

Revision Notes

Class 11 Chemistry

Chapter 3 – Classification of Elements and Periodicity in Properties

Introduction

A periodic table is a table that organises all known elements into groups based on their properties, with like elements grouped together in the same vertical column and dissimilar elements separated.

Historical Development of Periodic Table

The atomic weights of the elements were used in all previous attempts to classify them.

Dobereiner's Triads

Dobereiner defined triads as groups of three elements in 1829. A triad's chemical characteristics were comparable. The atomic weight of the middle element was found to be approximately equal to the arithmetic mean of the other two elements when the elements in a trio were ordered in order of increasing atomic weights.

Triad	Iron	Cobalt	Nickel	Mean of 1 st and 3 rd
Atomic weight	55.85	58.93	58.71	Atomic weights are nearly the same.
Triad	Lithium	Sodium	Potassium	$\frac{7+39}{2} = 23$
Atomic Weight	7	23	39	
Triad	Chlorine	Bromine	Iodine	$\frac{35.5+127}{2} = 81.25$

Atomic weight	35.5	80	127	
Triad	Calcium	Strontium	Barium	$\frac{40+177}{2} = 88.5$
Atomic weight	40	87.5	137	

Newland's Law of Octaves

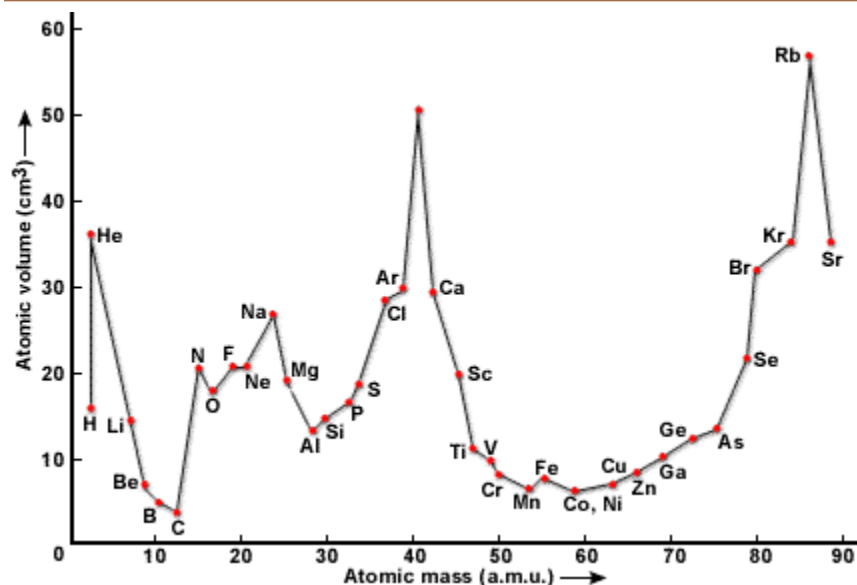
When the lighter elements were placed in order of increasing atomic weights, the properties of every eighth element were comparable to those of the first, as the eighth note of a musical scale, according to an English chemist named John Alexander Newlands in 1865. **Newlands' law of octaves** was named after this generalisation.

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	—	—

Lothar Meyer's Curve

“The atomic masses of elements are periodic functions of their physical properties.”

According to Lothar Meyer, atoms with similar properties occupy similar places in the atomic volume, as measured by the atomic mass curve.



Mendeleev's Periodic Law

Mendeleev organised atoms in a table's horizontal rows and vertical columns in order of increasing atomic weights, so that elements with comparable properties were grouped together in the same vertical column.

Modern Periodic Law

Henry Moseley, an English physicist, discovered regularities in the distinctive X-ray spectra of the elements in 1913. Instead of a straight line between $\sqrt{\nu}$ versus atomic mass, a plot of $\sqrt{\nu}$ (where ν is the frequency of X-rays emitted) against the atomic number (Z) produced a straight line.

As a result, Mendeleev's Periodic Law was altered correspondingly. The Modern Periodic Law can be stated as follows: **The physical and chemical properties of elements are periodic functions of their atomic numbers.**

PERIODIC TABLE OF ELEMENTS

5	Pent	p
6	Hex	h
7	Sept	s
8	Oct	o
9	enn	e

Table: Nomenclature of Elements with Atomic Number above 100

Atomic Number	Name	Symbol
101	Unnilunium	Unu
102	Unnibium	Unb
103	Unniltrium	Unt
104	Unnilquadium	Unq
105	Unnilpentium	Unp
106	Unnilhexium	Unh
107	Unnilseptium	Uns
108	Unniloctium	Uno

109	Unnilennium	Uue
110	Ununillium	Uun

Prediction of Block, Period, and Group

1. What electronic configuration
2. Block - last electron enters into which orbital
3. Period - Maximum value of principal quantum number

Group - *s* - block - number of valence electron

p - block - 10 + number of valence electron

d - block - $ns + \text{number of } (n-1)d \text{ electrons}$

f - block - III B

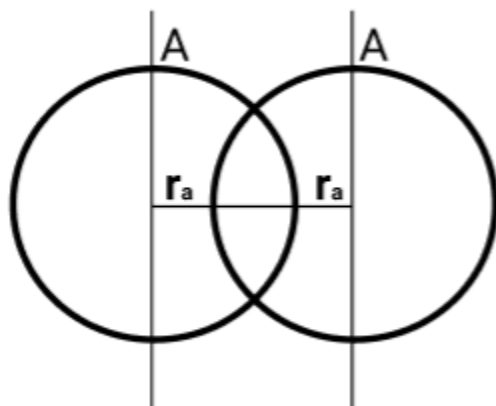
Properties of an Element

Atomic Radius

We can't measure an isolated atom's exact size since its outermost electron has a slim possibility of being located far away from the nucleus. So, depending on the atoms' surroundings, multiple forms of atomic radius can be utilised, such as covalent radius, van der Waals' radius, and metallic radius.

Covalent Radius

The covalent radius of a molecule is half the distance between the nuclei of two identical atoms linked by a single covalent bond.



So, covalent radius for A – A

$$r_A = \frac{d_{A-A}}{2}$$

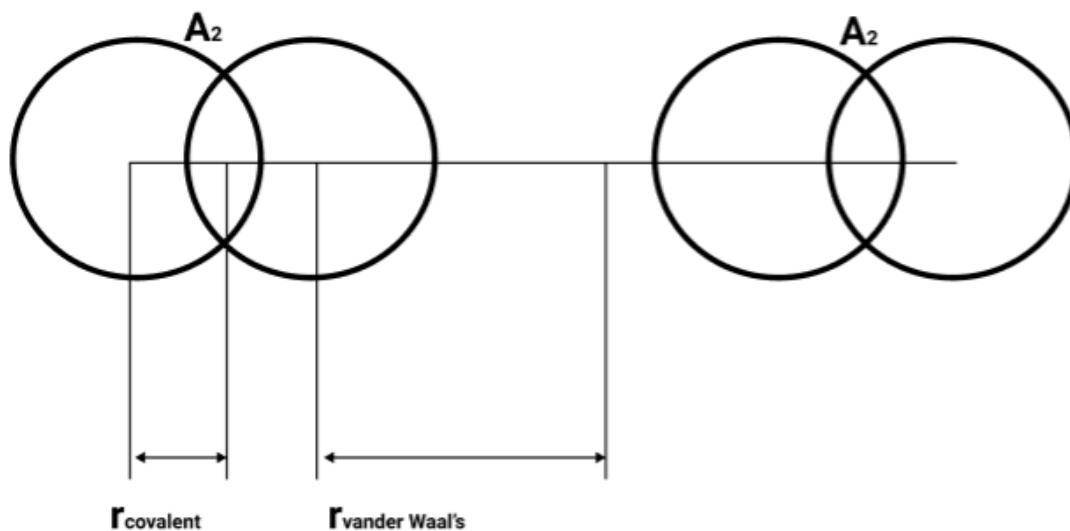
If covalent bond is formed between two different elements then

$$d_{A-B} = r_A + r_B - 0.09(\chi_A - \chi_B)$$

Where, χ_A and χ_B are electronegativities of A and B.

Vander Waal's Radius

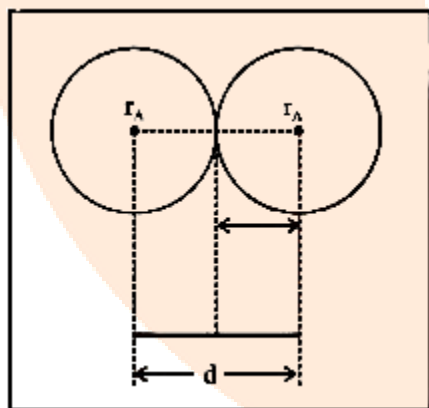
In the solid state, it is half of the internuclear distance between adjacent atoms of two neighbouring molecules.



$$r_{\text{vander waal}} = \frac{d_{A-A}}{2}$$

Metallic Radius (Crystal Radius)

In the metallic crystal lattice, it is half the distance between the nuclei of two neighbouring metal atoms.



So, metallic radius for A – A :

$$d = r_A + r_A$$

$$r_A = \frac{d}{2}$$

Therefore: $r_{\text{covalent}} < r_{\text{metallic}} < r_{\text{vander waal}}$

Variation of Atomic Radii in the Periodic Table

Atomic Radii

As we move from left to right in a period, the covalent and van der Waals radii generally decrease with increasing atomic number.

(a). Variation along a period: Because the outer electrons are in the same valence shell throughout the period, the effective nuclear charge increases as the atomic number grows, increasing the attraction of electrons to the nucleus.

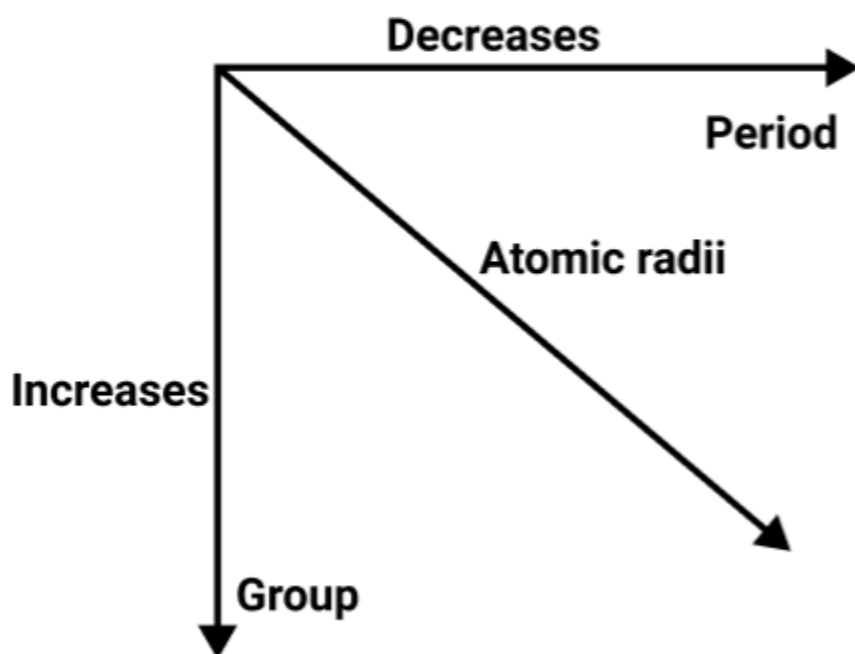
(b). Variation along with a group: The atomic radius of a group grows as the number of atoms in the group grows. The reason for this is that as the primary quantum number (n) increases inside the group, the valence electrons become further away from the nucleus.

(c). Ionic Radius: A cation is formed when an electron is removed from an atom, and an anion is formed when an electron is gained. In general, element ionic radii follow the same pattern as atomic radii. Because it has fewer electrons than its parent atom, a cation is smaller, but its nuclear charge remains the same. Because the addition of one or more electrons causes enhanced electron repulsion and a decrease in effective nuclear charge, the size of an anion will be bigger than that of the parent atom. The ionic radius of the fluoride ion (F^-), for example, is 136 pm, while the atomic radius of fluorine is just 64 pm. The atomic radius of sodium, on the other hand, is 186 pm, whereas the ionic radius of Na^+ is 95 pm.

(d). Isoelectronic Species: Isoelectronic species have the same number of electrons as each other. O^{2-} , F^- , Na^+ , and Mg^{2+} , for example, all have the same number of electrons (10). Because of their various nuclear charges, their radii would be varied. Because the electrons are more attracted to the nucleus, the cation with the higher positive charge will have a smaller radius. The radius of an anion with a higher negative charge will be larger. The net repulsion of the electrons will outweigh the nuclear charge in this situation, causing the ion to expand in size.

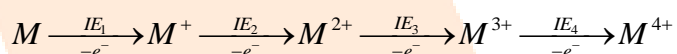
Order of Atomic radii is: $Mg^{2+} < Na^+ < F^- < O^{2-}$

General Trend



Ionization Energy

Ionization energy is the least amount of energy required to remove an electron from the outermost orbit of an isolated atom in the gaseous state.



IE_1 , IE_2 , IE_3 and IE_4 are successive ionization energies.

$$IE_4 > IE_3 > IE_2 > IE_1 \text{ or } \Delta_i H_4 > \Delta_i H_3 > \Delta_i H_2 > \Delta_i H_1$$

Variation of Ionisation Energy in Periodic Table

a. Variation along a period: The value of ionisation enthalpy grows from left to right over time, with intervals where the atoms' configurations are somewhat stable. Increased nuclear charge and decreased atomic radii are simple explanations for the observed trends. Both factors enhance the force of attraction towards the nucleus, requiring more and more energy to remove the electrons, resulting in higher ionisation enthalpies.

b. **Variation along a period:** The atomic size grows steadily as the group progresses due to the addition of one additional primary energy shell at each consecutive element. As a result, the ionisation enthalpy value drops as the force of attraction towards the valence electrons decreases.

Units of I.E/I.P

It is expressed in electron volts (eV) per atom, kilocalories per mole ($kcal\ mol^{-1}$), or kilojoules per mole ($kJ\ mol^{-1}$) units. The energy gained by an electron while travelling under a potential difference of one volt is known as an electron volt.

1 electron volt per atom:

$$= 3.83 \times 10^{-20}\ cal\ per\ atom$$

$$= 1.602 \times 10^{-19}\ J\ per\ atom$$

$$= 3.83 \times 10^{-20} \times 6.023 \times 10^{23}\ cal\ mol^{-1}$$

$$= 23.06\ kcal\ mol^{-1}$$

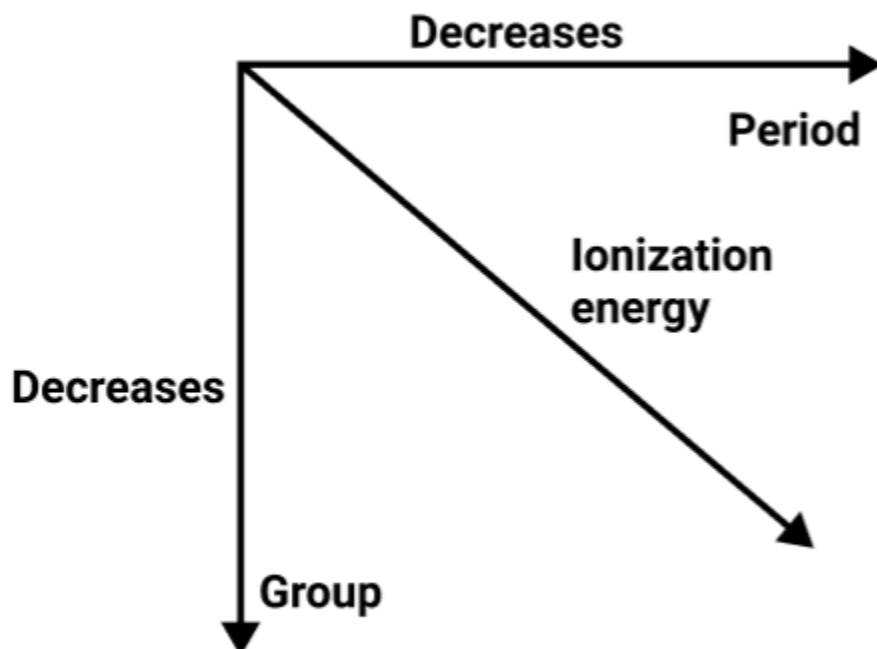
$$= 1.602 \times 10^{-19} \times 6.023 \times 10^{23}\ J\ mol^{-1}$$

$$= 96.49\ kJ\ mol^{-1}$$

$$1\ electron\ volt\ per\ atom = 23.06\ kcal\ mol^{-1} = 96.49\ kJ\ mol^{-1}$$

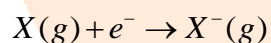
Important Points

- The ionisation energy of an atom or an ion increases as the size of the atom or ion decreases.
- As the screening effect decreases, the ionisation energy increases.
- With the increased nuclear charge, ionisation energy rises.
- If an atom's orbitals are half filled or entirely filled, the ionisation energy rises.
- The penetrating power of orbitals is in the order: $s > p > d > f$



Electron Gain Enthalpy

The Electron Gain Enthalpy is the enthalpy change that occurs when an electron is added to a neutral gaseous atom (X) to turn it into a negative ion. The ease with which an atom adds an electron to create an anion is measured by the electron gain enthalpy.

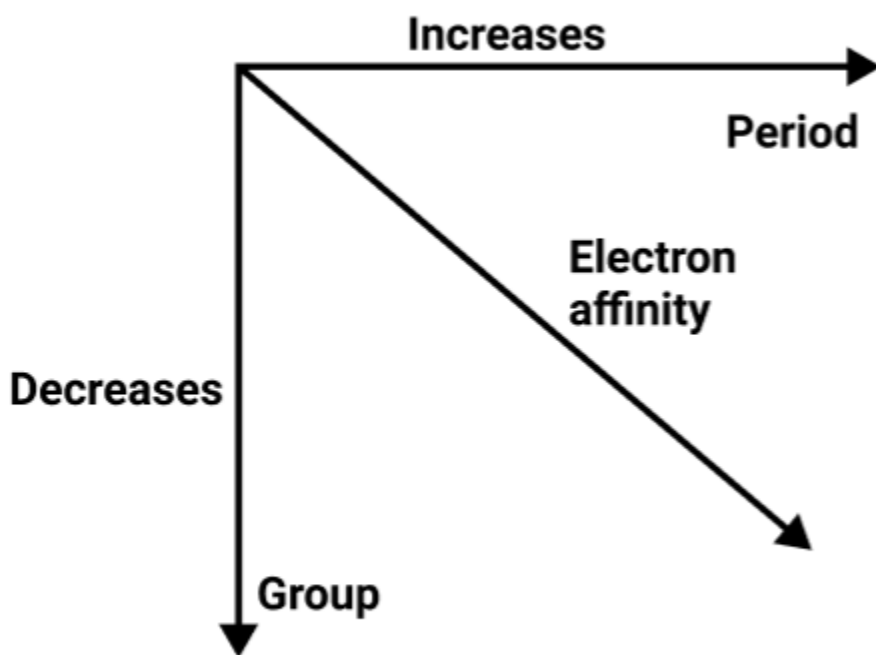


The process of adding an electron to an atom can be either endothermic or exothermic, depending on the element. When an electron is added to an atom, energy is dissipated, and the electron gain enthalpy is negative for many elements. Because they can achieve stable noble gas electrical configurations by taking up an electron, group 17 elements (the halogens) have exceptionally high negative electron gain enthalpies. Noble gases, on the other hand, have substantial positive electron gain enthalpies because the electron must enter the next higher main quantum level, resulting in an electronic configuration that is exceedingly unstable.

Variation of Electron Gain Enthalpy

a. **Variation along a period:** In a period, the electron gain enthalpy shifts from left to right, becoming increasingly negative. As the atomic number increases over time, the size of the nucleus shrinks and the nuclear charge increases. Due to the larger force of attraction by the nucleus for the incoming electron, both of these considerations favour the addition of an extra electron.

b. **Variation along with a group:** Going down from top to bottom in a group, the electron gain enthalpies become less negative. This is owing to the fact that as you move down a group, your size grows. In comparison to the other factor, which is an increase in nuclear charge, this factor is more important.



Electronegativity

Electronegativity refers to an atom's inclination to draw the shared pair of electrons towards itself. The electronegativity of fluorine is 4.0, according to Pauling, and the electronegativity of other elements can be derived as:

$$(\chi_A - \chi_B) = 0.208[E_{A-B} - (E_{A-A} \times E_{B-B})^{1/2}]^{1/2}$$

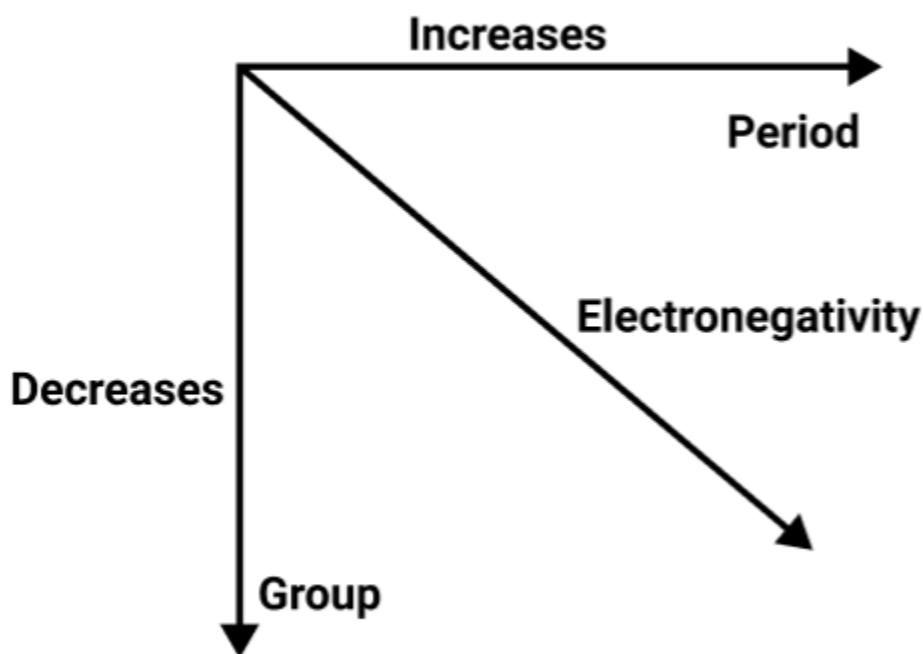
According to **Mulliken**:

$$\text{Electronegativity} = \frac{\text{IP} + \text{EA}}{2}$$

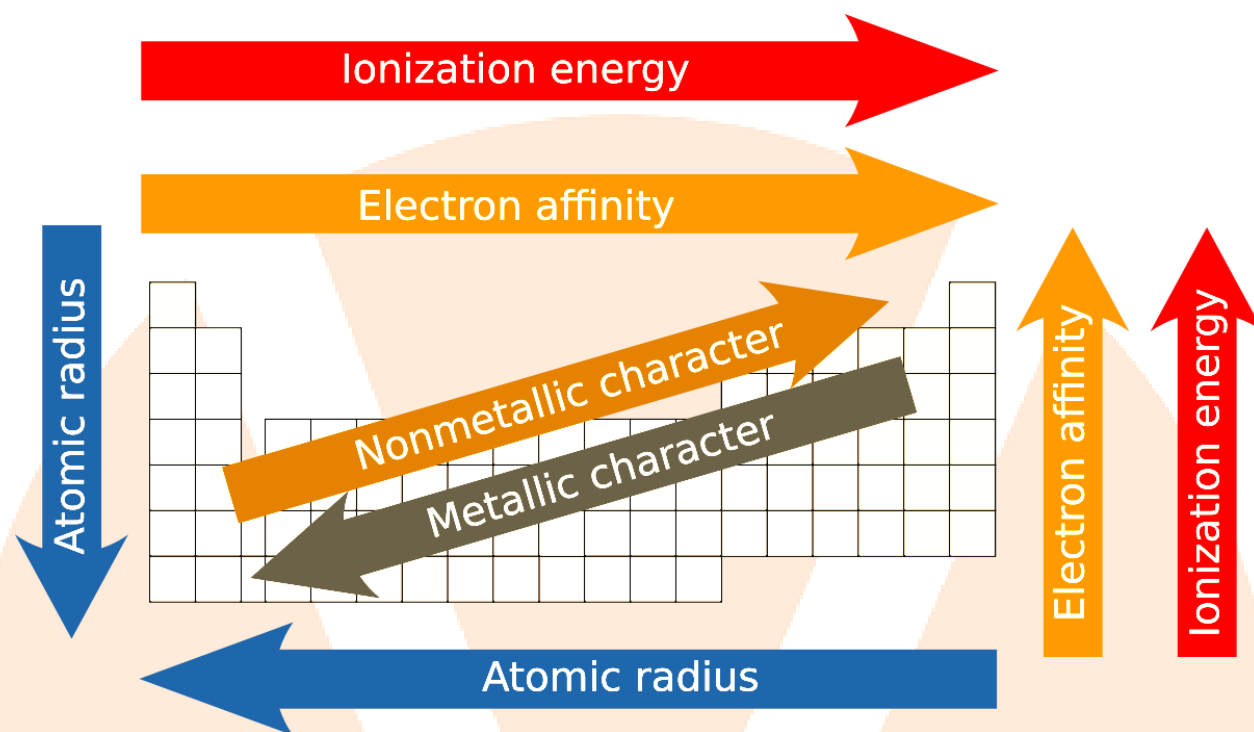
Where, IP is the Ionization potential and EA is the Electron Affinity

If IP and EA are taken as electron volt:

- Percentage ionic character = $16(\chi_A - \chi_B) + 3.5(\chi_A - \chi_B)^2$, where χ_A and χ_B are electronegativities of A and B.
- If the difference in the electronegativity of combining atoms is 1.7, the bond is 50% covalent and 50% ionic.
- The oxide takes on a basic character when the difference between the electronegativities of oxygen and the element is quite large.



The periodic trends of elements in the periodic table



Periodic Trends in Chemical Properties

Periodicity of Valence or Oxidation States

Valence electrons are the electrons found in an atom's outermost shell, and the number of these electrons determines the atom's valence or valency. Because of this, the atom's outermost shell is also known as the valence shell, and the orbitals that make up the valence shell are known as valence orbitals. The valence of an atom is generally equal to either the number of valence electrons (s - and p - block elements) or eight minus the number of valence electrons in representative elements.

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

Transition and inner transition elements, on the other hand, have varying valence due to the presence of not just valence electrons but also d - or f - electrons. Their most common valences, however, are 2 and 3.

Let us now look into valence periodicity over time and within a group.

a. **Variation along a period:** The number of valence electrons grows from 1 to 8 as we walk through a period from left to right. However, the valence of elements in relation to hydrogen or oxygen increases from 1 to 4 before decreasing to zero. The more electronegative oxygen receives two electrons from each of the two sodium atoms in the production of the Na_2O molecule, giving it an oxidation state of -2 . Sodium, on the other hand, loses one electron to oxygen and has an oxidation state of $+1$ because of its valence shell electrical configuration of $3s^1$. Thus, an element's oxidation state in a compound can be described as the charge obtained by its atom as a result of the electronegativity of the other atoms in the molecule.

b. **Variation within a group:** Because the number of valence electrons remains constant as we proceed down the group, all elements in a group have the same valence. All elements in group 1 (alkali metals), for example, have valence one, whereas all elements in group 2 (alkaline earth metals) have valence two. Noble gases of group 18 have zero valences, which means their valence is zero because they are chemically inert.

Anomalous Properties of Second Period Elements

Some elements from the second period appear to be identical to elements from the third period when put diagonally next to each other while belonging to separate groups. Lithium (group 1), for example, resembles magnesium (group 2), while beryllium (group 2) resembles aluminium (group 13), and so on. A diagonal relationship refers to the resemblance in attributes of elements that are arranged diagonally next to each other.



The elements' unusual behaviour is caused by their short size, huge charge/radius ratio, and strong electronegativity. Furthermore, the first member of the group only has four valence orbitals ($2s$ and $2p$) accessible for bonding, whereas the second member of the group has nine ($3s$, $3p$, $3d$). As a result, the maximum covalency of

the first member of each group is 4 (for example, boron can only create $[BF_4]^-$), whereas the other members of the groups can expand their valence shell to include more than four pairs of electrons (for example, aluminium can form $[AlF_6]^{3-}$). Furthermore, as compared to later members of the same groups, the first member of p-block elements has a larger potential to make 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., $C=O$, $C=N$, $C\equiv N$, $N=O$).

Periodic Trends and Chemical Reactivity

Reactivity of Metals

The ability of metals to lose electrons from their outermost shell is used to determine their reactivity.

In a period: Going from left to right in a period, an element's tendency to lose electrons decreases. As a result, metal reactivity reduces as you move from left to right. The reactivity of third-period elements, for example, follows the order:
 $Na > Mg > Al$

In a group: As we move down the group, the tendency to lose electrons rises. As a result, metal reactivity rises as one moves down the group. As a result, the reactivity in group 1 follows the order: $Li < Na < K < Rb < Cs$

Reactivity of Non – metals

The likelihood of a non-metal to gain electrons and form an anion is used to determine its reactivity.

In a period: Non-metals' reactivity rises from left to right over time. Non-metals tend to produce anions during reactions. The reactivity of non-metals, for example, increases in order in the second phase. $C < N < O < F$

In a group: As we move down the group, the reactivity of non-metals declines. This is due to the fact that as one moves down the group, the inclination to accept electrons decreases. The order of halogen reactivity is as follows: $F > Cl > Br > I$

The basic normal oxide generated by the element on the far left (e.g., Na_2O) is the most basic, whereas the acidic normal oxide formed by the element on the far right is the most acidic (e.g., Cl_2O_7). The oxides of the elements in the centre are either amphoteric (Al_2O_3 , As_2O_3) or neutral (**CO**, **NO**, N_2O). Amphoteric oxides are

acidic with bases and basic with acids, whereas neutral oxides are neither acidic nor basic.

Inert Pair Effect

As we progress through groups 13-16, the tendency of valence shell s-electrons to participate in bond formation declines. Lower oxidation states become more stable as a result of this.

Reason: The effect of the increasing nuclear charge overcomes the effect of the equivalent rise in atomic size as we move down these groups. As a result, the s - electrons become more securely held (penetrating) and hence less willing to participate in bond formation. The lower oxidation state becomes more stable as a result.

Summary and Important Points to Remember

1. The periodic table of Mendeleev was based on the atomic masses of the elements. Only 63 elements were known when Mendeleev presented the periodic chart. In the table, he left 29 spots for unknown elements.
2. The atomic numbers of the elements are used in the modern Mendeleev periodic table. "The physical and chemical properties of the elements are a periodic function of their atomic numbers," says the modern periodic law. A period is a horizontal row in the periodic table, and a group is a vertical column. The contemporary Mendeleev periodic table has seven periods and nine groups.
3. Seven eras and eighteen vertical columns make up the long or expanded periodic table (groups or families). A period's elements all have the same number of energy shells, or principal quantum number (n). These are numbered one through seven. There are now 114 elements known. The elements in a vertical column (group) have comparable valence shell electrical configurations and, as a result, chemical characteristics.
4. The s -, p -, d -, and f - blocks of elements are named for the orbital that receives the last electron. These blocks' electronic configuration is as follows:

s - block: [Noble gas] ns^{1-2} , but hydrogen has $1s^1$ configuration.

p - block: [Noble gas] ns^2np^{1-6}

d - block: [Noble gas] $(n-1)d^{1-10}ns^{1-2}$

f - block: [Noble gas] $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$

5. The ingredients are classified into three categories:

(i). Metals account for more than 78% of all known elements. Metals are the s - block, d - block, and f - block elements. Metals make up the higher components of the p - block.

(ii). The number of non-metals is less than twenty. (Non-metals include C , N , P , O , S , Se , H , F , Cl , Br , I , He , Ne , Ar , Kr , Xe , and Rn .)

(iii). Semimetals or metalloids are elements that exist on the boundary between metals and non-metals. Metalloids include B , Si , Ge , As , Se , Te , Po , and At .

6. The IUPAC has devised a new mechanism for giving newly discovered elements a temporary name. The name is derived directly from the element's atomic number. The names of the elements from atomic numbers 104 to 110 have been accepted by IUPAC.

7. Periodicity is defined as the occurrence of identical properties of the elements at regular intervals when the elements are grouped in the periodic table in order of increasing atomic numbers. The repetition of similar electronic configurations of the atom in the valence shell after certain specified periods is the cause of periodicity. These definite intervals are 2, 8, 8, 18, 18 and 32. These are referred to as "magic numbers." Periodicity can be seen in a variety of qualities that are either directly or indirectly related to the electronic configuration.

(i). Across a period, the effective nuclear charge increases.

(ii). Atomic radii tend to decrease as time passes.

(iii). As you progress through the groupings, the atomic radii tend to rise.

(iv). There are three types of atomic radius:

a. **Covalent radius:** The distance between the nuclei of two comparable atoms linked by a single covalent bond is half of this radius. Non-metals are typically described in this way.

b. **Crystal or metallic radius:** In the metallic lattice, it is half of the internuclear distance between the two closest atoms. It's a term that's commonly applied to metals.

c. **Van der Waals' radius:** In the solid-state, it is half of the internuclear distance between the nearest atoms of two neighbouring molecules.

For an atom:

Van der Waals' radius > Metallic radius > Covalent radius

- Anions are typically larger than cations.
- Cations are smaller and anions are larger than the elements' neutral atoms.

Cation size < Neutral atom size < Anion size

- Due to lanthanide contraction, elements in the 2nd and 3rd transition series belonging to the same vertical column are comparable in size and characteristics.
- The first element in each group of representative elements has aberrant features, i.e., it varies from the other elements in the group due to the atom's considerably smaller size.
- Isoelectronic ions are ions with the same amount of electrons but a different nuclear charge. The size of isoelectronic ions reduces as the atomic size grows, i.e., the larger the nuclear charge, the smaller the ion.
- Ionisation enthalpy is the amount of energy necessary to remove the most loosely held electron from a gaseous isolated atom.
- The values of ionisation enthalpy tend to rise over time.
- The enthalpy of ionisation decreases as you progress through the group.
- Higher energy is required to remove electrons from filled and half-filled shells. The ionisation enthalpy of nitrogen, for example, is larger than that of oxygen. High values are found in **Be**, **Mg**, and noble gases.
- Ionisation enthalpy values in metals are low, but non-metals have large ionisation enthalpy values.

- The values of successive ionisation enthalpies of an atom are higher.
 $IE_1 < IE_2 < IE_3 \dots$
- Electron gain enthalpy is the enthalpy change that occurs when an electron is added to an isolated gaseous atom of an element. The first electron gain enthalpy of most elements is negative since energy is released during the process, but in the case of atoms with stable configurations such as **Be**, **Mg**, **N**, noble gases, and so on, the values are positive or near zero.
- From left to right in a period, electron gain enthalpy grows more negative, and from top to bottom in a group, it becomes less negative.
- The enthalpies of successive electron gains are always positive.
- Negative electron gain enthalpy is higher in elements with higher ionisation enthalpy.
- An atom's tendency to attract the shared pair of electrons in a bond is known as electronegativity.
- Electronegativity rises over the periods and falls as you progress through the groups.
- Metals have low electronegativities, while non-metals have high ones.
- Metallic character decreases with time and increases as the group progresses.
- The elements' reducing nature decreases over time, whereas their oxidising nature increases.
- Moving from left to right in a period, the basic character of the oxides decreases while the acidic character grows.

Some Important Facts About Elements

- At room temperature, bromine is a non-metal that is liquid.
- At ambient temperature, mercury is the only metal that is liquid.
- Metals with low melting points include gallium (m.pt. 29.8°C), caesium (m.pt. 28.5°C), and francium (m.pt. 27°C).
- Tungsten (W) is the metal with the highest melting point 3380°C .
- Among nonmetals, carbon has the greatest melting point 4100°C .

- The most prevalent element on the planet is oxygen.
- The most plentiful metal is aluminium.
- The transition metal iron is the most common.
- Osmium (22.57 g cm^{-3}) and iridium (22.61 g cm^{-3}) have the highest density.
- The lightest metal is lithium. It has a density of 0.54 g cm^{-3} .
- The best conductor of electricity is silver.
- The hardest natural substance is diamond (carbon).
- The atomic volume of francium is the largest.
- Boron has the smallest atomic volume of all the elements.
- Nitrogen is the most abundant gas in the atmosphere.
- Fluorine is the most electronegative element in the periodic table.
- Chlorine has the highest negative electron gain enthalpy of any element.
- The ionisation enthalpy of helium is the highest.
- The ionisation enthalpy of cesium or francium is the lowest.
- The smallest and largest atoms are helium and francium, respectively.
- The smallest and largest anions are H^- and I^- ions, respectively.
- The smallest and largest cations are H^+ and Cs^+ , respectively.
- The most electropositive element is Cesium.
- Phosphorus, P_4 , is an element that is preserved in water (white or yellow).
- **Na**, **K**, **Rb**, **Cs** and other elements are stored in kerosene.
- Iodine is the element that transcends.
- The most prevalent element in the universe is hydrogen.
- The only coloured gas with a garlic odour is ozone.
- Electronegativity values for metalloids are closer to 2.0.
- Technetium is the first synthetic (i.e., man-made) element (At. No. 43).
- Plutonium is the most toxic metal.
- Astatine is the rarest element in the earth's crust.
- Transuranic elements are the elements that come after uranium. Transactinides, or super heavy elements, are those with $Z = 104-112, 114$, and 116 . All of these elements are synthetic, meaning they were created by humans. These are radioactive elements that do not occur naturally.
- Ruthenium (*Ru*), germanium (*Ge*), polonium (*Po*), and americium (*Am*) are elements named for the countries Ruthenia (Russia), Germany, Poland, and America, respectively.
- The actinide series contains radioactive elements, the majority of which are not found in nature.

- Rutherfordium (*Rf*,104) is also known as Kurchatovium (*Ku*) , and dubnium (*Db*,105) is also known as hahnium.
- Promethium (*Pm*,61) , a lanthanide element, does not exist in nature. It's a man-made substance.
- Members of these groups are given special names in the periodic table.

Group 1	Or	I A	Alkali metals
Group 2	Or	II A	Alkaline earth metals
Group 15	Or	V A	Pnicogens
Group 16	Or	VI A	Chalcogens
Group 17	Or	VII A	Halogens
Group 18	Or	VIII A	Inert or noble gases