

APPOLO STUDY CENTRE

Periodic table with properties

TEST 8

9th STD

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Unit- 12- Periodic Classification of Elements

Introduction

- We live in the world of substances with great diversity. Substances are formed by the combination of various elements. All the elements are unique in their nature and property. To categorize these elements according to their properties, scientists started to look for a way. In 1800, there were only 31 known elements. By 1865, their number became 63. Now 118 elements have been discovered. As different elements were being discovered, scientists gathered more and more information about the properties of these elements. They found it difficult to organize all that was known about the elements. They started looking for some pattern in their properties, on the basis of which they could study such a large number of elements with ease. Let us discuss the concepts of classification of elements proposed by various scientists from early to modern period.

Early Concepts of Classification of Elements

Dobereiner's Triads

- In 1817, Johann Wolfgang Dobereiner, a German chemist, suggested a method of grouping elements based on their relative atomic masses. He arranged the elements into groups containing three elements each. He called these groups as 'triads' (tri - three).

- Dobereiner showed that when the three elements in a triad are arranged in the ascending order of their atomic masses, the atomic mass of the middle element is nearly the same as average of atomic masses of other two elements. This statement is called the Dobereiner's law of triads. Table 12.1 shows the law of triads proposed by Dobereiner.
- Example: In the triad group (1), arithmetic mean of atomic masses of 1st and 3rd elements, $(6.9 + 39.1)/2 = 23$. So the atomic mass of Na (middle element) is 23.

Triad Group (1)		Triad Group (2)		Triad Group (3)	
Element	Atomic Mass	Element	Atomic Mass	Element	Atomic Mass
Li	6.9	Cl	35.5	Ca	40.1
Na	23	Br	79.9	Sr	87.6
K	39.1	I	126.9	Ba	137.3

Limitations:

- v Dobereiner could identify only three triads from the elements known at that time and all elements could not be classified in the form of triads.
- v The law was not applicable to elements having very low and very high atomic mass.

Newlands' Law of Octaves

- In 1866, John Newlands arranged 56 known elements in the increasing order of their atomic mass. He observed that every eighth element had properties similar to those of the first element like the eighth note in an octave of music is similar to the first. This arrangement was known as 'law of octaves'.
- The octave of Indian music system is sa, re, ga, ma, pa, da, ni, sa. The first and last notes of this octave are same i.e. sa. Likewise, in the Newlands' table of octaves, the element 'F' is eighth from the element 'H', thus they have similar properties.

Limitations:

- v There are instances of two elements being fitted into the same slot, e.g. cobalt and nickel.
- v Some elements, totally dissimilar in their properties, were fitted into the same group. (Arrangement of Co, Ni, Pd, Pt and Ir in the row of halogens)

- v The law of octaves was not valid for elements that had atomic masses higher than that of calcium.
- v Newlands' table was restricted to only 56 elements and did not leave any room for new elements.
- v Discovery of inert gases (Neon, Argon....) at later stage made the 9th element similar to the first one. Eg: Neon between Fluorine and Sodium.

Mendeleev's Periodic Table

- In 1869, Russian chemist, Dmitri Mendeleev observed that the elements of similar properties repeat at regular intervals when the elements are arranged in the order of their atomic masses. Based on this, he proposed the law of periodicity which states that "the physical and chemical properties of elements are the periodic functions of their atomic masses". He arranged 56 elements known at that time according to his law of periodicity. This was best known as the short form of periodic table.

Features of Mendeleev's Periodic Table:

- v It has eight vertical columns called 'groups' and seven horizontal rows called 'period'.
- v Each group has two subgroups 'A' and 'B'. All the elements appearing in a group were found to have similar properties.
- v For the first time, elements were comprehensively classified in such a way that elements of similar properties were placed in the same group
- v It was noticed that certain elements could not be placed in their proper groups in this manner. The reason for this was wrongly determined atomic masses. Consequently those wrong atomic masses were corrected. Eg: The atomic mass of beryllium was known to be 14. Mendeleev reassessed it as 9 and assigned beryllium a proper place.
- v Columns were left vacant for elements which were not known at that time and their properties also were predicted. This gave motivation to experiment in Chemistry. Eg: Mendeleev gave names EkaAluminium and Eka Silicon to those elements which were to be placed below Aluminium and Silicon respectively in the periodic table and predicted their properties. The discovery of Germanium later on, during his life time, proved him correct.

No.	No.	No.	No.	No.	No.	No.	No.	No.	No.
H 1	F 8	Cl 15	Co & Ni 22	Br 29	Pd 36	I 42	Pt & Ir 50		
Li 2	Na 9	K 16	Cu 23	Rb 30	Ag 37	Cs 44	Os 51		
G 3	Mg 10	Ca 17	Zn 24	Sr 31	Cd 38	Ba & V 45	Hg 52		
Bo 4	Al 11	Cr 19	Y 25	Ce & La 33	U 40	Ta 46	Ti 53		
C 5	Si 12	Ti 18	In 26	Zr 32	Sn 39	W 47	Pb 54		
N 6	P 13	Mn 20	As 27	Di & Mo 34	Sb 41	Nb 48	Bi 55		
O 7	S 14	Fe 21	Se 28	Ro & Ru 35	To 43	Au 49	Th 56		

Limitations:

- Elements with large difference in properties were included in the same group. Eg: Hard metals like copper (Cu) and silver (Ag) were included along with soft metals like sodium (Na) and potassium (K).
- No proper position could be given to the element hydrogen. Non-metallic hydrogen was placed along with metals like lithium (Li), sodium (Na) and potassium (K).
- The increasing order of atomic mass was not strictly followed throughout. Eg. Co & Ni, Te & I.
- No place for isotopes in the periodic table.

Properties of Germanium

Property	Mendeleev's prediction (1871)	Actual property (1886)
Atomic Mass	About 72	72.59
Specific Gravity	5.5	5.47
Colour	Dark grey	Dark grey
Formula of Oxide	EsO ₂	GeO ₂
Nature of Chloride	EsCl ₄	GeCl ₄

Modern Periodic Table

- In 1913, the English Physicist Henry Moseley, through his X-ray diffraction experiments, proved that the properties of elements depend on the atomic number and not on the atomic mass. Consequently, the modern periodic table was prepared by arranging elements in the increasing order of their atomic number.
- This modern periodic table is the extension of the original Mendeleev's periodic table and known as the long form of periodic table.

Modern Periodic Law

- Atomic number of an element (Z) indicates the number of protons (positive charge) or the number of electrons (negative charge). The physical and chemical properties of elements depend not only on the number of protons but also on the number of electrons and their arrangements (electronic configuration) in atoms. Hence, the modern periodic law can be stated as follows: "The chemical and physical properties of elements are periodic functions of their atomic numbers". Based on the modern periodic law, the modern periodic table is derived.

Features of Modern Periodic Table

- v All the elements are arranged in the increasing order of their atomic number.
- v The horizontal rows are called periods. There are seven periods in the periodic table.
- v The elements are placed in periods based on the number of shells in their atoms.
- v Vertical columns in the periodic table starting from top to bottom are called groups. There are 18 groups in the periodic table.
- v Based on the physical and chemical properties of elements, they are grouped into various families.

Groups in modern periodic table

Group	Families
1	Alkali metals
2	Alkaline earth metals
3 to 12	Transition metals
13	Boron Family
14	Carbon Family
15	Nitrogen Family
16	Oxygen Family (or) Chalcogen Family
17	Halogens
18	Noble gases

Classification of elements into blocks

- We know that the electrons in an atom are accommodated in shells around the nucleus. Each shell consists of one or more subshells in which the electrons are distributed in certain manner. These subshells are designated as s, p, d, and f. Based on the arrangement of electrons in subshells, the elements of periodic table are classified into four blocks namely s, p, d and f blocks.
- **s-Block Elements:** It includes group 1 (alkali metals) and group 2 (alkaline earth metals) elements. They are also called as representative elements. The elements of group 1 (except hydrogen) are metals. They react with water to

form solutions that change the colour of a vegetable dye from red to blue. These solutions are said to be highly alkaline or basic. Hence they are called alkali metals.

- The elements of group 2 are also metals. They combine with oxygen to form oxides, formerly called 'earths', and these oxides produce alkaline solutions when they are dissolved in water. Hence, these elements are called alkaline earth metals.
- **(2) p-Block Elements:** These elements are in group 13 to 18 in the periodic table. They include boron, carbon, nitrogen, oxygen, fluorine families in addition to noble gases (Except helium). They are also called as representative elements. The p-block is home to the biggest variety of elements and is the only block that contains all three types of elements: metals, nonmetals, and metalloids.
- **(3) d-Block Elements:** It includes group 3 to group 12 elements. They are found in the centre of the periodic table. Their properties are intermediate to that of s block and p block elements and so they are called transition elements.
- **(4) f - Block Elements:** It includes 14 elements after (Lanthanum) La (57), called Lanthanoides and 14 elements after (Actinium) Ac (89), called Actinoides. They are placed at the bottom of the periodic table. They are also called as inner Transition elements.

Advantages of the Modern Periodic Table

- ✓ The table is based on a more fundamental property i.e., atomic number.
- ✓ It correlates the position of the element with its electronic configuration more clearly.
- ✓ The completion of each period is more logical. In a period, as the atomic number increases, the energy shells are gradually filled up until an inert gas configuration is reached.
- ✓ It is easy to remember and reproduce.
- ✓ Each group is an independent group and the idea of subgroups has been discarded.
- ✓ One position for all isotopes of an element is justified, since the isotopes have the same atomic number.
- ✓ The position of the eighth group (in Mendeleev's table) is also justified in this table. All transition elements have been brought in the middle as the properties of transition elements are intermediate between left portion and right portion elements of the periodic table.
- ✓ The table completely separates metals from nonmetals. The nonmetals are present in upper right corners of the periodic table.

- v The positions of certain elements which were earlier misfit (interchanged) in the Mendeleev's periodic table are now justified because it is based on atomic number of the elements.
- v Justification has been offered for placing lanthanides and actinides at the bottom of the periodic table.

Position of hydrogen in the periodic table

- Hydrogen is the lightest, smallest and first element of the periodic table. Its electronic configuration ($1s^1$) is the simplest of all the elements. It occupies a unique position in the periodic table. It behaves like alkali metals as well as halogens in its properties.
- In the periodic table, it is placed at the top of the alkali metals.
 - (1) Hydrogen can lose its only electron to form a hydrogen ion (H^+) like alkali metals.
 - (2) It can also gain one electron to form the hydride ion (H^-) like halogens.
 - (3) Alkali metals are solids while hydrogen is a gas.
- Hence the position of hydrogen in the modern periodic table is still under debate as the properties of hydrogen are unique.

Position of Rare Gases

- The elements Helium, Neon, Argon, Krypton, Xenon and Radon of group 18 in the periodic table are called as Noble gases or Rare gases. They are monoatomic gases and do not react with other substances easily, due to completely filled subshells. Hence they are called as inert gases. They are found in very small quantities and hence they are called as rare gases.

Metals, Non-Metals and Metalloids

Metals

- Metals are typically hard, shiny, malleable (can be made as sheet), fusible and ductile (can be drawn into wire) with good electrical and thermal conductivity. Except mercury, most of the metals are solids at room temperature. Metals occupy larger area in the periodic table and are categorized as:
 - (1) Alkali metals. e.g. Lithium to Francium (top to bottom)
 - (2) Alkaline earth metals. e.g. Beryllium to Radium (top to bottom)
 - (3) Transition Metals. Group III B to II A
 - (4) P-Block metals. e.g. Al, Ga, In, Tl, Sn, Pb and Bi.

Non-metals

- A non-metal is an element that does not have the characters like hardness, shiny, malleable, suitable and ductile. In other words, a non-metal is an element that does not have the properties of metal. e.g. All non metals are arranged in P-Block only. P-Block non metals: C, N O, P, S, Se, Halogen (F, Cl, Br and I) and inert gases (He to Rn).

Metalloids

- Elements which have the properties of both metals and non-metals are called as metalloids. (eg) Boron, Arsenic.

Alloys

- During 3500 BC(BCE), people used an alloy named 'bronze'. The idea of making an alloy was quite old. The majority of the metallic substances used today are alloys. Alloys are mixtures of two or more metals and are formed by mixing molten metals thoroughly. Rarely nonmetals are also mixed with metals to produce alloys.
- It is generally found that alloying produces a metallic substance that has more useful properties than the original pure metals from which it is made. For example, the alloy brass is made from copper and zinc.

Advantages of alloys

- Alloys do not get corroded or get corroded to very less extent.
- They are harder and stronger than pure metals (Example: Gold is mixed with copper and it is harder than pure gold).
- They have less conductance than pure metals (Example: Copper is good conductor of heat and electricity whereas brass and bronze are not good conductors).
- Some alloys have lower melting point than pure metals (Example: Solder is an alloy of lead and tin which has lower melting point than each of the metals).
- When metal is alloyed with mercury, it is called amalgam.

10th STD
7.ATOMS AND MOLECULES

INTRODUCTION

- You have learnt, in your lower classes that matter is around us everywhere. Matter is made of atoms. Curiously the idea of atom was first proposed by the Greek philosophers in the fifth century BC (BCE). But, their theory was more philosophical than scientific.
- The first scientific theory of the atom was proposed by John Dalton. Few of the postulates of Dalton's theory about an atom were found incorrect by the later on studies made by J.J. Thomson, Rutherford, Neils Bohr and Schrodinger. In the light of the result of the researches most of the limitations of the Dalton's theory wereremoved and a new theory known as the modern atomic theory was put forward. **'The main postulates of modern atomic theory'** are as follows:
 1. An atom is no longer indivisible (after the discovery of the electron, proton, and neutron).
 2. Atoms of the same element may have different atomic mass. (discovery of isotopes $_{17}\text{Cl}^{35}$, $_{17}\text{Cl}^{37}$).
 3. Atoms of different elements may have same atomic masses (discovery of Isobars $_{18}\text{Ar}^{40}$, $_{20}\text{Ca}^{40}$).
 4. Atoms of one element can be transmuted into atoms of other elements. In other words, atom is no longer indestructible (discovery of artificial transmutation).
 5. Atoms may not always combine in a simple whole number ratio (E.g. Glucose $\text{C}_6\text{H}_{12}\text{O}_6$ C:H:O = 6:12:6 or 1:2:1 and Sucrose $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ C:H:O = 12:22:11).
 6. Atom is the **smallest particle that takes part in a chemical reaction.**
 7. The mass of an atom can be converted into energy ($E = mc^2$).
- The modern atomic theory is the basis for all the studies of chemical and physical processes that involve atoms. You have studied the most fundamental ideas about an atom in your lower classes. Let us discuss some more concepts about atoms in this lesson.

ATOM AND ATOMIC MASS

- As you know, anything that has mass and occupies space is called matter. Atoms are the building blocks of matter. Since matter has mass, it must be

due to its atoms. According to the modern atomic theory, an atom contains subatomic particles such as protons, neutrons and electrons. Protons and neutrons have considerable mass, but electrons don't have such a considerable mass. Thus, the mass of an atom is mainly contributed by its protons and neutrons and hence the sum of the number of protons and neutrons of an atom is called its mass number.

- Individual atoms are very small and it is difficult to measure their masses. You can measure the mass of macroscopic materials in gram or kilogram. The mass of an atom is measured in atomic mass unit (amu).
- Atomic mass unit is one-twelfth of the mass of a carbon-12 atom; an isotope of carbon, which contains 6 protons and 6 neutrons.
- (Note: The symbol 'amu' is no longer used in the modern system and instead, it uses the symbol 'u' to denote unified atomic mass. The mass of a proton or neutron is approximately 1 amu).

Relative Atomic Mass (RAM)

- As an atom is very small, its absolute mass cannot be determined directly. The early pioneers of chemistry used to measure the atomic mass of an atom relative to an atom of another element. They measured the masses of equal number of atoms of two or more elements at a time, to determine their relative masses. They established one element as a standard, gave it an arbitrary value of atomic mass and using this value they measured the relative mass of other elements. The mass obtained by this way is called relative atomic mass. In the beginning, the mass of hydrogen atom was chosen as a standard and masses of other atoms were compared with it, because of the existence of isotopic character of hydrogen (${}_1\text{H}^1$, ${}_1\text{H}^2$, ${}_1\text{H}^3$). Later hydrogen atom was replaced by oxygen atom as the standard. Now, the stable isotope of carbon (C-12) with atomic mass 12 is used as the standard for measuring the relative atomic mass of an element.
- Relative atomic mass of an element is the ratio between the average mass of its isotopes to $\frac{1}{12^{\text{th}}}$ part of the mass of a carbon-12 atom. It is denoted as Ar. It is otherwise called "Standard Atomic Weight".

Relative Atomic Mass

$$(\text{Ar}) = \frac{\text{Average mass of the isotopes of the element}}{\frac{1}{12^{\text{th}}} \text{ of the mass of one Carbon-12 atom}}$$

- Modern methods of determination of atomic mass by Mass Spectrometry uses C-12 as standard. For most of the elements, the relative atomic mass is very closer to a whole number and it is rounded off to a whole number, to make calculations easier.

Element	Symbol	A_r
Hydrogen	H	1
Carbon	C	12
Nitrogen	N	14
Oxygen	O	16
Sodium	Na	23
Magnesium	Mg	24
Sulphur	S	32

Relative Atomic Mass is only a ratio, so it has no unit. If the atomic mass of an element is expressed in grams, it is called as **Gram Atomic Mass**

Gram Atomic Mass of hydrogen = 1 g

Gram Atomic Mass of carbon = 12 g

Gram Atomic Mass of nitrogen = 14 g

Gram Atomic Mass of oxygen = 16 g

Average Atomic Mass (AAM)

- How can one measure the atomic mass of an element? It is somewhat more complicated because most of the naturally occurring elements exist as a mixture of isotopes, each of which has its own mass. Thus, it is essential to consider this isotopic mixture while calculating the atomic mass of an element.
- The average atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes.
- But, the abundance of isotopes of each element may differ. So, the abundance of all these isotopes are taken into consideration while calculating the atomic mass. Then, what do we mean by a weighted average? Let us consider an element which exists as a mixture of 50% of an isotope having a mass of 9 amu, and 50% of another isotope having a mass of 10 amu. Then, its average atomic mass is calculated by the following equation:

Average atomic mass

$$= (\text{Mass of 1st isotope} \times \% \text{ abundance of 1st isotope}) + (\text{Mass of 2nd isotope} \times \% \text{ abundance of 2nd isotope})$$

Thus, for the given element the average

- $$\text{atomic mass} = \left(9 \times \frac{50}{100}\right) + \left(10 \times \frac{50}{100}\right)$$

$$= 4.5 + 5 = 9.5 \text{ amu}$$
- (Note: In the calculations involving percentages, you need to convert percentage abundance into fractional abundance. For example, 50 percent is converted into 50/100 or 0.50 as shown in the a foresaid calculation.)
- The atomic masses of elements, given in the periodic table, are average atomic masses. Sometimes, the term atomic weight is used to mean average atomic mass. It is observed, from the periodic table that atomic masses of most of the elements are not whole numbers. For instance, the atomic mass of carbon given in the periodic table is 12.01 amu, not 12.00 amu. The reason is that while calculating the atomic mass of carbon, both of its natural isotopes such as carbon-12. and carbon-13 are considered. The natural abundance of C-12 and C-13 are 98.90 % and 1.10 % respectively. The average of the atomic mass of carbon is calculated as follows:

Average atomic mass of carbon

$$= \left(12 \times \frac{98.9}{100}\right) + \left(13 \times \frac{1.1}{100}\right)$$

$$= (12 \times 0.989) + (13 \times 0.011)$$

$$= 11.868 + 0.143 = 12.011 \text{ amu}$$

- So it is important to understand that if it is mentioned that the atomic mass of carbon is 12 amu, it refers to the average atomic mass of the carbon isotopes, not the mass of the individual atoms of carbon.

Atomic Number	Name	Symbol	Atomic Mas (amu)
1	Hydrogen	H	1.008
2	Helium	He	4.003
3	Lithium	Li	6.941
4	Beryllium	Be	9.012

5	Boron	B	10.811
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Calculation of average atomic mass - Solved Examples

- Example 1: Oxygen is the most abundant element in both the Earth's crust and the human body. It exists as a mixture of three stable isotopes in nature.

Isotope	Mass (amu)	% abundance
${}^8\text{O}^{16}$	15.9949	99.757
${}^8\text{O}^{17}$	16.9991	0.038
${}^8\text{O}^{18}$	17.9992	0.205

The atomic mass of

$$\begin{aligned} \text{oxygen} &= (15.9949 \times 0.99757) + (16.9991 \times 0.00038) + (17.9992 \times 0.00205) \\ &= 15.999 \text{ amu.} \end{aligned}$$

- Example 2: Boron naturally occurs as a mixture of boron-10 (5 protons + 5 neutrons) and boron-11 (5 protons + 6 neutrons) isotopes. The percentage abundance of B-10 is 20 and that of B-11 is 80. Then, the atomic mass of boron is calculated as follows:

Atomic mass of

$$\text{boron} = \left(10 \times \frac{20}{100} \right) + \left(11 \times \frac{80}{100} \right)$$

$$= (10 \times 0.20) + (11 \times 0.80)$$

$$= 2 + 8.8$$

$$= 10.8 \text{ amu}$$

MOLECULE AND MOLECULAR MASS

- Except noble gases, atoms of most of the elements are found in the combined form with itself or atoms of other elements. It is called as a molecule. A molecule is a combination of two or more atoms held together by strong chemical forces of attraction, i.e. chemical bonds.

Classification of molecules

- A molecule may contain atoms of the same element or may contain atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions. Thus, a molecule may be an element or a compound. If the molecule is made of similar kind of atoms, then it is called homoatomic molecule.
- The molecule that consist of atoms of different elements is called heteroatomic molecule. A compound is a heteroatomic molecule. The number of atoms present in the molecule is called its 'atomicity'.
- Let us consider oxygen. Oxygen gas exists in two allotropic forms: Oxygen (O_2) and Ozone (O_3). In oxygen molecule, there are two oxygen atoms. So its atomicity is two. Since both the atoms are similar, oxygen (O_2) is a homodiatomic molecule. Other elements that exist as diatomic molecules are hydrogen (H_2), nitrogen (N_2) and halogens: fluorine (F_2), chlorine (Cl_2), bromine (Br_2) and iodine (I_2).

Element	No. of Protons	No. of Neutrons	Mass Number	Stable Isotopes (abundance)	Atomic Mass (amu)
	7			N-14 (99.6 %)	
		8		N-15 (0.4 %)	
Sulphur	14		28	S-28 (92.2 %)	
	14			S-29 (4.7 %)	
		16		S-30 (3.1 %)	
	17			Cl-35 (75 %)	
	17			Cl-37 (25 %)	

- Ozone (O_3) contains three oxygen atoms and hence it is called homotriatomic molecule. If a molecule contains more than three atoms, then it is called polyatomic molecule.
- Consider hydrogen chloride. It consists of two atoms, but of different elements, i.e. hydrogen and chlorine. So, its atomicity is two. It is a heterodiatomic molecule. Similarly, the water molecule contains two hydrogen atoms and one oxygen atom. So its atomicity is three. It is a heterotriatomic molecule.

Classify the following molecules based on their atomicity and fill in the table:

- Fluorine (F₂), Carbon dioxide (CO₂), Phosphorous (P₄), Sulphur (S₈), Ammonia (NH₃), Hydrogen iodide (HI), Sulphuric Acid (H₂SO₄), Methane (CH₄), Glucose (C₆H₁₂O₆), Carbon monoxide (CO)

Relative Molecular Mass (RMM)

- As the molecules are made of atoms, they also have their own mass. The mass of the molecule of an element or compound is measured in the C-12 scale and hence called relative molecular mass.
- The Relative Molecular Mass of a molecule is the ratio between the mass of one molecule of the substance to $\frac{1}{12}^{\text{th}}$ mass of an atom of Carbon -12.

Relative Molecular Mass is only a ratio. So, it has no unit. If the molecular mass of a compound is expressed in grams, it is called Gram Molecular Mass.

Gram Molecular Mass of water = 18 g

Gram Molecular Mass of carbon dioxide = 44 g

Gram Molecular Mass of ammonia = 17 g

Gram Molecular Mass of HCl = 36.5 g

- The relative molecular mass is obtained by adding together the relative atomic masses of all the atoms present in a molecule.

Calculation of relative molecular mass - Solved examples:

- **Example 1:** Relative molecular mass of sulphuric acid (H₂SO₄) is calculated as follows: Sulphuric acid contains 2 atoms of hydrogen, 1 atom of sulphur and 4 atoms of oxygen.

$$\begin{aligned} \text{Therefore, Relative molecular mass of sulphuric acid} \\ &= (2 \times \text{mass of hydrogen}) + (1 \times \text{mass of sulphur}) + (4 \times \text{mass of oxygen}) \\ &= (2 \times 1) + (1 \times 32) + (4 \times 16) \\ &= 98 \end{aligned}$$

- i.e., one molecule of H₂SO₄ is 98 times as heavy as $\frac{1}{12}^{\text{th}}$ of the mass of a carbon -12.

- **Example 2:** Relative molecular mass of water (H₂O) is calculated as follows: A water molecule is made of 2 atoms of hydrogen and one atom of oxygen.

$$\begin{aligned} \text{So, the relative molecular mass of water} \\ &= (2 \times \text{mass of hydrogen}) + (1 \times \text{mass of oxygen}) \\ &= (2 \times 1) + (1 \times 16) \\ &= 18 \end{aligned}$$

- i.e., one molecule of H₂O is 18 times as heavy as $\frac{1}{12^{th}}$ of the mass of a carbon - 12.

DIFFERENCE BETWEEN ATOMS AND MOLECULES

- Even though atoms are the basic components of molecules, they differ in many aspects when compared to the molecules. Table 7.5 consolidates the major difference between atoms and molecules.

Atom	Molecule
An atom is the smallest particle of an element	A molecule is the smallest particle of an element or compound.
Atom does not exist in free state except in a noble gas	Molecule exists in free a state
Except some of noble gas, other atoms are highly reactive	Molecules are less reactive
Atom does not have a chemical bond	Atoms in a molecule are held by chemical bonds

MOLE CONCEPT

- So far we discussed about matters in terms of individual atoms and molecules. Atomic mass units provide a relative scale for the masses of the elements. Since the atoms have such small masses, no usable scale can be devised to weigh them in the calibrated units of atomic mass units. In any real situation, we deal with macroscopic samples containing enormous number of atoms. Therefore, it is convenient to have a special unit to describe a very large number of atoms. The idea of a 'unit' to denote a particular number of objects is not new. For example, the pair (2 items) and the dozen (12 items), are all familiar units. Chemists measure atoms and molecules in 'moles'. So, you can now understand that 'mole' denotes a number of particles.
- In the SI system, the mole (mol) is the amount of a substance that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in exactly 12 g (or 0.012 kg) of the carbon-12 isotope. The actual number of atoms in 12 g of carbon-12 is determined experimentally. This is called Avogadro's Number (N_A), named after an Italian scientist Amedeo Avogadro who proposed its significance. Its value is 6.023×10^{23} . So one mole of a substance contains 6.023×10^{23} entities. Thus, 5 moles of oxygen molecules contain $5 \times 6.023 \times 10^{23}$ molecules.

- Mole Concept: The study of the collection of particles by using mole as the counting unit, in order to express the mass and volume of such unit particles in a bulk of matter is known as mole concept.
- The number of moles of a substance can be calculated by various means depending on the data available, as follows:
 - v Number of moles of molecules.
 - v Number of moles of atoms.
 - v Number of moles of a gas (Standard molar volume at STP = 22.4 litre).
 - v Number of moles of ions

Mole of atoms:

- v One mole of an element contains 6.023×10^{23} atoms and it is equal to its gram atomic mass.
- v i.e., one mole of oxygen contains 6.023×10^{23} atoms of oxygen and its gram atomic mass is 16 g.

Mole of molecules:

- One mole of matter contains 6.023×10^{23} molecules and it is equal to its gram molecular mass.
- i.e., one mole of oxygen contains 6.023×10^{23} molecules of oxygen and its gram molecular mass is 32 g.

Molar volume:

- One mole of any gas occupies 22.4 litre or 22400 ml at S.T.P. This volume is called as molar volume.

Calculation of number of moles by Different modes

$$\begin{aligned}
 \text{Number of moles} &= \text{Mass} / \text{Atomic Mass} \\
 &= \text{Mass} / \text{Molecular mass} \\
 &= \text{Number of Atoms} / 6.023 \times 10^{23} \\
 &= \text{Number of Molecules} / 6.023 \times 10^{23}
 \end{aligned}$$

PERCENT COMPOSITION

- So far, we were dealing with the number of entities present in a given substance. But many times, the information regarding the percentage of a particular element present in a compound is required.

- The percentage composition of a compound represents the mass of each element present in 100 g of the compound.
- Let us understand the percentage composition of oxygen and hydrogen by taking the example of H₂O. It can be calculated using the formula

Mass % of an element

$$= \frac{\text{mass of that element in the compound}}{\text{molar mass of the compound}} \times 100$$

molar mass of H₂O = 2(1) + 16 = 18 g

$$\text{Mass \% of hydrogen} = \frac{2}{18} \times 100 = 11.11 \%$$

$$\text{Mass \% of oxygen} = \frac{16}{18} \times 100 = 88.89 \%$$

- This percentage composition is useful to determine the empirical formula and molecular formula.
- **Example 1:** Find the mass percentage composition of methane (CH₄).

Molar mass of CH₄ = 12 + 4 = 16 g

$$\text{Mass \% of carbon} = \frac{12}{16} \times 100 = 75 \%$$

$$\text{Mass \% of hydrogen} = \frac{4}{16} \times 100 = 25 \%$$

AVOGADRO HYPOTHESIS

- In 1811 Avogadro framed a hypothesis based on the relationship between the number of molecules present in equal volumes of gases in different conditions.
- The Avogadro's law states that "equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules"
- It follows that the volume of any given gas must be proportional to the number of molecules in it. If 'V' is the volume and 'n' is the number of molecules of a gas, then Avogadro law is represented, mathematically, as follows:

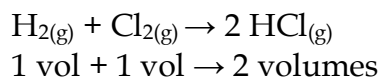
$$V \propto n$$

$$V = \text{constant} \times n$$

- Thus, one litre (1 dm³) of hydrogen contains the same number of molecules as in one litre of oxygen, i.e. the volume of the gas is directly proportional to the number of molecules of the gas.

Explanation

- Let us consider the reaction between hydrogen and chlorine to form hydrogen chloride gas



- According to Avogadro's law 1 volume of any gas is occupied by "n" number of molecules. n molecules + n molecules → 2n molecules

if n = 1 then

1 molecule + 1 molecule → 2 molecules.

½ molecule + ½ molecule → 1 molecule

- 1 molecule of hydrogen chloride gas is made up of ½ molecule of hydrogen and ½ molecule of chlorine. Hence, the molecules can be subdivided. This law is in agreement with Dalton's atomic theory.

APPLICATIONS OF AVOGADRO'S LAW

1. It explains Gay-Lussac's law.
2. It helps in the determination of atomicity of gases.
3. Molecular formula of gases can be derived using Avogadro's law
4. It determines the relation between molecular mass and vapour density.
5. It helps to determine gram molar volume of all gases (i.e, 22.4 litre at S.T.P)

RELATIONSHIP BETWEEN VAPOUR DENSITY AND RELATIVE MOLECULAR MASS

Relative molecular mass: (Hydrogen scale)

- The Relative Molecular Mass of a gas or vapour is the ratio between the mass of one molecule of the gas or vapour to mass of one atom of Hydrogen.

Vapour Density:

- Vapour density is the ratio of the mass of a certain volume of a gas or vapour, to the mass of an equal volume of hydrogen, measured under the same conditions of temperature and pressure.

Vapour Density (V.D.)

$$= \frac{\text{Mass of a given volume of gas or vapour at S.T.P.}}{\text{Mass of the same volume of hydrogen}}$$

- According to Avogadro's law, equal volumes of all gases contain equal number of molecules.
- Thus, let the number of molecules in one volume = n, then

$$= \frac{\text{Mass of 'n' molecules of a gas or vapour at S.T.P}}{\text{Mass of 'n' molecules of hydrogen}}$$

- Cancelling 'n' which is common, you get
V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{\text{Mass of 1 molecules of hydrogen}}$$

However, since hydrogen is diatomic

V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{\text{Mass of 2 atoms of hydrogen}}$$

- When you compare the formula of vapour density with relative molecular mass, they can be represented as

V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{2 \times \text{Mass of 1 atom of hydrogen}}$$

Relative molecular mass (hydrogen scale)

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at STP}}{\text{Mass of 1 atom of hydrogen}}$$

- You can therefore substitute the above equation to an Eqn 7.1 and arrive at the following formula

$$V.D. = \frac{\text{Relative molecular mass}}{2}$$

Now on cross multiplication, you have

$$2 \times \text{vapour density} = \text{Relative molecular mass of a gas (Or) Relative molecular mass} = 2 \times \text{Vapour density}$$

SOLVED PROBLEMS

I. Calculation of molar mass

- Calculate the gram molar mass of the following.
- H₂O
- CO₂
- Ca₃(PO₄)₂

Solution:

1) H₂O

Atomic masses of H = 1, O = 16

$$\text{Gram molar mass of H}_2\text{O} = (1 \times 2) + (16 \times 1) = 2 + 16$$

$$\text{Gram molar mass of H}_2\text{O} = 18 \text{ g}$$

2) CO₂

Atomic masses of C = 12, O = 16

$$\text{Gram molar mass of CO}_2 = (12 \times 1) + (16 \times 2) = 12 + 32$$

$$\text{Gram molar mass of CO}_2 = 44 \text{ g}$$

3) Ca₃(PO₄)₂

Atomic masses of Ca = 40, P = 30, O = 16.

$$\text{Gram molar mass of Ca}_3(\text{PO}_4)_2 = (40 \times 3) + [30 + (16 \times 4)] \times 2$$

$$= 120 + (94 \times 2) = 120 + 188$$

$$\text{Gram molar mass of Ca}_3(\text{PO}_4)_2 = 308 \text{ g}$$

Calculation based on number of moles from mass and volume

Calculate the number of moles in 46 g of sodium?

$$\text{Number of moles} = \frac{\text{Mass of the element}}{\text{Atomic mass of the element}}$$

$$= 46 / 23$$

$$= 2 \text{ moles of sodium}$$

5.6 litre of oxygen at S.T.P

$$\text{Number of moles} = \frac{\text{Given volume of O}_2 \text{ at S.T.P}}{\text{Molar volume at S.T.P}}$$

$$\text{Number of moles of oxygen} = \frac{5.6}{22.4} = 0.25 \text{ mole of oxygen}$$

Calculate the number of moles of a sample that contains 12.046×10^{23} atoms of iron ?

$$\text{Number of moles} = \frac{\text{Number of atoms of iron}}{\text{Avogadro's number}}$$

$$= \frac{12.046 \times 10^{23}}{6.023 \times 10^{23}}$$

$$= 2 \text{ moles of iron}$$

Calculation of mass from mole

Calculate the mass of the following

0.3 mole of aluminium (Atomic mass of Al = 27)

$$\text{Number of moles} = \frac{\text{Mass of Al}}{\text{Atomic mass of Al}}$$

$$\text{Mass} = \text{No. of moles} \times \text{atomic mass}$$

$$\text{So, mass of Al} = 0.3 \times 27$$

$$= 8.1 \text{ g}$$

2.24 litre of SO₂ gas at S.T.P

$$\text{Molecular mass of SO}_2 = 32 + (16 \times 2) = 32 + 32 = 64$$

$$\text{Number of moles of SO}_2 = \frac{\text{Given volume of SO}_2 \text{ at S.T.P}}{\text{Molar volume SO}_2 \text{ at S.T.P}}$$

$$\text{Number of moles of SO}_2 = \frac{2.24}{22.4} = 0.1 \text{ mole}$$

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$$

Mass = No. of moles \times molecular mass

Mass = 0.1×64

Mass of SO_2 = 6.4 g

1.51×10^{23} molecules of water Molecular mass of H_2O = 18

$$\text{Number of moles} = \frac{\text{Number of molecules of water}}{\text{Avogadro's number}}$$

$$= \frac{1.51 \times 10^{23}}{6.023 \times 10^{23}}$$

$$= 1 / 4$$

$$= 0.25 \text{ mole}$$

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$$

$0.25 = \text{mass} / 18$

Mass = 0.25×18

Mass = 4.5 g

5×10^{23} molecules of glucose ?

Molecular mass of glucose = 180

$$\text{Mass of glucose} = \frac{\text{Molecular mass} \times \text{number of particles}}{\text{Avogadro's number}}$$

$$= \frac{(180 \times 5 \times 10^{23})}{6.023 \times 10^{23}}$$

$$= 149.43 \text{ g}$$

Calculation based on number of atoms/ molecules.

Calculate the number of molecules in 11.2 litre of CO_2 at S.T.P

$$\text{Number of moles of } \text{CO}_2 = \frac{\text{Volume at S.T.P}}{\text{Molar volume}}$$

$$= 11.2 / 22.4$$

$$= 0.5 \text{ mole}$$

- Number of molecules of CO_2 = number of moles of CO_2 \times Avogadro's number

$$= 0.5 \times 6.023 \times 10^{23}$$

$$= 3.011 \times 10^{23} \text{ molecules of } \text{CO}_2$$

Calculate the number of atoms present in 1 gram of gold (Atomic mass of Au = 198)

$$\text{Number of atoms of Au} = \frac{\text{Mass of Au} \times \text{Avogadro's number}}{\text{Atomic mass of Au}}$$

$$\text{Number of atoms of Au} = \frac{1}{198} \times 6.023 \times 10^{23}$$

$$\text{Number of atoms of Au} = 3.042 \times 10^{21} \text{ g}$$

Calculate the number of molecules in 54 gm of H_2O ?

$$\text{Number of molecules} = \frac{(\text{Avogadro number} \times \text{Given mass})}{\text{Gram molecular mass}}$$

$$\text{Number of molecules of water} = 6.023 \times 10^{23} \times 54 / 18 = 18.069 \times 10^{23} \text{ molecules}$$

Calculate the number of atoms of oxygen and carbon in 5 moles of CO_2 .

1 mole of CO_2 contains 2 moles of oxygen
5 moles of CO_2 contain 10 moles of oxygen

Number of atoms of oxygen = Number of moles of oxygen \times Avogadro's number

$$= 10 \times 6.023 \times 10^{23}$$

$$= 6.023 \times 10^{24} \text{ atoms of Oxygen}$$

- v 1 mole of CO_2 contains 1 mole of carbon
- v 5 moles of CO_2 contains 5 moles of carbon

No. of atoms of carbon = No. of moles of carbon \times Avogadro's number

$$= 5 \times 6.023 \times 10^{23}$$

$$= 3.011 \times 10^{24} \text{ atoms of Carbon}$$

Calculation based on molar volume

Calculate the volume occupied by:

2.5 mole of CO₂ at S.T.P

$$\text{Number of moles of CO}_2 = \frac{\text{Given volume at S.T.P}}{\text{Molar volume at S.T.P}}$$

$$2.5 \text{ mole of CO}_2 = \frac{\text{Volume of CO}_2 \text{ at S.T.P}}{22.4}$$

Volume of CO₂ at S.T.P = 22.4 × 2.5 = 56 litres.

3.011 × 10²³ of ammonia gas molecules

$$\text{Number of moles} = \frac{\text{Number of molecules}}{\text{Avogadro's number}}$$

$$= \frac{3.011 \times 10^{23}}{6.023 \times 10^{23}}$$

$$= 2 \text{ moles}$$

Volume occupied by NH₃

$$= \text{number of moles} \times \text{molar volume}$$

$$= 2 \times 22.4$$

$$= 44.8 \text{ litres at S.T.P}$$

14 g nitrogen gas

$$\text{Number of moles} = \frac{14}{28} = 0.5 \text{ mole}$$

Volume occupied by N₂ at S.T.P

$$= \text{no. of moles} \times \text{molar volume}$$

$$= 0.5 \times 22.4$$

$$= 11.2 \text{ litres}$$

Calculation based on % composition

Calculate % of S in H₂SO₄

Molar mass of H₂SO₄

$$= (1 \times 2) + (32 \times 1) + (16 \times 4)$$

$$= 2 + 32 + 64$$

$$= 98 \text{ g}$$

$$\% \text{ of S in H}_2\text{SO}_4 = \frac{\text{Mass of sulphur}}{\text{Molar mass of H}_2\text{SO}_4} \times 100$$

$$\% \text{ of S in H}_2\text{SO}_4 = \frac{32}{98} \times 100 = 32.65 \%$$



8.PERIODIC CLASSIFICATION OF ELEMENTS

Introduction

- The eighteenth and nineteenth centuries witnessed a rapid development in chemistry in all spheres of scientific activities. By 1860, scientists had already discovered 60 elements and determined their atomic masses. They noticed that some elements had similar properties and hence arranged them into groups. During this period, several new elements were discovered. These elements were found to have different properties. It was realized that instead of studying the properties of all these elements individually, it would be more convenient to divide them into groups and periods in such a way that each group contained a certain number of elements (like an array of fruits and vegetables showing orderliness) with similar properties and periods showing a regular gradation. So, scientists made several attempts to arrange elements in a logical way.

MODERN PERIODIC LAW

- Mendeleev's periodic table had some discrepancies, which were difficult to overcome. For example, the atomic mass of argon (39.95 amu) is greater than that of potassium (39.10 amu), but argon comes before potassium in the periodic table. If elements were arranged solely according to increasing atomic mass, argon would appear in the position occupied by potassium in our modern periodic. No chemist would place argon, a gas with no tendency to react, in the same group as lithium and sodium, which are two highly reactive metals. This kind of discrepancies suggested that some fundamental property other than atomic mass must be the basis of periodicity. The fundamental property turned out to be the number of protons in an atom's nucleus, something that could not have been known by Mendeleev and his contemporaries.
- Henry Moseley, a British scientist in 1912, discovered a new property of elements called atomic number, which provided a better basis for the periodic arrangement of the elements. It is a well-known fact that atomic number of an element is equal to the number of protons or the number of electrons present in the neutral atom of an element. The periodic law was, therefore, modified to frame a modern periodic law, which states that "The physical and chemical properties of the elements are the periodic functions of their atomic numbers".

MODERN PERIODIC TABLE

- With reference to the modern periodic law, the elements were arranged in the increasing order of their atomic numbers to form the modern periodic table. The modern periodic table is a tabular arrangement of elements in rows and columns, highlighting the regular repetition of properties of the elements.

Features of Periods

- The **horizontal rows are called periods**. There are seven periods in the periodic table.
- **First period** (Atomic number 1 and 2): This is the shortest period. It contains only two elements (Hydrogen and Helium).
- **Second period** (Atomic number 3 to 10): This is a short period. It contains eight elements (Lithium to Neon).
- **Third period** (Atomic number 11 to 18): This is also a short period. It contains eight elements (Sodium to Argon).
- **Fourth period** (Atomic number 19 to 36): This is a long period. It contains eighteen elements (Potassium to Krypton). This includes 8 normal elements and 10 transition elements.
- **Fifth period** (Atomic number 37 to 54): This is also a long period. It contains 18 elements (Rubidium to Xenon). This includes 8 normal elements and 10 transition elements.
- **Sixth period** (Atomic number 55 to 86): This is the longest period. It contains 32 elements (Caesium to Radon). This includes 8 normal elements, 10 transition elements and 14 inner transition elements (Lanthanides).
- **Seventh period** (Atomic number 87 to 118): Like the sixth period, this period also accommodates 32 elements. Recently 4 elements have been included by IUPAC

Features of Groups

- The vertical columns in the periodic table starting from top to bottom are called groups. There are 18 groups in the periodic table.
- Based on the common characteristics of elements in each group, they can be grouped as various families.

Group Number	Family
1	

	Alkali Metals
2	Alkaline earth metals
3 to 12	Transition metals
13	Boron Family
14	Carbon Family
15	Nitrogen Family
16	Oxygen Family (or) Chalcogen family
17	Halogens
18	Noble gases

- The Lanthanides and Actinides, which form part of Group 3 are called inner transition elements.
- Except 'group 0', all the elements present in each group have the same number of electrons in their valence shell and thus have the same valency. For example, all the elements of group 1 have one electron in their valence shells (1s¹). So, the valency of all the alkali metals is '1'.
- As the elements present in a group have identical valence shell electronic configurations, they possess similar chemical properties.
- The physical properties of the elements in a group such as melting point, boiling point and density vary gradually.
- The atoms of the 'group 0' elements have stable electronic configuration in their valence shells and hence they are unreactive.

PERIODIC TRENDS IN PROPERTIES

- The electronic configurations of elements help us to explain the periodic recurrence of physical and chemical properties. Anything which repeats itself after a regular interval is called periodic and this behaviour is called periodicity. Some of the atomic properties of the elements are periodic.
- Properties such as atomic radius, ionic radius, ionisation energy, electronegativity, electron affinity, show a regular periodicity and hence they are called periodic properties. The main significance of the modern periodic table is that it gives a clear understanding of the general properties and trends within a group or a period to predict with considerable accuracy, the properties of any element, even though that element may be unfamiliar to us. Let us discuss the periodic trend of some of the properties.

Atomic Radius

- Atomic radius of an atom is defined as the distance between the centre of its nucleus and the outermost shell containing the valence electron. Direct measurement of the radius of an isolated atom is not possible. Except for noble gases, usually the atomic radius is referred to as covalent radius or metallic radius depending on the nature of the bonding between the concerned atoms. Atomic radius in metal atoms is known as metallic radius. It is defined as half the distance between the nuclei of adjacent metal atoms.
- In non-metallic elements, their atomic radius is known as Covalent radius. It is defined as half the distance between the nuclei of two covalently bonded atoms of the same element in a molecule. For example, let us consider H₂ molecule. The distance between the two hydrogen nuclei of the molecule is 0.74 Å. So its covalent radius is $0.74/2 = 0.37 \text{ Å}$.
- When you look at the variation of the atomic radii in the periodic table, there are two distinct trends. Along the period, from left to right, the atomic radius of the elements decreases whereas along the groups, from the top to bottom, the atomic radius increases. The increase, down a group, is due to the increase in the valence shell number down the group. As the shell number increases, the distance between the valence shell and the nucleus increases. In contrast, when you observe along the period, the shell number remains the same but the number of protons (i.e. atomic number) increases. More and more positive charges impose a strong attraction over the electrons and thus the electron cloud shrinks towards the nucleus, which results in the decrease in the atomic size.

Ionic Radii

- It is defined as the distance from the centre of the nucleus of the ion to the point where it exerts its influence on the electron cloud of the ion. You know that ions are formed when an atom loses or gains electrons. When a neutral atom loses an electron, it becomes a positively charged ion called cation, whereas the gain of an electron by a neutral atom forms a negatively charged ion called anion. The size of the ions is important to determine their behaviours in solutions and the structure of ionic solids. The size of a cation is always smaller than its corresponding neutral atom. But, the anion is larger than its neutral atom.
- For instance, lithium and sodium lose the single electron from their outermost energy level to form cations. The ions so formed are smaller because the remaining electrons are at an inner shell and are attracted more strongly by the nucleus. Fluorine and chlorine become negative ions by gaining an electron. When electrons are added, the charge on the nucleus is not great enough to hold the increased number of electrons as closely as it holds the electrons in

the neutral atom. So, as seen in atomic radius, ionic radii also decrease along the period from left to right and increase down the group.

Ionisation Energy

- Ionisation energy is the minimum energy required to remove an electron from a gaseous atom in its ground state to form a cation. It is otherwise called ionisation enthalpy. It is measured in kJ/mol. Higher the ionisation energy, it is more difficult to remove the electron.
- As the atomic size decreases from left to right in a period, more energy is required to remove the electrons. So, the ionisation energy increases along the period. But, down the group, the atomic size increases and hence the valence electrons are loosely bound. They require relatively less energy for the removal. Thus, ionisation energy decreases down the group in the periodic table.

As the positive charge increases the size of the cation decreases As the negative charge increases the size of the anion increases

Electron Affinity

- Electron affinity is the amount of energy released when a gaseous atom gains an electron to form its anion. It is also measured in kJ/mol and represented by the following equation:

$$A(g) + e^{-} \rightarrow A^{-}(g) + \text{Energy}$$

$$Cl(g) + e^{-} \rightarrow Cl^{-}(g) + \text{energy}$$
- Like ionisation energy, electron affinity also increases from left to right in a period and decreases from top to bottom in a group.

Electronegativity

- Electronegativity of an element is the measure of the tendency of its atom to attract the shared pair of electrons towards itself in a covalent bond. Let us consider HCl molecule. Both the hydrogen and chlorine atoms share one electron each to form the covalent bond between them. Chlorine atom has a higher electronegativity and hence it pulls the shared electrons towards itself more strongly than hydrogen. Thus, when the bond breaks, the bonding electrons are left with chlorine forming H⁺ and Cl⁻ ions.
- Electronegativity is based on various experimental data such as bond energy, ionization potential, electron affinity, etc.

- Pauling scale is the widely used scale to determine the electronegativity, which in turn predicts the nature of bonding (ionic or covalent) between the atoms in a molecule.

Electronegativity of some of the elements.

F = 4.0, Cl = 3.0, Br = 2.8, I = 2.5, H = 2.1, Na = 1

- If the difference in electronegativity between two elements is 1.7, the bond has 50% ionic character and 50% covalent character.
- If the difference is less than 1.7, the bond is considered to be covalent.
- If the difference is greater than 1.7, the bond is considered to be ionic.
- Along the period, from left to right in the periodic table, the electronegativity increases because of the increase in the nuclear charge which in turn attracts the electrons more strongly. On moving down a group, the electronegativity of the elements decreases because of the increased number of energy levels.

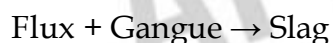
Periodic Property	In periods	In Groups
Atomic radius	Decreases	Increases
Ionic radius	Decreases	Increases
Ionisation energy	Increases	Decreases
Electron affinity	Increases	Decreases
Electronegativity	Increases	Decreases

METALLURGY

- Human life is associated with various metals. We use metals in our day to day activities. It is the utmost need to have some metals like sodium, potassium, calcium, iron, etc. in the human body. Deficiency of these metals affects the metabolic activities thereby causing diseases. So, metals play a vital role in our life. In this section, let us discuss how metals are obtained from various sources by the process of metallurgy.
- Metallurgy is a science of extracting metals from their ores and modifying the metals into alloys for various uses, based on their physical and chemical properties and their structural arrangement of atoms. A metallurgical process involves three main steps as follows:
 - **Concentration or Separation of the ore:** It is the process of removal of impurities from the ore.
 - **Production of the metal:** It is the conversion of the ore into metal.
 - **Refining of the metal:** It is the process of purification of the metal

Terminology in metallurgy

- **Minerals:** A mineral may be a single compound or a complex mixture of various compounds of metals found in the Earth.
- **Ore:** The mineral from which a metal can be readily and economically extracted on a large scale is said to be an ore.
- **For example:** Clay ($\text{Al}_2\text{O}_3 \cdot 2 \text{SiO}_2 \cdot 2 \text{H}_2\text{O}$) and bauxite ($\text{Al}_2\text{O}_3 \cdot 2 \text{H}_2\text{O}$) are the two minerals of aluminium, but aluminium can be profitably extracted only from bauxite. Hence, bauxite is an ore of aluminium and clay is its mineral.
- **Mining:** The process of extracting the ores from the Earth's crust is called mining.
- **Gangue or Matrix:** The rocky impurity associated with an ore is called gangue or matrix.
- **Flux:** It is the substance added to the ore to reduce the fusion temperature and to remove the impurities. E.g. Calcium oxide (basic), Silica (acidic). If the gangue is acidic, then basic flux is added and vice versa.
- **Slag:** It is the fusible product formed when a flux reacts with a gangue during the extraction of metals.



- **Smelting:** Smelting is the process of reducing the roasted metallic oxide from the metal in its molten condition. In this process, impurities are removed as slag by the addition of flux.

Types of separation or concentration of an ore

- There are four major types of separation of ores based on the nature of the ore. The different kinds of ores of metals.

Hydraulic (Gravity Separation) method

- **Principle:** The difference in the densities or specific gravities of the ore and the gangue is the main principle behind this method. Oxide ores are purified by this method. e.g., Haematite Fe_2O_3 the ore of iron.
- **Method:** The ore is poured over a sloping, vibrating corrugated table with grooves and a jet of water is allowed to flow over it. The denser ore particles settle down in the grooves and lighter gangue particles are washed down by water.

Magnetic separation method

- **Principle:** The magnetic properties of the ores form the basis of separation. When either the ore or the gangue is magnetic, this method is employed. e.g., Tin stone SnO_2 , the ore of tin.
- **Method:** The crushed ore is placed over a conveyor belt which rotates around two metal wheels, one of which is magnetic. The magnetic particles are attracted to the magnetic wheel and fall separately apart from the non-magnetic particles.

Froth floatation

- **Principle:** This process depends on the preferential wettability of the ore with oil (pine oil) and the gangue particles by water. Lighter ores, such as sulphide ores, are concentrated by this method. e.g., Zinc blende (ZnS).
- **Method:** The crushed ore is taken in a large tank containing oil and water and agitated with a current of compressed air. The ore is wetted by the oil and gets separated from the gangue in the form of froth. Since the ore is lighter, it comes on the surface with the froth and the impurities are left behind. e.g., Zinc blende (ZnS).

Chemical method or Leaching

This method is employed when the ore is in a very pure form.

Oxide Ores	Carbonate Ores	Halide Ores	Sulphide ores
Bauxite ($\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$)	Marble (CaCO_3)	Cryolite (Na_3AlF_6)	Galena (pbs)
Cuprite (Cu_2O)	Magnesite (MgCO_3)	Fluorspar (CaF_2)	Iron Pyrite (FeS_2)
Haematite (Fe_2O_3)	Siderite (FeCO_3)	Rock Salt (NaCl)	Zinc Blende (ZnS)

- The ore is treated with a suitable reagent such that the ore is soluble in it but the impurities are not. The impurities are removed by filtration. The solution of the ore, i.e., the filtrate is treated with a suitable reagent which precipitates the ore. E.g. Bauxite $\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$, the ore of aluminium.

OCCURRENCE OF ORES IN TAMIL NADU

- v **Lime stone:** Coimbatore, Cuddalore, Dindugul
- v **Gypsum:** Tiruchi and Coimbatore Distiricts
- v **Titanium minerals:** Kanyakumari, Tirunelveli and Tuticorin.

- v **Chromite:** Coimbatore and Salem district.
- v **Magnetite:** Dharmapuri, Erode, Salem, Thiruvannamalai.
- v **Tungsten:** Madurai and Dindugal.

PROPERTIES OF METALS

Physical properties

- **Physical state:** All metals are solids at room temperature except mercury and gallium.
- **Lustre:** Metals possess a high lustre (called metallic lustre).
- **Hardness:** Most of the metals are hard and strong (exceptions: sodium and potassium can be cut with a knife)
- **Melting point and Boiling point:** Usually, metals possess high melting and boiling points and vaporize only at high temperatures (exceptions: gallium, mercury, sodium and potassium).
- **Density:** Metals have a high density (exceptions: sodium and potassium are less dense than water).
- **Ductility:** Metals are usually ductile. In other words, they can be drawn into thin wires without breaking.
- **Malleability:** Metals are usually malleable, i.e, they can be beaten into thin sheets without cracking (except zinc and mercury).
- **Conduction of heat and electricity:** Metals are good conductors of heat and electricity; silver and copper excel in this property (exception: tungsten)
- **Solubility:** Usually, metals do not dissolve in liquid solvents.

Chemical Properties

- **Valence electrons:** Atoms of metals usually have 1,2 or 3 electrons in their outermost shell.
- **Formation of ions:** Metals form Positive ions by the loss of electrons and hence they are electro positive.
- **Discharge of ions:** Metals are discharged at the cathode during the electrolysis of their compounds.
- **Atomicity:** Molecules of metals in their vapour state are usually monoatomic.

- **Nature of oxides:** Oxides of metals are usually basic.

EXTRACTIVE METALLURGY OF ALUMINIUM

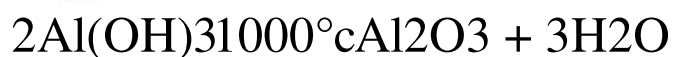
- v Aluminium is the metal found most abundantly in the Earth's crust. Since it is a reactive metal, it occurs in the combined state. The important ores of aluminium are as follows.

Ores of Aluminium	Formula
Bauxite	$\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$
Cryolite	Na_3AlF_6
Corundum	Al_2O_3

- Bauxite is the chief ore of aluminium. The extraction of aluminium from bauxite involves two steps:

Conversion of bauxite into alumina - Baeyer's Process

- The conversion of Bauxite into Alumina involves the following steps:
- Bauxite ore is finely ground and heated under pressure with a solution of concentrated caustic soda solution at 150°C to obtain sodium meta aluminate.
- On diluting sodium meta aluminate with water, a precipitate of aluminium hydroxide is formed.
- The precipitate is filtered, washed, dried and ignited at 1000°C to get alumina.



Electrolytic reduction of alumina - Hall's Process

- Aluminium is produced by the electrolytic reduction of fused alumina (Al_2O_3) in the electrolytic cell
- **Cathode:** Iron tank lined with graphite
- **Anode:** A bunch of graphite rods suspended in molten electrolyte.
- **Electrolyte:** Pure alumina + molten cryolite + fluorspar (fluorspar lowers the fusion temperature of electrolyte)

Temperature: $900 - 950^\circ\text{C}$

Voltage used: 5-6 V

Overall reaction: $2 \text{Al}_2\text{O}_3 \rightarrow 4 \text{Al} + 3 \text{O}_2 \uparrow$

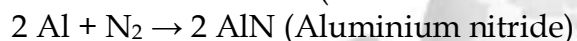
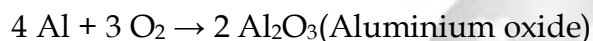
- Aluminium is deposited at the cathode and oxygen gas is liberated at the anode. Oxygen combines with graphite to form CO_2 .

Physical Properties of Aluminium

- It is a silvery white metal
- It has low density (2.7) and it is light
- It is malleable and ductile
- It is a good conductor of heat and electricity.
- Its melting point is 660°C .
- It can be polished to produce a shiny attractive appearance.

Chemical Properties of Aluminium

- Reaction with air:** It is not affected by dry air. On heating at 800°C , aluminium burns very brightly forming its oxide and nitride.



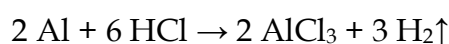
- Reaction with water:** Water does not react with aluminium due to the layer of oxide on it. When steam is passed over red hot aluminium, hydrogen is produced.



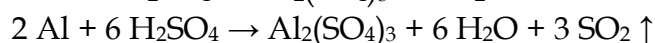
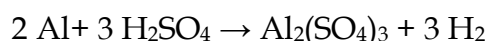
- Reaction with alkalis:** It reacts with strong caustic alkalis forming aluminates.



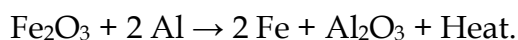
- Reaction with acids:** With dilute and con. HCl it liberates H_2 gas.



- Aluminium liberates hydrogen on reaction with dilute sulphuric acid. Sulphur dioxide is liberated with hot concentrated sulphuric acid



As reducing agent: Aluminium is a powerful reducing agent. When a mixture of aluminium powder and iron oxide is ignited, the latter is reduced to metal. This process is known as **aluminothermic process**.



Uses

Aluminium is used in

- v household utensils
- v electrical cable industry
- v making aeroplanes and other industrial machine parts

EXTRACTIVE METALLURGY OF COPPER

Occurrence:

- It was named as cuprum by the Romans because they got it from the Island of Cyprus. Copper is found in the native state as well as combined state.

Ores of copper

Copper pyrites

Cuprite or ruby copper

Copper glance

Formula

CuFeS_2

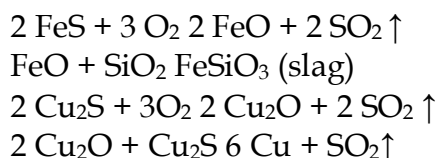
Cu_2O

Cu_2S

- The chief ore of copper is copper pyrite. It yields nearly 76% of the world production of copper. Extraction of copper from copper pyrites involves the following steps
- **Concentration of ore:** The ore is crushed and the concentrated by froth floatation process.
- **Roasting:** The concentrated ore is roasted in excess of air. During the process of roasting, the moisture and volatile impurities are removed. Sulphur, phosphorus, arsenic and antimony are removed as oxides. Copper pyrite is partly converted into sulphides of copper and iron.



- **Smelting:** The roasted ore is mixed with powdered coke and sand and is heated in a blast furnace to obtain matte ($\text{Cu}_2\text{S} + \text{FeS}$) and slag. The slag is removed as waste.
- **Bessemerisation:** The molten matte is transferred to Bessemer converter in order to obtain blister copper. Ferrous sulphide from matte is oxidized to ferrous oxide, which is removed as slag using silica.



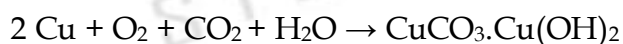
- **Refining:** Blister copper contains 98% of pure copper and 2% of impurities and is purified by **electrolytic refining**. This method is used to get metal of a high degree of purity. For electrolytic refining of copper, we use:
 - v **Cathode:** A thin plate of pure copper metal.
 - v **Anode:** A block of impure copper metal.
 - v **Electrolyte:** Copper sulphate solution acidified with sulphuric acid
- When electric current is passed through the electrolytic solution, pure copper gets deposited at the cathode and the impurities settle at the bottom of the anode in the form of sludge called anode mud.

Physical Properties of Copper

- Copper is a reddish brown metal, with high lustre, high density and high melting point (1356°C).

Chemical Properties of Copper

- **Action of Air and Moisture:** Copper gets covered with a green layer of basic copper carbonate in the presence of CO₂ and moisture.



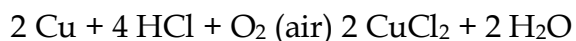
- **Action of Heat:** On heating at different temperatures in the presence of oxygen, copper forms two types of oxides CuO, Cu₂O.



Action of Acids:

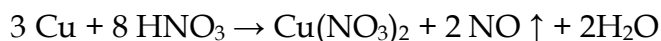
With dilute HCl and dilute H₂SO₄:

- Dilute acids such as HCl and H₂SO₄ have no action on these metals in the absence of air. Copper dissolves in these acids in the presence of air.



With dil. HNO₃:

- Copper reacts with dil. HNO₃ with the liberation of Nitric Oxide gas.



Action of Chlorine:

- Chlorine reacts with copper, resulting in the formation of copper(II) chloride.



Action of Alkalis:

- Copper is not attacked by alkalis.

Uses of Copper:

- ✓ It is extensively used in manufacturing electric cables and other electric appliances.
- ✓ It is used for making utensils, containers, calorimeters and coins,
- ✓ It is used in electroplating.
- ✓ It is alloyed with gold and silver for making coins and jewels

EXTRACTIVE METALLURGY OF IRON

Occurrence:

- Iron is the second most abundant metal available next to aluminium. It occurs in nature as oxides, sulphides and carbonates. The ores of iron are as follows:

Ores of iron	Formula
Haematite	Fe ₂ O ₃
Magnetite	Fe ₃ O ₄
Iron pyrite	FeS ₂

- Iron is chiefly extracted from haematite ore (Fe₂O₃)

Concentration by Gravity Separation:

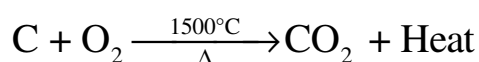
- The powdered ore is washed with a stream of water. As a result, the lighter sand particles and other impurities are washed away and the heavier ore particles settle down.

Roasting and Calcination:

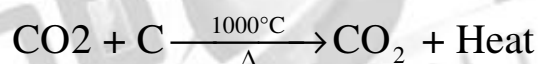
- The concentrated ore is strongly heated in a limited supply of air in a reverberatory furnace. As a result, moisture is driven out and sulphur, arsenic and phosphorus impurities are oxidized off.

Smelting (in a Blast Furnace):

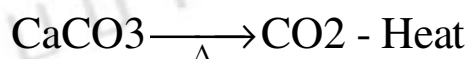
- The charge consisting of roasted ore, coke and limestone in the ratio 8:4:1 is smelted in a blast furnace by introducing it through the cup and cone arrangement at the top. There are three important regions in the furnace.
- The Lower Region (Combustion Zone)- The temperature is at 1500°C. In this region, coke burns with oxygen to form CO₂ when the charge comes in contact with a hot blast of air.



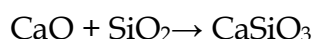
- It is an exothermic reaction since heat is liberated.
- The Middle Region (Fusion Zone) - The temperature prevails at 1000°C. In this region, CO₂ is reduced to CO.



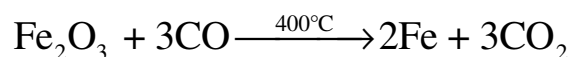
- Limestone decomposes to calcium oxide and CO₂



- These two reactions are endothermic due to absorption of heat. Calcium oxide combines with silica to form calcium silicate slag.



- The Upper Region (Reduction Zone)- The temperature prevails at 400°C . In this region carbon monoxide reduces ferric oxide to form a fairly pure spongy iron.



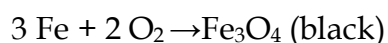
- The molten iron is collected at the bottom of the furnace after removing the slag.
- The iron thus formed is called pig iron. It is remelted and cast into different moulds. This iron is called cast iron.

Physical properties:

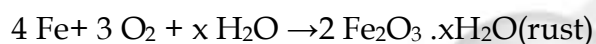
- v It is a lustrous metal, greyish white in colour.
- v It has high tensility, malleability and ductility.
- v It can be magnetized

Chemical properties:

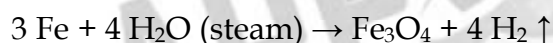
- **Reaction with air or oxygen:** Only on heating in air, iron forms magnetic oxide.



- **Reaction with moist air:** When iron is exposed to moist air, it forms a layer of brown hydrated ferric oxide on its surface. This compound is known as rust and the phenomenon of formation of rust is known as **rusting**.



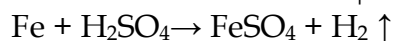
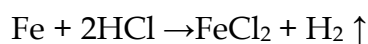
- **Reaction with steam:** When steam is passed over red hot iron, magnetic oxide is formed.



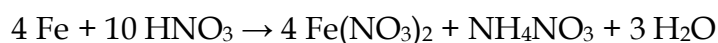
- **Reaction with chlorine:** Iron combines with chlorine to form ferric chloride.



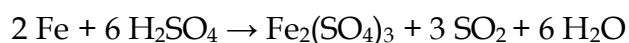
- **Reaction with acids:** With dilute HCl and dilute H₂SO₄ it liberates H₂ gas.



- With dilute HNO₃ in cold condition it gives ferrous nitrate.



- With con. H₂SO₄ it forms ferric sulphate.



- When iron is dipped in con. HNO₃ it becomes chemically passive or inert due to the formation of a layer of iron oxide (Fe₃O₄) on its surface.

Uses of iron

- **Pig iron** (Iron with 2-4.5% of carbon): It is used in making pipes, stoves, radiators, railings, manhole covers and drain pipes.
- **Steel** (Iron with < 0.25% of carbon): It is used in the construction of buildings, machinery, transmission cables and T.V towers and in making alloys.
- **Wrought iron** (Iron with 0.25-2% of wrought carbon): It is used in making springs, anchors and electromagnets.

ALLOYS

- An alloy is a homogeneous mixture of two or more metals or of one or more metals with certain non-metallic elements.
- The properties of alloys are often different from those of its components. Pure gold is too soft to be used. The addition of small percentage of copper enhances its strength and utility.

Amalgam

- An amalgam is an alloy of mercury with another metal. These alloys are formed through metallic bonding with the electrostatic force of attraction between the electrons and the positively charged metal ions. Silver tin amalgam is used for dental filling.

Reasons for alloying:

- ✓ To modify appearance and colour
- ✓ To modify chemical activity.
- ✓ To lower the melting point.
- ✓ To increase hardness and tensile strength.
- ✓ To increase resistance to electricity.

Method of making alloys

- By fusing the metals together. E.g. Brass is made by melting zinc and copper.
- By compressing finely divided metals. E.g. Wood metal: an alloy of lead, tin, bismuth and cadmium powder is a fusible alloy.

Alloys as solid solutions:

- Alloys can be considered solid solutions in which the metal with high concentration is solvent and other metals are solute.
- For example, brass is a solid solution of zinc (solute) in copper (solvent).

Types of Alloys

- Based on the presence or absence of Iron, alloys can be classified into:
 - ✓ Ferrous alloys: Contain Iron as a major component. A few examples of ferrous alloys are Stainless Steel, Nickel Steel etc.
 - Non-ferrous alloys: These alloys do not contain Iron as a major component. For example, Aluminium alloy, Copper alloy etc.

Copper Alloys (Non-ferrous)

Alloys	Uses
Brass (Cu, Zn)	Electrical fittings, medal, decorative items, hardware
Bronze (Cu, Sn)	Statues, coins, bells, gongs

Aluminium Alloys (Non-ferrous)

Alloys	Uses
Duralumin (Al, Mg, Mn, Cu)	Aircrafts, tools, pressure cookers
Magnalium (Al, Mg)	Aircraft, scientific instruments

Iron Alloys (Ferrous)

Alloys	Uses
Stainless steel (Fe,C, Ni,Cr)	Utensils, cutlery, automobile parts
Nickel steel (Fe,C,Ni)	Cables , aircraftparts, propeller

CORROSION

- It is the gradual destruction of metals by chemical or electrochemical reaction with the environment. It is a natural process which converts a metal into its oxide, hydroxide or sulphide so that it loses its metallic characteristics.
- Rust is chemically known as hydrated ferric oxide (it is formulated as $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$). Rusting results in the formation of scaling reddish brown hydrated ferric oxide on the surface of iron and iron containing materials.

Types of Corrosion

Dry Corrosion or Chemical Corrosion:

- The corrosive action in the absence of moisture is called dry corrosion. It is the process of a chemical attack on a metal by a corrosive liquids or gases such as O_2 , N_2 , SO_2 , H_2S etc. It occurs at high temperature. Of all the gases mentioned above O_2 is the most reactive gas to impart the chemical attack.

Wet Corrosion or Electrochemical Corrosion:

- The corrosive action in the presence of moisture is called wet corrosion. It occurs as a result of electrochemical reaction of metal with water or aqueous solution of salt or acids or bases.

Methods of preventing corrosion

1. **Alloying:** The metals can be alloyed to prevent the process of corrosion. E.g: Stainless Steel
2. **Surface Coating:** It involves application of a protective coating over the metal. It is of the following types:

Galvanization:

- It is the process of coating zinc on iron sheets by using electric current.

Electroplating:

- It is a method of coating one metal over another metal by passing electric current.

Anodizing:

- It is an electrochemical process that converts the metal surface into a decorative, durable and corrosion resistant. Aluminium is widely used for anodizing process.

Cathodic Protection:

- It is the method of controlling corrosion of a metal surface protected is coated with the metal which is easily corrodible. The easily corrodible metal is called Sacrificial metal to act as anode ensuring cathodic protection

PAMBAN BRIDGE

It is a railway bridge which connects the town of Rameshwaram on Pamban Island to mainland India. Opened on 1914, it was India's first sea bridge in India until the opening of the BandraWorli Sea Link in 2010. We can control the corrosion and renovation of historical pamban bridge by a periodical protective coating which will be the strong example for applied chemistry to uphold our history.



Periodic Classification of Elements

"An awareness of the periodic table is essential to anyone who wishes to disentangle the word and see how it is built up from the fundamental building blocks of the chemistry, the chemical elements "

- Glenn T. Seaborg

Glenn Theodore Seaborg received Nobel Prize in 1951 in chemistry for the discoveries of trans-uranium elements. He was the co-discoverer of plutonium and other trans uranium elements. He along with his colleagues has discovered over a hundred isotopes of other elements. He demonstrated that actinide elements are analogues to rare earth series of lanthanide elements.

Introduction

There are millions of chemical compounds existing in nature with different compositions and properties, formed from less than 100 naturally occurring elements.

The discovery of elements is linked with human civilization. In stone age man has used some metals to suit his needs without knowing that they are elements. Soon he learnt to extract elements from ores and fashion them into his daily life. Over the years, more and more elements were discovered. In 1789, Lavoisier from France, published the first list of chemical elements containing 23 elements after several experimental investigations.

Antoine Lavoisier classified the substances into four groups of elements namely acid-making elements, gas-like elements, metallic elements and earthy elements

Lavoisier table

acid-making elements	gas-like elements
Sulphur	Light
Phosphorus	Caloric (heat)
Charcoal (carbon)	Oxygen
	Azote (nitrogen)
	Hydrogen

metallic elements	earthy elements
cobalt, mercury, tin	lime (calcium oxide)
copper, nickel, iron	magnesia (magnesium oxide)
gold, lead, silver, zinc	barytes (barium)

	sulphate)
manganese, tungsten	argilla (aluminium oxide)
platina (platinum)	silex (silicon dioxide)

Classification of Elements

During the 19th century, scientists have isolated several elements and the list of known elements increased. Currently, we have 118 known elements. Out of 118 elements, 92 elements with atomic numbers 1 to 92 are found in nature. Scientists have found out there are some similarities in properties among certain elements. This observation has led to the idea of classification of elements based on their properties. In fact, classification will be beneficial for the effective utilization of these elements. Several attempts were made to classify the elements. However, classification based on the atomic weights led to the construction of a proper form of periodic table.

In 1817, J. W. Döbereiner classified some elements such as chlorine, bromine and iodine with similar chemical properties into the group of three elements called as triads. In triads, the atomic weight of the middle element nearly equal to the arithmetic mean of the atomic weights of the remaining two elements. However, only a limited number of elements can be grouped as triads.

Dobereiner Triads

S. No.	Elements in the Triad	Atomic weight of middle element	Average atomic weight of the remaining Elements
1.	Li, Na, K	23	$\frac{7 + 39}{2} = 23$
2.	Cl, Br, I	80	$\frac{35.5 + 127}{2} = 81.25$
3.	Ca, Sr, Ba	88	$\frac{40 + 137}{2} = 88.5$

This concept can not be extended to some triads which have nearly same atomic masses such as [Fe, Co, Ni], [Ru, Rh, Pd] and [Os, Ir, Pt].

In 1862, A. E. B. de Chancourtois reported a correlation between the properties of the elements and their atomic weights. He said 'the properties of bodies are the properties of numbers'. He intended the term numbers to mean the value of atomic weights. He designed a helix by tracing at an angle 45° to the vertical axis of a cylinder with circumference of 16 units. He arranged the elements in the increasing atomic weights along the helix on the surface of this cylinder. One complete turn of a helix corresponds to an atomic weight increase of 16. Elements which lie on the 16 equidistant vertical lines drawn on the surface of cylinder shows

similar properties. This was the first reasonable attempt towards the creation of periodic table. However, it did not attract much attention.

In 1864, J. Newland made an attempt to classify the elements and proposed the law of octaves. On arranging the elements in the increasing order of atomic weights, he observed that the properties of every eighth element are similar to the properties of the first element. This law holds good for lighter elements up to calcium.

Newlands' Octaves

⁷ Li	⁹ Be	¹¹ B	¹² C	¹⁴ N	¹⁶ O	¹⁹ F
²³ Na	²⁴ Mg	²⁷ Al	²⁹ Si	³¹ P	³² S	^{35.5} Cl
³⁹ K	⁴⁰ Ca					

Mendeleev's Classification

In 1868, Lothar Meyer had developed a table of the elements that closely resembles the modern periodic table. He plotted the physical properties such as atomic volume, melting point and boiling point against atomic weight and observed a periodical pattern.

During same period Dmitri Mendeleev independently proposed that “the properties of the elements are the periodic functions of their atomic weights” and this is called periodic law. Mendeleev listed the 70 known elements at that time in several vertical columns in order of increasing atomic weights. Thus, Mendeleev constructed the first periodic table based on the periodic law.

As shown in the periodic table, he left some blank spaces since there were no known elements with the appropriate properties at that time. He and others predicted the physical and chemical properties of the missing elements. Eventually these missing elements were discovered and found to have the predicted properties. For example, Gallium (Ga) of group III and germanium (Ge) of group IV were unknown at that time. But Mendeleev predicted their existence and properties. He referred the predicted elements as eka-aluminium and eka-silicon. After discovery of the actual elements, their properties were found to match closely to those predicted by Mendeleev.

Properties predicted for Eka-aluminium and Eka-silicon

S. NO.	Property	Ekaaluminium (Predicted)	Gallium (Observed)	Ekasilicon (Predicted)	Germanium (Observed)
1.	Atomic weight	68	70	72	72.59
2.	Density	5.9	5.94	5.5	5.35

Some elements with similar properties were placed in different groups and those with dissimilar properties were placed in same group.

Example: Tellurium (127.6) was placed in VI group but Iodine (127.0) was placed in VII group.

Similarly elements with higher atomic weights were placed before lower atomic weights based on their properties in contradiction to his periodic law. Example $^{59}\text{Co}_{27}$ was placed before $^{58.7}\text{Ni}_{28}$

Moseley's Work and Modern Periodic Law

In 1913, Henry Moseley studied the characteristic X-rays spectra of several elements by bombarding them with high energy electrons and observed a linear correlation between atomic number and the frequency of X-rays emitted which is given by the following expression.

$$\nu = a(Z - b)$$

Where, ν is the frequency of the X-rays emitted by the element with atomic number 'Z'; a and b are constants and have same values for all the elements.

The plot of $\sqrt{\nu}$ against Z gives a straight line. Using this relationship, we can determine the atomic number of an unknown (new) element from the frequency of X-ray emitted.

Based on his work, the modern periodic law was developed which states that, "the physical and chemical properties of the elements are periodic functions of their atomic numbers." Based on this law, the elements were arranged in order of their increasing atomic numbers. This mode of arrangement reveals an important truth that the elements with similar properties recur after regular intervals. The repetition of physical and chemical properties at regular intervals is called periodicity.

Modern Periodic Table

The physical and chemical properties of the elements are correlated to the arrangement of electrons in their outermost shell (valence shell). Different elements having similar outer shell electronic configuration possess similar properties. For example, elements having one electron in their valence shell s-orbital possess similar physical and chemical properties. These elements are grouped together in the modern periodic table as first group elements.

Electronic configuration of
alkali metals (ns¹)

Elements	Atomic	Number of	Valence
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in Group 1	number	electrons in various shells in the order K L M N P	shell configuration
Li	3	2,1	$2s^1$
Na	11	2,8,1	$3s^1$
K	19	2,8,8,1	$4s^1$
Rb	37	2,8,18,8,1	$5s^1$
Cs	55	2,8,18,18,8,1	$6s^1$
Fr	87	2,8,18,32,18,1	$7s^1$

Similarly, all the elements are arranged in the modern periodic table which contains 18 vertical columns and 7 horizontal rows. The vertical columns are called groups and the horizontal rows are called periods. Groups are numbered 1 to 18 in accordance with the IUPAC recommendation which replaces the old numbering scheme IA to VIIA, IB to VIIB and VIII.

Each period starts with the element having general outer electronic configuration ns^1 and ends with np^6 . Here 'n' corresponds to the period number (principal quantum number). The aufbau principle and the electronic configuration of atoms provide a theoretical foundation for the modern periodic table.

Nomenclature of Elements with Atomic Number Greater than 100

Usually, when a new element is discovered, the discoverer suggests a name following IUPAC guidelines which will be approved after a public opinion. In the meantime, the new element will be called by a temporary name coined using the following IUPAC rules, until the IUPAC recognises the new name.

1. The name was derived directly from the atomic number of the new element using the following numerical roots.

Notation for IUPAC Nomenclature of elements

Digital	0	1	2	3	4	5	6	7	8	9
Root	Nil	Un	Bi	Tri	Quad	Pent	Hex	Sept	Oct	enn
Abbreviation	N	U	B	T	Q	P	H	S	O	e

2. The numerical roots corresponding to the atomic number are put together and 'ium' is added as suffix
3. The final 'n' of 'enn' is omitted when it is written before 'nil' (enn + nil = enil) similarly the final 'i' of 'bi' and 'tri' is omitted when it is written before 'ium' (bi + ium = bium; tri + ium = trium)

4. The symbol of the new element is derived from the first letter of the numerical roots.

The following table illustrates these facts.

Name of elements with atomic number above 100

Atomic number	Temp. Name	Temp. Symbol	Name of the element	Symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnilium	Unn	Darmstadtium	Ds
111	Ununnilium	Uuu	Roentgenium	Rg
112	Ununbium	Uub	Copernicium	Cn
113	Ununtrium	Uut	Nihonium	Nh
114	Ununquadium	Uup	Flerovium	Lv
115	Ununpentium	Uup	Moscovium	Ts
116	Ununhexium	Uuh	Livermorium	Lv
117	Ununseptium	Uus	Tennessine	Ts
118	Ununoctium	Uco	Oganesson	Og

Grouping of Elements based on Electronic Configurations

In the modern periodic table, the elements are organised in 7 periods and 18 groups based on the modern periodic law. The placement of element in the periodic table is closely related to its outer shell electronic configuration. Let us analyse the change in the electronic configuration of elements along the periods and down the groups.

Variation of Electronic Configuration along the periods

We have already learnt that each period starts with the element having general outer electronic configuration ns^1 and ends with ns^2, np^6 where n is the period number. The first period starts with the filling of valence electrons in $1s$ orbital, which can accommodate only two electrons. Hence, the first period has two elements, namely hydrogen and helium. The second period starts with the filling of valence electrons in $2s$ orbital followed by three $2p$ orbitals with eight elements from lithium to neon. The third period starts with filling of valence electrons in the $3s$ orbital followed by $3p$ orbitals. The fourth period starts with filling of valence electrons from $4s$ orbital followed by $3d$ and $4p$ orbitals in accordance with Aufbau

principle. Similarly, we can explain the electronic configuration of elements in the subsequent periods.

Table 3.10 Electronic configuration of elements in a period

Period number (n)	Filling of electrons in orbitals		Number of elements	Outer shell Electronic configuration	
	Starts from	Ends with		First element	Last element
1	1s	1s	2	H - 1s ¹	He - 1s ²
2	2s	2p	8	Li - 2s ¹	Ar - 2s ² 2p ⁶
3	3s	3p	8	Na - 3s ¹	Ne - 3s ² 3p ⁶
4	4s	3d → 4p	18	K - 4s ¹	Kr - 4s ² 4p ⁶
5	5s	4d → 5p	18	Rb - 5s ¹	Xe - 5s ² 5p ⁶
6	6s	4f → 5d → 6p	32	Cs - 6s ¹	Rn - 6s ² 6p ⁶
7	7s	5f → 6d → 7p	32	Fr - 7s ¹	Og - 7s ² 7p ⁶

In the fourth period the filling of 3d orbitals starts with scandium and ends with zinc. These 10 elements are called first transition series. Similarly 4d, 5d and 6d orbitals are filled in successive periods and the corresponding series of elements are called second, third and fourth transition series respectively.

In the sixth period the filling of valence electrons starts with 6s orbital followed by 4f, 5d and 6p orbitals. The filling up of 4f orbitals begins with Cerium (Z=58) and ends at Lutetium (Z=71). These 14 elements constitute the first inner-transition series called Lanthanides. Similarly, in the seventh period 5f orbitals are filled, and its 14 elements constitute the second inner-transition series called Actinides. These two series are placed separately at the bottom of the modern periodic table.

Variation of Electronic Configuration in the Groups:

Elements of a group have similar electronic configuration in the outer shell. The general outer electronic configurations for the 18 groups are listed in the Table 3.11. The groups can be combined as s, p, d and f block elements on the basis of the orbital in which the last valence electron enters.

The elements of group 1 and group 2 are called s-block elements, since the last valence electron enters the ns orbital. The group 1 elements are called alkali metals while the group 2 elements are called alkaline earth metals. These are soft metals and possess low melting and boiling points with low ionisation enthalpies. They are highly reactive and form ionic compounds. They are highly electropositive in nature and most of the elements imparts colour to the flame. We will study the properties of these group elements in detail in subsequent chapters.

The elements of groups 13 to 18 are called p-block elements or representative elements and have a general electronic configuration ns², np¹⁻⁶. The elements of the

group 16 and 17 are called chalcogens and halogens respectively. The elements of 18th group contain completely filled valence shell electronic configuration (ns^2, np^6) and are called inert gases or nobles gases. The elements of p-block have high negative electron gain enthalpies. The ionisation energies are higher than that of s-block elements. They form mostly covalent compounds and shows more than one oxidation states in their compounds.

The elements of the groups 3 to 12 are called d-block elements or transition elements with general valence shell electronic configuration $ns^{1-2}, (n-1)d^{1-10}$. These elements also show more than one oxidation state and form ionic, covalent and coordination compounds. They can form interstitial compounds and alloys which can also act as catalysts. These elements have high melting points and are good conductors of heat and electricity.

The lanthanides ($4f^{1-14}, 5d^0-1, 6s^2$) and the actinides ($5f^{0-14}, 6d^0-, 7s^2$) are called f-block elements. These elements are metallic in nature and have high melting points. Their compounds are mostly coloured. These elements also show variable oxidation states.

Table 3.11 General outer electronic configuration of elements in groups:

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
ns^1	ns^2	$ns^2(n-1)d^1$	$ns^2(n-1)d^2$	$ns^2(n-1)d^3$	$ns^1(n-1)d^5$	$ns^2(n-1)d^5$	$ns^2(n-1)d^6$	$ns^2(n-1)d^7$	$ns^2(n-1)d^8$	$ns^1(n-1)d^{10}$	$ns^2(n-1)d^{10}$	$ns^2 np^1$	$ns^2 np^2$	$ns^2 np^3$	$ns^2 np^4$	$ns^2 np^5$	$ns^2 np^6$
s Block elements	d-Block elements										p-Block elements						
f block elements	Lanthanides $4f^{1-14} 5d^{0-1} 6s^2$ Actinides $5f^{0-14} 6d^{0-2} 7s^2$																

Periodic Trends in Properties

As discussed earlier, the electronic configuration of the elements shows a periodic variation with increase in atomic numbers. Similarly a periodic trend is observed in physical and chemical behaviour of elements. In this section, we will study the periodic trends in the following properties of elements.

1. Atomic radius
2. Ionic radius
3. Ionisation enthalpy (energy)
4. Electron gain enthalpy (electron affinity)
5. Electronegativity

Atomic radius

Atomic radius of an atom is defined as the distance between the centre of its nucleus and the outermost shell containing the valence electron.

It is not possible to measure the radius of an isolated atom directly. Except for noble gases, usually atomic radius is referred to as covalent radius or metallic radius depending upon the nature of bonding between the concerned atoms.

Covalent radius

It is one-half of the internuclear distance between two identical atoms linked together by a single covalent bond. Inter nuclear distance can be determined using x-ray diffraction studies.

Example:

The experimental internuclear distance in Cl₂ molecule is 1.98 Å. The covalent radius of chlorine is calculated as below.

$$\begin{aligned}
 d_{Cl-Cl} &= r_{Cl} + r_{Cl} \\
 \Rightarrow d_{Cl-Cl} &= 2r_{Cl} \\
 \Rightarrow r_{Cl} &= \frac{d_{Cl-Cl}}{2} \\
 &= \frac{1.98}{2} = 0.99 \text{ \AA}
 \end{aligned}$$

$$\begin{aligned}
 \text{The covalent radius of chlorine} &= \frac{198}{2} \text{ pm} \\
 &= 99 \text{ pm}
 \end{aligned}$$

The formation of covalent bond involves the overlapping of atomic orbitals and it reduces the expected internuclear distance. Therefore covalent radius is always shorter than the actual atomic radius.

The covalent radius of individual atom can also be calculated using the internuclear distance (d_{A-B}) between two different atoms A and B. The simplest method proposed by Schomaker and Stevenson is as follows.

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

where χ_A and χ_B are the electronegativities of A and B respectively in Pauling units. Here $\chi_A > \chi_B$ and radius is in Å.

Let us calculate the covalent radius of hydrogen using the experimental d_{H-Cl} value is 1.28 Å and the covalent radius of chlorine is 0.99 Å. In Pauling scale the electronegativity of chlorine and hydrogen are 3 and 2.1 respectively.

$$d_{\text{H-Cl}} = r_{\text{H}} + r_{\text{Cl}} - 0.09 (\chi_{\text{Cl}} - \chi_{\text{H}})$$

$$1.28 = r_{\text{H}} + 0.99 - 0.09 (3 - 2.1)$$

$$1.28 = r_{\text{H}} + 0.99 - 0.09 (0.9)$$

$$1.28 = r_{\text{H}} + 0.99 - 0.081$$

$$1.28 = r_{\text{H}} + 0.909$$

$$\therefore r_{\text{H}} = 1.28 - 0.909 = 0.371 \text{ Å}$$

Metallic radius

It is defined as one-half of the distance between two adjacent metal atoms in the closely packed metallic crystal lattice.

For example, the distance between the adjacent copper atoms in solid copper is 2.56 Å and therefore the metallic radius of copper is

$$\frac{2.56}{2} = 1.28 \text{ Å}$$

The metallic radius can be calculated using the unit cell length of the metallic crystal. You will study the detailed calculation procedure in XII standard solid state unit.

Periodic Trends in Atomic Radius

Variation in Periods

Atomic radius tends to decrease in a period. As we move from left to right along a period, the valence electrons are added to the same shell. The simultaneous addition of protons to the nucleus, increases the nuclear charge, as well as the electrostatic attractive force between the valence electrons and the nucleus. Therefore atomic radius decreases along a period.

Effective nuclear charge

In addition to the electrostatic forces of attraction between the nucleus and the electrons, there exists repulsive forces among the electrons. The repulsive force between the inner shell electrons and the valence electrons leads to a decrease in the electrostatic attractive forces acting on the valence electrons by the nucleus. Thus, the inner shell electrons act as a shield between the nucleus and the valence electrons. This effect is called shielding effect.

The net nuclear charge experienced by valence electrons in the outermost shell is called the effective nuclear charge. It is approximated by the below mentioned equation.

$$Z_{\text{eff}} = Z - S$$

Where Z is the atomic number and 'S' is the screening constant which can be calculated using Slater's rules as described below.

Step 1:

Write the electronic configuration of the atom and rearrange it by grouping ns and np orbitals together and others separately in the following form. (1s) (2s, 2p) (3s, 3p) (3d) (4s, 4p) (4d) (4f) (5s, 5p)...

Step 2:

Identify the group in which the electron of interest is present. The electron present right to this group does not contribute to the shielding effect.

Each of the electrons within the identified group (denoted by 'n') shields to an extent of 0.35 unit of nuclear charge. However, it is 0.30 unit for 1s electron.

Step 3 :

Shielding of inner shell electrons.

If the electron of interest belongs to either s or p orbital,

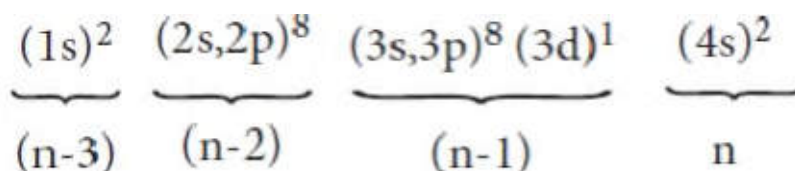
i) each electron within the (n-1) group shields to an extent of 0.85 unit of nuclear charge, and

ii) each electron within the (n-2) group (or) even lesser group (n-3, (n-4) etc... completely shields i.e. to an extent of 1.00 unit of nuclear charge. If the electron of interest belongs to d or f orbital, then each of electron left of the group of electron of interest shields to an extent of 1.00 unit of nuclear charge.

Step 4 :

Summation of the shielding effect of all the electrons gives the shielding constant 'S'

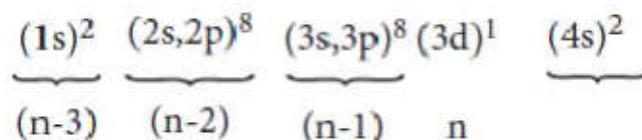
Example: Let us explain the calculation of effective nuclear charge on 4s electron and 3d electron in scandium. The electronic configuration of scandium is $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^1$. we can rearrange as below.



Group	number of electron in the group	contribution of each electron to 'S' value	contribution of a particular group to electrons to 'S' value
(n)	1	0.35	0.35
(n-1)	9	0.85	7.65
(n-2) & others	10	1	10.00
S value			18.00

$$Z_{\text{eff}} = Z - S \text{ i.e.} = 21 - 18 \therefore Z_{\text{eff}} = 3$$

Calculation of effective nuclear charge on 3d electron



Group	number of electron in the group	contribution of each electron to 'S' value	contribution of a particular group to electrons to 'S' value
n	0	0.35	0
(n-1) & others	18	1	18
S Value			18

$$\therefore Z_{\text{eff}} = Z - S \text{ i.e.} = 21 - 18 \therefore Z_{\text{eff}} = 3$$

Shielding effect from inner shell electrons (Slater's rules)

Electron Group	Electron of interest either S or P	Electron of interest either d or f
n	0.35 (0.30 for (S electron))	0.35
(n-1)	0.85	1.00
(n-2) and others	1.00	1.00

Atomic radius (covalent radius) of second period elements.

Elements	Effective nuclear charge	Covalent radius (pm)
Li ³	1.30	167
Be ⁴	1.95	112
C ⁶	2.60	87
N ⁷	3.25	67
O ⁸	3.25	56
F ⁹	4.55	48
Ne ¹⁰	5.85	38

Variation in Group

In the periodic table, the atomic radius of elements increases down the group. As we move down a group, new shells are opened to accommodate the newly added valence electrons. As a result, the distance between the centre of the nucleus and the outermost shell containing the valence electron increases. Hence, the atomic radius increases. The trend in the variation of the atomic radius of the alkali metals down the group is shown below.

Variation of covalent radius of group 1 elements

Element	Outermost shell containing valence electron	Covalent radius (Å)
Li	L (n=2)	1.35
Na	M (n=3)	1.54
K	N (n=4)	1.96
Rb	O (n=5)	2.11
Cs	P(n=6)	2.25

Ionic radius

It is defined as the distance from the centre of the nucleus of the ion up to which it exerts its influence on the electron cloud of the ion. Ionic radius of univalent crystal can be calculated using Pauling's method from the inter ionic distance between the nuclei of the cation and anion. Pauling assumed that ions

present in a crystal lattice are perfect spheres, and they are in contact with each other therefore,

Where d is the distance between the centre of the nucleus of cation C^+ and anion A^- and r_{C^+} , r_{A^-} are the radius of the cation and anion respectively.

Pauling also assumed that the radius of the ion having noble gas electronic configuration (Na^+ and Cl^- having $1s^2 2s^2, 2p^6$ configuration) is inversely proportional to the effective nuclear charge felt at the periphery of the ion.

$$i.e. r_{C^+} + \alpha \frac{1}{(Z_{eff})_{C^+}} \dots\dots\dots(1)$$

$$r_{A^-} - \alpha \frac{1}{(Z_{eff})_{A^-}} \dots\dots\dots(3)$$

Where Z_{eff} is the effective nuclear charge and $Z_{eff} = Z - S$

Dividing the equation 1 by 3

$$\frac{r_{C^+}}{r_{A^-}} = \frac{(Z_{eff})_{A^-}}{(Z_{eff})_{C^+}} \dots\dots\dots(4)$$

On solving equation (1) and (4) the values of r_{C^+} and r_{A^-} can be obtained

Let us explain this method by calculating the ionic radii of Na^+ and F^- in NaF crystal whose interionic distance is equal to 231 pm .

$$d = r_{Na^+} + r_{F^-} \dots\dots\dots(5)$$

$$i.e. r_{Na^+} + r_{F^-} = 231 pm$$

we know that

$$\frac{r_{Na^+}}{r_{F^-}} = \frac{(Z_{eff})_{F^-}}{(Z_{eff})_{Na^+}}$$

$$(Z_{eff})_{F^-} = Z - S$$

$$= 9 - 4.15$$

$$= 4.85$$

$$(Z_{eff})_{Na^+} = 11 - 4.15$$

$$= 6.85$$

$$\therefore \frac{r_{Na^+}}{r_{F^-}} = \frac{4.85}{6.85}$$

$$= 0.71$$

$$\Rightarrow r_{Na^+} = 0.71 \times r_{F^-}$$

Substituting (3) in (1)

$$(1) \Rightarrow 0.71r_{F^-} + r_{F^-} = 231 \text{ pm}$$

$$1.71r_{F^-} = 231 \text{ pm}$$

$$r_{F^-} = \frac{231}{1.71} = 135.1 \text{ pm}$$

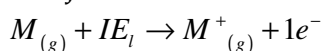
Substituting the value of r_{F^-} in equation (1)

$$r_{Na^+} + 135.1 = 231$$

$$r_{Na^+} = 95.9 \text{ pm}$$

Ionisation energy

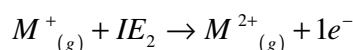
It is defined as the minimum amount of energy required to remove the most loosely bound electron from the valence shell of the isolated neutral gaseous atom in its ground state. It is expressed in kJ mol⁻¹ or in electron volts (eV).



Where IE_1 represents the first ionisation energy.

Successive Ionisation energies

The minimum amount of energy required to remove an electron from a unipositively charged ion is called second ionisation energy. It is represented by the following equation.



In this way we can define the successive ionisation energies such as third, fourth etc.

The total number of electrons are less in the cation than the neutral atom while the nuclear charge remains the same. Therefore the effective nuclear charge of the cation is higher than the corresponding neutral atom. Thus the successive ionisation energies, always increase in the following order

$$IE_1 < IE_2 < IE_3 < \dots$$

Periodic Trends in Ionisation Energy

The ionisation energy usually increases along a period with few exceptions. As discussed earlier, when we move from left to right along a period, the valence electrons are added to the same shell, at the same time protons are added to the nucleus. This successive increase of nuclear charge increases the electrostatic attractive force on the valence electron and more energy is required to remove the valence electron resulting in high ionisation energy.

Let us consider the variation in ionisation energy of second period elements. The plot of atomic number vs ionisation energy is given below.

In the following graph, there are two deviations in the trends of ionisation energy. It is expected that boron has higher ionisation energy than beryllium since it has higher nuclear charge. However, the actual ionisation energies of beryllium and boron are 899 and 800 kJ mol⁻¹ respectively contrary to the expectation. It is due to the fact that beryllium with completely filled 2s orbital, is more stable than partially filled valence shell electronic configuration of boron. (2s², 2p¹)

The electronic configuration of beryllium (Z=4) in its ground state is 1s², 2s² and that of boron (Z = 5) 1s² 2s² 2p¹

Similarly, nitrogen with 1s², 2s², 2p³ electronic configuration has higher ionisation energy (1402 kJ mol⁻¹) than oxygen (1314 kJ mol⁻¹). Since the half filled electronic configuration is more stable, it requires higher energy to remove an electron from 2p orbital of nitrogen. Whereas the removal of one 2p electron from oxygen leads to a stable half filled configuration. This makes it comparatively easier to remove 2p electron from oxygen.

Periodic variation in group

The ionisation energy decreases down a group. As we move down a group, the valence electron occupies new shells, the distance between the nucleus and the valence electron increases. So, the nuclear forces of attraction on valence electron decreases and hence ionisation energy also decreases down a group.

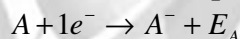
Ionisation energy and shielding effect

As we move down a group, the number of inner shell electron increases which in turn increases the repulsive force exerted by them on the valence electrons, i.e. the increased shielding effect caused by the inner electrons decreases the attractive force acting on the valence electron by the nucleus. Therefore the ionisation energy decreases.

Let us understand this trend by considering the ionisation energy of alkali metals.

Electron affinity

It is defined as the amount of energy released (required in the case noble gases) when an electron is added to the valence shell of an isolated neutral gaseous atom in its ground state to form its anion. It is expressed in kJ mol^{-1}



Variation of Electron Affinity in a period:

The variation of electron affinity is not as systematic as in the case of ionisation energy. As we move from alkali metals to halogens in a period, generally electron affinity increases, i.e. the amount of energy released will be more. This is due to an increase in the nuclear charge and decrease in size of the atoms. However, in case of elements such as beryllium ($1s^2, 2s^2$), nitrogen ($1s^2, 2s^2, 2p^3$) the addition of extra electron will disturb their stable electronic configuration and they have almost zero electron affinity.

Variation of electron affinity (electron gain energy) along I period

Noble gases have stable ns^2, np^6 configuration, and the addition of further electron is unfavourable and requires energy. Halogens having the general electronic configuration of ns^2, np^5 readily accept an electron to get the stable noble gas electronic configuration (ns^2, np^6), and therefore in each period the halogen has high electron affinity. (high negative values)

Variation of Electron affinity in a group:

As we move down a group, generally the electron affinity decreases. It is due to increase in atomic size and the shielding effect of inner shell electrons. However,

oxygen and fluorine have lower affinity than sulphur and chlorine respectively. The sizes of oxygen and fluorine atoms are comparatively small and they have high electron density. Moreover, the extra electron added to oxygen and fluorine has to be accommodated in the 2p orbital which is relatively compact compared to the 3p orbital of sulphur and chlorine so, oxygen and fluorine have lower electron affinity than their respective group elements sulphur and chlorine.

Electronegativity

It is defined as the relative tendency of an element present in a covalently bonded molecule, to attract the shared pair of electrons towards itself.

Electronegativity is not a measurable quantity. However, a number of scales are available to calculate its value. One such method was developed by Pauling, he assigned arbitrary value of electronegativities for hydrogen and fluorine as 2.2 and 4.0 respectively. Based on this the electronegativity values for other elements can be calculated using the following expression

$$(X_A - X_B) = 0.182\sqrt{E_{AB} - (E_{AA} * E_{BB})^{1/2}}$$

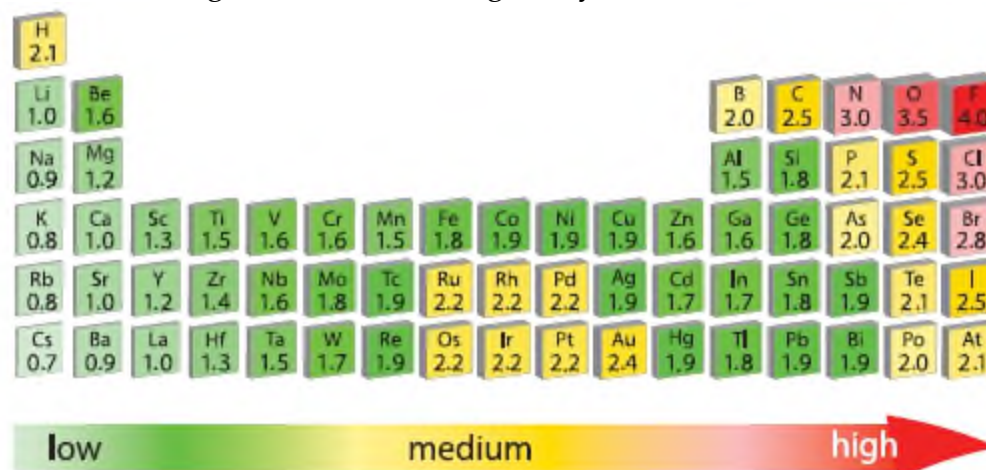
Where E_{AB} , E_{AA} and E_{BB} are the bond dissociation energies of AB, A_2 and B_2 molecules respectively.

The electronegativity of any given element is not a constant and its value depends on the element to which it is covalently bound. The electronegativity values play an important role in predicting the nature of the bond.

Variation of Electronegativity in a period:

The electronegativity generally increases across a period from left to right. As discussed earlier, the atomic radius decreases in a period, as the attraction between the valence electron and the nucleus increases. Hence the tendency to attract shared pair of electrons increases. Therefore, electronegativity also increases in a period.

Paulings scale of electronegativity value of elements



Variation of Electronegativity in a group:

The electronegativity generally decreases down a group. As we move down a group the atomic radius increases and the nuclear attractive force on the valence electron decreases. Hence, the electronegativity decreases.

Noble gases are assigned zero electronegativity. The electronegativity values of the elements of s-block show the expected decreasing order in a group. Except 13th and 14th group all other p-block elements follow the expected decreasing trend in electronegativity.

Periodic Trends in Chemical Properties:

So far, we have studied the periodicity of the physical properties such as atomic radius, ionisation enthalpy, electron gain enthalpy and electronegativity. In addition, the chemical properties such as reactivity, valence, oxidation state etc... also show periodicity to certain extent.

In this section, we will discuss briefly about the periodicity in valence (oxidation state) and anomalous behaviour of second period elements (diagonal relationship).

Valence or Oxidation States

The valence of an atom is the combining capacity relative to hydrogen atom. It is usually equal to the total number of electrons in the valence shell or equal to eight minus the number of valence electrons. It is more convenient to use oxidation state in the place of valence.

Periodicity of Valence or Oxidation States

The valence of an atom primarily depends on the number of electrons in the valence shell. As the number of valence electrons remains same for the elements in same group, the maximum valence also remains the same. However, in a period the number of valence electrons increases, hence the valence also increases.

Variation of valence in groups

Alkali Metals (Group 1)			Group 15		
Element	No. of electrons in valence shell	Valence	Element	No. of electrons in valence shell	Valence
Li	1	1	N	5	3,5
Na	1	1	P	5	3,5
K	1	1	As	5	3,5
Rb	1	1	Sb	5	3,5
Cs	1	1	Bi	5	3,5
Fr	1	1			

Variation of valence in period (1st period)

Element	Li	Be	B	C	N	O	F	Ne
No. of electrons in valence shell	1	2	3	4	5	6	7	8
Valence (Combining capacity)	1	2	3	4	5,3	6,2	7,1	8,0

In addition to that some elements have variable valence. For example, most of the elements of group 15 which have 5 valence electrons show two valences 3 and 5. Similarly transition metals and inner transition metals also show variable oxidation states.

Anomalous properties of second period elements:

As we know, the elements of the same group show similar physical and chemical properties. However, the first element of each group differs from other members of the group in certain properties. For example, lithium and beryllium

form more covalent compounds, unlike the alkali and alkali earth metals which predominantly form ionic compounds. The elements of the second period have only four orbitals (2s & 2p) in the valence shell and have a maximum co-valence of 4, whereas the other members of the subsequent periods have more orbitals in their valence shell and shows higher valences. For example, boron forms BF_4^- and aluminium forms AlF_6^{3-} .

Diagonal Relationship

On moving diagonally across the periodic table, the second and third period elements show certain similarities. Even though the similarity is not same as we see in a group, it is quite pronounced in the following pair of elements.

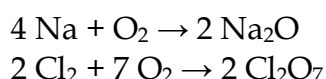
The similarity in properties existing between the diagonally placed elements is called 'diagonal relationship'.

Periodic Trends and Chemical Reactivity:

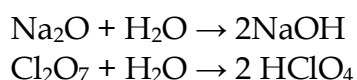
The physical and chemical properties of elements depend on the valence shell electronic configuration as discussed earlier. The elements on the left side of the periodic table have less ionisation energy and readily lose their valence electrons. On the other hand, the elements on right side of the periodic table have high electron affinity and readily accept electrons. As a consequence of this, elements of these extreme ends show high reactivity when compared to the elements present in the middle. The noble gases having completely filled electronic configuration neither accept nor lose their electron readily and hence they are chemically inert in nature.

The ionisation energy is directly related to the metallic character and the elements located in the lower left portion of the periodic table have less ionisation energy and therefore show metallic character. On the other hand the elements located in the top right portion have very high ionisation energy and are non-metallic in nature.

Let us analyse the nature of the compounds formed by elements from both sides of the periodic table. Consider the reaction of alkali metals and halogens with oxygen to give the corresponding oxides.



Since sodium oxide reacts with water to give strong base sodium hydroxide, it is a basic oxide. Conversely Cl_2O_7 gives strong acid called perchloric acid upon reaction with water So, it is an acidic oxide.



Thus, the elements from the two extreme ends of the periodic table behave differently as expected.

As we move down the group, the ionisation energy decreases and the electropositive character of elements increases. Hence, the hydroxides of these elements become more basic. For example, let us consider the nature of the second group hydroxides:

Be(OH)₂ amphoteric; Mg(OH)₂ weakly basic; Ba(OH)₂ strongly basic
Beryllium hydroxide reacts with both acid and base as it is amphoteric in nature.

