Some Basic Concepts of Chemistry

Changes Among Various States of Matter

Change of State-An Overview

In daily life, we see different kinds of changes in the states of matter. The formation of ice cubes from water in the refrigerator is an example of a change in the state of matter from liquid to solid. When water is boiled, vapours are formed. This is an example of change in the state of matter from liquid to gas.

The following terminologies are used to describe the changes in the states of matter.

- Change from the solid state to the liquid state is called **melting**.
- Change from the liquid state to the solid state is called **freezing**.
- Change from the liquid state to the gaseous state is called **vapourisation**.
- Change from the gaseous state to the liquid state is called **condensation**.

There are two other changes between the three states of matter—sublimation and deposition.

Sublimation: It is the process in which a substance changes directly from the solid state to the gaseous state without entering into the liquid state. The changing of snow into water vapour is an example of sublimation. Some common examples of substances that sublime are dry ice, camphor, and naphthalene.

Deposition: It is the process opposite to sublimation. In this, a substance changes directly from the gaseous state to the solid state. Frost is an example of deposition.

Did You Know?

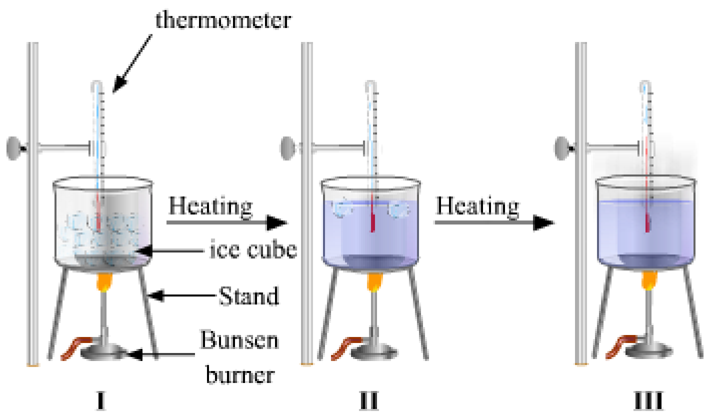
When we open the refrigerator, we see freezing fog. This is nothing but condensed water.

Air contains vapours. When we open the refrigerator, the temperature comes down. This condenses the vapours into tiny drops of water and produces freezing fog.

Temperature Affecting the Change of State

Let us perform an activity to understand the effect of temperature on the different states of matter.

Procedure: Take about 150 g of ice in a beaker and use a laboratory thermometer to note the temperature of ice. Start heating the beaker on a low flame and record the temperature when the ice starts melting. Observe the temperature when all the ice gets converted into water. Stir the water with a glass rod till it starts boiling.



Result: In the beginning, the temperature of ice is below 0°C. When ice begins melting, the temperature is recorded to be 0°C. Temperature remains constant at 0°C untill all the ice melts. The continued heating of water causes its temperature to rise.

Conclusion: It can be concluded from this activity that an increase in temperature changes a substance from its solid state to its liquid state, and further heating (i.e., further increase in temperature) changes the liquid so formed into vapour.

Temperature Affecting the Change of State

You know that matter, irrespective of its state, consists of particles. What happens to these particles of matter while it is undergoing a change in its state? For us to understand this, we need to first know that:

- The particles of matter possess kinetic energy.
- A force of attraction exists between any two particles.

Kinetic energy of the particles of matter: A moving particle/object possesses a certain amount of energy because of its motion. This energy is called kinetic energy. The particles of matter are in constant motion. Therefore, they possess kinetic energy.

Particle-particle force of attraction: Every particle of matter attracts the particles near it. An increase in the distance between particles decreases the force of attraction between them. Conversely, a decrease in distance increases this force of attraction.

The given figure shows the kinetic energy of particles and the particle-particle force of attraction in the three states of matter.

Kinetic energy of particles: Gas > Liquid > Solid

Particle–particle force of attraction: Solid > Liquid > Gas

When a solid substance is heated, there is an increase in the kinetic energy of its constituent particles. As a result, the particles start vibrating with greater speed. This extra energy helps the particles to overcome the particle-particle force of attraction. Soon, they leave their positions and start moving more freely. Consequently, the substance melts into its liquid state. This is known as **melting point**. The melting point of ice is 0°C.

Liquids have a characteristic temperature at which they turn into solids. This is called **freezing point**. The freezing point of water is 0°C.

Further heating increases the kinetic energy of the liquid particles. This increases the velocity of the particles. At a certain temperature, they obtain enough energy to break free from the particle-particle force of attraction. At this point, the liquid changes into its gaseous state. This is known as **boiling point**. The boiling point of water is 100°C.

During the conversion of ice into water, the temperature remains constant until all the ice melts into water. The supplied heat is used up for changing water from its solid state to its liquid state. The heat energy is absorbed by the ice without showing any rise in temperature. This heat energy is called **latent heat**.

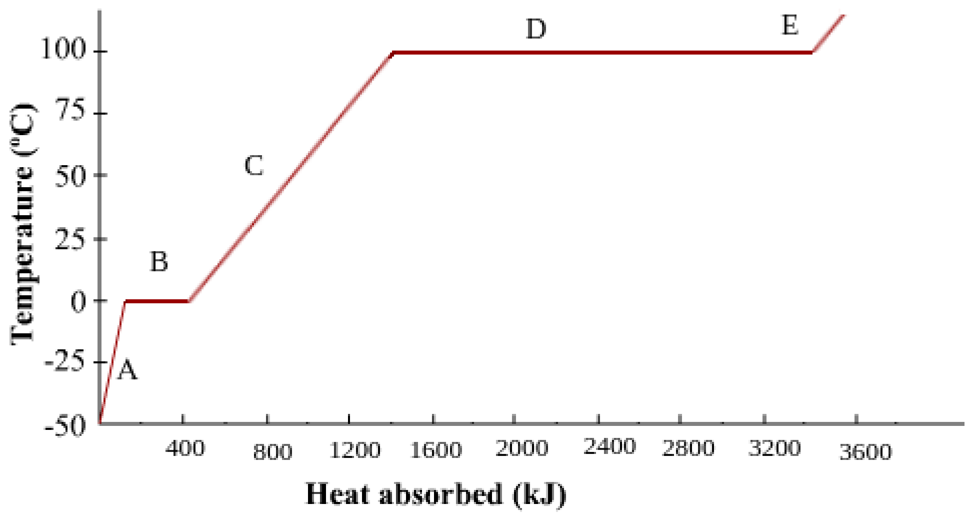
The amount of heat required to convert 1 kg of a solid into its liquid state without a change in temperature (i.e., at its melting point) is called **latent heat of fusion**. For ice, the latent heat of fusion is 334 kJ kg^{-1}. This implies 334 kJ of heat has to be provided to convert 1 kg of ice at 0°C into 1 kg of water at 0°C. Conversely, 334 kJ of heat is released when 1 kg of water freezes at 0°C to give 1 kg of ice at 0°C.

Know More

Latent heat of vapourization is the amount of heat required to convert 1 kg of a liquid into its vapour state without a change in temperature. For water, the latent heat of vapourization is 2260 kJ kg^{-1}. This means that 2260 kJ of heat must be provided to convert

1 kg of water at 100°C into 1 kg of vapour at 100°C. Conversely, 2260 kJ of heat is released when 1 kg of water vapour condenses at 100°C to give 1 kg of water at 100°C.

Heating curve



If the increase in temperature during heating and the absorbed heat are plotted on a graph, then the curvature which is formed is called the **heating curve**.

In the figure, 'A' represents the rise in the temperature of the substance in its solid state from
−50ºC to 0°C; 'B' shows the latent heat of fusion; 'C' shows the increase in the temperature of the substance in its liquid state from 0°C to 100°C; 'D' shows the latent heat of vapourisation, and 'E' shows the increase in the temperature of the substance in its gaseous state.

Solved Examples

Easy

Example 1: **If the melting point of a solid is high, then the _________ between the particles is stronger.**

Solution:

force of attraction

Medium

Example 2: **Which has more energy: solid wax at 42°C or liquid wax at 42°C?**

Solution:

Liquid wax at 42°C has more energy than solid wax at the same temperature.

Hard

Example 3: **Choose the process which will absorb heat/energy from the surroundings.**

A.Conversion of ice into water

B.Conversion of water vapour into snow

C.Precipitation of water vapour as rain

Solution:

The correct answer is A.

Measuring Temperature

Three scales are commonly used for measuring temperature, namely, the **Celsius scale**, the **Fahrenheit scale**, and the **Kelvin scale.**

The relation between the Celsius and the Kelvin scale can be expressed as C + 273 = K

The relation between the Celsius and the Fahrenheit scale can be expressed as follows:

$$\frac{C}{5} = \frac{F - 32}{9}$$

Example: 30°C can be expressed as 303 K and 86 ^{o}F.

Celsius to Kelvin: 30 + 273 = 303 K

Celsius to Fahrenheit:

$$\frac{30}{5}=\frac{F-32}{9}$$
$$\Rightarrow 6=\frac{F-32}{9}$$
$$\Rightarrow 54=F-32$$
$$\Rightarrow F=86$$

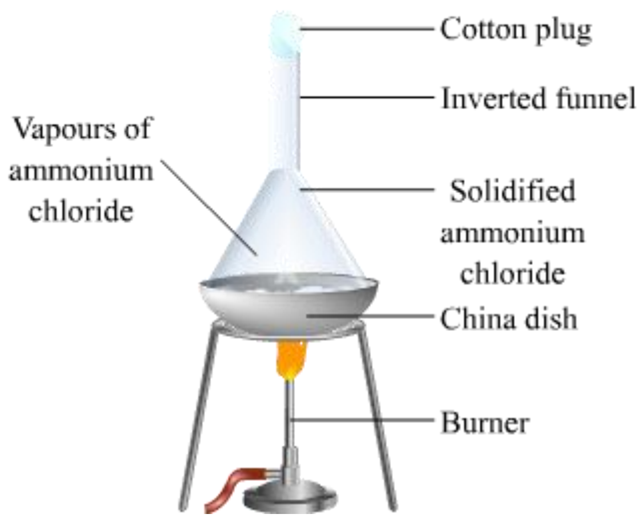
Did You Know?

Cool Facts

- The temperature zero Kelvin is known as absolute zero. Nothing can be colder than zero Kelvin.
- Dry ice is frozen carbon dioxide. Its temperature is −78.5°C. It turns directly into carbon dioxide gas without undergoing a liquid phase. Its sublimation characteristic and super-cold temperature make dry ice suitable for refrigeration. It is commonly used to export frozen materials across long distances.

Whiz Kid

Take some ammonium chloride salt in a china dish. Crush the salt and cover the dish with a funnel, as shown in the figure. Plug the stem of the funnel using some cotton. After this, start heating the dish slowly using a burner.



Result of the activity:

Upon heating, ammonium chloride will vapourise without transforming into its liquid form (**sublimation**). Later, the vapours will get cooled on the walls of the funnel and will directly convert into solid ammonium chloride (**deposition**).

Note: The same activity can be done using camphor or naphthalene.

We know that change in temperature affects the state of matter. Change in pressure, too, affects the state of matter. Let us see how.

We have a gas in a closed container. Say, we put some weight on the lid of the container. This increases the pressure on the container, which in turn causes the gas particles to come close to one another. As a result, the kinetic energy of the particles reduces. Nevertheless, the particles are still quite far away from one another and, hence, are still in the gaseous state. When the pressure on the container is increased further, the gas particles come very close to one another. Gradually, the gas **liquefies**.

Did You Know?

Water boils below 100°C (at approx. 92°C) in Mussoorie.

Mussoorie is a hill station set at a height of about 2000 m above sea level. Atmospheric pressure decreases as you go up from the sea level. Decrease in pressure lowers the boiling point of water below 100°C.

Whiz Kid

Liquid crystals are believed to be an independent state of matter as their properties lie in between those of liquids and solid crystals. They exist in a specific temperature range. They behave as solids below that temperature range and as liquids above that temperature range.

Know More

Why we need to liquefy gases

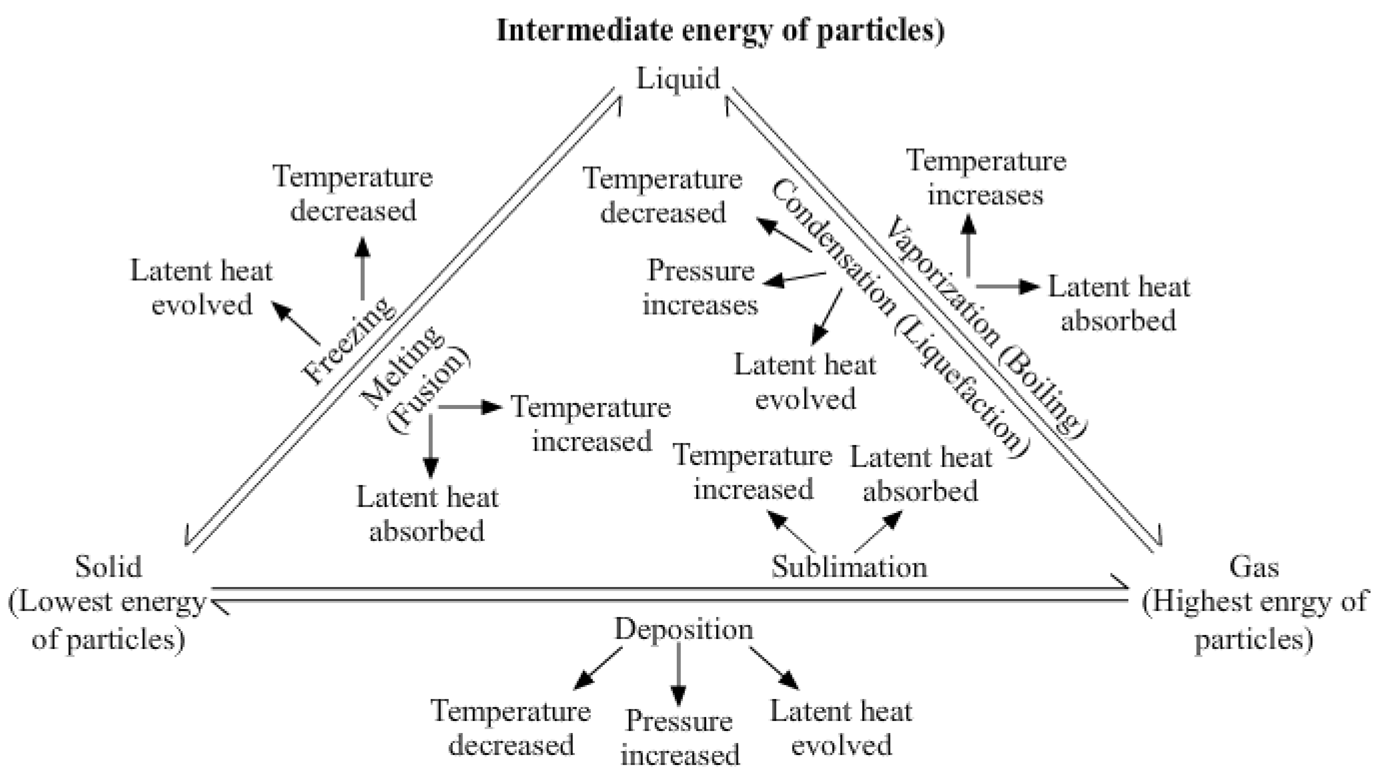
Together with low temperature, high pressure is generally used to liquefy gases.

A highly combustible gas is released during the fractional distillation of crude oil. This gas is known as petroleum gas. Petroleum gas is also trapped over the reserves of oil present beneath Earth's crust. Petroleum gas is liquefied by applying high pressure and low temperature. This is known as liquefied petroleum gas or LPG. LPG is used as a domestic fuel.

Other uses of liquefaction of gases

- Liquefaction of gases is helpful for their easy storage and transportation.
- Liquefied gases can be used in various fields; for example, in air conditioning and refrigeration systems (gases used are liquid ammonia and liquid sulphur dioxide).
- Liquid oxygen is supplied to hospitals for patients. It is also used as a rocket propellant.
- Liquid nitrogen is used in **cryosurgery**.
- Liquid chlorine is supplied to water treatment plants for purification of water.
- Liquid hydrogen in combination with liquid oxygen forms the fuel for rocket propulsion.

Inter-Conversion among Solids, Liquids, and Gases



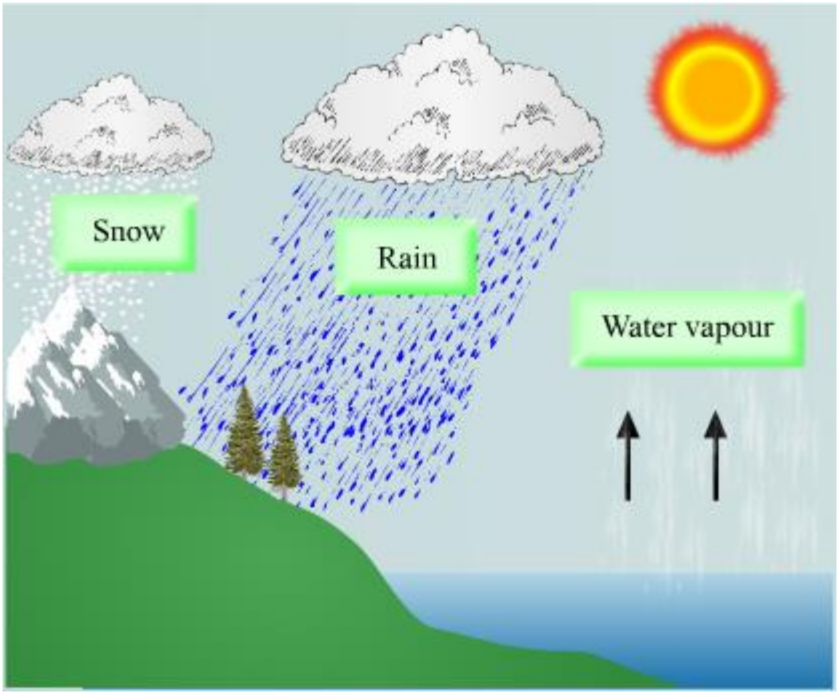
The Solid State

States of Matter-An Overview

We know that everything is made up of matter, yet things exist in different forms. What makes things look different from one another?

Matter is a broad umbrella covering different sub-categories which we know as the **states of matter**.

The view of the hills during winters is ideal for observing the three main states of matter—solid, liquid and gas. Here, you can see heavy clouds which are nothing but collections of vapourised water particles. You can also see liquid water falling from these same clouds as rain. And of course, there is the dusting of snow which is in fact solidified water.

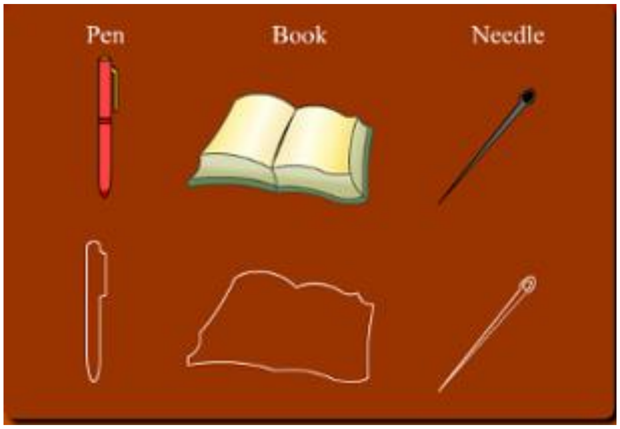


The different states of matter are:

- Solid
- Liquid
- Gaseous
- Plasma
- Bose-Einstein condensate

Matter is said to be solid if it has a fixed **shape** and a fixed **volume**. For example, a pen. It has a fixed shape and a fixed volume; hence, it is solid. Matter that does not have a fixed shape is not solid, as is the case with water. The particles of solids have a minimum or no kinetic energy and therefore the particles do not have any movement. The intermolecular spaces between the particles of solid are very small due to stronger attraction among the particles. Solids, therefore, cannot be compressed.

Activity Time



Procedure: Collect a pen, a book, and a needle. Trace the shapes of these materials in a notebook and compare the tracings. Also, try compressing each material.

Result: When you compare the tracings, you will observe that each material has a distinct shape and boundary. When you try compressing the materials, you will observe that each material has negligible [[mn: glossary]]compressibility[[/mn: glossary]].

Conclusions: The following conclusions can be made about a solid.

- It has a fixed shape, fixed volume, and a fixed boundary.
- There are very little **intermolecular** spaces in a solid. Hence, it has a tendency to maintain its shape. This means that it has negligible compressibility.
- It is **rigid**. It may break under force, but it is difficult to change its shape.
- It rarely **diffuses** in another solid. Example- Diffusion of chalk powder on a blackboard. This is the reason why it is difficult to clean (rub) a used blackboard that has not been cleaned for several days.

Whiz Kid

Solids have the following forms.

- **Crystalline**: Calcite (rhombic), fluorite (octahedral) and quartz (hexagonal) are crystalline solids.
- **Polycrystalline**: Metals are polycrystalline solids.
- **Amorphous**: Glass is an amorphous solid.
- **Polymeric**: Natural rubber is a polymeric solid.

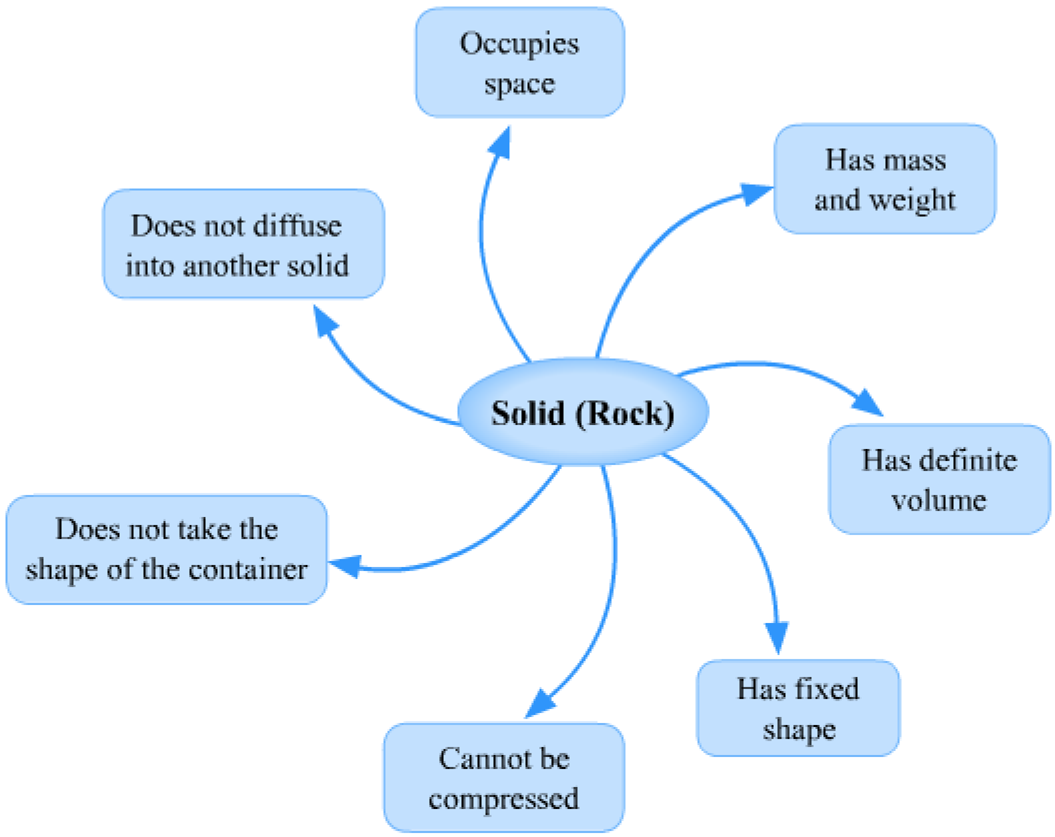
It is possible to stretch certain substances without breaking them. These substances are made up of long chains of atoms bonded together (usually carbon atoms bonded by covalent bonds). These substances are called polymeric substances. This is why the shape of rubber changes when stretched even though it is solid.

Did You Know?

Although sponge is solid, it can be bent and squeezed. Sponge has minute holes on its surface. Air is trapped in these holes. This air is expelled as we press or squeeze sponge. This makes it possible to bend and squeeze sponge.

The Solid State

The following diagram illustrates the properties of a solid.



The Liquid State

Unlike a solid, a liquid has no fixed shape. However, it does have a fixed volume. It takes the shape of the container in which it is kept. For example, water does not have a fixed shape, but its volume is fixed. When a certain volume of water is poured into a container, it takes the shape of the container, but its volume remains the same. On the other hand, a pen (which is a solid) has a fixed shape and volume. A liquid is not rigid, i.e., it flows freely. The intermolecular spaces in a liquid are greater than in case of a solid. Hence, a liquid has more compressibility than a solid. The particles of liquids have more kinetic energy than solid particles and therefore has greater speed than solid particles.

Characteristics of a liquid on the basis of the particle nature of matter

- A liquid does not have a fixed shape. It takes the shape of the container in which it is kept.
- A liquid has a fixed volume.
- It is not rigid, i.e., it flows freely.
- It has more compressibility than a solid. So, it can easily diffuse in other liquids.
- In most cases, the density of a substance in the liquid state is lesser than its density in the solid state.

Usually liquids have lower density than solids, yet ice floats in water. Can you say why?

Ice is lighter than water since a particular mass of ice occupies more space than the same mass of water. In ice, water molecules are closely packed because of the tight bonding between them. This makes ice lighter than water.

Know More

Solids, liquids and gases can diffuse in liquids. The dissolution of salt or sugar in water and the dissolution of ink in water are examples of the same. Gases such as oxygen and carbon dioxide diffuse and dissolve in water bodies. It is because of these gases that aquatic plants and animals are able to survive underwater.

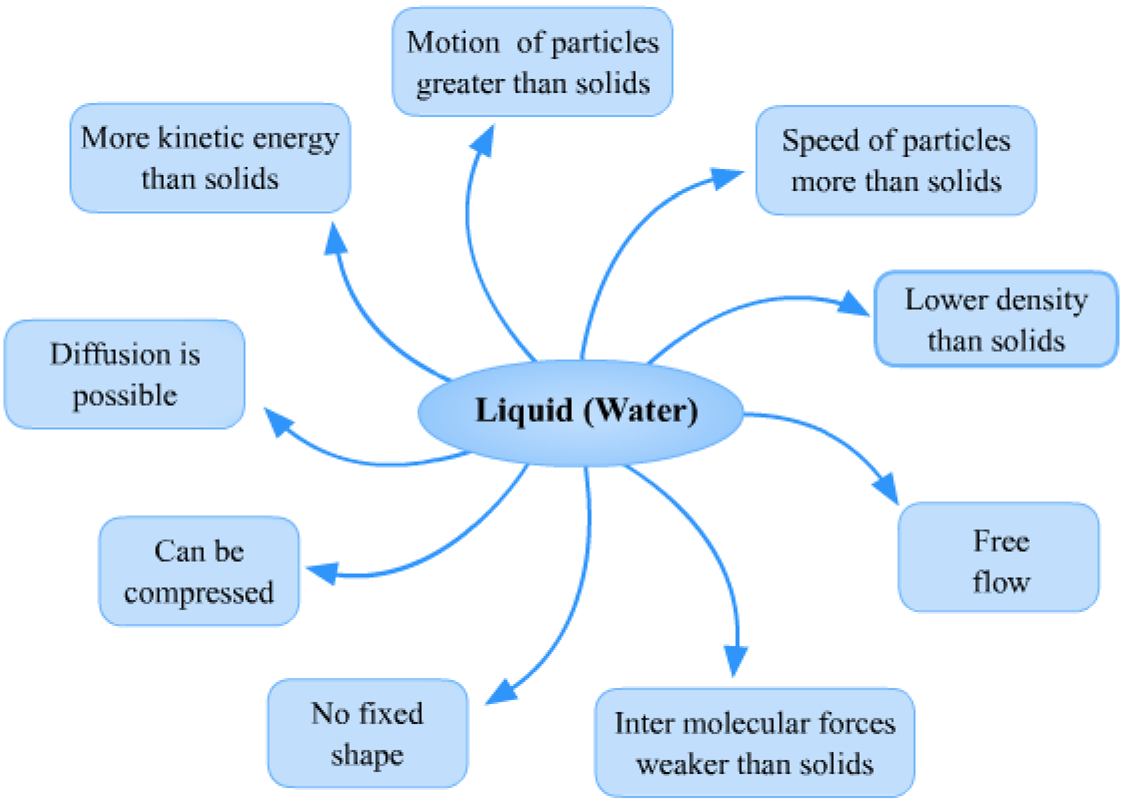
This high rate of diffusion in liquids is because of the fact that a liquid has larger intermolecular spaces.

Did You Know?

Bronze, an **alloy**, expands when its state changes from liquid to solid. This property of bronze is utilized in moulding statues.

The Liquid State

The following diagram illustrates the properties of a liquid.

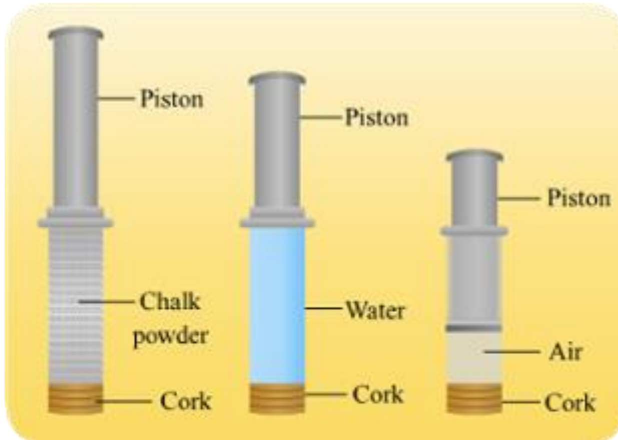


The Gaseous State

A gas neither has a fixed shape nor a fixed volume. Hence, it does not have a fixed boundary. It can flow in all directions and can be easily compressed. In a given space, the number of particles in a gas is lesser than in the case of a solid or a liquid. The constituent particles of a gas show a random motion because of the presence of large spaces between them. Consequently, the kinetic energy of the particles in a gas is more than in the case of a solid or a liquid. Due to the large distances between the particles, the forces of attraction between them are very low or negligible.

Activity Time

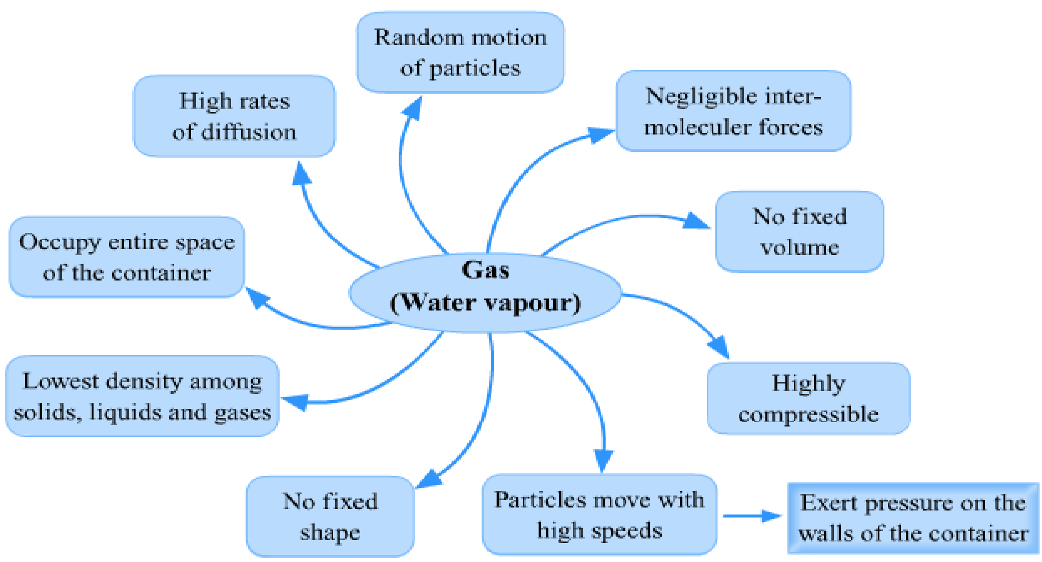
Procedure: Take three 100 mL syringes and remove their pistons. Close the nozzles of the syringes with rubber corks. Fill one syringe with chalk powder and another with water. Now, reinsert the pistons and push them.



Result: The force required to push the pistons of syringes containing chalk powder and water will be greater than that required to push the piston of the syringe containing air.

The Gaseous State

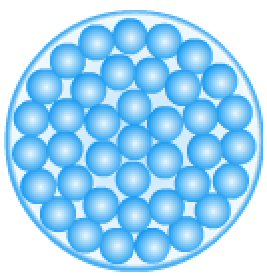
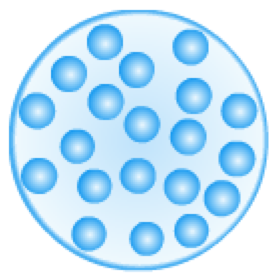
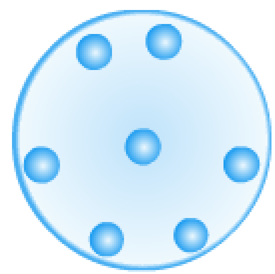
The following diagram illustrates the properties of a gas.



Differentiating between the Three States of Matter

Differentiating the Three States of Matter

Solid	Liquid	Gas

Definite shape Occupies space Definite volume Cannot be compressed Rigid Does not diffuse in other solids	No definite shape Occupies space Definite volume Slightly compressible Not rigid Can diffuse in other liquids	No definite shape Occupies space No definite volume Highly compressible Not rigid Can diffuse in other gases
 **Solid**	 **Liquid**	 **Gas**

Solved Examples

Easy

Example 1: **Answer the questions with a 'Yes' or a 'No' for each of the three states of matter.**

Questions	**Solid**	**Liquid**	**Gas**
Does it occupy space?			
Does it have a definite volume?			
Can it be compressed?			
Does it take the shape of the container enclosing it?			
Can it diffuse in a like state of matter?			

Solution:

Questions	**Solid**	**Liquid**	**Gas**
Does it occupy space?	Yes	Yes	Yes
Does it have a definite volume?	Yes	Yes	No
Can it be compressed?	No	Yes	Yes
Does it take the shape of the container enclosing it?	No	Yes	Yes
Can it diffuse in a like state of matter?	No	Yes	Yes

Solved Examples

Easy

Example 2: **Identify the state I'm in.**

Object	**State**
Glass	Bose-Einstein condensate
Welding arc	Solid
Liquid helium	Gas
Mercury	Plasma
Fog	Liquid

Solution:

i → b; ii → d; iii → a; iv → e; v → c

Medium

Example 3:

Guess who I am.

i) The container I'm placed in does not matter. My shape does not change. I'm _______.

ii) I'm flexible and particles can move with some speed. I'm _______.

iii) I possess highest kinetic energy and my particles move with high speed. I'm _______.

iv) I'm charged and have high temperature. I'm _______.

Solution:

(i) solid

(ii) liquid

(iii) gas

(iv) plasma

The Plasma and The Bose-Einstein Condensate

This state of matter comprises super-energetic and super-excited particles. It was discovered in Crookes tube by Sir William Crookes in 1879.

Here are a few examples of plasma.

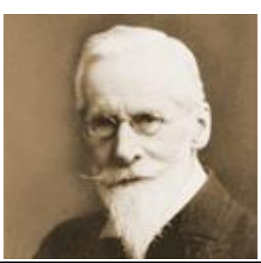
The sun, Fluorescent lights, Neon signs, Nebulae

Characteristics of plasma

- It is the most common state of matter found in the universe.
- Like a gas, it does not have a definite shape and volume.
- It is found in the sun, the stars, the interstellar space and the intergalactic space.
- High temperature is needed to maintain the ionization of the particles of plasma.
- Plasma is influenced by electric and magnetic fields.

Know Your Scientist

Sir William Crookes (1832–1919)



Sir William Crookes was a chemist and physicist of British origin. He is credited with the invention of Crookes tube. He discovered thallium and also identified the first known sample of helium. He developed Crookes tube to investigate cathode rays. In the course of his investigation, he discovered plasma—the fourth state of matter.

The Bose−Einstein condensate was discovered by Albert Einstein and S.N. Bose in 1924−25.

Characteristics of the Bose−Einstein condensate

- It consists of super-unenergetic and super-cold particles.
- It is formed when a gas of extremely low density is cooled to an extremely low pressure.
- It can be created with few elements. Rubidium is one such element.

The following properties are exhibited by Bose−Einstein condensate

- **Superfluidity:** It is the property by virtue of which matter shows frictionless flow at temperature near 0 K.
- **Superconductivity:** It is the property by virtue of which matter shows zero electrical resistance when cooled below a specified temperature.

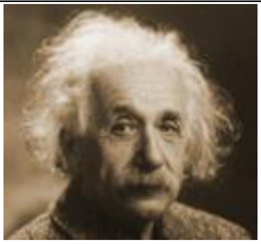
Examples of Bose-Einstein Condensate are:

- Helium-4 cooled below 2.17 K
- A gas of rubidium atoms

Know Your Scientist

Satyendra Nath Bose (1894−1974)

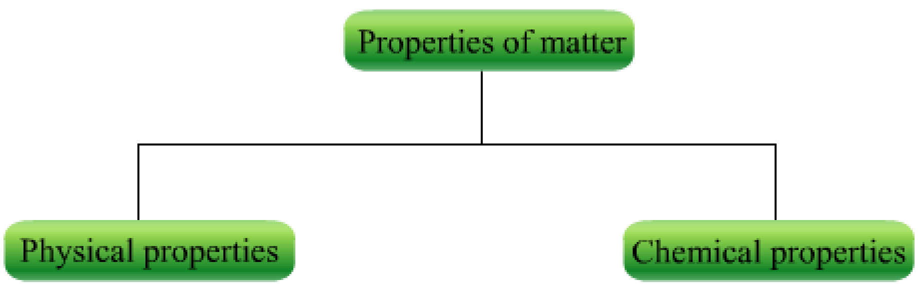

Satyendra Nath Bose was a physicist of Indian origin. He specialized in mathematical physics. He is known for his collaborative theory on the Bose−Einstein Condensate—the fifth state of matter. He was awarded the Padma Vibhushan in 1954.**Albert Einstein (1879−1955)**



Albert Einstein (1879–1955) was a physicist of German origin. He is credited with the development of the theory of relativity that revolutionized studies in physics. He is known as 'the father of modern physics'. His works include derivation of relationship between energy and mass as

$E = mc^2$ (where, E = energy, m = mass of a body and c = velocity of light). He received the Nobel Prize in Physics in the year 1921 for his contributions to physics and for the discovery of the laws of photoelectric effect.

Properties of Matter



- **Physical properties**
- Properties which can be measured or observed without changing the identity or composition of the substance.
- Example – Colour, odour, melting point, boiling point, density, etc.
- **Chemical Properties**
- Properties in which chemical change in the substance takes place.
- Examples – acidity, basicity, combustibility, characteristic reactions of substance with other elements and compounds
- Many properties of matter such as mass, volume, area etc are quantitative.

Measurement of properties

- **Systems of Measurement**
- English system
- Metric system
- International system of units (SI)
- **The International System of Units (SI)**
- Seven base units

Base Physical Quantity	***Symbol for Quantity***	***Name of SI Unit***	***Symbol for SI Unit***
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Temperature	T	kelvin	K
Amount of substance	n	mole	mol
Luminous intensity	lv	candela	cd

- Definitions of SI base units

Unit of length	metre	The metre is the length of the path travelled by light in vacuum during a time interval of 1/299 792 458 of a second.
Unit of mass	kilogram	The kilogram is the unit of mass; it is equal to the mass of the international prototype of the kilogram.
Unit of time	second	The second is the duration of 9 192 631 770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the caesium-133 atom.
Unit of electric current	ampere	The ampere is that constant current, which if maintained in two straight parallel conductors of infinite length of negligible circular cross-section and placed 1 metre apart in vacuum, would produce a force equal to 2 × 10−7 Newton per metre of length between these conductors.
Unit of thermodynamic temperature	kelvin	The kelvin, unit of thermodynamic temperature, is the fraction 1/273.16 of the thermodynamic temperature of the triple point of water.
Unit of amount of substance	mole	1. The mole is the amount of substance of a system, which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12; its symbol is "mol." 2. When the mole is used, the elementary entities must be specified and these may be atoms, molecules, ions, electrons, other particles, or specified groups of such particles.
Unit of luminous intensity	candela	The candela is the luminous intensity (in a given direction) of a source that emits monochromatic radiation of frequency 540 × 1012 hertz and that has a radiant intensity in that direction of 1/683 watt per steradian.

- Prefixes used to indicate the multiples or submultiples of a unit.

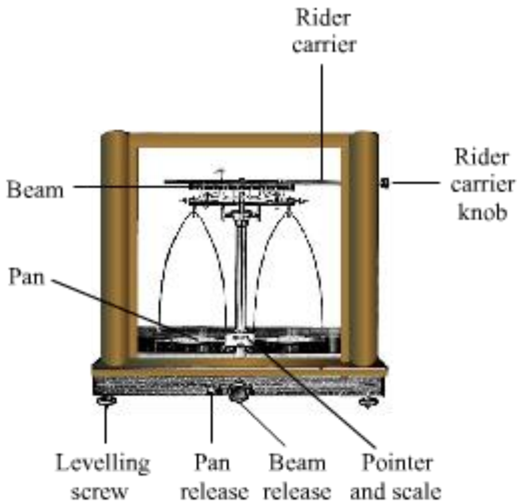
Multiple	**Prefix**	**Symbol**
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	a

10^{-15}	femto	f
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	μ
10^{-3}	milli	m
10^{-2}	centi	c
10^{-1}	deci	d
10	deca	da
10^{2}	hecto	h
10^{3}	kilo	k
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta	P
10^{18}	exa	E
10^{21}	zeta	Z
10^{24}	yotta	Y

- **Mass and Weight**

Mass	**Weight**
Amount of matter present in an object	Force exerted on an object by gravity
Constant, irrespective of the place	Varies from place to place due to change in gravity

- Mass can be determined accurately by using an analytical balance.



- SI unit of mass = Kilogram (kg)

- 1 kg = 1000 g = 10^6 mg

- **Volume**

- Amount of space occupied by an object

- Has the units of $(length)^3$

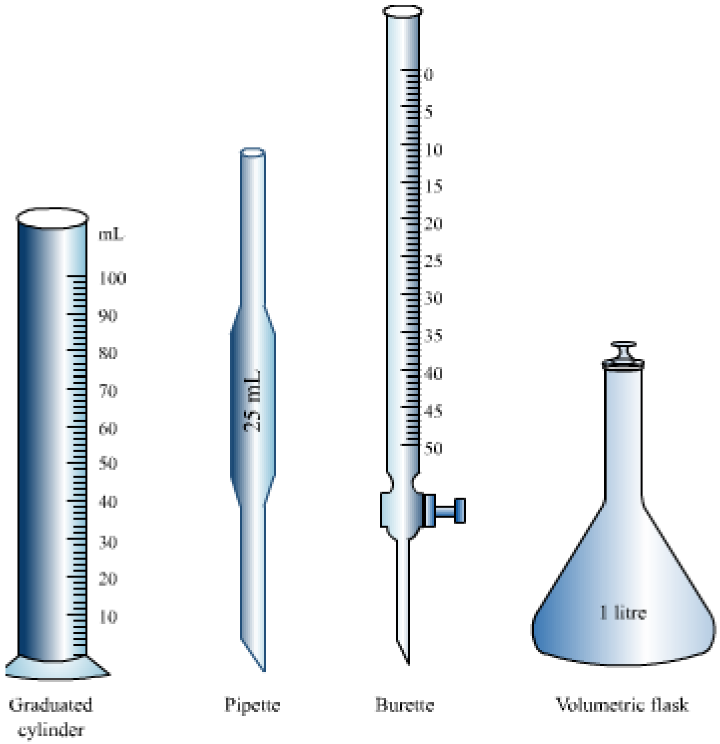
- SI unit = m^3

- Often used units = dm^3, L

  1 dm^3 = 1000 cm^3

  1 L = 1000 mL

- 1 Litre is equal to 1 dm^3.
- 1 Millilitre is equal to 1 cm^3.
- Measuring devices – Burette, pipette, graduated cylinder, volumetric flask



- **Density**
- Amount of mass per unit volume

  i.e. Density $= \dfrac{\text{Mass}}{\text{Volume}}$

- SI unit of density = $\dfrac{\text{SI unit of mass}}{\text{SI unit of volume}}$

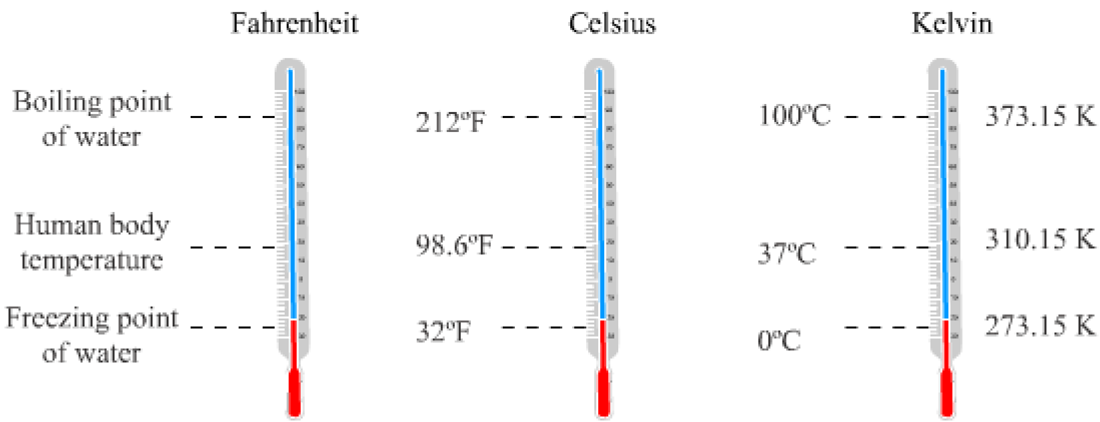
  $= \dfrac{\text{kg}}{\text{m}^3}$ or kg m^{-3}

- Often used unit = g cm^{-3}
- **Temperature**
- Three scales – degree Celsius (°C)

  degree Fahrenheit (°F)

  kelvin (K)

- SI unit = Kelvin (K)
- Thermometers using different temperature scales



- Relation between °F and °C scale

$$°F = \frac{9}{5}(°C) + 32$$

- Relation between K and °C scale

K = °C + 273.15

- Negative values of temperature are possible in °C scale, but not in °F and K scale.

Example

The boiling point of water at sea level is 212 °F. What is its equivalent in Kelvin scale?

Solution:

To convert temperature from Fahrenheit scale into Kelvin scale, the following equations are used.

$$°F = \frac{9}{5}(°C) + 32$$

$\Rightarrow$ °C = (°F − 32)/1.8

K = °C + 273.15

Converting °F into °C,

°C = (212 − 32)/1.8

= 100

Therefore, the boiling point of water is 100°C converting °C into K.

K = °C + 273.15

= 100 + 273.15

= 373.15

Hence, 212°F is equivalent to 373.15 K.

Uncertainty in Measurement

Scientific Notation

- Exponential notation in which a number can be represented in the form $N \times 10^n$, where n can be a positive or a negative number; N varies between 1 and 10.

Example
Express the following in the scientific notation.
(i) 0.00502
(ii) 32010.07
Solution:
(i) 5.02 × 10−3
(ii) 3.201007 × 104

- Multiplication and division

$(1.7 \times 10^{-7}) \times (5.9 \times 10^{2}) = (1.7 \times 5.9)\ (10^{-7+2})$

$= (1.7 \times 5.9) \times 10^{-5}$

$= 10.03 \times 10^{-5}$

$= 1.003 \times 10^{-4}$

- Addition and subtraction
- The numbers are written by keeping the exponent same.

  Example:

  $7.25 \times 10^{-3} + 4.12 \times 10^{-5}$

  $= 7.25 \times 10^{-3} + 0.0412 \times 10^{-3}$

  $= (7.25 + 0.0412) \times 10^{-3}$

  $= 7.2912 \times 10^{-3}$

  Example:

  $2.87 \times 10^{7} - 5.1 \times 10^{6}$

  $= 2.87 \times 10^{7} - 0.51 \times 10^{7}$

  $= (2.87 - 0.51) \times 10^{7}$

  $= 2.36 \times 10^{7}$

Significant Figures

Precision	**Accuracy**
Closeness of various measurements for the same quantity	Agreement of a particular value to the true value of the result
Expressed as the difference between a measured value and the arithmetic mean value for a series of measurements i.e., Precision = Individual value − Arithmetic mean value	Expressed as the difference between the experimental value or the mean value of a set of measurements and the true value i.e., Accuracy = Mean value − True value

Comprise meaningful digits which are known with certainty

- Indicated by writing the certain digits and the last uncertain digit
- Example, in a result 45.8 cm, 45 is certain and 8 is uncertain. The uncertainty would be ± 1 in the last digit.

- If not stated, then the uncertainty in the last digit is understood as ± 1.

- Rules to determine the number of significant figures:

- All non-zero digits are significant. Example: 145 mL has three significant figures.
- Zeroes preceding the first non-zero digit are not significant. Example: 0.04281 has four significant figures.
- Zeroes between two non-zero digits are significant. Example: 23.007 has five significant figures.
- Zeroes at the end or right of a number are significant when they are on the right side of the decimal point. Example: 0.8300 has four significant figures.
- Counting numbers of objects have infinite significant figures as these are exact numbers and can be represented by writing infinite number of zeros after placing a decimal. Example: 25 books can be represented as 25 = 25.00000000

Example

How many significant figures are present in 0.08700200?

Solution

There are seven significant figures.

$$0.0\underbrace{8700200}_{\text{Significant figures}}$$

- Addition and subtraction of significant figures:

- The result must not have more digits to the right of the decimal point than either of the original numbers.
- Example:

$$\begin{array}{r} 5.022 \\ 7.51 \\ 41.3 \\ \hline 53.832 \end{array}$$

Since 41.3 has only one digit to the right of the decimal point, the result will be 53.8.

- Multiplication and division of significant figures.

- The result must not have more significant figures than that of the original numbers with the few significant figures.
- Example: 1.1 × 2.134 = 2.3474

  Since 1.1 has two significant figures, the result will be 2.3.

- Rules for rounding off numbers:

- If greater than 5, then increased by 1
- If less than 5, then not changed
- If equal to 5, then increased by 1 (in case of odd number) and not changed (in case of even number)

Dimensional Analysis

- Method used to convert units from one system to other – Factor label method or unit factor method or dimensional analysis

Example

Calculate the volume of 3.5 L water in m^3.

Solution

We know that

$1\ L = 1000\ cm^3$

$= (10\ cm)^3$

Since 100 cm = 1 m,

10 cm = 0.1 m

Now, $1\ L = (10\ cm)^3$

$= (0.1\ m)^3$

$= (1 \times 10^{-1}\ m)^3$

$= 1 \times 10^{-3}\ m^3$

Therefore, $3.5\ L = 3.5 \times 1 \times 10^{-3}\ m^3$

$= 3.5 \times 10^{-3}\ m^3$

Laws of Chemical Combination

- **Law of Conservation of Mass:**

  Matter can be neither created nor destroyed. In other words, in a chemical reaction the mass of reactants is equal to the mass of the products.

- **Law of Definite and Multiple Proportions:**

  Law of definite proportions: A given compound always contains exactly the same proportion of elements by weight.

Law of multiple proportions: If two elements can combine to form more than one compound, then the masses of one element that combine with a fixed mass of the other element are in the ratio of small whole numbers.

- **Gay Lussac's Law of Gaseous Volumes:**

  When gases combine or are produced in a chemical reaction, they do so in a simple ratio by volume, provided all the gases are at the same temperature and pressure.

- **Avogadro Law:**

  Equal volumes of gases at the same temperature and pressure should contain equal number of molecules.

- **Postulates:**
- All matter is made of very tiny indivisible particles called atoms.
- All the atoms of a given element are identical in mass and chemical properties whereas those of different elements have different masses and chemical properties.
- Atoms of different elements combine in a fixed whole number ratio to form compounds.
- Chemical reactions involve reorganisation of atoms. Atoms are neither created nor destroyed in a chemical reaction.
- The laws of chemical combination could be explained by Dalton's atomic theory.

Atomic and Molecular Masses

- **Atomic mass:**
- The mass of an atom
- One atomic mass unit (1 amu) = Mass equal to one-twelfth of the mass of one carbon-12 atom

  1 amu = 1.66056×10^{-24} g

- Nowadays, 'u' (unified mass) has replaced 'amu'.
- Average atomic mass = $\sum$(Mass of isotope × Relative abundance)

Example

The relative abundance of two isotopes of copper, having atomic masses 62.93 u and 64.94 u, are 69.09% and 30.91% respectively. Calculate the average atomic mass of copper.

Solution:

Average atomic mass of copper = $\left(62.93\times\frac{69.09}{100}\right)+\left(64.94\times\frac{30.91}{100}\right)$

= 63.55 u

- **Molecular Mass:**
- Sum of the atomic masses of all the elements present in a molecule
- Example − Molecular mass of CO_2 = 1 × Atomic mass of carbon + 2 × Atomic mass of oxygen

  = (1 × 12.011 u) + (2 × 16.00 u)

  = 12.011 u + 32.00 u

  = 44.011 u
- **Formula Mass:**
- Sum of the masses of all the atoms present in a formula unit of a compound
- Used for compounds whose constituent particles are ions
- Example − Formula mass of sodium chloride (NaCl)

  = Atomic mass of sodium + Atomic mass of chlorine

  = 23.0 u + 35.5 u

  = 58.5 u

Know your scientists

Antoine Laurent Lavoisier (1743 – 1794): A French chemist and biologist postulated the "law of conservation of mass". He is known as "the father of modern chemistry".

Joseph Proust (1754 – 1856): A French Chemist immensely contributed to the chemical sciences by stating "the law of definite proportions".

John Dalton (1766 – 1844): This English chemist, meteorologist and physicist is known for his pioneering work towards the development of the modern atomic theory. He postulated "the law of multiple proportions".

Gay Lussac (1778 – 1850): A French chemist mathematically deduced volume relations and generalised his deductions as "the law of combining volumes".

Amedeo Avogadro (1776 – 1856): An Italian scientist is known worldwide for the term "mole". He defined the relationship between masses of same volume of different gases and their molecular weights. His law came to be known as Avogadro's law.

Mole Concept and Molar Masses

- 1 mole of any substance can be defined as:
- Amount of a substance that contains as many particles (atoms, molecules or ions) as there are atoms in 12 g of the ^{12}C isotope
- Avogadro number or Avogadro constant (N_A); it is equal to 6.022×10^{23} particles
- Example − 1 mole of oxygen atoms = 6.022×10^{23} atoms of oxygen

  1 mole of carbon dioxide molecules = 6.022×10^{23} molecules of carbon dioxide

  1 mole of sodium chloride = 6.022×10^{23} formula units of sodium chloride
- **Relative atomic mass:** Relative atomic mass of an element is the ratio of the average mass of one atom of an element to one-twelfth of the mass of an atom of carbon-12.

$$\text{Relative atomic mass} = \frac{\text{Average mass of Atom}}{\frac{1}{12} \times \text{Mass of carbon}\left(C^{12}\right)}$$

Molar mass of a substance can be defined as:

- Mass of one mole of a substance in grams

- Numerically equal to atomic/molecular/formula mass in u.

- Example − Molar mass of CO_2 = 44.011 g mol−1
- **Relative molecular mass:** It is defined as the ratio of the mass of a molecule to the atomic mass unit of the molecule. It is a unitless quantity.

$$\text{Relative molecular mass} = \frac{\text{Average mass of Molecule / Compound}}{\frac{1}{12} \times \text{Mass of carbon}\left(C^{12}\right)}$$

Molar mass of NaCl = 58.5 g mol^{-1}

Examples

1. What number of moles contains 3.011×10^{23} molecules of glucose?

Solution:

1 mole of glucose is equivalent to 6.022×10^{23} molecules of glucose.

Hence, 3.011×10^{23} molecules of glucose will be present in

$= \frac{1 \times 3.011 \times 10^{23}}{6.022 \times 10^{23}}$ mol = 0.5 mol (of glucose)

Thus, 0.5 mole of glucose contains 3.011×10^{23} molecules of glucose.

2. What is the mass of a fluorine molecule?

Solution:

1 mole of fluorine molecule contains 6.022×10^{23} molecules and weighs 38 g.

Therefore, mass of a fluorine molecule = $\frac{38}{6.022 \times 10^{23}}$ g

= 6.31×10^{-23} g

Atomicity

- It is defined as the total number of atoms of constituent elements which combine to form a molecule.
- One molecule of hydrogen combines with one molecule of chlorine to form two molecules of hydrogen chloride.
- One molecule of hydrogen or chlorine contains two atoms of each.

Percentage Composition

Mass percent of an element = $\frac{\text{Mass of that element in the compound} \times 100\%}{\text{Molar mass of the compound}}$

Example

What is the mass percent of oxygen in potassium nitrate? (Atomic mass of K = 39.10 u, atomic mass of N = 14.007 u, atomic mass of O = 16.00 u)

Solution:

Atomic mass of K = 39.10 u (Given)

Atomic mass of N = 14.007 u (Given)

Atomic mass of O = 16.00 u (Given)

Therefore, molar mass of potassium nitrate (KNO_3)

= 39.10 + 14.007 + 3(16.00)

= 101.107 g

Therefore, mass percent of oxygen in KNO_3

$= \frac{\text{Mass of oxygen in } KNO_3 \times 100\%}{\text{Molar mass of } KNO_3}$

$= \frac{3 \times 16.00}{101.107} \times 100\%$

= 47.47% (approx)

- Empirical formula and molecular formula:

Empirical formula	Molecular formula
Represents the simplest whole number ratio of various atoms present in a compound	Represents the exact number of different types of atoms present in a molecule of a compound

- Empirical formula is determined if mass % of various elements are known.
- Molecular formula is determined from empirical formula if molar mass is known.

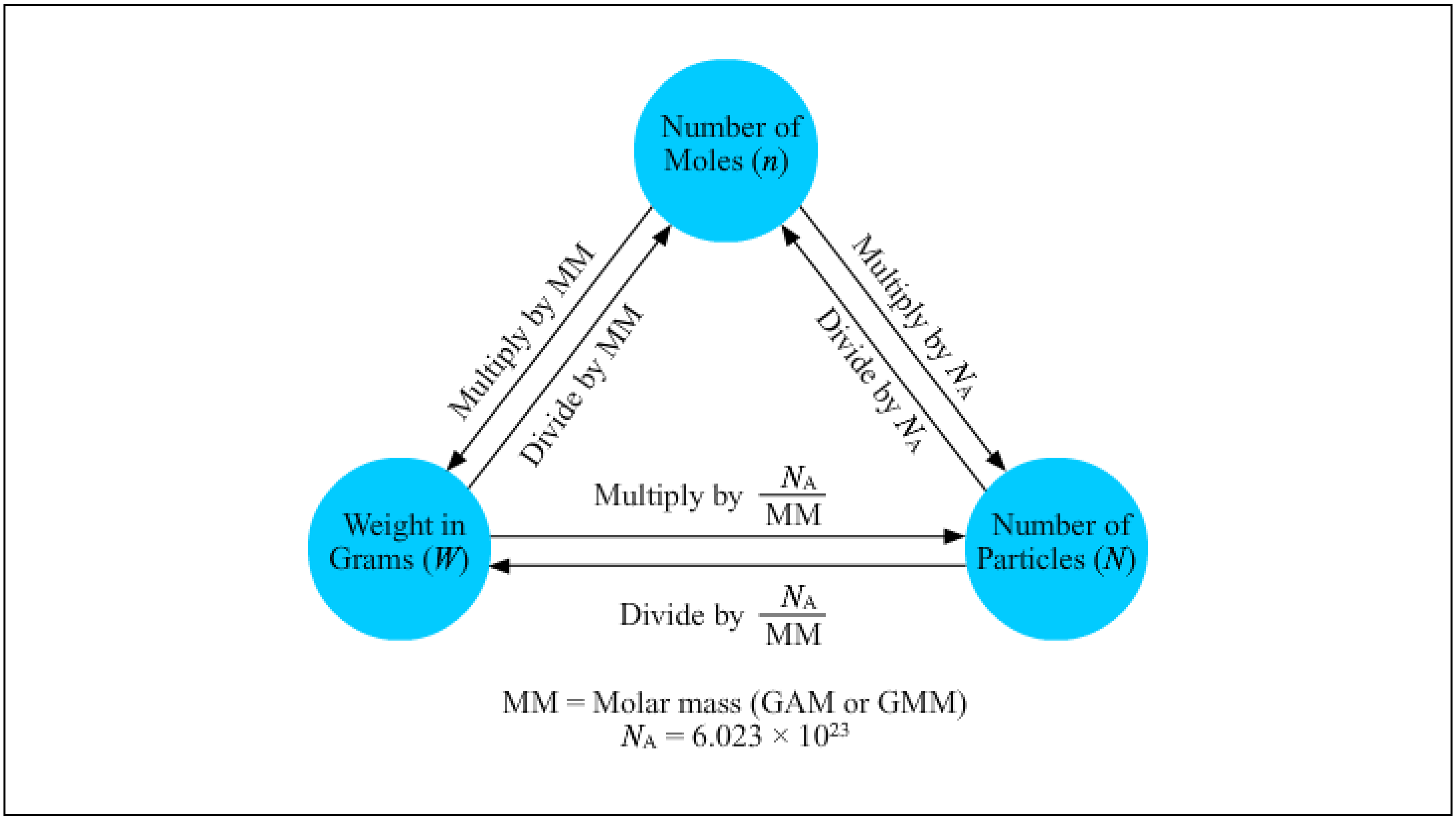
- **Example**

- A compound contains 92.26% carbon and 7.74% hydrogen. If the molar mass of the compound is 26.038 g mol^{-1}, then what are its empirical and molecular formulae?
- **Solution:**
- Mass percent of carbon (C) = 92.26% (Given)
- Mass percent of hydrogen (H) = 7.74% (Given)

  Therefore, 100 g of the compound contains 92.26 g and 7.74 g of of hydrogen

- Number of moles of carbon present in the compound = $\frac{92.26}{12.011}$
- = 7.68 mol
- Number of moles of hydrogen present in the compound $= \frac{7.74}{1.008}$
- = 7.68 mol
- Thus, in the given compound, carbon and hydrogen are present in the ratio C : H = 7.68 : 7.68
- = 1 : 1
- Therefore, the empirical formula of the compound is CH.
- Empirical formula mass of CH = (12.011 + 1.008)g
- = 13.019 g
- Molar mass of the compound = 26.038 g (Given)
- Therefore, n = $\frac{\text{Molar Mass}}{\text{Empirical formula mass}}$
- = $\frac{26.038\text{ g}}{13.019\text{ g}}$
- = 2
- Hence, the molecular mass of the compound is $(CH)_n$, i.e., $(CH)_2$ or C_2H_2.

Interconversion Among Number of Moles, Mass and Number of Molecules



Stoichiometric Calculations in Balanced Chemical Equations

- An example of a balanced chemical equation is given below.

$$C_3H_{8(g)} + 5O_{2(g)} \longrightarrow 3CO_{2(g)} + 4H_2O_{(l)}$$

From the above balanced chemical equation, the following information is obtained:

- One mole of $C_3H_8(g)$ reacts with five moles of $O_2(g)$ to give three moles of $CO_2(g)$ and four moles of $H_2O(l)$.

- One molecule of $C_3H_8(g)$ reacts with five molecules of $O_2(g)$ to give three molecules of $CO_2(g)$ and four molecules of $H_2O(l)$.

- 44 g of $C_3H_8(g)$ reacts with (5 × 32 = 160) g of $O_2(g)$ to give (3 × 44 = 132) g of $CO_2(g)$ and (4 × 18 = 72) g of $H_2O(l)$.

- 22.4 L of $C_3H_8(g)$ reacts with (5 × 22.4 = 112) L of $O_2(g)$ to give (3 × 22.4 = 67.2) L of $CO_2(g)$ and (4 × 22.4 = 89.6) L of $H_2O(l)$.

Example

Nitric acid (HNO_3) is commercially manufactured by reacting nitrogen dioxide (NO_2) with water (H_2O). The balanced chemical equation is represented as follows:

$$3NO_{2(g)} + H_2O_{(l)} \longrightarrow 2HNO_{3(aq)} + NO_{(g)}$$

Calculate the mass of NO_2 required for producing 5 moles of HNO_3.

Solution:

According to the given balanced chemical equation, 3 moles of NO_2 will produce 2 moles of HNO_3.

Therefore, 2 moles of HNO_3 require 3 moles of NO_2.

Hence, 5 moles of HNO_3 require $= \frac{3}{2} \times 5$ moles of NO_2

= 7.5 moles of NO_2

Molar mass of NO_2 = $(14 + 2 \times 16)$ g mol^{-1}

= 46 g mol^{-1}

Thus, required mass of NO_2 = (7.5×46) g

= 345 g

- Limiting reagent or limiting reactant:
- Reactant which gets completely consumed when a reaction goes to completion
- So called because its concentration limits the amount of the product formed

Example

Lead nitrate reacts with sodium iodide to give lead iodide and sodium nitrate in the following manner:

$$Pb(NO_3)_2 + 2NaI \longrightarrow PbI_2 + 2NaNO_3$$

What amount of sodium nitrate is obtained when 30 g of lead nitrate reacts with 30 g of sodium iodide?

Solution:

Molar mass of $Pb(NO_3)_2 = 207 + \left[\{14 + (16 \times 3)\} \times 2\right]$

= 331 g mol^{-1}

Molar mass of NaI = (23 + 127) = 150 g mol^{-1}

According to the given equation, 1 mole of $Pb(NO_3)_2$ reacts with 2 moles of NaI, i.e.

331 g of $Pb(NO_3)_2$ reacts with 300 g of NaI to give PbI_2 and $NaNO_3$

Thus, 30 g of $Pb(NO_3)_2$ will react with (30 × 300) / 331 g of NaI = 27.19g of NaI

However, we have 30 g of NaI. So, NaI is present in excess and $Pb(NO_3)_2$ is the limiting reagent.

Now, number of moles in 30 g of $Pb(NO_3)_2$ 30 g $= \frac{30}{331} = 0.09$ mole

According to the equation, 1 mole of $Pb(NO_3)_2$ gives 2 moles of $NaNO_3$. So

0.09 moles of $Pb(NO_3)_2$ will give (2 × 0.09) moles of $NaNO_3$ = 0.18 moles of $NaNO_3$.

- **Reactions in solutions:**

  Ways for expressing the concentration of a solution –

- Mass per cent or weight per cent (w/w%)

  Mass per cent $= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100\%$

Example

4.4 g of oxalic acid is dissolved in 200 mL of a solution. What is the mass per cent of oxalic acid in the solution? (Density of the solution = 1.1 g mL^{-1})

Solution:

Density of the solution = 1.1 g mL^{-1}

So the mass of the solution = (200 mL) × (1.1 g mL^{-1})

= 220 g

Mass of oxalic acid = 4.4 g

Therefore, mass per cent of oxalic acid in the solution

$$= \frac{\text{Mass of oxalic acid}}{\text{Mass of the solution}} \times 100\%$$
$$= \frac{4.4\text{g}}{220\text{g}} \times 100\%$$
$$= 2\%$$

- Mole fraction:

  If a substance 'A' dissolves in a substance 'B', then mole fraction of A $= \dfrac{\text{Number of moles of A}}{\text{Number of moles of the solution}}$

  $$= \frac{n_A}{n_A + n_B}$$

  Mole fraction of B $= \dfrac{\text{Number of moles of B}}{\text{Number of moles of the solution}}$

  $$= \frac{n_B}{n_A + n_B}$$

  n_A – Number of moles of A

  n_B – Number of moles of B

Example

A solution is prepared by dissolving 45 g of a substance **X** (molar mass = 25 g mol^{-1}) in 235 g of a substance **Y** (molar mass = 18 g mol^{-1}). Calculate the mole fractions of **X** and **Y.**

Solution:

Moles of X, $n_X = \frac{45}{25}$

= 1.8 mol

Moles of Y, $n_Y = \frac{235}{18}$

= 13.06 mol

Therefore, mole fraction of X, $n_X = \frac{1.8}{1.8+13.06}$

$= \frac{1.8}{14.86}$

$= 0.121$

And, mole fraction of Y, $n_Y = 1 - n_X$

= 1 − 0.121

= 0.879

- Molarity:

  Number of moles of solute in 1 L of solution

  $$\text{Molarity (M)} = \frac{\text{Number of moles of solute}}{\text{Volume of solution in litres}}$$

  For a given solution, the molarity equation is as follows:

  $M_1V_1 = M_2V_2$

M_1 = Molarity of a solution when its volume is V_1

M_2 = Molarity of the same solution when its volume is V_2

Examples

1. 10g of HCl is dissolved in enough water to form 500 mL of the solution. Calculate the molarity of the solution.

Solution:

Molar mass of HCl = 36.5 g mol^{-1}

So the moles of HCl = $\frac{10}{36.5}$ mol

= 0.274 mol

Volume of the solution = 500 mL = 0.5 L

Therefore, molarity = $\frac{\text{Number of moles of HCl}}{\text{Volume of solution in litres}}$

$= \frac{0.274 \text{ mol}}{0.5 \text{ L}}$

= 0.548 M

2. Commercially available concentrated HCl contains 38% HCl by mass. What volume of concentrated HCl is required to make 2.5 L of 0.2 M HCl? (Density of the solution = 1.19 g mL^{-1})

Solution:

38% HCl by mass means that 38g of HCl is present in 100 g of the solution.

Moles of HCl = $\frac{38}{36.5} = 1.04 \text{ mol}$

Volume of the solution $= \frac{\text{Mass}}{\text{Density}}$

$= \frac{100\text{g}}{1.19\text{g mL}^{-1}}$

= 84.03 mL

= 0.08403L

Therefore, molarity of the solution = $\frac{1.04}{0.08403\text{ L}}$

= 12.38 M

According to molarity equation,

$M_1V_1 = M_2V_2$

Here,

M_1 = 12.38 M

M_2 = 0.2 M

V_2 = 2.5 L

Now, $M_1V_1 = M_2V_2$

$$\Rightarrow V_1 = \frac{M_2V_2}{M_1}$$
$$= \frac{0.2 \times 2.5}{12.38}$$
$$= 0.0404 \text{ L} \quad \text{(approx)}$$

Hence, required volume of HCl = 0.0404 L

- Molality:

  Number of moles of solute present in 1 kg of solvent

  Molality (m) = $\frac{\text{Number of moles of solute}}{\text{Mass of solvent in kg}}$

Example

What is the molality of a solution of glucose in water, which is labelled as 15% (w/w)?

Solution:

15% (w/w) solution means that 15 g of glucose is present in 100 g of the solution, i.e. (100 − 15) g = 85 g of water = 0.085 kg of water

Moles of glucose = $\frac{15\text{g}}{180\,\text{g mol}^{-1}}$

= 0.083 mol

Therefore, molality of the solution $= \frac{0.083\,\text{mol}}{0.085\text{ kg}}$

= 0.976 m