

CHEMISTRY
REFERENCE STUDY MATERIAL

for

CLASS – IX

2017 – 18

**CHAPTER WISE CONCEPTS, FORMULAS AND
QUESTIONS INCLUDING HOTS QUESTIONS**

CHAPTER – 1

MATTER IN OUR SURROUNDINGS

All matters in the universe exist in three states. There are two ways of classification of matter.

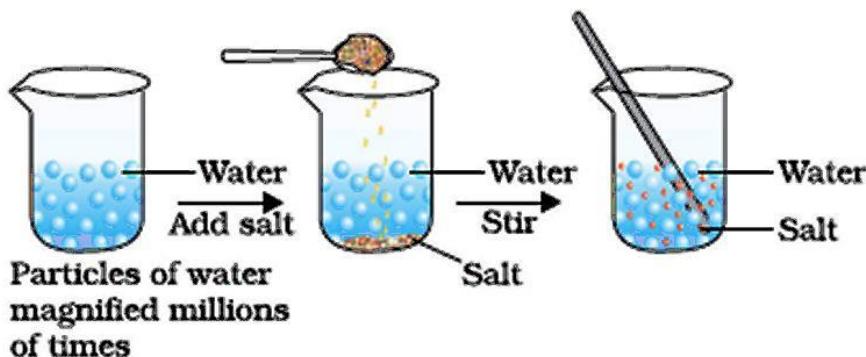
1. According to physical state as solid, liquid or gas.
2. According to its composition as element, compound or mixture.

CHARACTERISTICS OF PARTICLES OF MATTER

PARTICLES OF MATTER HAVE SPACE BETWEEN THEM- when we make tea, coffee or lemonade (nimbu paani), particles of one type of matter get into the spaces between particles of the other. This shows that there is enough space between particles of matter. Similarly particles of sugar, salt, Dettol, or potassium permanganate get evenly distributed in water.

PARTICLES OF MATTER ARE CONTINUOUSLY MOVING- Particles of matter are continuously moving, that is, they possess what we call the kinetic energy. As the temperature rises, particles move faster. So, we can say that with increase in temperature the kinetic energy of the particles also increases.

PARTICLES OF MATTER ATTRACT EACH OTHER- Particles of matter have force acting between them. This force keeps the particles together. The strength of this force of attraction varies from one kind of matter to another.



INTEXT QUESTIONS PAGE NO. 3

Q1. Which of the following are matter?

Chair, air, love, smell, hate, almonds, thought, cold, colddrink, smell of perfume.

Answer: Chair, air, smell, almonds, cold-drink and smell of perfume are matter because they have some weight and occupy space.

Q2. Give reasons for the following observation:

The smell of hot sizzling food reaches you several metres away, but to get the smell from cold food you have to go close.

Answer: Hot food evaporates easily. Its vapours diffuse between the air molecules and reach within a short time to a distant place. But the case is different with the cold food because it remains in solid form and does not mix with air molecules, so we have to go close to smell it.

Q3. A diver is able to cut through water in a swimming pool. Which property of matter does this observation show?

Answer: The phenomena of cutting the water by the diver show that matter has space between its particles.

Q4. What are the characteristics of the particles of matter?

Answer: Characteristics of particles of matter :

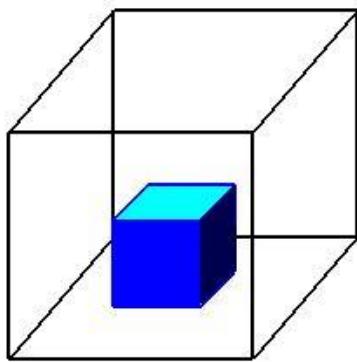
1. Particles of matter have space between them.
2. Particles of matter are continuously moving.
3. Particles of matter have an attraction force between them.
4. Particles of matter are very small in size.

STATES OF MATTER

1. THE SOLID STATE- all solids have a definite shape, distinct boundaries and fixed volumes, that is, have negligible compressibility. Solids have a tendency to maintain their shape when subjected to outside force. Solids may break under force but it is difficult to change their shape, so they are rigid.

2. THE LIQUID STATE- Liquids have no fixed shape but have a fixed volume. They take up the shape of the container in which they are kept. Liquids flow and change shape, so they are not rigid but can be called fluid. The rate of diffusion of liquids is higher than that of solids. This is due to the fact that in the liquid state, particles move freely and have greater space between each other as compared to particles in the solid state.

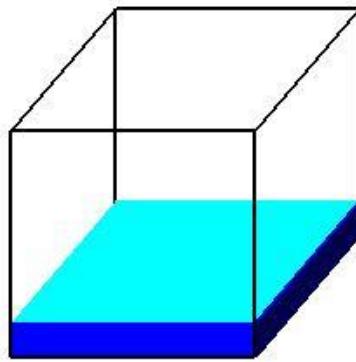
3. THE GASEOUS STATE- Gases are highly compressible as compared to solids and liquids. The liquefied petroleum gas (LPG) cylinder that we get in our home for cooking or the oxygen supplied to hospitals in cylinders is compressed gas. In the gaseous state, the particles move about randomly at high speed. Due to this random movement, the particles hit each other and also the walls of the container.



Solid

Holds Shape

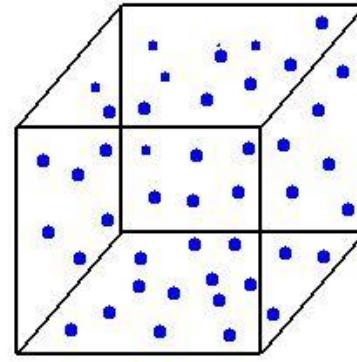
Fixed Volume



Liquid

**Shape of Container
Free Surface**

Fixed Volume



Gas

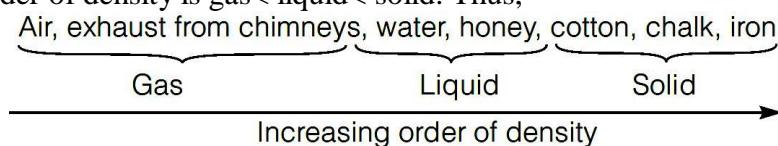
Shape of Container

Volume of Container

INTEXT QUESTIONS PAGE NO. 6

Q1. The mass per unit volume of a substance is called density. (density = mass/volume). Arrange the following in order of increasing density – air, exhaust from chimneys, honey, water, chalk, cotton and iron.

Answer: The order of density is gas < liquid < solid. Thus,



Q2. (a) Tabulate the differences in the characteristics of states of matter.

Answer:

(a) The differences in the characteristics of states of matter are given in the following table.

S. No.	Solids	Liquids	Gases
1	Definite shape and volume.	No definite shape. Liquids attain the shape of the vessel in which they are kept.	Gases have neither a definite shape nor a definite volume.
2	Incompressible	Compressible to a small extent.	Highly compressible
3	There is little space between the particles of a solid.	These particles have a greater space between them.	The space between gas particles is the greatest.
4	These particles attract each other very strongly.	The force of attraction between liquid particles is less than solid particles.	The force of attraction is least between gaseous particles.
5	Particles of solid cannot move freely.	These particles move freely.	Gaseous particles are in a continuous, random motion.

(b) Comment upon the following: rigidity, compressibility, fluidity, filling a gas container, shape, kinetic energy and density.

Answer:

(i) **Rigidity** The property due to which an object retains its shape and size is known as rigidity. Solids are rigid whereas liquids and gases are not.

(ii) **Compressibility** Compressibility is the property due to which a substance can be compressed, i.e., its volume can be decreased. Gases are compressible whereas solids and liquids are not.

(iii) **Fluidity** The property due to which a substance tends to flow is called fluidity. Gases and liquids are fluids, solids are not.

(iv) **Filling a gas container** A gas can be filled in a gas container by compressing it under high pressure. The property of compressibility (of gases) helps them in this regard.

(v) **Shape** The property of having a definite geometry is called shape of a particular substance. Solids have a definite shape whereas gases and liquids do not have.

(vi) **Kinetic energy** The energy possessed by an object or by the molecules of an object due to its state of motion is called kinetic energy. Molecules of gases possess highest kinetic energy. Increasing the temperature also increases the kinetic energy of a substance (or its molecules).

(vii) **Density** The mass per unit volume of a substance is called density.

Q3. Give reasons

(a) A gas fills completely the vessel in which it is kept.

Answer: There is little attraction between particles of gas. Thus, gas particles move freely in all directions. Therefore, gas completely fills the vessel in which it is kept.

(b) A gas exerts pressure on the walls of the container.

Answer: Particles of gas move randomly in all directions at high speed. As a result, the particles hit each other and also hit the walls of the container with a force. Therefore, gas exerts pressure on the walls of the container.

(c) A wooden table should be called a solid.

Answer: A wooden table has a definite shape and volume. It is very rigid and cannot be compressed i.e., it has the characteristics of a solid. Hence, a wooden table should be called a solid.

(d) We can easily move our hand in air but to do the same through a solid block of wood we need a karate expert.

Answer: Particles of air have large spaces between them. On the other hand, wood has little space between its particles. Also, it is rigid. For this reason, we can easily move our hands in air, but to do the same through a solid block of wood, we need a karate expert.

Q4. Liquids generally have lower density as compared to solids. But you must have observed that ice floats on water. Find out why.

Answer: The mass per unit volume of a substance is called density (density = mass/volume). As the volume of a substance increases, its density decreases. Though ice is a solid, it has large number of empty spaces between its particles. These spaces are larger as compared to the spaces present between the particles of water. Thus, the volume of ice is greater than that of water. Hence, the density of ice is less than that of water. A substance with lower density than water can float on water. Therefore, ice floats on water.

PROCESS OF CHANGE OF STATES OF MATTER:

- Vaporization
- Condensation
- Freezing
- Melting
- Sublimation
- Evaporation

VAPORIZATION:(CHANGE OF LIQUID INTO GAS):

The process of change of water into vapor is called vaporization. When water is heated after reaching at 100°C water starts boiling. At this temperature water turns into vapour. Since, water boils at 100°C , hence 100°C is called the boiling point of water.

BOILING POINTS

The common definition of boiling point is, the temperature at which a liquid boils is called its boiling point. Different liquid boils at different temperatures.

CONDENSATION:(CHANGE OF GAS INTO LIQUID)-

The change of vapor into water because of decrease in temperature is called condensation. Condensation is the reverse process of vaporization. When the temperature of vapor decreases it changes into water, this process is known as condensation.

Distilled water is manufactured by the condensation of vapor. The process of making of distilled water is known as distillation. In distillation first water is boiled to vaporize and the vapor is cooled, i.e. condensed to get distilled water.

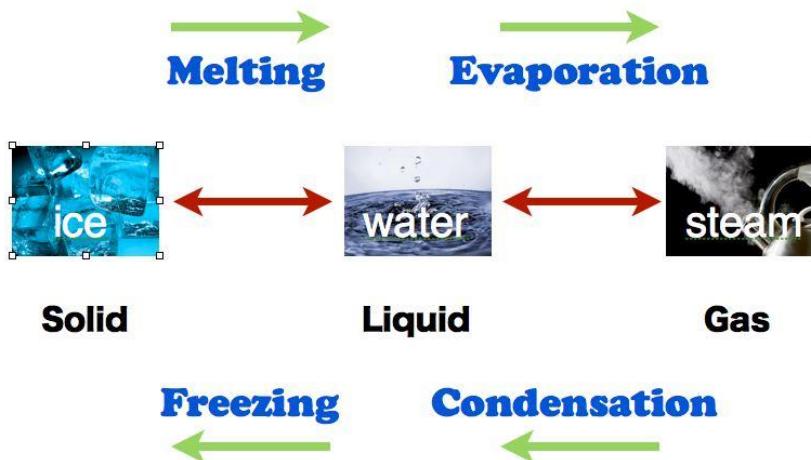
FREEZING:(CHANGE OF LIQUID INTO SOLID) -

The change of liquid into solid because of decrease in temperature is called freezing. Water changes into ice because of decrease in temperature, in other words water freezes into ice because of decrease in temperature. Water freezes at 0°C .

MELTING:(CHANGE OF SOLID INTO LIQUID)

The change of solid into liquid due to increase in temperature is known as melting. Ice, which is a solid melts, i.e. changes into water at 0°C .

Changes of State



LATENT HEAT

When water is heated up-to 100°C it starts boiling and changes into vapor. But we see that even after continuous supply of heat temperature does not rise above the 100°C while boiling of water.

Temperature supplied after 100°C to boiling water is used to change the water into vapor and temperature of water does not rise.

On the other hand we see that when heat is supplied to ice, temperature does not rise above the 0°C , until all ice melts. In this process also the heat supplied to ice after 0°C is used to change of ice into water and temperature of ice does not rise.

- Heat is used in these processes without rise in temperature is known as Latent Heat. The latent heat is used in change of states of matter, such as from solid to liquid or from liquid to gas without rising in temperature.
- Hence, Latent heat is the heat released or absorbed by a body during the process without change in temperature of the system. This happens while change of state of matter meaning a phase transition. Example - Melting of ice, boiling of water.

The particles of solid and liquid are bonded together with great force of attraction, because of which a matter exists in a particular state. When we supply heat to a solid or liquid, the heat is supplied without notice is used to break the force of attraction between particles and this heat is not used to increase the kinetic energy of particles. Since, kinetic energy of particles do not increase we do not see any rise in temperature of the system.

The word 'latent' is derived from the Latin word 'Latere' which means "to lie hidden". Joseph Black introduced the term Latent heat around 1762.

TYPES OF LATENT HEAT :

- Latent heat of fusion
- Latent heat of vaporization

LATENT HEAT OF FUSION (MELTING OR FREEZING):

When solid changes into liquid, the heat required changing the state without rising in temperature is called the Latent Heat of Fusion.

The change of solid to liquid state is an endothermic reaction as heat is required in it. The reaction in which heat is supplied or used is called endothermic reaction or process.

Let us take the example of melting of ice.

When heat is supplied to melt ice, temperature does not rise from 0°C even after continuous supply of heat till all ice melts. After melting of all ice temperature starts rising. The additional heat is required to melt the ice without coming into notice is the latent heat of fusion.

- The latent heat of fusion of ice is the energy which is used to change the state of ice (solid) to water (liquid).
- The quantity of heat required to convert 1 kilogram of solid to liquid without any change in temperature is called Latent Heat of fusion.
- The heat required is measured in joules (J).

3.34×10^5 joules of heat is required to convert 1 kilogram of ice into water at its melting point. Thus, the heat of fusion of ice at its melting point = 3.34×10^5 joules.

LATENT HEAT OF VAPORIZATION:

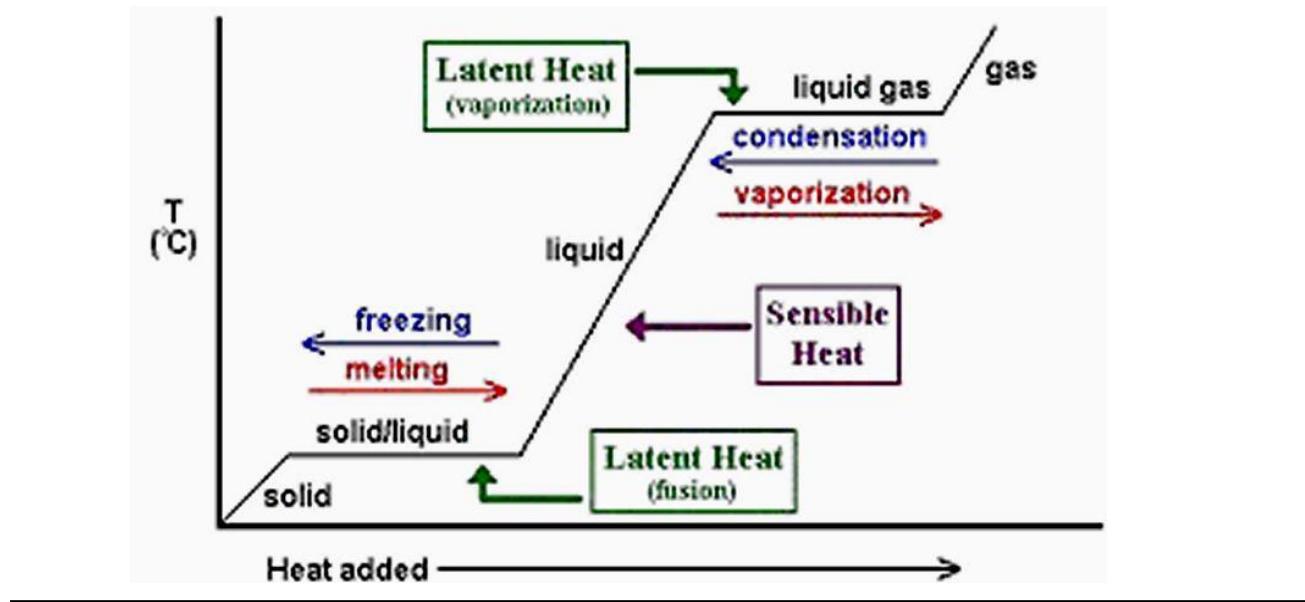
When liquid changes into gas because of rise in temperature, the heat required changing the state without rising in temperature is called the Latent Heat of vaporization.

The change of liquid to gaseous state is an endothermic reaction as heat is required in it.

Let us take the example of boiling of water.

Water boils at 100°C . When heat is supplied to water temperature does not rise after 100°C even after continuous supply of heat. The heat supplied at this stage is used to change water into vapor and hence does not come into notice. The additional heat is required to change the water into vapor without coming into notice is the latent heat of vaporization.

- The latent heat of vaporization of water is the energy which is used to change the state of water (liquid) to vapor (gas).
- 22.5×10^5 J energy is required to convert 1 kilogram of water into vapor. Hence, the latent heat of water is equal to 22.5×10^5 J per kilogram or it is written as 22.5×10^5 J/kg.
- Different liquid has different latent heat of vaporization.



SUBLIMATION:

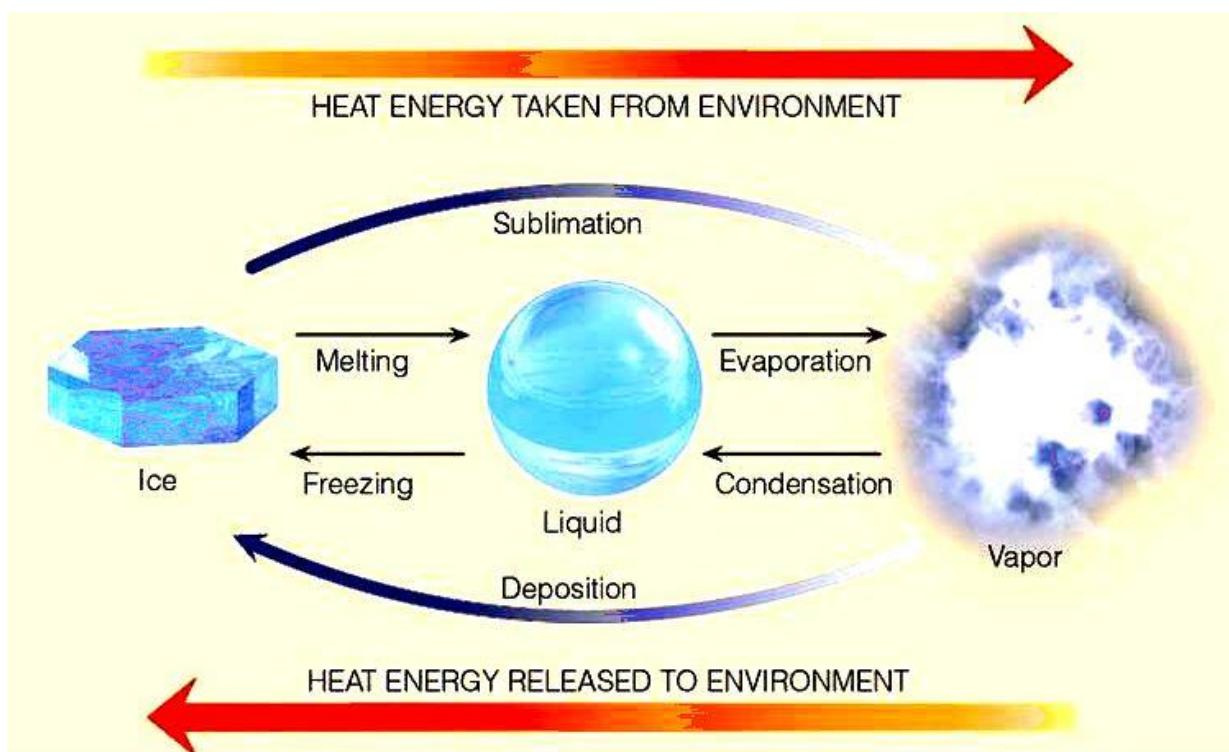
The process in which a solid changes into vapor without changing into liquid and from vapor changes into solid without changing into liquid is known as sublimation.

Generally solid first changes into liquid and then changes into gas because of rise in temperature. But there are many substances, which change into gas without changing into liquid and changes into solid from gas without changing into liquid. Such substances, which go under sublimation, are known as sublimate.

For example – camphor, naphthalene balls, ammonium chloride, iodine, dry ice, etc.

The solid obtained after cooling of the gas of sublimate is called Sublimate. The process of cooling of vapor of sublimate to get sublimate is also known as ‘sublimation’ although it is also known as deposition.

When camphor is heated it changes into vapor without changing into liquid. When the vapor of camphor is cooled it changes into solid without changing into liquid.



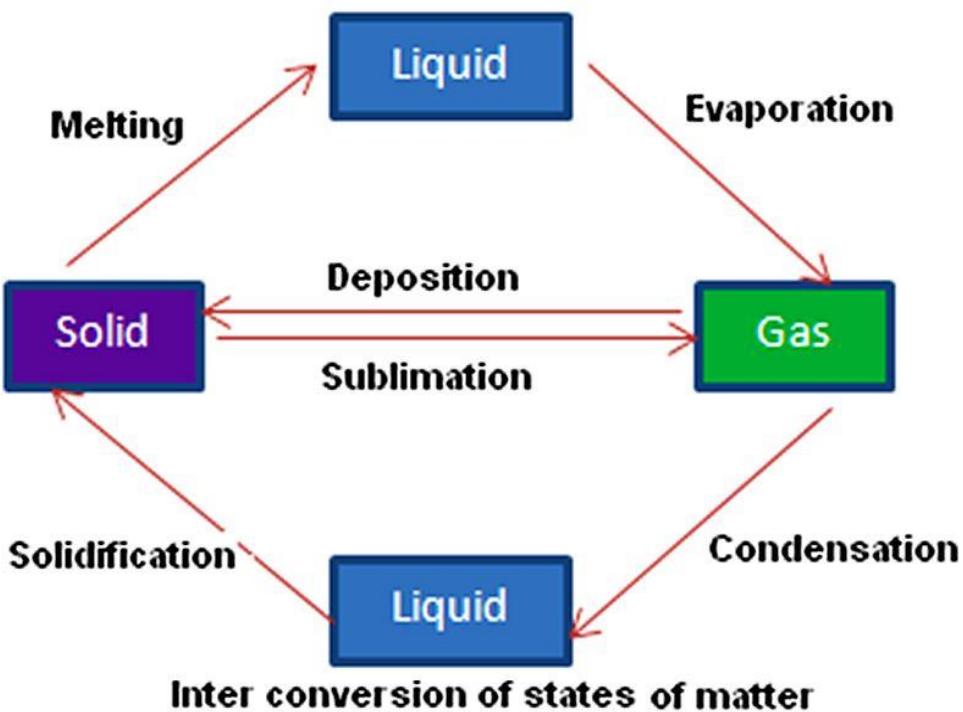
SUBLIMATION IN EVERYDAY LIFE:

The dry ice (solid carbon dioxide) turns into vapor without changing into liquid and is considered as sublimate. Because of this property dry ice is used to give the illusion of smoke or cloud on the stage in movies and stage shows.

Naphthalene balls are kept with cloths and documents to protect them from insects. Since naphthalene balls go under sublimation, hence it changes into vapor without changing into liquid and its vapor prevent the insects to come or stay in cloths or documents. This protects the documents and cloths to get destroyed.

We usually see that the size of naphthalene balls decrease gradually and finally disappeared when they are kept in open. This happens because of sublimation of naphthalene.

Naphthalene balls are used in toilets also as disinfectant and air freshener.



INTEXT QUESTIONS PAGE NO. 3

Q1. Convert the following temperature to celsius scale:

- a) 300 K b) 573 K.

Answer: (a) $300\text{ K} = (300 - 273)\text{ }^{\circ}\text{C} = 27\text{ }^{\circ}\text{C}$

(b) $573\text{ K} = (573 - 273)\text{ }^{\circ}\text{C} = 300\text{ }^{\circ}\text{C}$

Q2. What is the physical state of water at:

- a) 250°C b) 100°C ?

Answer:

(a) Water at 250°C exists in gaseous state.

(b) At 100°C , water can exist in both liquid and gaseous form. At this temperature, after getting the heat equal to the latent heat of vaporization, water starts changing from liquid state to gaseous state.

Q3. For any substance, why does the temperature remain constant during the change of state?

Answer: During a change of state, the temperature remains constant. This is because all the heat supplied to increase the temperature is utilised in changing the state by overcoming the forces of attraction between the particles. Therefore, this heat does not contribute in increasing the temperature of the substance.

Q4. Suggest a method to liquefy atmospheric gases.

Answer: Applying high pressure and cooling a gas to low temperature helps in the liquification of atmospheric gases. The reason is that under such conditions of temperature and pressure, the molecules of gases come closer, their kinetic energy becomes less and the gas is liquefied.

EVAPORATION

The change of liquid into vapor without reaching at its boiling point is called Evaporation. Evaporation takes place only at the surface of liquid while vaporization takes place on the whole mass of liquid.

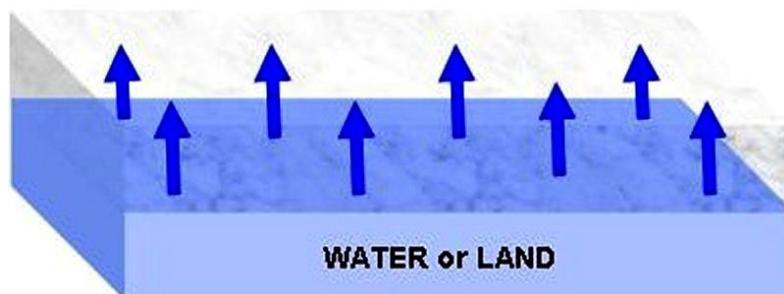
Evaporation takes place even at room temperature also. Evaporation speeds up with rise in temperature.

When water or other liquid is left in open, gradually its volume decreases. If you left some water in a pot in open, after two or three days water disappears. If the water left in garden or in a open balcony, it disappears quickly than kept in a room. This happens because of evaporation.

PROCESS OF EVAPORATION

Molecules at the surface of water, when exposed some temperature, their kinetic energy increases. Because of increase in kinetic energy those molecules become able to overcome the force of attraction between the particles of liquid. After getting required kinetic energy and decrease in force of attraction, they escape in the air in the form of vapor. Additionally those kinetic energy get some of the required kinetic energy from their neighboring molecules also because of which the temperature of the adjacent molecules decrease, which finally result in decrease of the temperature of surface of liquid.

**EVAPORATION CONTINUOUSLY MOVES
WATER FROM THE SURFACE TO THE ATMOSPHERE**



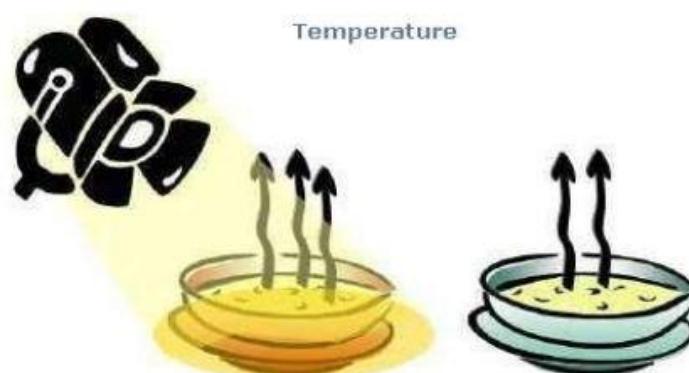
FACTORS AFFECTING THE EVAPORATION

- Temperature
- Pressure
- Surface area
- Humidity in air
- Wind speed

TEMPERATURE

Evaporation increases with increase in temperature and decreases with decrease in temperature. This means rate of evaporation is directly proportional to the temperature.

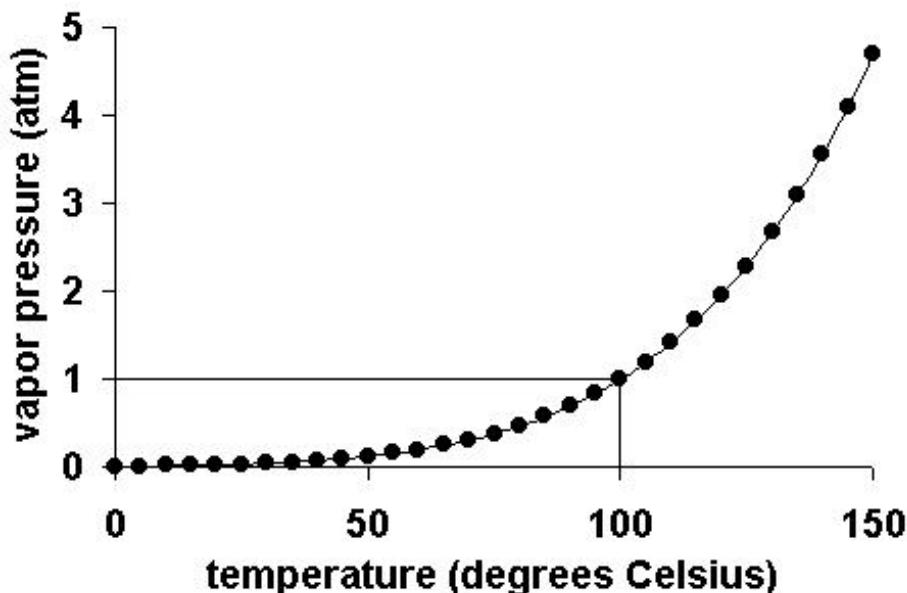
With increase in temperature the particles of liquid at surface get required kinetic energy to overcome the force of attraction and escape in air quickly. Hence, the increase in temperature increases the rate of evaporation.



PRESSURE

Evaporation decreases with increase in pressure and increases with decrease in pressure. This means the rate of evaporation is indirectly proportional to the pressure.

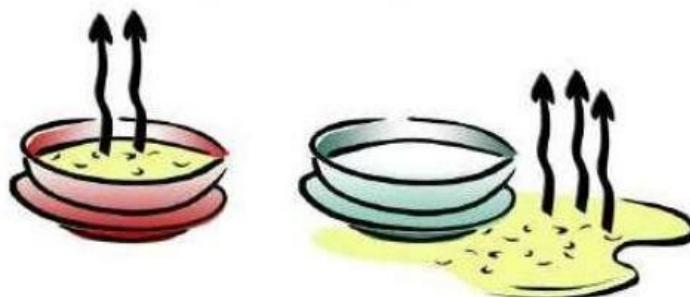
Particles at the surface of liquid require more kinetic energy to escape in air when there is more pressure over it and hence rate of evaporation will decrease. While if there is less pressure over the surface of liquid, the particles would require less kinetic energy to escape in air and hence rate of evaporation will increase. Therefore, increase in pressure slows down the rate of evaporation and decrease in pressure speeds up the rate of evaporation.



SURFACE AREA

Evaporation increases with increase in surface area and decreases with decrease in surface area.

Since evaporation takes place at the surface of liquid only, hence if the more surface of liquid is exposed to atmosphere more particles will receive the required temperature to get the required kinetic energy to escape in air. Therefore, evaporation takes place more rapidly with larger surface area. This means rate of evaporation increases with increase in surface area and decreases with decrease in surface area.



After rain roads are dried up quickly than pot holes. This happens because of increase in surface area of water. On roads water is spread over a large area, because of that large area of water exposed to atmosphere, and evaporation of water takes place quickly resulting in quickly drying of the roads. While in pot holes less water surface is exposed to air because of that less water area could come in contact with air and receives less temperature, resulting in delayed evaporation. That's why water dried from road quickly than in pot holes.

Wet clothes are spread up over the laundry line to get them dried up quickly. More surface area of water exposes to the air because of spreading of clothes this speeds up the rate of evaporation and clothes are dried up quickly. On the other hand if wet clothes are left even in the sun without spreading, they take more time to get dried because of less surface area

exposed to air. That's why wet clothes are kept spread over laundry line to get dried up quickly.

Water kept in a plate evaporates quickly than water kept in a tumbler (glass). This happens because in plate more surface area of water exposed to atmosphere which receives more heat and evaporates quickly. While in a glass less surface area of water exposed to atmosphere because of that less molecules of water receives heat from the atmosphere and evaporates slowly compare to the water exposed with large surface area.

Hence, rate of evaporation increases with increase in surface area and decreases with decrease in surface area.

HUMIDITY IN AIR AND EVAPORATION

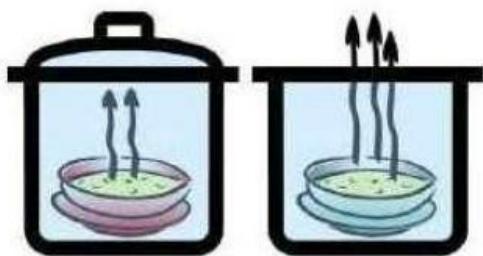
Evaporation decreases with increase in humidity and increases with decrease in humidity present in air. This means rate of evaporation is indirectly proportional to the humidity present in air.

Humidity is the amount of water vapor present in air. In weather reports, which are published in news paper or given on the TVs channels, the humidity percent is given, which shows the percent of water vapor present in air.

Because of more water vapor present in air the water holding capacity of atmosphere decreases which decrease the rate of evaporation. If air is dry then it can hold more water and thus in dry air rate of evaporation increases.

This is the cause that our cloths get dried up quickly in summer and winter than in rainy season. Because in rainy season there is more water vapor present in air, which decrease the water holding capacity of atmosphere resulting in decrease the rate of evaporation and our cloths do not dry up quickly in the rainy season.

Our sweat does not dry up quickly in rainy season. Because of that we feel uneasiness because of damp. This is because of higher percent of humidity present in air decreases the rate of evaporation in rainy season and our sweat does not evaporate quickly and we feel uneasy because of damp.



WIND SPEED AND EVAPORATION

Evaporation increases with the increase in wind speed and decreases with decrease in wind speed. This means rate of evaporation is directly proportional to the speed of wind.

Speedy wind propelled away some of the particles of water with it which speeds up the rate of evaporation. That's why speedy wind speeds up the rate of evaporation.



We see that wet cloth is dried up quickly in a windy day since the wind speeds up the rate of evaporation.

The wet clothes are given jerks before hanging them on laundry line because by giving jerks some of the water droplets propelled out. This reduces the presence of water in the wet cloths and they dried up quickly.

EVAPORATION IN EVERYDAY LIFE:

(a) Water from the surface of oceans, seas and other large water bodies evaporate continuously as they are exposed to atmosphere. The water vapor because of evaporation rises up in air and cumulates in the form of cloud, which makes the rain. Hence, evaporation is one of the essential parts of water cycle. Thus we can say that evaporation is one of the most natural phenomena for us.

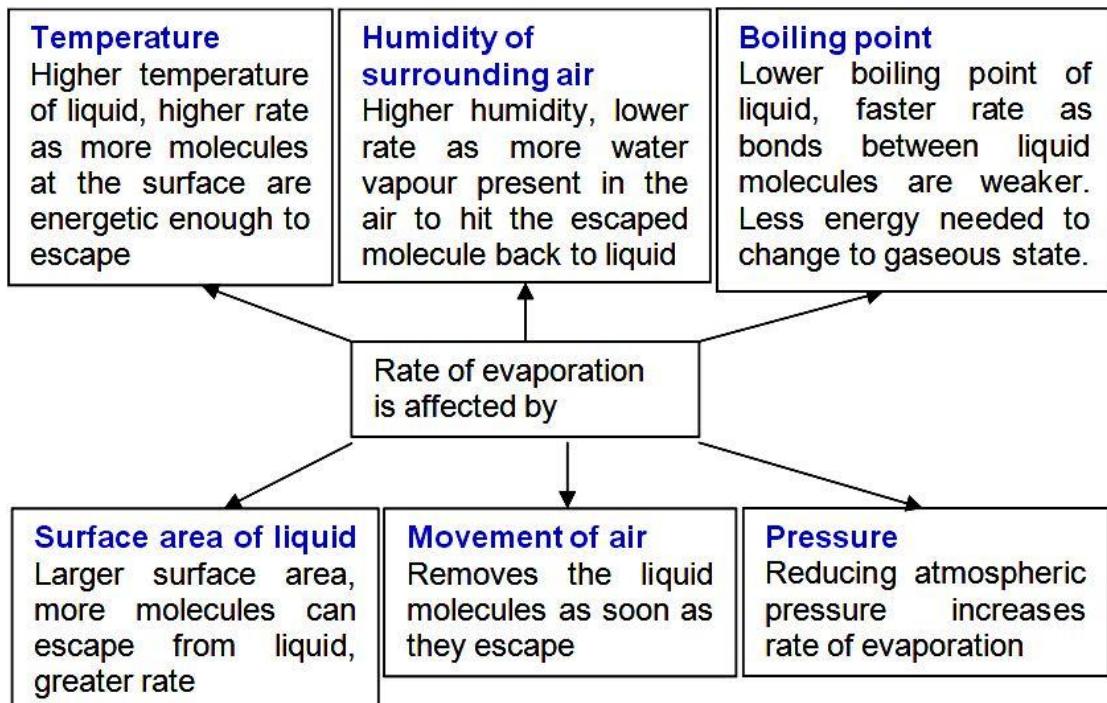
(b) In summer days sweats come out to regulate the temperature of our body. The sweat evaporates because of increase in surface area and getting the temperature from atmosphere. This is resulting in decrease in temperature of skin and finally our body, which gives relief to us in hot days. Sweating is a natural mechanism to keep cool the surface of our body in hot days. This is the cause that with increase in temperature our body sweats a lot especially in hot summer days.

(c) Water is kept in earthen pots to keep them cool. Earthen pot has lot of pores. Water kept in earthen pots evaporates from the pores of pots, which cools the neighboring molecules of water. This process continues and whole of the water kept in the earthen pots become cooler. Hence, water is kept in the earthen pot to keep them cool for drinking purpose.

(d) Wet clothes on the laundry lines are dried up because of evaporation.

(e) Sea water has lot of salt. Sea water is left in shallow pond. The water evaporates gradually because of heat of sun leaving the salt in the shallow pond. These salts are collected and used for with food after purification.

(f) Evaporative coolers are widely used in hot summer days. In evaporative coolers, dry air is blow over husk saturated of water. From the surface of husk water is evaporated resulted in cooling of husk. The water particles at the surface of wet husk evaporate and cool the rest portion of wet husk. Air blown from the cool husk is sent in the room, which cools the room.



INTEXT QUESTIONS PAGE NO. 6

Q1. Why does a desert cooler cool better on a hot dry day?

Answer: The rate of evaporation increases with increase in temperature and decrease in humidity. A desert cooler functions on the basis of evaporation. As evaporation increases when the day is hot and dry, so the desert cooler functions to a better extent.

Q2. How does the water kept in an earthen pot (matka) become cool during summer?

Answer: The surface of the earthen pot (matka) has tiny pores. The water stored in the earthen pot (matka) evaporates faster through these pores due to the increased exposed surface area. As the process of evaporation causes cooling, the stored water inside the earthen pot (matka) becomes cool.

Q3. Why does our palm feel cold when we put some acetone or petrol or perfume on it?

Answer: Acetone, petrol, perfume, etc., being volatile, evaporate very fast when exposed to larger surfaces. During the process they absorb the required latent heat of vaporisation from the palm (if kept on palm). So, the process causes cooling and the palm feels cool.

Q4. Why are we able to sip hot tea or milk faster from a saucer rather than a cup?

Answer: In a saucer, the exposed surface area of tea or milk is greater as compared to the cup. Therefore, the evaporation is faster and it is easier to sip colder tea or milk.

Q5. What type of clothes should we wear in summer?

Answer: We should wear cotton clothes in summers. During summers, we sweat more. On the other hand, cotton is a good absorber of water. Thus, it absorbs sweat from our body and exposes the liquid to the atmosphere, making evaporation faster. During this evaporation, particles on the surface of the liquid gain energy from our body surface, making the body cool.

EXERCISE QUESTIONS PAGE NO. 12

Q1. Convert the following temperatures to the Celsius scale.

- (a) 300 K (b) 573 K

Answer: Kelvin is an SI unit of temperature, where $0^{\circ}\text{C} = 273\text{ K}$

$$(a) 300\text{ K} = (300 - 273)\text{ }^{\circ}\text{C} = 27\text{ }^{\circ}\text{C}$$

$$(b) 573\text{ K} = (573 - 273)\text{ }^{\circ}\text{C} = 300\text{ }^{\circ}\text{C}$$

Q2. Convert the following temperatures to the Kelvin scale.

- (a) 25°C (b) 373°C .

Answer: Kelvin is an SI unit of temperature, where $0^{\circ}\text{C} = 273\text{ K}$

$$(a) 25\text{ }^{\circ}\text{C} = (25 + 273)\text{ K} = 298\text{ K}$$

$$(b) 373\text{ }^{\circ}\text{C} = (373 + 273)\text{ K} = 646\text{ K}$$

Q3. Give reason for the following observations.

- (a) Naphthalene balls disappear with time without leaving any solid.

- (b) We can get the smell of perfume sitting several metres away.

Answer:

(a) Naphthalene undergoes sublimation easily i.e., the change of state of naphthalene from solid to gas takes place easily. Thus, naphthalene balls disappear with time without leaving any solid.

(b) Gaseous particles possess high speed and large spaces between them. Particles of perfume diffuse into these gaseous particles at a very fast rate and reach our nostrils. This enables us to smell the perfume from a distance.

Q4. Arrange the following substances in increasing order of forces of attraction between the particles—

water, sugar, oxygen.

Answer:

Sugar is a solid; the forces of attraction between the particles of sugar are strong. Water is a liquid; the forces of attraction here are weaker than sugar. Oxygen is a gas; the forces of attraction are the weakest in gases.

Thus, the increasing order of forces of attraction between the particles of water, sugar and oxygen is

Oxygen < Water < Sugar

Q5. What is the physical state of water at—

- (a) 25°C (b) 0°C (c) 100°C ?

Answer:

(a) Water at 25°C is present in the liquid state.

(b) At 0°C , water can exist as both solid and liquid. At this temperature, after getting the heat equal to the latent heat of fusion, the solid form of water i.e., ice starts changing into its liquid form i.e., water.

(c) At 100°C , water can exist as both liquid and gas. At this temperature, after getting the heat equal to the latent heat of vaporization, water starts changing from its liquid state to its gaseous state, i.e., water vapours.

Q6. Give two reasons to justify—

- (a) **water at room temperature is a liquid.**
(b) **an iron almirah is a solid at room temperature.**

Answer:

(a) At room temperature (25°C), water is a liquid because it has the following characteristic of liquid:

- i). At room temperature, water has no shape but has a fixed volume that is, it occupies the shape of the container in which it is kept.
- ii). At room temperature, water flows.

(b) An iron almirah is a solid at room temperature (25°C) because:

- i). it has a definite shape and volume like a solid at room temperature.
- ii). it is rigid as solid at room temperature.

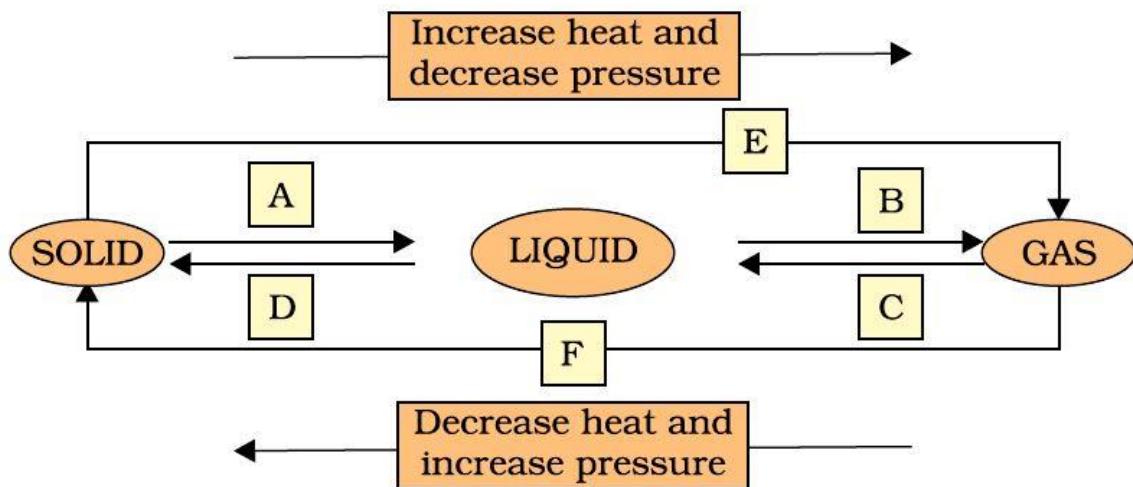
Q7. Why is ice at 273 K more effective in cooling than water at the same temperature?

Answer: When ice melts, it absorbs the energy equal to the latent heat of fusion too from the surroundings so it causes cooling more effectively than the water at same temperature (because water does not absorb energy from the surroundings).

Q8. What produces more severe burns, boiling water or steam?

Answer: Steam causes more severe burns than boiling water. The reason is that it releases the extra amount of heat (latent heat) which it has already taken during vaporisation (when the steam was formed from water).

Q9. Name A,B,C,D,E and F in the following diagram showing change in its state



Answer:

A = Melting or fusion, here the solid changes into liquid.

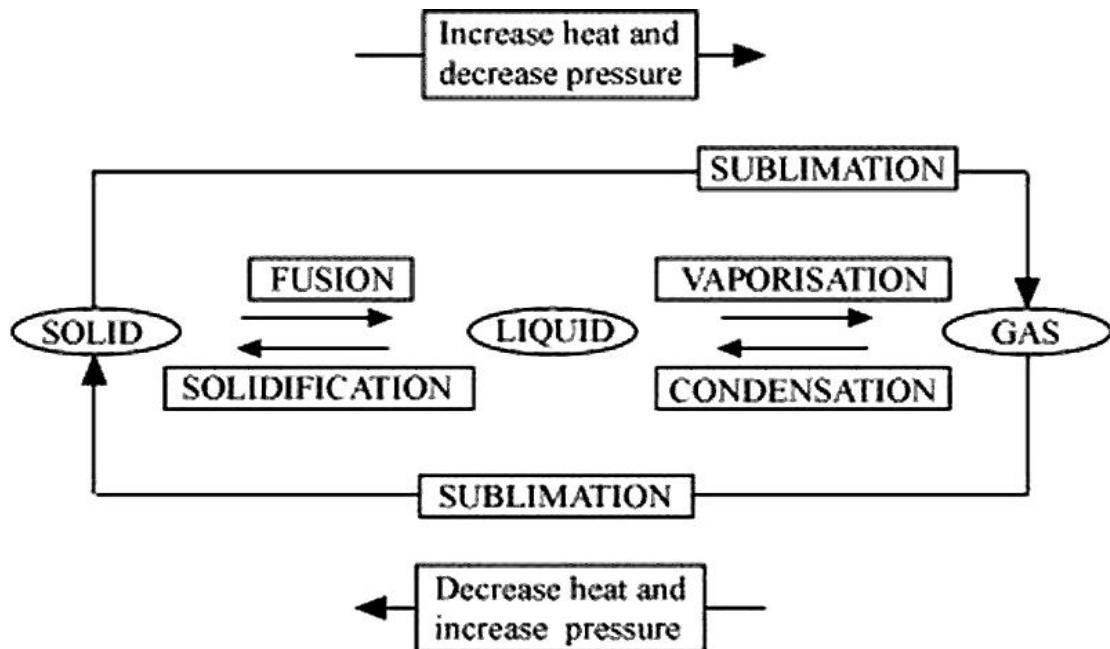
B = Evaporation or vaporisation, here the liquid changes into gas.

C = Condensation or liquification, here the gas changes into liquid.

D = Freezing or solidification, here the liquid changes into solid.

E = Sublimation, here solid directly changes into gas without coming in liquid state.

F = Sublimation, here gas changes into solid without coming to liquid state.



ASSIGNMENT QUESTIONS SET – 1
CHAPTER – 1
MATTER IN OUR SURROUNDINGS

- 1.** Fill in the blanks
 - a. The process of _____ causes cooling.
 - b. The process of cooling glass is known as _____.
 - c. Liquids have no fixed _____ but have fixed _____.
 - d. _____ exists in all three states of matter.
 - e. Carbon dioxide is a white solid called _____ at temperature below _____.

- 2.** State True or False
 - a. Evaporation of water is a bulk phenomenon.
 - b. Diffusion takes place in haphazard and random way.
 - c. SI unit of pressure is Pascal.
 - d. A gas is highly incompressible fluid.
 - e. Solids and liquids can be identified from their characteristic melting and boiling points.

- 3.** Zig-zag movement of the solute particle in a solution is known as
 - (a) Linear motion (b)
 - Circular motion (c)
 - Brownian motion (d)
 - Curved motion.

- 4.** Gases can be liquified by
 - (a) increasing pressure
 - (b) decreasing temperature
 - (c) both (a) and (b)
 - (d) decreasing pressure

- 5.** Density of a substance is defined as
 - (a) ratio of mass and volume
 - (b) product of mass and volume
 - (c) ratio of mass and temperature
 - (d) product of mass and temperature

- 6.** Which of the following is not matter
 - (a) Blood
 - (b) Humidity
 - (c) Electron
 - (d) Moon rock

- 7.** Which is more effective in cooling?
 - (a) Ice at 0°C
 - (b) Water at 0°C
 - (c) Water at 100°C
 - (d) Ice at 100°C

- 8.** 0°C temperature is equal to
(a) 0 K
(b) 273 K
(c) -273 K
(d) 300 K
- 9.** The process involving the change of state from solid to gas is called
(a) melting
(b) boiling
(c) sublimation
(c) fusion
- 10.** A solid has
(a) definite volume and no definite shape
(b) no definite volume no definite shape
(c) definite shape and volume
(d) definite shape but no definite volume
- 11.** A liquid has
(a) definite volume and no definite shape
(b) no definite volume no definite shape
(c) definite shape and volume
(d) definite shape but no definite volume
- 12.** A gas has
(a) definite volume and no definite shape
(b) no definite volume no definite shape
(c) definite shape and volume
(d) definite shape but no definite volume
- 13.** Which of the following is NOT a property of particles of a matter?
(a) The particles of matter are extremely small
(b) The particles of matter have spaces between them.
(c) The particles of matter are in stationery state.
(d) The particles of matter attract each other.
- 14.** Which of the following has minimum spaces among the particles?
(a) Solids
(b) Liquids
(c) Gases
(d) None of these
- 15.** During summer, water kept in an earthen pot becomes cool because of the phenomenon of
(a) diffusion
(b) transpiration
(c) osmosis
(d) evaporation
- 16.** Rate of diffusion is the fastest in
(a) Solids
(b) Liquids
(c) Gases
(d) None of these

- 17.** Thermal conduction takes place in
(a) solids only
(b) liquids only
(c) gases only
(d) solids, liquids and gases.
- 18.** Evaporation always causes
(a) thermal expansion
(b) Liquification
(c) Cooling down
(d) all of these
- 19.** A change of state directly from solid to gas without changing into liquid state (or vice versa) is called
(a) Eavalopartion
(b) Sublimation
(c) Diffusion
(d) Condensation
- 20.** The rate of evaporation decreases with
(a) increase in humidity
(b) increase of temperature
(c) increase in wind speed
(d) increase of surface area
- 21.** Expand CNG and LPG.
- 22.** Arrange the following substances in increasing order of force of attraction between the particles. (i) milk (ii) salt (iii) oxygen
- 23.** Why is sponge a solid though compressible?
- 24.** Write one important characteristic of matter.
- 25.** Why does a desert cooler cool better in a hot dry day?
- 26.** Convert: (a) 25°C into kelvin scale (b) 500 K into celsius scale
- 27.** Why does the smell of hot sizzling food reach you several metres away but to get the smell from cold food you have to go close?
- 28.** What is the term used for change of solid state to liquid state?
- 29.** Name the temperature at which solid and liquid states of matter can coexist.
- 30.** Define evaporation.
- 31.** "The wool being knitted into a sweater is a physical change." Justify the statement.
- 32.** Mention two ways to liquefy atmospheric gases.
- 33.** What is the value of boiling point of water on Kelvin Scale of temperature?
- 34.** What is dry ice?
- 35.** (a) Dry ice is compressed at high pressure. What happens when pressure is released?
(b) Suggest a method to liquefy atmospheric gases.

- 36.** (a) The melting points of 2 substances A & B are 280 K and 320 K respectively. Are these substances liquid at room temperature? Justify your answer.
(b) Give an example that shows the state of matter can be changed into another state by changing the temperature.
- 37.** How will you show that matter is composed of tiny particles?
- 38.** Define (i) Latent heat of fusion and (ii) latent heat of vapourisation.
- 39.** Explain how the following factors affect the rate of evaporation of a liquid:
(i) temperature of the liquid.
(ii) area of the exposed surface.
(iii) moisture in the surrounding air.
(iv) increase in wind speed.
- 40.** When a bottle of scent is opened in one corner of a room, it immediately spreads throughout the room. What property of matter is responsible for this observation? Explain.
- 41.** (a) Conversion of solid to vapour is called sublimation. Name the term used to denote the conversion of vapour to solid.
(b) Conversion of solid state to liquid state is called fusion; what is meant by latent heat of fusion?
- 42.** Both boiling and evaporation convert a liquid into vapours. What is the difference between the two processes?
- 43.** A sample of water under study was found to boil at 102°C at normal pressure. Is the water pure? Will this water freeze at 0°C ? Comment.
- 44.** Why does the temperature of a substance remain constant during its m.pt. or boiling point?
- 45.** Answer the following questions:
(i) Arrange the following substances in increasing order of force of attraction between the particles.
 (i) water (ii) hydrogen (iii) sand
(ii) Why does the temperature remain constant at the melting point?
(iii) Which property of gases makes it possible to fill large volume of gases in small cylinders?
- 46.** Answer the following questions:
(a) Why is ice at 273 K more effective in cooling than water at the same temperature?
(b) Name the two gases which are supplied in compressed form in homes and hospitals.
- 47.** You want to wear your favourite shirt to a party but the problem is that it is still wet after a wash. What steps would you take to dry it fast?
- 48.** Give reasons:
(a) Steam produces more severe burns than boiling water.
(b) We are able to sip hot tea faster from a saucer rather than from a cup.
(c) Water kept in an earthen pot becomes cool during summer.

49. Why do cotton clothes suit best in summer?

50. Classify the following into osmosis and diffusion:

- (a) Swelling up of a resin on keeping in water.
 - (b) Spreading of virus on sneezing.
 - (c) Earthworm dying on coming in contact with common salt.
 - (d) Shrinking of grapes kept in thick sugar syrup.
 - (e) Preserving pickles in salt.
 - (f) Aquatic animals using oxygen dissolved in water during respiration.
-

ASSIGNMENT QUESTIONS SET – 2
CHAPTER – 1
MATTER IN OUR SURROUNDINGS

- 1.** What is a matter?
- 2.** Sodium chloride and sugar have same appearance. Are they same or different?
- 3.** All substances around us are alike. How?
- 4.** How can we say that air is a matter?
- 5.** State the characteristics of matter?
- 6.** What are the intensive properties of matter?
- 7.** What are the extensive properties of matter?
- 8.** State the characteristics of particles of matter.
- 9.** What is the effect of temperature on a matter?
- 10.** The smell of hot sizzling food reaches us several metres away. Why?
- 11.** What is diffusion?
- 12.** If a bottle of perfume is opened in one corner of a room, it immediately spreads throughout the room. Why?
- 13.** Name the three states of matter with examples.
- 14.** State the characteristics of solids.
- 15.** What are the characteristics of liquids?
- 16.** What are the characteristics of gases?
- 17.** What are fluids?
- 18.** Which of the following substances is most compressible? CO_2 , H_2O , NaCl .
- 19.** Which property of a gas results in steady pressure of the gas ?
- 20.** In which of the following substances, weakest inter molecular force is expected: H_2O ,
 CH_3OH ,
- 21.** One gas mixes with another gas easily. What is this property called ?
- 22.** Describe briefly (i) Melting point and (ii) Boiling point.
- 23.** How would you find out whether a sample of sodium chloride is pure or impure ?
- 24.** How will you find out whether a sample of water is pure or impure ?
- 25.** Why do solids have a fixed shape and gases have neither a fixed shape nor a fixed volume?
- 26.** What is Vaporization ?
- 27.** What is Sublimation ?
- 28.** What is Condensation ?
- 29.** What is Deposition ?
- 30.** What is Liquefaction ?
- 31.** What is Solidification ?

- 32.** What is difference between vapour and gas.
- 33.** Why do the three states of matter differ ?
- 34.** Why does the temperature remain constant until whole of the solid changes into liquid, though the heat energy is constantly supplied ?
- 35.** Why does the temperature remain constant during boiling though heat is constantly supplied ?
- 36.** Why does a gas fill a vessel completely ?
- 37.** How does the state of matter changes from solid to liquid and then to gas on heating ?
- 38.** Why evaporation is called surface phenomenon?
- 39.** List two processes from which it may be concluded that the particles of a gas move continuously.
- 40.** At what temperature does solid ice and liquid water co-exist together?
- 41.** What is common among the three states of matter?
- 42.** Which property of gas is used in supplying oxygen cylinders to hospitals?
- 43.** A substance x was highly compressible and could easily be liquefied. it could also take the shape of the container. Predict the nature of the substance
- 44.** What is the state of water at 100 degree celsius, zero degree celsius and 4 degree celsius?
- 45.** Can a liquid turn into vapor without heating?
- 46.** What do you mean by Latent heat of Fusion?
- 47.** What is compressibility? How it is negligible in solids?
- 48.** Two cubes of ice are pressed hard between two palms. After releasing the pressure, the cubes join together. Why?
- 49.** Explain why ice has lower density than water?
- 50.** Give one similarity between a liquid and a gas and one dissimilarity.
- 51.** What property of gas is utilized when natural gas is supplied for vehicles.
- 52.** What are ‘intermolecular forces’? How are these related to the three states of matter ?
- 53.** Separate the following substances in groups of high and low intermolecular force: Ice, sulphur vapour, nitrogen, sugar, copper, air, salt, plastic.
- 54.** Which of the following substances you expect strongest and in which weakest intermolecular force: Water, alcohol, sugar, sodium chloride, carbon dioxide.
- 55.** Why are gases compressible but liquids not ?
- 56.** Compare the process of boiling and vaporization.
- 57.** How is pressure developed in a container full of a gas ?
- 58.** What are the applications of interconversion of states of matter ?
- 59.** What happens to a gas if its inter molecular space is reduced ?
- 60.** Explain how lumps of ice are more effective cooling than water at 273K.

ASSIGNMENT QUESTIONS SET – 3
CHAPTER – 1
MATTER IN OUR SURROUNDINGS

1. Conversion of solid to vapour is called sublimation. Name the term used to denote the conversion of vapour to solid.
2. Conversion of solid state to liquid state is called fusion; what is meant by latent heat of fusion?
3. Fill in the blanks:
 - (a) Evaporation of a liquid at room temperature leads to a_____ effect.
 - (b) At room temperature the forces of attraction between the particles of solid substances are_____than those which exist in the gaseous state.
 - (c) The arrangement of particles is less ordered in the _____ state. However, there is no order in the _____ state.
 - (d) _____ is the change of gaseous state directly to solid state without going through the _____ state.
 - (e) The phenomenon of change of a liquid into the gaseous state at any temperature below its boiling point is called_____.
4. Match the physical quantities given in column A to their S I units given in column B :

(A)	(B)
(a) Pressure	(i) cubic metre
(b) Temperature	(ii) kilogram
(c) Density	(iii) pascal
(d) Mass	(iv) kelvin
(e) Volume	(v) kilogram per cubic metre
5. The non S I and S I units of some physical quantities are given in column A and column B respectively. Match the units belonging to the same physical quantity:

(A)	(B)
(a) degree celsius	(i) kilogram
(b) centimetre	(ii) pascal
(c) gram per centimetre cube	(iii) metre
(d) bar	(iv) kelvin
(e) milligram	(v) kilogram per metre cube
6. Classify the following into osmosis/diffusion
 - (a) Swelling up of a raisin on keeping in water.
 - (b) Spreading of virus on sneezing.
 - (c) Earthworm dying on coming in contact with common salt.
 - (d) Shrinking of grapes kept in thick sugar syrup.
 - (e) Preserving pickles in salt.
 - (f) Spreading of smell of cake being baked through out the house.
 - (g) Aquatic animals using oxygen dissolved in water during respiration.

- 7.** Which one of the following sets of phenomena would increase on raising the temperature?
- (a) Diffusion, evaporation, compression of gases
 - (b) Evaporation, compression of gases, solubility
 - (c) Evaporation, diffusion, expansion of gases
 - (d) Evaporation, solubility, diffusion, compression of gases
- 8.** Seema visited a Natural Gas Compressing Unit and found that the gas can be liquefied under specific conditions of temperature and pressure. While sharing her experience with friends she got confused. Help her to identify the correct set of conditions
- (a) Low temperature, low pressure
 - (b) High temperature, low pressure
 - (c) Low temperature, high pressure
 - (d) High temperature, high pressure
- 9.** The property to flow is unique to fluids. Which one of the following statements is correct?
- (a) Only gases behave like fluids
 - (b) Gases and solids behave like fluids
 - (c) Gases and liquids behave like fluids
 - (d) Only liquids are fluids
- 10.** During summer, water kept in an earthen pot becomes cool because of the phenomenon of
- (a) diffusion
 - (b) transpiration
 - (c) osmosis
 - (d) evaporation
- 11.** A few substances are arranged in the increasing order of ‘forces of attraction’ between their particles. Which one of the following represents a correct arrangement?
- (a) Water, air, wind
 - (b) Air, sugar, oil
 - (c) Oxygen, water, sugar
 - (d) Salt, juice, air
- 12.** On converting 25°C , 38°C and 66°C to kelvin scale, the correct sequence of temperature will be
- (a) 298 K, 311 K and 339 K
 - (b) 298 K, 300 K and 338 K
 - (c) 273 K, 278 K and 543 K
 - (d) 298 K, 310 K and 338 K
- 13.** The boiling points of diethyl ether, acetone and *n*-butyl alcohol are 35°C , 56°C and 118°C respectively. Which one of the following correctly represents their boiling points in kelvin scale?
- (a) 306 K, 329 K, 391 K
 - (b) 308 K, 329 K, 392 K
 - (c) 308 K, 329 K, 391 K
 - (d) 329 K, 392 K, 308 K
- 14.** Which condition out of the following will increase the evaporation of water?
- (a) Increase in temperature of water
 - (b) Decrease in temperature of water
 - (c) Less exposed surface area of water
 - (d) Adding common salt to water

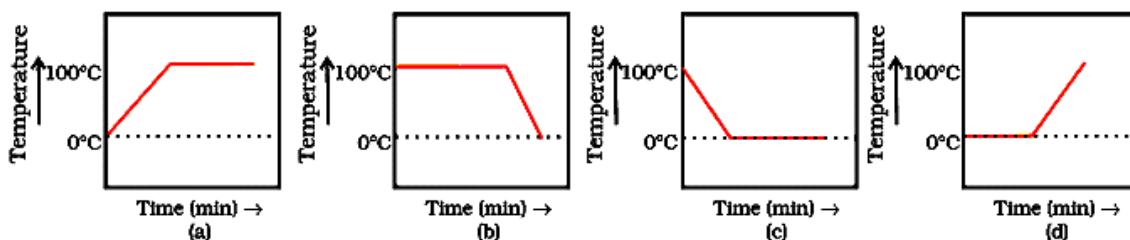
15. Choose the correct statement of the following

- (a) conversion of solid into vapours without passing through the liquid state is called vapourisation.
- (b) conversion of vapours into solid without passing through the liquid state is called sublimation.
- (c) conversion of vapours into solid without passing through the liquid state is called freezing.
- (d) conversion of solid into liquid is called sublimation.

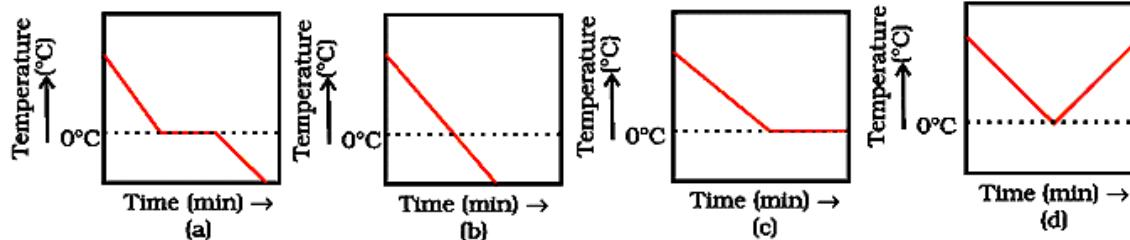
16. In which of the following conditions, the distance between the molecules of hydrogen gas would increase?

- (i) Increasing pressure on hydrogen contained in a closed container
 - (ii) Some hydrogen gas leaking out of the container
 - (iii) Increasing the volume of the container of hydrogen gas
 - (iv) Adding more hydrogen gas to the container without increasing the volume of the container
- (a) (i) and (iii)
 - (b) (i) and (iv)
 - (c) (ii) and (iii)
 - (d) (ii) and (iv)

17. A student heats a beaker containing ice and water. He measures the temperature of the content of the beaker as a function of time. Which of the following graph would correctly represent the result? Justify your choice.



18. A glass tumbler containing hot water is kept in the freezer compartment of a refrigerator ($\text{temperature} < 0^\circ\text{C}$). If you could measure the temperature of the content of the tumbler, which of the following graphs would correctly represent the change in its temperature as a function of time.

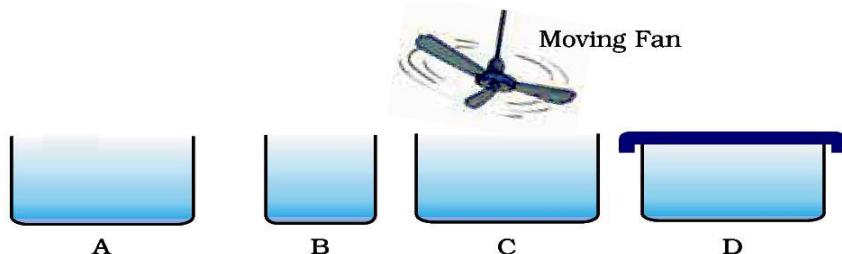


19. A sample of water under study was found to boil at 102°C at normal temperature and pressure. Is the water pure? Will this water freeze at 0°C ? Comment.

20. ‘Osmosis is a special kind of diffusion’. Comment.

21. Water as ice has a cooling effect, whereas water as steam may cause severe burns. Explain these observations.

22. Alka was making tea in a kettle. Suddenly she felt intense heat from the puff of steam gushing out of the spout of the kettle. She wondered whether the temperature of the steam was higher than that of the water boiling in the kettle. Comment.
23. Look at below figure and suggest in which of the vessels A,B, C or D the rate of evaporation will be the highest? Explain.



24. You are provided with a mixture of naphthalene and ammonium chloride by your teacher. Suggest an activity to separate them with well labelled diagram.
25. It is a hot summer day, Priyanshi and Ali are wearing cotton and nylon clothes respectively. Who do you think would be more comfortable and why?
26. You want to wear your favourite shirt to a party, but the problem is that it is still wet after a wash. What steps would you take to dry it faster?
27. Comment: Evaporation produces cooling.
28. Comment: Rate of evaporation of an aqueous solution decreases with increase in humidity.
29. Comment: Sponge though compressible is a solid.
30. Why does the temperature of a substance remain constant during its melting point or boiling point?
-

CHAPTER – 2

IS MATTER AROUND US PURE

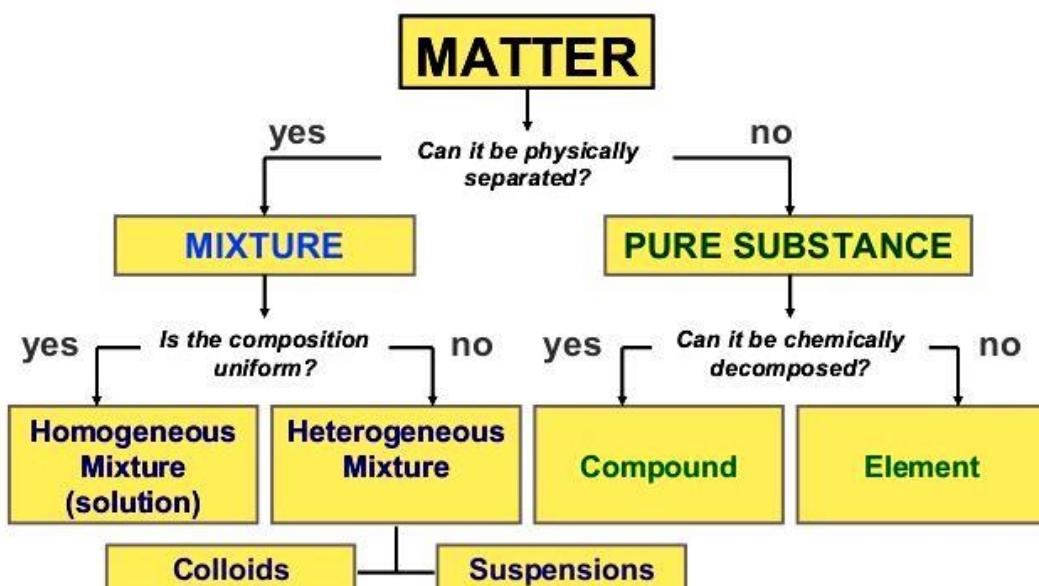
Matters can be classified into two types – Pure substances and Impure substances.

Pure substances – Pure substances are of two types – Elements and Compounds.

Impure substances – All mixture are considered as impure substances.

Colour, odour, density, melting point and boiling point are often treated as physical properties of matter. The physical properties of a substance can be observed or measured without changing its composition.

The term “impure” is different from adulteration. According to scientists, the term “pure” means single form of matter.



Most of the substances in our surroundings are not in their pure form and are called mixture. Substances which are made of two or more matters and which can be separated by physical methods are known as mixtures, such as mixture of salt and water, mixture of sugar and water, mixture of different gases, air, etc.

In a mixture, components do not combine chemically or through any chemical change. In a mixture, components do not lose their properties.

Mixtures are of two types on the basis of their composition - Homogeneous mixture and Heterogeneous mixture.

HOMOGENEOUS MIXTURE

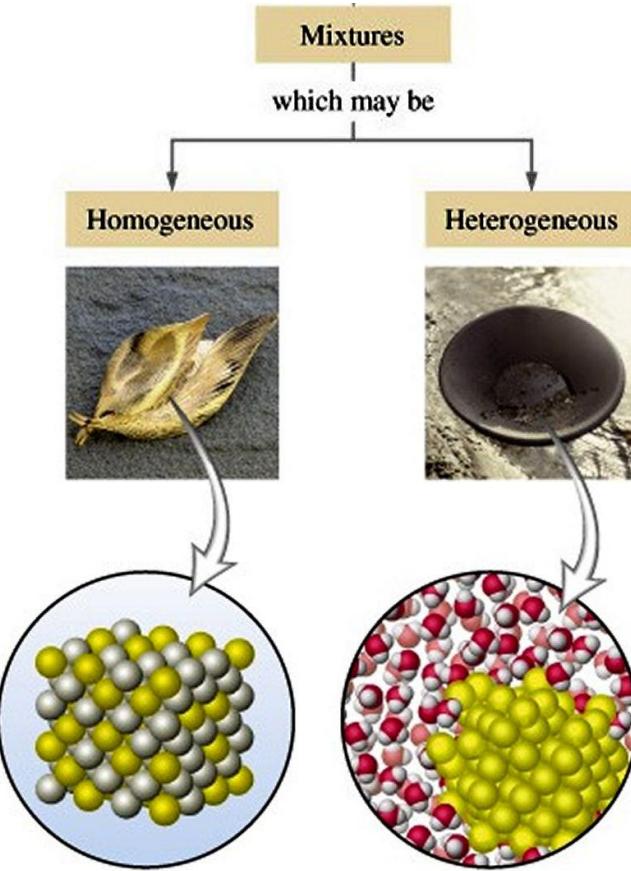
Mixtures which have uniform composition throughout are called Homogeneous Mixture. For example – mixture of salt and water, mixture of sugar and water, air, lemonade, soda water, etc.

Mixture of salt in water is an example of homogeneous mixture. In this mixture, the boundary of salt and water cannot be differentiated. When a ray of light is passed through the mixture of salt and water, the path of light is not seen.

General Properties of Homogeneous Mixture:

- All solutions are the examples of homogeneous mixture.
- The particles of a homogeneous mixture are less than one nanometer.

- A homogenous mixture does not show Tyndall effect.
- The boundaries of particles cannot be differentiated.
- The constituent particles of homogenous mixture cannot be separated using centrifugation or decantation.
- Alloys are the examples of solution.



HETEROGENEOUS MIXTURE

Mixtures which do not have uniform composition throughout are called Heterogeneous Mixture. For example – mixture of soil and sand, mixture of sulphur and iron fillings, mixture of oil and water etc. The boundaries of constituent particles of a homogeneous mixture can be identified easily; as a homogeneous mixture has two or more distinct phases.

General Properties of Heterogeneous Mixture:

- Most of the mixtures are heterogeneous except solutions and alloys.
- The constituent particles are present uniformly in a heterogeneous mixture.
- The components of a heterogeneous mixture can be identified easily.
- Generally, two or more phases are present in a heterogeneous mixture.
- Particles of a heterogeneous mixture are sized between one nanometer and one micrometer or more.
- Heterogeneous mixtures show Tyndall effect.

INTEXT QUESTIONS PAGE NO. 15

Q1. What is meant by a pure substance?

Answer: A pure substance is the one that consists of a single type of particles, i.e., all constituent particles of the substance have the same chemical nature. Pure substances can be classified as elements or compounds.

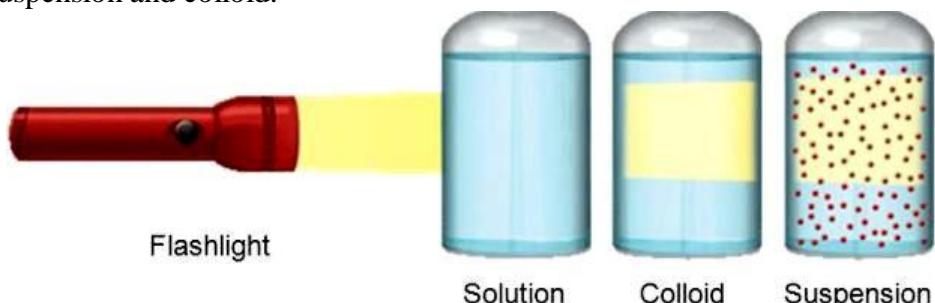
Q2. List the points of differences between homogeneous and heterogeneous mixtures.

Answer:

- A homogeneous mixture is a mixture having a uniform composition throughout the mixture. For example: salt in water, sugar in water, copper sulphate in water
- A heterogeneous mixture is a mixture having a non-uniform composition throughout the mixture. For example: sodium chloride and iron filings, salt and sulphur, oil and water

TYPES OF MIXTURE

Mixture can be categorised in three types on the basis of their particles' size. These are; solution, suspension and colloid.



SOLUTION

Mixture of two or more substances with one phase only, i.e. having no distinct boundary of constituent particles are called solution. For example, solution of sugar and water, solution of salt and water, lemonade, soft drinks, etc. Solution is a homogeneous mixture of two or more substances.

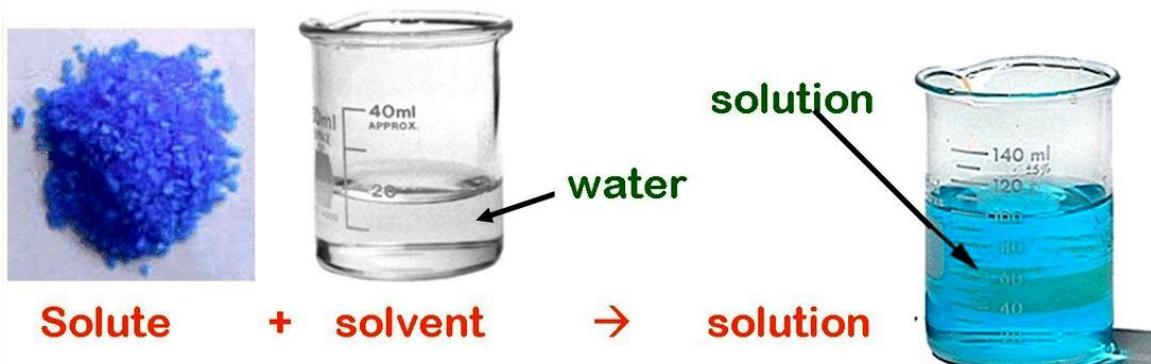
In a solution, components are mixed in such a way that they appear as only one phase. Seeing by naked eye, constituent particles of a solution cannot be identified because particles are mixed evenly throughout.

In a solution, there are two types of components – one is called solute and other is called solvent.

Solute – Substance which is present in smaller quantity in a mixture is called solute.

Solvent – Substance in a mixture which is present in larger quantity in a mixture is called solvent.

Example: In the solution of salt and water, salt is present in small quantity while water is present in larger quantity. Here salt is solute and water is solvent.



Solution of tincture iodine is made by dissolving iodine in alcohol. In this solution, iodine is solute and alcohol is solvent.

Air is mixture of many gases. Since air consists of only one phase, thus it is considered as solution. Air consists about 78% of nitrogen and 22% rest of other gases. Thus in the case of air, nitrogen can be called as solvent and rest other gases as solutes. Solvent and solute can be solid, liquid or gas.

TYPES OF SOLUTION

Solid - solid solution – Solution of two or more solids are generally known as solid-solid solution. For example – alloys. Alloy is a homogeneous mixture of two or more metals and non metals or two metals or two non-metals. The components of an alloy cannot be separated by physical methods, their boundaries are not distinct and they can have variable compositions, thus alloy is considered as solution.

Solid – Liquid solution – Solution of solid and liquid is called solid-liquid solution. For example - solution of salt and water.

Liquid – liquid solution – Solution of two miscible liquids are called liquid-liquid solution, such as solution of water and acetic acid. The solution of acetic acid in water is known as vinegar.

Gas - liquid solution – Solution of gas into liquid is called gas-liquid solution. For example – Soft drink. In soft drink, carbon dioxide is usually dissolved in liquid, because of which a hiss sound comes while opening the cap of the bottle.

Gas-gas solution – Solution of two or more gas is called gas-gas solution. For example – air, which is the solution of many gases, such as hydrogen, oxygen, carbon dioxide, etc.

State of Solvent	State of Solute	State of Solution	Examples
Gas	Gas	Gas	Air, natural gas
Liquid	Liquid	Liquid	Alcoholic beverages, Antifreeze solution;
Liquid	Solid	Liquid	Seawater, sugar solution
Liquid	Gas	Liquid	Carbonated water (soda) Ammonia solution;
Solid	Solid	Solid	Metal alloys: brass, bronze,..
Solid	Gas	Solid	Hydrogen in platinum

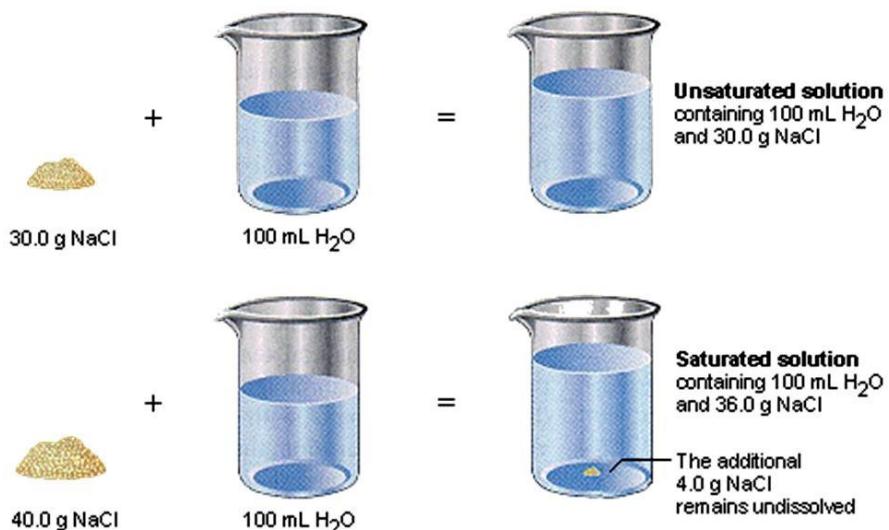
General Characteristics of Solution:

- Solutions are homogeneous mixture of two or more substances.
- Constituent particles of a solution are mixed evenly throughout.
- There is only one phase in a solution.
- Boundaries of constituent particles cannot be distinguished.
- The size of particles of solution is smaller than one nanometer.
- Solutions do not show Tyndall effect, because, small particles of solution do not scatter the ray of light.
- Solute cannot be separated by using filtration or decantation.
- Solutions are stable, since when left undisturbed the particles do not settle in bottom.

SATURATED AND UNSATURATED SOLUTIONS

Saturated Solution: When a solution cannot dissolve more solute at a given temperature, the point is called saturation point of the solution and solution is called saturated solution. This means, no more solute can be dissolved in a saturated solution at a given temperature.

Unsaturated Solution: Solution in which more solution can dissolved at a given temperature, is called unsaturated solution.



Solubility: Solubility is the amount of solute in a saturated solution at a given temperature. In other words, maximum capacity to dissolve a solute in a solution at a given temperature is called solubility.

Different solvents can dissolve different amount of solute. This means different solvents have different solubility. Solubility increases with increase in temperature.

Concentration: Concentration is the amount of solute present in a given amount of solvent or solution.

$$\text{This means Concentration} = \frac{\text{Amount of solute}}{\text{Amount of Solvent}} \text{ or } \frac{\text{Amount of solute}}{\text{Amount of Solution}}$$

Thus, concentration is the ratio of amount of solute and amount of solvent.

Concentration can be expressed in mass percentage or volume percentage of a solution.

(a) Mass percentage of a solution

$$\text{Concentration} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

This means when concentration is expressed in mass percentage, it is called concentration by mass percentage.

(b) Volume percentage of a solution

$$\text{Concentration} = \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100$$

This means when concentration is expressed in volume percentage, it is called concentration by volume percentage.

SUSPENSION

A suspension is a heterogeneous mixture in which the solute particles do not dissolve but remain suspended throughout the bulk of the medium. Particles of a suspension are visible to the naked eye.

Properties of a Suspension

- Suspension is a heterogeneous mixture.
- The particles of a suspension can be seen by the naked eye.
- The particles of a suspension scatter a beam of light passing through it and make its path visible.
- The solute particles settle down when a suspension is left undisturbed, that is, a suspension is unstable. They can be separated from the mixture by the process of filtration.

COLLOIDAL SOLUTIONS

A colloidal solution, occasionally identified as a colloidal suspension, is a mixture in which a substances regularly suspended in a fluid. A colloid is a minutely small material that is regularly spread out all through another substance.

Properties of a colloid

- A colloid is a heterogeneous mixture.
- The size of particles of a colloid is too small to be individually seen by naked eyes.
- Colloids are big enough to scatter a beam of light passing through it and make its path visible.
- They do not settle down when left undisturbed, that is, a colloid is quite stable.
- They cannot be separated from the mixture by the process of filtration.

The components of a colloidal solution are the dispersed phase and the dispersion medium. The solute-like component or the dispersed particles in a colloid form the dispersed phase, and the component in which the dispersed phase is suspended is known as the dispersing medium. Colloids are classified according to the state (solid, liquid or gas) of the dispersing medium and the dispersed phase.

Colloidal solutions have three sub-classifications: **Foams, emulsions and sol.** **Foam** in this setting is created by ensnaring a gas in a liquid. The substance being dispersed would be the gas, triggering the fluid to become frothy and foamy. A sample of this would be shaving cream. **An emulsion** is a combination of liquids; it's basically when one liquid is consistently dispersed all through another liquid. A sample of this would be mayonnaise or milk. The third form is called **a sol**, which is when a solid is evenly dispersed throughout a fluid. Samples of sols include paint, blood and silver aquasols.

Common examples of colloids

Dispersed phase	Dispersing Medium	Type	Example
Liquid	Gas	Aerosol	Fog, clouds, mist
Solid	Gas	Aerosol	Smoke, automobile exhaust
Gas	Liquid	Foam	Shaving cream
Liquid	Liquid	Emulsion	Milk, face cream
Solid	Liquid	Sol	Milk of magnesia, mud
Gas	Solid	Foam	Foam, rubber, sponge, pumice
Liquid	Solid	Gel	Jelly, cheese, butter
Solid	Solid	Solid Sol	Coloured gemstone, milky glass

INTEXT QUESTIONS PAGE NO. 18

Q1. Differentiate between homogeneous and heterogeneous mixtures with examples.

Answer:

- A homogeneous mixture is a mixture having a uniform composition throughout the mixture. For example, mixtures of salt in water, sugar in water, copper sulphate in water, iodine in alcohol, alloy, and air have uniform compositions throughout the mixtures.
- On the other hand, a heterogeneous mixture is a mixture having a non-uniform composition throughout the mixture. For example, composition of mixtures of sodium chloride and iron filings, salt and sulphur, oil and water, chalk powder in water, wheat flour in water, milk and water are not uniform throughout the mixtures.

Q2. How are sol, solution and suspension different from each other?

Answer:

- **Sol** is a heterogeneous mixture. In this mixture, the solute particles are so small that they cannot be seen with the naked eye. Also, they seem to be spread uniformly throughout the mixture. The Tyndall effect is observed in this mixture. For example: milk of magnesia, mud
- **Solution** is a homogeneous mixture. In this mixture, the solute particles dissolve and spread uniformly throughout the mixture. The Tyndall effect is not observed in this mixture. For example: salt in water, sugar in water, iodine in alcohol, alloy
- **Suspensions** are heterogeneous mixtures. In this mixture, the solute particles are visible to the naked eye, and remain suspended throughout the bulk of the medium. The Tyndall effect is observed in this mixture. For example: chalk powder and water, wheat flour and water

Q3. To make a saturated solution, 36 g of sodium chloride is dissolved in 100 g of water at 293 K. Find its concentration at this temperature.

Answer:

Mass of solute (sodium chloride) = 36 g (Given)

Mass of solvent (water) = 100 g (Given)

Then, mass of solution = Mass of solute + Mass of solvent = (36 + 100) g = 136 g

Therefore, concentration (mass by mass percentage) of the solution

$$= \frac{\text{Mass of solute}}{\text{Mass of solvent}} \times 100\% = \frac{36}{136} \times 100\% = 26.47\%$$

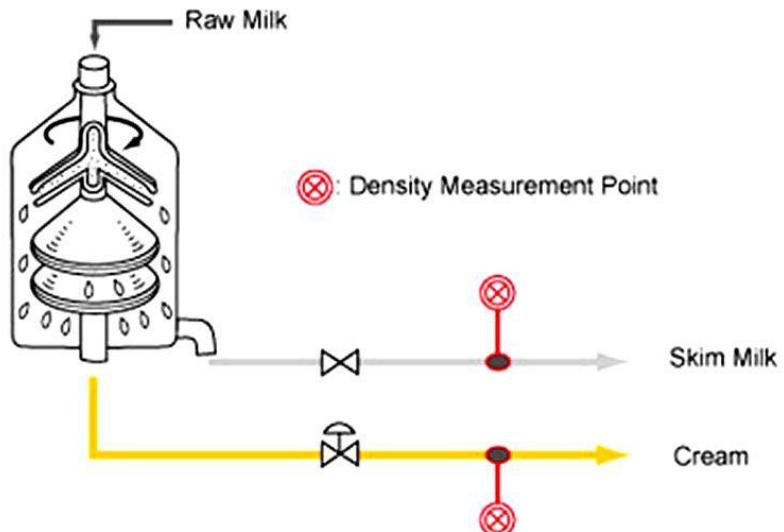
SEPARATION OF COMPONENTS OF MIXTURE

CENTRIFUGATION

In the method of centrifugation, the centripetal and centrifugal forces are used to separate lighter and heavier components of mixture of two immiscible liquids. This process is used to separate very small solids particles from a liquid mixture.

Example – Milk is the mixture of fat, water, and other constituents. Using the method of centrifugation, most of the fat can be separated from milk. In milk, fat is suspended throughout the milk which is separated out using the method of centrifugation.

When milk is churned rapidly, water which is heavier than fat, migrates away from the centre of centrifuge while fat is forced towards the bottom, which is drained out.



Application of centrifugation –

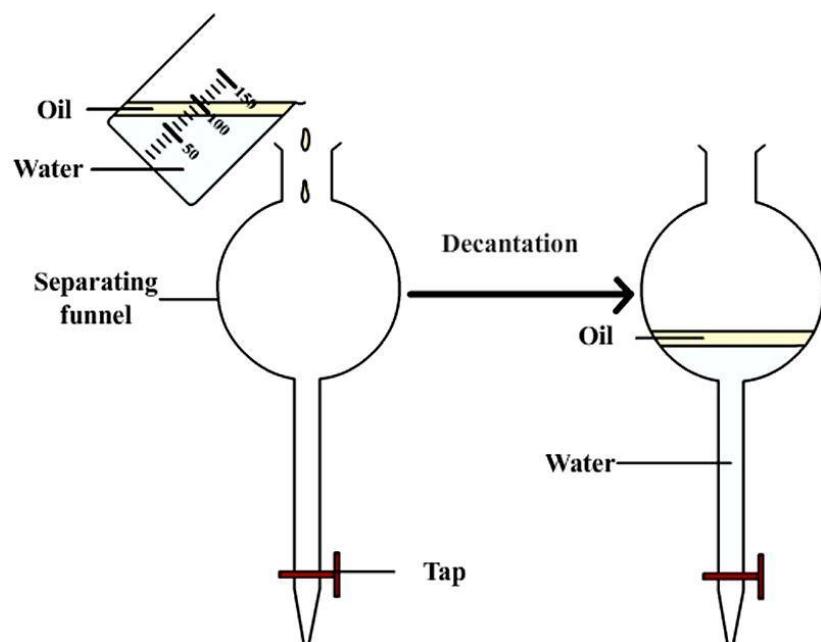
- In pathological test of blood and urine.
- In separation of fat from milk.
- In washing machines to squeeze the water from wet clothes.

DECANTATION

Decantation is used to separate the components from a mixture of two immiscible liquids, such as mixture of oil and water. In a mixture of two immiscible liquids, lighter one and heavier one form separate layer. The lighter one can be decanted after settling of mixture, carefully in other container.

In the process of decantation some of the heavier liquid also poured out with lighter one. Therefore, components from a mixture of two immiscible liquids; can be separated more easily and accurately using a separating funnel.

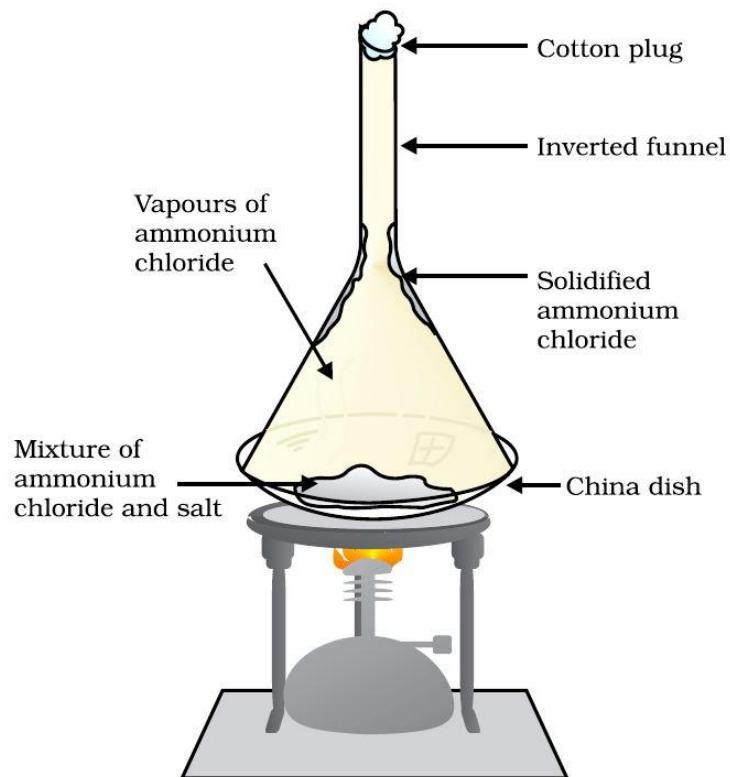
A separating funnel is usually made of glass with a stop cork with drain pipe at bottom. The heavier liquid which is settled at bottom is drained out from the mixture of two immiscible liquids by opening of stop cork from a separating funnel.



SUBLIMATION

There are many substances which are converted into gas from solid when heated, and converted from gas to solid when cooled without converting into liquid. Such substances are known as sublimate. For example – ammonium chloride, naphthalene balls, camphor, etc. Therefore, mixture of one sublimate and other substance can be separated using the method of sublimation.

The mixture of ammonium chloride and common salt can be separated out using the process of sublimation. For this, the mixture is heated in a China dish. The China dish is covered by an inverted funnel. Cotton is used for plugging the opening of the funnel. After heating, ammonium chloride is converted into vapour and gets deposited over the inner surface of funnel; due to cooling. This leaves the common salt in China dish. Ammonium chloride can be taken out by scratching from the inner wall of funnel.

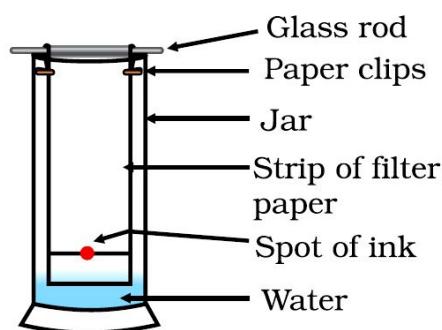
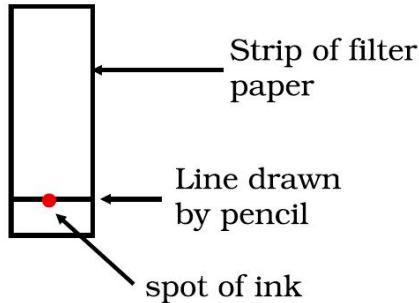


CHROMATOGRAPHY

Chromatography is a method of separation which works on the principle of travel speed of components of a mixture. This method is used for separating dyes and pigments from a mixture. Ink is the mixture of dyes of different colours.

There are many types of chromatography. The dyes from an ink can be separated using paper chromatography.

For this, a strip of filter paper is dipped in the ink. Particles of dye start rising on filter paper; along with water. Different dyes rise with different speed because of different types of solubility in water and go up to certain heights.

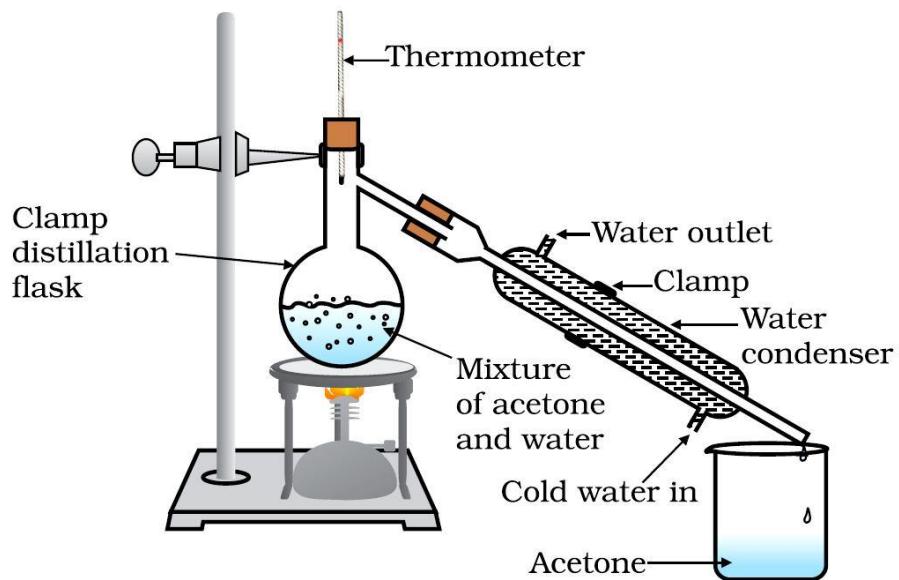


Application of chromatography -

- In the separation of colours from a dyes.
- In the separation of pigments from natural colours.
- In the separation of drugs from blood for pathological tests.

DISTILLATION

The process of distillation is used to separate two miscible liquids. The technique of distillation is based on the difference in boiling points of components of mixture of miscible liquids. Distillation is to separate the liquids which do not decompose even upto their boiling points and should boil at more than 25°C .

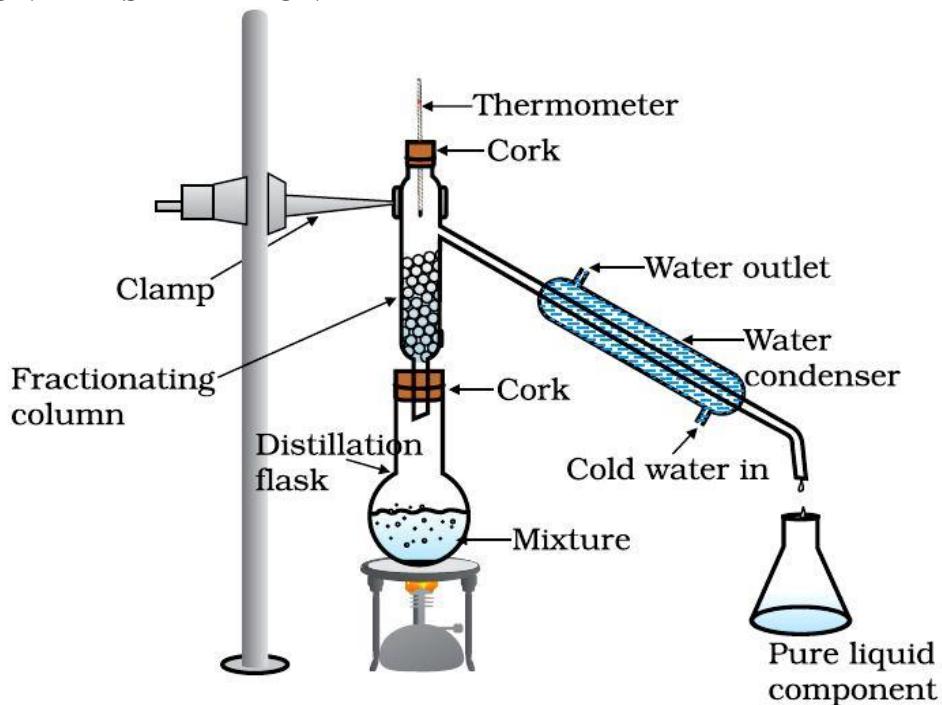


In the process of distillation, the mixture is heated after keeping in a retort or distillation flask. The liquid which boils at lower temperature is vaporized at lower temperature. The vapour so obtained is passed through a tube and gets condensed in a separate container; leaving liquid with higher boiling point in the retort or distillation flask.

Distillation is used to separate the components of the mixture of two miscible liquids that boils without decomposition and have sufficient difference in their boiling points.

The process of distillation is used to purify many liquids, such as water.

FRACTIONAL DISTILLATION



Fractional distillation is the process of separation of components of mixture into parts or fraction on the basis of fractional differences in their boiling points.

Fractional distillation is done when the difference in boiling points of the components of miscible liquids is less than 25°C . In the process of fractional distillation, a fractional column is used along with retort or distillation flask.

Fractional column is a tube which contains glass beads, which facilitate surface for the vapour to cool and condense repeatedly.

Example – Ethanol and water are separated from their mixture using fractional distillation. The boiling point of water is 100°C while the boiling point of ethanol is 78.4°C . Since the difference of their boiling point is less than 25°C , thus they are separated using fractional distillation.

Some of the Applications of Fractional Distillation:

- In petroleum refineries, petrochemical and chemical plants, natural gas processing and cryogenic air separation plants.
- In oil refineries to separate crude oil into useful substances (or fractions).
- In the process of organic juice.
- In the separation of oxygen, liquid nitrogen and argon from air.

SEPARATION OF DIFFERENT GASES FROM AIR

Air comprises of nitrogen, oxygen, carbon dioxide and argon as major components. Since air is the cheapest source of these gases, thus these are extracted from air at large scale

After liquefaction of air by repeated compression and cooling; nitrogen, oxygen, carbon dioxide and argon are extracted using fractional distillation.

Liquid nitrogen has boiling point equal to -190°C and thus turns into gas first and separated from air.

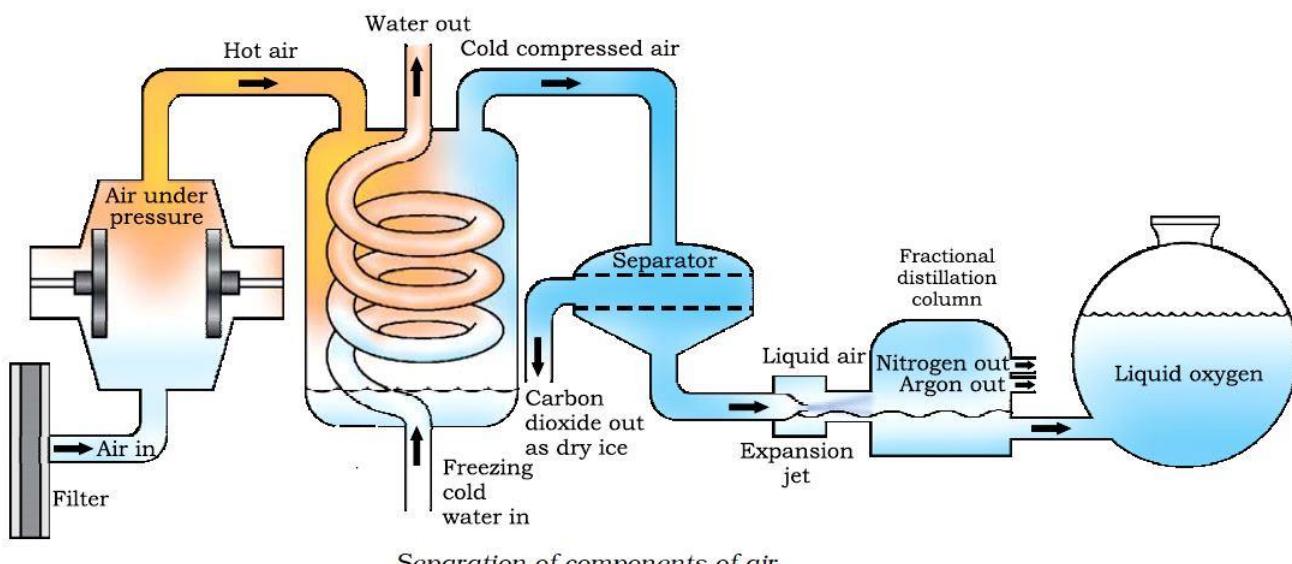
The boiling point of argon is -186°C , therefore it is extracted after argon.

The boiling point of oxygen is -183°C , thus it is collected after the extraction of argon.

Carbon dioxide turns into solid at a temperature of -97°C , therefore, it is removed while air is put under liquefaction.

USE

Nitrogen is used as fertilizers, oxygen is used in hospitals and argon is used in bulbs.



CRYSTALLIZATION

Crystallisation is a process that separates a pure solid in the form of its crystals from a solution. The crystallisation method is used to purify solids. For example, the salt we get from sea water can have many impurities in it. To remove these impurities, the process of crystallisation is used.

Crystallisation technique is better than simple evaporation technique as –

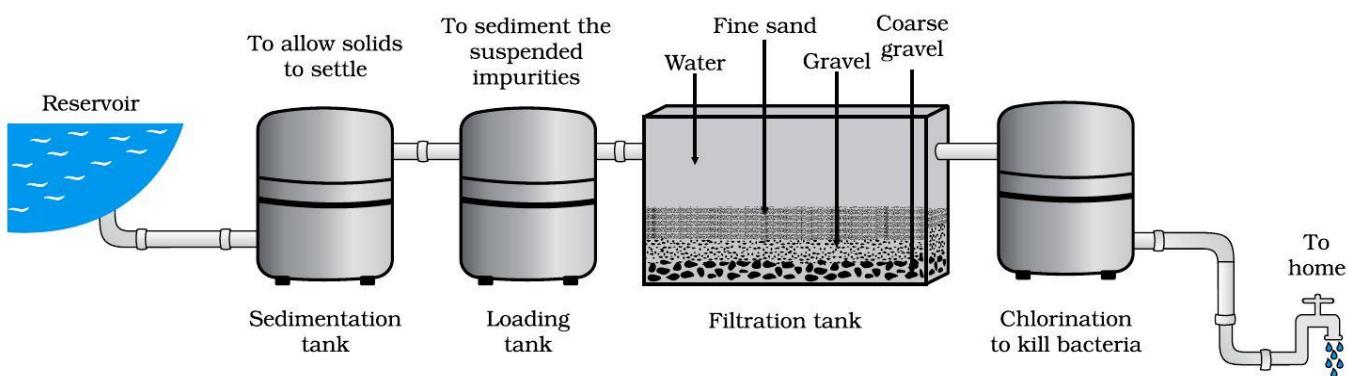
- some solids decompose or some, like sugar, may get charred on heating to dryness.
- some impurities may remain dissolved in the solution even after filtration. On evaporation these contaminate the solid.

APPLICATIONS

- Purification of salt that we get from sea water.
- Separation of crystals of alum (*phitkari*) from impure samples.

Thus, by choosing one of the above methods according to the nature of the components of a mixture, we get a pure substance. With advancements in technology many more methods of separation techniques have been devised.

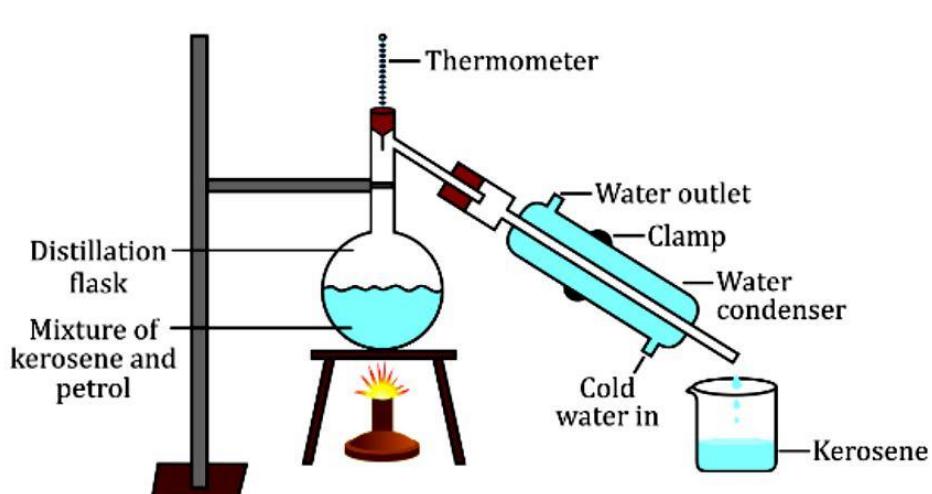
In cities, drinking water is supplied from water works. A flow diagram of a typical water works is shown in below figure. From this figure write down the processes involved to get the supply of drinking water to your home from the water works and discuss it in your class.



INTEXT QUESTIONS PAGE NO. 24

Q1. How will you separate a mixture containing kerosene and petrol (difference in their boiling points is more than 25°C), which are miscible with each other?

Answer:



A mixture of two miscible liquids having a difference in their boiling points more than 25°C can be separated by the method of distillation. Thus, kerosene and petrol can be separated by distillation.

In this method, the mixture of kerosene and petrol is taken in a distillation flask with a thermometer fitted in it. We also need a beaker, a water condenser, and a Bunsen burner. The apparatus is arranged as shown in the above figure. Then, the mixture is heated slowly. The thermometer should be watched simultaneously. Kerosene will vaporize and condense in the water condenser. The condensed kerosene is collected from the condenser outlet, whereas petrol is left behind in the distillation flask.

Q2. Name the technique to separate

- (i) butter from curd,
- (ii) salt from sea-water,
- (iii) camphor from salt.

Answer:

- (i) Butter can be separated from curd by centrifugation.
- (ii) Salt can be separated from sea-water by evaporation.
- (iii) Camphor can be separated from salt by sublimation.

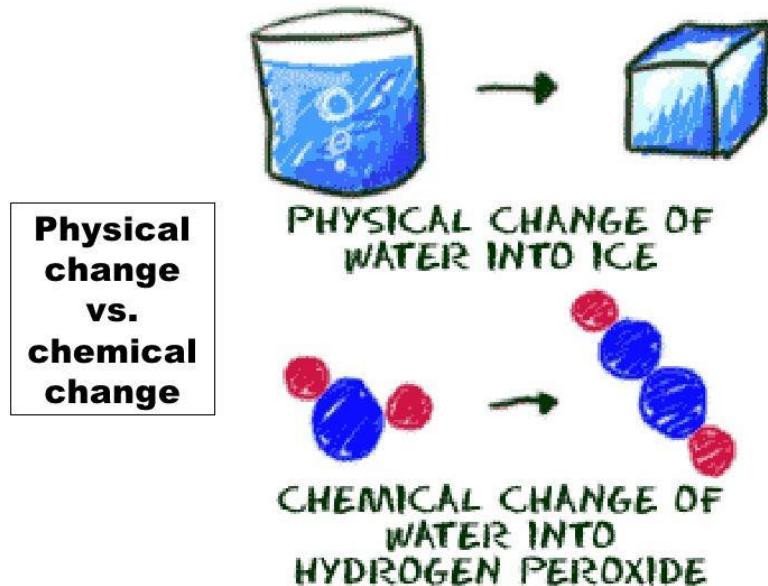
Q3. What type of mixtures are separated by the technique of crystallisation?

Answer:

By the technique of crystallization, pure solids are separated from impurities. For example, salt obtained from sea is separated from impurities; crystals of alum (Phitkari) are separated from impure samples.

PHYSICAL CHANGE: The change in which no new substance is formed is called a physical change. During a physical change, chemical properties do not change but physical properties do change.

CHEMICAL CHANGE: The change in which a new substance is formed is called a chemical change. During a chemical change, chemical properties change.



PURE SUBSTANCES

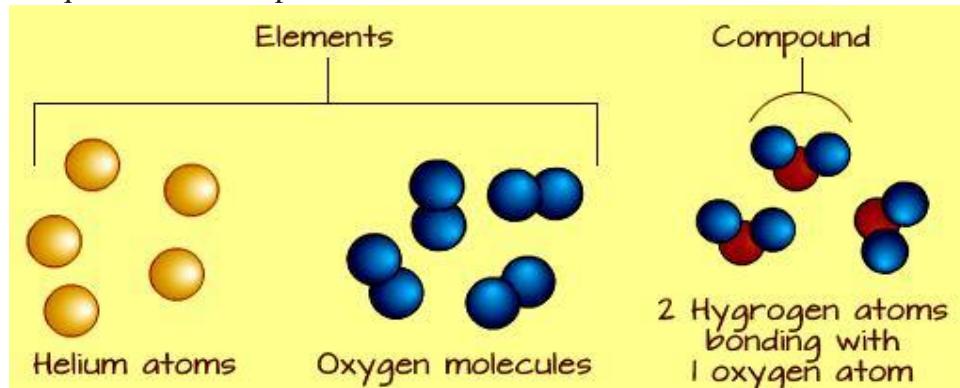
Elements and Compounds are considered as pure substances.

Elements – Substances that is made of only one element are called elements, such as hydrogen, carbon, oxygen, silver, gold, etc.

Elements can be normally divided into metals, non-metals and metalloids. Metals usually show some or all of the following properties:

- They have a lustre (shine).
- They have silvery-grey or golden-yellow colour.
- They conduct heat and electricity.
- They are ductile (can be drawn into wires).
- They are malleable (can be hammered into thin sheets).
- They are sonorous (make a ringing sound when hit).

Examples of metals are gold, silver, copper, iron, sodium, potassium etc. Mercury is the only metal that is liquid at room temperature.



Non-metals usually show some or all of the following properties:

- They display a variety of colours.
- They are poor conductors of heat and electricity.
- They are not lustrous, sonorous or malleable.

Examples of non-metals are hydrogen, oxygen, iodine, carbon (coal, coke), bromine, chlorine etc. Some elements have intermediate properties between those of metals and non-metals, they are called metalloids; examples are boron, silicon, germanium etc.

Compounds – Substances that is made of one or more elements by chemical combination are called compounds, such as water, carbon dioxide, copper oxide, hydrochloric acid, etc.

A compound does not contain the properties of its constituent elements and shows quite different characteristics.

COMPARISON BETWEEN COMPOUND AND ELEMENT

	Compound	Element
Distinguishing Feature	Compounds contain different elements in a fixed ratio arranged in a defined manner through chemical bonds.	Elements are distinguished by their atomic number (number of protons in their nucleus).
Ability to Breakdown	A compound can be separated into simpler substances by chemical methods/reactions.	Elements cannot be broken down into simpler substances by chemical reactions.
Types	The list of compounds is endless.	There are about 117 elements that have been observed. Can be classified as metal, non-metal or metalloid.
Representation	A compound is represented using a formula.	An element is represented using symbols.
Examples	Water (H_2O), Sodium chloride ($NaCl$), Sodium bicarbonate ($NaHCO_3$) etc.	Iron, copper, silver, gold, nickel etc.

COMPARISON BETWEEN MIXTURE AND COMPOUND

Mixture	Compound
Elements are physically mixed in any ratio and no new compound is formed.	Elements are chemically combined in a fixed ratio to form a new compound.
They have no sharp or definite melting point, boiling point, density etc.	They have definite melting point, boiling point, density etc.
A mixture exhibits the properties of its constituent or component elements.	Property of a compound is different from its constituent or component elements.
They are either homogeneous or heterogeneous in nature.	They are always homogeneous in nature.
Constituents of a mixture can be separated by physical methods like filtration, magnetic separation etc.	Constituents of a compound cannot be separated by physical methods.

INTEXT QUESTIONS PAGE NO. 24

Q1. Classify the following as chemical or physical changes:

- cutting of trees,
- melting of butter in a pan,
- rusting of almirah,
- boiling of water to form steam,
- passing of electric current, through water and the water breaking down into hydrogen and oxygen gases,
- dissolving common salt in water,
- making a fruit salad with raw fruits, and
- burning of paper and wood.

Answer:

- Cutting of trees → Physical change
- Melting of butter in a pan → Physical change
- Rusting of almirah → Chemical change
- Boiling of water to form steam → Physical change
- Passing of electric current through water, and water breaking down into hydrogen and oxygen gas → Chemical change
- Dissolving common salt in water → Physical change
- Making a fruit salad with raw fruits → Physical change
- Burning of paper and wood → Chemical change

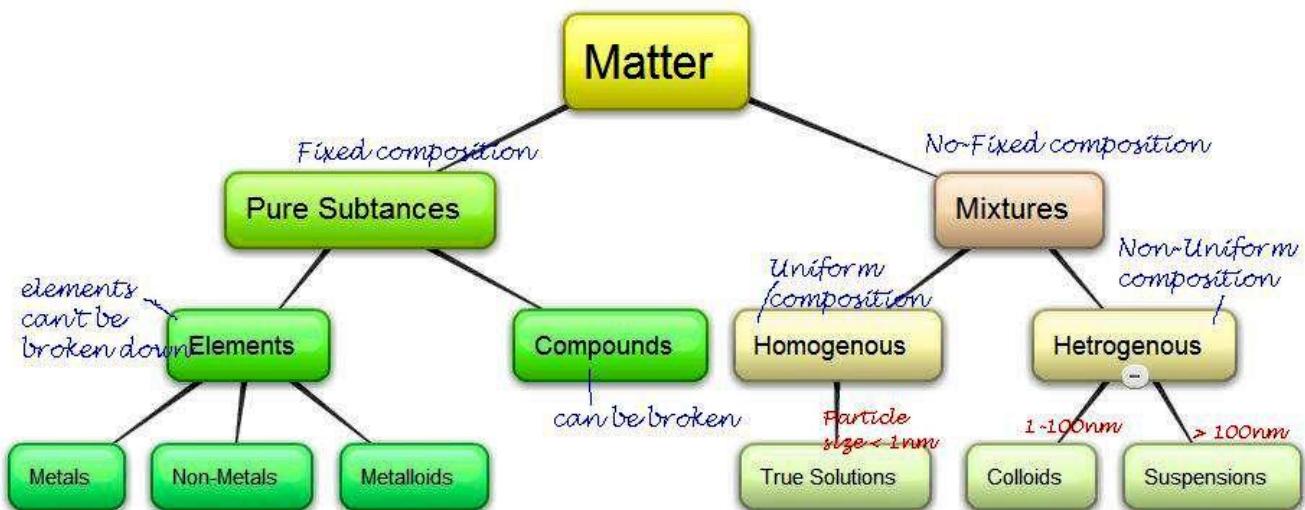
Q2. Try segregating the things around you as pure substances or mixtures.

Answer:

Pure substance: Water, salt, sugar

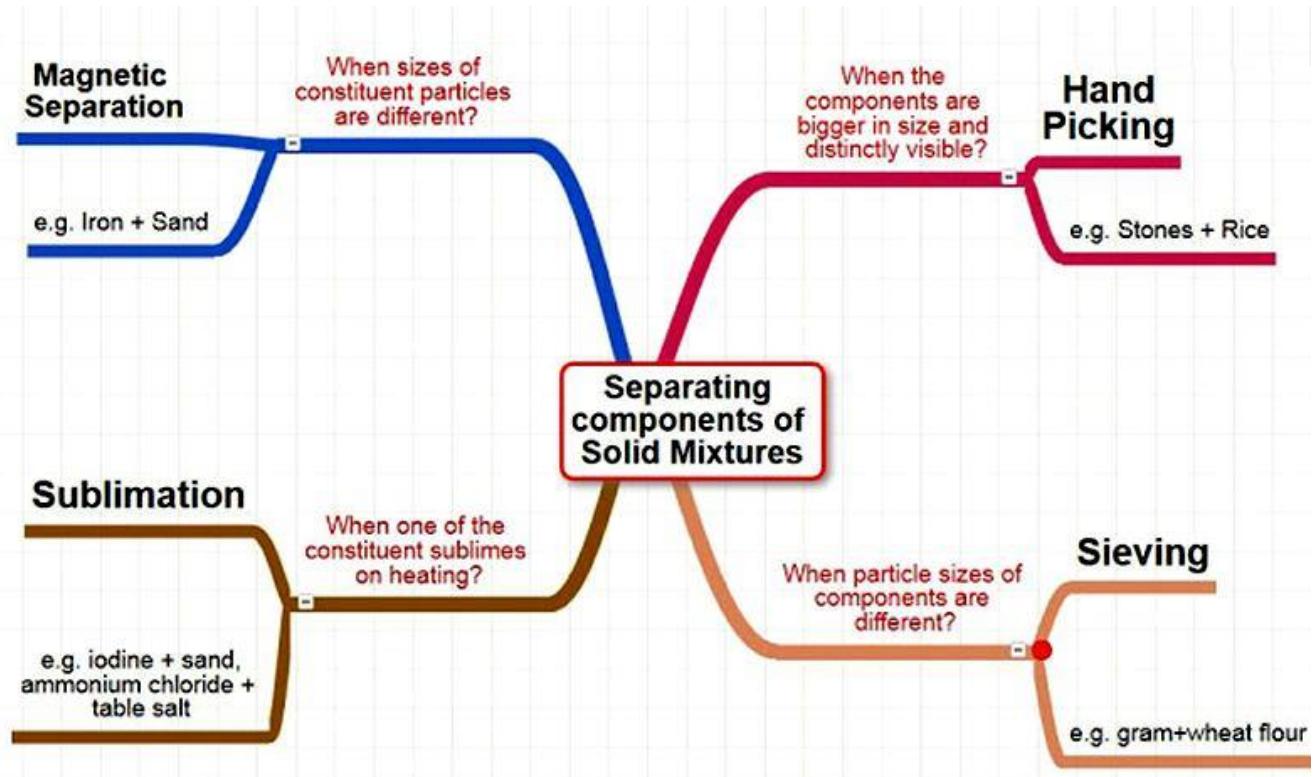
Mixture: Salt water, soil, wood, air, cold drink, rubber, sponge, fog, milk, butter, clothes, food

IMPORTANT CONCEPT MAPS



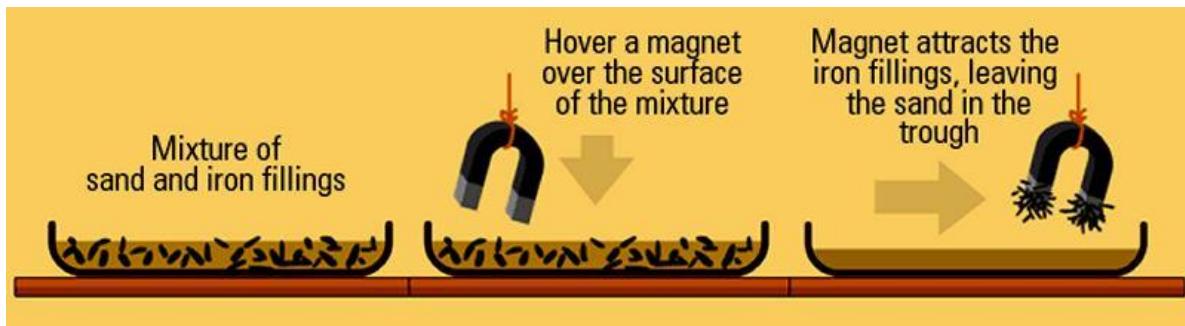
DIFFERENT WAYS TO SEPARATE SOLID MIXTURES

- Hand Picking
- Sieving
- Magnetic Separation Method
- Sublimation



MAGNETIC SEPARATION METHOD

Magnetic Separation Method is ideal for separating mixtures of two solids with one part having **magnetic** properties. Some metals like iron, nickel and cobalt have magnetic properties while gold, silver and aluminum do not. Magnetic elements are attracted to a magnet.



It works like this: Let us take a mixture of sand and iron filing for example.

To separate this, spread out the mixture on a flat surface. Run a magnet bar over the surface. You will notice that the magnetic elements (iron filings) will be attracted to the magnet over it. After a number of runs, all the sand will be free from any iron filing.

SIEVING METHOD

When the sizes of the components of a mixture are big enough, they can be separated with the help of sieve. A sieve is a simple mechanical device in which a mesh is attached to a frame. When the mixture is placed on the mesh and is stirred, particles of smaller size pass through the mesh while the bigger particles of the other component remain above the mesh. E.g. gram can be separated wheat, sieving of sand at construction site etc.

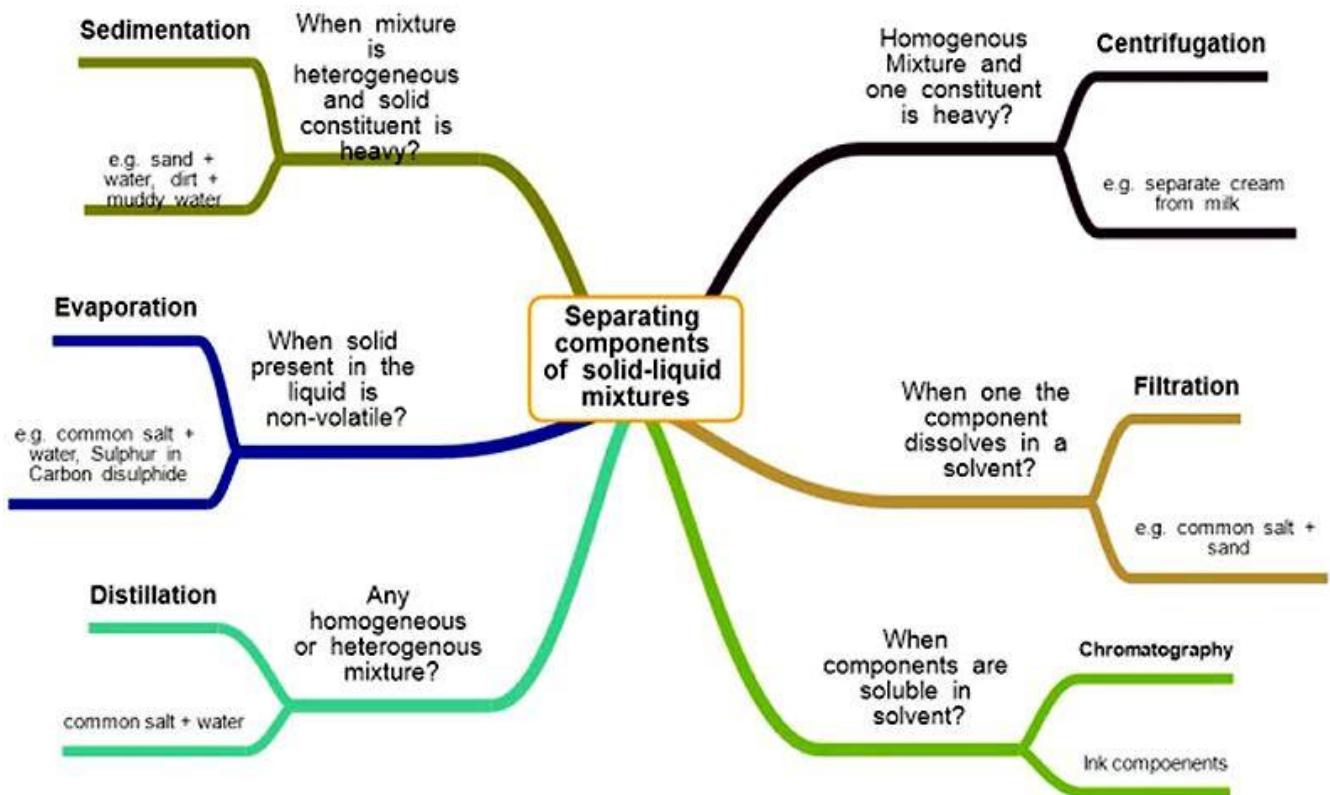


HAND PICKING METHOD

It involves simply picking out substances by hand and separating them from others. The substances being separated may be impurities that have to be thrown away or it may be that both the substances being separated are useful – such as if you separate green grapes from black ones from a mixture of the two.

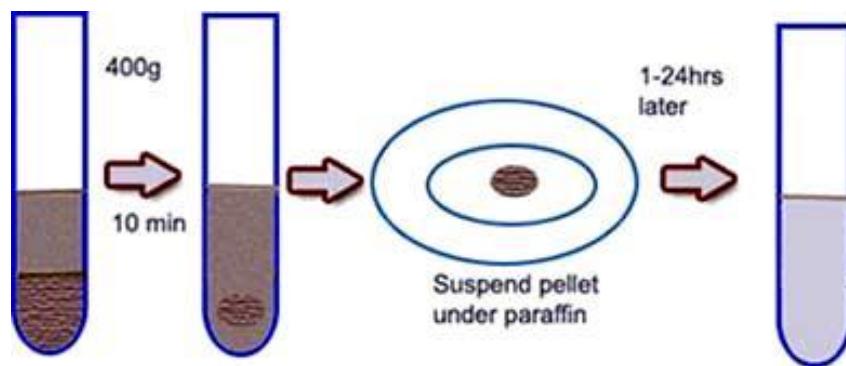


DIFFERENT WAYS TO SEPARATE SOLID-LIQUID MIXTURES



SEDIMENTATION METHOD

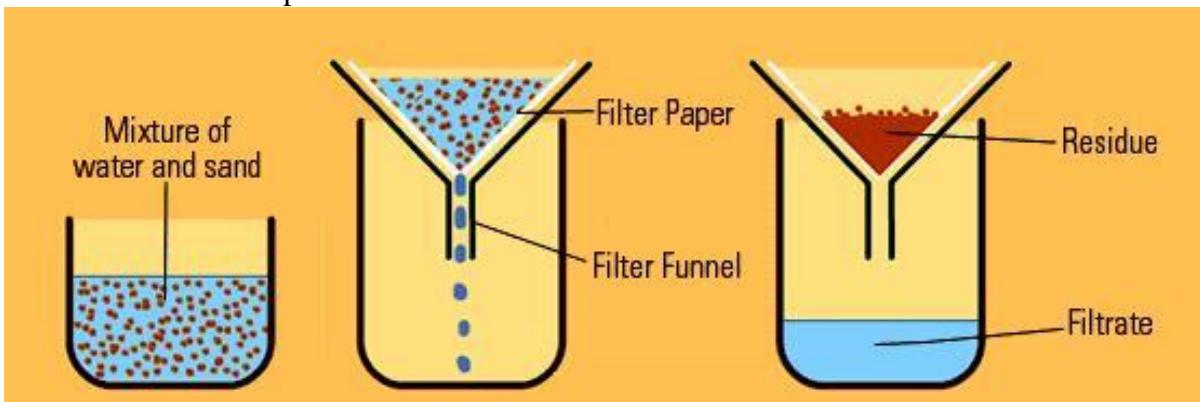
Sedimentation is the tendency for particles in suspension to settle out of the fluid in which they are entrained and come to rest against a barrier. This is due to their motion through the fluid in response to the forces acting on them: these forces can be due to gravity, centrifugal acceleration, or electromagnetism. In geology, sedimentation is often used as the opposite of erosion, i.e., the terminal end of sediment transport. In that sense, it includes the termination of transport by saltation or true bedload transport. Settling is the falling of suspended particles through the liquid, whereas sedimentation is the termination of the settling process.



FILTRATION METHOD

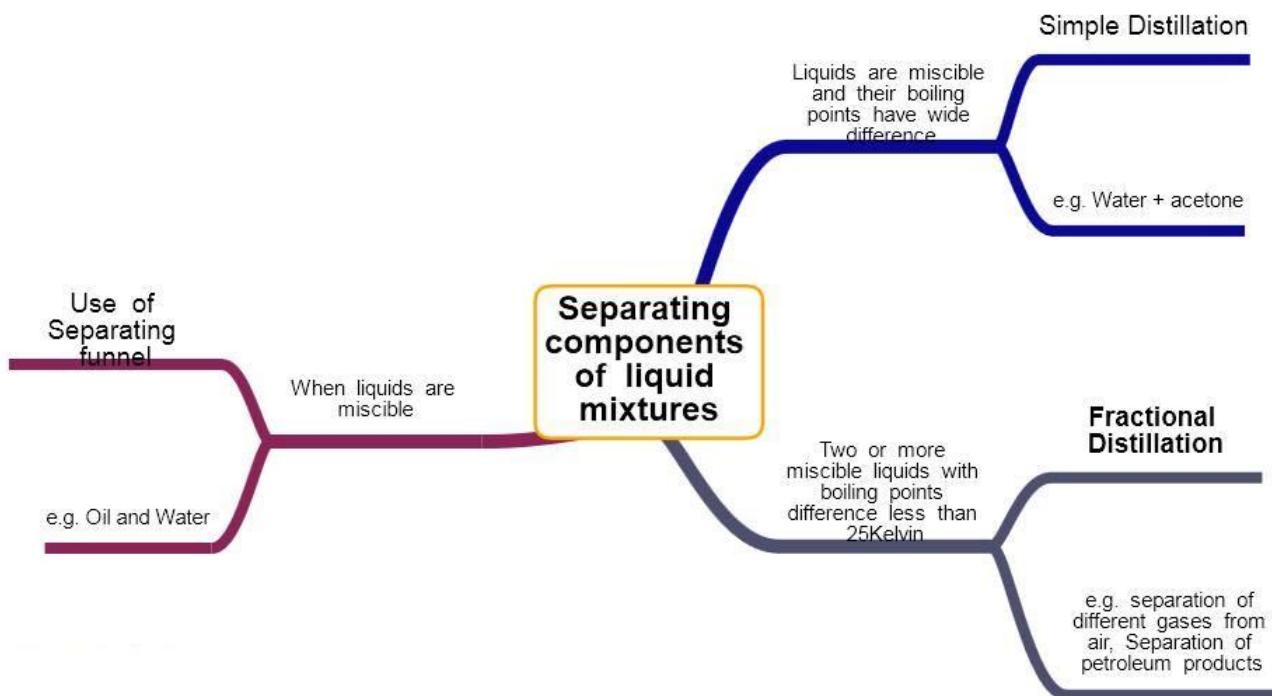
This is a more common method of separating an insoluble solid from a liquid. An example of such a mixture is sand and water. Filtration is used in water treatment plants, where water from the river is filtered to remove solid particles.

Here is a basic lab setup for filtration:



This process involves the use of a filter paper placed in a filter funnel. The funnel is placed in a beaker and the mixture of water and sand is poured into the funnel. The liquid part drains through the filter paper into the beaker, leaving the solid sand particles trapped on the filter. In filtration, the liquid part collected is called the filtrate and the solid bit that remained on the filter paper is called the residue.

DIFFERENT WAYS TO SEPARATE LIQUID-LIQUID MIXTURES



- Fractional distillation is the process of separating two or more miscible liquids by a modified distillation process, in which the distillates are collected as fractions having different boiling points. The separation of the liquids by this method is based on the difference in their boiling points.
- Fractional distillation makes use of a fractionating column or distillation column, a tube which provides different temperature zones inside it during distillation, the temperature decreasing from bottom to top. It provides surfaces on which condensations (of less volatile liquids) and vaporizations (of more volatile liquids) can occur before the vapours enter the condenser in order to concentrate the more volatile liquid in the first fractions and the less volatile components in the later fractions.
- Fractional distillation is very effective in separating mixtures of volatile components, and is widely used in laboratories and industries.

EXERCISE QUESTIONS PAGE NO. 28 to 30

Q1. Which separation techniques will you apply for the separation of the following?

- (a) Sodium chloride from its solution in water.
- (b) Ammonium chloride from a mixture containing sodium chloride and ammonium chloride.
- (c) Small pieces of metal in the engine oil of a car.
- (d) Different pigments from an extract of flower petals.
- (e) Butter from curd.
- (f) Oil from water.
- (g) Tea leaves from tea.
- (h) Iron pins from sand.
- (i) Wheat grains from husk.
- (j) Fine mud particles suspended in water.

Answer:

- (a) Sodium chloride from its solution in water → Evaporation
- (b) Ammonium chloride from a mixture containing sodium chloride and ammonium chloride → Sublimation
- (c) Small pieces of metal in the engine oil of a car → Centrifugation or filtration or decantation
- (d) Different pigments from an extract of flower petals → Chromatography
- (e) Butter from curd → Centrifugation
- (f) Oil from water → Using separating funnel
- (g) Tea leaves from tea → Filtration
- (h) Iron pins from sand → Magnetic separation
- (i) Wheat grains from husk → Winnowing
- (j) Fine mud particles suspended in water → Centrifugation

Q2. Write the steps you would use for making tea. Use the words solution, solvent, solute, dissolve, soluble, insoluble, filtrate and residue.

Answer: First, water is taken as a solvent in a saucer pan. This water (solvent) is allowed to boil. During heating, milk and tea leaves are added to the solvent as solutes. They form a solution. Then, the solution is poured through a strainer. The insoluble part of the solution remains on the strainer as residue. Sugar is added to the filtrate, which dissolves in the filtrate. The resulting solution is the required tea.

Q3. Pragya tested the solubility of three different substances at different temperatures and collected the data as given below (results are given in the following table, as grams of substance dissolved in 100 grams of water to form a saturated solution).

Substance Dissolved	Temperature in K				
	283	293	313	333	353
Potassium nitrate	21	32	62	106	167
Sodium chloride	36	36	36	37	37
Potassium chloride	35	35	40	46	54
Ammonium chloride	24	37	41	55	66

- (a) What mass of potassium nitrate would be needed to produce a saturated solution of potassium nitrate in 50 grams of water at 313 K?
- (b) Pragya makes a saturated solution of potassium chloride in water at 353 K and leaves the solution to cool at room temperature. What would she observe as the solution cools? Explain.
- (c) Find the solubility of each salt at 293 K. Which salt has the highest solubility at this temperature?
- (d) What is the effect of change of temperature on the solubility of a salt?

Answer:

(a) At 313 K, 62 grams of Potassium nitrate dissolved in 100 grams of water. So to produce a saturated solution of potassium nitrate in 50 grams of water, we need $\frac{62}{100} \times 50 = 31$ grams of potassium nitrate

(b) Some soluble potassium chloride will separate out in the form of crystals at room temperature because the solubility of potassium chloride will decrease with decrease in temperature.

(c)

- (i) Solubility of Potassium nitrate at 293 K is 32 grams.
- (ii) Solubility of Sodium chloride at 293 K is 36 grams.
- (iii) Solubility of Potassium chloride at 293 K is 35 grams.
- (iv) Solubility of Ammonium chloride at 293 K is 37 grams.

The solubility of Ammonium chloride is highest at this temperature.

(d) The solubility of salt increases with increase in temperature.

Q4. Explain the following giving examples.

(a) saturated solution

(b) pure substance

(c) colloid

(d) suspension

Answer:

(a) Saturated solution

A saturated solution is a solution in which the maximum amount of solute has been dissolved at a given temperature. The solution cannot dissolve beyond that amount of solute at that temperature. Any more solute added will settle down at the bottom of the container as a precipitate.

Suppose 500 g of a solvent can dissolve a maximum of 150 g of a particular solute at 40°C. Then, the solution obtained by dissolving 150 g of that solute in 500 g of that solvent at 300 K is said to be a saturated solution at 300 K.

(b) Pure substance

A pure substance is a substance consisting of a single type of particles i.e., all constituent particles of the substance have the same chemical properties.

For example, salt, sugar, water are pure substances.

(c) Colloid

A colloid is a heterogeneous mixture. The size of the solutes in this mixture is so small that they cannot be seen individually with naked eyes, and seems to be distributed uniformly throughout the mixture. The solute particles do not settle down when the mixture is left undisturbed. This means that colloids are quite stable. Colloids cannot be separated by the process of filtration. They can be separated by centrifugation. Colloids show the Tyndall effect. For example, milk, butter, foam, fog, smoke, clouds.

(d) Suspension

Suspensions are heterogeneous mixtures. The solute particles in this mixture remain suspended throughout the bulk of the medium. The particles can be seen with naked eyes. Suspension shows the Tyndall effect. The solute particles settle down when the mixture is left undisturbed. This means that suspensions are unstable. Suspensions can be separated by the method of filtration. For example, mixtures of chalk powder and water, wheat flour and water.

Q5. Classify each of the following as a homogeneous or heterogeneous mixture. soda water, wood, air, soil, vinegar, filtered tea.

Answer:

Homogeneous mixtures: Soda water, air, vinegar

Heterogeneous mixtures: Wood, soil, filtered tea

Q6. How would you confirm that a colourless liquid given to you is pure water?

Answer: Every liquid has a characteristic boiling point. Pure water has a boiling point of 100°C (373 K) at 1 atmospheric pressure. If the given colourless liquid boils at even slightly above or below 100°C, then the given liquid is not pure water. It must boil at sharp 100°C. Thus, by observing the boiling point, we can confirm whether a given colourless liquid is pure water or not.

Q7. Which of the following materials fall in the category of a “pure substance”?

- (a) Ice
- (b) Milk
- (c) Iron
- (d) Hydrochloric acid
- (e) Calcium oxide
- (f) Mercury
- (g) Brick
- (h) Wood
- (i) Air.

Answer: The following materials fall in the category of a “pure substance”:

- | | | |
|-------------------|-------------|-----------------------|
| (a) Ice | (c) Iron | (d) Hydrochloric acid |
| (e) Calcium oxide | (f) Mercury | |

Q8. Identify the solutions among the following mixtures.

- (a) Soil
- (b) Sea water
- (c) Air
- (d) Coal
- (e) Soda water.

Answer: The following mixtures are solutions:

- | | | |
|---------------|---------|----------------|
| (b) Sea water | (c) Air | (e) Soda water |
|---------------|---------|----------------|

Q9. Which of the following will show “Tyndall effect”?

- (a) Salt solution
- (b) Milk
- (c) Copper sulphate solution
- (d) Starch solution.

Answer: Milk and starch solution will show the “Tyndall effect”.

Q10. Classify the following into elements, compounds and mixtures.

- | | | |
|-------------|-----------------------|--------------------|
| (a) Sodium | (b) Soil | (c) Sugar solution |
| (d) Silver | (e) Calcium carbonate | (f) Tin |
| (g) Silicon | (h) Coal | (i) Air |
| (j) Soap | (k) Methane | (l) Carbon dioxide |
| (m) Blood | | |

Answer:

Elements

- | | | | |
|------------|------------|---------|-------------|
| (a) Sodium | (d) Silver | (f) Tin | (g) Silicon |
|------------|------------|---------|-------------|

Compounds

(e) Calcium carbonate (k) Methane (l) Carbon dioxide

Mixtures

(b) Soil (c) Sugar solution (h) Coal
(i) Air (j) Soap (m) Blood

Q11. Which of the following are chemical changes?

- (a) Growth of a plant
- (b) Rusting of iron
- (c) Mixing of iron filings and sand
- (d) Cooking of food
- (e) Digestion of food
- (f) Freezing of water
- (g) Burning of a candle.

Answer: The following changes are chemical changes:

- (a) Growth of a plant
 - (b) Rusting of iron
 - (d) Cooking of food
 - (e) Digestion of food
 - (g) Burning of candle
-

ASSIGNMENT QUESTIONS SET – 1
CHAPTER – 2
IS MATTER AROUND US PURE

- 1.** On the basis of composition, how matter is classified?
- 2.** Mention whether the following statements are true or false. Correct the false statements.
 - a) An aqueous solution of copper sulphate is homogeneous.
 - b) Milk is a pure substance.
 - c) A molecule of sulphur is monoatomic.
- 3.** What is meant by a pure substance?
- 4.** What are the characteristics exhibited by a pure substance?
- 5.** What are different categories of pure substance?
- 6.** Name two properties of a substance to check its purity?
- 7.** Define mixture.
- 8.** What are the kinds of mixture?
- 9.** What are the characteristics of mixture?
- 10.** List the points of differences between homogeneous and heterogeneous mixtures.
- 11.** Identify which of the following is homogenous mixture or heterogeneous one. Also identify the type of constituents in mixture (e.g. gas in gas, gas in liquid, gas in solid etc.)
 - a. Air
 - b. Water and Oil ($N_2 + O_2$)
 - c. Hydrogen in Palladium
 - d. Aerated Water ($CO_2 + H_2O$)
 - e. Chalk in water
 - f. Ethyl Alcohol in Water
 - g. Alloys e.g. brass
 - h. Dust (e.g. fine sand) in water
 - i. Sand + iron fillings
 - j. Sand + ammonium chloride
 - h. milk
 - i. Mercury in amalgamated Zinc
- 12.** What are the constituents of brass?
- 13.** Alloys cannot be separated by physical means, though it is considered mixture, Why?
- 14.** What are elements?
- 15.** How elements are further classified?

- 16.** What is a compound? Give an example.
- 17.** State the differences between compounds and mixtures.
- 18.** What is a solution? What are the properties of a solution?
- 19.** What is meant by solute and solvent?
- 20.** Identify solute and solvent in the following solutions. Also mention the physical state of solute and solvent.
- (a) Sugar in water
 - (b) Urea in water
 - (c) Ammonium chloride in water
 - (d) Ethyl alcohol in water
 - (e) Carbon Di-Oxide in water (soda water)
- 21.** What is meant by Solubility?
- 22.** What factors affect the solubility of solvent and solute?
- 23.** Why do fish go in deep waters during day light?
- 24.** Based on the type of solvent, how solutions are classified?
- 25.** Based on the amount of solute in the given solution, how solutions are classified?
- 26.** What are aqueous solutions?
- 27.** What are non-aqueous solutions?
- 28.** When we open the cap of a cola drink (or any carbonated beverage), why does excess of bubbles come out?
- 29.** Why air is a mixture not a compound? Give reasons.
- 30.** What are the advantages of preparing solutions?
- 31.** What is an unsaturated solution?
- 32.** What is saturated solution? Explain with an example.
- 33.** What is supersaturated solution?
- 34.** What are suspensions? Explain with an example.
- 35.** What are the properties of suspensions?
- 36.** Explain with an example what is a colloid?
- 37.** What are the physical states of dispersed phase and dispersion medium of a cloud?
- 38.** What are the physical states of dispersed phase and dispersion medium of a fog?
- 39.** What are the properties of a colloid?
- 40.** How are sol, solution and suspension different from each other?
- 41.** What is Tyndall effect? Does true solution exhibit Tyndall effect.
- 42.** "Tyndall effect can be observed when sunlight passes through the canopy of dense forest. or we see a rich red sunset." Explain how this occurs

- 43.** What do you mean by strength of the solution?
- 44.** What are the various methods to express concentration of a solution?
- 45.** To make a saturated solution, 36 g of sodium chloride is dissolved in 100 g of water at 293K. Find its concentration at this temperature.
- 46.** Calculate the mass of glucose and mass of water required to make 200g of 25% solution of glucose.
- 47.** A solution contains 40 mL of ethyl alcohol mixed with 100mL of water. What is the concentration of the solution in terms of volume by volume percentage?
- 48.** What are different ways to separate solid mixtures?
- 49.** Describe sieving method.
- 50.** Winnowing works on what property?
- 51.** What are the reasons for separating the constituents of a mixture?
- 52.** When it is useful to apply sedimentation and decantation? Give an example where these methods are used? Explain the process.
- 53.** What principle is applied in centrifugation? Give examples where this method is applied to separate mixtures.
- 54.** How crystallization is better than evaporation?
- 55.** Name the technique to separate (i) butter from curd (ii) salt from sea-water (iii) camphor from salt
- 56.** What is Chromatography?
- 57.** A good method to separate alum (phitkari) from impure samples is (a) Filtration (b) Sedimentaion (c) Crystallization (d) Sublimation
- 58.** A boy buys common salt from the market which is contaminated with Ammonium Chloride (NH_4Cl) and sand. The procedure he should adopt to obtain pure NaCl is the following :
- to mix the sample in water and evaporate the solution
 - to mix the sample in water and evaporate the decanted solution
 - to mix the sample in acetone and evaporate the decanted solution
 - to heat the sample, then mix in water and evaporate the decanted solution
- 59.** Define Brownian movement in colloids.
- 60.** Name the following :
- a lustrous liquid metal.
 - a liquid non-metal
 - a metal which can be cut with a knife
 - a non-metal which is good conductor of electricity.

- (e) an element which melts when kept on the palm.
- (f) the best conductor of heat.

- 61.** How many elements are there which are in gaseous state at room temperature?
 - 62.** Name the elements are in liquid state at room temperature.
 - 63.** Who used the term 'element' first time?
 - 64.** Who gave the first explanatory definition of 'element'?
 - 65.** Identify the following as mixture or compound. (i) blood (ii) common salt (iii) sugar (iv) brass
 - 66.** Sasha heats a container carrying Nitrogen and Oxygen. After heating at very high temperature, it gives Nitric oxide. Identify what are mixtures and/or compounds before and after the reaction.
 - 67.** In beaker A, sugar cubes are dissolved into water while in beaker B, crushed cubes are taken. In which beaker the rate of dissolution is faster?
 - 68.** Identify solute and solvent in the following solutions :
 - (i) aerated drinks
 - (ii) tincture of iodine
 - (iii) lemon water
 - 69.** State the principle of each of the following methods of separation of mixtures.
 - (i) centrifugation method.
 - (ii) separation using separating funnel.
 - 70.** Why solutions do not exhibit Tyndall effect?
 - 71.** What is an emulsion? Give examples
 - 72.** What are the differences and similarities between concentration and solubility?
 - 73.** Give examples of liquids that are
 - (i) completely miscible
 - (ii) partially miscible
 - (iii) practically immiscible
-
-

ASSIGNMENT QUESTIONS SET – 2
CHAPTER – 2
IS MATTER AROUND US PURE

1. Salt can be recovered from its solution by evaporation. Suggest some other technique for the same?
2. While diluting a solution of salt in water, a student by mistake added acetone (boiling point 56°C). What technique can be employed to get back the acetone? Justify your choice.
3. Classify the following as element, compound or mixture. i) Zinc amalgam (ii) Sea water (iii) Iodine vapour (iv) Gold coin (v) Water.
4. How would you separate a mixture of ammonia and hydrogen?
5. Give one example for each of the following mixtures: i) Solid/solid (homogeneous) ii) Solid/solid (heterogeneous) iii) Liquid/liquid (homogeneous) iv) Liquid/liquid (heterogeneous) v) Gas/liquid (homogeneous).
6. Explain why particles of a colloidal solution do not settle down when left undisturbed, while in the case of a suspension they do.
7. Smoke and fog both are aerosols. In what way are they different?
8. Name the process associated with the following
 - (a) Dry ice is kept at room temperature and at one atmospheric pressure.
 - (b) A drop of ink placed on the surface of water contained in a glass spreads throughout the water.
 - (c) A potassium permanganate crystal is in a beaker and water is poured into the beaker with stirring.
 - (d) A acetone bottle is left open and the bottle becomes empty. (e) Milk is churned to separate cream from it.
 - (e) Settling of sand when a mixture of sand and water is left undisturbed for some time.
 - (f) Fine beam of light entering through a small hole in a dark room, illuminates the particles in its paths.
9. The teacher instructed three students ‘A’, ‘B’ and ‘C’ respectively to prepare a 50% (mass by volume) solution of sodium hydroxide (NaOH). ‘A’ dissolved 50g of NaOH in 100 mL of water, ‘B’ dissolved 50g of NaOH in 100g of water while ‘C’ dissolved 50g of NaOH in water to make 100 mL of solution. Which one of them has made the desired solution and why?
10. Why is gold alloyed with copper or silver for the purpose of making ornaments?
11. Give some examples of Tyndall effect observed in your surroundings?

- 12.** Calculate the mass of sodium sulphate required to prepare its 20% (mass percent) solution in 100g of water?
- 13.** How would you separate a mixture of ammonia and hydrogen?
- 14.** Action of heat on blue vitriol is a physical as well as chemical change. Justify.
- 15.** How would you separate a mixture of NH_4Cl and I_2 ?
- 16.** Describe a method for separation of the constituents of gunpowder.
- 17.** Describe how you would obtain the substances mentioned below, from the given mixtures.
a) Iodine from tincture of iodine. b) Lead chloride from a mixture of lead chloride and silver chloride
- 18.** Write the characteristics of Suspension method of separation of mixture.
- 19.** Briefly describe how to separate, i) Sulphur from a mixture of sulphur and sand. ii) Black CuO from a mixture of CuO and ZnO .
- 20.** Fill in the blanks
- A colloid is a _____ mixture and its components can be separated by the technique known as _____.
 - Ice, water and water vapour look different and display different _____ properties but they are _____ the same.
 - A mixture of chloroform and water taken in a separating funnel is mixed and left undisturbed for some time. The upper layer in the separating funnel will be of _____ and the lower layer will be that of _____.
 - A mixture of two or more miscible liquids, for which the difference in the boiling points is less than 25 K can be separated by the process called _____.
 - When light is passed through water containing a few drops of milk, it shows a bluish tinge. This is due to the _____ of light by milk and the phenomenon is called _____. This indicates that milk is a _____ solution.
- 21.** Which of the following is NOT true about colloids?
(a) Particles in a colloid can pass through filter paper.
(b) A colloid is a homogeneous solution.
(c) Colloidal particles exhibit Brownian motion.
(d) Colloidal particles exhibit electrophoresis.
- 22.** Which of the following is a heterogeneous mixture?
(a) air
(b) brass
(c) sugar dissolved in water
(d) lime water
- 23.** How can we separate cream from milk?
(a) Centrifugation

- (b) Chromatography
- (c) Sublimation
- (d) Distillation

24. Removal of clear layer of the liquid without disturbing the settled solid is known as _____.

- (a) Sedimentation
- (b) Decantation
- (c) Filtration
- (d) Evaporation

25. A mixture of water and silver chloride can be separated by

- (a) Centrifugation
- (b) Sedimentation
- (c) Filtration
- (d) Sublimation

26. When two liquids in a mixture differ by their boiling points, which of the following is the best method to separate these liquids?

- (a) Evaporation
- (b) Distillation
- (c) Chromatography
- (d) Filtration

27. Solution which has uniform composition throughout is called _____.

- (a) homogeneous solution
- (b) heterogeneous solution
- (c) Colloidal solution
- (d) none of these

28. The particles of a suspension will be of size _____

- (a) less than 1nm (nm = nano metre)
- (b) between 1nm to 100 nm
- (c) greater than 100nm
- (d) less than 0.1 nm

29. Which of the following is a chemical change?

- (a) Melting of ice
- (b) Dissolving salt in water
- (c) Rusting of iron
- (d) Boiling of water into steam

30. Which one of the following is TRUE for compounds?

- (a) It shows properties of its constituents.
- (b) It may be homogeneous or heterogeneous.
- (c) It can be separated by physical methods.
- (d) It has fixed melting and boiling point.



ASSIGNMENT QUESTIONS SET – 3
CHAPTER – 2
IS MATTER AROUND US PURE

- 1.** Which of the following statements are true for pure substances?
 - (i) Pure substances contain only one kind of particles
 - (ii) Pure substances may be compounds or mixtures
 - (iii) Pure substances have the same composition throughout
 - (iv) Pure substances can be exemplified by all elements other than nickel
 - (a) (i) and (ii)
 - (b) (i) and (iii)
 - (c) (iii) and (iv)
 - (d) (ii) and (iii)
- 2.** Why do we call sugar a pure substance?
- 3.** What are saturated and unsaturated solutions?
- 4.** Define a solution.
- 5.** What is a suspension? Give its example and properties.
- 6.** Define concentration of a solution.
- 7.** What is Tyndall effect?
- 8.** What is the difference between True solution and colloids?
- 9.** What are alloys? Why are alloys called as mixture?
- 10.** Write the characteristics of brass.
- 11.** Define solute and solvent.
- 12.** What is solubility?
- 13.** Give properties of a true solution.
- 14.** Rusting of an article made up of iron is called
 - (a) corrosion and it is a physical as well as chemical change
 - (b) dissolution and it is a physical change
 - (c) corrosion and it is a chemical change
 - (d) dissolution and it is a chemical change
- 15.** A mixture of sulphur and carbon disulphide is
 - (a) heterogeneous and shows Tyndall effect
 - (b) homogeneous and shows Tyndall effect
 - (c) heterogeneous and does not show Tyndall effect
 - (d) homogeneous and does not show Tyndall effect
- 16.** Tincture of iodine has antiseptic properties. This solution is made by dissolving
 - (a) iodine in potassium iodide
 - (b) iodine in vaseline
 - (c) iodine in water
 - (d) iodine in alcohol
- 17.** Why do we need to separate mixtures?
- 18.** How can we separate cream from milk?

19. Which of the following are homogeneous in nature?

- (i) ice (ii) wood (iii) soil (iv) air
- (a) (i) and (iii)
- (b) (ii) and (iv)
- (c) (i) and (iv)
- (d) (iii) and (iv)

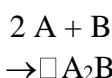
20. Which of the following are physical changes?

- (i) Melting of iron metal
- (ii) Rusting of iron
- (iii) Bending of an iron rod
- (iv) Drawing a wire of iron metal
- (a) (i), (ii) and (iii)
- (b) (i), (ii) and (iv)
- (c) (i), (iii) and (iv)
- (d) (ii), (iii) and (iv)

21. Which of the following are chemical changes?

- (i) Decaying of wood
- (ii) Burning of wood
- (iii) Sawing of wood
- (iv) Hammering of a nail into a piece of wood
- (a) (i) and (ii)
- (b) (ii) and (iii)
- (c) (iii) and (iv)
- (d) (i) and (iv)

22. Two substances, A and B were made to react to form a third substance, A_2B according to the following reaction



Which of the following statements concerning this reaction are incorrect?

- (i) The product A_2B shows the properties of substances A and B
- (ii) The product will always have a fixed composition
- (iii) The product so formed cannot be classified as a compound
- (iv) The product so formed is an element
- (a) (i), (ii) and (iii),
- (b) (ii), (iii) and (iv)
- (c) (i), (iii) and (iv)
- (d) (ii), (iii) and (iv)

23. Two chemical species X and Y combine together to form a product P which contains both X and Y



X and Y cannot be broken down into simpler substances by simple chemical reactions.

Which of the following concerning the species X, Y and P are correct?

- (i) P is a compound
- (ii) X and Y are compounds
- (iii) X and Y are elements
- (iv) P has a fixed composition
- (a) (i), (ii) and (iii),
- (b) (i), (ii) and (iv)

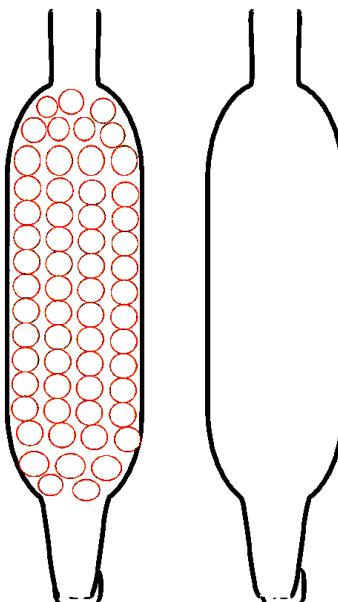
(c) (ii), (iii) and (iv)

(d) (i), (iii) and (iv)

24. Suggest separation technique(s) one would need to employ to separate the following mixtures.

- (a) Mercury and water
- (b) Potassium chloride and ammonium chloride
- (c) Common salt, water and sand
- (d) Kerosene oil, water and salt

25. Which of the tubes in the below figure (a) and (b) will be more effective as a condenser in the distillation apparatus?



26. Salt can be recovered from its solution by evaporation. Suggest some other technique for the same?

27. The ‘sea-water’ can be classified as a homogeneous as well as heterogeneous mixture. Comment.

28. While diluting a solution of salt in water, a student by mistake added acetone (boiling point 56°C). What technique can be employed to get back the acetone? Justify your choice.

29. What would you observe when (a) a saturated solution of potassium chloride prepared at 60°C is allowed to cool to room temperature. (b) an aqueous sugar solution is heated to dryness. (c) a mixture of iron filings and sulphur powder is heated strongly.

30. Explain why particles of a colloidal solution do not settle down when left undisturbed, while in the case of a suspension they do.

31. Smoke and fog both are aerosols. In what way are they different?

32. Classify the following as physical or chemical properties

- (a) The composition of a sample of steel is: 98% iron, 1.5% carbon and 0.5% other elements.
- (b) Zinc dissolves in hydrochloric acid with the evolution of hydrogen gas.
- (c) Metallic sodium is soft enough to be cut with a knife.
- (d) Most metal oxides form alkalis on interacting with water.

- 33.** The teacher instructed three students ‘A’, ‘B’ and ‘C’ respectively to prepare a 50% (mass by volume) solution of sodium hydroxide (NaOH). ‘A’ dissolved 50g of NaOH in 100 mL of water, ‘B’ dissolved 50g of NaOH in 100g of water while ‘C’ dissolved 50g of NaOH in water to make 100 mL of solution. Which one of them has made the desired solution and why?
- 34.** Name the process associated with the following
- Dry ice is kept at room temperature and at one atmospheric pressure.
 - A drop of ink placed on the surface of water contained in a glass spreads throughout the water.
 - A potassium permanganate crystal is in a beaker and water is poured into the beaker with stirring.
 - A acetone bottle is left open and the bottle becomes empty.
 - Milk is churned to separate cream from it.
 - Settling of sand when a mixture of sand and water is left undisturbed for some time.
 - Fine beam of light entering through a small hole in a dark room, illuminates the particles in its paths.
- 35.** Write the applications of centrifugation.
- 36.** How can we separate a mixture of salt and ammonium chloride? Draw a diagram.
- 37.** What is chromatography? Explain the process.
- 38.** How can you separate copper sulphate from an impure sample?
- 39.** You are given two samples of water labelled as ‘A’ and ‘B’. Sample ‘A’ boils at 100°C and sample ‘B’ boils at 102°C. Which sample of water will not freeze at 0°C? Comment.
- 40.** What are the favourable qualities given to gold when it is alloyed with copper or silver for the purpose of making ornaments?
- 41.** An element is sonorous and highly ductile. Under which category would you classify this element? What other characteristics do you expect the element to possess?
- 42.** Give an example each for the mixture having the following characteristics. Suggest a suitable method to separate the components of these mixtures
- A volatile and a non-volatile component.
 - Two volatile components with appreciable difference in boiling points.
 - Two immiscible liquids.
 - One of the components changes directly from solid to gaseous state.
 - Two or more coloured constituents soluble in some solvent.
- 43.** Fill in the blanks
- A colloid is a _____ mixture and its components can be separated by the technique known as _____.
 - Ice, water and water vapour look different and display different _____ properties but they are _____ the same.
 - A mixture of chloroform and water taken in a separating funnel is mixed and left undisturbed for some time. The upper layer in the separating funnel will be of _____ and the lower layer will be that of _____.
 - A mixture of two or more miscible liquids, for which the difference in the boiling points is less than 25 K can be separated by the process called _____.

(e) When light is passed through water containing a few drops of milk, it shows a bluish tinge. This is due to the _____ of light by milk and the phenomenon is called _____. This indicates that milk is a _____ solution.

44. Sucrose (sugar) crystals obtained from sugarcane and beetroot are mixed together. Will it be a pure substance or a mixture? Give reasons for the same.

45. Give some examples of Tyndall effect observed in your surroundings?

46. Can we separate alcohol dissolved in water by using a separating funnel? If yes, then describe the procedure. If not, explain.

47. On heating calcium carbonate gets converted into calcium oxide and carbon dioxide.

(a) Is this a physical or a chemical change?

(b) Can you prepare one acidic and one basic solution by using the products formed in the above process? If so, write the chemical equation involved.

48. Non metals are usually poor conductors of heat and electricity. They are non-lustrous, non-sonorous, non-malleable and are coloured.

(a) Name a lustrous non-metal.

(b) Name a non-metal which exists as a liquid at room temperature.

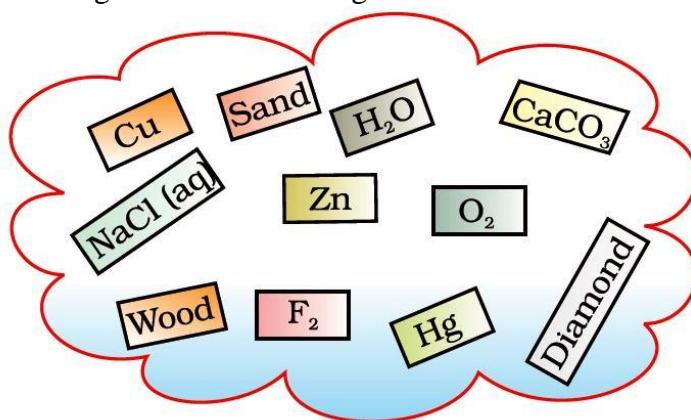
(c) The allotropic form of a non-metal is a good conductor of electricity. Name the allotrope.

(d) Name a non-metal which is known to form the largest number of compounds.

(e) Name a non-metal other than carbon which shows allotropy.

(f) Name a non-metal which is required for combustion.

49. Classify the substances given in the below figure into elements and compounds



50. Which of the following are not compounds?

- (a) Chlorine gas
- (b) Potassium chloride
- (c) Iron
- (d) Iron sulphide
- (e) Aluminium
- (f) Iodine
- (g) Carbon
- (h) Carbon monoxide
- (i) Sulphur powder

51. Fractional distillation is suitable for separation of miscible liquids with a boiling point difference of about 25 K or less. What part of fractional distillation apparatus makes it

efficient and possess an advantage over a simple distillation process. Explain using a diagram.

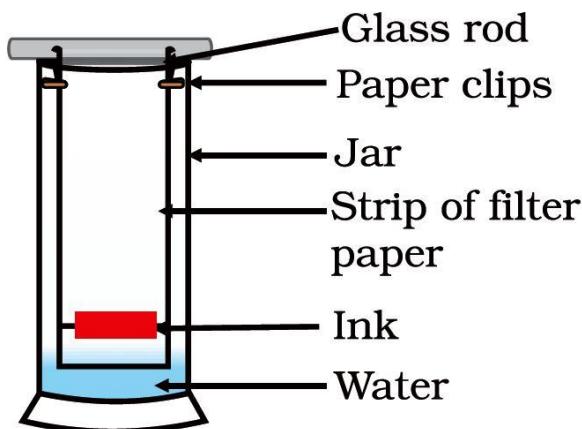
52. (a) Under which category of mixtures will you classify alloys and why?

(b) A solution is always a liquid. Comment.

(c) Can a solution be heterogeneous?

53. Iron filings and sulphur were mixed together and divided into two parts, 'A' and 'B'. Part 'A' was heated strongly while Part 'B' was not heated. Dilute hydrochloric acid was added to both the Parts and evolution of gas was seen in both the cases. How will you identify the gases evolved?

54. A child wanted to separate the mixture of dyes constituting a sample of ink. He marked a line by the ink on the filter paper and placed the filter paper in a glass containing water as shown in below figure. The filter paper was removed when the water moved near the top of the filter paper.

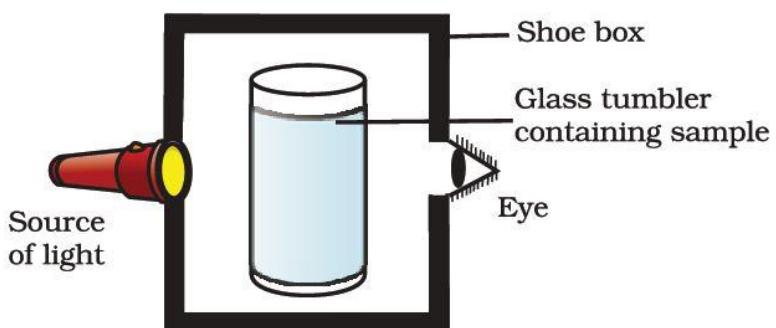


(i) What would you expect to see, if the ink contains three different coloured components?

(ii) Name the technique used by the child.

(iii) Suggest one more application of this technique.

55. A group of students took an old shoe box and covered it with a black paper from all sides. They fixed a source of light (a torch) at one end of the box by making a hole in it and made another hole on the other side to view the light. They placed a milk sample contained in a beaker/tumbler in the box as shown in the below figure. They were amazed to see that milk taken in the tumbler was illuminated. They tried the same activity by taking a salt solution but found that light simply passed through it?



(a) Explain why the milk sample was illuminated. Name the phenomenon involved.

(b) Same results were not observed with a salt solution. Explain.

(c) Can you suggest two more solutions which would show the same effect as shown by the milk solution?

- 56.** Classify each of the following, as a physical or a chemical change. Give reasons.
- (a) Drying of a shirt in the sun.
 - (b) Rising of hot air over a radiator.
 - (c) Burning of kerosene in a lantern.
 - (d) Change in the colour of black tea on adding lemon juice to it.
 - (e) Churning of milk cream to get butter.
- 57.** During an experiment the students were asked to prepare a 10% (Mass/Mass) solution of sugar in water. Ramesh dissolved 10g of sugar in 100g of water while Sarika prepared it by dissolving 10g of sugar in water to make 100g of the solution.
- (a) Are the two solutions of the same concentration
 - (b) Compare the mass % of the two solutions.
- 58.** You are provided with a mixture containing sand, iron filings, ammonium chloride and sodium chloride. Describe the procedures you would use to separate these constituents from the mixture?
- 59.** Arun has prepared 0.01% (by mass) solution of sodium chloride in water. Which of the following correctly represents the composition of the solutions?
- (a) 1.00 g of NaCl + 100g of water
 - (b) 0.11g of NaCl + 100g of water
 - (c) 0.0l g of NaCl + 99.99g of water
 - (d) 0.10 g of NaCl + 99.90g of water
- 60.** Calculate the mass of sodium sulphate required to prepare its 20% (mass percent) solution in 100g of water?
-
-

CHAPTER – 3

ATOMS AND MOLECULES

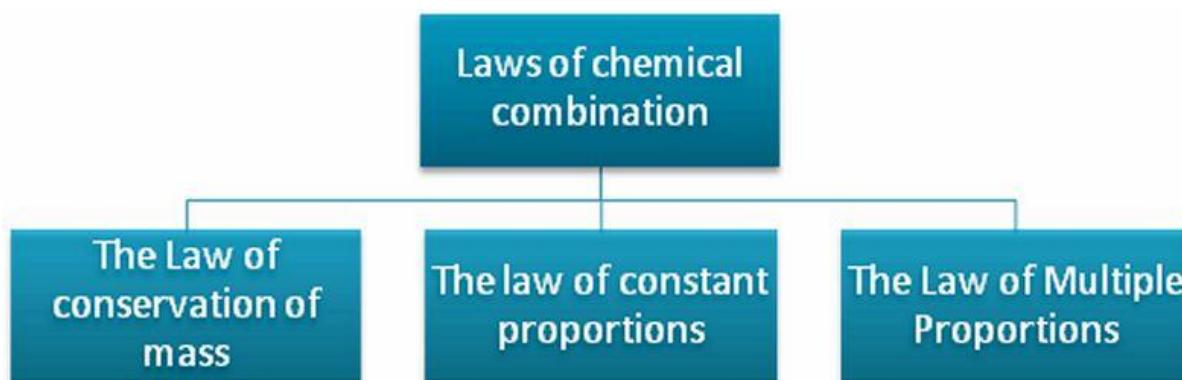
ATOMS AND MOLECULES

LAWS OF CHEMICAL COMBINATIONS

Before Dalton concept of atom was mere philosophical. Dalton explained about atom on the basis of Laws of Chemical Combinations.

There are three laws of chemical combination.

1. Law of Conservation of Mass
2. Law of Constant Proportions
3. Law of Multiple Proportions



LAW OF CONSERVATION OF MASS

Antoine L. Lavoisier, a French scientist, established the theory of Law of Conservation of Mass. The law of conservation of mass states, “**Mass can neither be created nor destroyed in a chemical reaction**”.

All matters in the universe exist in three states. There are two ways of classification of matter.

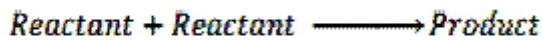
1. According to physical state as solid, liquid or gas.
2. According to its composition as element, compound or mixture.

According to this law mass of an isolated system will remain constant over time. This means when mass is enclosed in a system and none is allowed in or out, its quantity will never change. That is mass will be conserved, and hence this is called Law of Conservation of Mass. This means total mass of products is always equal to the total mass of reactants. As there is no loss of mass of substances, i.e. mass is conserved, that's why Lavoisier called this the law of conservation of mass.

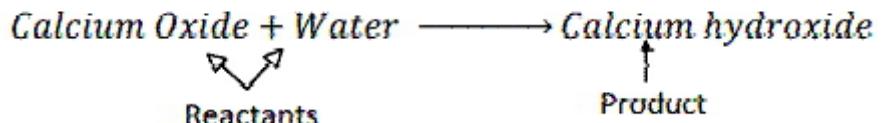
REACTANTS AND PRODUCTS:

In a chemical reaction the substances that combine or react are known as reactants and the new substance/substances formed are called product or products.

A chemical reaction can be represented in general as follows:

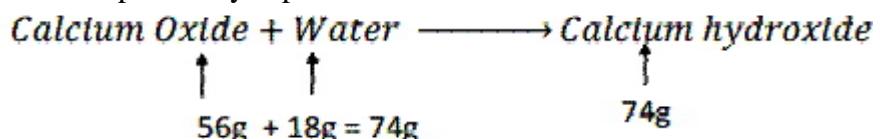


Example: When calcium oxide is dissolved in water calcium hydroxide is formed. The reaction involve in this can be written as:

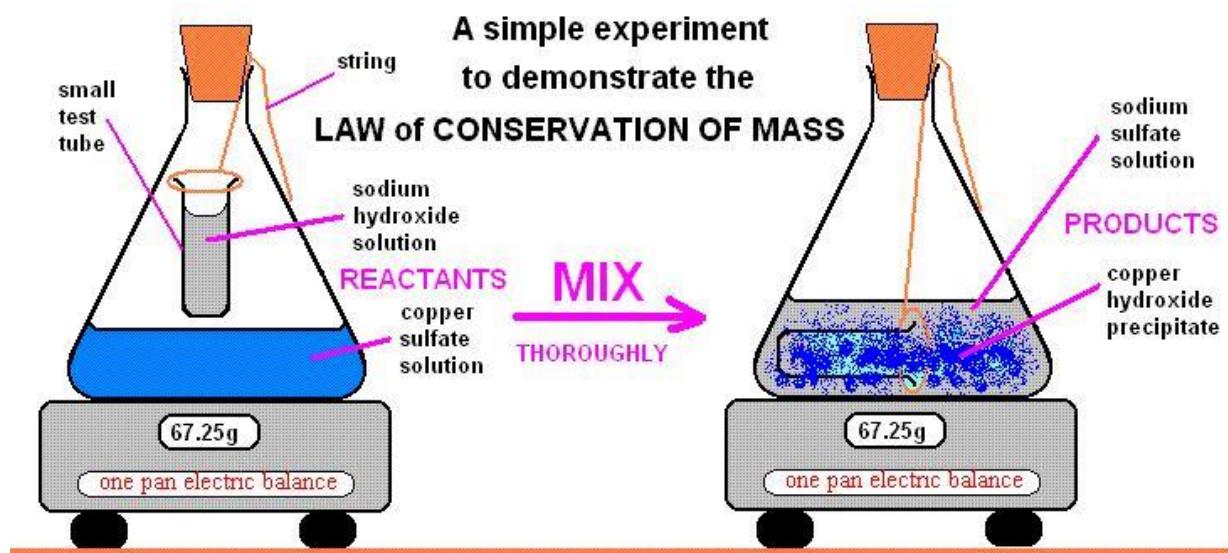


In this reaction calcium oxide and water are reactants while calcium hydroxide is product.

In this reaction 74 g of calcium hydroxide is obtained when 56 g of calcium oxide reacts with 18 g of water, which is proved by experiment.



Here the total mass of reactants, i.e. calcium oxide and water is equal to 74 g. And the mass of product, i.e. calcium hydroxide is also equal to 74g. This proves that the total mass of reactants is always equal to the total mass of product, which proves the Law of Conservation of Mass.



LAW OF CONSTANT PROPORTIONS

Law of Constant Proportion states that **a chemical compound always contains exactly the same proportion of elements by mass.**

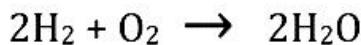
This law is also known as Law of definite proportions. Joseph Louis Proust gave this law hence, this law is also known as Proust's Law.

Explanation of the law:-

Compounds are formed by the combination of two or more elements. In a compound the ratio of the atoms or element by mass remains always same irrespective of the source of compound. This means a certain compound always formed by the combination of atoms in same ratio by mass. If the ratio of mass of constituent atoms will be altered the new compound is formed.

Examples:-

Water is formed by the combination of hydrogen and oxygen. The ratio of masses of hydrogen and oxygen is always in 1:8 in water irrespective of source of water. Whether you collect the water from a well, river, pond or from anywhere the ratio of their constituent atoms by mass will always same.



$$\begin{array}{cc} \downarrow & \downarrow \\ 4 & 32 \\ \downarrow & \downarrow \\ 1 : 8 \text{ (By Mass)} \end{array}$$

Nitrogen dioxide is a compound, which is formed by the combination of nitrogen and oxygen. The ratio of nitrogen and oxygen by mass in nitrogen dioxide is in 7:16.

Nitrous oxide is a compound which is also formed by the combination of nitrogen and oxygen. The ratio of nitrogen and oxygen in nitrous oxide is in 28:16.

Nitric oxide is a compound, which is also formed by the combination of nitrogen and oxygen. The ratio of nitrogen and oxygen in nitric oxide is in 7:8.

From the above three examples it is clear that if the ratio of the atoms by mass is altered then the new compound is formed, such as in the case of nitrogen dioxide, nitrous oxide, nitric oxide. These three compounds are formed by the combination of same atoms but because of combination of the constituent atoms in different ratios by mass new compound is formed.

Law of Conservation of mass:

**proposed by the French chemist
Antoine Lavoisier (1774)**

Mass can neither be created nor destroyed in a chemical reaction.

OR

For any chemical process in a closed system, the mass of the reactants must be equal the mass of the products.



$$12\text{g} + 32\text{g} = 44\text{g}$$

Law of definite proportion:

proposed by Louis Proust (1799)

'A chemical compound always consists of the same elements combined together in the same ratio,

Irrespective of the method of preparation or the source from where it is taken'.

One molecule of a compound water always contains same ratio of Hydrogen and Oxygen by mass i.e.

$$\text{H}_2\text{O} = 2 : 16 = 1 : 8$$

DALTON'S ATOMIC THEORY

John Dalton, a British Chemists and scientists gave the Atomic Theory in 1808. This theory is popularly known as Dalton's Atomic Theory in the honour of John Dalton. He gave the theory on the basis of Laws of Chemical Combination and explains them properly. In his theory he explains about atom.

Main postulates of Dalton's atomic theory

1. Elements are made of extremely small particles called atoms.
2. Atoms of a given element are identical in size, mass, and other properties;
3. Atoms of different elements differ in size, mass, and other properties.
4. Atoms cannot be subdivided, created, or destroyed.
5. Atoms of different elements combine in simple whole-number ratios to form chemical compounds.
6. In chemical reactions, atoms are combined, separated, or rearranged.

Dalton's Atomic Theory



1. Each element is composed of extremely small particles called atoms.

An atom of the element oxygen An atom of the element nitrogen

2. All atoms of a given element are identical, but the atoms of one element are different from the atoms of all other elements.

Oxygen Nitrogen

3. Atoms of one element cannot be changed into atoms of a different element by chemical reactions; atoms are neither created nor destroyed in chemical reactions.

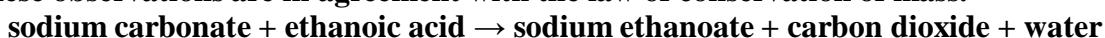
Oxygen → Nitrogen

4. Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kind of atoms.

Elements Compound

INTEXT QUESTIONS PAGE NO. 32

Q1. In a reaction, 5.3 g of sodium carbonate reacted with 6 g of ethanoic acid. The products were 2.2 g of carbon dioxide, 0.9 g water and 8.2 g of sodium ethanoate. Show that these observations are in agreement with the law of conservation of mass.



Answer:

In the given reaction, sodium carbonate reacts with ethanoic acid to produce sodium ethanoate, carbon dioxide, and water.



Mass of sodium carbonate = 5.3 g (Given)

Mass of ethanoic acid = 6 g (Given)

Mass of sodium ethanoate = 8.2 g (Given)

Mass of carbon dioxide = 2.2 g (Given)

Mass of water = 0.9 g (Given)

Now, total mass before the reaction = $(5.3 + 6)$ g = 11.3 g

And, total mass after the reaction = $(8.2 + 2.2 + 0.9)$ g = 11.3 g

Therefore, Total mass before the reaction = Total mass after the reaction

Hence, the given observations are in agreement with the law of conservation of mass.

Q2. Hydrogen and oxygen combine in the ratio of 1:8 by mass to form water. What mass of oxygen gas would be required to react completely with 3 g of hydrogen gas?

Answer:

It is given that the ratio of hydrogen and oxygen by mass to form water is 1:8.

Then, the mass of oxygen gas required to react completely with 1 g of hydrogen gas is 8 g.

Therefore, the mass of oxygen gas required to react completely with 3 g of hydrogen gas is $8 \times 3 \text{ g} = 24 \text{ g}$.

Q3. Which postulate of Dalton's atomic theory is the result of the law of conservation of mass?

Answer:

The postulate of Dalton's atomic theory which is a result of the law of conservation of mass is:

Atoms are indivisible particles, which can neither be created nor destroyed in a chemical reaction.

Q4. Which postulate of Dalton's atomic theory can explain the law of definite proportions?

Answer:

The postulate of Dalton's atomic theory which is a result of the law of conservation of mass is:

Atoms are indivisible particles, which can neither be created nor destroyed in a chemical reaction.

ATOMS

On the basis of Dalton's Atomic On the basis of Dalton's Atomic Theory atom can be defined as the smallest particles of matter are called atoms.

Characteristics of atoms:

- Atom is the smallest particle of matter.
- All elements are made of tiny particles called atom.
- Atoms are very small in size and cannot be seen through naked eyes.
- Atom does not exist in free-state in nature. But atom takes part in a chemical reaction.
- The properties of a matter depend upon the characteristics of atoms.
- Atoms are the building block of an element similar to a brick which combine together to make a building.
- The size of atoms is indicated by its radius.
- In ancient time atoms was considered indivisible.

SYMBOLS OF ATOMS OF ELEMENTS

Dalton was the first scientist to use the symbols for elements in a very specific sense. When he used a symbol for an element he also meant a definite quantity of that element, that is, one atom of that element. Berzilius suggested that the symbols of elements be made from one or two letters of the name of the element.

Many of the symbols are the first one or two letters of the element's name in English. The first letter of a symbol is always written as a capital letter (uppercase) and the second letter as a small letter (lowercase).

For convenience elements are represented by unique symbols. For example: Hydrogen is represented by 'H'. Oxygen is represented 'O'. Nitrogen is represented by 'N'. Iron is represented by 'Fe'. ments are represented by unique symbols. For example: Hydrogen is

represented by ‘H’. Oxygen is represented ‘O’. Nitrogen is represented by ‘N’. Iron is represented by ‘Fe’.

	Hydrogen		Carbon		Oxygen
	Phosphorus		Sulphur		Iron
	Copper		Lead		Silver
	Gold		Platina		Mercury

His work proved as boon to science. For his marvelous work Berzilius, together with John Dalton, Antoine Lavoisier, and Robert Boyle is considered as the Father of Modern Chemistry.

Symbol and Name of some elements					
Element	Symbol	Element	Symbol	Element	Symbol
Hydrogen	H	Sodium	Na	Cromium	Cr
Helium	He	Magnesium	Mg	Mangese	Mn
Lithium	Li	Aluminium	Al	iron	Fe
Beryllium	Be	Silicon	Si	Cobalt	Co
Boron	B	Phosphorous	P	Nickel	Ni
Carbon	C	Sulphur	S	Copper	Cu
Nitrogen	N	Chlorine	Cl	Zinc	Zn
Oxygen	O	Argon	Ar	Silver	Ag
Fluorine	F	Potassium	K	Gold	Au
Neon	Ne	Calcium	Ca	Mercury	Hg

Symbol of many elements are taken from their English name, while symbol of many elements are taken from their Greek or Latin names.

Symbol of some element which are derived from their Latin name

Several elements are named after the place where they discovered, such as ‘Copper’ which was taken from Cyprus. Some elements are named after their colour, such as ‘Gold’ which means yellow.

Symbols of some elements taken from their Latin Name		
English Name of Elements	Symbol	Latin Name of Elements
Sodium	Na	Natrium
Potassium	K	Kalium
Iron	Fe	Ferrum
Copper	Cu	Cuprum
Silver	Ag	Argentum
Gold	Au	Aurum
Mercury	Hg	Hydragyrum
Lead	Pb	Plumbum
Tin	Sn	Stannum

ATOMIC MASS

Mass of atom is called atomic mass. Since, atoms are very small consequently actual mass of an atom is very small. For example the actual mass of one atom of hydrogen is equal to 1.673×10^{-24} g. This is equal to 0.0000000000000000000000001673 gram. To deal with such small number is very difficult. Thus for convenience relative atomic mass is used.

Carbon-12 is considered as unit to calculate atomic mass. Carbon-12 is an isotope of carbon. The relative mass of all atoms are found with respect to C-12.

One atomic mass = 1/12 of the mass of one atom of C-12.

$$\text{This means atomic mass unit} = \frac{1}{12} \text{ th of Carbon - 12}$$

Atomic Mass of some elements					
Element	Symbol	Atomic Mass	Element	Symbol	Atomic Mass
Hydrogen	H	1u	Sodium	Na	23u
Helium	He	4u	Magnesium	Mg	24u
Lithium	Li	7u	Aluminium	Al	27u
Beryllium	Be	9u	Silicon	Si	28u
Boron	B	11u	Phosphorous	P	31u
Carbon	C	12u	Sulphur	S	32u
Nitrogen	N	14u	Chlorine	Cl	35u
Oxygen	O	16u	Potassium	K	39u
Fluorine	F	19u	Calcium	Ca	40u
Neon	Ne	20u	Iron	Fe	56

Thus atomic mass is the relative atomic mass of an atom with respect to 1/12 th of the mass of carbon-12 atom. ‘amu’ is the abbreviation of Atomic mass unit, but now it is denoted just by ‘u’.

The atomic mass of hydrogen atom = 1u.

This means one hydrogen atom is 1 times heavier than $1/12^{\text{th}}$ of the carbon atom.

The atomic mass of oxygen is 16u, this means one atom of oxygen is 16 times heavier than $1/12^{\text{th}}$ of carbon atom.

Absolute mass or Actual atomic mass:

It is found that, the actual atomic mass of a carbon-12 atom is equal to 1.9926×10^{-23} g.

$$\text{Therefore, } 1\text{u} = \frac{1.9926 \times 10^{-23}}{12} \text{g} = 1.6605 \times 10^{-24} \text{g}$$

Thus by multiplying the relative atomic mass with 1.6605×10^{-24} g we can get the absolute or actual mass of an atom.

Example - 1 - Find the absolute mass oxygen.

Solution:

The atomic mass of oxygen is 16u

We know $1\text{u} = 1.6605 \times 10^{-24}$ g

$$\begin{aligned}\text{Therefore, Absolute mass of oxygen} &= 1.6605 \times 10^{-24} \times 16 \text{ g} \\ &= 26.568 \times 10^{-24} = 2.6568 \times 10^{-25} \text{ g}\end{aligned}$$

Example – 2 – Find the absolute mass of Sodium.

Solution:

The atomic mass of Sodium is 23u

We know $1\text{u} = 1.6605 \times 10^{-24}$ g

$$\begin{aligned}\text{Therefore, Absolute mass of Sodium} &= 1.6605 \times 10^{-24} \times 23 \text{ g} \\ &= 38.191 \times 10^{-24} = 3.8191 \times 10^{-25} \text{ g}\end{aligned}$$

Example – 3 – Calculate the absolute mass of hydrogen atom.

Solution:

The atomic mass of hydrogen is 1u

We know $1\text{u} = 1.6605 \times 10^{-24}$ g

$$\text{Therefore, Absolute mass of hydrogen} = 1.6605 \times 10^{-24} \times 1 \text{ g} = 1.6605 \times 10^{-24} \text{ g}$$

Example – 4 - Calculate the absolute or actual mass of Nitrogen atom.

Solution:

The atomic mass of Nitrogen is 14u

We know $1\text{u} = 1.6605 \times 10^{-24}$ g

$$\begin{aligned}\text{Therefore, Absolute mass of hydrogen} &= 1.6605 \times 10^{-24} \times 14 \text{ g} \\ &= 23.247 \times 10^{-24} = 2.3247 \times 10^{-25} \text{ g}\end{aligned}$$

EXISTENCE OF ATOMS

Atoms of most of the elements exist in the form of molecule or ion, since they are most reactive. For example, hydrogen, oxygen, chlorine, etc. However, atoms of some elements, which are non-reactive, exist in free-state in nature. For example helium, neon, argon, etc.

Usually atoms are exist in following two forms -

- In the form of molecules
- In the form of ions

INTEXT QUESTIONS PAGE NO. 35

Q1. Define the atomic mass unit.

Answer:

Mass unit equal to exactly one-twelfth $\left(\frac{1}{12^{th}}\right)$ the mass of one atom of carbon-12 is called one atomic mass unit. It is written as ‘u’.

Q2. Why is it not possible to see an atom with naked eyes?

Answer:

The size of an atom is so small that it is not possible to see it with naked eyes. Also, the atom of an element does not exist independently.

MOLECULE

Molecule is the smallest particle of a compound.

Atoms exist in free states in the form of molecule.

- A molecule may be formed by the combination of two or more similar atoms of an element, such as oxygen molecule is formed by the combination of two oxygen atoms, molecule of hydrogen which is formed by the combination of two hydrogen atoms.
- Molecules may be formed by the combination of atoms of two or more different elements. For example molecule of water. It is formed by the combination of two atoms of hydrogen and one atom of oxygen. Molecule of Nitric oxide or nitrogen monoxide. It is formed by the combination of one nitrogen atom and one oxygen atom.
- A molecule takes part in chemical reaction.

Most of the atoms exist in the form of molecule. Molecules are formed by the combination of two or more elements.

Example: Molecule of hydrogen (H_2), Molecule of oxygen (O_2), Molecule of nitrogen (N_2), etc.

- Molecules of elements
- Molecules of Compounds

MOLECULES OF ELEMENTS

When two or more atoms of same element combine to form a molecule these are called molecules of element.

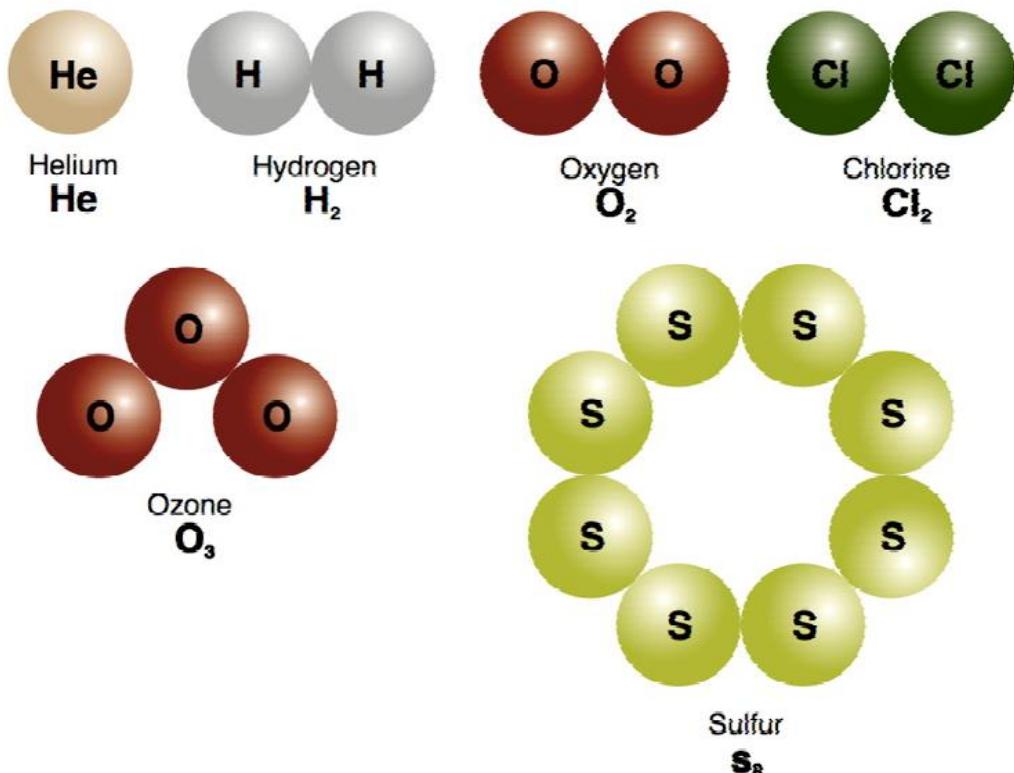
Example:

Hydrogen molecule (H_2). Hydrogen molecule (H_2). Molecule of hydrogen is formed by the combination of two hydrogen atoms.

Oxygen molecule (O_2). Molecule of oxygen is formed by the combination of two oxygen atoms.

Sulphur molecule (S_8). Molecule of sulphur is formed by the combination of eight sulphur atoms.

Phosphorous molecule (P_4). Molecule of phosphorous is formed by the combination of four phosphorous atoms.



Molecules of some non-reactive elements are formed by single atom. For example – helium, neon, argon, etc. molecules: Molecules of metals formed as big cluster of atoms. They are represented by their symbols simply. For example: Iron (Fe), Copper (Cu), Zinc (Zn), etc. These molecules are known as giant molecules.

Carbon is a non-metal, but it also exists as giant molecule and represented by its symbol ‘C’.

ATOMICITY

Monoatomic:

When molecule is formed by single atom only, it is called monoatomic molecule. Generally noble gas forms monoatomic molecules. For example: Helium (He), Neon (Ne), Argon (Ar), Kr (Krypton), Xenon (Xe), Randon (Rn).

Diatomc

When molecule is formed by the combination of two atoms it is called diatomic molecule. For example: Hydrogen (H₂), Oxygen (O₂), Nitrogen (N₂), Chlorine (Cl₂), etc.

Triatomic

When molecule is formed by the combination of three atoms it is called triatomic molecule. For example: molecule of ozone (O₃)

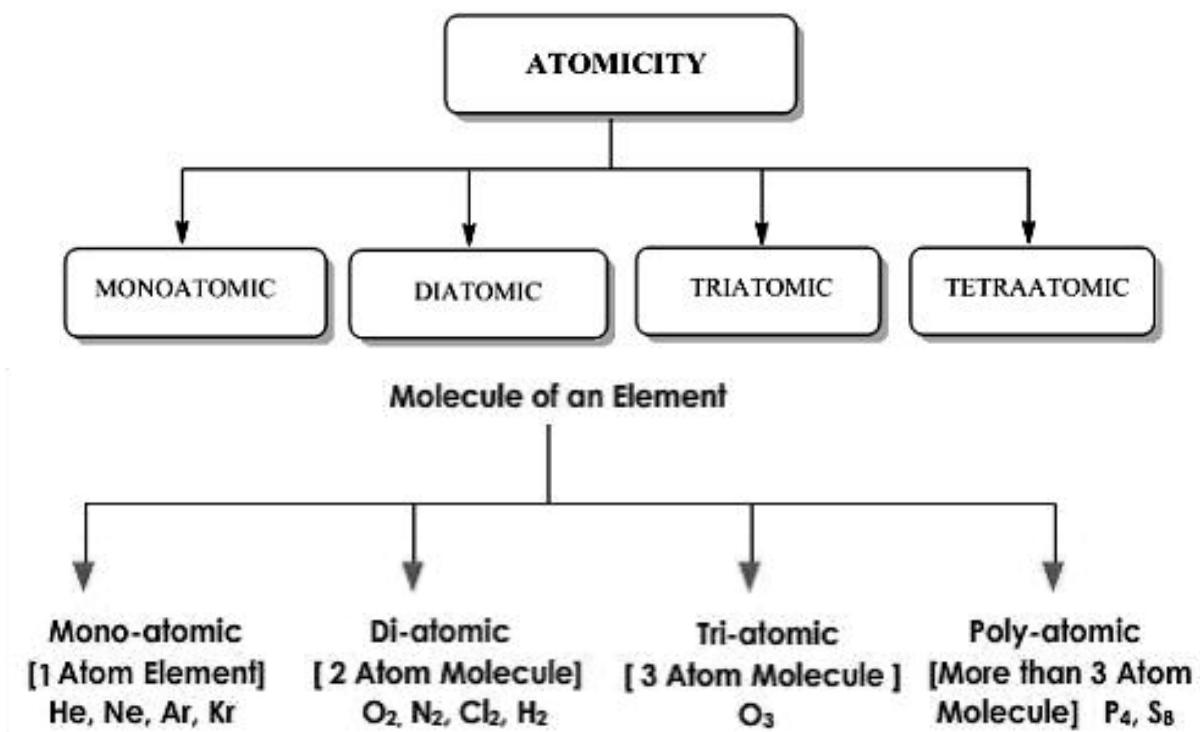
Tetra-atomic

When molecule is formed by the combination of four atoms it is called tetra-atomic molecule. For example: Phosphorous molecule (P₄)

Polyatomic

When molecule is formed by the combination of more than two atoms, it is called polyatomic molecule. For example: Sulphur molecule (S₈)

Atomicity of some elements		
Name	Atomicity	Formula
Argon	Monoatomic	Ar
Helium	Monoatomic	He
Oxygen	Diatomeric	O ₂
Hydrogen	Diatomeric	H ₂
Nitrogen	Diatomeric	N ₂
Chlorine	Diatomeric	Cl ₂
Phosphorous	Tetra-atomic	P ₄
Sulphur	Poly-atomic	S ₈

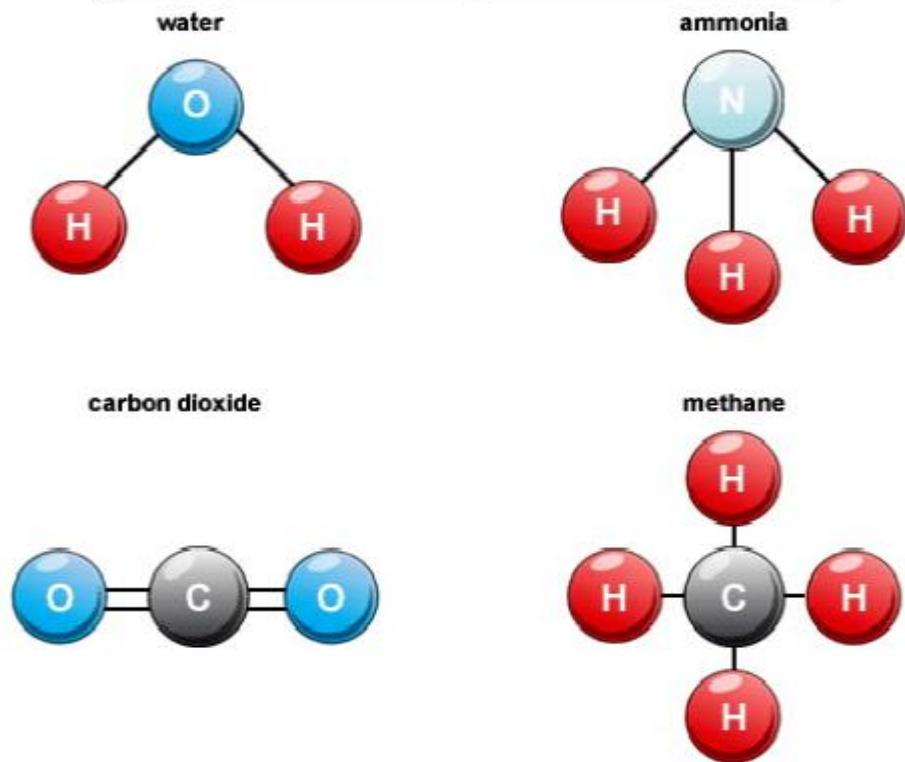


MOLECULES OF COMPOUNDS

When molecule is formed by the combination of two or more atoms of different elements, it is called the molecule of compound.

Example: Molecule of water (H₂O). Molecule of water is formed by the combination of two hydrogen and one oxygen atoms.

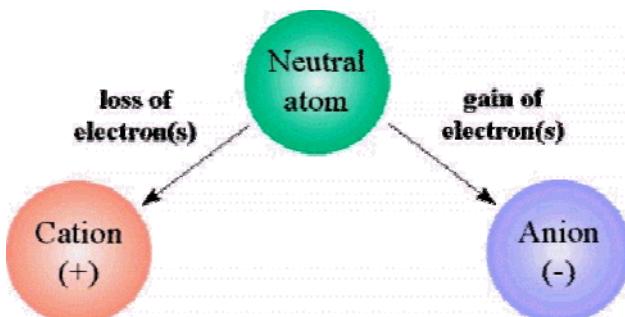
Molecules of some compounds	
Compound	Combining Elements
Water (H_2O)	Hydrogen, Oxygen
Ammonia (NH_3)	Nitrogen, hydrogen
Carbon dioxide(CO_2)	Carbon, oxygen
Hydrogen Chloride (HCl)	Hydrogen, Chlorine
Methane (CH_4)	Carbon, Hydrogen
Ehtane (C_2H_6)	Carbon, hydrogen
Sodium chloride ($NaCl$)	Sodium, chlorine.
Copper oxide (CuO)	Copper and oxygen



IONS

Atoms of several elements exists in the form of ion. Atoms or molecule with negative or positive charge over them are called ions.

For example: Sodium ion (Na^+), postassium ion (K^+), Chlorine ion (Cl^-), Fluoride ion (F^-)etc.



Cations:

Ions having positive charge over them are called cations.

For example: sodium ion (Na^+), potassium ion (K^+), etc

Anions:

Ions having negative charge over them are called anions.

For example: Chloride ion (Cl^-), Fluoride ion (F^-), etc

Monoatomic ions:

Ions formed by one atom are called monoatomic ions.

For example: sodium ion (Na^+), potassium ion (K^+), Chloride ion (Cl^-), Fluoride ion (F^-), etc.

Polyatomic ions:

Ions formed by two or more atoms are called polyatomic ions. These are group of atoms of different elements which behave as single units, and are known as polyatomic ions.

For example: Ammonium ion (NH_4^+), Hydroxide ion (OH^-), etc

Some Common ions					
Cations		Anions		Polyatomic ions	
Lithium ion	Li^+	Chloride ion	Cl^-	Hydroxide	OH^-
Sodium ion	Na^+	Fluorine	F^-	Ammonium	NH_4^+
Potassium ion	K^+	Iodide	I^-	Nitrate	NO_3^-
Silver ion	Ag^+	Hydride	H^-	Bicarbonate or Hydrogen carbonate	HCO_3^-
Copper ion	Cu^+	Oxide ion	O^{2-}		
Hydrogen ion	H^+	Sulphide	S^{2-}		
Magnesium ion	Mg^{++}	Nitride	N^{3-}	Sulphate	SO_4^{2-}
Calcium ion	Ca^{++}			Carbonate	CO_3^{2-}
Iron ion	Fe^{++}			Sulphite	SO_3^{2-}
Zinc ion	Zn^{++}			Phosphate	PO_4^{3-}
Copper ion	Cu^{++}				
Aluminium ion	Al^{+++}				

WRITING CHEMICAL FORMULA

Chemical formula of the compound is the symbolic representation of its composition. To write chemical formula of a compound, symbols and valencies of constituent elements must be known. The valency of atom of an element can be thought of as hands or arms of that atom.

Points to remember

- The symbols or formulas of the component radicals of the compound are written side by side.
- Positive radicals are written left and negative radicals on the right.
- The valencies of the radicals are written below the respective symbols.
- The criss-cross method is applied to exchange the numerical value of valency of each radical. It is written as subscript of the other radical.

- The radical is enclosed in a bracket and the subscript is placed outside the lower right corner.
- The common factor is removed.
- If the subscript of the radical is one, it is omitted.

The rules that you have to follow while writing a chemical formula are as follows:

- the valencies or charges on the ion must balance.
- when a compound consists of a metal and a non-metal, the name or symbol of the metal is written first. For example: calcium oxide (CaO), sodium chloride (NaCl), iron sulphide (FeS), copper oxide (CuO) etc., where oxygen, chlorine, sulphur are non-metals and are written on the right, whereas calcium, sodium, iron and copper are metals, and are written on the left.
- in compounds formed with polyatomic ions, the ion is enclosed in a bracket before writing the number to indicate the ratio.

The simplest compounds, which are made up of two different elements are called binary compounds. While writing the chemical formulae for compounds, we write the constituent elements and their valencies as shown below. Then we must crossover the valencies of the combining atoms.

The formulae of ionic compounds are simply the whole number ratio of the positive to negative ions in the structure. For magnesium chloride, we write the symbol of cation (Mg^{2+}) first followed by the symbol of anion (Cl^-). Then their charges are criss-crossed to get the formula.

EXAMPLES

Formula of Sodium oxide

write symbol	Na	O
Remember charge	+1	-2
Cross the charge	Na_2O	

Formula of Sodium hydroxide

write symbol	Na	OH
Remember charge	+1	-1
Cross the charge	NaOH	

Formula of Sodium Chloride

write symbol	Na	Cl
Remember charge	+1	-1
Cross the charge	NaCl	

Formula of Calcium chloride

write symbol	Ca	Cl
Remember charge	+2	-1
Cross the charge	CaCl_2	

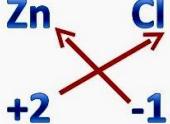
Formula of Potassium hydroxide

write symbol	K	OH
Remember charge	+1	-1
Cross the charge	KOH	

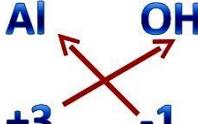
Formula of Zinc hydroxide

write symbol	Zn	OH
Remember charge	+2	-1
Cross the charge	Zn(OH)_2	

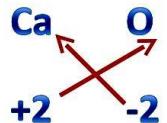
Formula of Zinc chloride

write symbol	Zn Cl +2 -1 
Remember charge	
Cross the charge	$ZnCl_2$

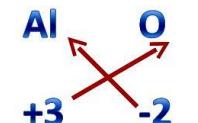
Formula of Aluminium hydroxide

write symbol	Al OH +3 -1 
Remember charge	
Cross the charge	$Al(OH)_3$

Formula of Calcium oxide

write symbol	Ca O +2 -2 
Remember charge	
Cross the charge	CaO

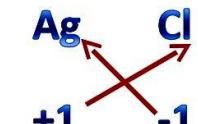
Formula of Aluminium oxide

write symbol	Al O +3 -2 
Remember charge	
Cross the charge	Al_2O_3

Formula of Silver oxide

write symbol	Ag O +1 -2 
Remember charge	
Cross the charge	Ag_2O

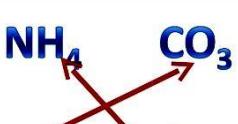
Formula of Silver chloride

write symbol	Ag Cl +1 -1 
Remember charge	
Cross the charge	$AgCl$

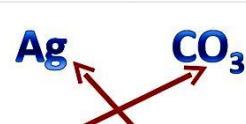
Formula of Ammonium Chloride

write symbol	NH ₄ Cl +1 -1 
Remember charge	
Cross the charge	NH_4Cl

Formula of Ammonium carbonate

write symbol	NH ₄ CO ₃ +1 -2 
Remember charge	
Cross the charge	$(NH_4)_2CO_3$

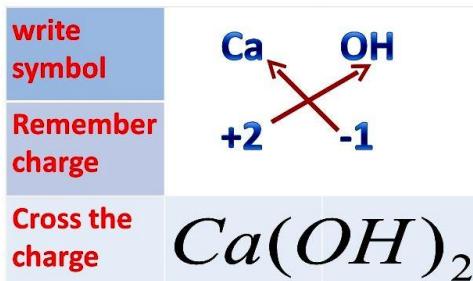
Formula of Silver Carbonate

write symbol	Ag CO ₃ +1 -2 
Remember charge	
Cross the charge	Ag_2CO_3

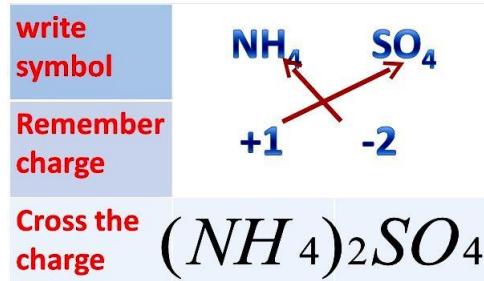
Formula of Silver Sulphate

write symbol	Ag SO ₄ +1 -2 
Remember charge	
Cross the charge	Ag_2SO_4

Formula of Calcium hydroxide



Formula of Ammonium sulphate



INTEXT QUESTIONS PAGE NO. 39

Q1. Write down the formulae of

- (i) sodium oxide
- (ii) aluminium chloride
- (iii) sodium sulphide
- (iv) magnesium hydroxide

Answer:

- (i) Sodium oxide → Na_2O
- (ii) Aluminium chloride → AlCl_3
- (iii) Sodium sulphide → Na_2S
- (iv) Magnesium hydroxide → $\text{Mg}(\text{OH})_2$

Q2. Write down the names of compounds represented by the following formulae:

- (i) $\text{Al}_2(\text{SO}_4)_3$
- (ii) CaCl_2
- (iii) K_2SO_4
- (iv) KNO_3
- (v) CaCO_3 .

Answer:

- (i) $\text{Al}_2(\text{SO}_4)_3$ → Aluminium Sulphate
- (ii) CaCl_2 → Calcium Chloride
- (iii) K_2SO_4 → Potassium sulphate
- (iv) KNO_3 → Potassium nitrate
- (v) CaCO_3 . → Calcium carbonate

Q3. What is meant by the term chemical formula?

Answer:

The chemical formula of a compound means the symbolic representation of the composition of a compound. From the chemical formula of a compound, we can know the number and kinds of atoms of different elements that constitute the compound.

For example, from the chemical formula CO_2 of carbon dioxide, we come to know that one carbon atom and two oxygen atoms are chemically bonded together to form one molecule of the compound, carbon dioxide.

Q4. How many atoms are present in a

- (i) H_2S molecule and
- (ii) PO_4^{3-} ion?

Answer:

- (i) In an H_2S molecule, three atoms are present; two of hydrogen and one of sulphur.
- (ii) In a PO_4^{3-} ion, five atoms are present; one of phosphorus and four of oxygen.

MOLECULAR MASS

Atomic mass: The atomic mass of an element is the mass of one atom of that element in atomic mass units or (u).

Atomic mass unit (amu): 1/12th of the mass of an atom of carbon-12 is called atomic mass unit. It is a unit of mass used to express atomic masses and molecular masses.

Molar mass: The molar mass of an element is equal to the numerical value of the atomic mass. However, in case of molar mass, the units change from 'u' to 'g'. The molar mass of an atom is also known as gram atomic mass.

For example, the atomic mass of carbon = 12 atomic mass units. So, the gram atomic mass of carbon = 12 grams.

Molecular mass of the molecule: The sum of the atomic masses of all the atoms in a molecule of a substance is called the molecular mass of the molecule.

Molecular mass - calculation: Generally we use relative atomic masses of atoms for calculating the molecular mass of 1 mole of any molecular or ionic substances.

Example: Molecular mass of H₂SO₄

Atomic mass of Hydrogen = 1

Atomic mass of sulphur = 32

Atomic mass of oxygen = 16

Molecular mass of H₂SO₄ = 2(Atomic mass of Hydrogen) + 1 (Atomic mass of sulphur) + 4 (Atomic mass of oxygen) = 2×1 + 32 + 4×16 = 98 u.

Calculation of molecular mass of hydrogen chloride:

Atomic mass of hydrogen + Atomic mass of chlorine = 1 + 35.5 = 36.5 u.

FORMULA UNIT MASS

The formula unit mass of a substance is the sum of the atomic masses of all atoms in a formula unit of a compound. The term 'formula unit' is used for those substances which are made up of ions.

Formula unit mass of NaCl: 1 x Atomic mass of Na + 1 x Atomic mass of Cl
1 x 23 + 1 x 35.5 = 58.5 atomic mass units.

Formula unit mass of ZnO:

= 1 x Atomic mass of Zn + 1 x Atomic mass O

= 1 x 65 + 1 x 16 = 81 u.

INTEXT QUESTIONS PAGE NO. 40

Q1. Calculate the molecular masses of H₂, O₂, Cl₂, CO₂, CH₄, C₂H₆, C₂H₄, NH₃, CH₃OH.

Answer:

Molecular mass of O₂ = 2 × Atomic mass of O = 2 × 16 = 32 u

Molecular mass of Cl₂ = 2 × Atomic mass of Cl = 2 × 35.5 = 71 u

Molecular mass of CO₂ = Atomic mass of C + 2 × Atomic mass of O = 12 + 2 × 16 = 44 u

Molecular mass of CH₄ = Atomic mass of C + 4 × Atomic mass of H = 12 + 4 × 1 = 16 u

Molecular mass of C₂H₆ = 2 × Atomic mass of C + 6 × Atomic mass of H = 2 × 12 + 6 × 1 = 30 u

Molecular mass of C₂H₄ = 2 × Atomic mass of C + 4 × Atomic mass of H = 2 × 12 + 4 × 1 = 28 u

Molecular mass of NH_3 = Atomic mass of N + $3 \times$ Atomic mass of H = $14 + 3 \times 1 = 17$ u
Molecular mass of CH_3OH = Atomic mass of C + $4 \times$ Atomic mass of H + Atomic mass of O
 $= 12 + 4 \times 1 + 16 = 32$ u

Q2. Calculate the formula unit masses of ZnO , Na_2O , K_2CO_3 , given atomic masses of Zn = 65 u, Na = 23 u, K = 39 u, C = 12 u, and O = 16 u.

Answer:

Formula unit mass of ZnO = Atomic mass of Zn + Atomic mass of O = $65 + 16 = 81$ u

Formula unit mass of Na_2O = $2 \times$ Atomic mass of Na + Atomic mass of O
 $= 2 \times 23 + 16 = 62$ u

Formula unit mass of K_2CO_3 = $2 \times$ Atomic mass of K + Atomic mass of C + $3 \times$ Atomic mass of O = $2 \times 39 + 12 + 3 \times 16 = 138$ u

MOLE CONCEPT

Mole: Mole is the measurement in chemistry. It is used to express the amount of a chemical substance.

One mole is defined as the amount of substance of a system which contains as many entities like, atoms, molecules and ions as there are atoms in 12 grams of carbon - 12".

Avogadro number: The number of the particles present in one mole of any substance is equal to 6.022×10^{23} . This is called avogadro's number or avogadro's constant.

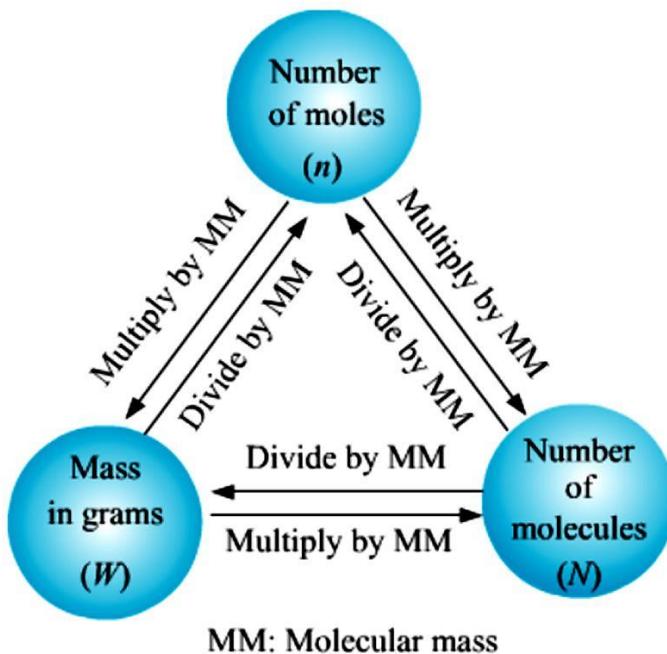
Number of particles in 1 mole:

1 mole of hydrogen atoms represents 6.022×10^{23} hydrogen atoms.

1 mole of hydrogen molecules represents 6.022×10^{23} hydrogen molecules.

1 mole of water molecules represents 6.022×10^{23} water molecules.

Conversion of moles to mass and vice-versa



The key concept used in these kind of problems is that a mole of any substance contains gram formula mass or molecular mass of that substance i.e. molecular mass of Hydrogen is 2 a.m.u.

so mass of 1 mole of hydrogen which is also known as molar mass will be 2 gram. Similarly if we have 2 moles of hydrogen, it will weigh 2×2 grams which is equal to 4 grams.

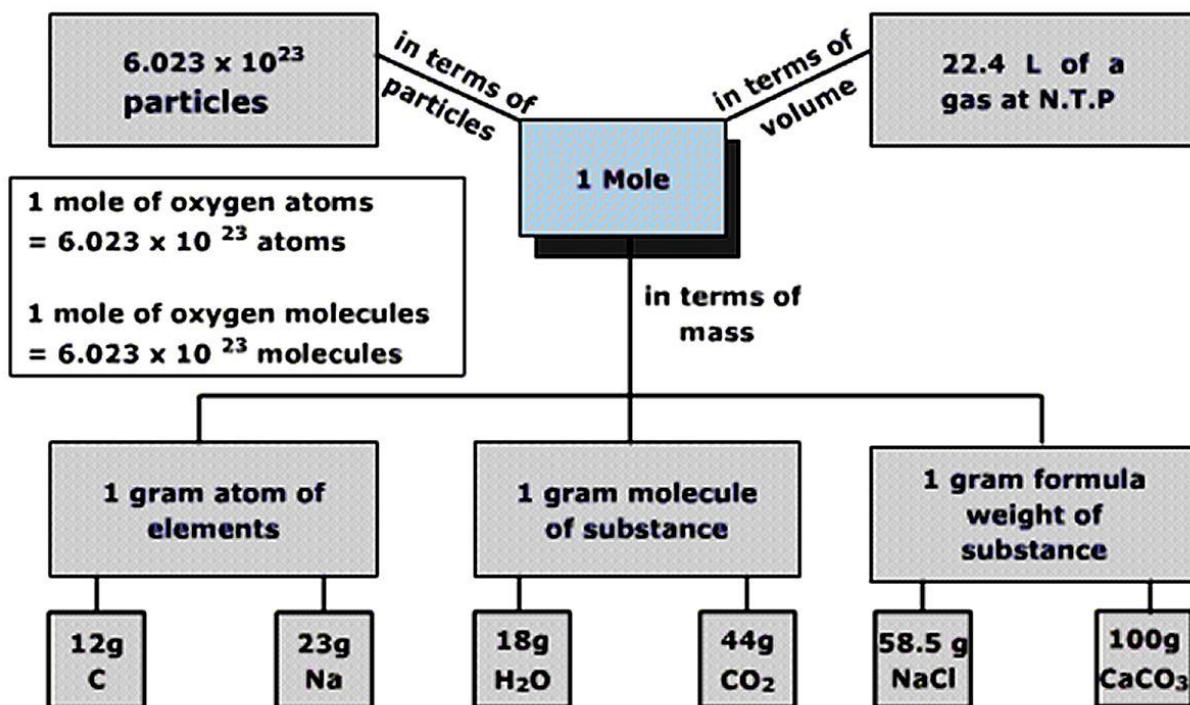
MOLE CONCEPT CALCULATION

This is the most basic and the most used calculation that a student comes across while solving a mole concept problem. Most of the times, moles or number of atoms or molecules are given in the question and the mass is needed to be calculated. In that case proceed as shown in the above example. In rest of the cases, mass will be given and moles or number will be needed to be calculated. In those questions also, proceed by:

STEP 1:- Establishing relationship between molar mass and the number (N_A) or moles of that particular entity (atom, molecule or ion).

STEP 2:- Use unitary method to calculate what is asked in the question.

NOTE: – When we say oxygen gas weighs 32 gram then we mean to say that 1 mole of oxygen molecule (O_2) weighs 32 grams and not 1 mole of oxygen atom which is O. This is because in natural form, oxygen exists as O_2 molecule.



PROBLEMS (BASED ON MOLE CONCEPT)

1. When the mass of the substance is given:

$$\text{Number of moles} = \frac{\text{given mass}}{\text{atomic mass}}$$

Example 1. Calculate the number of moles in 81g of aluminium

$$\text{Number of moles} = \frac{\text{given mass}}{\text{atomic mass}} = \frac{81}{27} = 3 \text{ moles of aluminium}$$

Example 2. Calculate the mass of 0.5 mole of iron

Solution: mass = atomic mass x number of moles
= $55.9 \times 0.5 = 27.95 \text{ g}$

2. Calculation of number of particles when the mass of the substance is given:

$$\text{Number of particles} = \frac{\text{Avogadro number} \times \text{given mass}}{\text{gram molecular mass}}$$

Example 1. Calculate the number of molecules in 11g of CO₂

Solution: gram molecular mass of CO₂ = 44g

$$\text{Number of particles} = \frac{6.023 \times 10^{23} \times 11}{44} = 1.51 \times 10^{23} \text{ molecules}$$

3. Calculation of mass when number of particles of a substance is given:

$$\text{Mass of a substance} = \frac{\text{gram molecular mass} \times \text{number of particles}}{6.023 \times 10^{23}}$$

Example 1. Calculate the mass of 18.069×10^{23} molecules of SO₂

Sol: Gram molecular mass SO₂ = 64g

$$\text{Mass of SO}_2 = \frac{64 \times 18.069 \times 10^{23}}{6.023 \times 10^{23}} = 192 \text{ g}$$

Example 2. Calculate the mass of glucose in 2×10^{24} molecules

Gram molecular mass of glucose = 180g

$$\text{Mass of glucose} = \frac{180 \times 2 \times 10^{23}}{6.023 \times 10^{23}} = 597.7 \text{ g}$$

4. Calculation of number of moles when you are given number of molecules:

$$\text{Number of moles of atom} = \frac{\text{Number of molecules}}{\text{Avogadro Number}}$$

Example 1. Calculate number of moles in 12.046×10^{22} atoms of copper

$$\text{Number of moles of atom} = \frac{\text{Number of molecules}}{\text{Avogadro Number}} = \frac{12.046 \times 10^{22}}{6.023 \times 10^{23}} = 0.2 \text{ moles}$$

INTEXT QUESTIONS PAGE NO. 42

Q1. If one mole of carbon atoms weighs 12 gram, what is the mass (in gram) of 1 atom of carbon?

Answer:

One mole of carbon atoms weighs 12 g (Given)

i.e., mass of 1 mole of carbon atoms = 12 g

Then, mass of 6.022×10^{23} number of carbon atoms = 12 g

$$\text{Therefore, mass of 1 atom of carbon} = \frac{12}{6.022 \times 10^{23}} \text{ g} = 1.9926 \times 10^{-23} \text{ g}$$

Q2. Which has more number of atoms, 100 grams of sodium or 100 grams of iron (given, atomic mass of Na = 23 u, Fe = 56 u)?

Answer:

Atomic mass of Na = 23 u (Given)

Then, gram atomic mass of Na = 23 g

Now, 23 g of Na contains $= 6.022 \times 10^{23}$ number of atoms

Thus, 100 g of Na contains $= \frac{6.022 \times 10^{23}}{23} \times 100 = 2.6182 \times 10^{24}$ number of atoms

Again, atomic mass of Fe = 56 u(Given)

Then, gram atomic mass of Fe = 56 g

Now, 56 g of Fe contains $= 6.022 \times 10^{23}$ number of atoms

Thus, 100 g of Fe contains $= \frac{6.022 \times 10^{23}}{56} \times 100 = 1.0753 \times 10^{24}$ number of atoms

Therefore, 100 grams of sodium contain more number of atoms than 100 grams of iron.

EXERCISE QUESTIONS PAGE NO. 43, 44

Q1. A 0.24 g sample of compound of oxygen and boron was found by analysis to contain 0.096 g of boron and 0.144 g of oxygen. Calculate the percentage composition of the compound by weight.

Answer:

Mass of boron = 0.096 g (Given)

Mass of oxygen = 0.144 g (Given)

Mass of sample = 0.24 g (Given)

Thus, percentage of boron by weight in the compound $= \frac{0.096}{0.24} \times 100 = 40\%$

And, percentage of oxygen by weight in the compound $= \frac{0.144}{0.24} \times 100 = 60\%$

Q 2. When 3.0 g of carbon is burnt in 8.00 g oxygen, 11.00 g of carbon dioxide is produced. What mass of carbon dioxide will be formed when 3.00 g of carbon is burnt in 50.00 g of oxygen? Which law of chemical combination will govern your answer?

Answer:

Carbon + Oxygen \rightarrow Carbon dioxide

3 g of carbon reacts with 8 g of oxygen to produce 11 g of carbon dioxide.

If 3 g of carbon is burnt in 50 g of oxygen, then 3 g of carbon will react with 8 g of oxygen.

The remaining 42 g of oxygen will be left un-reactive.

In this case also, only 11 g of carbon dioxide will be formed.

The above answer is governed by the law of constant proportions.

Q 3. What are polyatomic ions? Give examples.

Answer:

A polyatomic ion is a group of atoms carrying a charge (positive or negative). For example, ammonium ion (NH_4^+), hydroxide ion (OH^-), carbonate ion (CO_3^{2-}), sulphate ion(SO_4^{2-})

Q 4. Write the chemical formulae of the following.

(a) Magnesium chloride

(b) Calcium oxide

(c) Copper nitrate

(d) Aluminium chloride

(e) Calcium carbonate.

Answer:

(a) Magnesium chloride $\rightarrow \text{MgCl}_2$

- (b) Calcium oxide → CaO
- (c) Copper nitrate → Cu(NO₃)₂
- (d) Aluminium chloride → AlCl₃
- (e) Calcium carbonate → CaCO₃

Q 5. Give the names of the elements present in the following compounds.

- (a) Quick lime
- (b) Hydrogen bromide
- (c) Baking powder
- (d) Potassium sulphate.

Answer:

- (a) Quick lime

Chemical formula : CaO

Elements present: Calcium, Oxygen

- (b) Hydrogen bromide

Chemical formula : HBr

Elements present: Hydrogen, Bromine

- (c) Baking powder

Chemical formula : NaHCO₃

Elements present: Sodium, Hydrogen, Carbon, Oxygen

- (d) Potassium sulphate.

Chemical formula : K₂SO₄

Elements present: Potassium, Sulphur, Oxygen

Q 6. Calculate the molar mass of the following substances.

- (a) Ethyne, C₂H₂
- (b) Sulphur molecule, S₈
- (c) Phosphorus molecule, P₄ (Atomic mass of phosphorus = 31)
- (d) Hydrochloric acid, HCl
- (e) Nitric acid, HNO₃

Answer:

(a) Molar mass of ethyne, C₂H₂ = 2 × 12 + 2 × 1 = 26 g

(b) Molar mass of sulphur molecule, S₈ = 8 × 32 = 256 g

(c) Molar mass of phosphorus molecule, P₄ = 4 × 31 = 124 g

(d) Molar mass of hydrochloric acid, HCl = 1 + 35.5 = 36.5 g

(e) Molar mass of nitric acid, HNO₃ = 1 + 14 + 3 × 16 = 63 g

Q 7. What is the mass of _____

- (a) 1 mole of nitrogen atoms?
- (b) 4 moles of aluminium atoms (Atomic mass of aluminium = 27)?
- (c) 10 moles of sodium sulphite (Na₂SO₃)?

Answer:

(a) The mass of 1 mole of nitrogen atoms is 14 g.

(b) The mass of 4 moles of aluminium atoms is (4 × 27) g = 108 g

(c) The mass of 10 moles of sodium sulphite (Na₂SO₃) is

10 × [2 × 23 + 32 + 3 × 16] g = 10 × 126 g = 1260 g

Q 8. Convert into mole.

- (a) 12 g of oxygen gas
- (b) 20 g of water
- (c) 22 g of carbon dioxide.

Answer:

(a) 32 g of oxygen gas = 1 mole

$$\text{Then, } 12 \text{ g of oxygen gas} = \frac{12}{32} \text{ mole} = 0.375 \text{ mole}$$

(b) 18 g of water = 1 mole

$$\text{Then, } 20 \text{ g of water} = \frac{20}{18} \text{ mole} = 1.11 \text{ moles (approx)}$$

(c) 44 g of carbon dioxide = 1 mole

$$\text{Then, } 22 \text{ g of carbon dioxide} = \frac{22}{44} \text{ mole} = 0.5 \text{ mole}$$

Q 9. What is the mass of:

(a) 0.2 mole of oxygen atoms?

(b) 0.5 mole of water molecules?

Answer:

(a) Mass of one mole of oxygen atoms = 16 g

Then, mass of 0.2 mole of oxygen atoms = $0.2 \times 16 \text{ g} = 3.2 \text{ g}$

(b) Mass of one mole of water molecule = 18 g

Then, mass of 0.5 mole of water molecules = $0.5 \times 18 \text{ g} = 9 \text{ g}$

Q 10. Calculate the number of molecules of sulphur (S_8) present in 16 g of solid sulphur.**Answer:**

1 mole of solid sulphur (S_8) = $8 \times 32 \text{ g} = 256 \text{ g}$

i.e., 256 g of solid sulphur contains = 6.022×10^{23} molecules

$$\text{Then, } 16 \text{ g of solid sulphur contains} = \frac{6.022 \times 10^{23}}{256} \times 16 \text{ molecules}$$
$$= 3.76 \times 10^{22} \text{ molecules (approx)}$$

Q 11. Calculate the number of aluminium ions present in 0.051 g of aluminium oxide.

(Hint: The mass of an ion is the same as that of an atom of the same element. Atomic mass of Al = 27 u)

Answer:

1 mole of aluminium oxide (Al_2O_3) = $2 \times 27 + 3 \times 16 = 102 \text{ g}$

i.e., 102 g of Al_2O_3 = 6.022×10^{23} molecules of Al_2O_3

$$\text{Then, } 0.051 \text{ g of } Al_2O_3 \text{ contains} = \frac{6.022 \times 10^{23}}{102} \times 0.051 \text{ molecules}$$
$$= 3.011 \times 10^{20} \text{ molecules of } Al_2O_3$$

The number of aluminium ions (Al^{3+}) present in one molecule of aluminium oxide is 2.

Therefore, the number of aluminium ions (Al^{3+}) present in 3.011×10^{20} molecules (0.051 g) of aluminium oxide (Al_2O_3) = $2 \times 3.011 \times 10^{20} = 6.022 \times 10^{20}$

ASSIGNMENT QUESTIONS SET – 1
CHAPTER – 3
ATOMS AND MOLECULES

Calculate the mass of one mole of these substances.

- | | | | | | |
|---|--|------------------------------------|--|---|--|
| 1. AlCl ₃ | 14. Ba(SCN) ₂ | 27. LiH | 40. Ba(BrO ₃) ₂ | 53. AlBr ₃ | 66. HCl |
| 2. TeF ₄ | 15. K ₂ S | 28. CO | 41. Hg ₂ Cl ₂ | 54. P ₂ O ₅ | 67. K ₂ SO ₄ |
| 3. PbS | 16. NH ₄ Cl | 29. SnI ₄ | 42. Cr ₂ (SO ₃) ₃ | 55. NH ₄ NO ₃ | 68. NaCl |
| 4. Cu ₂ O | 17. KH ₂ PO ₄ | 30. KOH | 43. Al(MnO ₄) ₃ | 56. Ba(OH) ₂ | 69. LiI |
| 5. AgI | 18. C ₂ H ₅ NBr | 31. K ₂ O | 44. CoSO ₄ | 57. PbSO ₄ | 70. Hg ₂ O |
| 6. N ₂ O | 19. Ba(ClO ₃) ₂ | 32. H ₂ SO ₄ | 45. Ca(NO ₃) ₃ | 58. Ba ₃ (PO ₄) ₂ | 71. HF |
| 7. MoCl ₅ | 20. Fe(OH) ₃ | 33. Hg ₃ N ₂ | 46. NaH ₂ PO ₄ | 59. NaC ₂ H ₃ O ₂ | 72. FeCl ₃ |
| 8. Hg ₂ Br ₂ | 21. (NH ₄) ₂ S | 34. SiF ₄ | 47. (NH ₄) ₃ PO ₄ | 60. Ba(OH) ₂ | 73. NaHSO ₄ |
| 9. Ta ₂ O ₅ | 22. CoCl ₂ | 35. NH ₄ OH | 48. KAl(SO ₄) ₂ | 61. NaHCO ₃ | 74. Ag ₂ O |
| 10. HgF ₂ | 23. KMnO ₄ | 36. N ₂ O ₅ | 49. Hg ₂ SO ₄ | 62. Al(OH) ₃ | 75. Pb(ClO ₂) ₂ |
| 11. KCl | 24. CaSO ₄ | 37. SnCrO ₄ | 50. Al ₂ (SO ₄) ₃ | 63. NH ₄ MnO ₄ | 76. CoF ₃ |
| 12. KF | 25. H ₂ CO ₃ | 38. Al ₂ O ₃ | 51. FePO ₄ | 64. Fe ₂ O ₃ | 77. Al(C ₂ H ₃ O ₂) ₃ |
| 13. ZnO | 26. CO ₂ | 39. CuCO ₃ | 52. Ca(C ₂ H ₃ O ₂) ₂ | 65. CaCO ₃ | |
| 78. Na ₂ Al ₂ (SO ₄) ₄ | | | | | |
| 79. (HOOCCH ₂) ₂ NCH ₂ CH ₂ N(CH ₂ COOH) ₂ | | | | | |
| 80. (NH ₄) ₂ CH(CH ₂) ₅ COOH | | | | | |

Answers (each answer has the units g/mol)

- | | | | | | |
|-------------|------------|-------------|--------------|-------------|-------------|
| 1. 133.34 | 14. 255.26 | 27. 7.95 | 40. 393.1314 | 53. 266.69 | 66. 36.461 |
| 2. 203.59 | 15. 110.26 | 28. 28.01 | 41. 472.09 | 54. 141.944 | 67. 174.25 |
| 3. 239.3 | 16. 53.49 | 29. 626.31 | 42. 344.1666 | 55. 80.04 | 68. 58.443 |
| 4. 143.09 | 17. 136.08 | 30. 56.106 | 43. 383.788 | 56. 171.34 | 69. 133.846 |
| 5. 234.77 | 18. 122.97 | 31. 94.20 | 44. 154.99 | 57. 303.26 | 70. 417.179 |
| 6. 44.01 | 19. 304.23 | 32. 98.07 | 45. 226.09 | 58. 601.93 | 71. 20.006 |
| 7. 273.20 | 20. 106.87 | 33. 629.78 | 46. 119.977 | 59. 82.03 | 72. 162.206 |
| 8. 560.98 | 21. 68.14 | 34. 104.08 | 47. 149.087 | 60. 171.34 | 73. 120.055 |
| 9. 441.89 | 22. 129.84 | 35. 35.046 | 48. 258.195 | 61. 84.007 | 74. 231.74 |
| 10. 238.59 | 23. 158.03 | 36. 108.01 | 49. 497.24 | 62. 78.00 | 75. 342.10 |
| 11. 74.55 | 24. 136.14 | 37. 234.68 | 50. 342.136 | 63. 136.97 | 76. 115.928 |
| 12. 58.10 | 25. 62.02 | 38. 101.96 | 51. 150.82 | 64. 159.69 | 77. 204.12 |
| 13. 81.38 | 26. 44.01 | 39. 123.555 | 52. 158.169 | 65. 100.09 | |
| 78. 484.173 | | | | | |
| 79. 292.246 | | | | | |
| 80. 164.248 | | | | | |

- 1.** Who established the two important laws of chemical combinations?
- 2.** What are the laws of chemical combinations?
- 3.** What is the law of conservation of mass?
- 4.** Give an example to show Law of conservation of mass applies to physical change also.
- 5.** Explain with example that law of conservation of mass is valid for chemical reactions.
- 6.** The 2.8 g of nitrogen gas was allowed to react with 0.6 g of hydrogen gas to produce 3.4 g of ammonia. Show that these observations are in agreement with the law of Conservation of mass.
- 7.** If 12 g of carbon is burnt in the presence of 32 g of oxygen, how much carbon dioxide will be formed?
- 8.** Who proposed Law of Definite Proportions (or Law of Constant Composition)?
- 9.** State Law of constant proportions. Explain with an example.
- 10.** Show that water illustrates the law of constant proportions.
- 11.** Hydrogen and oxygen combine in the ratio of 1:8 by mass to form water. What mass of oxygen gas would be required to react completely with 3 g of hydrogen gas?
- 12.** A 0.24 g sample of compound of oxygen and boron was found by analysis to contain 0.096 g of boron and 0.144 g of oxygen. Calculate the percentage composition of the compound by weight.
- 13.** When 3.0 g of carbon is burnt in 8.00 g oxygen, 11.00 g of carbon dioxide is produced. What mass of carbon dioxide will be formed when 3.00 g of carbon is burnt in 50.00 g of oxygen? Which law of chemical combination will govern your answer?
- 14.** Magnesium and oxygen combine in the ratio of 3 : 2 by mass to form magnesium oxide. How much oxygen is required to react completely with 12 g of magnesium?
- 15.** What are the postulates of Dalton's atomic theory?
- 16.** Which of the following statements is NOT true about an atom?
 - (a) Atoms are the building blocks from which molecules and ions are formed.
 - (b) Atoms cannot exist independently.
 - (c) Atoms are neutral in nature
 - (d) Atoms combine together to form matter that we can see, feel or touch.
- 17.** What is an atom?
- 18.** Why is it not possible to see an atom with naked eyes?
- 19.** Who proposed the chemical notation based on first two letters of the name of the element?
- 20.** Name the international organization who approves names of elements.
- 21.** What is the chemical symbol for iron?
- 22.** Name five elements have single letter chemical symbol.
- 23.** Name the element having following Latin names
 - (i) Stibium

- (ii) Cuprum
- (iii) Argentum
- (iv) Natrium
- (v) Stannum
- (vi) Wolfram
- (vii) plumbum
- (viii) Kalium

24. Write the chemical symbols of the following:

- (i) Gold
- (ii) Iron
- (iii) Chlorine
- (iv) Mercury

25. How will you define chemical symbol?

26. What is the significance of a chemical symbol?

27. Can atoms of an element exist independently? Give examples of elements which exist in atomic form. Give examples of elements that do not exist in atomic form.

28. Why do atoms of the most of the elements not exist independently?

29. Which element has the smallest atom in size?

30. What is the atomic mass unit?

31. Magnesium is two times heavier than C-12 atom, what shall be the mass of Mg atom in terms of atomic mass units? (Given mass of C-12 atom = 12u)

32. What is relative atomic mass of an element? How it is related to atomic mass unit?

33. Define molecule. What are its important properties?

34. Based on type of substance, how molecules are classified?

35. What is atomicity?

36. Based on atomicity, how molecules are categorized?

37. Give three examples of monoatomic molecules.

38. Give four examples of diatomic molecules.

39. (i) What is the chemical formula of Water molecule? (ii) What is its atomicity? (iii) Calculate the ratio of masses of atoms of elements present in water molecule. (iv) Calculate the ratio by number of atoms of elements present in water molecule.

40. What is an ion?

41. What are polyatomic ions? Give examples?

42. Give examples of triatomic molecules.

43. What is valency of an element?

44. What is meant by the term chemical formula?

45. Write down the formulae of

- (i) sodium oxide

- (ii) aluminium chloride
- (iii) sodium sulphide
- (iv) magnesium hydroxide

46. Write down the names of compounds represented by the following formulae:

- (i) $\text{Al}_2(\text{SO}_4)_3$
- (ii) CaCl_2
- (iii) K_2SO_4
- (iv) KNO_3
- (v) CaCO_3

47. Write the chemical formulae of the following. Also identify the ions present.

- (a) Magnesium chloride
- (b) Calcium oxide
- (c) Copper nitrate
- (d) Aluminium chloride
- (e) Calcium carbonate.

48. Give the names of the elements present in the following compounds.

- (a) Quick lime
- (b) Hydrogen bromide
- (c) Baking powder
- (d) Potassium sulphate.

49. How many atoms are present in a (i) H_2S molecule and (ii) PO_4^{3-} ion?

50. (a) Write a chemical formula of a compound using zinc ion and phosphate ion. (b) Calculate the ratio by mass of atoms present in a molecule of carbon dioxide.
(Given C = 12, O = 16)

51. What is the molecular mass of a substance?

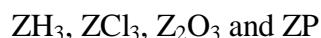
52. What is Formula Unit Mass? How it is different from molecular mass?

53. Calculate the formula unit masses of ZnO , Na_2O , K_2CO_3 , given atomic masses of Zn = 65 u, Na = 23 u, K = 39 u, C = 12 u, and O = 16 u.

54. What are ionic compounds?

55. How do we know the presence of atoms if they do not exist independently for most of the elements?

56. An element 'Z' forms the following compound when it reacts with hydrogen, chlorine, oxygen and phosphorus.



- (a) What is the valency of element Z?
- (b) Element 'Z' is metal or non-metal?

57. Name one element each which forms diatomic and tetra atomic molecule.

58. Name one element which forms diatomic and triatomic molecule.

- 59.** What is gram-atomic mass of an element?
- 60.** What is gram-molecular mass of a substance?
- 61.** Define mole. What is its significance?
- 62.** What is molar mass?
- 63.** Who introduced the term 'mole' in chemistry?
- 64.** When 'mole' was chosen internationally standard way to express larger number of chemical units?
- 65.** How many moles are there in 4.6 gms of Sodium(Na)?
- 66.** If one mole of carbon atoms weighs 12 gram, what is the mass (in gram) of 1 atom of carbon?
- 67.** Which has more number of atoms, 100 grams of sodium or 100 grams of iron (given, atomic mass of Na = 23 u, Fe = 56 u)?
- 68.** What is the mass of
(a) 1 mole of nitrogen atoms?
(b) 4 moles of aluminium atoms (Atomic mass of aluminium = 27)?
(c) 10 moles of sodium sulphite (Na_2SO_3)?
- 69.** Convert into moles:
(a) 12 g of oxygen gas
(b) 20 g of water
(c) 22 g of carbon dioxide
- 70.** What is the mass of: (a) 0.2 mole of oxygen atoms? (b) 0.5 mole of water molecules?
- 71.** Find out number of atoms in 15 moles of He.
- 72.** Calculate the number of molecules of sulphur (S_8) present in 16 g of solid sulphur.
- 73.** Calculate the number of aluminium ions present in 0.051 g of aluminium oxide.
(Hint: The mass of an ion is the same as that of an atom of the same element. Atomic mass of Al = 27 u)
- 74.** Calculate the mass percentage of Carbon(C), Hydrogen (H) and Oxygen (O) in one molecule of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$).
(Atomic mass of C = 12u, H = 1u and O = 16u)
- 75.** Calculate the number of molecules of phosphorus (P_4) present in 31 gram of phosphorus.

ASSIGNMENT QUESTIONS SET – 2
CHAPTER – 3
ATOMS AND MOLECULES

MOLECULAR MASS AND MOLE CALCULATION PROBLEMS

1. Calculate the number of moles in i) 4.6g sodium ii) 5.1g of Ammonia iii) 90g of water iv) 2g of NaOH
2. Calculate the number of molecules in 360g of glucose.
3. Find the mass of 2.5 mole of oxygen atoms
4. Calculate the mass of 12.046×10^{23} molecules in CaO.
5. Calculate the number of moles in 24.092×10^{22} molecules of water.
6. Calculate the number of moles in a) 12.046×10^{23} atoms of copper b) 27.95g of iron c) 1.51×10^{23} molecules of CO₂
7. If 3.0115×10^{23} particles are present in CO₂. Find the number of moles.
8. Find the number of moles present in 24.088×10^{23} particles of carbon dioxide
9. Calculate the number of atoms in 48g of Mg
10. Calculate the number of molecules in 3.6 g of water
11. Calculate the number of atoms in 0.5 moles of carbon
12. Calculate the number of moles in 12g of oxygen gas
13. Calculate the number of moles present in 14g of carbon monoxide .
14. Find the mass of 5 moles of aluminium atoms?
15. Calculate the molar mass of sulphur.
16. Calculate the mass of 0.2 mole of water molecules.
17. Which has greater number of atoms, 100g of sodium or 100g of iron?
18. How many atoms of oxygen are present in 300 grams of CaCO₃?
19. The mass of one atom of an element 'A' is 2.65×10^{-23} g. Calculate its atomic mass and name the element.
20. Calculate the number moles of magnesium in 0.478g of magnesium ?
21. In which of the following cases the number of hydrogen atoms is more ? Two moles of HCl or one mole of NH₃.
22. Calculation of number of hydrogen atoms present in 1 mole of NH₃
23. Find the number of oxygen atoms in 88g of CO₂?
24. Calculate the number of water molecules contained in a drop of water weighing 0.06g ?
25. Find the number of aluminium ions present in 0.051g of aluminium oxide(Al₂O₃). (Atomic masses: Al= 27u; O= 16u)
26. Calculate the mass of 1.000 mole of CaCl₂
27. Calculate grams in 3.0000 moles of CO₂
28. Calculate number of moles in 32.0 g of CH₄

- 29.** Determine mass in grams of 40.0 moles of Na_2CO_3
- 30.** Calculate moles in 168.0 g of HgS
- 31.** Calculate moles in 510.0 g of Al_2S_3
- 32.** How many moles are in 27.00 g of H_2O
- 33.** Determine the mass in grams of Avogadro number of $\text{C}_{12}\text{H}_{22}\text{O}_{11}$
- 34.** Find mass in grams of 9.03 moles of H_2S
- 35.** Determine grams in 1.204 mole of NH_3

Consider the molecule CuNH_4Cl_3 as you answer 11 - 19.

- 36.** Name the elements present.
- 37.** How many atoms form the molecule?
- 38.** How many of each atom in the molecule?
- 39.** How many hydrogen atoms in one mole of molecules?
- 40.** How many chlorine atoms in six moles of molecules?
- 41.** What is the molar mass of this molecule?
- 42.** Name this molecule.
- 43.** What is the mass in grams of one molecule?
- 44.** How many moles would be in 6.84 g of this substance?
- 45.** You need 0.0100 mole of lead (II) chromate. How much should you weigh on the scale?
- 46.** Given 6.40 g of HBr. How many moles is this?
- 47.** Write the correct formula for calcium acetate and then answer 23 - 25 based on it.
- 48.** What is the mass of exactly one mole of calcium acetate?
- 49.** How many moles are contained in 1.58 g of the substance in #23?
- 50.** How much does 0.400 mole of #23 weigh?
- 51.** Write the formula for oxygen gas.
- 52.** How many atoms (and moles) are represented by the formula in #26?
- 53.** What is the mass of Avogadro Number of oxygen molecules?
- 54.** Calculate the molar mass of HNO_3 . [N = 14, O = 16, H = 1]
- 55.** Calculate the formula mass of CaCl_2 . [Ca = 40, Cl = 35.5]
- 56.** A certain non-metal X forms two oxides I and II. The mass percentage of oxygen in oxide I (X_4O_6) is 43.7, which is same as that of X in oxide II. Find the formula of the second oxide.
- 57.** Calculate the mass of 0.2 moles of water (O=16, H=1).
- 58.** What is the volume of 7.1 g of chlorine (Cl=35.5) at S.T.P.
- 59.** The reaction between aluminium carbide and water takes place according to the following equation: $\text{Al}_4\text{C}_3 + 12\text{H}_2\text{O} \longrightarrow 3\text{CH}_4 + 4\text{Al(OH)}_3$. Calculate the volume of CH_4 released from 14.4 g of Al_4C_3 by excess water at S.T.P. (C = 12, Al = 27)

- 60.** A compound of sodium, sulphur and oxygen has the following percentage composition. Na=29.11%, S=40.51%, O=30.38%. Find its empirical formula (O=16, Na=23, S=32).
- 61.** Solid ammonium dichromate with relative molecular mass of 252 g decomposes according to the equation: $(NH_4)Cr_2O_7 \longrightarrow N_2 + Cr_2O_3 + 4H_2O$. (i) What volume of nitrogen at S.T.P will be evolved when 63 g of $(NH_4)_2Cr_2O_7$ is decomposed? (ii) If 63 g of $(NH_4)_2Cr_2O_7$ is heated above 100^0C , what will be the loss of mass? (H=1, N=14, O=16, Cr=52).
- 62.** How many litres of ammonia are present in 3.4 kg of it? (N = 14, H = 1)
- 63.** About 640 mL of carbon monoxide is mixed with 800 mL of oxygen and ignited in an enclosed vessel. Calculate the total volume of gases after the burning is completed. All volumes are measured at S.T.P.
- 64.** Calculate the number of moles of ammonium sulphate present in 15.84 kg of it. (H=1, N=14, O=16, S=32)
-

ASSIGNMENT QUESTIONS SET – 3

CHAPTER – 3

ATOMS AND MOLECULES

- 11.** Which of the following has maximum number of atoms?
- (a) 18g of H_2O
 - (b) 18g of O_2
 - (c) 18g of CO_2
 - (d) 18g of CH_4
- 12.** Which of the following contains maximum number of molecules?
- (a) 1g CO_2
 - (b) 1g N_2
 - (c) 1g H_2
 - (d) 1g CH_4
- 13.** Mass of one atom of oxygen is
- (a) $\frac{16}{6.023 \times 10^{23}}$ g
 - (b) $\frac{32}{6.023 \times 10^{23}}$ g
 - (c) $\frac{1}{6.023 \times 10^{23}}$ g
 - (d) $8u$
- 14.** 3.42 g of sucrose are dissolved in 18g of water in a beaker. The number of oxygen atoms in the solution are
- (a) 6.68×10^{23}
 - (b) 6.09×10^{22}
 - (c) 6.022×10^{23}
 - (d) 6.022×10^{21}
- 15.** A change in the physical state can be brought about
- (a) only when energy is given to the system
 - (b) only when energy is taken out from the system
 - (c) when energy is either given to, or taken out from the system
 - (d) without any energy change
- 16.** Which of the following represents a correct chemical formula? Name it.
- (a) CaCl
 - (b) BiPO_4
 - (c) NaSO_4
 - (d) NaS
- 17.** Write the molecular formulae for the following compounds
- (a) Copper (II) bromide
 - (b) Aluminium (III) nitrate
 - (c) Calcium (II) phosphate
 - (d) Iron (III) sulphide
 - (e) Mercury (II) chloride
 - (f) Magnesium (II) acetate
- 18.** Write the molecular formulae of all the compounds that can be formed by the combination of following ions
- $\text{Cu}^{2+}, \text{Na}^+, \text{Fe}^{3+}, \text{Cl}^-, \text{SO}_4^{2-}, \text{PO}_4^{3-}$
- 19.** Write the cations and anions present (if any) in the following compounds
- (a) CH_3COONa
 - (b) NaCl
 - (c) H_2
 - (d) NH_4NO_3
- 20.** Give the formulae of the compounds formed from the following sets of elements
- (a) Calcium and fluorine

- (b) Hydrogen and sulphur
- (c) Nitrogen and hydrogen
- (d) Carbon and chlorine
- (e) Sodium and oxygen
- (f) Carbon and oxygen

21. Which of the following symbols of elements are incorrect? Give their correct symbols

- (a) Cobalt CO
- (b) Carbon c
- (c) Aluminium AL
- (d) Helium He
- (e) Sodium So

22. Give the chemical formulae for the following compounds and compute the ratio by mass of the combining elements in each one of them.

- (a) Ammonia
- (b) Carbon monoxide
- (c) Hydrogen chloride
- (d) Aluminium fluoride
- (e) Magnesium sulphide

23. State the number of atoms present in each of the following chemical species

- (a) CO_3^{2-}
- (b) PO_4^{3-}
- (c) P_2O_5
- (d) CO

24. Find the ratio by mass of the combining elements in the compound – $\text{C}_2\text{H}_5\text{OH}$.

25. Give the formula of the compound formed by the elements calcium and fluorine.

26. What is the acid radical present in sodium peroxide?

27. Carbon and silicon have the same valency. What is the formula of sodium silicate?

28. What is the ratio by number of atoms in mercurous chloride?

29. Name the element whose Latin name is Stibium.

30. What is the valency of a sulphide ion?

31. How many atoms of oxygen are present in 50g of CaCO_3 ?

32. How many molecules are present in 1 ml of water?

33. What is the unit of measurement of atomic radius?

34. Name the international organization who approves names of elements.

35. How do we know the presence of atoms if they do not exist independently for most of the elements?

36. Give an example to show Law of conservation of mass applies to physical change also.

37. Explain with example that law of conservation of mass is valid for chemical reactions.

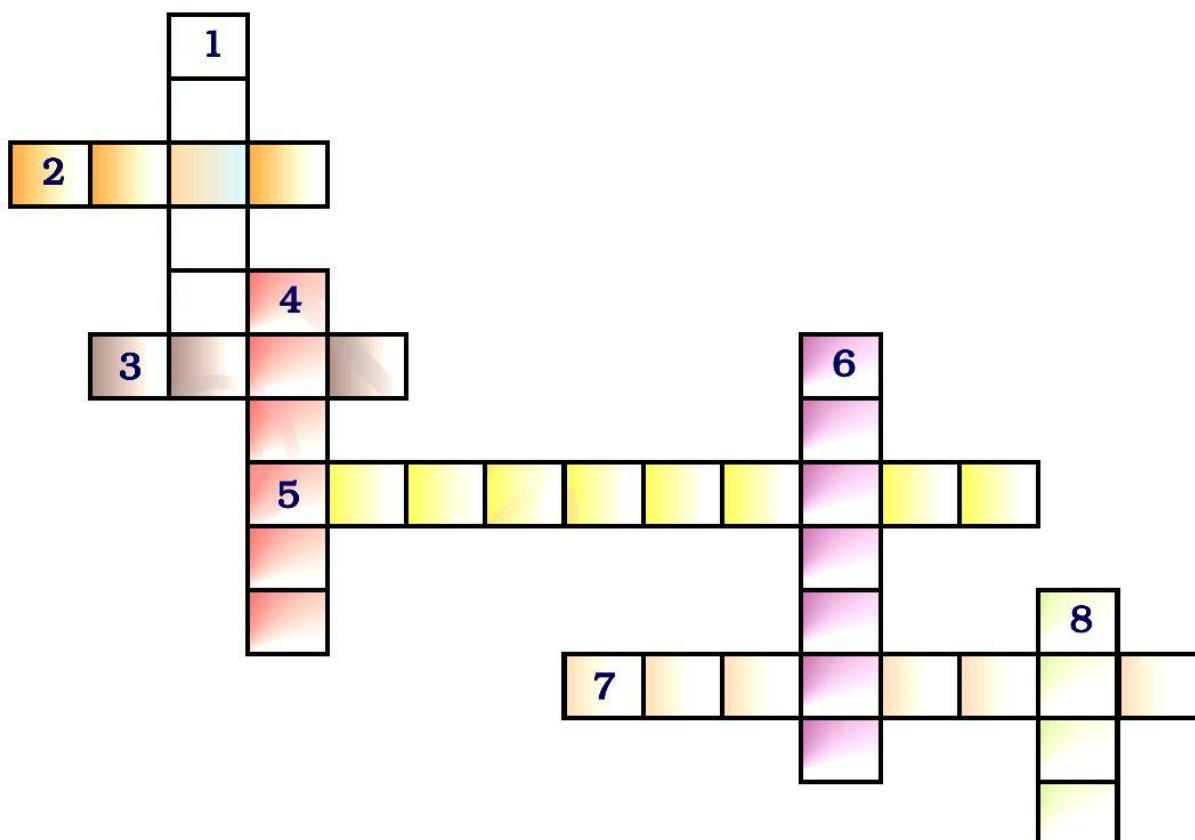
38. Is there any exception to law of conservation of mass?

- 39.** In a reaction, 5.3 g of sodium carbonate reacted with 6 g of ethanoic acid. The products were 2.2 g of carbon dioxide, 0.9 g water and 8.2 g of sodium ethanoate. Show that these observations are in agreement with the law of conservation of mass.
sodium carbonate + ethanoic acid → sodium ethanoate + carbon dioxide + water
- 40.** If 12 g of carbon is burnt in the presence of 32 g of oxygen, how much carbon dioxide will be formed?
- 41.** A 0.24 g sample of compound of oxygen and boron was found by analysis to contain 0.096 g of boron and 0.144 g of oxygen. Calculate the percentage composition of the compound by weight.
- 42.** When 3.0 g of carbon is burnt in 8.00 g oxygen, 11.00 g of carbon dioxide is produced. What mass of carbon dioxide will be formed when 3.00 g of carbon is burnt in 50.00 g of oxygen? Which law of chemical combination will govern your answer?
- 43.** Magnesium and oxygen combine in the ratio of 3 : 2 by mass to form magnesium oxide. How much oxygen is required to react completely with 12 g of magnesium?
- 44.** Why are Dalton's symbol not used in chemistry?
- 45.** What is the fraction of the mass of water due to neutrons?
- 46.** Does the solubility of a substance change with temperature? Explain with the help of an example.
- 47.** Classify each of the following on the basis of their atomicity.
(a) F₂ (b) NO₂ (c) N₂O (d) C₂H₆ (e) P₄ (f) H₂O₂
(g) P₄O₁₀ (h) O₃ (i) HCl (j) CH₄ (k) He (l) Ag
- 48.** You are provided with a fine white coloured powder which is either sugar or salt. How would you identify it without tasting?
- 49.** Calculate the number of moles of magnesium present in a magnesium ribbon weighing 12 g. Molar atomic mass of magnesium is 24 g mol⁻¹.
- 50.** Verify by calculating that
(a) 5 moles of CO₂ and 5 moles of H₂O do not have the same mass.
(b) 240 g of calcium and 240 g magnesium elements have a mole ratio of 3:5.
- 51.** Find the ratio by mass of the combining elements in the following compounds.
(a) CaCO₃ (d) C₂H₅OH
(b) MgCl₂ (e) NH₃
(c) H₂SO₄ (f) Ca(OH)₂
- 52.** Calcium chloride when dissolved in water dissociates into its ions according to the following equation.
 $\text{CaCl}_2 \text{ (aq)} \rightarrow \square \text{Ca}^{2+} \text{ (aq)} + 2\text{Cl}^- \text{ (aq)}$
Calculate the number of ions obtained from CaCl₂ when 222 g of it is dissolved in water.
- 53.** The difference in the mass of 100 moles each of sodium atoms and sodium ions is 5.48002 g. Compute the mass of an electron.

- 54.** Complete the following crossword puzzle (below Figure) by using the name of the chemical elements. Use the data given in below Table

Table 3.2

Across	Down
2. The element used by Rutherford during his α -scattering experiment	1. A white lustrous metal used for making ornaments and which tends to get tarnished black in the presence of moist air
3. An element which forms rust on exposure to moist air	4. Both brass and bronze are alloys of the element
5. A very reactive non-metal stored under water	6. The metal which exists in the liquid state at room temperature
7. Zinc metal when treated with dilute hydrochloric acid produces a gas of this element which when tested with burning splinter produces a pop sound.	8. An element with symbol Pb



- 55.** Fill in the missing data in the below Table

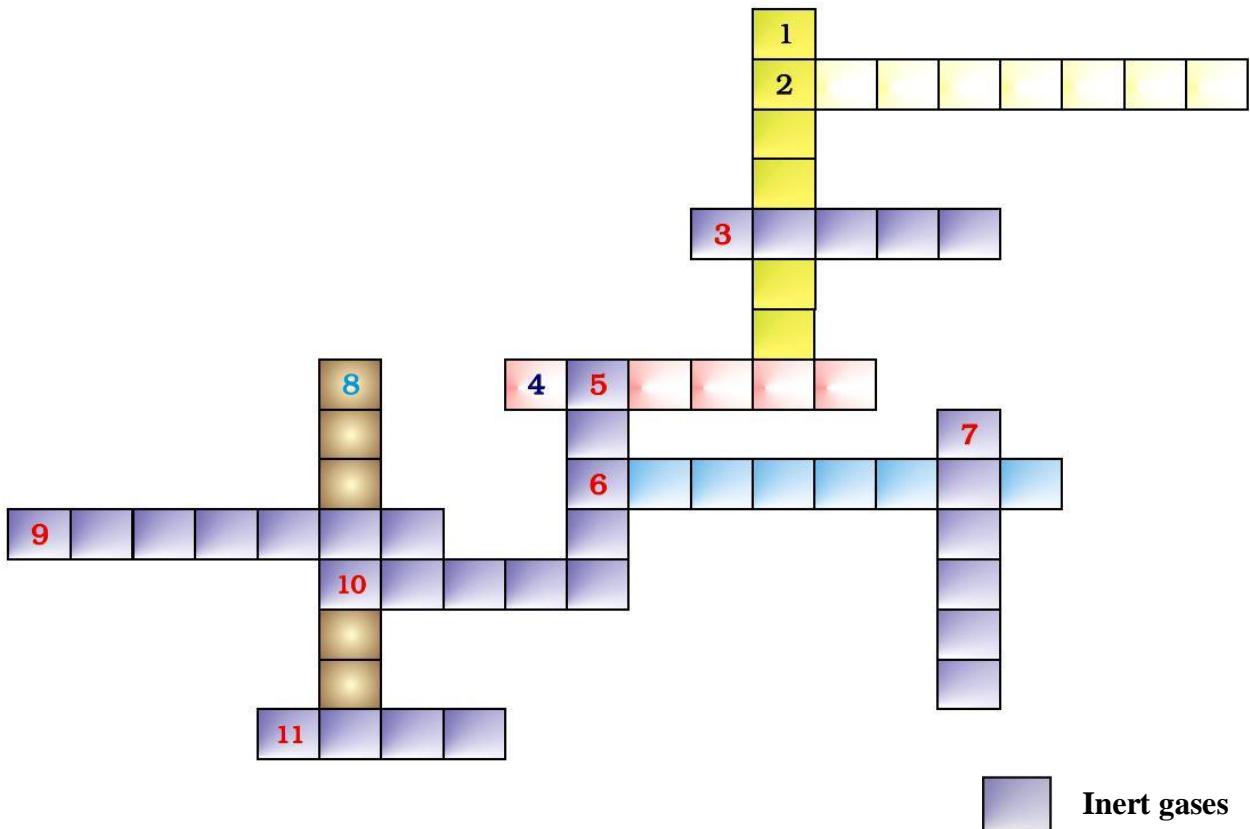
Species	H_2O	CO_2	Na atom	MgCl_2
Property				
No of moles	2			0.5
No. of particles		3.011×10^{23}		
Mass	36g		115g	

- 56.** The visible universe is estimated to contain 10^{22} stars. How many moles of stars are present in the visible universe?

- 57.** What is the SI prefix for each of the following multiples and submultiples of a unit?
 (a) 10^3 (b) 10^{-1} (c) 10^{-2} (d) 10^{-6} (e) 10^{-9} (f) 10^{-12}

58. (a) In this crossword puzzle (Fig 3.2), names of 11 elements are hidden. Symbols of these are given below. Complete the puzzle.

- | | |
|-------|--------|
| 1. Cl | 7. He |
| 2. H | 8. F |
| 3. Ar | 9. Kr |
| 4. O | 10. Rn |
| 5. Xe | 11. Ne |
| 6. N | |



(b) Identify the total number of inert gases, their names and symbols from this cross word puzzle.

59. Express each of the following in kilograms

- 5.84×10^{-3} mg
- 58.34 g
- 0.584g
- 5.873×10^{-21} g

60. Compute the difference in masses of 103 moles each of magnesium atoms and magnesium ions. (Mass of an electron = 9.1×10^{-31} kg)

61. Which has more number of atoms? 100g of N_2 or 100 g of NH_3

62. Compute the number of ions present in 5.85 g of sodium chloride.

63. A gold sample contains 90% of gold and the rest copper. How many atoms of gold are present in one gram of this sample of gold?

- 64.** Cinnabar (HgS) is a prominent ore of mercury. How many grams of mercury are present in 225 g of pure HgS ? Molar mass of Hg and S are 200.6 g mol^{-1} and 32 g mol^{-1} respectively.
- 65.** The mass of one steel screw is 4.11g. Find the mass of one mole of these steel screws. Compare this value with the mass of the Earth ($5.98 \times 10^{24} \text{ kg}$). Which one of the two is heavier and by how many times?
- 66.** A sample of vitamic C is known to contain 2.58×10^{24} oxygen atoms. How many moles of oxygen atoms are present in the sample?
- 67.** Raunak took 5 moles of carbon atoms in a container and Krish also took 5 moles of sodium atoms in another container of same weight. (a) Whose container is heavier? (b) Whose container has more number of atoms?
- 68.** What are ionic and molecular compounds? Give examples.
- 69.** Compute the difference in masses of one mole each of aluminium atoms and one mole of its ions. (Mass of an electron is $9.1 \times 10^{-28} \text{ g}$). Which one is heavier?
- 70.** A silver ornament of mass ‘m’ gram is polished with gold equivalent to 1% of the mass of silver. Compute the ratio of the number of atoms of gold and silver in the ornament.
- 71.** A sample of ethane (C_2H_6) gas has the same mass as 1.5×10^{20} molecules of methane (CH_4). How many C_2H_6 molecules does the sample of gas contain?
- 72.** Fill in the blanks
(a) In a chemical reaction, the sum of the masses of the reactants and products remains unchanged. This is called _____.
(b) A group of atoms carrying a fixed charge on them is called _____.
(c) The formula unit mass of $\text{Ca}_3(\text{PO}_4)_2$ is _____.
(d) Formula of sodium carbonate is _____ and that of ammonium sulphate is -----.
- 73.** Write the formulae for the following and calculate the molecular mass for each one of them.
(a) Caustic potash
(b) Baking powder
(c) Lime stone
(d) Caustic soda
(e) Ethanol
(f) Common salt
- 74.** In photosynthesis, 6 molecules of carbon dioxide combine with an equal number of water molecules through a complex series of reactions to give a molecule of glucose having a molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$. How many grams of water would be required to produce 18 g of glucose? Compute the volume of water so consumed assuming the density of water to be 1 g cm^{-3} .

CHAPTER – 4

STRUCTURE OF THE ATOM

STRUCTURE OF THE ATOM

Atoms are the basic units of matter and the defining structure of elements. Matter are made of tiny particles called atom. Atom is made of three particles; electron, proton and neutron. These particles are called fundamental particles of an atom or sub atomic particles.

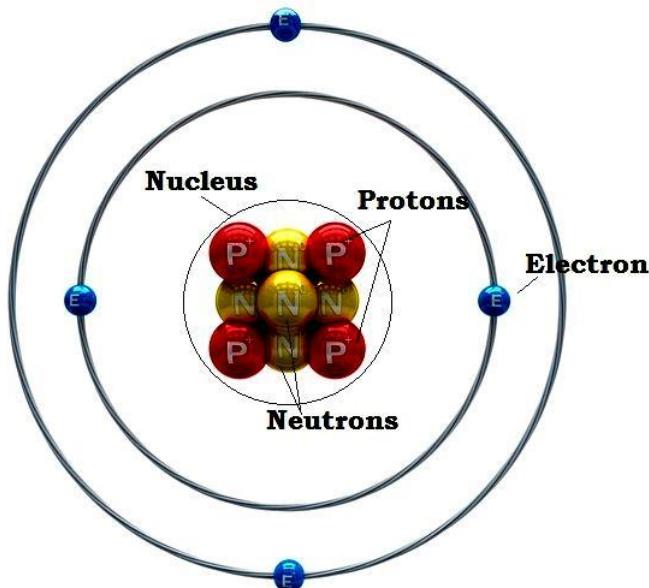
Electron (e^-) - Electron is denoted by ‘e’ and is a negatively charged particle. The absolute charge over an electron is equal to 1.6×10^{-19} of negative charge and is considered equal to – 1. The relative mass of electron is 1/1836. Since the mass of an electron is very small, thus it is considered equal to 0. Electrons revolve round the nucleus of atoms.

Proton (p^+) - Proton is denoted by ‘p’ and is positively charged particle. The absolute charge over proton is 1.6×10^{-19} coulomb of positive charge and it is considered as unit positive charge. Thus absolute charge over a proton is equal to +1. The absolute mass of a proton is equal to 1.6×10^{-24} g and considered equal to 1 as it is equal to the mass of 1 hydrogen atom. Proton is present in the nucleus of atom.

Neutron (n) – Neutron is denoted by ‘n’ and is a neutral particle.

The absolute mass of neutron is 1.6×10^{-24} g. The relative mass of neutron is equal to 1. Neutron is present in the nucleus of atom.

Nucleus – The centre of atom is called nucleus. Nucleus comprises of neutron and proton. Nucleus of an atom contains the whole mass of an atom.



INTEXT QUESTIONS PAGE NO. 47

Q1. What are canal rays?

Answer:

Canal rays are positively charged radiations that can pass through perforated cathode plate. These rays consist of positively charged particles known as protons.

Q2. If an atom contains one electron and one proton, will it carry any charge or not?

Answer:

An electron is a negatively charged particle, whereas a proton is a positively charged particle. The magnitude of their charges is equal. Therefore, an atom containing one electron and one proton will not carry any charge. Thus, it will be a neutral atom.

Discovery of Electron

In 1897; J. J. Thomson, a British physicist, proposed that atom contains at least one negatively charged particle. Later this particle was named as electron. Thomson called those particles 'corpuscles'.

Discovery of Proton:

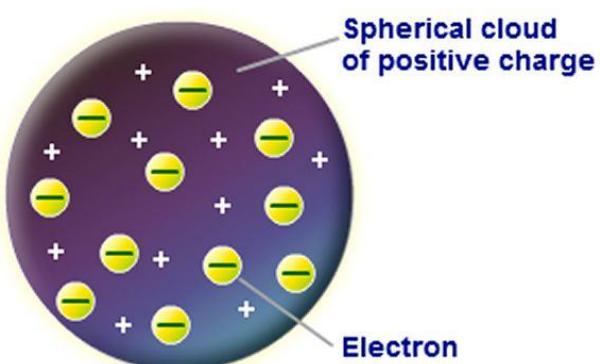
Ernest Goldstein in 1886 discovered the presence of new radiation in gas discharge tube even before the identification of electron. He called these rays as Canal Rays. His experiment led to the discovery of proton.

Discovery of Neutron:

In 1932 J. Chadwick discovered another subatomic particle called neutron. Neutron is present in the nucleus of all atoms.

THOMSON'S MODEL OF ATOM

J. J. Thomson proposed the model of atom similar to a Christmas Pudding or similar to a water melon. His model of atom is generally called plum and pudding model of atom.



Thomson's Plum pudding model

He proposed that electrons are embedded the way black seeds of water melon are embedded; in the sphere of positive charge. According to Thomson

- (a) An atom consists of positively charged sphere in which electrons are embedded.
- (b) The quanta of negative and positive charges are equal. The equal number of negative charge and positive charge makes an atom electrically neutral.

RUTHERFORD'S MODEL OF ATOM

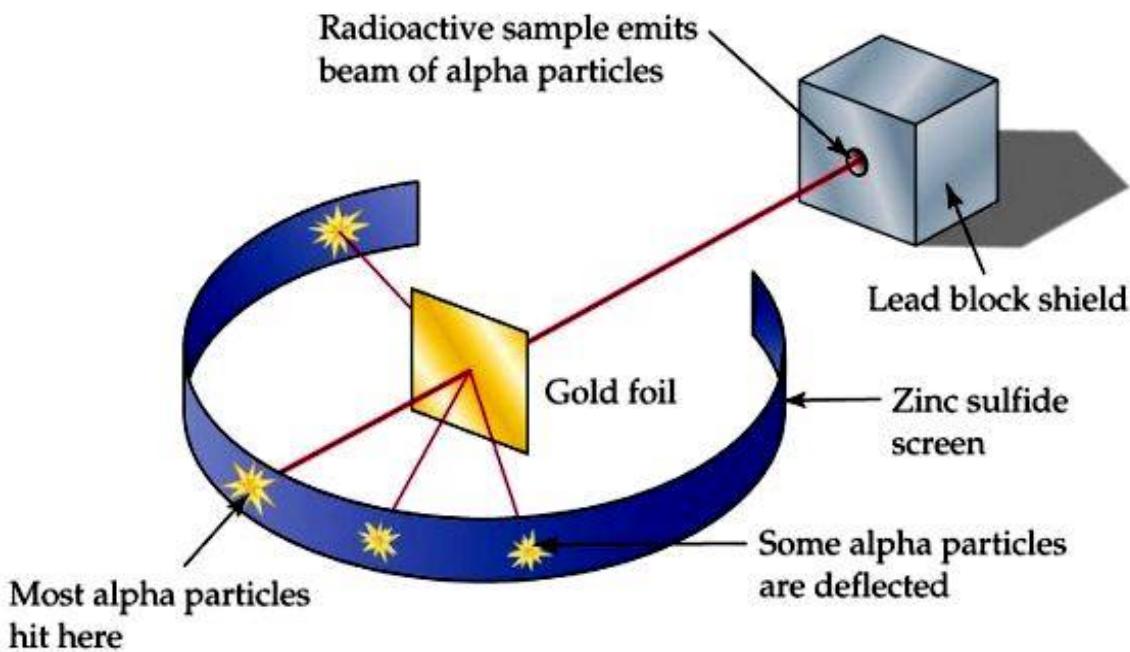
Ernest Rutherford in 1909 with his team bombarded very thin gold foil with α – particles. He found that

- (a) Most of the α – particles passed without any hindrance.
- (b) Some of the α – particles deflected from their original path at noticeable angle.
- (c) Very few of the α – particles bounced back at their original path.

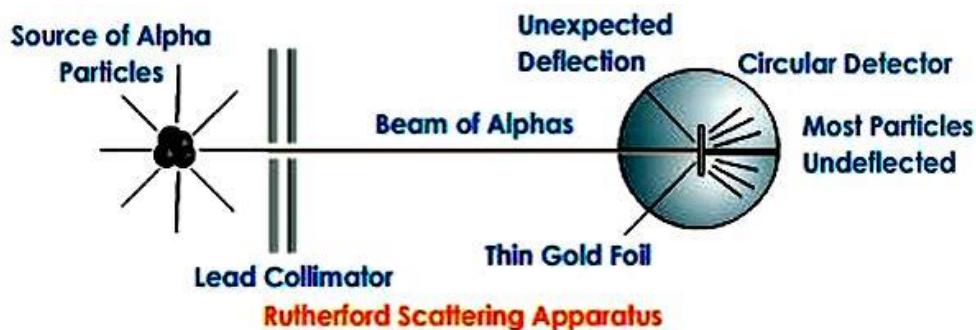
On the basis of his observation, he proposed the model of atom. The Rutherford's Model of Atom is as follows:

- (a) Most of the part in an atom is empty.
- (b) There is a positively charged center in atom, which contains nearly the whole mass of atom. The centre is called nucleus.
- (c) The size of nucleus is very small compared to an atom.

(d) Electrons revolve round the nucleus.



The Rutherford's Experiment is also known as Geiger-Marsden Experiment.



DRAWBACKS OF RUTHERFORD MODEL

(a) According to Rutherford's Model, electron revolves round the positively charged nucleus which is not expected to be stable. But a charged particle in an accelerated motion along a circular path would lose energy because of radiation and finally would fall into nucleus. This makes an atom unstable while atoms are quite stable.

If atoms were not stable no matter would exist in nature.

(b) Rutherford model could not solve the problem of atomic mass of atom as it proposed only the existence of protons in the nucleus.

However, the problem of atomic mass could be solved after the discovery of neutron.

BOHR'S MODEL OF ATOM

Neils Bohr, a Danish physicist, in 1913 proposed model of atom which rectified the problems left by Rutherford's Model. He proposed that

(a) Electrons revolve round the nucleus in a fixed orbit.

(b) He called these orbits as 'stationary orbit'.

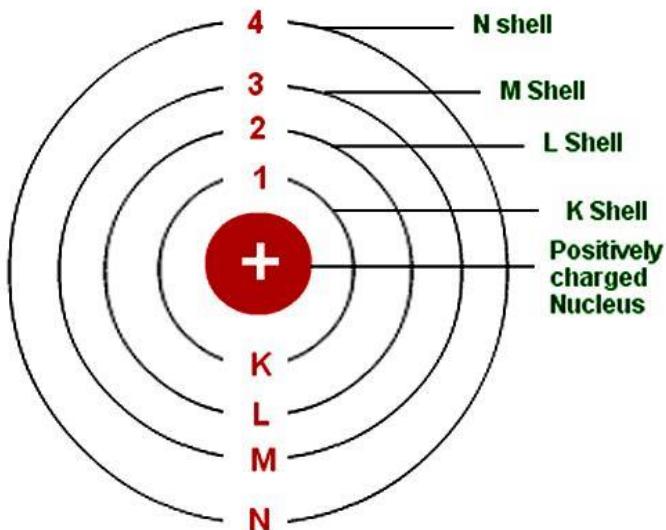
(c) Each stationary orbit is associated with fixed amount of energy, thus electrons do not radiate energy as long as they keep on revolving around the nucleus in fixed orbit.

The circular path around the nucleus is called orbit, energy level or shell. Energy level are represented by letter – K, L, M, N, and so on.

Therefore,

- 1st orbit is denoted by – K
- 2nd orbit is denoted by – L
- 3rd orbit is denoted by – M, and so on.

The orbits are denoted by 1, 2, 3, and so on.



INTEXT QUESTIONS PAGE NO. 49

Q1. On the basis of Thomson's model of an atom, explain how the atom is neutral as a whole.

Answer:

As per Thomson's model of the atom, an atom consists both negative and positive charges which are equal in number and magnitude. So, they balance each other as a result of which atom as a whole is electrically neutral.

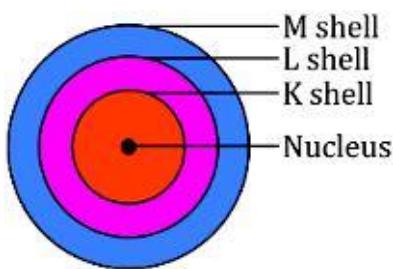
Q2. On the basis of Rutherford's model of an atom, which subatomic particle is present in the nucleus of an atom?

Answer:

On the basis of Rutherford's model of an atom, protons are present in the nucleus of an atom.

Q3. Draw a sketch of Bohr's model of an atom with three shells.

Answer:



Q4. What do you think would be the observation if the α -particle scattering experiment is carried out using a foil of a metal other than gold?

Answer:

If α -particle scattering experiment is carried out using a foil of any metal as thin as gold foil used by Rutherford, there would be no change in observations. But since other metals are not so malleable so, such a thin foil is difficult to obtain. If we use a thick foil, then more α -

particles would bounce back and no idea about the location of positive mass in the atom would be available with such a certainty.

INTEXT QUESTIONS PAGE NO. 49

Q1. Name the three sub-atomic particles of an atom.

Answer:

The three sub-atomic particles of an atom are:

- (i) Protons
- (ii) Electrons, and
- (iii) Neutrons

Q2. Helium atom has an atomic mass of 4 u and two protons in its nucleus. How many neutrons does it have?

Answer:

Number of neutrons = Atomic mass - Number of protons

Therefore, the number of neutrons in the atom = $4 - 2 = 2$

DISTRIBUTION OF ELECTRONS IN ORBIT OR SHELL:

The distribution of electrons in an orbit is obtained by $2n^2$, where 'n' is number of orbit.

Therefore,

Number of electrons in K-shell i.e. in 1st orbit.

Here $n = 1$

Therefore,

$$2n^2 = 2 \times 1^2 = 2$$

Thus, maximum number of electrons in K-shell i.e. 1st shell = 2

Number of electrons in L-shell, i.e. in 2nd orbit

Here $n = 2$, therefore,

$$2n^2 = 2 \times 2^2 = 8$$

Thus, maximum number of electrons in L-shell = 8

Number of electrons in M-shell, i.e. in 3rd orbit

Here $n = 3$, therefore,

$$2n^2 = 2 \times 3^2 = 18$$

Thus, maximum number of electrons in M-shell = 18

Number of electrons in N-shell, i.e. in 4th shell

Here $n = 4$, therefore,

$$2n^2 = 2 \times 4^2 = 32$$

Thus, maximum number of electrons in N-shell = 32

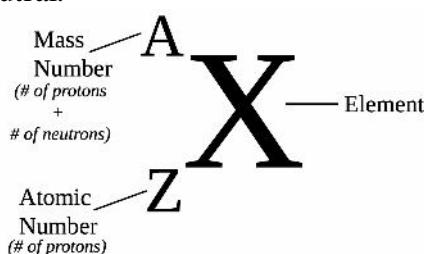
In similar way maximum number of electrons in any shell can be calculated.

ATOMIC NUMBER

Atomic number is the fundamental properties of an atom. Every atom is identified by its unique atomic number. Atomic number is denoted by 'z'.

Atomic number is equal to the number of protons present in an atom.

Since an atom is electrically neutral, thus number of protons and number of electrons are equal to make an atom electrically neutral.



Atomic number = Number of protons = Number of electrons

Example :-

The atomic number of Hydrogen is 1, helium is 2, lithium is 3, beryllium is 4, boron is 5, carbon is 6, nitrogen is 7, oxygen is 8, etc.

Sample exercise:

(1) Atomic number of calcium is 20. Calculate the number of electrons and protons in calcium.

Solution:

Since, Atomic number = Number of protons = Number of electrons

Therefore,

Number of electrons in calcium = 20

Number of protons in calcium = 20

(2) Number of protons in sodium atom is 11, find the atomic number and number of electrons in a sodium atom.

Solution,

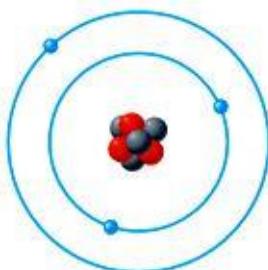
Since, Atomic number = Number of protons = Number of electrons

Therefore,

Atomic number of sodium = 11

Number of electrons in sodium = 11

Lithium, Li



Atomic number: 3

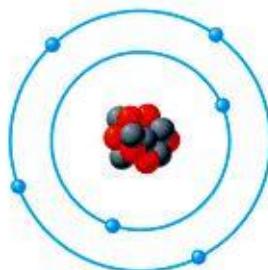
Average atomic mass: 6.941 amu

$$3 \text{ protons} = 3 \text{ amu}$$

$$+ 4 \text{ neutrons} = 4 \text{ amu}$$

$$\underline{\text{atomic mass} = 7 \text{ amu}}$$

Carbon, C



Atomic number: 6

Average atomic mass: 12.01 amu

$$6 \text{ protons} = 6 \text{ amu}$$

$$+ 6 \text{ neutrons} = 6 \text{ amu}$$

$$\underline{\text{atomic mass} = 12 \text{ amu}}$$

MASS NUMBER OR ATOMIC MASS

Mass number of an atom is defined as the sum of the number of protons and number of neutrons. Mass number is nearly equal to the atomic mass of an atom. Since, protons and neutrons reside in the nucleus, thus they are also known as nucleons.

This means

Mass number of an atom = Number of protons + Number of neutrons

Example

(1) Atomic mass of aluminium is 27 u and atomic number is 13, find the number of protons and number of neutrons in aluminium.

Solution:

Since,

Atomic number = 13

Therefore, number of proton = 13

We know that; Atomic mass (Mass number) = Number of protons + Number of neutrons

Therefore,

$$27 \text{ u} = 13 + n$$

$$\text{Or, } n = 27 - 13 = 14$$

Therefore, number of proton = 13 and number of neutron = 14

(2) The atomic number of carbon is 6 and number of neutron is equal to 6. Find the atomic mass or mass number of carbon.

Solution:

Since atomic number of carbon = 6

Therefore, number of proton = 6

Now, Atomic mass = number of proton + number of neutron

Or, Atomic mass or mass number = $6 + 6 = 12 \text{ u}$

Thus, mass number or atomic mass of carbon = 12u

ARRANGEMENT OF ELECTRONS IN AN ATOM – ELECTRONIC CONFIGURATION

The maximum number of electrons can be obtained by $2n^2$; where 'n' is the orbit number. Thus after knowing the maximum number of electrons for a particular shell, the arrangement of electrons in an atom can be identified. It is called Bohr Bury Schemes.

Rules to write the electronic configuration of an atom

- Maximum number of electrons in an orbit is calculated by $2n^2$, where 'n' is number of orbit and may be equal to 1, 2, 3, ...
- Electrons occupy the next orbit only after filling the inner orbit completely.
- The maximum number of electrons in outermost orbit will not be more than 8.

Electronic configuration of Hydrogen

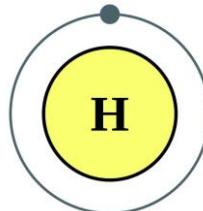
Atomic number of hydrogen = 1

Therefore number of electrons = 1

Maximum number of electrons in 1st orbit = 2

Since, hydrogen has only one electron, therefore, it will reside in 1st orbit.

Thus electronic configuration of hydrogen



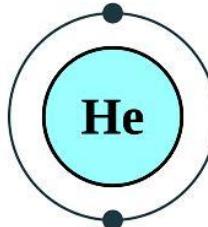
Number of orbit present in hydrogen = 1

Electronic configuration of Helium

Atomic number of helium = 2

Therefore number of electrons = 2

Therefore, electronic configuration of helium is



Number of orbit in helium atom = 1

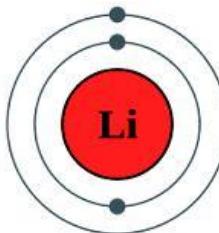
Electronic configuration of Lithium

Atomic number of Lithium = 3

Therefore number of electrons = 3

Since the maximum number of electrons in 1st orbit is equal to 2, therefore, after accommodating 2 electrons in 1st orbit, the third electron will go in 2nd orbit.

Thus, electronic configuration of lithium is



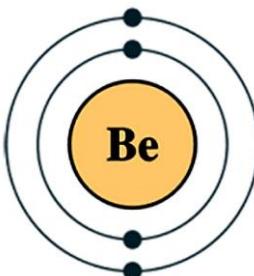
Number of orbit in Lithium atom = 3.

Electronic configuration of Beryllium

Atomic number of beryllium = 4.

Therefore number of electrons = 4.

Thus, electronic configuration of Beryllium is



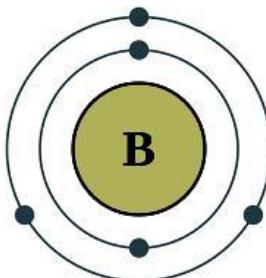
Number of orbit in beryllium = 2

Electronic configuration of Boron

Atomic number of boron = 5

Therefore number of electrons = 5

Thus, electronic configuration of boron is



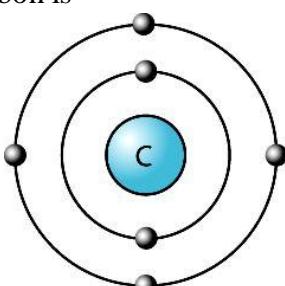
Number of orbit in boron = 2

Electronic configuration of Carbon

Atomic number of carbon = 6

Therefore number of electrons = 6

Thus, electronic configuration of carbon is



Number of orbit in carbon = 2

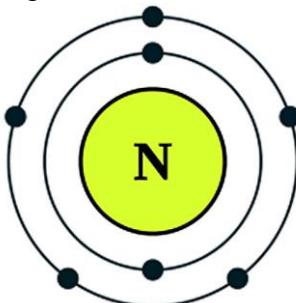
ELECTRONIC CONFIGURATION OF ELEMENTS – FROM NITROGEN (N) TO SODIUM (NA)

Electronic configuration of Nitrogen

Atomic number of nitrogen = 7.

Therefore number of electrons = 7

Thus, electronic configuration of nitrogen is



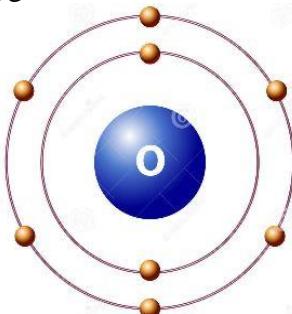
Number of orbit in nitrogen = 2

Electronic configuration of Oxygen

Atomic number of oxygen = 8.

Therefore number of electrons = 8.

Thus, electronic configuration of oxygen is



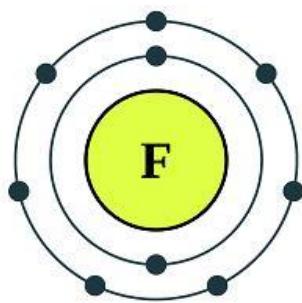
Number of orbit in oxygen = 2

Electronic configuration of Fluorine

Atomic number of fluorine = 9

Therefore number of electrons = 9

Thus, electronic configuration of fluorine is



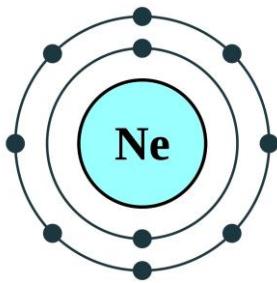
Number of orbit in fluorine = 2

Electronic configuration of Neon

Atomic number of neon = 10

Therefore number of electrons = 10

Thus, electronic configuration of neon is



Number of orbit in Neon = 2

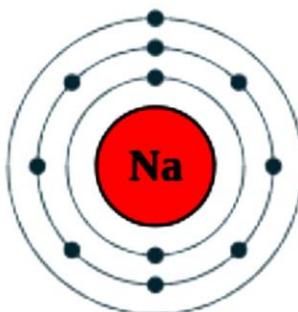
Electronic configuration of Sodium

Atomic number of sodium = 11

Therefore number of electrons = 11

Since, in ^{2nd} orbit the maximum number of electrons is equal to 8 and there are 11 electrons in sodium atom, thus the eleventh electron will go in third orbit.

Thus, electronic configuration of sodium is



Number of orbit in sodium = 3

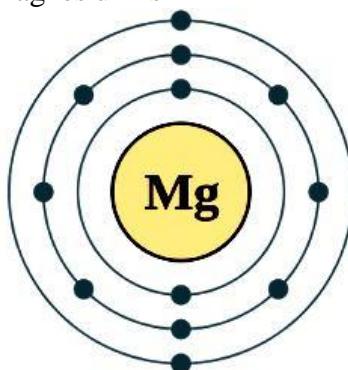
ELECTRONIC CONFIGURATION OF ELEMENTS – FROM MAGNESIUM TO CALCIUM

Electronic configuration of Magnesium

Atomic number of magnesium = 12

Therefore number of electrons = 12

Thus, electronic configuration of magnesium is



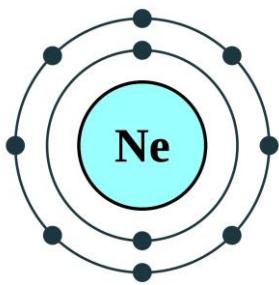
Number of orbit in magnesium = 3.

Electronic configuration of Aluminium

Atomic number of aluminium = 13.

Therefore number of electrons = 13.

Thus, electronic configuration of aluminium is



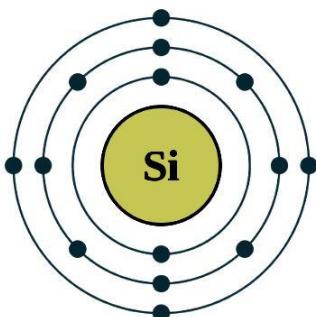
Number of orbit in aluminium = 3

Electronic configuration of Silicon

Atomic number of silicon = 14

Therefore number of electrons = 14

Thus, electronic configuration of silicon is



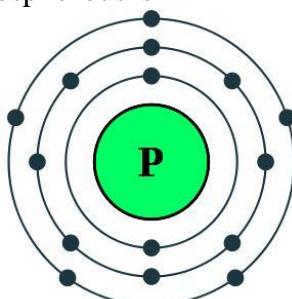
Number of orbit in silicon = 3

Electronic configuration of Phosphorous (P)

Atomic number of phosphorous = 15

Therefore number of electrons = 15

Thus, electronic configuration of phosphorous is



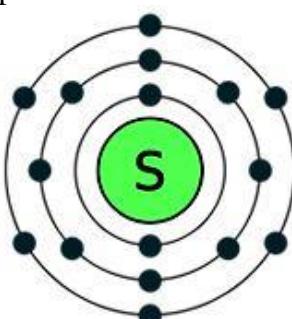
Number of orbit in phosphorous = 3

Electronic configuration of Sulphur (S)

Atomic number of sulphur = 16

Therefore number of electrons = 16

Thus, electronic configuration of sulphur is



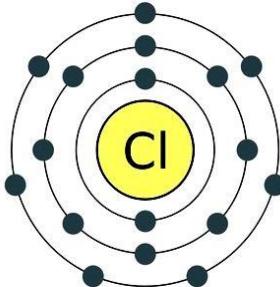
Number of orbit in sulphur = 3

Electronic configuration of Chlorine (Cl)

Atomic number of chlorine = 17

Therefore number of electrons = 17

Thus, electronic configuration of chlorine is



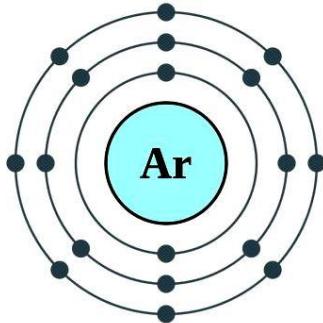
Number of orbit in chlorine = 3

Electronic configuration of Argon (Ar)

Atomic number of argon = 18

Therefore number of electrons = 18

Thus, electronic configuration of argon is



Number of orbit in argon = 3

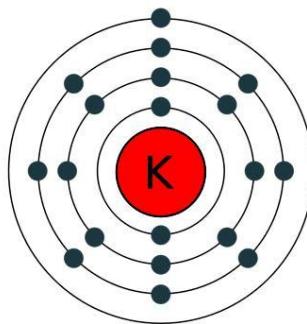
Electronic configuration of Potassium (K)

Atomic number of potassium = 19

Therefore number of electrons = 19

Since, maximum number of electrons in outermost orbit will not be more than 8, thus the 19th electron of potassium atom will reside in 4th orbit.

Thus, electronic configuration of potassium is



Number of orbit in potassium = 4

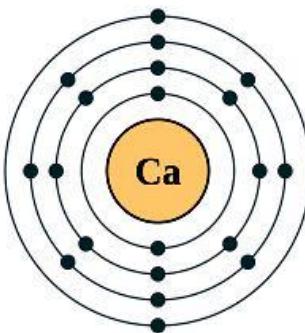
Electronic configuration of Calcium (Ca)

Atomic number of calcium = 20

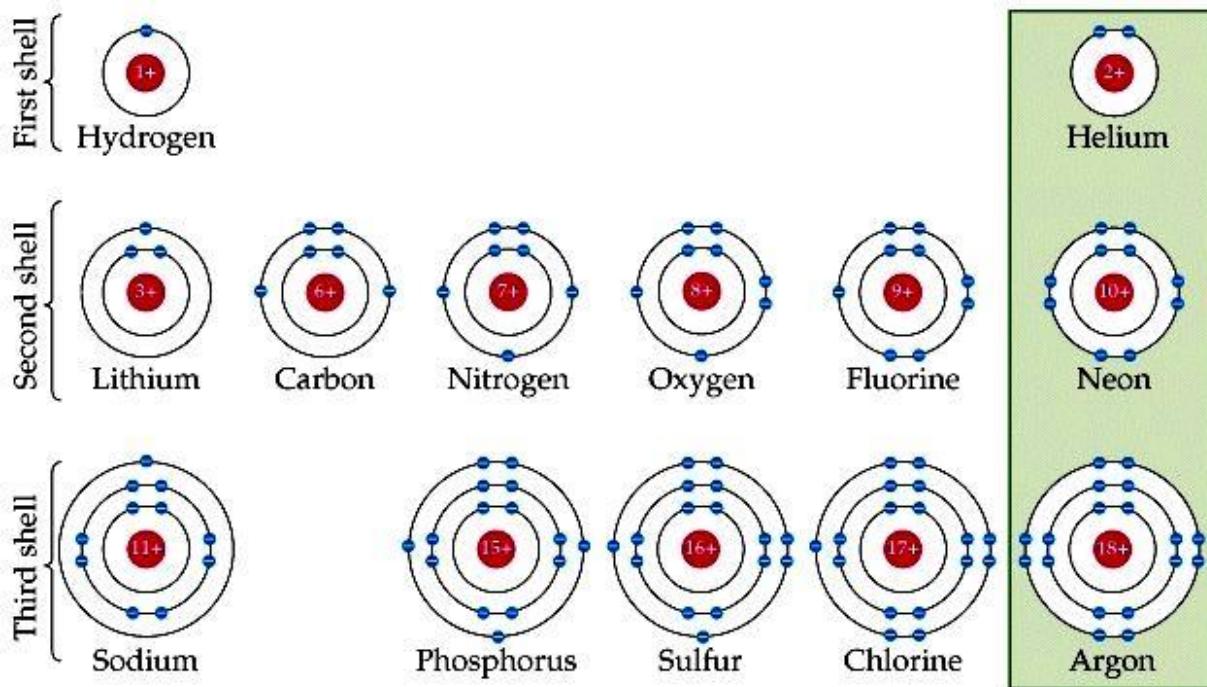
Therefore number of electrons = 20

Thus, electronic configuration of calcium is

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Number of orbit in calcium = 4



INTEXT QUESTIONS PAGE NO. 50

Q1. Write the distribution of electrons in carbon and sodium atoms.

Answer:

Carbon: The total number of electrons in a carbon atom is 6. The distribution of electrons in carbon atom is given by:

First orbit or K-shell = 2 electrons

Second orbit or L-shell = 4 electrons

Or, we can write the distribution of electrons in a carbon atom as 2, 4.

Sodium: The total number of electrons in a sodium atom is 11. The distribution of electrons in sodium atom is given by:

First orbit or K-shell = 2 electrons

Second orbit or L-shell = 8 electrons

Third orbit or M-shell = 1 electron

Or, we can write distribution of electrons in a sodium atom as 2, 8, 1.

Q2. If K and L shells of an atom are full, then what would be the total number of electrons in the atom?

Answer:

The maximum capacity of K shell is 2 electrons and L shell can accommodate maximum 8 electrons in it. Therefore, there will be ten electrons in the atom.

VALENCY

Noble gases have fully filled outermost shell. Due to this, they are stable and they do not react with other elements. Other elements also tend to attain stable configuration by completing the octet in their outermost orbit. This is important to note that, the number of electrons in the outermost orbit of an element is closer to octet. An element can lose or gain electron in order to complete the octet. This tendency of losing or gaining electrons imparts valency to an element. Let us take example of hydrogen. Hydrogen can readily lose or gain an electron. So, its valency is one. Now, let us take example of Hydrochloric Acid (HCl). One atom of chlorine combines with one atom of hydrogen to form hydrochloric acid. In this case, hydrogen loses one electron and thus gets +1 charge. On the other hand, chlorine gains an electron and thus gets -1 charge. So, valency of hydrogen and chlorine are one.

Valecy can be defined as combining capacity of an atom.

Example :

Hydrogen molecule - Hydrogen has only one electron in its outermost orbit, thus it requires one more electrons to complete its outermost orbit. Therefore, in order to complete outermost orbit, hydrogen shares one electron with another hydrogen atom and form H₂ (hydrogen molecule).

In the case of LiCl (Lithium chloride) - Lithium has three electrons in its outermost orbit and chlorine has seven electrons in its outermost orbit. Thus in order to make outermost orbit completely filled lithium loses one electrons and chlorine gains one electron. After losing one electron, lithium has two electrons in its outermost orbit and after gaining one electron, chlorine has eight electrons in its outermost orbit. And they form LiCl (Lithium chloride)

Name, Symbol, Atomic number, Number of electrons, Distribution of electrons in shells (electronic configuration) and Valency of some elements (From Hydrogen to Calcium)

Elements	Symbol	Atomic Number	No. of electron	Distribution of electron				Valency
				K	L	M	N	
Hydrogen	H	1	1	1				1
Helium	He	2	2	2				0
Lithium	Li	3	3	2	1			1
Beryllium	Be	4	4	2	2			2
Boron	B	5	5	2	3			3
Carbon	C	6	6	2	4			4
Nitrogen	N	7	7	2	5			3
Oxygen	O	8	8	2	6			2
Fluorine	F	9	9	2	7			1
Neon	Ne	10	10	2	8			0
Sodium	Na	11	11	2	8	1		1
Magnesium	Mg	12	12	2	8	2		2
Aluminium	Al	13	13	2	8	3		3
Silicon	Si	14	14	2	8	4		4
Phosphorous	P	15	15	2	8	5		3
Sulphur	S	16	16	2	8	6		2
Chlorine	Cl	17	17	2	8	7		1
Argon	Ar	18	18	2	8	8		0
Potassium	K	19	19	2	8	8	1	1
Calcium	Ca	20	20	2	8	8	2	2

INTEXT QUESTIONS PAGE NO. 52

Q1. How will you find the valency of chlorine, sulphur and magnesium?

Answer:

If the number of electrons in the outermost shell of the atom of an element is less than or equal to 4, then the valency of the element is equal to the number of electrons in the outermost shell. On the other hand, if the number of electrons in the outermost shell of the atom of an element is greater than 4, then the valency of that element is determined by subtracting the number of electrons in the outermost shell from 8.

The distribution of electrons in chlorine, sulphur, and magnesium atoms are 2, 8, 7; 2, 8, 6 and 2, 8, 2 respectively.

Therefore, the number of electrons in the outer most shell of chlorine, sulphur, and magnesium atoms are 7, 6, and 2 respectively.

Thus, the valency of chlorine = $8 - 7 = 1$

The valency of sulphur = $8 - 6 = 2$

The valency of magnesium = 2

INTEXT QUESTIONS PAGE NO. 52

Q1. If number of electrons in an atom is 8 and number of protons is also 8, then (i) what is the atomic number of the atom? and (ii) what is the charge on the atom?

Answer:

(i) The atomic number is equal to the number of protons. Therefore, the atomic number of the atom is 8.

(ii) Since the number of both electrons and protons is equal, therefore, the charge on the atom is 0.

Q2. With the help of Table 4.1, find out the mass number of oxygen and sulphur atom.

Answer:

Mass number of oxygen = Number of protons + Number of neutrons = $8 + 8 = 16$

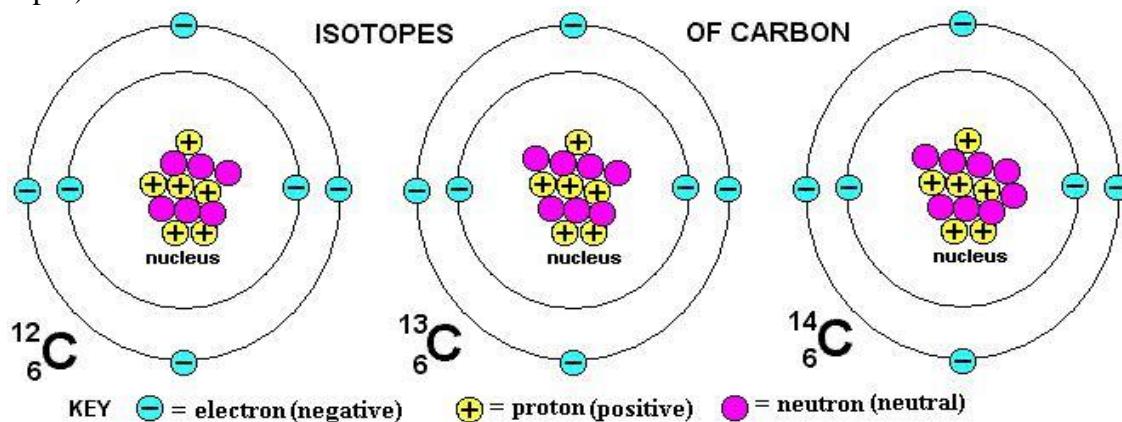
Mass number of sulphur = Number of protons + Number of neutrons = $16 + 16 = 32$

ISOTOPES

Elements having same atomic number but different atomic masses are known as Isotopes.

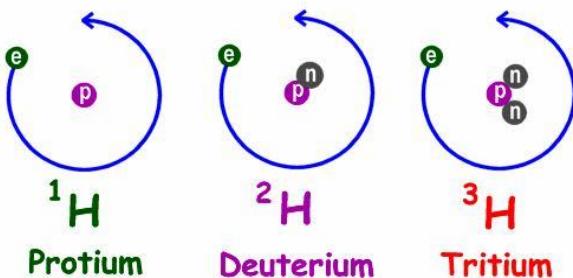
Example –

Carbon-12, Carbon-13, Carbon-14 are thee isotopes of carbon atom. Here 12, 13 and 14 are the atomic masses of isotopes of carbon respectively. Since, atomic number is the unique property of an atom, thus the atomic number of carbon is 6 even in the case of three types of carbon (isotopes)



Hydrogen -1 , Deuterium – 2, Tritium -3 are three isotopes of hydrogen.
The isotopes of hydrogen are written as:

Three Isotopes of Hydrogen



Use of Isotopes:

Carbon – 14 is used in carbon dating.

An isotope of uranium is used as fuel in nuclear reactor.

An isotope of cobalt is used in treatment of cancer.

An isotope of iodine is used in treatment of goitre.

ISOBARS

Atoms having same atomic mass and different atomic numbers are known as Isobars.

Example – $^{40}_{18}Ar$ (argon) and $^{40}_{20}Ca$ (calcium)

Both the elements have same atomic mass equal to 40 but different atomic numbers, i.e. argon has atomic number equal to 18 and calcium has atomic number equal to 20.

INTEXT QUESTIONS PAGE NO. 53

Q1. For the symbol H,D and T tabulate three sub-atomic particles found in each of them.
Answer:

Symbol	Proton	Neutron	Electron
H	1	0	1
D	1	1	1
T	1	2	1

Q2. Write the electronic configuration of any one pair of isotopes and isobars.

Answer:

$^{12}C_6$ and $^{14}C_6$ are isotopes, have the same electronic configuration as (2, 4).

$^{22}Ne_{10}$ and $^{22}Ne_{11}$ are isobars. They have different electronic configuration as given below:

$^{22}Ne_{10}$ – 2, 8

$^{22}Ne_{11}$ – 2, 8, 1

EXERCISE QUESTIONS PAGE NO. 55, 56

Q1. Compare the properties of electrons, protons and neutrons.

Answer:

Particle	Nature of Charge	Mass	Location
Electron	Electrons are negatively charged.	9×10^{-31} kg	Extra nuclear part distributed in different shell or orbits.
Proton	Protons are positively charged.	1.672×10^{-27} kg (1 μ) (approx. 2000 times that of the electron)	Nucleus
Neutron	Neutrons are neutral.	Equal to mass of proton	Nucleus

Q2. What are the limitations of J.J. Thomson's model of the atom?

Answer:

The limitations of J.J. Thomson's model of the atom are:

- It could not explain the result of scattering experiment performed by rutherford.
- It did not have any experiment support.

Q3. What are the limitations of Rutherford's model of the atom?

Answer:

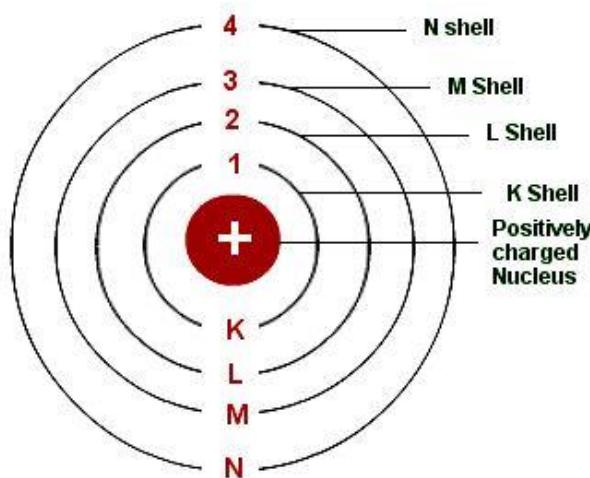
The limitations of Rutherford's model of the atom are

- It failed to explain the stability of an atom.
- It doesn't explain the spectrum of hydrogen and other atoms.

Q4. Describe Bohr's model of the atom.

Answer:

- The atom consists of a small positively charged nucleus at its center.
- The whole mass of the atom is concentrated at the nucleus and the volume of the nucleus is much smaller than the volume of the atom.
- All the protons and neutrons of the atom are contained in the nucleus.
- Only certain orbits known as discrete orbits of electrons are allowed inside the atom.
- While revolving in these discrete orbits electrons do not radiate energy. These orbits or cells are represented by the letters K, L, M, N etc. or the numbers, n = 1, 2, 3, 4, . . . as shown in below figure.



Q5. Compare all the proposed models of an atom given in this chapter.**Answer:**

Thomson's model	Rutherford's model	Bohr's model
<ul style="list-style-type: none"> → An atom consists of a positively charged sphere and the electrons are embedded in it. → The negative and positive charges are equal in magnitude. As a result the atom is electrically neutral. 	<ul style="list-style-type: none"> → An atom consists of a positively charged center in the atom called the nucleus. The mass of the atom is contributed mainly by the nucleus. → The size of the nucleus is very small as compared to the size of the atom. → The electrons revolve around the nucleus in well-defined orbits. 	<ul style="list-style-type: none"> → Bohr agreed with almost all points as said by Rutherford except regarding the revolution of electrons for which he added that there are only certain orbits known as discrete orbits inside the atom in which electrons revolve around the nucleus. → While revolving in its discrete orbits the electrons do not radiate energy.

Q6. Summarise the rules for writing of distribution of electrons in various shells for the first eighteen elements.**Answer:**

The rules for writing of the distribution of electrons in various shells for the first eighteen elements are given below.

→ If n gives the number of orbit or energy level, then $2n^2$ gives the maximum number of electrons possible in a given orbit or energy level. Thus,

First orbit or K-shell will have 2 electrons,

Second orbit or L-shell will have 8 electrons,

Third orbit or M-shell will have 18 electrons.

→ If it is the outermost orbit, then it should have not more than 8 electrons.

→ There should be step-wise filling of electrons in different orbits, i.e., electrons are not accompanied in a given orbit if the earlier orbits or shells are incompletely filled.

Q7. Define valency by taking examples of silicon and oxygen.**Answer:**

The valency of an element is the combining capacity of that element. The valency of an element is determined by the number of valence electrons present in the atom of that element.

→ Valency of Silicon: It has electronic configuration: 2,8,4

Thus, the valency of silicon is 4 as these electrons can be shared with others to complete octet.

→ Valency of Oxygen: It has electronic configuration: 2,6

Thus, the valency of oxygen is 2 as it will gain 2 electrons to complete its octet.

Q8. Explain with examples (i) Atomic number, (ii) Mass number, (iii) Isotopes and iv) Isobars. Give any two uses of isotopes.**Answer:****(i) Atomic number**

The atomic number of an element is the total number of protons present in the atom of that element. For example, nitrogen has 7 protons in its atom. Thus, the atomic number of nitrogen is 7.

(ii) Mass number

The mass number of an element is the sum of the number of protons and neutrons present in the atom of that element. For example, the atom of boron has 5 protons and 6 neutrons. So, the mass number of boron is $5 + 6 = 11$.

(iii) Isotopes

Isotopes are atoms of the same element having the same atomic number, but different mass numbers. For example, hydrogen has three isotopes. They are protium (${}_1^1H$), deuterium (${}_1^2H$) and tritium (${}_1^3H$).

(iv) Isobars

Isobars are atoms having the same mass number, but different atomic numbers i.e., isobars are atoms of different elements having the same mass number. For example, ${}_{20}^{40}Ca$ and ${}_{18}^{40}Ar$ are isobars.

Two uses of isotopes are:

- One isotope of uranium is used as a fuel in nuclear reactors.
- One isotope of cobalt is used in the treatment of cancer.

Q9. Na^+ has completely filled K and L shells. Explain.

Answer:

An atom of Na has a total of 11 electrons. Its electronic configuration is 2, 8, 1. But, Na^+ ion has one electron less than Na atom i.e., it has 10 electrons. Therefore, 2 electrons go to K-shell and 8 electrons go to L-shell, thereby completely filling K and L shells.

Q10. If bromine atom is available in the form of, say, two isotopes ${}_{35}^{79}Br$ (49.7%) and ${}_{35}^{81}Br$ (50.3%), calculate the average atomic mass of bromine atom.

Answer:

It is given that two isotopes of bromine are ${}_{35}^{79}Br$ (49.7%) and ${}_{35}^{81}Br$ (50.3%). Then, the average atomic mass of bromine atom is given by:

$$79 \times \frac{49.7}{100} + 81 \times \frac{50.3}{100} = \frac{3926.3}{100} + \frac{4074.3}{100} \\ = \frac{8000.6}{100} = 80.006u = 80u(\text{approx})$$

Q11. The average atomic mass of a sample of an element X is 16.2 u. What are the percentages of isotopes ${}_{8}^{16}X$ and ${}_{8}^{18}X$ in the sample?

Answer:

It is given that the average atomic mass of the sample of element X is 16.2 u.

Let the percentage of isotope ${}_{8}^{18}X$ be y%. Thus, the percentage of isotope ${}_{8}^{16}X$ will be $(100 - y)\%$.

$$\text{Therefore, } 18 \times \frac{y}{100} + 16 \times \frac{(100-y)}{100} = 16.2 \\ \Rightarrow \frac{18y}{100} + \frac{16(100-y)}{100} = 16.2 \\ \Rightarrow \frac{18y+16(100-y)}{100} = 16.2 \Rightarrow \frac{18y+1600-16y}{100} = 16.2 \\ \Rightarrow 18y+1600-16y = 1620 \\ \Rightarrow 2y+1600 = 1620 \\ \Rightarrow 2y = 1620 - 1600 = 20 \\ \Rightarrow y = 10$$

Therefore, the percentage of isotope ${}_{8}^{18}X$ is 10%.

And, the percentage of isotope ${}_{8}^{16}X$ is $(100 - 10)\% = 90\%$.

Q12. If Z = 3, what would be the valency of the element? Also, name the element.

Answer:

By Z = 3, we mean that the atomic number of the element is 3. Its electronic configuration is 2, 1. Hence, the valency of the element is 1 (since the outermost shell has only one electron). Therefore, the element with Z = 3 is lithium.

Q13. Composition of the nuclei of two atomic species X and Y are given as under

	X	Y
Protons =	6	6
Neutrons =	6	8

Give the mass numbers of X and Y. What is the relation between the two species?

Answer:

Mass number of X = Number of protons + Number of neutrons = 6 + 6 = 12

Mass number of Y = Number of protons + Number of neutrons = 6 + 8 = 14

These two atomic species X and Y have the same atomic number, but different mass numbers. Hence, they are isotopes.

Q14. For the following statements, write T for True and F for False.

- (a) J.J. Thomson proposed that the nucleus of an atom contains only nucleons.
- (b) A neutron is formed by an electron and a proton combining together. Therefore, it is neutral.
- (c) The mass of an electron is about $\frac{1}{2000}$ times that of proton.
- (d) An isotope of iodine is used for making tincture iodine, which is used as a medicine.

Answer:

- (a) False
- (b) False
- (c) True
- (d) False

Put tick (✓) against correct choice and cross (X) against wrong choice in questions Q15, Q16 and Q17

Q15. Rutherford's alpha-particle scattering experiment was responsible for the discovery of

- (a) Atomic Nucleus
- (b) Electron
- (c) Proton
- (d) Neutron

Answer: (a) Atomic nucleus

Q16. Isotopes of an element have

- (a) the same physical properties
- (b) different chemical properties
- (c) different number of neutrons
- (d) different atomic numbers.

Answer: (c) different number of neutrons

Q17. Number of valence electrons in Cl^- ion are:

- (a) 16
- (b) 8
- (c) 17
- (d) 18

Answer: (b) 8

Q18. Which one of the following is a correct electronic configuration of sodium?

- (a) 2,8
- (b) 8,2,1
- (c) 2,1,8
- (d) 2,8,1

Answer: (d) 2, 8, 1

Q19. Complete the following table.

Atomic number	Mass number	No. of neutrons	No. of Protons	No. of electrons	Name of the Atomic Species
9		10			
16	32				Sulphur
	24		12		
	2		1		
	1	0	1	0	

Answer:

Atomic number	Mass number	No. of neutrons	No. of Protons	No. of electrons	Name of the Atomic Species
9	19	10	9	9	Fluorine
16	32	16	16	16	Sulphur
12	24	12	12	12	Magnesium
1	2	1	1	1	Deuterium
1	1	0	1	0	Hydrogen ion



ASSIGNMENT QUESTIONS SET – 1
CHAPTER – 4
STRUCTURE OF ATOMS

1. Plum pudding Model of atom was discovered by _____
2. Combining capacity of an atom is called _____
3. Alpha - particle scattering experiment of Rutherford led to discovery of _____
4. Number of neutrons in $^{81}_{35}Br$ is _____
5. _____ are atoms having the same mass number but different atomic number.
6. The charge on the electron is found to be _____ coulombs.
7. The electronic configuration of silicon is _____. (Atomic number = 14)
8. _____ electrons are responsible for the chemical properties of an atom.
9. Rutherford's model of an atom was modified by _____.
10. _____ is an isotope of carbon used in determining the age of dead plants.
11. An atom with 3 protons and 4 neutrons will have a valency of _____.
12. What do you understand by valency of an element? What is valency of boron?
13. List the features of Rutherford's nuclear model of atom.
14. What are the postulates of Bohr Model of an atom?
15. Define Valency. What is the valency of chlorine, sulphur and magnesium?
16. Cl^- has completely filled K&L shells. Explain.
17. Na^+ is possible but Cl^+ is not possible. What is the reason?
18. What are isotopes? Give two examples .
19. What were the observations of Rutherford's Alpha particles scattering experiment?
20. What are the drawbacks of Rutherford's model of atom?
21. From what observations do you derive the following inferences?
 - (i) The most of the space inside the atom is empty.
 - (ii) The volume of the nucleus is very small.
 - (iii) Anode rays consist of positively charged particles.
22. Name the fundamental particle not present in the nucleus of hydrogen atom
23. Who discovered electrons?
24. Who discovered protons?
25. Who discovered neutrons?
26. Who discovered nucleus of an atom?
27. What are canal rays?
28. What are the properties of anode rays?
29. What are cathode rays?
30. Who discovered X-Rays?

31. If an atom contains one electron and one proton, will it carry any charge or not?

32. Complete the following table.

Particle	Electron	Proton	Neutron
(i) Symbol	_____	_____	_____
(ii) Nature	_____	_____	_____
(iii) Relative Charge	_____		
(iv) Absolute Charge	_____		
(v) Relative Mass	_____		
(vi) Absolute Mass	_____		

33. What is the mass of proton as compared to electron?

34. Describe briefly Thomson's model of an atom.

35. Write the limitations of J.J. Thomson's model of an atom.

36. What are α -particles?

37. On the basis of Rutherford's model of an atom, which subatomic particle is present in the nucleus of an atom?

38. Who is known as 'Father of Nucleus Physics'?

39. What were the observations of Rutherford's α -scattering experiment?

40. What were the important features of atomic model based on Rutherford's scattering experiment?

41. What are the limitations of Rutherford's model of the atom?

42. Draw a sketch of Bohr's model of an atom with three shells.

43. Describe Bohr's model of the atom.

44. What do you think would be the observation if the α -particle scattering experiment is carried out using a foil of a metal other than gold?

45. Write a short note on Nucleus.

46. What is Atomic Number? Who coined this term?

47. What is mass number?

48. If $A = 23$ and $Z = 11$ for Na atom, how many protons, electrons and neutrons present in Na atom?

49. What is ionization energy?

50. In a sample of ethyl ethanoate ($\text{CH}_3\text{COOC}_2\text{H}_5$) the two oxygen atoms have the same number of electrons but different number of neutrons. What can be the reason for it?

51. In the atom of element X, 6 electrons are present in the outermost shell. If it requires an octet configuration by accepting requisite number of electrons, then what would be the charge on the ion so formed?

52. Why do helium, neon & argon have a zero valency?

53. Write the atomic number and the symbol of an element which has mass number 32 and the number of neutrons 16 in the nucleus.

- 54.** Helium atom has an atomic mass of 4 u and two protons in its nucleus. How many neutrons does it have?
- 55.** If number of electrons in an atom is 8 and number of protons is also 8, then (i) what is the atomic number of the atom? (ii) what is the charge on the atom?
- 56.** What are the limitations of J. J. Thomson's model of the atom?
- 57.** Why was Thomson's model of atom discarded and replaced by Rutherford's model? Why is Rutherford's model also called the nuclear model of atom
- 58.** The average atomic mass of a sample of an element X is 16.2u. What are the percentage of isotopes $^{16}_8X$ and $^{18}_8X$ in the sample?
- 59.** What are the salient features of Bohr's Atomic Model? How is it advantageous over Rutherford's Nuclear Model?
- 60.** What is the charge and mass of the anode rays emitted when, hydrogen gas is enclosed in the discharge tube experiment? What is the name of these particles?
- 61.** Do the anode rays always consist of protons whatever be the gas enclosed in the discharge tube? Explain.
- 62.** What is the other name of X-rays?
- 63.** How are x-rays produced? Why was this name given?
- 64.** How are x-rays made tools in diagnostic purposes in the medical field?
- 65.** What is phosphorescence? How does it differ from fluorescence?
- 66.** Who discovered the nuclear model of the atom?
- 67.** Explain how Rutherford's atomic structure cannot explain the stability of atom?
- 68.** Explain how Bohr's atomic model explains the emission and absorption of radiation.
- 69.** What is the charge and mass of a β -particle? From which part of the atom are β -particles emitted?
- 70.** The two elements, A and B have 17 electrons each in their atom. Element A has 35 nucleons while element B has 33 nucleons (nucleons are the total number of protons and neutrons in the nucleus). What is the relationship between these two elements? Explain. The element A reacts with hydrogen in diffused light. Does the element B also react with hydrogen?
- 71.** One isotope of carbon with atomic mass 12, occupies group 14 in the 2nd period in the long form periodic Table. Predict the position of another radioactive isotope of carbon with atomic mass 14?
- 72.** What is the name given to the rays, traveling from cathode to anode?
- 73.** Do the cathode rays always consist of electrons only, whatever be the gas enclosed in the discharge tube?
- 74.** Why are electrons in the outermost shell known as valence electrons?

ASSIGNMENT QUESTIONS SET – 2

CHAPTER – 4

STRUCTURE OF ATOMS

- 22.** In a given electric field, β - particles are deflected more than α - particles inspite of the fact that α - particles have larger charge, why?
- 23.** What are valence electrons? What is their significance?
- 24.** What would be the observation if the α - particle scattering experiment is carried out using a foil of a metal other than gold?
- 25.** Electronic configuration of Potassium is 2,8,8,1 and Calcium 2,8,8,2, when M shell can have maximum of 18 electrons then why next element Scandium has electronic configuration 2,8,9,2 and not 2,8,8,3 ?
- 26.** What are isotopes and Isobars? What are two isotopes of chlorine? Calculate the average atomic mass of a chlorine atom?
- 27.** What is present concept of an atom? Explain in detail? Why this model is considered to be the most appropriate model?
- 28.** Explain the Rutherford's alpha particle scattering experiment. What were the main conclusions drawn from this experiment?
- 29.** A naturally occurring sample of :-
(i) 69.2% of ^{63}Cu & 30.8% of ^{65}Cu . Find the average atomic mass of a naturally occurring sample of copper.
(ii) 7.42% of ^{6}Li & 92.58% of ^{7}Li . Find the average atomic mass of a naturally occurring sample of Lithium.
- 30.** Calculate the no. of atoms of each element present in 9.8 g of sulphuric acid, H_2SO_4 . (H=1, S=32, O=16)
- 31.** How to calculate the atomicity and the atomic mass of an atom?
- 32.** To weight BaCl_2 or Na_2SO_4
a) use a polythene bags and spring balance
b) use a watch glass and spring balance
c) use a polythene bags and physical balance
d) use a watch glass and physical balance
- 33.** To weight sodium sulphate or barium chloride it should be in form of
a) saturated solution b) large crystals c) small crystals d) fine powder
- 34.** What is the distribution of electrons in an atom of Phosphorus and how can it have two valencies
- 35.** The maximum no. of electrons present in shell is given by the formula $2n^2$ and the maximum no. of electrons filled in M shell is 18. But in the element calcium we only fill 8 electrons in the M shell and move on to the N shell. why ?
- 36.** What are alpha particles?
- 37.** Write difference between atomic mass and mass number?
- 38.** An ion (M^{2+}) contain 10 electrons and 12 neutrons what is the atomic number and mass number of the element M.

- 39.** An ion M^{3+} contains 10 electrons and 14 neutrons. What are the atomic mass and mass number of the element M ? Name the element.

40. 10 gm of silver nitrate solution is added to 10 gm of sodium chloride solution. What change in mass do you expect after the reaction and why?

41. Write the atomicity of the following molecules: (i) H_2SO_4 (ii) CCl_4

42. Define the term mole.

43. What is the law of constant proportions?

44. What is molar mass? What are its units?

45. Define atomicity.

46. Calculate the formula unit mass of Na_2CO_3 .
[Atomic mass of Na = 23 u, C = 12 u, O = 16 u]

47. Give the definition of a cation and an anion.

48. An element X has a valency 3. Write the formula of its oxide.

49. Calculate the number of moles is 52 g of He (Helium)

50. If 12 gm of carbon is burnt in the presence of 32 gm of oxygen, how much carbon dioxide will be formed ?

51. Calculate the number of moles in 17 gm of H_2O_2 . (Atomic weight of H = 1 u, O = 16 u).

52. Define and explain atomic mass of an element.

53. If one mole of carbon atoms weighs 12 g. What is the mass of 1 atom of carbon ?

54. Calculate the mass of 1 molecule of oxygen gas.

55. The mass of single atom of an element is 2.65×10^{-23} g. Calculate its atomic mass. [NA = $6.022 \times 10^{23} \text{ mol}^{-1}$]

56. Calculate percentage composition of glucose ($C_6H_{12}O_6$).

57. John placed 10 moles of sulphur molecules (S_8) and 5 moles of glucose ($C_6H_{12}O_6$) in the two different pans of a physical balance. Find which pan of the balance would be heavier, support your answer with calculations.
(atomic mass : O = 16 u, C = 12 u, H = 1 u, S = 32 u)

58. What is Avogadro constant ?

59. Calculate the number of particles present in 56 gm of N_2 molecule.

60. Calculate the number of moles present in (a) 60 g of calcium. (b) 3.011×10^{23} number of oxygen atoms.
[Given that Ca = 40u; Avogadro number, NA = 6.022×10^{23} per mole]

61. Give an account of the ‘mole concept’.

62. Calculate the ratio by number of atoms for Magnesium sulphide.

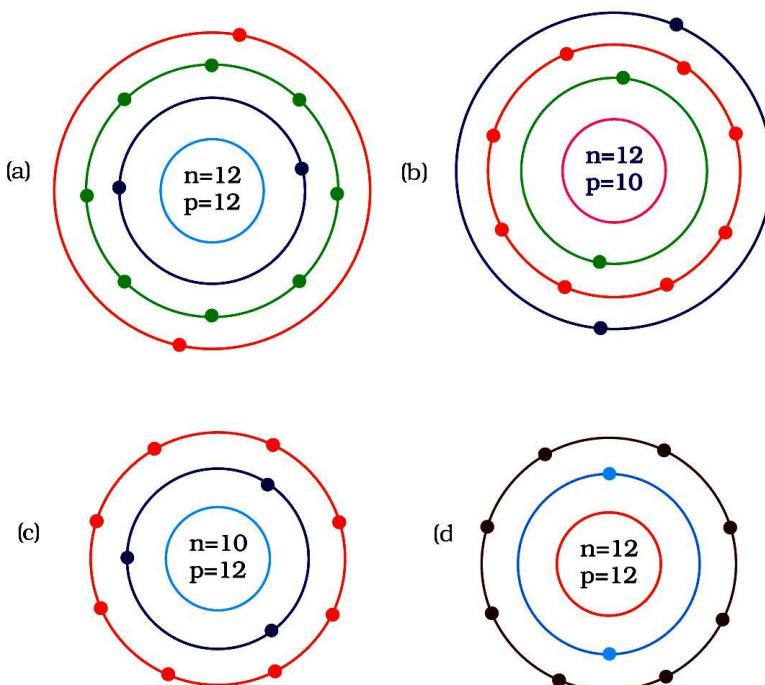
63. (a) The average atomic mass of a sample of an element X is 35.5u. What are the percentages of isotopes $^{37}_{17}X$ and $^{35}_{17}X$ in the sample ?
(b) Write any two applications of isotopes.

- 64.** Nitu presented a silver lamp to her mother on her birthday. The lamp contained 3.011×10^{23} atoms of silver in it. What is the mass of silver lamp and the cost of it if 1 gm silver costs Rs 60. Atomic mass of Ag = 108 u, $N_0 = 6.022 \times 10^{23}$ per mole.
- 65.** Verify by calculating that 5 mole of CO₂ and 5 mole of H₂S do not have the same mass.
(Atomic mass of C = 12 u, O = 16 u, H = 1 u, S = 32 u)
- 66.** A solution is made by dissolving sodium chloride in water and its concentration is expressed as 0.9% by mass. Calculate : (i) the number of moles, and (ii) number of molecules present in NaCl for this solution.
[Given : mass % = (Mass of solute $\times 100$) / Mass of solution]
(Atomic mass of Na = 23.0 u, Cl = 35.5 u)
- 67.** (i) What do the following symbols / formulae stand for :
(a) 2O (b) O₂ (c) O₃ (d) H₂O
(ii) Give the chemical formula of the following compounds :
(a) Potassium carbonate (b) Calcium chloride (iii)
Calculate the formula unit mass of Al₂(SO₄)₃. (Given :
Atomic mass of Al = 27 u, S = 32 u, O = 16 u)
- 68.** An element ${}_{\frac{1}{7}}A$ exists as diatomic gas in nature which is relatively inert and forms 78% of earth's atmosphere.
(a) Identify the gas and write its molecular formula. Write the formulae of its nitrite and nitrate ions.
(b) How many moles of this gas would contain 12.044×10^{23} atoms of this element ?
(Avogadro's no. = 6.022×10^{23})
(c) Calculate the molecular mass of (a) NH₄NO₃ and (b) HNO₃
(Given atomic masses N = 14 u, O = 16 u, H = 1u)
- 69.** (a) In ammonia, nitrogen and hydrogen are always present in the ratio 14 : 3 by mass. State the law which explains the above statement.
(b) During the formation of ammonia, what mass of hydrogen gas would be required to react completely with 42 g of nitrogen gas ?
- 70.** (a) Calculate the number of moles in 112 g of iron.
(b) Calculate the mass of 0.5 moles of sugar (C₁₂H₂₂O₁₁).
(c) Define the term molecular mass.
(d) Determine the molecular mass of ZnSO₄.
[Atomic mass of Zn = 65 u, S = 32 u, O = 16 u]
(e) Calculate the number of molecules of carbon dioxide, present in 4.4 g of CO₂.

ASSIGNMENT QUESTIONS SET – 3
CHAPTER – 4
STRUCTURE OF ATOMS

1. Which of the following correctly represent the electronic distribution in the Mg atom?
(a) 3, 8, 1
(b) 2, 8, 2
(c) 1, 8, 3
(d) 8, 2, 2
2. Rutherford's 'alpha (α) particles scattering experiment' resulted in to discovery of
(a) Electron
(b) Proton
(c) Nucleus in the atom
(d) Atomic mass
3. The number of electrons in an element X is 15 and the number of neutrons is 16. Which of the following is the correct representation of the element?
(a) $^{31}_{15}X$ (b) $^{31}_{16}X$ (c) $^{16}_{15}X$ (d) $^{15}_{16}X$
4. Dalton's atomic theory successfully explained
(i) Law of conservation of mass
(ii) Law of constant composition
(iii) Law of radioactivity
(iv) Law of multiple proportion
(a) (i), (ii) and (iii)
(b) (i), (iii) and (iv)
(c) (ii), (iii) and (iv)
(d) (i), (ii) and (iv)
5. Which of the following statements about Rutherford's model of atom are correct?
(i) considered the nucleus as positively charged
(ii) established that the α -particles are four times as heavy as a hydrogen atom
(iii) can be compared to solar system
(iv) was in agreement with Thomson's model
(a) (i) and (iii)
(b) (ii) and (iii)
(c) (i) and (iv)
(d) only (i)
6. Which of the following are true for an element?
(i) Atomic number = number of protons + number of electrons
(ii) Mass number = number of protons + number of neutrons
(iii) Atomic mass = number of protons = number of neutrons
(iv) Atomic number = number of protons = number of electrons
(a) (i) and (ii)
(b) (i) and (iii)
(c) (ii) and (iii)
(d) (ii) and (iv)

7. In the Thomson's model of atom, which of the following statements are correct?
- the mass of the atom is assumed to be uniformly distributed over the atom
 - the positive charge is assumed to be uniformly distributed over the atom
 - the electrons are uniformly distributed in the positively charged sphere
 - the electrons attract each other to stabilise the atom
- (i), (ii) and (iii)
 - (i) and (iii)
 - (i) and (iv)
 - (i), (iii) and (iv)
8. Rutherford's α -particle scattering experiment showed that
- electrons have negative charge
 - the mass and positive charge of the atom is concentrated in the nucleus
 - neutron exists in the nucleus
 - most of the space in atom is empty
- Which of the above statements are correct?
- (i) and (iii)
 - (ii) and (iv)
 - (i) and (iv)
 - (iii) and (iv)
9. Identify the Mg^{2+} ion from the Fig.4.1 where, n and p represent the number of neutrons and protons respectively



10. In a sample of ethyl ethanoate ($CH_3COOC_2H_5$) the two oxygen atoms have the same number of electrons but different number of neutrons. Which of the following is the correct reason for it?
- One of the oxygen atoms has gained electrons
 - One of the oxygen atoms has gained two neutrons
 - The two oxygen atoms are isotopes
 - The two oxygen atoms are isobars.

11. The ion of an element has 3 positive charges. Mass number of the atom is 27 and the number of neutrons is 14. What is the number of electrons in the ion?

- (a) 13
- (b) 10
- (c) 14
- (d) 16

12. Elements with valency 1 are

- (a) always metals
- (b) always metalloids
- (c) either metals or non-metals
- (d) always non-metals

13. The first model of an atom was given by

- (a) N. Bohr
- (b) E. Goldstein
- (c) Rutherford
- (d) J.J. Thomson

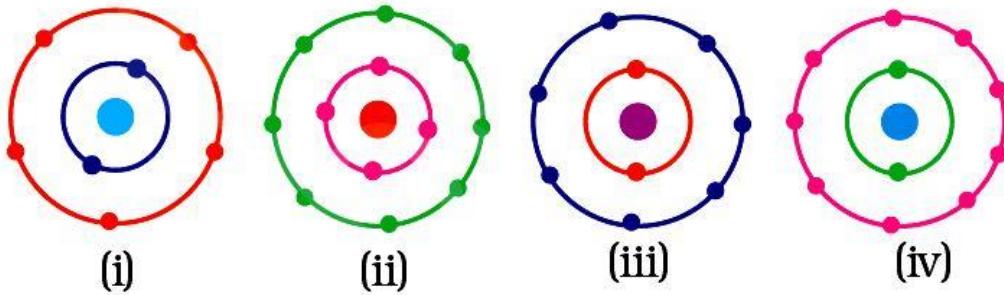
14. An atom with 3 protons and 4 neutrons will have a valency of

- (a) 3
- (b) 7
- (c) 1
- (d) 4

15. The electron distribution in an aluminium atom is

- (a) 2, 8, 3
- (b) 2, 8, 2
- (c) 8, 2, 3
- (d) 2, 3, 8

16. Which of the following in Fig. 4.2 do not represent Bohr's model of an atom correctly?



- (a) (i) and (ii)
- (b) (ii) and (iii)
- (c) (ii) and (iv)
- (d) (i) and (iv)

17. Which of the following statement is always correct?

- (a) An atom has equal number of electrons and protons.
- (b) An atom has equal number of electrons and neutrons.
- (c) An atom has equal number of protons and neutrons.
- (d) An atom has equal number of electrons, protons and neutrons.

- 18.** Atomic models have been improved over the years. Arrange the following atomic models in the order of their chronological order

 - (i) Rutherford's atomic model
 - (ii) Thomson's atomic model
 - (iii) Bohr's atomic model
 - (a) (i), (ii) and (iii)
 - (b) (ii), (iii) and (i)
 - (c) (ii), (i) and (iii)
 - (d) (iii), (ii) and (i)

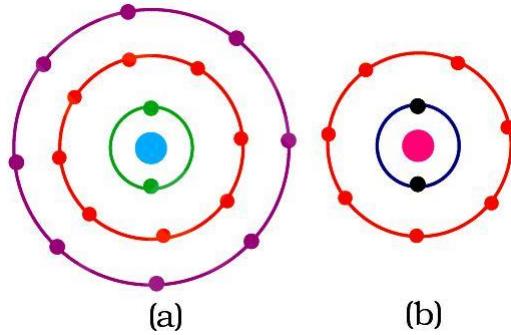
19. Is it possible for the atom of an element to have one electron, one proton and no neutron. If so, name the element.

20. Write any two observations which support the fact that atoms are divisible.

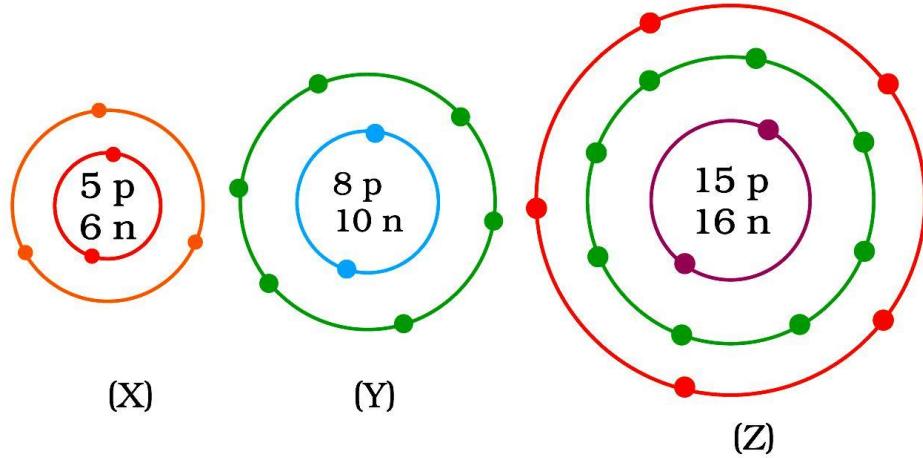
21. Will ^{35}Cl and ^{37}Cl have different valencies? Justify your answer.

22. Why did Rutherford select a gold foil in his α -ray scattering experiment?

23. Find out the valency of the atoms represented by the Fig. 4.3 (a) and (b).



- 24.** What information do you get from the Fig. 4.4 about the atomic number, mass number and valency of atoms X, Y and Z? Give your answer in a tabular form.



- 25.** In response to a question, a student stated that in an atom, the number of protons is greater than the number of neutrons, which in turn is greater than the number of electrons. Do you agree with the statement? Justify your answer.

26. Calculate the number of neutrons present in the nucleus of an element X which is represented as $^{31}_{15}X$

- 27.** One electron is present in the outer most shell of the atom of an element X. What would be the nature and value of charge on the ion formed if this electron is removed from the outer most shell?
- 28.** Write down the electron distribution of chlorine atom. How many electrons are there in the L shell? (Atomic number of chlorine is 17).
- 29.** In the atom of an element X, 6 electrons are present in the outermost shell. If it acquires noble gas configuration by accepting requisite number of electrons, then what would be the charge on the ion so formed?
- 30.** Match the names of the Scientists given in column A with their contributions towards the understanding of the atomic structure as given in column B
- | (A) | (B) |
|-----------------------|-----------------------------|
| (a) Ernest Rutherford | (i) Indivisibility of atoms |
| (b) J.J.Thomson | (ii) Stationary orbits |
| (c) Dalton | (iii) Concept of nucleus |
| (d) Neils Bohr | (iv) Discovery of electrons |
| (e) James Chadwick | (v) Atomic number |
| (f) E. Goldstein | (vi) Neutron |
| (g) Mosley | (vii) Canal rays |
- 31.** The atomic number of calcium and argon are 20 and 18 respectively, but the mass number of both these elements is 40. What is the name given to such a pair of elements?
- 32.** Complete the Table 4.1 on the basis of information available in the symbols given below
- (a) $^{35}_{17}Cl$ (b) $^{12}_6C$ (c) $^{81}_{35}Br$
- | Element | n _p | n _n |
|---------|----------------|----------------|
| | | |
- 33.** Helium atom has 2 electrons in its valence shell but its valency is not 2, Explain.
- 34.** Fill in the blanks in the following statements
- Rutherford's α -particle scattering experiment led to the discovery of the _____.
 - Isotopes have same _____ but different _____.
 - Neon and chlorine have atomic numbers 10 and 17 respectively. Their valencies will be _____ and _____ respectively.
 - The electronic configuration of silicon is _____ and that of sulphur is _____.
- 35.** An element X has a mass number 4 and atomic number 2. Write the valency of this element?
- 36.** Why do Helium, Neon and Argon have a zero valency?
- 37.** The ratio of the radii of hydrogen atom and its nucleus is ~ 105 . Assuming the atom and the nucleus to be spherical, (i) what will be the ratio of their sizes? (ii) If atom is represented by planet earth ' $R_e = 6.4 \times 10^6$ m', estimate the size of the nucleus.
- 38.** Enlist the conclusions drawn by Rutherford from his α -ray scattering experiment.

- 39.** In what way is the Rutherford's atomic model different from that of Thomson's atomic model?
- 40.** What were the drawbacks of Rutherford's model of an atom?
- 41.** What are the postulates of Bohr's model of an atom?
- 42.** Show diagrammatically the electron distributions in a sodium atom and a sodium ion and also give their atomic number.
- 43.** In the Gold foil experiment of Geiger and Marsden, that paved the way for Rutherford's model of an atom, $\sim 1.00\%$ of the α -particles were found to deflect at angles $> 50^\circ$. If one mole of α -particles were bombarded on the gold foil, compute the number of α -particles that would deflect at angles less than 50° .
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