electrons are added to a neutral gaseous atom to form a gaseous anion.

**114**

**DEFINITION**

**Pauling scale**

Linus Pauling was the original scientist to describe the phenomena of electronegativity. The best way to describe his method is to look at a hypothetical molecule that we will call XY. By comparing the measured X-Y bond energy with the theoretical X-Y bond energy (computed as the average of the X-X bond energy and the Y-Y bond energy), we can describe the relative affinities of these two atoms with respect to each other.

ΔΔ Bond Energies = (X-Y)measured - (X-Y)expected

if the electonegativities of X and y are the same, then we would expect the measured bond energy to equal the theoretical (expected) bond energy and therefore the bond energies would be zero. If the electronegativities of these atoms are not the same, we would see a polar molecule where one atom would start to pull electron density toward itself, causing it to become partially negative.By doing some careful experiments and calculations, Pauling came up with a slightly more sophisticated equation for the relative electronegativities of two atoms in a molecule:

EN(X) - EN(Y) = 0.102 (ΔΔ1/2).1

In that equation, the factor 0.102 is simply a conversion factor between kJ and eV to keep the units consistent with bond energies.

**115**

**DEFINITION**

**Mullikan scale**

### Mulliken's Scale

In this scale, electronegativity is taken as the average of ionization potential and electron affinity of the element in eV.

$$\text{Electronegativity, } \chi_A = \frac{IE_A + EA_A}{2}$$

Pauling and Mulliken values of electronegativity are related as

$$\chi_{Pauling} = 0.336\,[\chi_{Mulliken} - 0.615]$$

**116**

**DEFINITION**

**Periodicity of valence or oxidation state within a group**

When we move down the group, the number of valence electrons remain same therefore all the electrons have the same valence.

**117**

**DEFINITION**

**Periodicity of oxidation state along a period**

As we move across a period from left to right, the number of valence electrons increases from 1 to 8. But the valence of elements w.r.t H or O first increases from 1 to 4 and then decreases to zero.

**118**

**DEFINITION**

**Anomalous Behaviour in Periodic Trend**

According to Hund's rule, it is difficult to remove electrons from fully or half-filled atoms. They are very stable and thus requires more energy to remove an electron. We know that ionization energy (IP) increase across the period. But when it comes across fully or half-filled atoms, IP changes drastically. Thus, it shows an anomalous behaviour.

**119**

**DEFINITION**

**Variation of melting and boiling point across a period**

Melting and boiling point across period increases because in a period electron enters in same shell hence the shell become heavy.

**120**

**DEFINITION**

**Variation of melting and boiling point down a group**

Down the group the melting and boiling point decreases because electrons enters into new groups hence accumulation in shell become less and trend decrease.

# Concepts

## Points to Remember

1 DEFINITION

**Significance of classification**

Classification of elements is important because it is easy for one to know that this element has certain properties and is same for all the elements in the same period. It makes the study easy and helps in discovering new elements.

2

DEFINITION

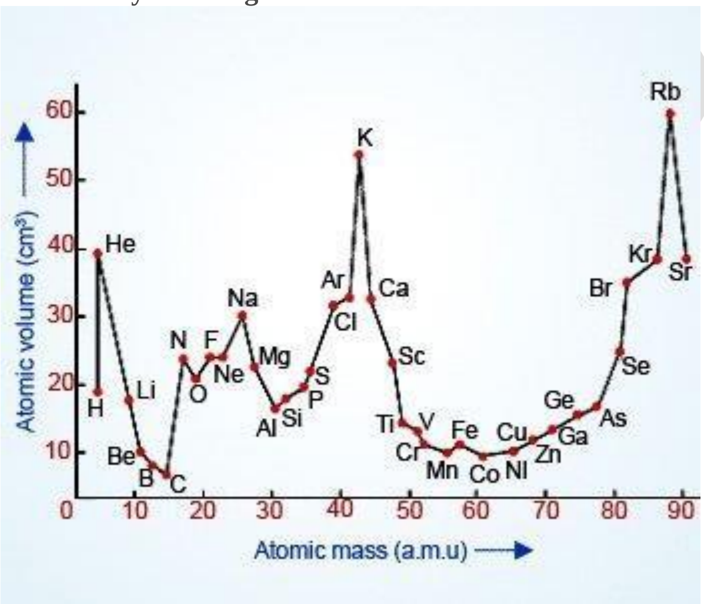
**Periodic table**

A table of the chemical elements arranged in order of atomic number, usually in rows, so that elements with similar atomic structure (and hence similar chemical properties) appear in vertical columns.

3

DEFINITION

**Lothar Meyer arrangment**




Meyer considered the volume taken up by fixed weights of the various elements. Each weight contained the same number of atoms of its particular element (Avogadro's number). This meant that the ratio of the volumes of the various elements was equal to the ratio of the volumes of single atoms of the various elements. Thus Lothar Meyer could determine the atomic volumes of elements. If the atomic volumes of the elements were plotted against the atomic weight, a series of peaks were produced. The peaks had alkali metals: sodium, potassium, rubidium, and cesium.

Each fall and rise to a peak, corresponded to a period like the waves. In each period a number of physical properties other than atomic volume also fell and rose, such as valence and melting point. The second and third period in Meyer's table included seven elements each, and duplicated Newlands' law of octaves.

4 DEFINITION

**Dobereiner's Triad**

Dobereiner arranged the element in increasing order of atomic masses. He found that the atomic mass of the middle element was approximately equal to the arithmetic mean (average) of the atomic masses of the other two elements of that triad when they are arranged in their increasing order of atomic mass,

e.g., Li, Na, K. The atomic mass of Li is 9 and K is 39. The average of two atomic number is 23 that is the atomic number of sodium.

5

DEFINITION

**Newland's octaves**

| 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 La | Zr | — | — |

According to Newland's law of octaves, when elements are arranged by increasing atomic mass, the properties of every eighth element starting from any element are a repetition of the properties of the starting element.

6

DEFINITION

**Telluric screw**

Alexander-Emile, a French geologist, was the first person to notice the periodicity of the elements, similar elements occurring at regular intervals when they are ordered by their atomic weights. In 1862 he gave an early form of periodic table, which he named the 'telluric helix', after the element Tellurium, which fell near the center of the diagram. With the elements arranged in a spiral on a cylinder by order of increasing atomic weight, Chancourtois saw that elements with similar properties lined up vertically.

Chancourtois plotted the atomic weights on the surface of a cylinder with a circumference of 16 units, the approximate atomic weight of oxygen. The resulting helical curve, which Chancourtois called a square circle triangle, brought similar elements onto corresponding points above or below one another on the cylinder. He was the first scientist to see the periodicity of elements when they were arranged in order of their atomic weights. He found out that the similar elements occurred at regular atomic weight intervals.

**7**

**DEFINITION**

**Mendeleev's periodic law**

Mendeleev stated that the periodic properties of elements are the periodic function of atomic mass.

**8**

**DEFINITION**

**Mendeleev's Periodic Table**

| Group | I | II | III | IV | V | VI | VII | VIII |
|---|---|---|---|---|---|---|---|---|
| Oxide : Hydride: | $R_2O$ RH | RO $RH_4$ | $R_2O_3$ $RH_4$ | $RO_2$ $RH_4$ | $R_2O_5$ $RH_3$ | $RO_3$ $RH_2$ | $R_2O_7$ 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 22.99 | Al 24.31 | Si 28.09 | P 30.974 | S 32.06 | Cl 35.453 | |
| 4 First series Second series | K 39.102 Cu 63.54 | Ca 40.08 Zn 65.54 | Sc 44.96 Ga 69.72 | Ti 47.90 Ge 72.59 | V 50.94 As 74.92 | Cr 50.20 Se 78.96 | Mn 54.94 Br 79.909 | Fe Co Ni 55.85 58.93 58.71 |
| 5 First series Second series | Rb 85.47 Ag 107.87 | Sr 87.62 Cd 112.40 | Y 88.91 In 114.82 | Zr 91.22 Sn 118.69 | Nb 92.91 Sb 121.60 | Mo 95.94 Te 127.60 | Tc 99 I 126.90 | Ru Rh Pd 101.07 102.91 106.4 |
| 6 First series Second series | Cs 132.90 Au 196.97 | Ba 137.34 Hg 200.59 | La 138.91 Tl 204.37 | Hf 178.40 Pb 207.19 | Ta 180.95 Bi 208.98 | W 183.85 | | Ru Rh Pd 190.2 192.2 195.09 |

Mendeleev's Periodic law: Physical ,and chemical properties of elements are a periodic function of their atomic masses. Periodic table is a chart of elements prepared in such a way that elements with similar properties occur in the same vertical column. It has seven horizontal rows and eight vertical columns.

9 DEFINITION

**Characteristics of mandeleev's periodic table**

1. Elements were arranged in order of their increasing atomic weights in horizontal rows called periods.
2. Elements were arranged in vertical columns called groups according to their resemblance in properties.
3. The periodic table had 8 groups and 7 periods.
4. Since only 63 elements were known at that time, he had left gaps for undiscovered elements.

10 DEFINITION

**Achievements of Mendeleev's Periodic Table**

1. There was grouping of elements according to the chemical properties.
2. The periodic table has gaps for the undiscovered elements.
3. The prediction of properties of undiscovered elements were correct.
4. It helped in systematic study of elements

11 DEFINITION

**Limitation of Mendeleev's Periodic Table**

1. There were anamalous pairs of elements.
2. Position of hydrogen was not fixed.
3. There were grouping of chemically dissimilar elements.
4. There were separation of chemically similar elements.

12 EXAMPLE

**Limitation of Mendeleev's classification**

Mendeleev's periodic table suffered few defects as follows:

The position of hydrogen was not correctly defined.

In some cases, Mendeleev placed elements according to their similarities in properties and not in increasing order of their atomic masses. Thus, the position of these elements was not justified

Isotopes were not given separate places in the periodic table although Mendeleev's classification is based on the atomic masses.

Some similar elements were grouped separately while some dissimilar elements were grouped together.
Mendeleev could not explain the cause of periodicity in the elements.
The position for lanthanides and actinides were not included in this table.

**13 DEFINITION**

**Moseley's Periodic table**
Moseley found and measured a property linked to Periodic Table position. Hence atomic number became more meaningful and the three pairs of elements that seemed to be in the wrong order could be explained. Moseley used what was then brand-new technology in his experiments. A device now called an electron gun had just been developed. He used this to fire a stream of electrons (like machine gun bullets) at samples of different elements. He found that the elements gave off X-rays. (This is how the X-rays used in hospitals are produced.)
Moseley measured the frequency of the X-rays given off by different elements. Each element gave a different frequency and he found that this frequency was mathematically related to the position of the element in the Periodic Table he could actually measure atomic number. Moseley plotted the *square root* of the X-ray frequency against atomic number.

**14 DEFINITION**

**Justification of Modern periodic law**
When elements are arranged in order of their atomic number, there is a periodicity in properties of elements. The anomalies seen in Mendeleev's periodic table was removed. The isotopes were given positions. The atomic weights are arranged in order.

**15**
**DEFINITION**

**Cause of periodicity**
The modern periodic table is based on the electronic configuration of the elements. The properties of an element are determined largely by the electrons in its outermost or valence shell. Valence electrons interact with other atoms and take part in all chemical reactions, while inner shell electrons have little influence on the properties of elements. When elements are placed in the order of their increasing atomic number, the elements having the same number of valence shell electrons is repeated in such a way, so as to fall under the same group. Since, the electronic configuration of the valence shell electrons is same they show similar properties.

**16** LAW

**Modern periodic law**

The law that the properties of the elements are periodic functions of their atomic numbers.

**17** DEFINITION

**Long form of periodic table**

The periodic table is a tabular arrangement of the chemical elements, ordered by their atomic number , electron configuration, and recurring chemical properties. This ordering shows periodic trends such as elements with similar behavior in the same column. It also shows four rectangular blocks with some approximately similar chemical properties. In general, within one row (period) the elements are metals on the lefthand side, and non-metals on the righthand side.
The rows of the table are called periods, the columns are called groups. Six groups (columns) have names as well as numbers: for example, group 17 elements are the halogens and group 18, the noble gases. The periodic table can be used to derive relationships between the properties of the elements, and predict the properties of new elements yet to be discovered or synthesized.

**19** LAW

**Modern Periodic Law**

The properties of elements are the periodic function of their atomic number i.e number of protons.

**20** DEFINITION

**Features of long form of periodic table**

1. 18 vertical columns known as groups.
2. 7 Horizontal rows known as periods.
3. Light metals These are elements of periodic table of group 1 and 2.
4. Heavy metals or Transition metals - These are elements of periodic table of group 3, 4, 5, 6, 7, 8, 9, 10, 11 and 12.
5. Non-Metals These are elements of periodic table of group 13, 14, 15, 16 and 17.
6. Zero group These are elements of periodic table of group 18.

21 **DEFINITION**

**Merits of Modern Periodic Table over Mendeleev's periodic table**

The advantages of Modern periodic table over Mendeleev's periodic table are : Modern periodic table is based on the most fundamental property, the atomic number of elements, while Mendeleev's periodic table is based upon the atomic masses of elements. In the modern periodic table, elements are arranged in accordance with their electronic configurations. The elements having similar electronic configurations are placed in the same group. Hence, elements in given group show similar properties. Elements with different electronic configurations are grouped separately, hence they show different properties. Mendeleev's periodic table does not provide any reason for the similarity and the difference in properties of elements. In Mendeleev's periodic table there are several anomalies, e.g. the position of isotopes, wrong order of atomic masses of some elements etc. In the long form of the periodic table, these anomalies have been removed. In the long form of the periodic table, elements have been clearly separated as normal elements, transition elements and noble gases. Metals and non-metals are also separated. But in Mendeleev's periodic table there is no such separation of different types of elements.

22

**DEFINITION**

**Defects of Modern Periodic table**

- Position of hydrogen still dicey. It is not fixed till now.
- Position of lanthanides and actinides has not been given inside the main body of periodic table.

- It does not reflect the exact distribution of electrons of some of transition and inner transition elements.

**23**

**DEFINITION**

**Group in a periodic table**

A group (also known as a family) is a column of elements in the periodic table of the chemical elements. There are 18 numbered groups in the periodic table, but the f-block columns (between groups 2 and 3) are not numbered.

**24**

**DEFINITION**

**Period in a periodic table**

In the periodic table of the elements, each numbered row is a period. In the periodic table of the elements, elements are arranged in a series of rows (or periods) so that those with similar properties appear in a column.

**25**

**SHORTCUT**

**Determine the period number of elements**

All of the elements in a period have the samenumber of atomic orbitals. For example, every element in the top row (the first period) has one orbital for its electrons. All of the elements in the second row (the second period) have two orbitals for their electrons.

**26**

**SHORTCUT**

**Determine the group number of elements**

The group number is an identifier used to describe the column of the standard periodic table in which the element appears. Groups 1-2 (except hydrogen) and 13-18 are termed main group elements. Groups 3-11 are termed transition elements.

Electronic configuration of Sodium is 2, 8, 1. So the number of group is 1