

SLO 4: Periodic Table and Periodicity

4.1 Periodic Table

4.1.1 State The Modern Periodic Law

In 1869 Mendeleev Proposed a periodic law on the basis of physical and chemical properties empirically. Periodic law states that “The Properties of the elements are a periodic Function of their atomic weight”. In 1913 Moseley discovered that atomic number is the basic property of an atom. He proposed a modern periodic law. The Moseley states that “The Physical and chemical Properties of elements are the periodic function of their atomic numbers”.

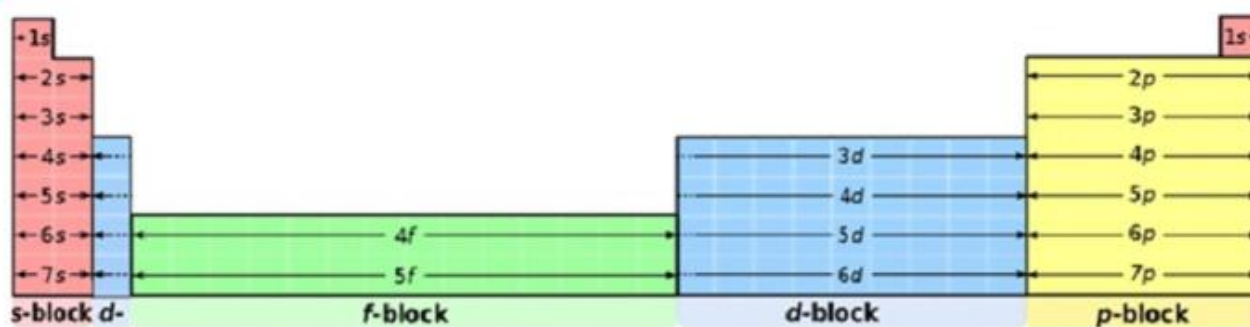
4.1.2 Differentiate Between A Period And A Group In The Periodic Table

FEATURE	PERIOD	GROUP
Definition	A horizontal row in the periodic table.	A vertical column in the periodic table.
Direction	Left to right.	Top to bottom.
Number	7 periods.	18 groups (1 to 18).
Element Arrangement	Increasing atomic number and number of electron shells.	Similar number of valence electrons and similar chemical properties.
Electron Shells	Same number of electron shells.	Shell increase down the group.
Atomic Number	Increases from left to right	Increases from top to bottom
Valence Electrons	Varies across the period	Remains the same in the same group.
Chemical Properties	Change gradually across a period.	Are similar within a group.
Reactivity Trend	Metals become less reactive and non-metals become more reactive across a period.	In metals, reactivity increases down the group; in non-metals, it decreases.
Block Composition	Can include elements from different blocks (s, p, d, f).	Consists of elements from the same block (main group, transition, etc.).
Metallic Character	Decreases from left to right.	Increases from top to bottom.
Size (Atomic Radius)	Decreases across the period (due to increasing nuclear charge).	Increases down the group (due to added electron shells).

4.1.3 Determine The Group, Period And Block Of An Element Using Its Electronic Configuration (First Twenty Elements)

4.1.4 Describe The Demarcation Of The Periodic Table Into S, P, D, And F-Blocks

The periodic table has been divided into four blocks, s, p, d, and f based on electronic configuration.



- Nobel Gases: They are colorless, unreactive and diamagnetic, they are placed in zero group. Their electronic configuration is ns^2, np^6 and are exceptionally stable.
- Representative Elements: It includes metals and nonmetals. Some are diamagnetic and some are paramagnetic and marked as S block and P block elements.

- s-Block Elements: electrons occupy in ns orbital. The elements of group IA and IIA are s-block elements. Their electronic configuration varies ns^1 to ns^2 .
- p-Block Elements: electrons begin to fill np^1 to np^6 . Elements of group IIIA to VIIA and zero group except He are also p-block elements.
- d-block Elements (Outer Transition Elements): These elements are metals but their properties are different from metals of representative elements i.e. Melting point, Boiling Point, variable oxidation state color compounds etc. In these elements electron fills in $ns^2 (n-1) d^{1-10}$ orbital. d-block elements consist of three series.
- f-Block Elements (Inner Transition Elements): The elements in which inner f-orbital is filled, are called f block elements. They exhibit electronic configuration $ns^2(n-1) d^1(n-2) f^{1-14}$. There are two series called Lanthanides and Actinides.

4.1.5 Determine The Location Of Families On The Periodic Table Based On Their Characteristics And Electronic Configuration (Representative Elements)

Family Name	Group Number	Block	Electron Configuration	Common Characteristics	Examples
Alkali Metals	Group 1 (IA)	s-block	ns^1	- Soft, highly reactive metals - React violently with water - Form +1 ions	Li, Na, K, Rb, Cs, Fr
Alkaline Earth Metals	Group 2 (IIA)	s-block	ns^2	- Harder than alkali metals - Less reactive - Form +2 ions	Be, Mg, Ca, Sr, Ba, Ra
Boron Family	Group 13 (IIIA)	p-block	$ns^2 np^1$	- Mostly metals (except B) - Form +3 ions	B, Al, Ga, In, Tl
Carbon Family	Group 14 (IVA)	p-block	$ns^2 np^2$	- Varying properties - Includes nonmetals, metalloids, and metals	C, Si, Ge, Sn, Pb
Nitrogen Family	Group 15 (VA)	p-block	$ns^2 np^3$	- Tend to form -3 ions - Show multiple oxidation states	N, P, As, Sb, Bi
Oxygen Family (Chalcogens)	Group 16 (VIA)	p-block	$ns^2 np^4$	- Reactive nonmetals - Commonly form -2 ions	O, S, Se, Te, Po
Halogens	Group 17 (VIIA)	p-block	$ns^2 np^5$	- Very reactive nonmetals - Form salts with metals - Form -1 ions	F, Cl, Br, I, At
Noble Gases	Group 18 (VIIIA)	p-block	$ns^2 np^6$ (except He: $1s^2$)	- Inert gases - Very stable - Full outer shell	He, Ne, Ar, Kr, Xe, Rn

4.2 Periodic Properties

4.2.1 Define The Following Terms:

Shielding Effect

Electrons present in the inner shells shield the force of attraction of nucleus felt by the valence shell electrons is called Shielding effect.

Atomic Radii

Half of the distance between the centers of two bonded atoms of any elements is considered to be the radius of the atom. It is measured in Angstrom unit (A) $1A = 10^{-8}$ cm.

Electron Affinity

The electron affinity is amount of energy released when an electron is added in the outermost shell of a gaseous atom. It is calculated in KJ/mol. Affinity means attraction; therefore, electron affinity means tendency to accept electron to form anion.

Ionization Energy (1st And 2nd)

The ionization energy is minimum amount of energy required to remove an electron from a gaseous state and measured in joule/mole. The ionization energy depends upon atomic size and nuclear charge. The higher ionization energy means removal of electron is more difficult.

Electronegativity

The ability of an atom to attract the shared pair of electrons towards itself in a molecule is called electronegativity.

4.2.2 Explain The Periodic Trend Within A Group And A Period Of The Periodic Table**Shielding Effect**

The electrons present between the nucleus and valence shell of atom reduce the nuclear charge effect on electrons present in outermost shell. As a result, valence electron experience less nuclear charge than the actual charge. Therefore, The Shielding effect increases down the group in periodic table and remain same in period from left to right.

Electronegativity

It increases from left to right in period due to increase in nuclear charge which decrease the distance from nucleus to shared electron pair. It increases the power to attract the shared pair of electrons. In group electronegativity decreases because size of Atom increases and attraction for shared electron pair decreases.

Atomic Radii

In the periodic table, that atomic radius increases from top to bottom within a group due to increase in number of shells. However, as the atomic number increases from left to right, the atomic radius decrease. This gradual decrease in the radius is due to increase in the positive charge on the nucleus. As the positive nuclear charge increases, the negatively charged electrons in the shells are pulled closer to the nucleus. Thus, the size of the outermost shell becomes gradually smaller.

Electron Affinity

In a period, electron affinity increases from left to right due to decrease of atomic size because when size of atom decreases the attraction between nucleus and incoming electron increases and more energy is released.

In a group electron affinity values decrease from Top to bottom, because the size of atom increases. Down the group attraction of incoming electron and Nucleus decreases and less energy released. For example, the size of iodine is bigger than bromine, so its electron affinity is less than bromine.

Ionization Energy (1st And 2nd)

If we move from left to right in periods the value of ionization energy increases. It's because of size of atoms reduces and electrons are held strongly by the attractive force of nucleus. Due to this elements on the left side have less ionization energy.

As we move down the group ionization energy decreases from top to bottom due to additions of shells. Increase in number of shells reduce the electrostatic force between electrons of valence shell and nucleus.

4.2.3 Compare The Chemical Reactivity And Physical Properties Of Elements In The Same Family Of Elements (Representative Elements)

Property	Group Trend	Explanation
Chemical Reactivity	Increases in metals Decreases in nonmetals	Metals lose electrons more easily Non-metals gain electrons less easily.
Metallic Character	Increases	Atoms lose electrons more readily due to weaker nuclear attraction on outer electrons.
Physical State	Metals: Solid throughout Non-metals: From gas to solid	Non-metals become heavier and more complex molecules
Electrical Conductivity	Remains high in metals; may vary slightly	Metals maintain free electrons for conduction.
Density	Generally, increases	Heavier atoms and tighter packing increase density down the group.
Melting Point	Decreases in alkali metals Varies in other groups	Weaker metallic bonding due to larger atomic radii (Cs < Na).
Boiling Point	Decreases in alkali metals Varies in other groups	Weaker forces as atoms get larger.

4.3 Properties of Elements in a Group

4.3.1 Determine Elements As An Alkali Metal, An Alkaline Earth Metal, A Halogen And A Noble Gas Based On Their Electronic Configuration

Element Type	Group	Valence Shell Configuration	General Rule
Alkali Metals	Group 1	ns^1	1 electron in the outermost s-orbital
Alkaline Earth Metals	Group 2	ns^2	2 electrons in the outermost s-orbital
Halogens	Group 17	$ns^2 np^5$	7 valence electrons: 2 in s, 5 in p
Noble Gases	Group 18	$ns^2 np^6$ (He: $1s^2$)	Full valence shell

4.3.2 Compare The General Properties Of Metals And Non-Metals In Terms Of

Property	Metals	Non-Metals
Physical States	Solid at room temperature	Mostly gases or solids
Density	Generally high density	Generally low density
Malleability	Malleable – can be hammered into thin sheets	Brittle – break or shatter when hammered (if solid)
Ductility	Ductile – can be drawn into wires	Non-ductile – break when stretched
Melting and Boiling Points	Usually, high	Usually low (except carbon and silicon which have high melting points)
Electrical Conductivity	Good conductors of electricity (due to free-moving electrons)	Poor conductors (except graphite)
Thermal Conductivity	Good conductors of heat	Poor conductors of heat
Sonority	Sonorous – produce ringing sound when struck	Non-sonorous – dull sound
Appearance	Lustrous – shiny appearance	Dull (except iodine, which is lustrous)
Hardness	Generally hard and strong	Generally soft and brittle
Nature of Oxides	Form basic oxides (e.g., Na_2O , CaO) Some form amphoteric oxides (e.g., Al_2O_3 , ZnO)	Form acidic oxides (e.g., CO_2 , SO_2) Some may form neutral oxides (e.g., CO , NO)

4.3.3 Explain The Following Properties Of Group I And II Elements:

Occurrence In Combined State In Nature

- Found only in combined form, not free, because they are highly reactive.
- React easily with water, oxygen, and acids.
- Found as salts (e.g., NaCl, CaCO₃, KNO₃, CaSO₄).
- Must be extracted by electrolysis or other methods.

Softness Of Metals

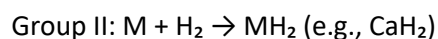
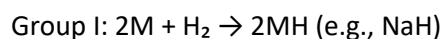
- Group I metals (e.g., Li, Na, K) are very soft – can be cut with a knife.
- Group II metals (e.g., Mg, Ca) are harder than Group I but still softer than transition metals.

Reaction With Water

- Group I metals react violently with cold water. They form a strong base (alkaline solution) and hydrogen gas. $\text{Na} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2$
- Group II metals react less vigorously, especially Be and Mg (which are slow or unreactive with cold water). $\text{Ca} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{H}_2\uparrow$

Reaction With Hydrogen

- Both groups form metal hydrides when heated in hydrogen gas. These hydrides are ionic, and they release hydrogen gas when reacting with water.



Reaction With Oxygen

- Group I metals form oxides, peroxides, or superoxide, depending on the metal:
 - Li forms oxide (Li₂O)
 - Na forms peroxide (Na₂O₂)
 - K forms superoxide (KO₂)
- Group II metals form basic oxides (e.g., MgO, CaO)

Reaction With Dilute Acids

- Both Group I and II metals react vigorously with dilute acids to form salt + hydrogen gas: $\text{M} + \text{HCl} \rightarrow \text{MCl} + \text{H}_2\uparrow$
- Group I reacts more violently. Group II reactions are slightly less vigorous but still exothermic.

4.3.4 Explain The Following Properties Of Group VII Elements:

Existence As Diatomic Molecules

Halogens naturally exist as diatomic molecules, meaning two atoms bonded together (F₂, Cl₂, Br₂, I₂, At₂). This is because they need one electron to complete their outer shell, and by sharing one electron each, two halogen atoms achieve a stable noble gas configuration via a single covalent bond.

Appearance (Color And State)

Halogens show a gradual change in physical state and color as we go down the group due to increasing molecular mass and intermolecular forces. Fluorine (Pale Yellow), Chlorine (Greenish-Yellow), Bromine (Reddish-Brown), Iodine (Purple-Black), Astatine (Black).

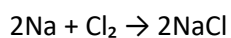
Displacement Reactions With Other Halogens

More reactive halogens can displace less reactive halogens from their salt solutions. Chlorine can displace bromine and iodine, Bromine can displace iodine but not chlorine, and Iodine cannot displace Br₂ or Cl₂.

Reaction With Metals

Halogens react vigorously with metals, especially alkali and alkaline earth metals, to form ionic halide salts. These salts are usually white crystalline solids and soluble in water. The reactivity decreases down the group, so fluorine reacts most violently, while iodine reacts less readily.

Metal + Halogen → Metal Halide



4.3.5 Explain The Following Properties Of Group VIII Elements:

Existence As Monoatomic Gases Or Free State In Nature

Noble gases (He, Ne, Ar, Kr, Xe, Rn) exist as monoatomic gases—each atom exists independently, not bonded to others. This is because they have complete outermost (valence) electron shells making them stable and not needing to bond with other atoms. As a result, noble gases are found in free state in nature (not combined with other elements) and are chemically unreactive.

Chemical Inertness

Noble gases are chemically inert (unreactive) because they already have a stable electronic configuration (full valence shell). They do not gain, lose, or share electrons easily, meaning they rarely form chemical bonds. This makes them nonreactive with most other elements, especially at room temperature and pressure.

Importance Of Noble Gas Electronic Configuration In The Formation Of Ions

Atoms of other elements tend to lose or gain electrons to achieve the stable noble gas configuration (also known as the octet rule). This concept is fundamental in the formation of ions:

- Metals: Lose electrons to get noble gas configuration
- Non-metals: Gain electrons to get noble gas configuration

4.3.6 Explain The Following Properties Of Transition Metals:

1. Relative Hardness: Transition metals are generally harder and stronger than s-block metals due to strong metallic bonding and closely packed atoms. The presence of unpaired d-electrons allows formation of strong interatomic bonds, increasing rigidity.
2. Density: Transition metals have high densities because they have small atomic radii and high atomic masses. Their atoms are tightly packed, which contributes to a greater mass per unit volume.
3. Melting and boiling points: They have high melting and boiling points due to strong metallic bonds formed by delocalized d-electrons. These electrons create strong attraction between metal ions, requiring more energy to break. (Exception: Mercury is liquid at room temperature).
4. Variable oxidation states: Transition metals exhibit more than one oxidation state (e.g., Fe^{2+} and Fe^{3+} , Cu^+ and Cu^{2+}) because both the outermost s- and d-electrons can participate in bonding. This variability gives them flexibility in forming compounds.
5. Formation of colored compounds: Transition metal ions form colored compounds due to d-d electron transitions. When light hits the compound, some d-orbitals absorb specific wavelengths, and the remaining light is reflected as color. Example: Cu^{2+} (blue), Cr^{3+} (green), Fe^{3+} (yellow-brown).