GATE Material for Chemistry

Chemistry

- Introduction
- Physical Chemistry
- Organic Chemistry
- Inorganic Chemistry
- Appendix A Conversion Factors
- Appendix B Values of some Constants
- Appendix C Basic and Derived S. I. Units

1. **INTRODUCTION TO CHEMISTRY:**

Science has been classified into various branches for simplification of study. It is broadly classified into **Chemistry**, **Biology** and **Physics**. Mathematics is not branched as a science but considered as the basis of all science.

Chemistry may be defined as the science that is concerned with the characterization and classification of matter and the chemical transformations which it undergoes. However, this may not be an adequate definition as each new chemical concept stimulates fresh observation and experimentation leading to growing refinement as well as development of other concepts.

From the above it follows that, an understanding of chemistry starts with characterization (i.e. composition and structure) of different types of matter. As all matter (available naturally or synthetically prepared) is made up of atoms of different elements, study of elements and its atoms constitutes a fundamental branch of chemistry. The chemical behavior of matter can be explained using atomic structure as its basis. Similarly the physical behavior of matter provides information regarding molecular structure.

Chemistry is further divided into **three branches -**

1.	**Organic Chemistry**	Study of carbon compounds.
2.	**Inorganic Chemistry**	Study of elements (metals or non-metals) and compounds other than those studied in organic chemistry.
3.	**Physical Chemistry**	Deals with structure of matter and the laws governing chemical reactions of organic or inorganic matter. (i.e. chemical bonding, electrolysis, gas laws, chemical thermodynamics etc.)

2. **Physical Chemistry:**

1. Classification of Elements:

a.ELEMENTS:

An **element** is a pure substance that cannot be decomposed into any pure substance and which forms basic building blocks of any matter.

Classification of Elements

Class	Position in periodic table	Properties	Examples
Noble gases	Group 'O'	• Gases • Colorless • Chemically Unreactive • Diamagnetic (weakly repelled by magnetic field)	Neon Argon Krypton
Represe-ntative elements	Group 'A'	• Metallic or nonmetallic • Wide range of chemical behavior • Range of physical characteristics • Diamagnetic or paramagnetic (drawn in magnetic field)	Nitrogen Bromine Boron
Transition elements	Group 'B'	• Metallic • Paramagnetic • Forms highly colored compounds	Titanium Iron Zinc
Inner Transition elements	Bottom of periodic table	• Properties similar to elements of group III B • Covered over two periods 6th and 7^{th}	Thorium

		• 6th period elements are called Lanthanide series • 7th period elements are called Actinide Series • Metallic • In general paramagnetic	Terbium Curium

Macro and Micro Economics:

Introduction to the Periodic Table

The Periodic table is an arrangement of elements in the order of their increasing atomic numbers to show that elements have related properties. Earlier tables, such as those of **Dmitry Mandeleev** (photograph) (1869) and **Lother Meyer** (1869-70) were based on atomic weights which are measured as bulk properties and valency relationships. At that time the concept of atomic numbers was unknown. The main purpose of the periodic table

was:

1. Classification elements into groups with similar properties.

2. To predict the possibilities of new elements based on their properties.

The **modern form of the periodic law** states that properties of the elements are the periodic function of their atomic numbers and the properties of the elements depend on their electronic configuration.

The modern periodic table consists of arrangements of elements in three broad categories.

- metals (located on LHS)

- non-metals (located on RHS)

- metalloids (in between)

Atomic Structure:

ATOMIC STRUCTURE:

Dalton and his contemporaries considered the atom as the smallest indivisible particle of matter. However, the concept of indivisibility was proved to be incorrect by many experiments and observed facts.

Constituents of Atom

The atom is further made of smaller particles named as subatomic or ultimate particles. The atom has been split by physicists mainly into three subatomic particles namely **protons**, **neutrons** and **electrons** (Table 1).

Table 1 Fundamental Units of Atom

Particle	Relative wt.	Electric charge	Atomic location	Symbol
Proton	1.0073	1	nucleus	+ or p
Neutron	1.0087	0	nucleus	n
Electron	0.0005	-1	orbital	- or e-

A **proton** carries a positive charge. It is placed in the nucleus of an atom and has mass nearly equal to 1 a.m.u.

A **neutron** is chargeless. It is placed in the nucleus of an atom and has mass also nearly equal to 1 a.m.u.

An **electron** is negatively charged. It is placed in the orbital and has negligible mass.

Isotopes:

Different atoms of the same element possessing different atomic masses but having same atomic number are known as **Isotopes.** Since the isotopic atoms have the same atomic number, they must contain an equal number of protons.

As their atomic masses are different, there must be a difference in the number of neutrons they possess. They will show similar chemical properties but their physical properties will be different due to different masses. Also since all isotopes of the same element have the same atomic number they occupy the same place in the periodic table.

Fractional atomic masses can be explained from the study of isotopes.

Chlorine has two major isotopes having atomic masses 35 and 37 and present in the ratio 3 : 1. Hence atomic weight of chlorine is neither 35 or 37 but the average of the two.

i.e. atomic weight of chlorine = $\frac{(35 \times 3) + (37 \times 1)}{4}$

= 35.5

Examples of isotopes of some common elements are illustrated in following table.

Table 3 Isotopic Elements						
Element	Atomic No.	Mass No.	Symbol	No. of Protons	No. of Neutrons	Abudance
Hydrogen	1	1.0078	$_1H^1$	1	0	99.99%

Deuterium	1	2.0141	$_1D^2$	1	1	0.02%
Tritium	1	3.0161	$_1T^3$	1	2	rare
Oxygen	8	16	$_8O^{16}$	8	8	99.76%
	8	17	$_8O^{17}$	8	9	0.04%
	8	18	$_8O^{18}$	8	10	0.20%
Chlorine	17	34.97	$_{17}C1^{35}$	17	18	76%
	17	36.97	$_{17}C1^{37}$	17	20	26%
Neon	16	20	$_{16}Ne^{20}$	16	4	90%
	16	22	$_{16}Ne^{22}$	16	6	10%

The above table can be used in the following way to find out the atomic weight of naturally occurring elements.

Example :

$$\text{Hydrogen} = \frac{\text{\% Natural abundance}}{100} \times \text{Atomic weight}$$

$$^1H$$

$$+ \frac{\text{\% Natural abundance}}{100} \times \text{Atomic weight}$$

$$^2H$$

$$= \left(\frac{99.985}{100} \times 1.0078\right) + \left(\frac{0.015}{100} \times 2.0141\right)$$

$$= 1.0079$$

Here tritium is not considered as its abundance is very less.

Here tritium is not considered as its abundance is very less.

Chemical Bonding:

CHEMICAL BONDAGE:

Introduction

The union of two or more atoms involving redistribution of electrons either by transfer or sharing between themselves, so that all of them acquire the stable noble gas configuration of minimum energy is known as **chemical bonding**.

The attraction between atoms within a molecule is called a **chemical bond**.

Electronic theory of valency :

The electronic theory leads to the following types of combination between atoms.

1) Ionic or Electrovalent Bond :
Involving transfer of one or more electrons from one atom to another.
2) Covalent Bond :
Mutual sharing of electrons between two atoms.

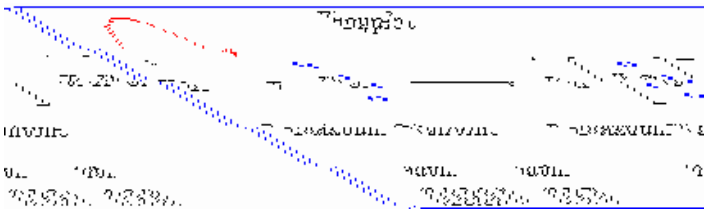
Octet rule : Formation of many ionic and covalent compounds can be explained on the basis of the octet rule which states that "Atoms of elements tend to acquire an octet of electrons (i.e. eight electrons) in their outermost shell through combination with other atoms." In doing so they tend to achieve electronic configuration of the nearest inert gas.

However there are some limitations with the octet rule as it fails to explain incomplete and expanded octets.

The Ionic Bonds:

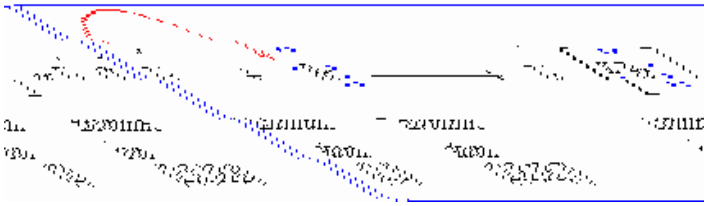
An ionic bond is formed by transfer of electrons from one atom to another atom in the molecule.

Consider the formation of potassium chloride. An electron is transferred from potassium atom to chlorine atom. The resulting K^+ and Cl- ions possessing configuration of argon (2,8,8) combine to form an electrovalent compound as shown below :



Figures in the parenthesis indicate electron distribution in different orbitals (Electronic configuration) in ground state.

Another example :



The total number of electrons lost by Lithium must equal the total number of electrons gained by Bromine. Thus the number of Lithium ions produced is the same as the number of bromide ions produced. These ions attract each other to form crystals as shown below.

States of Matter:

The Physical Status of Matter

There are mainly three different kinds of the physical states of matter, namely:
- Solid
- Liquid
- Gas

Plasma a fourth class of matter has also been identified. These states of matter are also termed as phases.

1) Solid : Solids are characterized by their definite shape and also their considerable mechanical strength and rigidity. Solids tend to resist the deformation of their shape due to strong intra molecular forces and absence of the translatory motion of the structural units (atoms, ions etc). A solid is relatively non compressible, i.e. temperature and pressure have only a slight effect on its volume.

Solids are broadly classified as crystalline or amorphous.

Crystalline solids : Here the atoms are arranged in a definite pattern which is constantly

repeated.

Amorphous solids : These have no definite geometrical form.

2) Liquid : A liquid has no definite shape and it takes the shape of the vessel containing it. Like solids, the volume of a liquid is slightly altered by variations in temperature and pressure. Liquids have three typical physical properties, namely:

i) Vapor pressure : A Liquid when kept in a closed container vaporizes into the free space above it. The process of vaporization will continue till the equilibrium is reached between liquid and vapor. The pressure at which the liquid and vapor can co-exist is called the vapor pressure of the liquid at a given temperature.

ii) Surface tension : The surface of a liquid is always in a state of tension because a molecule at the surface is attracted towards the bulk by a force much greater than that drawing it toward the vapor where the attracting molecules are more widely spread. Due to this, a certain force is required to penetrate along any line in the surface. This force is called surface tension.

iii) Viscosity : It determines the flow of the liquid. It is the internal friction between layers of the liquid. Higher the rate of friction, greater the viscosity of the liquid and its flow will be retarded. Conversely, a lower rate of friction lessens the viscosity and makes the liquid more fluid.

3) Gases : A gas has no bounding surface at all and will occupy completely any vessel in which it is filled. It has no definite volume or shape and can be easily expanded or compressed.

Laws governing behavior of gases will be dealt with in detail in the next chapter.

Water is the ideal example to show the different states of matter.

Water when cooled to 0^0C becomes solid. When the temperature of solid water is raised it becomes liquid. If the liquid is heated to 100^0C it gets converted to steam or vapor (The Gas phase).

Almost all chemical substances can exist in more than one physical state (phase) depending on external pressure and temperature.

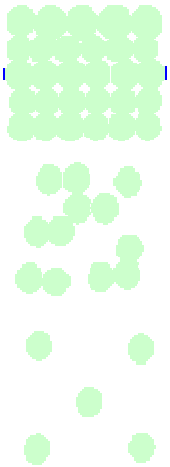
The following table illustrates different states of matter and their physical properties.

Table 9

	Property	States of matter		
		Solid	Liquid	Gas(Vapor)

i)	Shape	definite	indefinite	indefinite
ii)	Volume	definite	definite	indefinite
iii)	Molecular Bonding	very strong	strong	Weak
iv)	Examples	NaCl	H_2O	H_2 , CO_2
		$ZnSO_4$	Petrol	LPG

The Molecular mode of Solids, Liquids, Gases



Solids:- Molecules closely packed
- Strong intermolecular attraction
- Molecules with restricted movement
- Intermolecular space is minimal
- Not compressible

Liquids:- Molecules not very closely held
- Less intermolecular attraction
- Molecules with some movement
- More Intermolecular space
- Compressible

Gases: - Molecules are far apart
- Weak intermolecular attraction
- Molecules with free movement
- More Intermolecular space
- Easily compressible

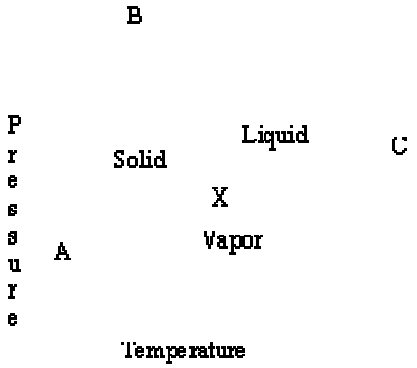
Plasma : This is the fourth state of matter. It is a type of gas containing positively and negatively charged particles in approximately equal numbers and present in the sun and most stars.

Phases

A phase is defined as a physically distinct but homogenous part of a system separated from other parts by boundary surface.

Phase diagrams : The effect of changes in temperature or in pressure or in the composition on the changes of the phase in a system can be understood easily when it is presented graphically. The conditions under which the different phases can co-exist are conveniently described by graphs called **phase diagrams**.

Figure 17



The line XB is the locus of points (P, T) at which a solid and a liquid can co-exist. Similarly co-existence of other phases namely liquid-gas and solid-gas can be represented by lines XC and XA respectively. The point of intersection of any two curves will indicate the co-existence of all the three phases solid-liquid-vapor in equilibrium. This point is called a **triple point**.

Phase diagrams are classified based on the number of components as

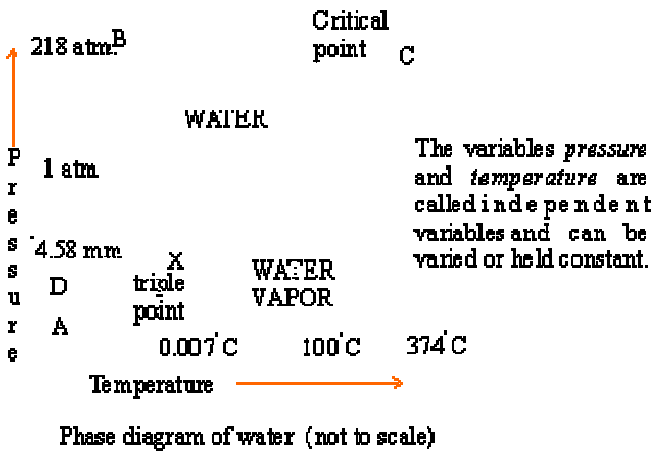
- one component system

- two component system

One component system :

The Water System : The simplest and typical example of a one component system is that of water. Three possible phases depending on pressure and temperature are ice, water and water vapor.

Figure 18



Phase diagram of water (not to scale)

The line XC is the vapor pressure curve for water which increases by increasing temperature.

Each of the three lines XA, XB and XC indicates the **two phase equilibria** and represents **univariant states**. The areas enclosed between these lines represent regions of a single phase, i.e. **bivariant states**. Thus the region AXB is ice, BXC is water and CXA is vapor.

X is the triple point where all the three phases (ice, water and vapor) co-exist. The pressure at this point is found to be 4.58 mm. The freezing point is 0^0C at 760 mm and changes by 1^0 when pressure changes by 140 atm. Hence at 4.58 mm the freezing point is 0.007^0C

Gases: The Gaseous States Properties:

As studied in States of Matter a gas has no bounding surface and thus has no definite volume or shape.

In this chapter, we examine and learn to use the laws governing the behavior of gases. These laws, combined with the other properties of gases, form the basis for the **Kinetic molecular theory of gases**.

Properties of Gaseous State

Following are the **five properties of gases** which can be measured experimentally.

1.	Gases are easily compressible.	The molecules in a gas are very far apart due to very low intermolecular forces. By external pressure they can be brought closer, thereby compressing the gas.
2.	Gases do not have definite volume.	Molecules move easily and occupy the entire volume of the container and take the shape of the container.
3.	Gases exert pressure in all directions.	Molecules are always in a state of rapid zig zag motion, colliding with each other and with the walls of the container. Collisions account for pressure. As the same number of molecules strike a given surface in unit time, gases exert pressure uniformly in all directions.
4.	Gases diffuse easily.	Molecules of other substances can fill a large space between gaseous molecules. This is nothing but diffusion.
5.	Gases have low density.	Due to large space, the number of units of molecules per unit volume of gas is very low compared to solids or liquids.

Gas Laws :

The quantitative relationship between volume, pressure, temperature and the rate of diffusion for a given quantity of gas are termed as ’Gas Laws’.

These Laws are

* Boyle’s law
* Charles’ law
* Pressure-temperature law
* Avogadro’s law
* Graham’s law of diffusion

Boyle’s Law :

Robert Boyle, in 1662, published a mathematical statement on the relation between the volume and pressure of gas at constant temperature called **Boyle’s Law**. This law states that "At constant temperature, volume of a definite mass of a dry gas is inversely proportional to its pressure."



Robert Boyle

It can be mathematically expressed as

$$V \propto \frac{1}{P} \qquad \therefore \qquad V = \frac{\text{constant}}{P}$$

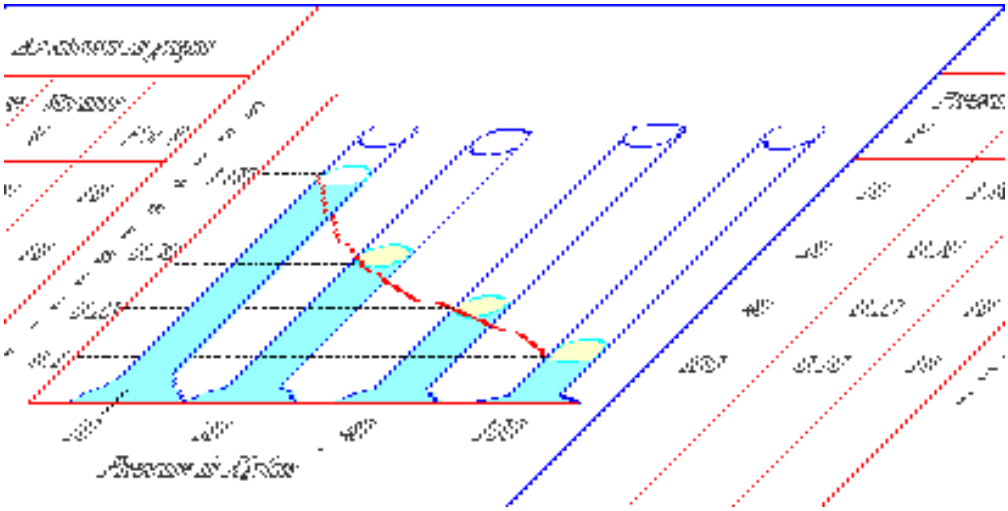
$$\therefore \quad \text{constant} = PV$$

The magnitude of constant depends on temperature, mass and nature of a gas.

Boyle's law can be useful in calculating the volume of a gas at any required pressure if the volume at some other pressure is known by using the following equation.

$P_1V_1 = P_2V_2 = K$

This can be graphically represented as follows :



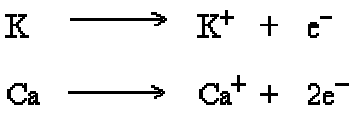
Redox Reactions:

Introduction

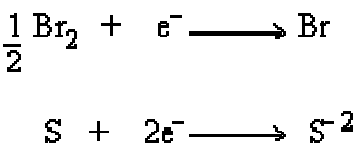
Practically, all chemical reactions are carried out for the purpose of generating energy. Combusion of fuels, generation of electricity by batteries and even metabolism of food are **redox (oxidation-reduction) reactions**.

Oxidation : It is a process in which a substance (an ion or an atom) loses one or

more electrons.



Reduction : It is a process in which a substance (an ion or an atom) gains one or more electrons.



Redox reaction : Electrons are not transferred to the surroundings in ordinary chemical reactions but electrons lost by one species, must be gained by another. Such chemical reactions involving loss of electrons by one species and the simultaneous gain by another constitute (oxidation - reduction) redox reaction.

Any redox reaction comprises two half reactions: one corresponding to oxidation and the other to reduction.

Examples

1) $2Ca + O_2 \rightarrow 2CaO$ Half reaction

$2Ca - 4e^- \rightarrow 2Ca^{+2}$ (Oxidation)

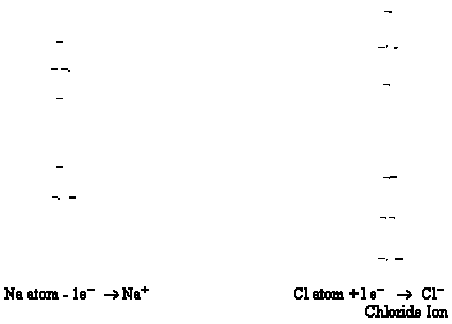
$O_2 + 4e^- \rightarrow 2O^{-2}$ (Reduction)

2) $Na + \frac{1}{2} Cl_2 \rightarrow NaCl$

$Na - e^- \rightarrow Na^+$ (Oxidation)

$\frac{1}{2} Cl_2 + e^- \rightarrow Cl^-$ (Reduction)

A Redox reaction can be represented as



Oxidation Number or Oxidation State :

The **oxidation number** is defined as a positive or negative number that represents a charge that an atom appears to have in a given species when the bonding electrons are counted as per the prescribed set of rules.

Rule 1 : The oxidation number of an atom in an element in its free uncombined state is zero.

Rule 2 : The oxidation number of a monoatomic ion is same as the charge on the ion.

The oxidation number for the Calcium ion is +2 and that of Sulphur in S^{-2} is -2

Rule 3 : Oxidation numbers conventionally assigned to atoms in their chemical compounds are as follows:

a) Oxygen = -2 (except in peroxides -1)
b) Hydrogen = +1 (except in metallic hydrides -1)
c) Group IA elements (alkali metals) = +1
d) Group IIA elements (alkaline earth metals) = +2
e) Halides = -1

Rule 4 : The algebraic sum of positive and negative oxidation numbers in compound is zero.

Rule 5 : The algebraic sum of positive and negative oxidation numbers of atoms in a polyatomic ion is equal to the charge on the ion.

Examples : What is the oxidation number of

1) Mn in Mn_2O_7

2) S in H_2SO_4
3) S in S_8
4) Br in BrO_3^-

Considering ‘x’ as oxidation number in each case.

1) 2 (x) + 7 (-2) = 0
$\therefore$ 2x = 14 and x = 7

2) 2 (+1) + x + 4 (-2) = 0
+ 2 + x - 8 = 0 $\therefore$ x = 6

3) x = 0 (rule 1)

4) x + 3 (-2) = -1
$\therefore$ x = -1 + 6 = + 5

b. Macro and Micro Economics:

Introduction to the Periodic Table

The Periodic table is an arrangement of elements in the order of their increasing atomic numbers to show that elements have related properties. Earlier tables, such as those of **Dmitry Mandeleev** (photograph) (1869) and **Lother Meyer** (1869-70) were based on atomic weights which are measured as bulk properties and valency relationships. At that time the concept of atomic numbers was unknown. The main purpose of the periodic table was:

1. Classification elements into groups with similar properties.

2. To predict the possibilities of new elements based on their properties.

The **modern form of the periodic law** states that properties of the elements are the periodic function of their atomic numbers and the properties of the elements depend on their electronic configuration.

The modern periodic table consists of arrangements of elements in three broad categories.

- metals (located on LHS)

- non-metals (located on RHS)

- metalloids (in between)

2. **Introduction to the Periodic Table**

The Periodic table is an arrangement of elements in the order of their increasing atomic numbers to show that elements have related properties. Earlier tables, such as those of **Dmitry Mandeleev** (photograph) (1869) and **Lother Meyer** (1869-70) were based on atomic weights which are measured as bulk properties and valency relationships. At

that time the concept of atomic numbers was unknown. The main purpose of the periodic table was:

1. Classification elements into groups with similar properties.
2. To predict the possibilities of new elements based on their properties.

The **modern form of the periodic law** states that properties of the elements are the periodic function of their atomic numbers and the properties of the elements depend on their electronic configuration.

The modern periodic table consists of arrangements of elements in three broad categories.

- metals (located on LHS)
- non-metals (located on RHS)
- metalloids (in between)

3. Atomic Structure:

a. The Atom:

b. ATOMIC STRUCTURE

c. Dalton and his contemporaries considered the atom as the smallest indivisible particle of matter. However, the concept of indivisibility was proved to be incorrect by many experiments and observed facts.

d. Constituents of Atom

e. The atom is further made of smaller particles named as subatomic or ultimate particles. The atom has been split by physicists mainly into three subatomic particles namely **protons**, **neutrons** and **electrons** (Table 1).

Table 1 Fundamental Units of Atom

Particle	Relative wt.	Electric charge	Atomic location	Symbol
Proton	1.0073	1	nucleus	+ or p
Neutron	1.0087	0	nucleus	n
Electron	0.0005	-1	orbital	- or e-

f. A **proton** carries a positive charge. It is placed in the nucleus of an atom and has mass nearly equal to 1 a.m.u.
g. A **neutron** is chargeless. It is placed in the nucleus of an atom and has mass also nearly equal to 1 a.m.u.
h. An **electron** is negatively charged. It is placed in the orbital and has negligible mass.

b. Isotopes:

Isotopes

Different atoms of the same element possessing different atomic masses but having same atomic number are known as **Isotopes.** Since the isotopic atoms have the same atomic number, they must contain an equal number of protons.

As their atomic masses are different, there must be a difference in the number of neutrons they possess. They will show similar chemical properties but their physical properties will be different due to different masses. Also since all isotopes of the same element have the same atomic number they occupy the same place in the periodic table.

Fractional atomic masses can be explained from the study of isotopes.

Chlorine has two major isotopes having atomic masses 35 and 37 and present in the ratio 3 : 1. Hence atomic weight of chlorine is neither 35 or 37 but the average of the two.

i.e. atomic weight of chlorine = $\frac{(35 \times 3) + (37 \times 1)}{4}$

= 35.5

Examples of isotopes of some common elements are illustrated in following table.

Table 3 Isotopic Elements						
Element	Atomic No.	Mass No.	Symbol	No. of Protons	No. of Neutrons	Abudance
Hydrogen	1	1.0078	$_1H^1$	1	0	99.99%

Deuterium	1	2.0141	$_1D^2$	1	1	0.02%
Tritium	1	3.0161	$_1T^3$	1	2	rare
Oxygen	8	16	$_8O^{16}$	8	8	99.76%
	8	17	$_8O^{17}$	8	9	0.04%
	8	18	$_8O^{18}$	8	10	0.20%
Chlorine	17	34.97	$_{17}C1^{35}$	17	18	76%
	17	36.97	$_{17}C1^{37}$	17	20	26%
Neon	16	20	$_{16}Ne^{20}$	16	4	90%
	16	22	$_{16}Ne^{22}$	16	6	10%

The above table can be used in the following way to find out the atomic weight of naturally occurring elements.

Example :

$$\text{Hydrogen} = \frac{\text{\% Natural abundance}}{100} \times \text{Atomic weight}$$

$$^1H$$

$$+ \frac{\text{\% Natural abundance}}{100} \times \text{Atomic weight}$$

$$^2H$$

$$= \left(\frac{99.985}{100} \times 1.0078\right) + \left(\frac{0.015}{100} \times 2.0141\right)$$

$$= 1.0079$$

Here tritium is not considered as its abundance is very less.

Here tritium is not considered as its abundance is very less.

c. Atomic Weight:

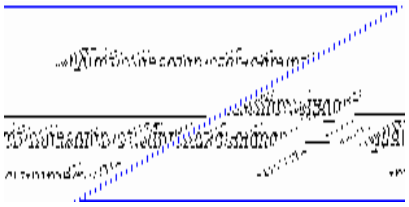
Atomic Weights of An Elements

The **atomic weight** of an element is defined as a number which expresses the ratio of weight of one atom of an element to the weight of one atom of hydrogen.

$$\text{Atomic weight of an element} = \frac{\text{Weight of one atom of an element}}{\text{Weight of one atom of hydrogen}}$$

For e.g. The atomic weight of nitrogen is 14 which implies that an atom of nitrogen is 14 times heavier than one atom of hydrogen.

In recent times, the naturally occurring isotope of carbon C_{12} is taken as standard.



Thus atomic weight is only a number. It has no units and it does not provide any information regarding the actual weight of an atom of an element.

Fractional Atomic Weight :

A large number of naturally occurring elements consists of a mixture of isotopes (atoms of the same elements with the same atomic number but different atomic masses). The properties of these isotopes are constant and hence an element has a fixed atomic weight. The atomic weight of an element thus represents the average of atomic masses of different isotopes of the element. This in some cases leads to fractional atomic weight.

For e.g. Chlorine possesses two isotopes with atomic masses 35 & 37 in the proportion of 3 : 1.

Hence,

$$\text{Atomic weight of Chlorine} = \frac{(3 \times 35) + (1 \times 37)}{4}$$

$$= 35.5$$

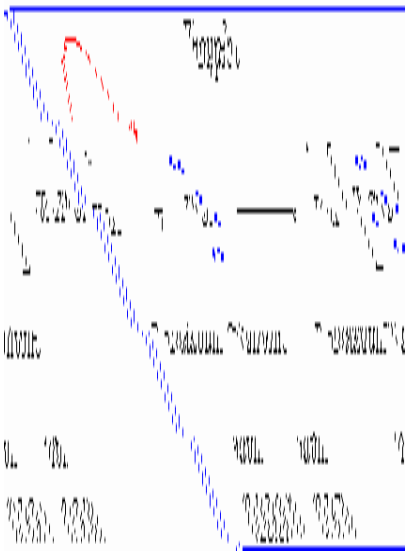
Gram atomic weight :

The atomic weight of an element expressed in grams is known as **gram atomic weight** or **1 gm. atom of an element**.

e.g. 1 gm. atom of Bromine = 79.9 gms. of Bromine

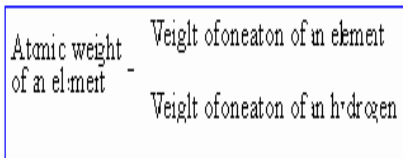
1 gm. atom of Sodium = 22.98 gms of Sodium

At a glance



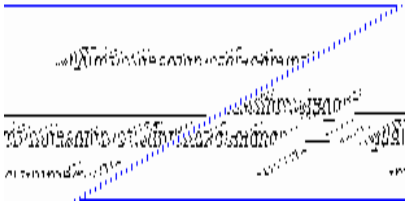
3. **Atomic Weights of An Elements**

The **atomic weight** of an element is defined as a number which expresses the ratio of weight of one atom of an element to the weight of one atom of hydrogen.



For e.g. The atomic weight of nitrogen is 14 which implies that an atom of nitrogen is 14 times heavier than one atom of hydrogen.

In recent times, the naturally occurring isotope of carbon C_{12} is taken as standard.



Thus atomic weight is only a number. It has no units and it does not provide any information regarding the actual weight of an atom of an element.

Fractional Atomic Weight :

A large number of naturally occurring elements consists of a mixture of isotopes (atoms of the same elements with the same atomic number but different atomic masses). The properties of these isotopes are constant and hence an element has a fixed atomic weight. The atomic weight of an element thus represents the average of atomic masses of different isotopes of the element. This in some cases leads to fractional atomic weight.

For e.g. Chlorine possesses two isotopes with atomic masses 35 & 37 in the proportion of 3 : 1.

Hence,

$$\text{Atomic weight of Chlorine} = \frac{(3 \times 35) + (1 \times 37)}{4}$$

$$= 35.5$$

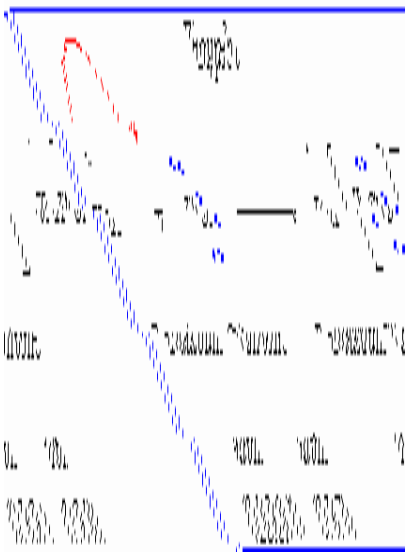
Gram atomic weight :

The atomic weight of an element expressed in grams is known as **gram atomic weight** or **1 gm. atom of an element**.

e.g. 1 gm. atom of Bromine = 79.9 gms. of Bromine

1 gm. atom of Sodium = 22.98 gms of Sodium

At a glance



4. Chemical Bonding:

a.CHEMICAL BONDAGE

Introduction

The union of two or more atoms involving redistribution of electrons either by transfer or sharing between themselves, so that all of them acquire the stable noble gas configuration of minimum energy is known as **chemical bonding**.

The attraction between atoms within a molecule is called a **chemical bond**.

Electronic theory of valency :

The electronic theory leads to the following types of combination between atoms.

1) Ionic or Electrovalent Bond : Involving transfer of one or more electrons from one atom to another. **2) Covalent Bond :** Mutual sharing of electrons between two atoms.

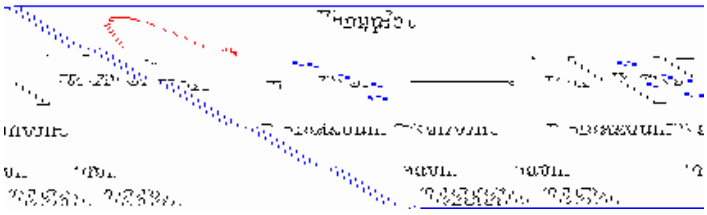
Octet rule : Formation of many ionic and covalent compounds can be explained on the basis of the octet rule which states that "Atoms of elements tend to acquire an octet of electrons (i.e. eight electrons) in their outermost shell through combination with other atoms." In doing so they tend to achieve electronic configuration of the nearest inert gas.

However there are some limitations with the octet rule as it fails to explain incomplete and expanded octets.

b. **The Ionic Bond :**

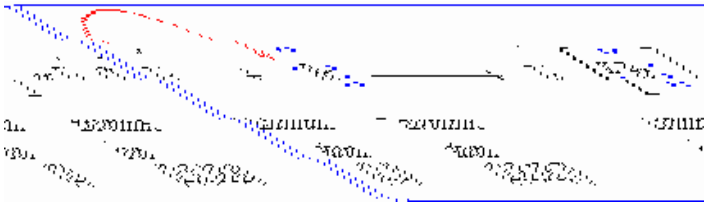
An ionic bond is formed by transfer of electrons from one atom to another atom in the molecule.

Consider the formation of potassium chloride. An electron is transferred from potassium atom to chlorine atom. The resulting K^+ and Cl- ions possessing configuration of argon (2,8,8) combine to form an electrovalent compound as shown below :



Figures in the parenthesis indicate electron distribution in different orbitals (Electronic configuration) in ground state.

Another example :



The total number of electrons lost by Lithium must equal the total number of electrons gained by Bromine. Thus the number of Lithium ions produced is the same as the number of bromide ions produced. These ions attract each other to form crystals as shown below

5. Solids, Liquids and Gases:

a.STATES OF MATTER

The Physical Status of Matter

There are mainly three different kinds of the physical states of matter, namely:
- Solid
- Liquid
- Gas

Plasma a fourth class of matter has also been identified. These states of matter are also termed as phases.

1) Solid : Solids are characterized by their definite shape and also their considerable mechanical strength and rigidity. Solids tend to resist the deformation of their shape due to strong intra molecular forces and absence of the translatory motion of the

structural units (atoms, ions etc). A solid is relatively non compressible, i.e. temperature and pressure have only a slight effect on its volume.

Solids are broadly classified as crystalline or amorphous.

Crystalline solids : Here the atoms are arranged in a definite pattern which is constantly repeated.

Amorphous solids : These have no definite geometrical form.

2) Liquid : A liquid has no definite shape and it takes the shape of the vessel containing it. Like solids, the volume of a liquid is slightly altered by variations in temperature and pressure. Liquids have three typical physical properties, namely:

i) Vapor pressure : A Liquid when kept in a closed container vaporizes into the free space above it. The process of vaporization will continue till the equilibrium is reached between liquid and vapor. The pressure at which the liquid and vapor can co-exist is called the vapor pressure of the liquid at a given temperature.

ii) Surface tension : The surface of a liquid is always in a state of tension because a molecule at the surface is attracted towards the bulk by a force much greater than that drawing it toward the vapor where the attracting molecules are more widely spread. Due to this, a certain force is required to penetrate along any line in the surface. This force is called surface tension.

iii) Viscosity : It determines the flow of the liquid. It is the internal friction between layers of the liquid. Higher the rate of friction, greater the viscosity of the liquid and its flow will be retarded. Conversely, a lower rate of friction lessens the viscosity and makes the liquid more fluid.

3) Gases : A gas has no bounding surface at all and will occupy completely any vessel in which it is filled. It has no definite volume or shape and can be easily expanded or compressed.

Laws governing behavior of gases will be dealt with in detail in the next chapter.

Water is the ideal example to show the different states of matter.

Water when cooled to 0^0C becomes solid. When the temperature of solid water is raised it becomes liquid. If the liquid is heated to 100^0C it gets converted to steam or vapor (The Gas phase).

Almost all chemical substances can exist in more than one physical state (phase) depending on external pressure and temperature.

The following table illustrates different states of matter and their physical properties.

6. **Table 9**

	Property	States of matter		
		Solid	Liquid	Gas(Vapor)
i)	Shape	definite	indefinite	indefinite
ii)	Volume	definite	definite	indefinite
iii)	Molecular Bonding	very strong	strong	Weak
iv)	Examples	NaCl	H_2O	H_2 , CO_2
		$ZnSO_4$	Petrol	LPG

The Molecular mode of Solids, Liquids, Gases

Solids:- Molecules closely packed
- Strong intermolecular attraction
- Molecules with restricted movement
- Intermolecular space is minimal
- Not compressible

Liquids:- Molecules not very closely held
- Less intermolecular attraction
- Molecules with some movement
- More Intermolecular space
- Compressible

Gases: - Molecules are far apart
- Weak intermolecular attraction
- Molecules with free movement
- More Intermolecular space
- Easily compressible

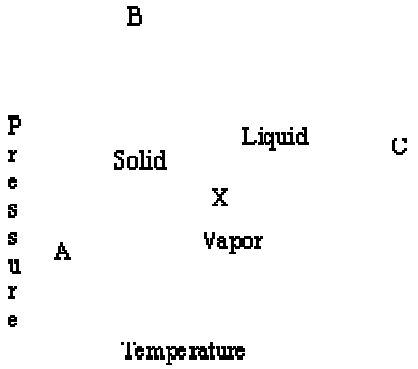
Plasma : This is the fourth state of matter. It is a type of gas containing positively and negatively charged particles in approximately equal numbers and present in the sun and most stars.

b. **Phases**

A phase is defined as a physically distinct but homogenous part of a system separated from other parts by boundary surface.

Phase diagrams : The effect of changes in temperature or in pressure or in the composition on the changes of the phase in a system can be understood easily when it is presented graphically. The conditions under which the different phases can co-exist are conveniently described by graphs called **phase diagrams**.

Figure 17



The line XB is the locus of points (P, T) at which a solid and a liquid can co-exist. Similarly co-existence of other phases namely liquid-gas and solid-gas can be represented by lines XC and XA respectively. The point of intersection of any two curves will indicate the co-existence of all the three phases solid-liquid-vapor in equilibrium. This point is called a **triple point**.

Phase diagrams are classified based on the number of components as

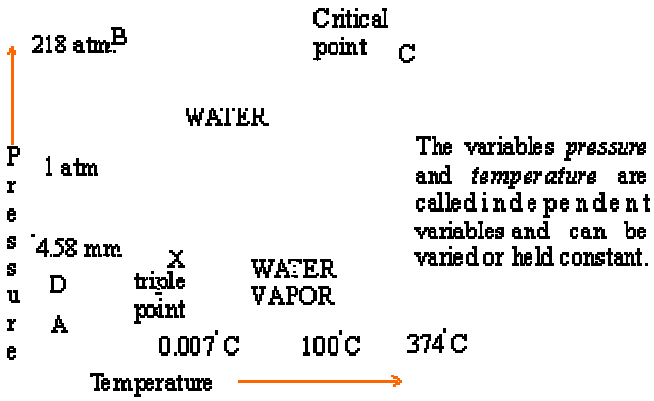
- one component system

- two component system

One component system :

The Water System : The simplest and typical example of a one component system is that of water. Three possible phases depending on pressure and temperature are ice, water and water vapor.

Figure 18



Phase diagram of water (not to scale)

The line XC is the vapor pressure curve for water which increases by increasing temperature.

Each of the three lines XA, XB and XC indicates the **two phase equilibria** and represents **univariant states**. The areas enclosed between these lines represent regions of a single phase, i.e. **bivariant states**. Thus the region AXB is ice, BXC is water and CXA is vapor.

X is the triple point where all the three phases (ice, water and vapor) co-exist. The pressure at this point is found to be 4.58 mm. The freezing point is 0^0C at 760 mm and changes by 1^0 when pressure changes by 140 atm. Hence at 4.58 mm the freezing point is 0.007^0C

Nature of Heat :

When a change in the state of a system occurs, energy is transferred to or from the surroundings. Energy consumed or lost is commonly studied as **heat** which is a measure of motions of molecules in a given system. Heat is best understood in terms of rise or fall of temperature of a system.

The sign convention : According to MKS System dq or q is negative (-ve) if heat is added to the system, dq or q is positive (+ve) if heat is removed from the system where q is the symbol of heat.

Units of heat : The unit of heat, which was used for many years is **calorie (Cal).** It is defined as the quantity of heat required to raise the temperature of one gram of water at 14.5^0C by one degree.

SI unit of heat is the joule (J).

1 joule = 0.2390 calories
1 calorie = 4.184 J

1K cal = 1000 calories
1K cal = 4.184 KJ

Heat Capacity : It is the amount of heat required to raise the temperature of the system by one degree.

If the system is one gram of pure substance it is called **specific heat**, but if one mole of the pure substance is considered, it is called **molar heat capacity**.

Units of Heat Capacity : The usual units of the molar capacity are calorie per degree per mole (Cal K^{-1} mol^{-1}) or joules per degree per mole (J $K^{-1}mol^{-1}$) the latter being the SI unit.

The following chart shows the heat capacities of several elements and compounds in calories per degree per gram as well as per mole and their inter conversion.

Table 10

Elements Compounds	State	Mol.wt/ atomic wt.	Cal deg.$^{-1}$ gram^{-1}	Cal deg.$^{-1}$ mol^{-1} A × mol. Wt
H_2O	Liquid	18.015	1.000	18.015
H_2O	Solid	18.015	0.485	8.737
CaCO3	Solid	100.087	0.205	20.518
CO_2	Gas	44.009	0.158	6.9788
O_2	Gas	31.999	0.203	6.500
MgO	Solid	40.304	0.208	8.383
Al	Solid	26.982	0.213	5.747

Based on the above table let's calculate heat capacity value.

Amount of heat required in calories to heat 100 gm of $H_2O_{(S)}$ from 0.2 to 0.8^{0}C repesented as follows :

= (0.8 - 0.2) deg × 100 grams × 0.485 Cal $deg^{-1}gram^{-1}$

= 0.6 × 100 × 0.485

= 29.1 calories

Similarly heat released into the surrounding while cooling 100 grams of $H_2O_{(S)}$from 0.8 to 0.2^{0}C will be 29% calories.

5. Gases:

a. The Gaseous States Properties:

GASES

As studied in States of Matter a gas has no bounding surface and thus has no definite volume or shape.

In this chapter, we examine and learn to use the laws governing the behavior of gases. These laws, combined with the other properties of gases, form the basis for the **Kinetic molecular theory of gases**.

Properties of Gaseous State

Following are the **five properties of gases** which can be measured experimentally.

1.	Gases are easily compressible.	The molecules in a gas are very far apart due to very low intermolecular forces. By external pressure they can be brought closer, thereby compressing the gas.
2.	Gases do not have definite volume.	Molecules move easily and occupy the entire volume of the container and take the shape of the container.
3.	Gases exert pressure in all directions.	Molecules are always in a state of rapid zig zag motion, colliding with each other and with the walls of the container. Collisions account for pressure. As the same number of molecules strike a given surface in unit time, gases exert pressure uniformly in all directions.
4.	Gases diffuse easily.	Molecules of other substances can fill a large space between gaseous molecules. This is nothing but diffusion.
5.	Gases have low density.	Due to large space, the number of units of molecules per unit volume of gas is very low compared to solids or liquids.

Gas Laws :

The quantitative relationship between volume, pressure, temperature and the rate of diffusion for a given quantity of gas are termed as ’Gas Laws’.

These Laws are

* Boyle’s law:
* Charles’ law
* Pressure-temperature law
* Avogadro’s law
* Graham’s law of diffusion

Boyle’s Law


Robert Boyle

Robert Boyle, in 1662, published a mathematical statement on the relation between the volume and pressure of gas at constant temperature called **Boyle’s Law**. This law states that "At constant temperature, volume of a definite mass of a dry gas is inversely proportional to its pressure."

It can be mathematically expressed as

$$V \propto \frac{1}{P} \quad \therefore \quad V = \frac{\text{constant}}{P}$$

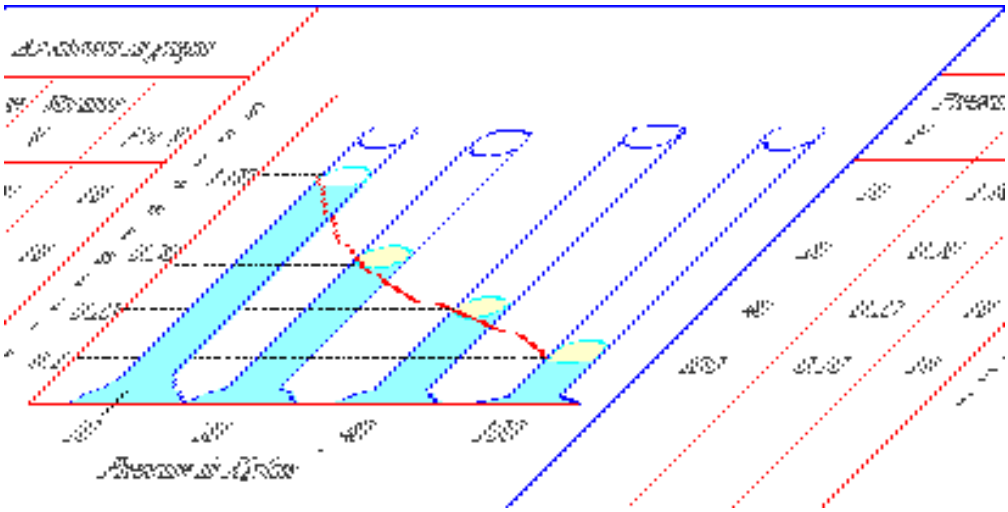
$$\therefore \quad \text{constant} = PV$$

The magnitude of constant depends on temperature, mass and nature of a gas.

Boyle’s law can be useful in calculating the volume of a gas at any required pressure if the volume at some other pressure is known by using the following equation.

$P_1V_1 = P_2V_2 = K$

This can be graphically represented as follows :



Charle's Law:



Jacques Charles

Jacques Charles, in 1787, formulated the relationship between the volume and temperature of a given mass of a dry gas at constant pressure called Charles' Law which states that "At constant pressure, the volume of a fixed mass of dry gas is directly proportional to the absolute temperature;

It can be mathematically expressed as

$$V \propto T \quad \therefore V = \text{constant} \times T$$

$$\therefore \frac{V}{T} = \text{constant}$$

The magnitude of constant depends on pressure, mass and nature of a gas.

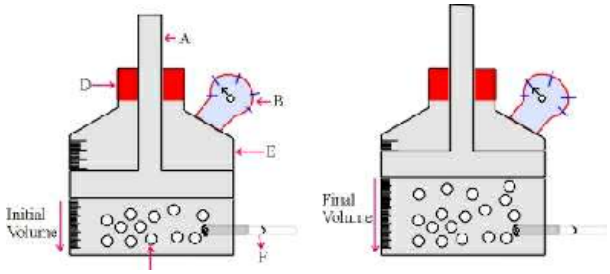
Charles' law is useful for calculating the volume of a gas at any required temperature if the volume at some other temperature is known by using the following equation.

$V_1 = T_1$

$V_2 = T2$

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} = K$$

Illustration of Charles' Law:



A : Piston

B : Pressure gauge

C : Gas Molecule

D : Rubber cork

E : Vessel

F : Thermometer

Example :

A gas at constant pressure is kept at 100^0C. On decreasing the temperature to 50^0C, the gas occupies a volume of 800 ml. Find the initial volume of the gas.

V_1 = ? T_1 = 100^0C = 100 + 273 K = 373 K

V_2 = 800 ml T_2 = 50^0C = 50 + 273 K = 323 K

According to Charles' Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\therefore V_1 = \frac{800 \times 373}{323} = 923.8 \text{ ml}$$

Initial volume of the gas was 923.8 ml.

6. **Charle's Law**



Jacques Charles

Jacques Charles, in 1787, formulated the relationship between the volume and temperature of a given mass of a dry gas at constant pressure called Charles' Law which states that "At constant pressure, the volume of a fixed mass of dry gas is directly proportional to the absolute temperature;

It can be mathematically expressed as

$$V \propto T \quad \therefore V = \text{constant} \times T$$

$$\therefore \frac{V}{T} = \text{constant}$$

The magnitude of constant depends on pressure, mass and nature of a gas.

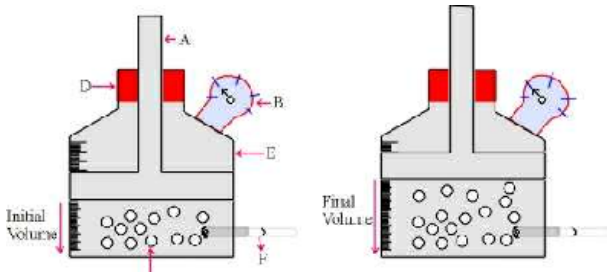
Charles' law is useful for calculating the volume of a gas at any required temperature if the volume at some other temperature is known by using the following equation.

$V_1 = T_1$

$V_2 = T2$

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} = K$$

Illustration of Charles' Law:



A : Piston

B : Pressure gauge

C : Gas Molecule

D : Rubber cork

E : Vessel

F : Thermometer

Example :

A gas at constant pressure is kept at 100^0C. On decreasing the temperature to 50^0C, the gas occupies a volume of 800 ml. Find the initial volume of the gas.

V_1 = ? T_1 = 100^0C = 100 + 273 K = 373 K

V_2 = 800 ml T_2 = 50^0C = 50 + 273 K = 323 K

According to Charles' Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\therefore V_1 = \frac{800 \times 373}{323} = 923.8 \text{ ml}$$

Initial volume of the gas was 923.8 ml.

6. **Charle's Law**



Jacques Charles

Jacques Charles, in 1787, formulated the relationship between the volume and temperature of a given mass of a dry gas at constant pressure called Charles' Law which states that "At constant pressure, the volume of a fixed mass of dry gas is directly proportional to the absolute temperature;

It can be mathematically expressed as

$$V \propto T \quad \therefore V = \text{constant} \times T$$

$$\therefore \frac{V}{T} = \text{constant}$$

The magnitude of constant depends on pressure, mass and nature of a gas.

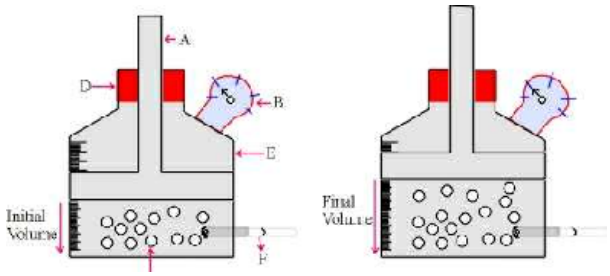
Charles' law is useful for calculating the volume of a gas at any required temperature if the volume at some other temperature is known by using the following equation.

$V_1 = T_1$

$V_2 = T2$

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} = K$$

Illustration of Charles' Law:



A : Piston

B : Pressure gauge

C : Gas Molecule

D : Rubber cork

E : Vessel

F : Thermometer

Example :

A gas at constant pressure is kept at 100^0C. On decreasing the temperature to 50^0C, the gas occupies a volume of 800 ml. Find the initial volume of the gas.

$V_1 = ?$ $T_1 = 100^0C = 100 + 273$ K = 373 K

$V_2 = 800$ ml $T_2 = 50^0C = 50 + 273$ K = 323 K

According to Charles' Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\therefore V_1 = \frac{800 \times 373}{323} = 923.8 \text{ ml}$$

Initial volume of the gas was 923.8 ml.

6.Redox ReactionsP:

REDOX REACTIONS

Introduction

Practically, all chemical reactions are carried out for the purpose of generating energy. Combusion of fuels, generation of electricity by batteries and even metabolism of food are **redox (oxidation-reduction) reactions**.

Oxidation : It is a process in which a substance (an ion or an atom) loses one or more electrons.

$$K \longrightarrow K^+ + e^-$$

$$Ca \longrightarrow Ca^+ + 2e^-$$

Reduction : It is a process in which a substance (an ion or an atom) gains one or more electrons.

$$\frac{1}{2} Br_2 + e^- \longrightarrow Br$$

$$S + 2e^- \longrightarrow S^{-2}$$

Redox reaction : Electrons are not transferred to the surroundings in ordinary chemical reactions but electrons lost by one species, must be gained by another. Such chemical reactions involving loss of electrons by one species and the simultaneous gain by another constitute (oxidation - reduction) redox reaction.

Any redox reaction comprises two half reactions: one corresponding to oxidation and the other to reduction.

Examples

1) $2Ca + O_2 \rightarrow 2CaO$ Half reaction

$2Ca - 4e^- \rightarrow 2Ca^{+2}$ (Oxidation)

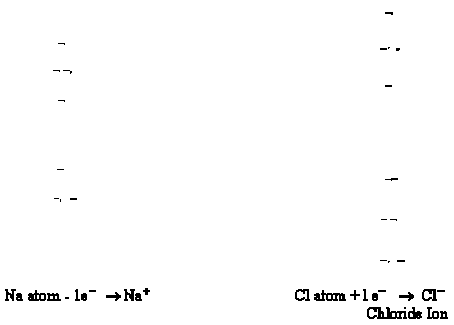
$O_2 + 4e^- \rightarrow 2O^{-2}$ (Reduction)

2) $Na + \frac{1}{2} Cl_2 \rightarrow NaCl$

$Na - e^- \rightarrow Na^+$ (Oxidation)

$\frac{1}{2} Cl_2 + e^- \rightarrow Cl^-$ (Reduction)

A Redox reaction can be represented as



B.Oxidation Number or Oxidation State :

The **oxidation number** is defined as a positive or negative number that represents a charge that an atom appears to have in a given species when the bonding electrons are counted as per the prescribed set of rules.

Rule 1 : The oxidation number of an atom in an element in its free uncombined state is zero.

Rule 2 : The oxidation number of a monoatomic ion is same as the charge on the ion.

The oxidation number for the Calcium ion is +2 and that of Sulphur in S^{-2} is -2

Rule 3 : Oxidation numbers conventionally assigned to atoms in their chemical compounds are as follows:

a) Oxygen = -2 (except in peroxides -1)
b) Hydrogen = +1 (except in metallic hydrides -1)
c) Group IA elements (alkali metals) = +1
d) Group IIA elements (alkaline earth metals) = +2
e) Halides = -1

Rule 4 : The algebraic sum of positive and negative oxidation numbers in compound is zero.

Rule 5 : The algebraic sum of positive and negative oxidation numbers of atoms in a polyatomic ion is equal to the charge on the ion.

Examples : What is the oxidation number of

1) Mn in Mn_2O_7
2) S in H_2SO_4
3) S in S_8
4) Br in BrO_3^-

Considering 'x' as oxidation number in each case.

1) $2(x) + 7(-2) = 0$
$\therefore 2x = 14$ and $x = 7$

2) $2(+1) + x + 4(-2) = 0$
$+2 + x - 8 = 0 \therefore x = 6$

3) $x = 0$ (rule 1)

4) $x + 3(-2) = -1$
$\therefore x = -1 + 6 = +5$

Balancing Redox Reactions :

A chemical reaction is treated as a balanced one when the number of atoms of each element is equal in the reactants and products.

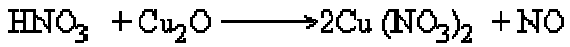
How to balance redox reactions

1) Write the unbalanced equation and identify the elements which undergo change in oxidation number.

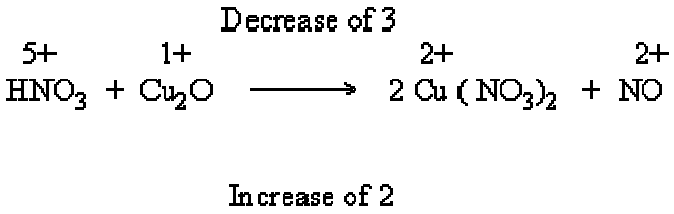
For convenience draw a line between them



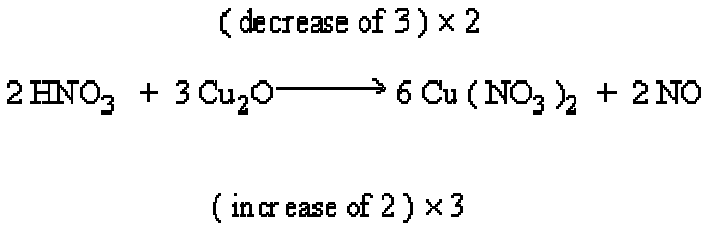
2) Balance the number of atoms of each element undergoing change in oxidation number



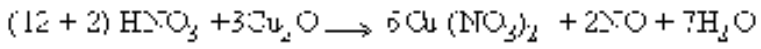
3) Indicate total change in oxidation number



4) Multiply the total charges as given below to equal the gain and loss of electrons



5) Balance the remainder of atoms by inspection, if necessary add acid or base and water as needed.



The balanced equation is :-

$$14HNO_3 + 3Cu_2O \longrightarrow 6cu\ (NO_3)_2 + 2\ NO + 7H_2O$$

Solutions:

Nature of Solutions:

SOLUTIONS

Solutions are homogenous mixtures. They may be classified according to their physical state: gaseous, solid and liquid solutions.

Gaseous solutions, of which air is an example, are mixtures of molecules of two or more gases.

Certain alloys are **solid solutions** e.g. Brass is a solid solution of zinc and copper.

Liquid solutions are the most common.

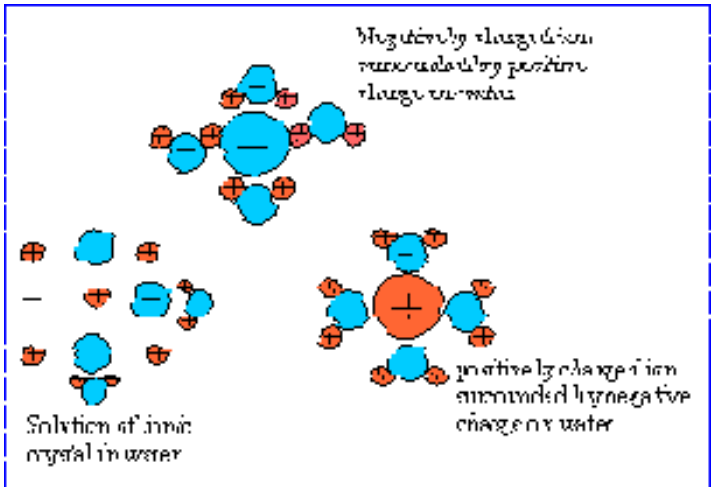
8.1 Nature of Solution

Solutions are obtained when solutes are dissolved in solvents. The component of the solution that is present in greater quantity is called the **solvent** and the other component is called the **solute**.

The **solubility** of a substance in a particular solvent at a specified temperature is the maximum amount of the solute that will dissolve in a definite amount of the solvent and produce a stable system.

For a given solution, the amount of solute dissolved in a unit volume of solution (or unit amount of solvent) is the **concentration** of the solute.

The following figures show the solution of an ionic crystal in water. The ions in the center of the crystal are attracted equally in all directions by oppositely charged ions of the crystal.



Solutions with a low concentration of solute are called **dilute solutions** and those with a high concentration of solute are called **concentrated solutions**.

If an excess of solute is added to the solvent then the equilibrium rate at which the solute dissolves equals to the rate at which the solute **precipitates** in the solution. Hence the amount of solute at a given time remains constant and such solutions are called **saturated solutions**.

When less amount of solute is dissolved in a solvent it becomes an **unsaturated solution**.

Heat of solutions:

Heat of solutions relates to a process in which one mole of solute is dissolved in an infinitely

large quantity of solvent.

Table 11

Heat of solution is the **net energy** required to break apart solute-solute (crystal lattice H positive) and solvent-solvent attraction or bonds and the energy released by the formation of new solute-solvent bond (heat of hydration; H negative).			
Example: *Substance*	*Lattice Energy* Kcal/mol	*Heat of Hydration* Kcal/mol	*Heat of Solution* Kcal/mol
KC1	+ 167.6	- 163.5	+ 4.1
AgF	+ 217.7	- 222.6	- 4.9

Concentration of Solutions:

Concentration of Solutions can be expressed in different ways as follows :

i)**Weight percentage of solute** is the weight of solute
(× 100), divided by the total weight of the solution.

i.e. 10% w/w of NaCl means 10 gms of NaCl is present in a 100 g of solution.

mathematically $\frac{10\times 100}{100} = 10\%\ w/w$

ii) The molarity, M, of a solution is the number of moles of solute per kg of solvent.

20 M of sodium chloride solution means 20 moles (20 × 55.5 gms) of sodium chloride dissolved in a solvent to get one liter of solution.

iii) The molality, m, of a solution is the number of moles of solute per kg of solution.

12 m of sodium chloride solution means 12 moles (12 × 55.5 gms) of sodium chloride dissolved in a solvent to get 1 kg of solution.

iv) The normality, N, of a solution is the number of gram equivalent weight of solute per liter of solution.

5 N of sulphuric acid solution means (5 × 49 gms) (∴ eq. wt = mol.wt / no.of equivalents = 98/2 = 49) of sulphuric acid dissolved in solvent to get one liter of solution.

v) The mole fraction, X, is the ratio of the number of moles of solute (or solvent) to the total number of moles of all components in a solution.

In a 10% HCl solution in H2O

$$X_{HCl} = \frac{\text{No. of moles of HCl}}{\text{Total no. of moles (HCl} + H_2O)}$$

$$= \frac{10/36.5}{(10/36.5) + (90/18)}$$

$$\text{and } X_{H_2O} = \frac{90/18}{(90/18) + (10/36.5)}$$

$$\therefore \quad X_{solvent} + X_{solution} = 1 \text{ always}$$

At a glance the concentration of solutions:

No.	Name	Symbol	Explanation
1	Weight percentage	% w/w	Weight of solute in percentage in 100 gm of solution.
2	Molarity	M	Number of moles of solute per liter of solution.
3	Molality	m	Number of moles of solute per kg of solution.
4	Normality	N	Number of gram equivalent weights of solute per liter of solution.

5	Mole fraction	X	Ratio of the number of moles of solute (or solvent) to the total number of moles of all components in a solution.

Concentration of Solutions can be expressed in different ways as follows :

i)**Weight percentage of solute** is the weight of solute
(× 100), divided by the total weight of the solution.

i.e. 10% w/w of NaCl means 10 gms of NaCl is present in a 100 g of solution.

mathematically $\frac{10 \times 100}{100} = 10\%\ w/w$

ii) The molarity, M, of a solution is the number of moles of solute per kg of solvent.

20 M of sodium chloride solution means 20 moles (20 × 55.5 gms) of sodium chloride dissolved in a solvent to get one liter of solution.

iii) The molality, m, of a solution is the number of moles of solute per kg of solution.

12 m of sodium chloride solution means 12 moles (12 × 55.5 gms) of sodium chloride dissolved in a solvent to get 1 kg of solution.

iv) The normality, N, of a solution is the number of gram equivalent weight of solute per liter of solution.

5 N of sulphuric acid solution means (5 × 49 gms) (∴ eq. wt = mol.wt / no.of equivalents = 98/2 = 49) of sulphuric acid dissolved in solvent to get one liter of solution.

v) The mole fraction, X, is the ratio of the number of moles of solute (or solvent) to the total number of moles of all components in a solution.

In a 10% HCl solution in H2O

$$X_{HCl} = \frac{\text{No. of moles of HCl}}{\text{Total no. of moles (HCl} + H_2O)}$$

$$= \frac{10/36.5}{(10/36.5) + (90/18)}$$

$$\text{and } X_{H_2O} = \frac{90/18}{(90/18) + (10/36.5)}$$

$$\therefore \quad X_{solvent} + X_{solution} = \text{always one}$$

At a glance the concentration of solutions:

No.	Name	Symbol	Explanation
1	Weight percentage	% w/w	Weight of solute in percentage in 100 gm of solution.
2	Molarity	M	Number of moles of solute per liter of solution.
3	Molality	m	Number of moles of solute per kg of solution.
4	Normality	N	Number of gram equivalent weights of solute per liter of solution.
5	Mole fraction	X	Ratio of the number of moles of solute (or solvent) to the total number of moles of all components in a solution.

Electrochemistry:

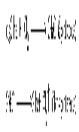
ELECTROCHEMISTRY

Introduction

Electrochemistry is that branch of chemistry which deals with the relationship between chemical energy and electrical energy. Chemical reactions produce electrical energy and conversely, electrical energy can carry out chemical reactions. These transformations take place through the flow of electrons. Electrons are evolved at one place (oxidation), transferred through a conductor (metal wire) and absorbed at another place (reduction).

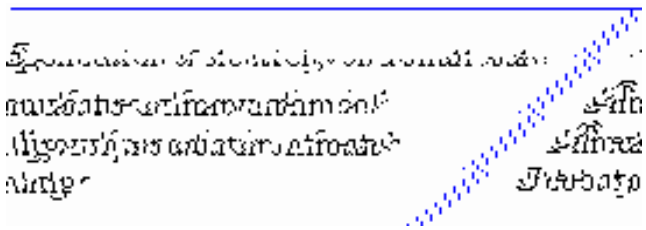
A chemical reaction can generate electrical energy if it takes place spontaneously, when carried out under appropriate conditions. Here conditions mean proper device, concentration and temperature or pressure. The reaction of zinc with copper sulfate is one such example which produces heat energy when reacted upon in a simple vessel. But when the same reaction is carried out in a special device called Daniell cell it produces electrical energy.

Reactions that do not take place spontaneously, can be made to take place by applying an external source of electrical energy.



The former reaction generates heat (energy) whereas the latter reaction requires electrical energy under appropriate conditions.

Electrochemistry finds its application in many fields like



ELECTROCHEMISTRY

Introduction

Electrochemistry is that branch of chemistry which deals with the relationship between chemical energy and electrical energy. Chemical reactions produce electrical energy and conversely, electrical energy can carry out chemical reactions. These transformations take place through the flow of electrons. Electrons are evolved at one place (oxidation), transferred through a conductor (metal wire) and absorbed at another place (reduction).

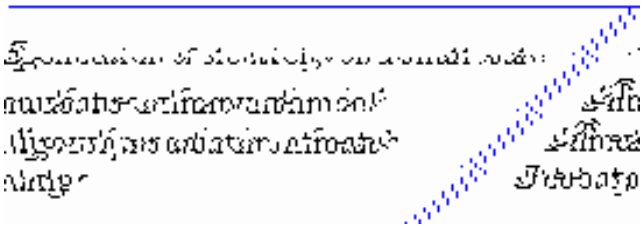
A chemical reaction can generate electrical energy if it takes place spontaneously, when carried out under appropriate conditions. Here conditions mean proper device, concentration and temperature or pressure. The reaction of zinc with copper sulfate is one such example which

produces heat energy when reacted upon in a simple vessel. But when the same reaction is carried out in a special device called Daniell cell it produces electrical energy.

Reactions that do not take place spontaneously, can be made to take place by applying an external source of electrical energy.

The former reaction generates heat (energy) whereas the latter reaction requires electrical energy under appropriate conditions.

Electrochemistry finds its application in many fields like

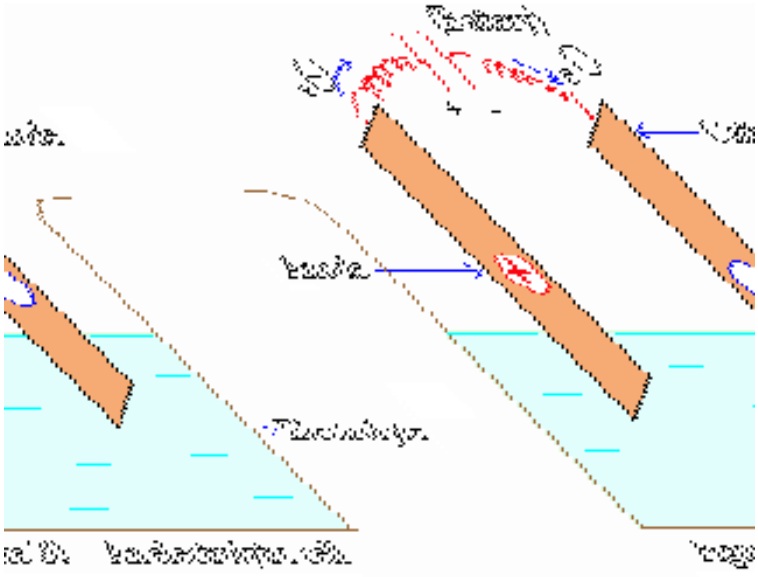


Electrolytic Cell:

Electrolytic Cell

A device which converts chemical energy to electrical energy is called **electrolytic cell**, and the process of conversion is called **electrolysis**. e.g. Voltameter used for electrolysis of water. Electricity is conducted by an electrolyte which is nothing but a liquid solution present in an electrochemical cell.

Electrolytic cell : It consists of a vessel with two electrodes dipped in an electrolyte. The electrode attached to the negative terminal of an external source of energy **(Battery)** is called the **cathode** and the one which connects to the positive terminal is called the **anode**.



- Elements of Chemical Thermodynamic:

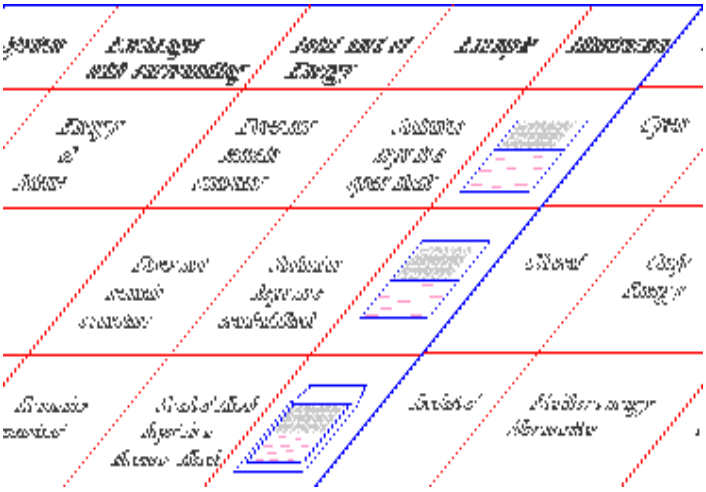
Introduction

Thermodynamics is the study of energy and energy changes. **Energy** is the capacity to work. It is present in various forms like potential energy, kinetic energy, light energy, electrical energy etc.

Energy and energy changes are always calculated in terms of heat energy as there is a natural tendency that all other forms of energy get easily and finally converted into heat energy.

The branch of thermodynamics dealing with energy changes during chemical transformation is called **chemical thermodynamics**.

Types of thermodynamic systems



Homogenous system : A system is homogenous when it has some chemical composition throughout. e.g. mixture of gases or true solution of solid in liquid.

Heterogenous system : Two or more different phases which are homogenous but separated by a boundary. e.g. Ice in water.

The First Law of Thermodynamics: Chemical Equilibrium

The First Law of Thermodynamics is the law of conservation of energy. It can be stated in various ways as follows :

1. Energy can neither be created nor destroyed although it can be converted from one form into another.
2. Whenever a quantity of one form of energy disappears, an equivalent amount of energy of another kind makes its appearance.
3. The total energy of an isolated system remains constant whatever changes may take place in the system.
4. It is not possible for any machine to produce work without consuming energy.
5. The energy of the universe remains constant.

Mathematical Expression of the First Law of Thermodynamics

Let q be the heat absorbed by a system. Part of it may be utilized for increasing the internal energy i.e. ΔE.

$$\Delta E = E_2 - E_1$$

and part for external work.

According to the first law of thermodynamics, the total energy of a system remains constant.

Heat absorbed = Energy change + work done

$q = \Delta E + W$

$\therefore \Delta E = q - W$

If work W is only pressure - volume type

$W = P \Delta V$

$\therefore \Delta E = q - P \Delta V$ or $q = \Delta E + P \Delta V$

Chemical Equilibrium:

CHEMICAL EQUILIBRIUM

Many physical and chemical changes which we encounter are reversible. For e.g. water changes into ice when temperature is lowered; increasing the temperature reverses this process. Conversely, on heating water, it converts into steam; lowering its temperature reverses the process.

$$H_2O\ (solid) \underset{cooling}{\overset{heating}{\rightleftharpoons}} H_2O\ (liquid) \underset{cooling}{\overset{heating}{\rightleftharpoons}} H_2O\ (vapor)$$

When the conditions are such that forward (represented by ⟶) and backward (represented by ⟵)reactions can both occur to a noticeable extent, the process is described as a **reversible reaction**.

It has been found that after a certain time interval, reversible reactions attain a state of chemical equilibrium i.e. a state in which no further change in composition with time can be detected, provided the temperature and pressure are not altered.

11.1 Types of Reactions

a) Reversible reactions

Reactions in which products themselves react (or decompose) to produce the starting reactants, are termed as reversible reactions. Such reactions are denoted by the use of

two half arrows.

e.g. $N_2\,(g) + 3H_2\,(g) \underset{\text{backward reaction}}{\overset{\text{forward reaction}}{\rightleftharpoons}} 2NH_3\,(g)$

e.g. $H_2\,(g) + I_2\,(g) \rightleftharpoons 2HI\,(g)$

For a reverse reaction to occur, the reaction should be carried out in a closed container when one of the products is gaseous.

b) Irreversible reactions

The reactions in which products do not recombine to form the original reactants is called an irreversible reaction.

e.g.

$2KBrO_3 \xrightarrow{\Delta} 2KBr + 3O_2$

$Zn + H_2SO_4\,(dil) \longrightarrow ZnSO_4 + H_2$

c) Exothermic reactions

eg.

$N_2\,(g) + 3H_2\,(g)2 \rightleftharpoons NH_3\,(g) + 96.2\,KJ$

$2SO_2\,(g) + O_2 \rightleftharpoons 2SO_3\,(g) + 188.3\,KJ$ It is a chemical reaction which proceeds with the liberation of heat energy.

d) Endothermic reactions

It is a chemical reaction which proceeds with the absorption of heat energy.

eg. $3O_2 \rightleftharpoons 2O_3 \quad -288.7\,KJ$

Chemical Equilibrium

Consider a reversible homogenous reaction occurring in a closed container

$$A + B \rightleftharpoons C + D$$

Initially the molar concentration of 'A' & 'B' being maximum the rate of forward reaction is maximum. As the reaction proceeds 'C' & 'D' appear in the reaction and the backward reaction sets in but at much lower rate.

The changing rates of forward and backward reactions with time can be represented as follows :

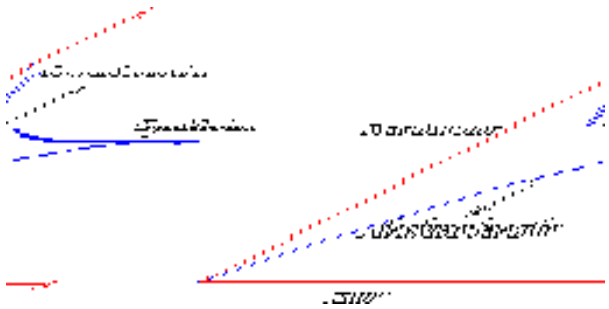


Figure 22 Equilibrium graph

Equilibrium Law Expression

Let us consider a reversible homogenous reaction.

$$A + B \rightleftharpoons C + D$$

According to the law of mass action, the rate of forward reaction (r_f) is

$r_f = K_f [A] [B]$

and that of the backward reaction is

$r_b = K_b [C] [D]$

K_b and K_f are the rate constants.

The rate of the net reaction is :

$r_n = r_f$ - rb

At equilibrium, there is no net reaction i.e.

$r_f - r_b = 0 \therefore r_f = rb$

$\backslash K_f [A][B] = K_b [C][D]$

$$\therefore \quad K_f = \frac{[C][D]}{[A][B]} \quad \therefore \quad K_c = \frac{[C][D]}{[A][B]}$$

K_c = equilibrium constant

e.g. $2SO_2(g) + O_2 \rightleftharpoons 2SO_3(g)$

$$K_c = \frac{[SO_3]^2}{[SO_2]^2\ [O_2]}$$

The equilibrium constant expressed in terms of partial pressures (Kp).

Equilibrium constants for gaseous reaction can be based on the partial pressure of a gas as the partial pressure of a gas is directly proportional to its molar concentration for a given temperature.

if $A(g) + B(g) \rightleftharpoons C(g) + D(g)$

then $$K_p = \frac{P_C \cdot P_D}{P_A \cdot P_B}$$

Acids, Bases and Salts:

ACIDS, BASES AND SALTS

12.1 Lowry and Bronsted Concepts of Acids and Bases

The classification of substances of acids was at first suggested by their sour taste (Latin *acidus* = Sour) and alkali (Arabic *alkali* = ashes of a plant) were taken as those substances that could reverse or neutralize the action of

acids.

The classical definition of

ACID : as a substance whose water solution

i) turns blue litmus red

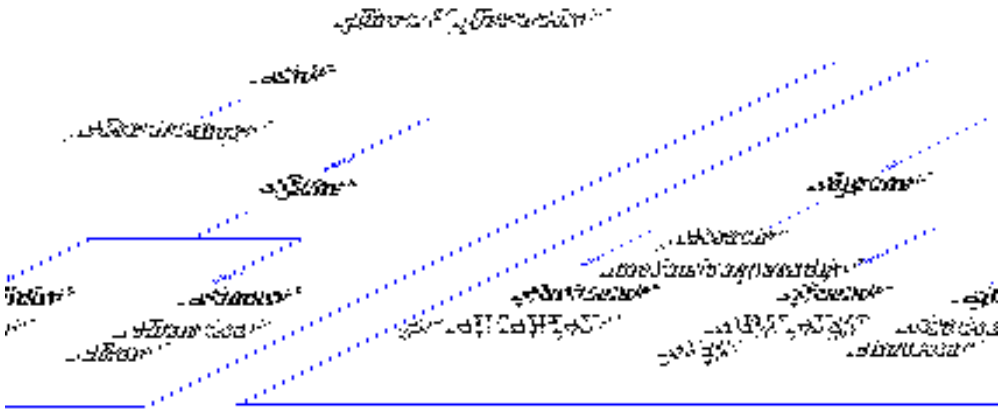
ii) neutralizes base

iii) reacts with active metals like Zn, Mg, Cu etc with evolution of H_2

iv) having sour taste

v) decomposes carbonates

The following chart shows different types of acid.



BASE : as a substance whose water solution

i) turns red litmus blue

ii) neutralizes acid to produce salts

iii) tastes bitter

iv) soapy to touch

v) reacts with few metals like Zn, Al, Pb etc liberating hydrogen.

Examples : NaOH, KOH, NH_4OH

In 1923, J. N. Bronsted and T. M. Lowry proposed definitions of acids and bases in aqueous as well as non aqueous solutions according to which

An acid is defined as a substance having a tendency to lose or to donate one or more protons and,

A base is defined as a substance having tendency to accept or add a proton.

Example :

NH_3 (g) + HCl (g) $\rightleftharpoons$ NH_4^+ + Cl-

Base Acid

Here HCl is proton donor (hence acid) and ammonia is a proton acceptor (hence base)

Other examples are :

1) NH_3 (aq) + H_2O (*l*) $\rightleftharpoons$ NH_4^+ (aq) + OH^- (aq)

Base Acid

2) NH_3 (aq) + H_3O^+ (aq) $\rightleftharpoons$ NH_4^+ (aq) + H_2O (*l*)

Base Acid

Conjugate Acid Base Pair

Acid $\rightleftharpoons$ H^+ (proton) + Base

An acid and a base which differ by a proton are said to form a conjugate acid base pair *or* the pairs of substances which can be formed from one another by the gain or loss of protons are known as conjugate acid base pairs.

Table 16 Conjugate acids and bases.

$Acid_1$ + $Base_2$	$Acid_2$ + $Base_1$
$HCl + H_2O$	$H_3O^+ + Cl^-$
$HNO_3 + H_2O$	$H_3O^+ + NO_3$
$HCN + H_2O$	$H_3O^+ + CN^-$
$HS^- + H_2O$	$H_3O^+ + S^-2$
$H_2O + H_2O$	$H_3O^+ + OH^-$

Organic Chemistry:

Introduction

A wide range of compounds existing either naturally or synthetically prepared are known to possess the element carbon. This forms a branch of chemistry, known as Organic Chemistry. Over three million different organic compounds have been characterized & every year tens of thousands of new substances are added to the list.



Antoine Lavoisier

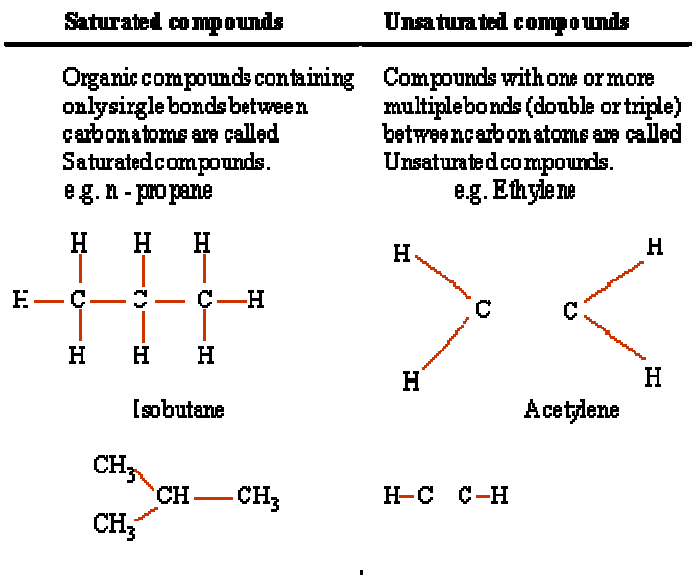
In 1774, Antoine Lavoisier showed that compounds obtained from vegetable and animal sources always contained at least carbon and hydrogen.

Carbon is not the only element found in organic compounds. Hydrogen atoms are almost always present. In addition organic compounds often contain atoms of oxygen, nitrogen, phosphorus, sulfur or the halogens. The bonds formed by carbon with other carbon atoms or other atoms can be single, double or triple bonds.

Standard conventions used to represent these bonds are as follows:

No. of lines	**Type of bond**	**Example**

Single	Single	C – C, C – H
Double	double	C = C, C = O
Triple	triple	C ≡ C, C ≡ N



Organic compounds differ considerably from inorganic compounds. The following comparison illustrates the difference between organic and inorganic compounds. Remember that there are exceptions to every point of comparison

	Organic compounds	Inorganic compounds
(i)	Relatively few elements (mainly C, H, O, N, S, P, F, Cl, Br, I) are involved.	All elements are involved.
(ii)	Bonds are covalent.	Bonds are ionic or electrovalent.
(iii)	Sparingly soluble or insoluble in water but soluble in non-aqueous solvents (organic solvents).	Soluble in water but insoluble in organic solvents.
(iv)	Volatile in nature.	Non-volatile in nature.

(v)	Non electrolytes	Electrolytes
(vi)	Rate of reaction is slow and a catalyst is needed.	Rate of reaction is fast and a catalyst is not needed.
(vii)	Mostly inflammable	Not inflammable
(viii)	Complex structure	Simple structure
(ix)	Isomerism is very common.	Isomerism is not very common.

Functional Groups

Functional groups : is a structural unit consisting of an atom or group of atoms that determines the properties of organic compounds.

The following Table 20 gives the list of common functional groups in organic compounds.

Table 20 Functional groups in organic compounds

Class of compounds	Functional group	General Structure	Example
Alkenes (olefines)	Double bond		Ethylene Butylene
Alkynes	Triple bond	$-C \equiv C-$	Acetylene
Alkyl halides	Halogen	R - X Ar- X	Propyl bromide Bromobenzene
Alcohols	Hydroxy	R-OH	Methanol Ethanol
Phenols	"	Ar-OH	Phenol Resorcinol
Aldehydes	aldehyde or formyl	R-CHO	Acetaldehyde
		Ar - CHO	Benzaldehyde
Ketones	Keto or carbonyl	O R - C - R' O	Acetone
		Ar - C - R	Acetophenone
Carboxylic acids	Carboxyl	R - COOH Ar - COOH	Acetic acid Benzoic acid
Esters	Ester	R - COOR' Ar-COOR	Ethyl acetate Methyl benzoate
Acid Chlorides	Acid chloride or Acyl chloride	R - COCl	Acetyl chloride
		Ar - COCl	Benzoyl chloride
		Ar - C≡N	Benzonitrile
Amide	Amide	R - $CONH_2$ Ar - $CONH_2$	Acetamide Benzamide
Ethers	Ether	R - O - R' Ar - O - R	Diethyl ether Anisole
Amines	Amino	R - NH_2 Ar - NH_2	Methylamine Aniline
Nitriles	Cyano	R - C N Ar - C N	Acetonitrile Benzonitrile
Isocyanides	isocyano	R - N CH Ar - N CH	Ethyl isocyanide Phenyl isocyanide
Nitro	nitro	R - NO_2 Ar- NO_2	Nitromethane Nitrobenzene
Thiols	thiol or mercapto	R - SH Ar - SH	Methyl mercaptan Thiophenol
Sulphides	thioether	R - S - R'	Dimethyl sulphide

Formula of Organic Compounds

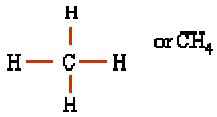
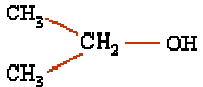
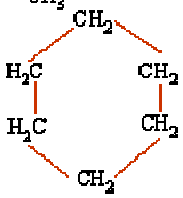
An organic compound can be designated by molecular or structural formula. It gives a concise and instantly recognizable description of the compound.

Molecular formula : It is the formula which expresses the exact number of atoms of each element present in the molecule and does not distinguish between isomers.

Structural formula : It describes the actual attachment of all atoms with one another in the molecule and distinguishes between isomers.

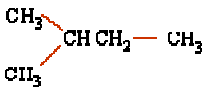
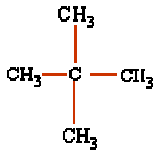
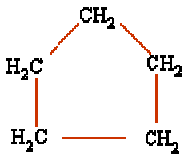
The following table illustrates how these formulae are used to represent the compound.

Table 21

Compound name	Molecular formula	Condensed / Structural formula
Methane	CH_4	 or CH_4
n - propyl alcohol	C_3H_8OH	
isopropyl alcohol	C_3H_8OH	
Cyclohexane	C_6H_{12}	
Benzene	C_6H_6	 or

Note that entries 2 and 3 have similar molecular formulae but different structural formula. Such compounds are called isomers. Structure affects both the chemical reactivity and physical properties of compounds as illustrated in following

Table 22 Structure & Physical properties of isomers

		mp^{0}C	bp^{0}C	Density g / ml 20
Pentane	CH_3-$CH_2CH_2CH_2$-CH_3	-130	36	0.626
Isopentane		-160	28	0.620
Neopentane		-17	10	0.614
Cyclopentane		-94	49	0.746

Inorganic Chemistry:

Introduction

Inorganic Chemistry, study of the structure, properties, and reactions of the chemical elements and their compounds. Inorganic chemistry does not include the investigation of hydrocarbons—compounds composed of carbon and hydrogen that are the parent material of all other organic compounds. The study of organic compounds is called organic chemistry.

Inorganic chemists have made significant advances in understanding the minute particles that compose our world. These particles, called atoms, make up the elements, which are the building blocks of all the compounds and substances in the world around us. Just as the entire English language is constructed from combinations of the 26 letters in the alphabet, all chemical substances are made from combinations of the 112 chemical elements found on the periodic table.

Ninety elements are known to occur in nature, and 22 more have been made artificially. Elements—which include substances such as oxygen, nitrogen, and sulfur—cannot be broken into more elementary substances by ordinary chemical means. The elements are arranged in the periodic table in rows from the lightest element (hydrogen) to the heaviest (ununbium). These rows are split so that elements with similar chemical properties fall in the same columns (for more information, see the *Periodic Law* section of this article).

The smallest representative unit of an element is an atom. (For example, the smallest representative of the element helium (He) is a helium atom.) When atoms that come in close contact have a sufficiently large attractive force, a chemical bond, or binding link, forms between them. The combination of two or more atoms bonded together is called a molecule. A molecule is the smallest particle of a substance possessing the specific chemical properties of that substance. For example, an atom of oxygen (O) combines with two atoms of hydrogen (H) to form a water molecule (H_2O). While molecules of H_2O possess the properties of water, individual oxygen and hydrogen atoms do not.

Important Inorganic Compounds

Advances in inorganic chemistry have made significant contributions to modern living. For instance, synthetic fertilizers manufactured from inorganic chemicals have increased worldwide crop production. Inorganic substances used to fabricate silicon chips help power the global information age. Engineers use metal alloys in automobiles and aircraft to make them lighter and stronger.

Companies also use inorganic compounds to fabricate concrete, steel, and glass—materials used to construct buildings, infrastructure, and other public works around the world.

In the United States, 10 of the 11 most commonly produced chemicals are derived from inorganic elements. These 10 inorganic chemicals (presented below in descending order of production) are used in a wide variety of applications. Sulfuric acid (H_2SO_4) is used to make fertilizers, synthetic fibers, and metals. Nitrogen (N_2) is used in recovering underground petroleum deposits, in the production of ammonia (NH_3), and as a blanketing material for shipping perishables such as fruits and vegetables. Oxygen (O_2) is used in the production of steel and plastics, in medical applications, and in rocketry. Lime (CaO) is used in the manufacture of steel and cement. Ammonia (NH3) is combined with sulfuric acid to make ammonium sulfate (NH_4SO_4), the most important of the synthetic fertilizers.

The remaining five most commonly produced inorganic chemicals (which frequently interchange rankings in production volume) are also used in a wide variety of applications. Sodium hydroxide (NaOH), commonly called lye, is used in the manufacture of paper, soap, detergents, and synthetic fibers, and is also a caustic material used as a drain cleaner. Chlorine (Cl_2) is used to manufacture vinyl chloride plastic, to disinfect drinking water, and to

bleach paper during manufacturing. Phosphoric acid (H_3PO_4) is used to give soft drinks a tart flavor and to make fertilizers. Sodium carbonate (Na_2CO_3), more commonly known as soda ash, is used in the production of glass, paper, and textiles. Nitric acid (HNO_3) is used to make synthetic fibers, such as nylon; explosives, such as nitroglycerin and TNT (trinitrotoluene); and is also combined with ammonia to make fertilizer.

Periodic Law:

Modern inorganic chemistry can be traced to the work of Russian chemist Dmitry Ivanovich Mendeleyev and German physicist Julius Lothar Meyer, who independently developed the periodic law of the chemical elements at about the same time in the late 19th century. Mendeleyev is generally credited with the findings, because he established the periodic law in 1869, and Meyer established this chemical law a year later.

Both scientists, however, discovered that arranging the elements in order of increasing atomic mass produced a table of chemical properties and reactivity patterns that were regularly repeated. This phenomenon—known as the periodic law—is most often represented in the periodic table of the elements.

By arranging the elements into rows of increasing atomic mass, Mendeleyev observed that elements with similar properties fell into the same vertical columns, called groups. For example, members of the alkali metals—lithium (Li), sodium (Na), potassium (K), rubidium (Rb), and cesium (Cs)—are all are extremely reactive, bursting into flames when they are brought in contact with water.

Appendix A Conversion Factors:

Appendix A

Conversion Factors		
Multiply	**By**	**To obtain**

Liters	0.22009	Gallons (U.K.)
Liters	1.76072	Pints
Meters	3.2808	Feet
Meters	1.093614	Yards
Metric Tonnes	0.98421	Tons (U.K.)
Miles	1.609343	Kilometers
Millimeters	0.03937	Inches
Ounces	28.34953	Grams
Pints	0.56795	Liters
Pounds	0.4536	Kilograms
Square Centimeters	0.15500	Square Inches
Square Feet	0.09290	Square Meters
Square Inches	6.45163	Sq. Centimeters
Square Kilometers	0.386102	Square Miles
Square Meters	0.0002471	Acres
Square Meters	10.76387	Square Feet
Square Meters	1.19599	Square Yards
Square Miles	640.00	Acres
Square Miles	2.58999	Square Kilometers
Square Yards	0.83613	Square Meters
Tons (U.K.)	1.01605	Metric Tonnes
Tons (U.K.)	1016.0470	Kilograms
Yards	0.914399	Meters

Acres	0.404686	Hectares
Acres	4047.00	Square Meters
Acres	0.001562	Square Miles
Centimeters	0.393701	Inches
Cubic Centimeters	0.061024	Cubic Inches
Cubic Feet	6.228	Gallons
Cubic Feet	0.02832	Cubic Meters
Cubic Inches	16.3871	Cubic Centimeters
Cubic Meters	35.31445	Cubic Feet
Cubic Meters	1.30794	Cubic Yards
Cubic Yards	0.76456	Cubic Meters
Feet	0.3048	Meters
Gallons	0.1337	Cubic Feet
Gallons (U.K.)	4.5461	Litres
Grams	0.03527	Ounces
Hectares	2.47105	Acres
Horse power	0.7457	Kilowatts
Inches	2.539998	Centimeters
Inches	25.39998	Millimeters
Kilograms	2.204620	Pounds

Appendix B Values of some Constants: Appendix B

Values of some constants:

1. Speed of light (c) = 2.9979×10^{8} ms^{-1} (metres per second)

2. Avogadro's number (N) = 6.022×10^{23} mol^{-2} (per mole)

3. Mass of an electron (me) = 9.1095×10^{-31} Kg

4. Mass of a proton (mp) = 1.6726×10^{-27} Kg

5. Mass of a neutron (mn) = 1.6748×10^{-27} Kg

6. Gas constant (R) = 8.3144 JK^{-1} mol^{-1} (joules per kelvin per mol)

Appendix C Basic and Derived S. I. Units

Appendix C

Basic units

Physical Quantity	Unit	Symbol
Length	meter	m
Mass	Kilogram	kg
time	Second	s
Temperature	Kelvin	K
Amount of substance	mole	mol

Derived units

Physical Quantity	Unit	Symbol
area	square meter	m^2
volume	cubic meter	m^3
density	Kilogram per cubic meter	$Kg.m^{-3}$
molar mass	Kilogram per mole	$Kg.mol^{-1}$
molar volume	cubic meter per mole	$m^3 mol^{-}$
molar concentration	mole per cubic meter	$mol\ m^{-3}$
Pressure	Pascal (Kilo Pascal)	Pa (KPa)
Energy	Joule	J
Electrical potential difference	Volt	V
Heat Capacity	Joule per Kelvin	JK^{-1}