INTRODUCTION TO PERIODIC CHEMISTRY

Periodic chemistry involves the study of how **physical** and **chemical properties** of elements **vary** with their different **electron arrangement**. Atomic number is the number of protons in the nucleus of an atom and it equal to the number of electrons in a neutral atom.

Atoms of different element have different atomic number and hence have different number of electrons. Different number of electrons means different electron arrangements. Hence, elements with different electron arrangement have different properties and element with similar electron arrangement have similar properties.

ARRANGEMENT OF ELECTRON IN THE SHELL

1. Electrons in an atom can **orbit** in a maximum of seven shells or energy levels. The shells are represented by K,L,M,N,O,P and Q.
2. A shell is a volume of space around the nucleus in which there is high probability of locating and electron.
3. The maximum number of electrons that a shell can accommodate is known by using the formula $\mathbf{2n^2}$ where n is the shell number.
4. The maximum number of electrons possible in an outermost shell of an atom is 8
5. The electrons in the outermost shell are called **valence electrons** and the shell is called **valence shell.**
6. The **valence electrons** are responsible for the chemical properties of elements.
7. **Core electrons (inner electrons)** are electrons between the nucleus and the valence electrons. They do not participate in chemical reactions but do influence the behaviour of valence electron.

ELEMENT AND THEIR SYMBOLS

Name of the Element	Symbol of the Element	Atomic Number
Hydrogen	H	1
Helium	He	2
Lithium	Li	3
Beryllium	Be	4
Boron	B	5
Carbon	C	6
Nitrogen	N	7
Oxygen	O	8
Fluorine	F	9
Neon	Ne	10
Sodium	Na	11

Magnesium	Mg	12
Aluminium	Al	13
Silicon	Si	14
Phosphorus	P	15
Sulfur	S	16
Chlorine	Cl	17
Argon	Ar	18
Potassium	K	19
Calcium	Ca	20
Scandium	Sc	21
Titanium	Ti	22
Vanadium	V	23
Chromium	Cr	24

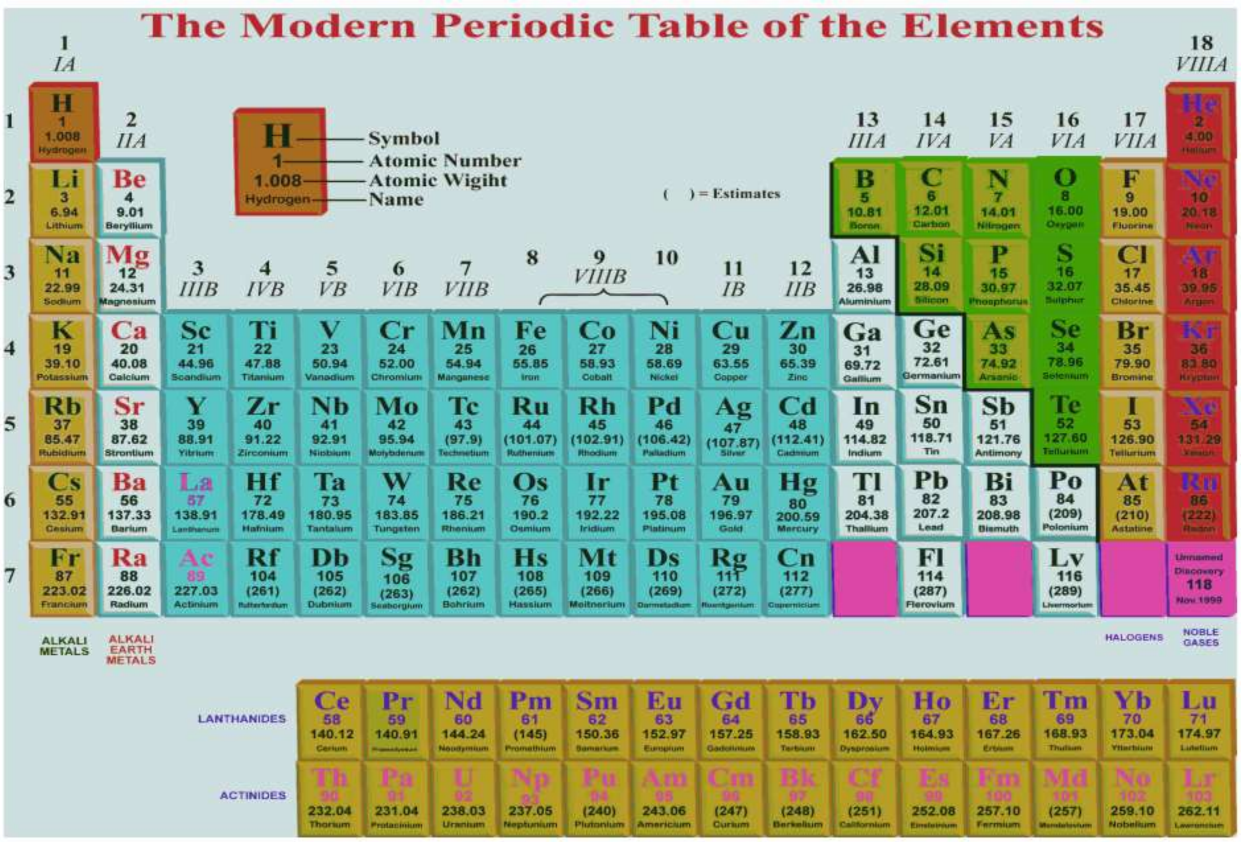
Manganese	Mn	25
Iron	Fe	26
Cobalt	Co	27
Nickel	Ni	28
Copper	Cu	29
Zinc	Zn	30
Gallium	Ga	31
Germanium	Ge	32
Arsenic	As	33
Selenium	Se	34
Bromine	Br	35
Krypton	Kr	36
Rubidium	Rb	37

Strontium	Sr	38
Yttrium	Y	39
Zirconium	Zr	40
Niobium	Nb	41
Molybdenum	Mo	42
Technetium	Tc	43
Ruthenium	Ru	44
Rhodium	Rh	45
Palladium	Pd	46
Silver	Ag	47

PERIODIC LAW: it states that the properties of element are a periodic function of their atomic number.

THE PERIODIC TABLE

- ✓ This is a table showing all the known elements arrange systematically in order of increasing atomic number such that their outer electron configuration and recur at interval.
- ✓ Elements with the same number of valence electron constitute a group.



Characteristics of the groups

GROUP ONE (I) ELEMENT (**Alkali Metals)**

- Less dense than other metals
- One loosely bound valence electron
- Highly reactive, with reactivity increasing moving down the group

- The largest atomic radius of elements in their period
- Low ionization energy
- Low electronegativity
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GROUP TWO (II) ELEMENT (**Alkaline Earth Metals)**

- Two electrons in the valence shell
- Readily form divalent cations
- Low electron affinity
- Low electronegativity

Transition Metals

The lanthanides (rare earth) and actinides are also transition metals. The basic metals are similar to transition metals but tend to be softer and to hint at nonmetallic properties. In their pure state, all of these elements tend to have a shiny, metallic appearance. While there are radioisotopes of other elements, all of the actinides are radioactive.

- Very hard, usually shiny, ductile, and malleable
- High melting and boiling points
- High thermal and electrical conductivity
- Form cations (positive oxidation states)
- Tend to exhibit more than one oxidation state
- Low ionization energy

Metalloids or Semimetals

- Electronegativity and ionization energy intermediate between that of metals and nonmetals
- May possess a metallic luster
- Variable density, hardness, conductivity, and other properties
- Often make good semiconductors
- Reactivity depends on the nature of other elements in the reaction

Non-metals

The halogens and noble gases are nonmetals, although they have their own groups, too.

- High ionization energy
- High electronegativity
- Poor electrical and thermal conductors
- Form brittle solids
- Little if any metallic luster
- Readily gain electrons

GROUP SEVEN (VII) ELEMENTS **(Halogens)**

The halogens exhibit different physical properties from each other but do share chemical properties.

- Extremely high electronegativity
- Very reactive
- Seven valence electrons, so elements from this group typically exhibit a -1 oxidation state

GROUP EIGHT (VIII) ELEMENTS **(Noble Gases)**

The noble gasses have complete valence electron shells, so they act differently. Unlike other groups, noble gasses are unreactive and have very low electronegativity or electron affinity.

PERIODIC PROPERTIES

A periodic property is a periodic variation in physical and chemical properties of elements with increasing in atomic number OR it is any property of elements (either physical or chemical) which changes regularly with increase in atomic number. Examples of periodic properties to be considered for study are,

i. Atomic radius or ionic radius
ii. Electronegativity
iii. Ionization energy
iv. Electron affinity

ATOMIC RADIUS OR SIZE

Atomic radius is half the internuclear distance between two Identical atoms in a covalent bond. Size of an atom depends on

a. Number of protons (i.e nuclear charge)

b. Number of core shells

Atomic radius depends on how strongly the proton in the nucleus attract or pull the outermost electrons. Given two atoms with the same number of core shell, the one with greater number of protons will have greater pull of its outermost electrons to the nucleus leading to smaller size. The Inner or core electrons in the core shells which are negatively charged tend to shield or screen the outermost shell electrons from the full impact of attraction by the positively charged protons in the nucleus. The attraction by the positive protons (also called nuclear charge) for electrons in the outermost shell is usually reduced due to the presence of the core electrons in the core shells which cause **screening** or **shielding effects.**

Effective nuclear charge is therefore is the **net positive** charge which is felt by electrons in the outermost shell through attraction by the positive proton (i.e effective nuclear charge = Nuclear charge – Shielding effects). *Atomic radii decrease across the period from Lithium to Fluorine and increase down the group from Lithium to Rubidium.*

Example: arrange the following in order of increasing size K, Li, and Na. give reason for your order.

IONIC RADIUS OR SIZE

CATIONS: positive ion or cation has smaller ionic radius than its original neutral atom.

ANIONS: A negative ion or anion has a larger radius than that of the parent atom.

IONIZATION ENERGY

It is the minimum energy required to remove an electron from the outermost shell of a free or isolated gaseous atom to form a gaseous singly charge cation. *Generally, ionization energy increase across the period from left to right and decrease down the group.*

Factors that influence ionization energy

a. Atomic radius or distance of the outermost electron to the nucleus.
b. Stability of the electron configuration
c. Shielding effects of inner electrons

d. Size of the positive nuclear charge

Example: Arrange the following elements in decreasing order of first ionization energies of their atoms. Give reason for your answer. Na, Cl, and K.

ELECTRON AFFINITY E.A.

First electron affinity is the energy lost or gained when 1 mole of atoms in the free gaseous state gain 1 mole of electrons into the outermost shells to form 1 mole of gaseous singly charge anions. Generally, electron affinity increases across the period from left to right and decrease down the group.

Factors influencing EA.

a. Atomic size or effective nuclear charge
b. Stability of the electron configuration
c. Charge carried by the anion

ELECTRONEGATIVITY

It is the ability or power of an atom to attract or pull the shared pair of electrons to itself when it is in a bonded molecule. The smaller the atomic size the higher the electronegativity. Electronegativity therefore increases across the period from left to right and decrease down the group.

TO BE CONTINUE……