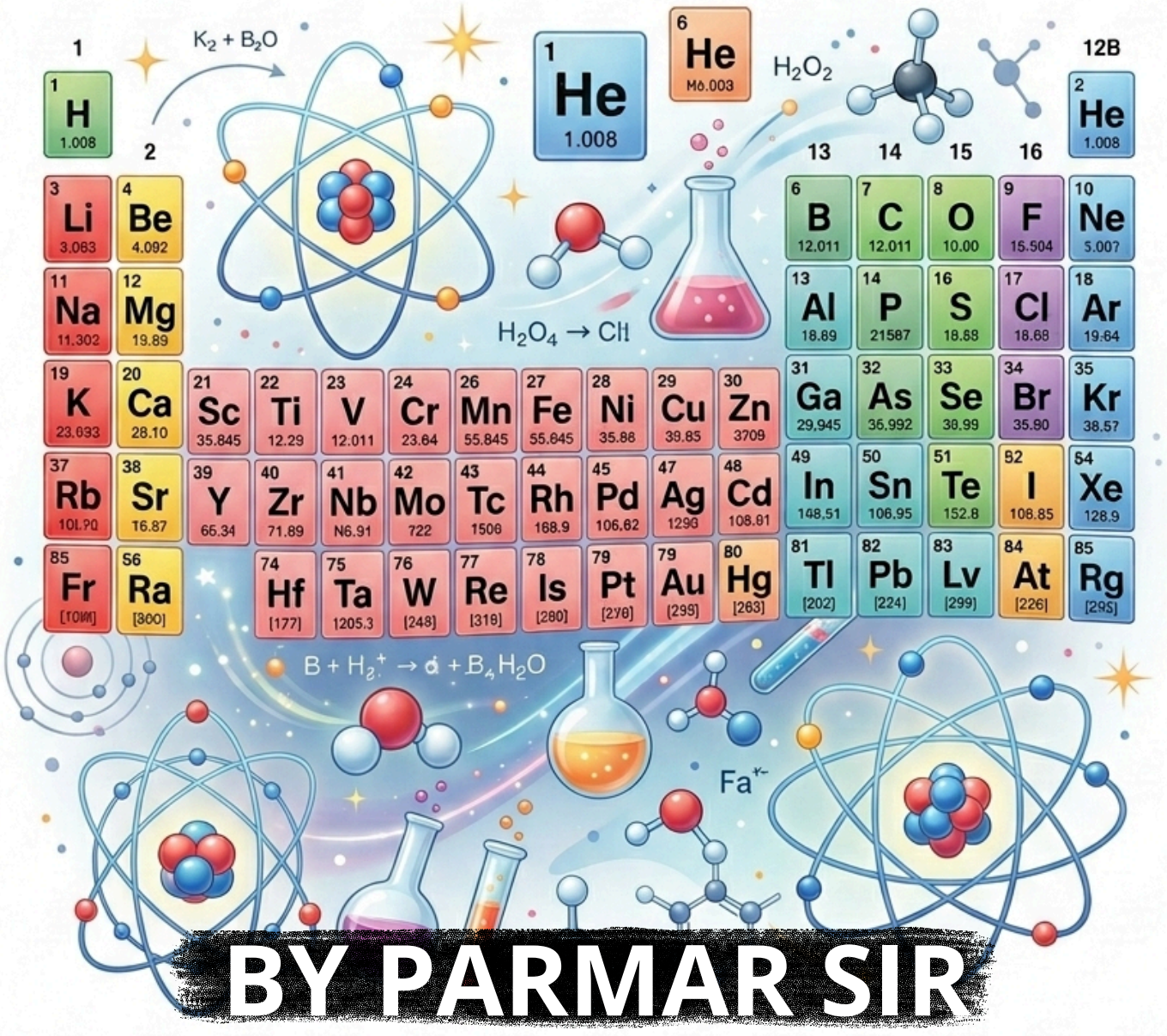
PARMAR SSC

PERIODIC TABLE

Lecture - 4



EARLY ATTEMPTS AT CLASSIFICATION OF ELEMENTS

Dobereiner's Law of Triads (1817)

- Dobereiner observed that when groups of three elements with similar properties (which he named "triads") were arranged by increasing atomic mass, the atomic mass of the middle element was approximately the average of the atomic masses of the 1st and 3rd elements

Elements	Atomic Mass	Average
Lithium (Li) Sodium (Na) Potassium (K)	6.9 23.0 39.0	$\frac{6.9 + 39.0}{2} = 22.95$
Calcium (Ca) Strontium (Sr) Barium (Ba)	40.1 87.6 137.3	$\frac{40.1 + 137.3}{2} = 88.7$
Chlorine (Cl) Bromine (Br) Iodine (I)	35.5 79.9 126.9	$\frac{35.5 + 126.9}{2} = 81.2$
Sulfur (S) Selenium (Se) Tellurium (Te)	32.1 78.96 127.6	$\frac{32.1 + 127.6}{2} = 79.85$
Iron (Fe) Cobalt (Co) Nickel (Ni)	55.8 58.9 58.7	$\frac{55.8 + 58.7}{2} = 57.25$

Newland's Law of Octaves (1865)

- Newland discovered that when elements are arranged in increasing order of their atomic mass, the 1st and the 8th elements showed similar properties and characteristics.

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	-	-

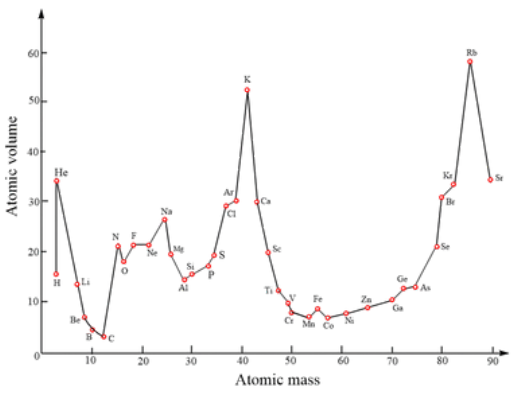
- Newland arranged the elements in horizontal rows, with each row having 7 elements and compared this to the eight octaves of music (do, re, mi, fa, so, la, ti, do).
- He published his concept in 1864 and recognition was given in 1865.
- **Last element:** Thorium
- **Total number of elements:** 56

Limitations of Newland's Model

1. The law was not true for elements beyond Calcium. After calcium, **every eighth element did not share similar properties to the first one.**
2. Those elements that were discovered after Newland's octaves **did not follow the law of octave.**
3. To fit elements into his table, Newland sometimes placed two elements in the same position, even when they had different chemical and physical properties. **E.g.** Cobalt and nickel were grouped together in a single slot, despite their distinct characteristics. At the same time, iron—an element with properties closely resembling those of cobalt and nickel—was placed far away.
4. The periodic table did not include noble gases because they were not discovered then.

Lothar Meyer

- Meyer published his **first periodic table** in 1862, which included 28 elements.
- He published an **updated version in 1864, which included 50 elements.**
- He was the first to identify periodic trends in the properties of elements.
- He plotted the atomic volume of an element against its atomic weight to show this pattern.



Mendeleev's Periodic Table (1869)

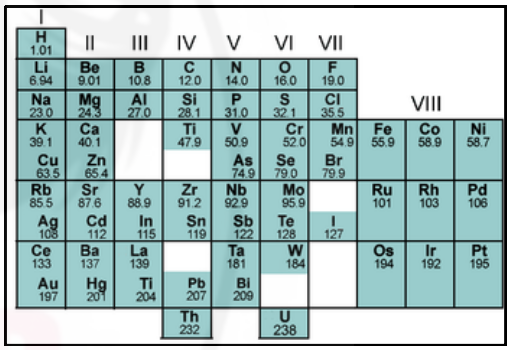
- The periodic table was created in 1869 by Dmitri Mendeleev a Russian chemist and inventor.
- He is known as the Father of the Periodic Table.
- He arranged the 63 known elements at the time in order of their increasing relative atomic masses.
- He divided the table into 8 groups and 7 periods.

Mendeleev's Periodic Law

- Mendeleev's law states that the physical and chemical properties of elements are periodic functions of their atomic masses, meaning that when elements are arranged in order of increasing atomic mass, elements with similar properties repeat themselves at regular intervals.

Element	Atomic Mass
Eka Boron	44 (Scandium) → 44.955912 u
Eka Silicon	72 (Germanium) → 72.64 u
Eka Aluminium	68 (Gallium) → 69.723 u

Mendeleev's Given Name	Modern Name
Eka-aluminium	Gallium
Eka-boron	Scandium
Eka-silicon	Germanium
Eka-manganese	Technetium
Tri-manganese	Rhenium
Dvi-tellurium	Polonium
Dvi-caesium	Francium
Eka-tantalum	Protactinium

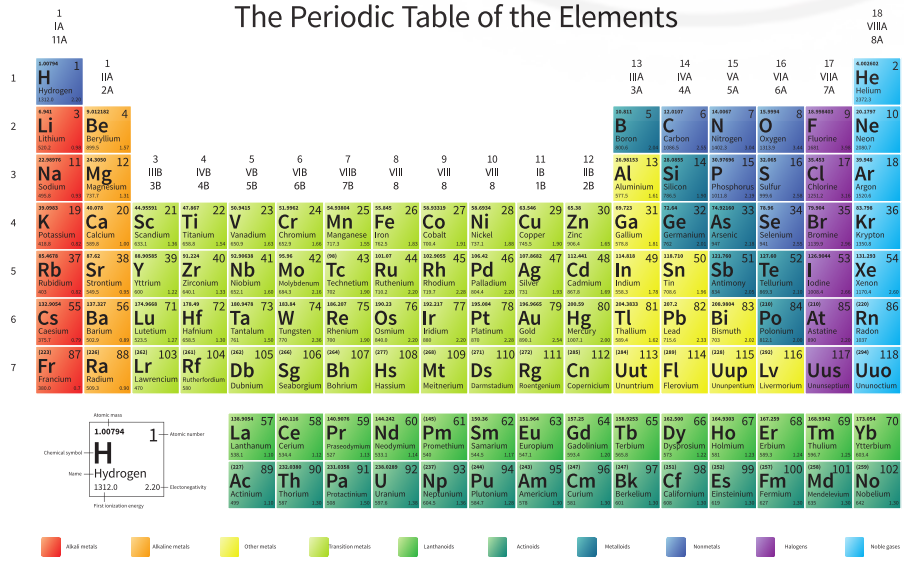


Limitations of Mendeleev's Periodic Table

1. The position of hydrogen is not fixed.
2. The increasing order of atomic weight is not maintained.
3. Some elements in the same group differ in their properties.
4. Lanthanides and Actinides were not included in the table.

Modern Periodic Table

- It was given by **Henry Moseley in 1913.**
- Moseley discovered that elements are best characterized by their atomic number (the number of protons in the nucleus) and not by atomic weight.
- **The modern periodic table contains:** 18 Groups and 7 Periods.



Classification of Groups

Group 1: Alkali metals
Group 2: Alkaline Earth metals
Group 15: Pnictogens
Group 16: Chalcogens
Group 17: Halogens
Group 18: Noble gas/Inert gas

Blocks and Elements

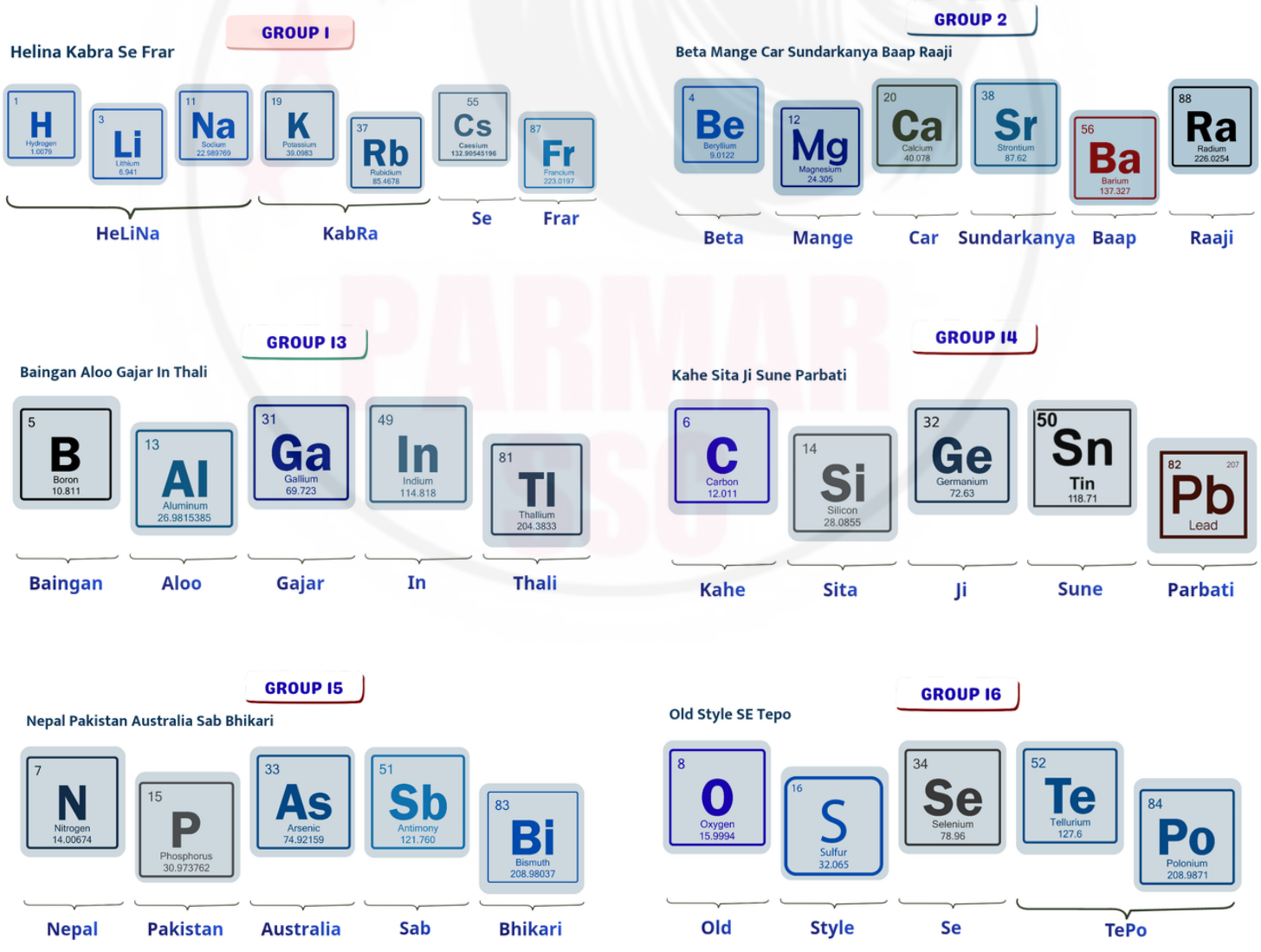
- **s-block**: Group 1 & 2.
- **p-block**: Group 13 to 18 (excluding Helium).
- **d-block**: Group 3 to 12 (Transition Metals).
- **f-block**: Inner Transition Metals (Lanthanides and Actinides).
- **Representative Elements**: s and p-block elements collectively.

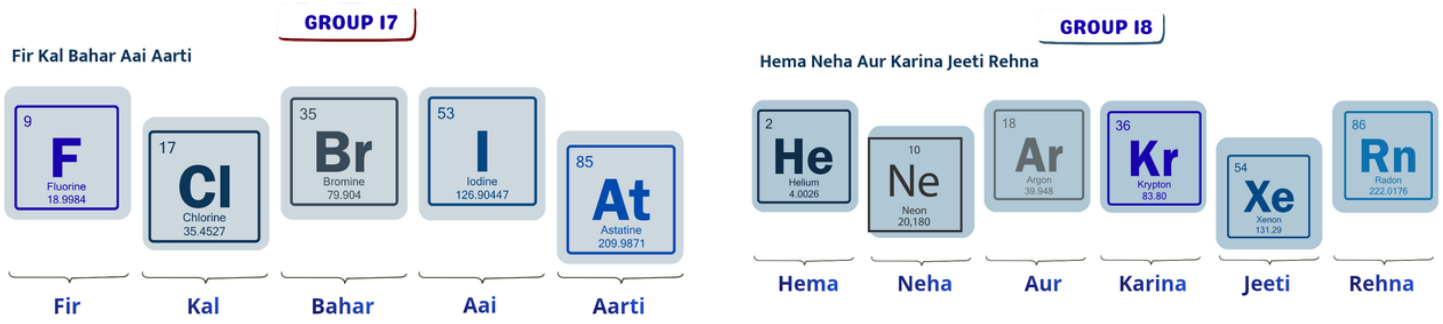
- **Metalloids (7):** Germanium, Arsenic, Antimony, Tellurium, Boron, Polonium, and Silicon (Only in p-block).
- **Naturally occurring elements:** 94
- **1st man-made element:** Technetium (Z = 43)
- **Heaviest naturally occurring element:** Uranium (Z = 92)
- **Highest Density Element:** Osmium
- **Valence electron:** e^- present in the last orbit.

Some Special Elements

- Seaborgium >>> Z = 106
- Mendelevium >>> Z = 101
- Rutherfordium >>> Z = 104

TRICK TO REMEMBER GROUPS





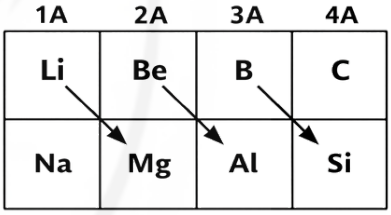
Lanthanides and Actinides

57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

They are collectively called as Inner Transition Metals (f-block).
Lanthanides: 57–71 → Period 6
Actinides: 89–103 → Period 7

Diagonal Relationship

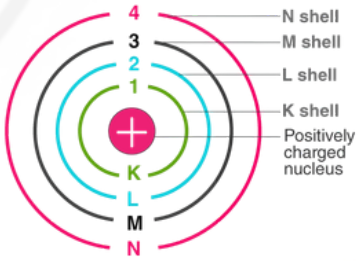
- This relationship is prominent among the lighter members of the 2nd and 3rd periods.
- A diagonal relationship in the periodic table is when two elements that are diagonally adjacent in the 2nd and 3rd periods have similar properties.



Some examples of diagonal relationships include

- Boron and silicon (they are both semiconductors).
- → Li and Mg.
- → Be and Al.
- → C and P.

Electronic Configuration



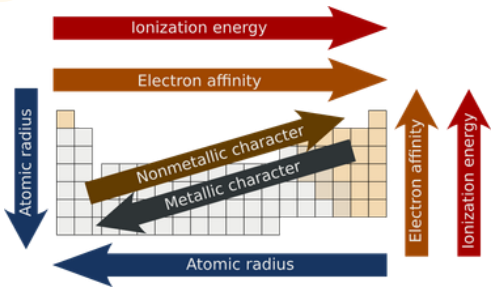
- The electronic configuration is distribution of electrons of an atom or molecule in atomic or molecular orbitals.
- The K, L, M, and N shells represent the first four energy levels (or principal quantum numbers, n = 1, 2, 3, 4) where electrons are distributed, with K holding a maximum of 2, L holding 8, M holding 18, and N holding 32 electrons.

Octet Rule

- The Octet rule refers to the tendency of atoms to prefer to have 8 electrons in the valence or outermost shell. When atoms have fewer than 8 electrons, they tend to react to achieve a stable electronic configuration.
- Octet Rule was given by **Lewis** and **Kossel** in **1916**.

Elements	K	L	M	N
^{3} Li	2	1		
^{11} Na	2	8	1	
^{8} O	2	6		
^{10} Ne	2	8		
^{37} Rb	2	8	18	8, 1

TRENDS IN PERIODIC TABLE



- In the Periodic Table, if a property increases from top to bottom, it decreases from left to right. Conversely, if a property decreases from top to bottom, it increases from left to right.
- **Metallic Character:** Metallic character is essentially defined as the ability of an atom to lose its valence electrons and form a cation, which is known as electropositivity.
- **With respect to Valence Electrons:**
  - **General Rule:** Most metals have 1, 2, or 3 electrons in their outermost shell.

Valency

- Valency is the combining capacity of an atom, indicating how many electrons an atom can lose, gain, or share to form a stable compound.

- **^{11}Na (Sodium)** → 2, 8, 1 → Loses 1e$^-$ to achieve stability → Na^+ (Cation) → Loss of electron.
- **^{17}Cl (Chlorine)** → 2, 8, 7 → Gains 1e$^-$ to complete octet → Cl^- (Anion) → Gain of electron.

- **Metals:** These always lose electrons and form cations. This property is known as electropositivity.
- **Non-metals:** These usually gain electrons and form anions. This property is known as electronegativity.

Variable Valency

- It is the ability of certain elements, particularly transition metals, to exhibit more than one oxidation state or valency in different compounds.

- **Fe** → 2, 3; Ag; Cu (all transition metal show variable valency due to incomplete filling of d-orbitals).

Valence Electrons

- Valence electrons are the electrons present in the outermost shell (valence shell) of an atom. These electrons determine the chemical properties and reactivity of the element.

Duet Rule

- Duet Rule states that atoms like Hydrogen (H) and Helium (He) achieve stability when they have two electrons in their valence shell (K shell). To attain this stable configuration, atoms either gain, lose, or share electrons by forming bonds with other atoms.
- **E.g.:** In a water (H_2O) molecule, hydrogen shares electrons with oxygen to achieve a stable duet configuration.
  - Small elements like Hydrogen (H), Helium (He), Lithium (Li), and Beryllium (Be) are unable to complete their octet (8 electrons) during chemical reactions. Therefore, they achieve stability by completing 2 electrons.

Nuclear Charge → The total positive charge (+) present in the nucleus of an atom is called the Nuclear Charge.

1. **Left to Right: Stronger Nuclear Control**
   - As the atomic number increases, the nuclear charge (number of protons) increases, but no new shells are added.
   - **Result:** The nucleus pulls the electrons more strongly toward itself, which causes the atomic size to decrease.
2. **Top to Bottom: Decrease in Control**
   - As we move from top to bottom, the number of protons also increases (nuclear charge increases), yet the nuclear control decreases. This happens due to two main reasons:
   - **New Shells:** At every step, a new electronic shell is added, which makes the outermost electrons very far from the nucleus.
   - **Shielding Effect:** Electrons in the inner orbits act like a "screen" or "curtain," preventing the nuclear pull from reaching the outermost electrons effectively.

Ionisation Energy

- Ionization energy is the energy required to remove an electron from a neutral atom in its gaseous state.

Element with the Highest Ionization Energy: Helium (He)

Across a period (left to right)

- It increases across a period because it becomes harder to remove an electron from an atom with a higher nuclear charge and smaller atomic radius.

Down the group (top to bottom)

- It decreases down a group as the outermost electrons are farther from the nucleus and more easily removed.

Electronegativity

- Electronegativity is the tendency of an atom of a given chemical element to attract shared electrons (or electron density) when forming a chemical bond.

Across a period (left to right)

- It generally increases as atoms become more effective at attracting electrons due to increasing nuclear charge.

Down the group (top to bottom)

- It generally decreases as the outermost electrons are further from the nucleus and less strongly attracted.

Electron Affinity

- The energy released when an electron is added to a neutral atom to form a negative ion is called electron affinity.

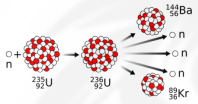
Across a period (left to right)

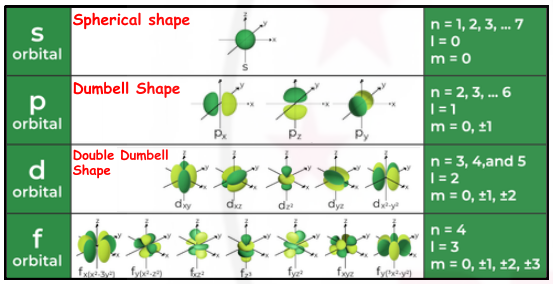
- It generally increases (becomes more negative) because of increasing nuclear charge and decreasing atomic size, making it easier for the atom to attract an extra electron.

Down the group (top to bottom)

- It decreases (becomes less negative) down a group due to increasing atomic size and shielding effect.

- **Elements with highest electron affinity: Chlorine.**
- **Elements with highest electronegativity: Fluorine.**

Noble Gas/Inert Gas/Group 18	Discovered By
Helium (used in balloons)	Pierre Janssen and Norman Lockyer
Neon (used in light bulbs)	William Ramsay and Morris Travers
Argon (used in double pane window glasses)	William Ramsay and Lord Rayleigh
Krypton (used in nuclear fission) 	William Ramsay and Morris Travers
Xenon	William Ramsay and Morris Travers
Radon	Friedrich Ernst Dorn



Number	Symbol	Possible Values
Principal Quantum Number	n	$1, 2, 3, 4, \ldots$
Azimuthal / Angular Momentum Quantum Number	ℓ	$0, 1, 2, 3, \ldots, (n-1)$
Magnetic Quantum Number	m_1	$-\ell, \ldots, -1, 0, 1, \ldots, \ell$
Spin Quantum Number	m_s	$+1/2, -1/2$

- **K Shell (n=1):** 1s
- **L Shell (n=2):** 2s, 2p
- **M Shell (n=3):** 3s, 3p, 3d
- **N Shell (n=4):** 4s, 4p, 4d, 4f

Isoelectronic Species

- Isoelectronic species are atoms, ions, or molecules that possesses the same number of electrons, and therefore the same electronic structure, even if they differ in nuclear charge.
- **E.g.** Oxygen, Fluorine, Magnesium.

$_8O^{2-}$	$_9F^-$	$_{12}Mg^{2+}$
10	10	10

Calculation of Atomic Mass (A)

- When protons (Z) is even, the number of neutrons (N) → 2Z.
- When protons (Z) is odd, the number of neutrons (N) → 2Z + 1.

Exceptions

$_1H \rightarrow 1,\ _4Be \rightarrow 9,\ _7N \rightarrow 14,\ _{18}Ar \rightarrow 40$

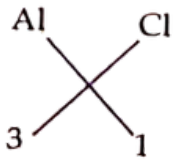
Valency Based on Valence Electrons

- If an element has 1, 2, 3 or 4 valence electrons → its valency is equal to the number of valence electrons.
- **E.g.** Elements in Group 1 (like hydrogen and sodium) have 1 valence electrons, so their valency is 1.
- If an element has 5, 6, 7 or 8 valence electrons → its valency is calculated as:
- 8 – No. of valence electrons.
- **E.g.** Elements in Group 16 (like oxygen and sulphur) have 6 valence electrons, so their valency is 8 – 6 = 2.

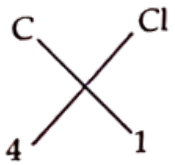
Atomic Number	Name of the Element	Electronic Configuration (subshell wise)	Valency	Electronic Configuration (shell wise)
1	Hydrogen (H)	$1s^1$	1	1
2	Helium (He)	$1s^2$	0	2
3	Lithium (Li)	$1s^2\ 2s^1$	1	2, 1
4	Beryllium (Be)	$1s^2\ 2s^2$	2	2, 2
5	Boron (B)	$1s^2\ 2s^2\ 2p^1$	3	2, 3
6	Carbon (C)	$1s^2\ 2s^2\ 2p^2$	4	2, 4
7	Nitrogen (N)	$1s^2\ 2s^2\ 2p^3$	3	2, 5
8	Oxygen (O)	$1s^2\ 2s^2\ 2p^4$	2	2, 6
9	Fluorine (F)	$1s^2\ 2s^2\ 2p^5$	1	2, 7
10	Neon (Ne)	$1s^2\ 2s^2\ 2p^6$	0	2, 8
11	Sodium (Na)	$1s^2\ 2s^2\ 2p^6\ 3s^1$	1	2, 8, 1
12	Magnesium (Mg)	$1s^2\ 2s^2\ 2p^6\ 3s^2$	2	2, 8, 2
13	Aluminium (Al)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	3	2, 8, 3
14	Silicon (Si)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$	4	2, 8, 4
15	Phosphorus (P)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$	5, 3	2, 8, 5
16	Sulphur (S)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$	2	2, 8, 6
17	Chlorine (Cl)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	1	2, 8, 7
18	Argon (Ar)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	0	2, 8, 8
19	Potassium (K)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	1	2, 8, 8, 1
20	Calcium (Ca)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$	2	2, 8, 8, 2
21	Scandium (Sc)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$	3	2, 8, 9, 2
22	Titanium (Ti)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$	4	2, 8, 10, 2
23	Vanadium (V)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3\ 4s^2$	5, 4	2, 8, 11, 2
24	Chromium (Cr)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$	2	2, 8, 13, 1
25	Manganese (Mn)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$	7, 4, 2	2, 8, 13, 2
26	Iron (Fe)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$	2, 3	2, 8, 14, 2
27	Cobalt (Co)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$	3, 2	2, 8, 15, 2
28	Nickel (Ni)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$	2	2, 8, 16, 2
29	Copper (Cu)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$	2, 1	2, 8, 18, 1
30	Zinc (Zn)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$	2	2, 8, 18, 2

Writing Chemical Formula of a Compound

Aluminium Chloride: $AlCl_3$



Carbon Tetrachloride: CCl_4



CONCEPT OF ATOMICITY

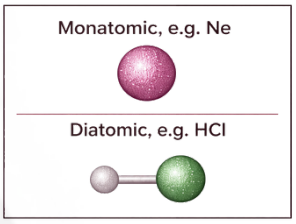
Atomicity

- Atomicity is the number of atoms present in a single molecule of an element or compound.

Diatomic Molecules

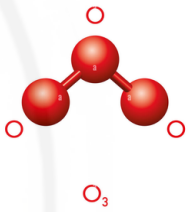
- The substances that have two atoms in their molecules are called diatomic molecules.
- There are only 7 elements that form stable homonuclear diatomic molecules and mostly exists as gases.
- **They are:** Oxygen, Nitrogen, Hydrogen, Chlorine, and Fluorine, Bromine, Iodine. So diatomic is atomicity of non-metals too.

Monoatomic Molecules



- The substances that are made up of single atom and lacks any covalent bond are called monoatomic molecules.
- **E.g.** Helium, Neon, Argon, Krypton, Xenon, Radon

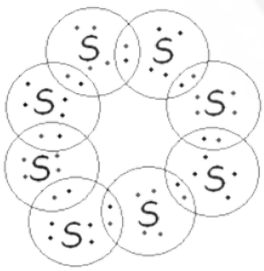
Triatomic Molecules

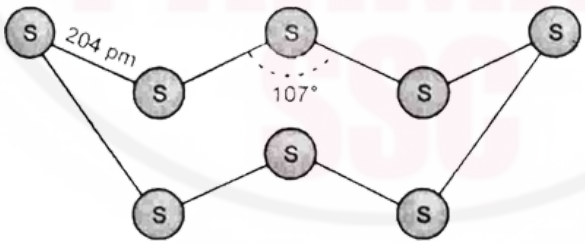


- The substances formed of molecules that are made up of three atoms are called triatomic molecules.
- **E.g.** Water, Carbon dioxide, Hydrogen cyanide.

Polyatomic Molecules

- All the molecules that are made up of four or more atoms are known as polyatomic molecules.
- **E.g.** One molecule of sulphur is made up of 8 molecules, and one molecule of phosphorus is made up of 4 molecules. So both sulphur and phosphorus are polyatomic molecules.





- **Ring structure of S_8 molecules**
- **Crown shape of S_8 molecule**



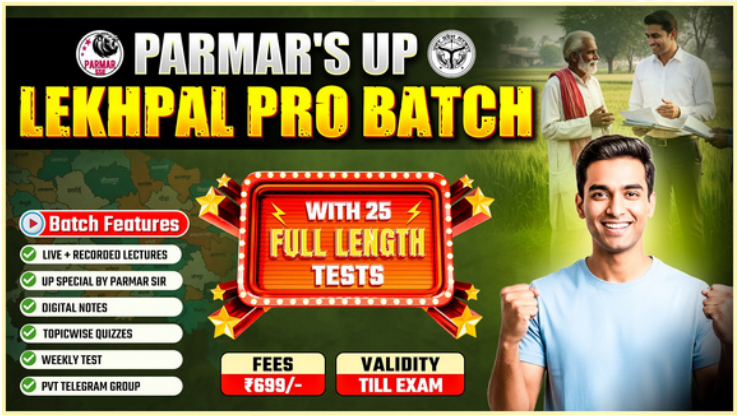


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