FOR CLASS NINE

SCIENCE

PUNJAB SCHOOL EDUCATION BOARD

Contents

Chapter 1

MATTER IN OUR SURROUNDINGS

As we look at our surroundings, we see a large variety of things with different shapes, sizes and textures. Everything in this universe is made up of material which scientists have named "matter". The air we breathe, the food we eat, stones, clouds, stars, plants and animals, even a small drop of water or a particle of sand – every thing is matter. We can also see as we look around that all the things mentioned above occupy space and have mass. In other words, they have both mass* and volume**.

Since early times, human beings have been trying to understand their surroundings. Early Indian philosophers classified matter in the form of five basic elements – the "*Panch Tatva*"– air, earth, fire, sky and water. According to them everything, living or non-living, was made up of these five basic elements. Ancient Greek philosophers had arrived at a similar classification of matter.

Modern day scientists have evolved two types of classification of matter based on their physical properties and chemical nature.

In this chapter we shall learn about matter based on its physical properties. Chemical aspects of matter will be taken up in subsequent chapters.

1.1 Physical Nature of Matter

1.1.1 MATTER IS MADE UP OF PARTICLES

For a long time, two schools of thought prevailed regarding the nature of matter. One school believed matter to be continuous like a block of wood, whereas, the other thought that matter was made up of particles like sand. Let us perform an activity to decide about the nature of matter – is it continuous or particulate?

Activity _______________ *1.1*

- Take a 100 mL beaker.
- Fill half the beaker with water and mark the level of water.
- Dissolve some salt/ sugar with the help of a glass rod.
- Observe any change in water level.
- What do you think has happened to the salt?
- Where does it disappear?
- Does the level of water change?

In order to answer these questions we need to use the idea that matter is made up of particles. What was there in the spoon, salt or sugar, has now spread throughout water. This is illustrated in Fig. 1.1.

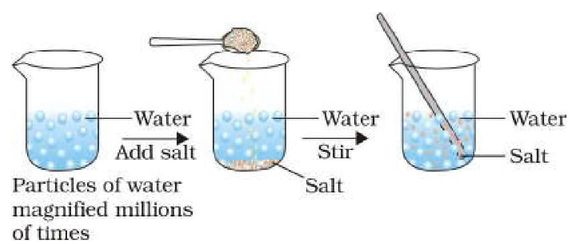


Fig. 1.1: *When we dissolve salt in water, the particles of salt get into the spaces between particles of water.*

1.1.2 HOW SMALL ARE THESE PARTICLES OF MATTER?

Activity _______________ *1.2*

- Take 2-3 crystals of potassium permanganate and dissolve them in 100 mL of water.

---

* The SI unit of mass is kilogram (kg).

** The SI unit of volume is cubic metre (m^3). The common unit of measuring volume is litre (L) such that $1L = 1\ dm^3$, $1L = 1000\ mL$, $1\ mL = 1\ cm^3$.

- Take out approximately 10 mL of this solution and put it into 90 mL of clear water.
- Take out 10 mL of this solution and put it into another 90 mL of clear water.
- Keep diluting the solution like this 5 to 8 times.
- Is the water still coloured ?

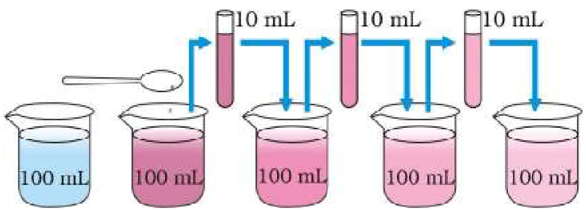


Fig. 1.2: *Estimating how small are the particles of matter. With every dilution, though the colour becomes light, it is still visible.*

This experiment shows that just a few crystals of potassium permanganate can colour a large volume of water (about 1000 L). So we conclude that there must be millions of tiny particles in just one crystal of potassium permanganate, which keep on dividing themselves into smaller and smaller particles.

The same activity can be done using 2 mL of Dettol instead of potassium permanganate. The smell can be detected even on repeated dilution.

> The particles of matter are very small – they are small beyond our imagination!!!!

1.2 Characteristics of Particles of Matter

1.2.1 Particles of matter have space between them

In activities 1.1 and 1.2 we saw that particles of sugar, salt, Dettol, or potassium permanganate got evenly distributed in water. Similarly, 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.

1.2.2 Particles of matter are continuously moving

Activity _______________ 1.3

- Put an unlit incense stick in a corner of your class. How close do you have to go near it so as to get its smell?
- Now light the incense stick. What happens? Do you get the smell sitting at a distance?
- Record your observations.

Activity _______________ 1.4

- Take two glasses/beakers filled with water.
- Put a drop of blue or red ink slowly and carefully along the sides of the first beaker and honey in the same way in the second beaker.
- Leave them undisturbed in your house or in a corner of the class.
- Record your observations.
- What do you observe immediately after adding the ink drop?
- What do you observe immediately after adding a drop of honey?
- How many hours or days does it take for the colour of ink to spread evenly throughout the water?

Activity _______________ 1.5

- Drop a crystal of copper sulphate or potassium permanganate into a glass of hot water and another containing cold water. Do not stir the solution. Allow the crystals to settle at the bottom.
- What do you observe just above the solid crystal in the glass?
- What happens as time passes?
- What does this suggest about the particles of solid and liquid?
- Does the rate of mixing change with temperature? Why and how?

From the above three activities (1.3, 1.4 and 1.5), we can conclude the following:

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.

In the above three activities we observe that particles of matter intermix on their own with each other. They do so by getting into the spaces between the particles. This intermixing of particles of two different types of matter on their own is called diffusion. We also observe that on heating, diffusion becomes faster. Why does this happen?

1.2.3 Particles of matter attract each other

Activity 1.6

- Play this game in the field— make four groups and form human chains as suggested:
- The first group should hold each other from the back and lock arms like Idu-Mishmi dancers (Fig. 1.3).

Fig. 1.3

- The second group should hold hands to form a human chain.
- The third group should form a chain by touching each other with only their finger tips.
- Now, the fourth group of students should run around and try to break the three human chains one by one into as many small groups as possible.
- Which group was the easiest to break? Why?
- If we consider each student as a particle of matter, then in which group the particles held each other with the maximum force?

Activity 1.7

- Take an iron nail, a piece of chalk and a rubber band.
- Try breaking them by hammering, cutting or stretching.
- In which of the above three substances do you think the particles are held together with greater force?

Activity 1.8

- Take some water in a container, try cutting the surface of water with your fingers.
- Were you able to cut the surface of water?
- What could be the reason behind the surface of water remaining together?

The above three activities (1.6, 1.7 and 1.8) suggest that 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.

Questions

1. *Which of the following are matter?*
   Chair, air, love, smell, hate, almonds, thought, cold, lemon water, smell of perfume.
2. *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.
3. *A diver is able to cut through water in a swimming pool. Which property of matter does this observation show?*
4. *What are the characteristics of the particles of matter?*

1.3 States of Matter

Observe different types of matter around you. What are its different states? We can see that matter around us exists in three different states– solid, liquid and gas. These states of matter arise due to the variation in the characteristics of the particles of matter.

Now, let us study about the properties of these three states of matter in detail.

1.3.1 The solid state

Activity 1.9

- Collect the following articles— a pen, a book, a needle and a piece of wooden stick.
- Sketch the shape of the above articles in your notebook by moving a pencil around them.
- Do all these have a definite shape, distinct boundaries and a fixed volume?
- What happens if they are hammered, pulled or dropped?
- Are these capable of diffusing into each other?
- Try compressing them by applying force. Are you able to compress them?

All the above are examples of solids. We can observe that all these 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.

Consider the following:

(a) What about a rubber band, can it change its shape on stretching? Is it a solid?

(b) What about sugar and salt? When kept in different jars these take the shape of the jar. Are they solid?

(c) What about a sponge? It is a solid yet we are able to compress it. Why?

All the above are solids as:

- A rubber band changes shape under force and regains the same shape when the force is removed. If excessive force is applied, it breaks.
- The shape of each individual sugar or salt crystal remains fixed, whether we take it in our hand, put it in a plate or in a jar.
- A sponge has minute holes, in which air is trapped, when we press it, the air is expelled out and we are able to compress it.

1.3.2 The liquid state

Activity 1.10

- Collect the following:
  (a) water, cooking oil, milk, juice, a cold drink.
  (b) containers of different shapes. Put a 50 mL mark on these containers using a measuring cylinder from the laboratory.
- What will happen if these liquids are spilt on the floor?
- Measure 50 mL of any one liquid and transfer it into different containers one by one. Does the volume remain the same?
- Does the shape of the liquid remain the same ?
- When you pour the liquid from one container into another, does it flow easily?

We observe that 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.

Refer to activities 1.4 and 1.5 where we saw that solids and liquids can diffuse into liquids. The gases from the atmosphere diffuse and dissolve in water. These gases, especially oxygen and carbon dioxide, are essential for the survival of aquatic animals and plants.

All living creatures need to breathe for survival. The aquatic animals can breathe under water due to the presence of dissolved oxygen in water. Thus, we may conclude that solids, liquids and gases can diffuse into liquids. 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.

1.3.3 The gaseous state

Have you ever observed a balloon seller filling a large number of balloons from a single cylinder of gas? Enquire from him how many balloons is he able to fill from one cylinder. Ask him which gas does he have in the cylinder.

Activity 1.11

- Take three 100 mL syringes and close their nozzles by rubber corks, as shown in Fig.1.4.
- Remove the pistons from all the syringes.
- Leaving one syringe untouched, fill water in the second and pieces of chalk in the third.
- Insert the pistons back into the syringes. You may apply some vaseline on the pistons before inserting them into the syringes for their smooth movement.
- Now, try to compress the content by pushing the piston in each syringe.

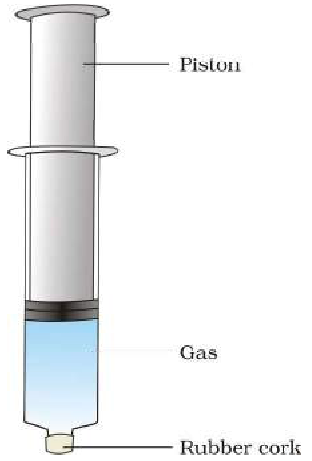


Fig. 1.4

- What do you observe? In which case was the piston easily pushed in?
- What do you infer from your observations?

We have observed that 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. Compressed natural gas (CNG) is used as fuel these days in vehicles. Due to its high compressibility, large volumes of a gas can be compressed into a small cylinder and transported easily.

We come to know of what is being cooked in the kitchen without even entering there, by the smell that reaches our nostrils. How does this smell reach us? The particles of the aroma of food mix with the particles of air spread from the kitchen, reach us and even farther away. The smell of hot cooked food reaches us in seconds; compare this with the rate of diffusion of solids and liquids. Due to high speed of particles and large space between them, gases show the property of diffusing very fast into other gases.

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. The pressure exerted by the gas is because of this force exerted by gas particles per unit area on the walls of the container.

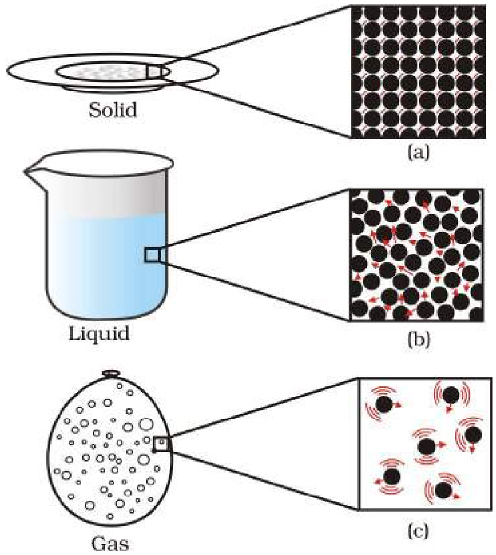


Fig.1.5: *a, b and c show the magnified schematic pictures of the three states of matter. The motion of the particles can be seen and compared in the three states of matter.*

Questions

1. *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.*
2. *(a) Tabulate the differences in the characterisitcs of states of matter.*
   (b) Comment upon the following: rigidity, compressibility, fluidity, filling a gas container, shape, kinetic energy and density.
3. *Give reasons*
   (a) A gas fills completely the vessel in which it is kept.
   (b) A gas exerts pressure on the walls of the container.
   (c) 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.
4. *Liquids generally have lower density as compared to solids. But you must have observed that ice floats on water. Find out why.*

1.4 Can Matter Change its State?

We all know from our observation that water can exist in three states of matter–

- solid, as ice,
- liquid, as the familiar water, and
- gas, as water vapour.

What happens inside the matter during this change of state? What happens to the particles of matter during the change of states? How does this change of state take place? We need answers to these questions, isn't it?

1.4.1 Effect of change of temperature

Activity _______________ 1.12

- Take about 150 g of ice in a beaker and suspend a laboratory thermometer so that its bulb is in contact with the ice, as in Fig. 1.6.

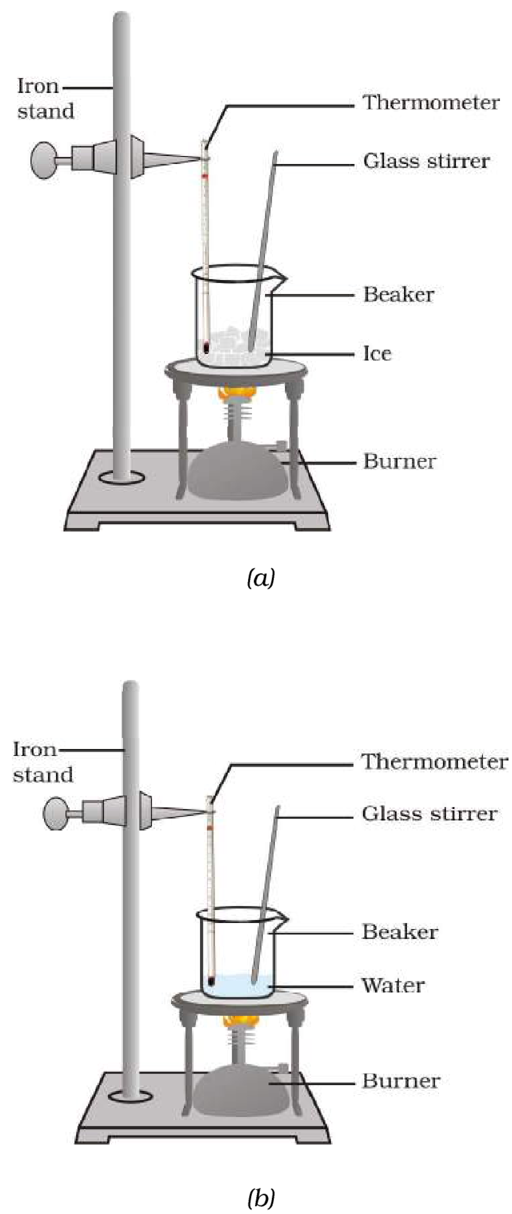


Fig. 1.6: *(a) Conversion of ice to water, (b) conversion of water to water vapour*

- Start heating the beaker on a low flame.
- Note the temperature when the ice starts melting.
- Note the temperature when all the ice has converted into water.
- Record your observations for this conversion of solid to liquid state.
- Now, put a glass rod in the beaker and heat while stirring till the water starts boiling.
- Keep a careful eye on the thermometer reading till most of the water has vaporised.
- Record your observations for the conversion of water in the liquid state to the gaseous state.

On increasing the temperature of solids, the kinetic energy of the particles increases. Due to the increase in kinetic energy, the particles start vibrating with greater speed. The energy supplied by heat overcomes the forces of attraction between the particles. The particles leave their fixed positions and start moving more freely. A stage is reached when the solid melts and is converted to a liquid. The minimum temperature at which a solid melts to become a liquid at the atmospheric pressure is called its melting point.

> The melting point of a solid is an indication of the strength of the force of attraction between its particles.

The melting point of ice is 273.15 K*. The process of melting, that is, change of solid state into liquid state is also known as fusion. **When a solid melts, its temperature remains the same, so where does the heat energy go?**

You must have observed, during the experiment of melting, that the temperature of the system does not change after the melting point is reached, till all the ice melts. This happens even though we continue to heat the beaker, that is, we continue to supply heat. This heat gets used up in changing the state by overcoming the forces of attraction between the particles. As this heat energy is absorbed by ice without showing any rise in temperature, it is considered that it gets hidden into the contents of the beaker and is known as the latent heat. The word latent means hidden. The amount of heat energy that is required to change 1 kg of a solid into liquid at atmospheric pressure at its melting point is known as the latent heat of fusion. So, particles in water at 0^0 C (273 K) have more energy as compared to particles in ice at the same temperature.

When we supply heat energy to water, particles start moving even faster. At a certain temperature, a point is reached when the particles have enough energy to break free from the forces of attraction of each other. At this temperature the liquid starts changing into gas. The temperature at which a liquid starts boiling at the atmospheric pressure is known as its boiling point. Boiling is a bulk phenomenon. Particles from the bulk of the liquid gain enough energy to change into the vapour state.

For water this temperature is 373 K ($100\ ^0$C = 273 + 100 = 373 K).

Can you define the latent heat of vaporisation? Do it in the same way as we have defined the latent heat of fusion. Particles in steam, that is, water vapour at 373 K (100^0 C) have more energy than water at the same temperature. This is because particles in steam have absorbed extra energy in the form of latent heat of vaporisation.

$$\text{Solid State} \underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}} \text{Liquid State} \underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}} \text{Gaseous State}$$

So, we infer that the state of matter can be changed into another state by changing the temperature.

We have learnt that substances around us change state from solid to liquid and from liquid to gas on application of heat. But there

*Note: *Kelvin* is the SI unit of temperature, 0^0 C =273.15 K. For convenience, we take 0^0 C = 273 K after rounding off the decimal. To change a temperature on the Kelvin scale to the Celsius scale you have to subtract 273 from the given temperature, and to convert a temperature on the Celsius scale to the Kelvin scale you have to add 273 to the given temperature.

are some that change directly from solid state to gaseous state and vice versa without changing into the liquid state.

Activity _____________ 1.13

- Take some camphor or ammonium chloride. Crush it and put it in a china dish.
- Put an inverted funnel over the china dish.
- Put a cotton plug on the stem of the funnel, as shown in Fig. 1.7.

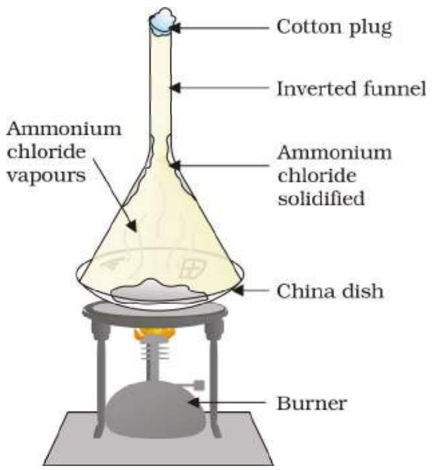


Fig. 1.7: *Sublimation of ammonium chloride*

- Now, heat slowly and observe.
- What do you infer from the above activity?

A change of state directly from solid to gas without changing into liquid state is called sublimation and the direct change of gas to solid without changing into liquid is called deposition.

1.4.2 EFFECT OF CHANGE OF PRESSURE

We have already learnt that the difference in various states of matter is due to the difference in the distances between the constituent particles. What will happen when we start putting pressure and compress a gas enclosed in a cylinder? Will the particles come closer? Do you think that increasing or decreasing the pressure can change the state of matter?

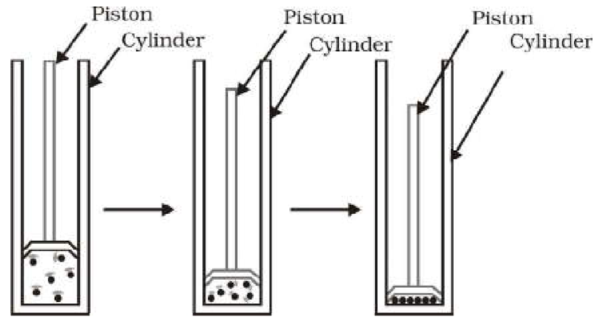


Fig. 1.8: *By applying pressure, particles of matter can be brought close together.*

Applying pressure and reducing temperature can liquefy gases.

Have you heard of solid carbon dioxide (CO_2)? It is stored under high pressure. Solid CO_2 gets converted directly to gaseous state on decrease of pressure to 1 atmosphere* without coming into liquid state. This is the reason that solid carbon dioxide is also known as dry ice.

Thus, we can say that pressure and temperature determine the state of a substance, whether it will be solid, liquid or gas.

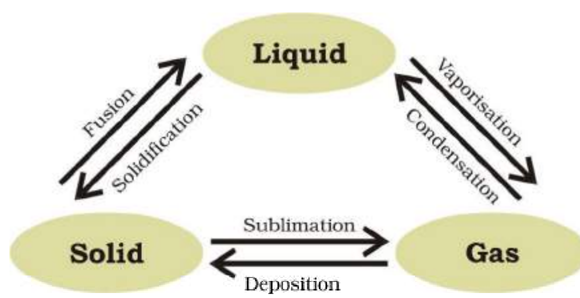


Fig. 1.9: *Interconversion of the three states of matter*

* *atmosphere (atm)* is a unit of measuring pressure exerted by a gas. The unit of pressure is Pascal (Pa): 1 atmosphere = 1.01×10^5 Pa. The pressure of air in atmosphere is called atmospheric pressure. The atmospheric pressure at sea level is 1 atmosphere, and is taken as the normal atmospheric pressure.

Questions

1. *Convert the following temperature to celsius scale:*
   a. 300 K b. 573 K.
2. *What is the physical state of water at:*
   a. 250ºC b. 100ºC ?
3. *For any substance, why does the temperature remain constant during the change of state?*
4. *Suggest a method to liquefy atmospheric gases.*

1.5 Evaporation

Do we always need to heat or change pressure for changing the state of matter? Can you quote some examples from everyday life where change of state from liquid to vapour takes place without the liquid reaching the boiling point? Water, when left uncovered, slowly changes into vapour. Wet clothes dry up. What happens to water in the above two examples?

We know that particles of matter are always moving and are never at rest. At a given temperature in any gas, liquid or solid, there are particles with different amounts of kinetic energy. In the case of liquids, a small fraction of particles at the surface, having higher kinetic energy, is able to break away from the forces of attraction of other particles and gets converted into vapour. This phenomenon of change of a liquid into vapours at any temperature below its boiling point is called evaporation.

1.5.1 Factors affecting evaporation

Let us understand this with an activity.

Activity 1.14

- Take 5 mL of water in a test tube and keep it near a window or under a fan.
- Take 5 mL of water in an open china dish and keep it near a window or under a fan.
- Take 5 mL of water in an open china dish and keep it inside a cupboard or on a shelf in your class.
- Record the room temperature.
- Record the time or days taken for the evaporation process in the above cases.
- Repeat the above three steps of activity on a rainy day and record your observations.
- What do you infer about the effect of temperature, surface area and wind velocity (speed) on evaporation?

You must have observed that the rate of evaporation increases with–

- an increase of surface area:
  We know that evaporation is a surface phenomenon. If the surface area is increased, the rate of evaporation increases. For example, while putting clothes for drying up we spread them out.
- an increase of temperature:
  With the increase of temperature, more number of particles get enough kinetic energy to go into the vapour state.
- a decrease in humidity:
  Humidity is the amount of water vapour present in air. The air around us cannot hold more than a definite amount of water vapour at a given temperature. If the amount of water in air is already high, the rate of evaporation decreases.
- an increase in wind speed:
  It is a common observation that clothes dry faster on a windy day. With the increase in wind speed, the particles of water vapour move away with the wind, decreasing the amount of water vapour in the surrounding.

1.5.2 How does evaporation cause cooling?

In an open vessel, the liquid keeps on evaporating. The particles of liquid absorb energy from the surrounding to regain the energy lost during evaporation. This absorption of energy from the surroundings make the surroundings cold.

What happens when you pour some acetone (nail polish remover) on your palm? The particles gain energy from your palm or surroundings and evaporate causing the palm to feel cool.

After a hot sunny day, people sprinkle water on the roof or open ground because the large latent heat of vaporisation of water helps to cool the hot surface.

Can you cite some more examples from daily life where we can feel the effect of cooling due to evaporation?

Why should we wear cotton clothes in summer?

During summer, we perspire more because of the mechanism of our body which keeps us cool. We know that during evaporation, the particles at the surface of the liquid gain energy from the surroundings or body surface and change into vapour. The heat energy equal to the latent heat of vaporisation is absorbed from the body leaving the body cool. Cotton, being a good absorber of water helps in absorbing the sweat and exposing it to the atmosphere for easy evaporation.

Why do we see water droplets on the outer surface of a glass containing ice-cold water?

Let us take some ice-cold water in a tumbler. Soon we will see water droplets on the outer surface of the tumbler. The water vapour present in air, on coming in contact with the cold glass of water, loses energy and gets converted to liquid state, which we see as water droplets.

Questions

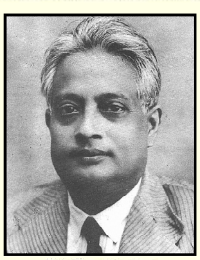
1. *Why does a desert cooler cool better on a hot dry day?*
2. *How does the water kept in an earthen pot (matka) become cool during summer?*
3. *Why does our palm feel cold when we put some acetone or petrol or perfume on it?*
4. *Why are we able to sip hot tea or milk faster from a saucer rather than a cup?*
5. *What type of clothes should we wear in summer?*

More to know

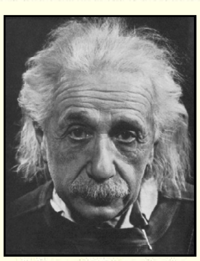
Now scientists are talking of five states of matter: Solid, Liquid, Gas, Plasma and Bose-Einstein Condensate.

Plasma: The state consists of super energetic and super excited particles. These particles are in the form of ionised gases. The fluorescent tube and neon sign bulbs consist of plasma. Inside a neon sign bulb there is neon gas and inside a fluorescent tube there is helium gas or some other gas. The gas gets ionised, that is, gets charged when electrical energy flows through it. This charging up creates a plasma glowing inside the tube or bulb. The plasma glows with a special colour depending on the nature of gas. The Sun and the stars glow because of the presence of plasma in them. The plasma is created in stars because of very high temperature.

Bose-Einstein Condensate: In 1920, Indian physicist Satyendra Nath Bose had done some calculations for a fifth state of matter. Building on his calculations, Albert Einstein predicted a new state of matter – the Bose-Einstein Condensate (BEC). In 2001, Eric A. Cornell, Wolfgang Ketterle and Carl E. Wieman of USA received the Nobel prize in physics for achieving "Bose-Einstein condensation". The BEC is formed by cooling a gas of extremely low density, about one-hundred-thousandth the density of normal air, to super low temperatures. You can log on to www.chem4kids.com to get more information on these fourth and fifth states of matter.



S.N. Bose (*1894-1974*)



Albert Einstein (*1879-1955*)

What you have learnt

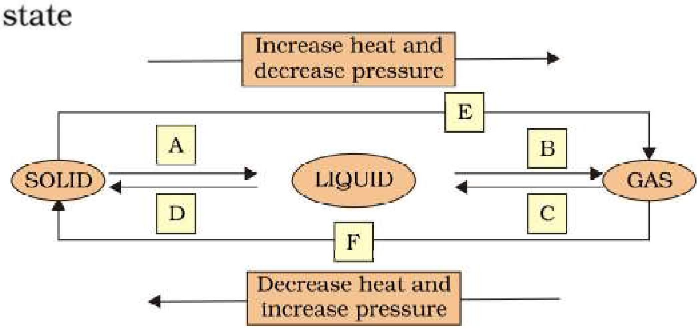
- Matter is made up of small particles.
- The matter around us exists in three states— solid, liquid and gas.
- The forces of attraction between the particles are maximum in solids, intermediate in liquids and minimum in gases.
- The spaces in between the constituent particles and kinetic energy of the particles are minimum in the case of solids, intermediate in liquids and maximum in gases.
- The arrangement of particles is most ordered in the case of solids, in the case of liquids layers of particles can slip and slide over each other while for gases, there is no order, particles just move about randomly.
- The states of matter are inter-convertible. The state of matter can be changed by changing temperature or pressure.
- Sublimation is the change of solid state directly to gaseous state without going through liquid state.
- Deposition is the change of gaseous state directly to solid state without going through liquid state.
- Boiling is a bulk phenomenon. Particles from the bulk (whole) of the liquid change into vapour state.
- Evaporation is a surface phenomenon. Particles from the surface gain enough energy to overcome the forces of attraction present in the liquid and change into the vapour state.
- The rate of evaporation depends upon the surface area exposed to the atmosphere, the temperature, the humidity and the wind speed.
- Evaporation causes cooling.
- Latent heat of vaporisation is the heat energy required to change 1 kg of a liquid to gas at atmospheric pressure at its boiling point.
- Latent heat of fusion is the amount of heat energy required to change 1 kg of solid into liquid at its melting point.

- Some measurable quantities and their units to remember:

Quantity	Unit	Symbol
Temperature	kelvin	K
Length	metre	m
Mass	kilogram	kg
Weight	newton	N
Volume	cubic metre	m^3
Density	kilogram per cubic metre	$kg\ m^{-3}$
Pressure	pascal	Pa

Exercises

1. Convert the following temperatures to the celsius scale.
   (a) 293 K (b) 470 K.
2. Convert the following temperatures to the kelvin scale.
   (a) 25°C (b) 373°C.
3. 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.
4. Arrange the following substances in increasing order of forces of attraction between the particles— water, sugar, oxygen.
5. What is the physical state of water at—
   (a) 25°C (b) 0°C (c) 100°C ?
6. Give two reasons to justify—
   (a) water at room temperature is a liquid.
   (b) an iron almirah is a solid at room temperature.
7. Why is ice at 273 K more effective in cooling than water at the same temperature?
8. What produces more severe burns, boiling water or steam?
9. Name A,B,C,D,E and F in the following diagram showing change in its state



Group Activity

Prepare a model to demonstrate movement of particles in solids, liquids and gases.

For making this model you will need

- A transparent jar
- A big rubber balloon or piece of stretchable rubber sheet
- A string
- Few chick-peas or black gram or dry green peas.

How to make?

- Put the seeds in the jar.
- Sew the string to the centre of the rubber sheet and put some tape to keep it tied securely.
- Stretch and tie the rubber sheet on the mouth of the jar.
- Your model is ready. Now run your fingers up and down the string by first tugging at it slowly and then rapidly.

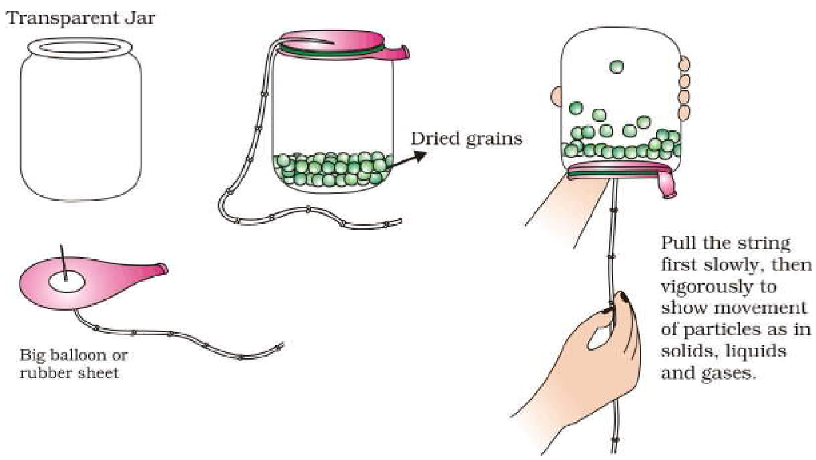


Fig. 1.10: *A model for converting of solid to liquid and liquid to gas.*

C hapter 2

Is Matter Around Us Pure?

How do we judge whether milk, ghee, butter, salt, spices, mineral water or juice that we buy from the market are pure?

Fig. 2.1: *Some consumable items*

Have you ever noticed the word 'pure' written on the packs of these consumables? For a common person pure means having no adulteration. But, for a scientist all these things are actually mixtures of different substances and hence not pure. For example, milk is actually a mixture of water, fat, proteins etc. When a scientist says that something is pure, it means that all the constituent particles of that substance are the same in their chemical nature. A pure substance consists of a single type of particles. In other words, a substance is a pure single form of matter.

As we look around, we can see that most of the matter around us exist as mixtures of two or more pure components, for example, sea water, minerals, soil etc. are all mixtures.

2.1 What is a Mixture?

Mixtures are constituted by more than one kind of pure form of matter. We know that dissolved sodium chloride can be separated from water by the physical process of evaporation. However, sodium chloride is itself a pure substance and cannot be separated by physical process into its chemical constituents. Similarly, sugar is a substance which contains only one kind of pure matter and its composition is the same throughout.

Soft drink and soil are not single pure substances. Whatever the source of a substance may be, it will always have the same characteristic properties.

Therefore, we can say that a mixture contains more than one pure substance.

2.1.1 Types of mixtures

Depending upon the nature of the components that form a mixture, we can have different types of mixtures.

Activity _______________ 2.1

- Let us divide the class into groups A, B, C and D.
- Group A takes a beaker containing 50 mL of water and one spatula full of copper sulphate powder. Group B takes 50 mL of water and two spatula full of copper sulphate powder in a beaker.
- Groups C and D can take different amounts of copper sulphate and potassium permanganate or common salt (sodium chloride) and mix the given components to form a mixture.
- Report the observations on the uniformity in colour and texture.
- Groups A and B have obtained a mixture which has a uniform composition throughout. Such mixtures are called homogeneous mixtures or solutions. Some other examples of such mixtures are: (i) salt dissolved in water and (ii) sugar

dissolved in water. Compare the colour of the solutions of the two groups. Though both the groups have obtained copper sulphate solution but the intensity of colour of the solutions is different. This shows that a homogeneous mixture can have a variable composition.

- Groups C and D have obtained mixtures, which contain physically distinct parts and have non-uniform compositions. Such mixtures are called heterogeneous mixtures. Mixtures of sodium chloride and iron filings, salt and sulphur, and oil and water are examples of heterogeneous mixtures.

Activity ______________ *2.2*

- Let us again divide the class into four groups – A, B, C and D.
- Distribute the following samples to each group:
  - Few crystals of copper sulphate to group A.
  - One spatula full of copper sulphate to group B.
  - Chalk powder or wheat flour to group C.
  - Few drops of milk or ink to group D.
- Each group should add the given sample in water and stir properly using a glass rod. Are the particles in the mixture visible?
- Direct a beam of light from a torch through the beaker containing the mixture and observe from the front. Was the path of the beam of light visible?
- Leave the mixtures undisturbed for a few minutes (and set up the filtration apparatus in the meantime). Is the mixture stable or do the particles begin to settle after some time?
- Filter the mixture. Is there any residue on the filter paper?

  Discuss the results and form an opinion.
- *Groups A and B have got a solution.*
- *Group C has got a suspension.*
- *Group D has got a colloidal solution.*

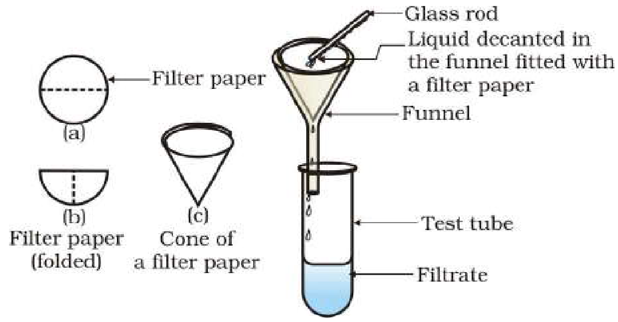


Fig. 2.2: *Filtration*

Now, we shall learn about solutions, suspensions and colloidal solutions in the following sections.

uestions

1. *What is meant by a substance?*
2. *List the points of differences between homogeneous and heterogeneous mixtures.*

2.2 What is a Solution?

A solution is a homogeneous mixture of two or more substances. You come across various types of solutions in your daily life. Lemonade, soda water etc. are all examples of solutions. Usually we think of a solution as a liquid that contains either a solid, liquid or a gas dissolved in it. But, we can also have solid solutions (alloys) and gaseous solutions (air). In a solution there is homogeneity at the particle level. For example, lemonade tastes the same throughout. This shows that particles of sugar or salt are evenly distributed in the solution.

More to know

Alloys: Alloys are mixtures of two or more metals or a metal and a non-metal and cannot be separated into their components by physical methods. But still, an alloy is considered as a mixture because it shows the properties of its constituents and can have variable composition. For example, brass is a mixture of approximately 30% zinc and 70% copper.

A solution has a solvent and a solute as its components. The component of the solution that dissolves the other component in it (usually the component present in larger amount) is called the solvent. The component of the solution that is dissolved in the solvent (usually present in lesser quantity) is called the solute.

Examples:

(i) A solution of sugar in water is a solid in liquid solution. In this solution, sugar is the solute and water is the solvent.

(ii) A solution of iodine in alcohol known as 'tincture of iodine', has iodine (solid) as the solute and alcohol (liquid) as the solvent.

(iii) Aerated drinks like soda water etc., are gas in liquid solutions. These contain carbon dioxide (gas) as solute and water (liquid) as solvent.

(iv) Air is a mixture of gas in gas. Air is a homogeneous mixture of a number of gases. Its two main constituents are: oxygen (21%) and nitrogen (78%). The other gases are present in very small quantities.

Properties of a solution

- A solution is a homogeneous mixture.
- The particles of a solution are smaller than 1 nm (10^{-9} metre) in diameter. So, they cannot be seen by naked eyes.
- Because of very small particle size, they do not scatter a beam of light passing through the solution. So, the path of light is not visible in a solution.
- The solute particles cannot be separated from the mixture by the process of filtration. The solute particles do not settle down when left undisturbed, that is, a solution is stable.

2.2.1 Concentration of a solution

In activity 2.2, we observed that groups A and B obtained different shades of solutions. So, we understand that in a solution the relative proportion of the solute and solvent can be varied. Depending upon the amount of solute present in a solution, it can be called a dilute, concentrated or a saturated solution. Dilute and concentrated are comparative terms. In activity 2.2, the solution obtained by group A is dilute as compared to that obtained by group B.

Activity _______________ 2.3

- Take approximately 50 mL of water each in two separate beakers.
- Add salt in one beaker and sugar or barium chloride in the second beaker with continuous stirring.
- When no more solute can be dissolved, heat the contents of the beaker to raise the temperature by about 5°C.
- Start adding the solute again.

Is the amount of salt and sugar or barium chloride, that can be dissolved in water at a given temperature, the same?

At any particular temperature, a solution that has dissolved as much solute as it is capable of dissolving, is said to be a saturated solution. In other words, when no more solute can be dissolved in a solution at a given temperature, it is called a saturated solution. The amount of the solute present in the saturated solution at this temperature is called its solubility.

If the amount of solute contained in a solution is less than the saturation level, it is called an unsaturated solution.

What would happen if you were to take a saturated solution at a certain temperature and cool it slowly.

We can infer from the above activity that different substances in a given solvent have different solubilities at the same temperature.

The concentration of a solution is the amount (mass or volume) of solute present in a given amount (mass or volume) of solution.

There are various ways of expressing the concentration of a solution, but here we will learn only three methods.

(i) Mass by mass percentage of a solution

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

(ii) Mass by volume percentage of a solution

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

(iii) Volume by volume percentage of a solution

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

Example 2.1 A solution contains 40 g of common salt in 320 g of water. Calculate the concentration in terms of mass by mass percentage of the solution.

Solution:

Mass of solute (salt) = 40 g
Mass of solvent (water) = 320 g

We know,

Mass of solution = Mass of solute + Mass of solvent
= 40 g + 320 g
= 360 g

Mass percentage of solution

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

$$= \frac{40}{360} \times 100 = 11.1\%$$

2.2.2 What is a suspension?

Non-homogeneous systems, like those obtained by group C in activity 2.2, in which solids are dispersed in liquids, are called suspensions. 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. When the particles settle down, the suspension breaks and it does not scatter light any more.

2.2.3 What is a colloidal solution?

The mixture obtained by group D in activity 2.2 is called a colloid or a colloidal solution. The particles of a colloid are uniformly spread throughout the solution. Due to the relatively smaller size of particles, as compared to that of a suspension, the mixture appears to be homogeneous. But actually, a colloidal solution is a heterogeneous mixture, for example, milk.

Because of the small size of colloidal particles, we cannot see them with naked eyes. But, these particles can easily scatter a beam of visible light as observed in activity 2.2. This scattering of a beam of light is called the Tyndall effect after the name of the scientist who discovered this effect.

Tyndall effect can also be observed when a fine beam of light enters a room through a small hole. This happens due to the scattering of light by the particles of dust and smoke in the air.

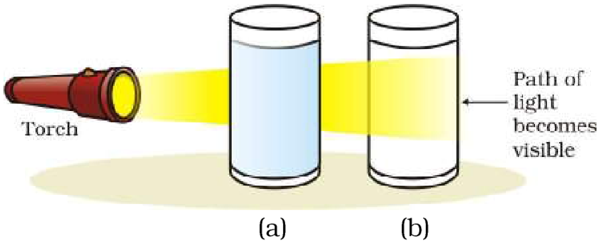


Fig. 2.3: *(a) Solution of copper sulphate does not show Tyndall effect, (b) mixture of water and milk shows Tyndall effect.*

Tyndall effect can be observed when sunlight passes through the canopy of a dense forest. In the forest, mist contains tiny droplets of water, which act as particles of colloid dispersed in air.

Fig. 2.4: *The Tyndall effect*

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. But, a special technique of separation known as centrifugation (perform activity 2.5), can be used to separate the colloidal particles.

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. A few common examples are given in Table 2.1. From this table you can see that they are very common everyday life.

Questions

1. *Differentiate between homogeneous and heterogeneous mixtures with examples.*
2. *How are sol, solution and suspension different from each other?*
3. *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.*

Table 2.1: 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

2.3 Separating the Components of a Mixture

We have learnt that most of the natural substances are not chemically pure. Different methods of separation are used to get individual components from a mixture. Separation makes it possible to study and use the individual components of a mixture.

Heterogeneous mixtures can be separated into their respective constituents by simple physical methods like handpicking, sieving, filtration that we use in our day-to-day life. Sometimes special techniques have to be used for the separation of the components of a mixture.

2.3.1 How can we obtain coloured component (dye) from blue/black ink?

Activity _______________ 2.4

- Fill half a beaker with water.
- Put a watch glass on the mouth of the beaker (Fig. 2.5).
- Put few drops of ink on the watch glass.
- Now start heating the beaker. We do not want to heat the ink directly. You will see that evaporation is taking place from the watch glass.
- Continue heating as the evaporation goes on and stop heating when you do not see any further change on the watch glass.
- Observe carefully and record your observations.

Fig. 2.5: *Evaporation*

Now answer

- *What do you think has got evaporated from the watch glass?*
- *Is there a residue on the watch glass?*
- *What is your interpretation? Is ink a single substance (pure) or is it a mixture?*

We find that ink is a mixture of a dye in water. Thus, we can separate the volatile component (solvent) from its non-volatile solute by the method of evaporation.

2.3.2 How can we separate cream from milk?

Now-a-days, we get full-cream, toned and double-toned varieties of milk packed in poly-packs or tetra packs in the market. These varieties of milk contain different amounts of fat.

Activity _______________ 2.5

- Take some full-cream milk in a test tube.
- Centrifuge it by using a centrifuging machine for two minutes. If a centrifuging machine is not available in the school, you can do this activity at home by using a milk churner, used in the kitchen.
- If you have a milk dairy nearby, visit it and ask (i) how they separate cream from milk and (ii) how they make cheese (*paneer*) from milk.

Now answer

- *What do you observe on churning the milk?*
- *Explain how the separation of cream from milk takes place.*

Sometimes the solid particles in a liquid are very small and pass through a filter paper. For such particles the filtration technique cannot be used for separation. Such mixtures are separated by centrifugation. The principle is that the denser particles are forced to the bottom and the lighter particles stay at the top when spun rapidly.

Applications

- Used in diagnostic laboratories for blood and urine tests.
- Used in dairies and home to separate butter from cream.
- Used in washing machines to squeeze out water from wet clothes.

2.3.3 How can we separate a mixture of two immiscible liquids?

Activity _______________ *2.6*

- Let us try to separate kerosene oil from water using a separating funnel.
- Pour the mixture of kerosene oil and water in a separating funnel (Fig. 2.6).
- Let it stand undisturbed for sometime so that separate layers of oil and water are formed.
- Open the stopcock of the separating funnel and pour out the lower layer of water carefully.
- Close the stopcock of the separating funnel as the oil reaches the stop-cock.

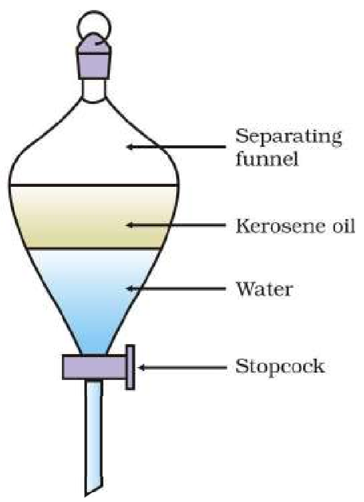


Fig. 2.6: *Separation of immiscible liquids*

Applications

- To separate mixture of oil and water.
- In the extraction of iron from its ore, the lighter slag is removed from the top by this method to leave the molten iron at the bottom in the furnace.

The principle is that immiscible liquids separate out in layers depending on their densities.

2.3.4 How can we separate a mixture of salt and camphor?

We have learnt in chapter 1 that camphor changes directly from solid to gaseous state on heating. So, to separate such mixtures that contain a sublimable volatile component from a non-sublimable impurity, the sublimation process is used (Fig. 2.7). Some examples of solids which sublime are ammonium chloride, naphthalene and anthracene.

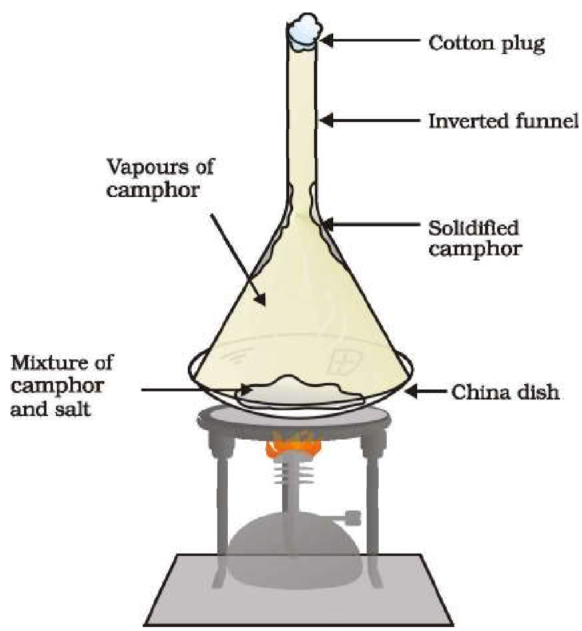


Fig. 2.7: *Separation of camphor and salt by sublimation*

2.3.5 Is the dye in black ink a single colour?

Activity 2.7

- Take a thin strip of filter paper.
- Draw a line on it using a pencil, approximately 3 cm above the lower edge [Fig. 2.8 (a)].
- Put a small drop of ink (water soluble, that is, from a sketch pen or fountain pen) at the centre of the line. Let it dry.
- Lower the filter paper into a jar/glass/beaker/test tube containing water so that the drop of ink on the paper is just above the water level, as shown in Fig. 2.8(b) and leave it undisturbed.
- Watch carefully, as the water rises up on the filter paper. Record your observations.

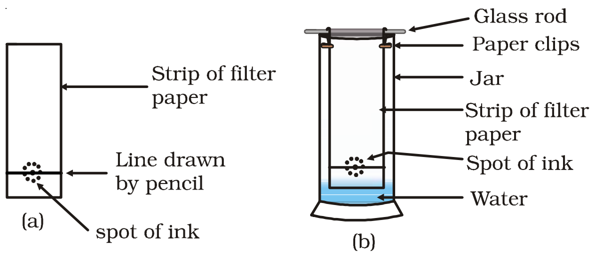


Fig. 2.8: *Separation of dyes in black ink using chromatography*

Now answer

- *What do you observe on the filter paper as the water rises on it?*
- *Do you obtain different colours on the filter paper strip?*
- *What according to you, can be the reason for the rise of the coloured spot on the paper strip?*

The ink that we use has water as the solvent and the dye is soluble in it. As the water rises on the filter paper it takes along with it the dye particles. Usually, a dye is a mixture of two or more colours. The coloured component that is more soluble in water, rises faster and in this way the colours get separated.

This process of separation of components of a mixture is known as chromatography. *Kroma* in Greek means colour. This technique was first used for separation of colours, so this name was given. Chromatography is the technique used for separation of those solutes that dissolve in the same solvent.

With the advancement in technology, newer techniques of chromatography have been developed. You will study about chromatography in higher classes.

Applications

To separate
- colours in a dye
- pigments from natural colours
- drugs from blood.

2.3.6 How can we separate a mixture of two miscible liquids?

Activity 2.8

- Let us try to separate acetone and water from their mixture.
- Take the mixture in a distillation flask. Fit it with a thermometer.
- Arrange the apparatus as shown in Fig. 2.9.
- Heat the mixture slowly keeping a close watch at the thermometer.
- The acetone vaporises, condenses in the condenser and can be collected from the condenser outlet.
- Water is left behind in the distillation flask.

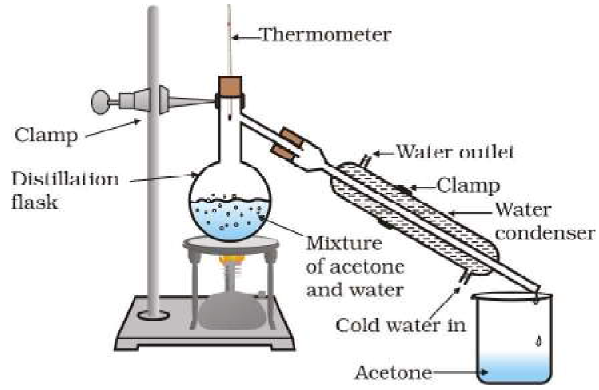


Fig.2.9: *Separation of two miscible liquids by distillation*

Now answer

- *What do you observe as you start heating the mixture?*
- *At what temperature does the thermometer reading become constant for some time?*
- *What is the boiling point of acetone?*
- *Why do the two components separate?*

This method is called distillation. It is used for the separation of components of a mixture containing two miscible liquids that boil without decomposition and have sufficient difference in their boiling points.

To separate a mixture of two or more miscible liquids for which the difference in boiling points is less than 25 K, fractional distillation process is used, for example, for the separation of different gases from air, different factions from petroleum products etc. The apparatus is similar to that for simple distillation, except that a fractionating column is fitted in between the distillation flask and the condenser.

A simple fractionating column is a tube packed with glass beads. The beads provide surface for the vapours to cool and condense repeatedly, as shown in Fig. 2.10.

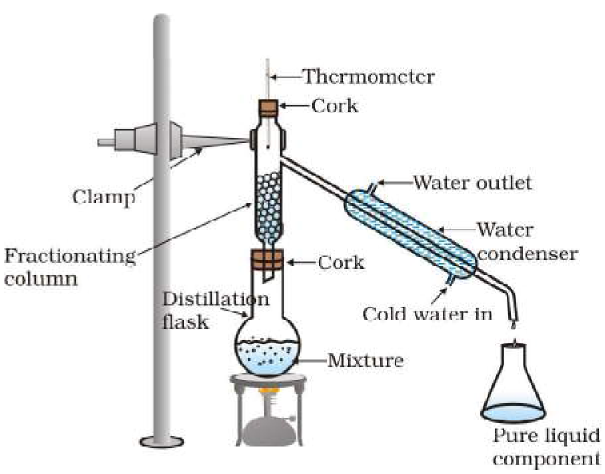
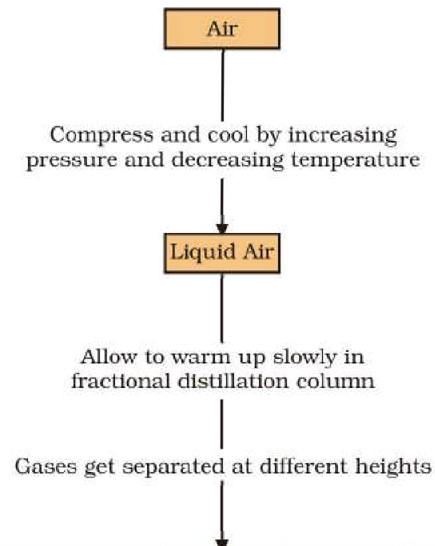


Fig. 2.10: *Fractional distillation*

2.3.7 How can we obtain different gases from air ?

Air is a homogeneous mixture and can be separated into its components by fractional distillation. The flow diagram (Fig. 2.11) shows the steps of the process.



	Oxygen	Argon	Nitrogen
Boiling Point (°C)	–183	–186	–196
% Air by Volume	20.9	0.9	78.1

Fig. 2.11: *Flow diagram shows the process of obtaining gases from air*

If we want oxygen gas from air (Fig. 2.12), we have to separate out all the other gases present in the air. The air is compressed by increasing the pressure and is then cooled by decreasing the temperature to get liquid air. This liquid air is allowed to warm-up slowly in a fractional distillation column, where gases get separated at different heights depending upon their boiling points.

Answer the following:

- *Arrange the gases present in air in increasing order of their boiling points.*
- *Which gas forms the liquid first as the air is cooled?*

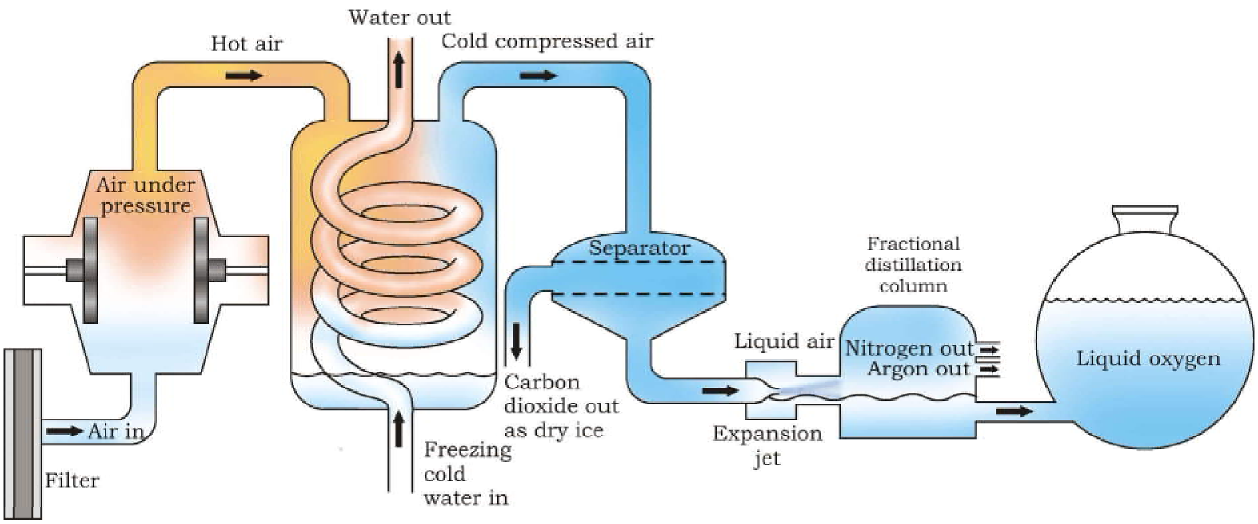


Fig. 2.12: *Separation of components of air*

2.3.8 How can we obtain pure copper sulphate from an impure sample?

Activity _______________ 2.9

- Take some (approximately 5 g) impure sample of copper sulphate in a china dish.
- Dissolve it in minimum amount of water.
- Filter the impurities out.
- Evaporate water from the copper sulphate solution so as to get a saturated solution.
- Cover the solution with a filter paper and leave it undisturbed at room temperature to cool slowly for a day.
- You will obtain the crystals of copper sulphate in the china dish.
- This process is called crystallisation.

Now answer

- *What do you observe in the china dish?*
- *Do the crystals look alike?*
- *How will you separate the crystals from the liquid in the china dish?*

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 is a process that separates a pure solid in the form of its crystals from a solution. 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 Fig. 2.13. 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.

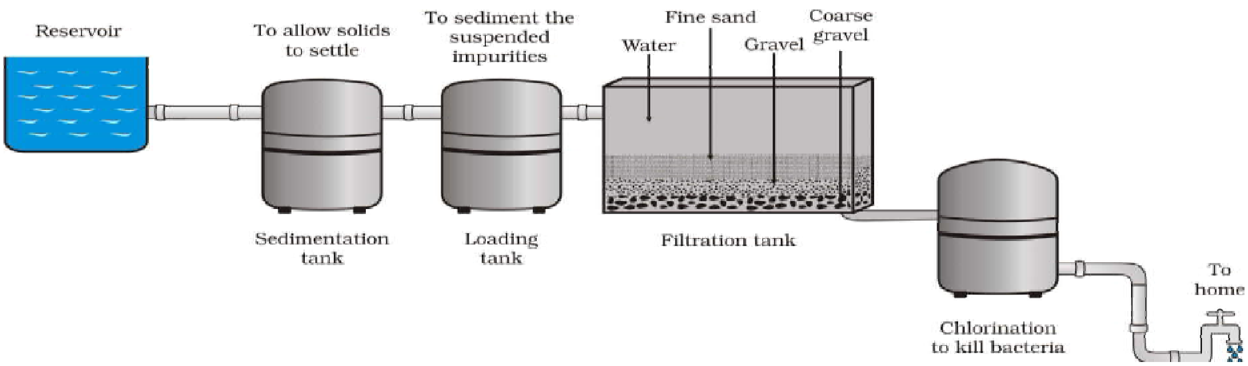


Fig. 2.13: *Water purification system in water works*

Questions

1. *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?*
2. *Name the technique to separate*
   (i) butter from curd,
   (ii) salt from sea-water,
   (iii) camphor from salt.
3. *What type of mixtures are separated by the technique of crystallisation?*

2.4 Physical and Chemical Changes

In the previous chapter, we have learnt about a few physical properties of matter. The properties that can be observed and specified like colour, hardness, rigidity, fluidity, density, melting point, boiling point etc. are the physical properites.

The interconversion of states is a physical change because these changes occur without a change in composition and no change in the chemical nature of the substance. Although ice, water and water vapour all look different and display different physical properties, they are chemically the same.

Both water and cooking oil are liquid but their chemical characteristics are different. They differ in odour and inflammability. We know that oil burns in air whereas water extinguishes fire. It is this chemical property of oil that makes it different from water. Burning is a chemical change. During this process one substance reacts with another to undergo a change in chemical composition. Chemical change brings change in the chemical properties of matter and we get new substances. A chemical change is also called a chemical reaction.

During burning of a candle, both physical and chemical changes take place. Can you distinguish these?

Questions

1. *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.*
2. *Try segregating the things around you as pure substances or mixtures.*

2.5 What are the Types of Pure Substances?

On the basis of their chemical composition, substances can be classified either as elements or compounds.

2.5.1 Elements

Robert Boyle was the first scientist to use the term element in 1661. Antoine Laurent Lavoisier (1743-94), a French chemist, was the first to establish an experimentally useful definition of an element. He defined an element as a basic form of matter that cannot be broken down into simpler substances by chemical reactions.

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.

More to know

- The number of elements known at present are more than 100. Ninety-two elements are naturally occurring and the rest are man-made.
- Majority of the elements are solid.
- Eleven elements are in gaseous state at room temperature.
- Two elements are liquid at room temperature–mercury and bromine.
- Elements, gallium and cesium become liquid at a temperature slightly above room temperature (303 K).

2.5.2 Compounds

A compound is a substance composed of two or more elements, chemically combined with one another in a fixed proportion.

What do we get when two or more elements are combined?

Activity 2.10

- Divide the class into two groups. Give 5 g of iron filings and 3 g of sulphur powder in a china dish to both the groups.

Group I

- Mix and crush iron filings and sulphur powder.

Group II

- Mix and crush iron filings and sulphur powder. Heat this mixture strongly till red hot. Remove from flame and let the mixture cool.

Groups I and II

- Check for magnetism in the material obtained. Bring a magnet near the material and check if the material is attracted towards the magnet.
- Compare the texture and colour of the material obtained by the groups.
- Add carbon disulphide to one part of the material obtained. Stir well and filter.
- Add dilute sulphuric acid or dilute hydrochloric acid to the other part of

Table 2.2: Mixtures and Compounds

Mixtures	Compounds
1. Elements or compounds just mix together to form a mixture and no new compound is formed.	1. Elements react to form new compounds.
2. A mixture has a variable composition.	2. The composition of each new substance is always fixed.
3. A mixture shows the properties of the constituent substances.	3. The new substance has totally different properties.
4. The constituents can be seperated fairly easily by physical methods.	4. The constituents can be separated only by chemical or electrochemical reactions.

the material obtained.*(Note: teacher supervision is necessary for this activity).*

- Perform all the above steps with both the elements (iron and sulphur) separately.

Now answer

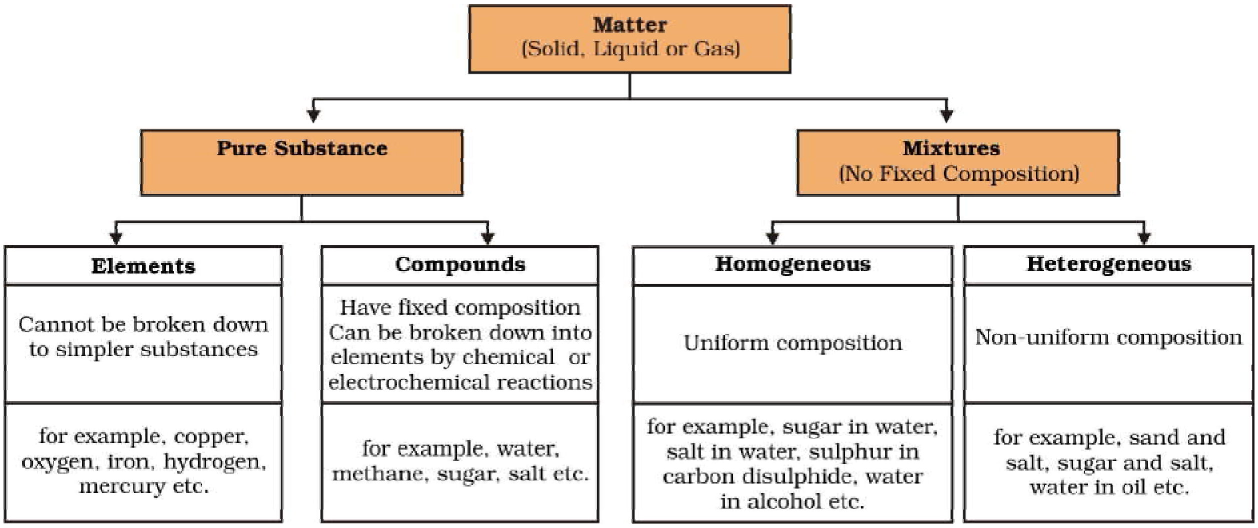
- *Did the material obtained by the two groups look the same?*
- *Which group has obtained a material with magnetic properties?*
- *Can we separate the components of the material obtained?*
- *On adding dilute sulphuric acid or dilute hydrochloric acid, did both the groups obtain a gas? Did the gas in both the cases smell the same or different?*

The gas obtained by Group I is hydrogen, it is colourless, odourless and combustible– it is not advised to do the combustion test for hydrogen in the class. The gas obtained by Group II is hydrogen sulphide. It is a colourless gas with the smell of rotten eggs.

You must have observed that the products obtained by both the groups show different properties, though the starting materials were the same. Group I has carried out the activity involving a physical change whereas in case of Group II, a chemical change (a chemical reaction) has taken place.

- The material obtained by group I is a mixture of the two substances. The substances given are the elements– iron and sulphur.
- The properties of the mixture are the same as that of its constituents.
- The material obtained by group II is a compound.
- On heating the two elements strongly we get a compound, which has totally different properties compared to the combining elements.
- The composition of a compound is the same throughout. We can also observe that the texture and the colour of the compound are the same throughout.

Thus, we can summarise the physical and chemical nature of matter in the following graphical organiser :



What you have learnt

- A mixture contains more than one substance (element and/or compound) mixed in any proportion.
- Mixtures can be separated into pure substances using appropriate separation techniques.
- A solution is a homogeneous mixture of two or more substances. The major component of a solution is called the solvent, and the minor, the solute.
- The concentration of a solution is the amount of solute present per unit volume or per unit mass of the solution.
- Materials that are insoluble in a solvent and have particles that are visible to naked eyes, form a suspension. A suspension is a heterogeneous mixture.
- Colloids are heterogeneous mixtures in which the particle size is too small to be seen with the naked eye, but is big enough to scatter light. Colloids are useful in industry and daily life. The particles are called the dispersed phase and the medium in which they are distributed is called the dispersion medium.
- Pure substances can be elements or compounds. An element is a form of matter that cannot be broken down by chemical reactions into simpler substances. A compound is a substance composed of two or more different types of elements, chemically combined in a fixed proportion.
- Properties of a compound are different from its constituent elements, whereas a mixture shows the properties of its constituting elements or compounds.

Exercises

1. 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.
2. Write the steps you would use for making tea. Use the words solution, solvent, solute, dissolve, soluble, insoluble, filtrate and residue.
3. 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**
	Solubility				
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?

4. Explain the following giving examples.
   (a) saturated solution
   (b) pure substance
   (c) colloid
   (d) suspension
5. Classify each of the following as a homogeneous or heterogeneous mixture.
   soda water, wood, air, soil, vinegar, filtered tea.
6. How would you confirm that a colourless liquid given to you is pure water?
7. 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.
8. Identify the solutions among the following mixtures.
   (a) Soil
   (b) Sea water
   (c) Air
   (d) Coal
   (e) Soda water.
9. Which of the following will show “Tyndall effect”?
   (a) Salt solution
   (b) Milk
   (c) Copper sulphate solution
   (d) Starch solution.
10. 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

11. 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.

Group Activity

Take an earthen pot (*mutka*), some pebbles and sand. Design a small-scale filtration plant that you could use to clean muddy water.

Chapter 3

Atoms and Molecules

Ancient Indian and Greek philosophers have always wondered about the unknown and unseen form of matter. The idea of divisibility of matter was considered long back in India, around 500 BC. An Indian philosopher Maharishi Kanad, postulated that if we go on dividing matter (*padarth*), we shall get smaller and smaller particles. Ultimately, a stage will come when we shall come across the smallest particles beyond which further division will not be possible. He named these particles *Parmanu*. Another Indian philosopher, Pakudha Katyayama, elaborated this doctrine and said that these particles normally exist in a combined form which gives us various forms of matter.

Around the same era, ancient Greek philosophers – Democritus and Leucippus suggested that if we go on dividing matter, a stage will come when particles obtained cannot be divided further. Democritus called these indivisible particles atoms (meaning indivisible). All this was based on philosophical considerations and not much experimental work to validate these ideas could be done till the eighteenth century.

By the end of the eighteenth century, scientists recognised the difference between elements and compounds and naturally became interested in finding out how and why elements combine and what happens when they combine.

Antoine L. Lavoisier laid the foundation of chemical sciences by establishing two important laws of chemical combination.

3.1 Laws of Chemical Combination

The following two laws of chemical combination were established after much experimentations by Lavoisier and Joseph L. Proust.

3.1.1 Law of conservation of mass

Is there a change in mass when a chemical change (chemical reaction) takes place?

Activity 3.1

- Take one of the following sets, X and Y of chemicals–

	X	Y
(i)	copper sulphate 1.25 g	sodium carbonate 1.43 g
(ii)	barium chloride 1.22 g	sodium sulphate 1.53 g
(iii)	lead nitrate 2.07 g	sodium chloride 1.17 g

- Prepare separately a 5% solution of any one pair of substances listed under X and Y each in 10 mL in water.
- Take a little amount of solution of Y in a conical flask and some solution of X in an ignition tube.
- Hang the ignition tube in the flask carefully; see that the solutions do not get mixed. Put a cork on the flask (see Fig. 3.1).

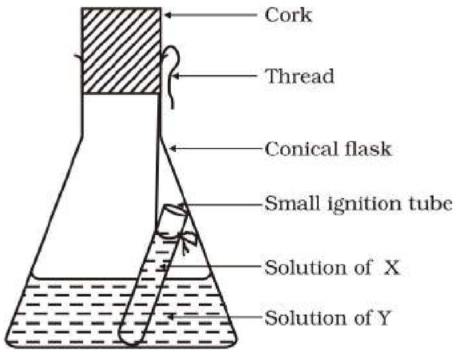


Fig. 3.1: *Ignition tube containing solution of X, dipped in a conical flask containing solution of Y.*

- Weigh the flask with its contents carefully.
- Now tilt and swirl the flask, so that the solutions X and Y get mixed.
- Weigh again.
- What happens in the reaction flask?
- Do you think that a chemical reaction has taken place?
- Why should we put a cork on the mouth of the flask?
- Does the mass of the flask and its contents change?

Law of conservation of mass states that mass can neither be created nor destroyed in a chemical reaction.

3.1.2 Law of constant proportions

Lavoisier, along with other scientists, noted that many compounds were composed of two or more elements and each such compound had the same elements in the same proportions, irrespective of where the compound came from or who prepared it.

In a compound such as water, the ratio of the mass of hydrogen to the mass of oxygen is always 1:8, whatever the source of water. Thus, if 9 g of water is decomposed, 1 g of hydrogen and 8 g of oxygen are always obtained. Similarly in ammonia, nitrogen and hydrogen are always present in the ratio 14:3 by mass, whatever the method or the source from which it is obtained.

This led to the law of constant proportions which is also known as the law of definite proportions. This law was stated by Proust as "*In a chemical substance the elements are always present in definite proportions by mass*".

The next problem faced by scientists was to give appropriate explanations of these laws. British chemist John Dalton provided the basic theory about the nature of matter. Dalton picked up the idea of divisibility of matter, which was till then just a philosophy. He took the name 'atoms' as given by the Greeks and said that the smallest particles of matter are atoms. His theory was based on the laws of chemical combination. Dalton's atomic theory provided an explanation for the law of conservation of mass and the law of definite proportions.

John Dalton

John Dalton was born in a poor weaver's family in 1766 in England. He began his career as a teacher at the age of twelve. Seven years later he became a school principal. In 1793, Dalton left for Manchester to teach mathematics, physics and chemistry in a college. He spent most of his life there teaching and researching. In 1808, he presented his atomic theory which was a turning point in the study of matter.

According to Dalton's atomic theory, all matter, whether an element, a compound or a mixture is composed of small particles called atoms. The postulates of this theory may be stated as follows:

(i) All matter is made of very tiny particles called atoms, which participate in chemical reactions.

(ii) Atoms are indivisible particles, which cannot be created or destroyed in a chemical reaction.

(iii) Atoms of a given element are identical in mass and chemical properties.

(iv) Atoms of different elements have different masses and chemical properties.

(v) Atoms combine in the ratio of small whole numbers to form compounds.

(vi) The relative number and kinds of atoms are constant in a given compound.

You will study in the next chapter that all atoms are made up of still smaller particles.

Questions

1. *In a reaction, 5.3 g of sodium carbonate reacted with 6 g of acetic acid. The products were 2.2 g of carbon dioxide, 0.9 g water and 8.2 g of sodium acetate. Show that these*

observations are in agreement with the law of conservation of mass.
sodium carbonate + acetic acid → sodium acetate + carbon dioxide + water

2. *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?*
3. *Which postulate of Dalton's atomic theory is the result of the law of conservation of mass?*
4. *Which postulate of Dalton's atomic theory can explain the law of definite proportions?*

3.2 What is an Atom?

Have you ever observed a mason building walls, from these walls a room and then a collection of rooms to form a building? What is the building block of the huge building? What about the building block of an ant-hill? It is a small grain of sand. Similarly, the building blocks of all matter are atoms.

How big are atoms?

Atoms are very small, they are smaller than anything that we can imagine or compare with. More than millions of atoms when stacked would make a layer barely as thick as this sheet of paper.

Atomic radius is measured in nanometres.

$$1/10^{9}\ \mathrm{m} = 1\ \mathrm{nm}$$
$$1\ \mathrm{m} = 10^{9}\ \mathrm{nm}$$

Relative Sizes	
Radii (in m)	**Example**
10^{-10}	Atom of hydrogen
10^{-9}	Molecule of water
10^{-8}	Molecule of haemoglobin
10^{-4}	Grain of sand
10^{-3}	Ant
10^{-1}	Apple

We might think that if atoms are so insignificant in size, why should we care about them? This is because our entire world is made up of atoms. We may not be able to see them, but they are there, and constantly affecting whatever we do. Through modern techniques, we can now produce magnified images of surfaces of elements showing atoms.

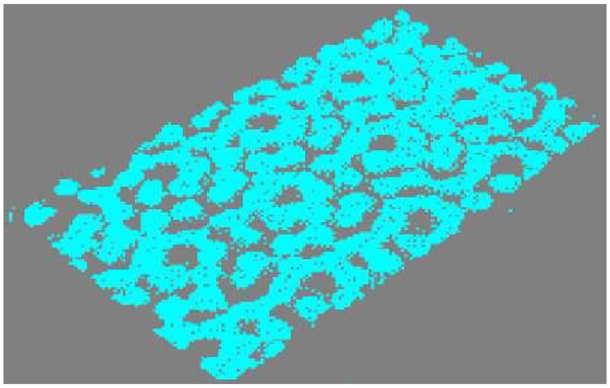


Fig. 3.2: *An image of the surface of silicon*

3.2.1 What are the modern day symbols of atoms of different 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.

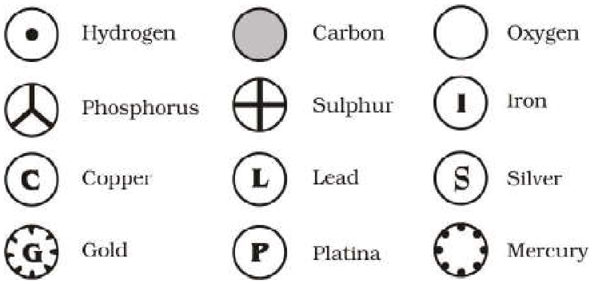


Fig. 3.3: *Symbols for some elements as proposed by Dalton*

In the beginning, the names of elements were derived from the name of the place where they were found for the first time. For example, the name copper was taken from Cyprus. Some names were taken from specific colours. For example, gold was taken from the English word meaning yellow. Now-a-days, IUPAC (International Union of Pure and Applied Chemistry) is an international scientific organisation which approves names of elements, symbols and units. 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 example

(i) hydrogen, H
(ii) aluminium, Al and not AL
(iii) cobalt, Co and not CO.

Symbols of some elements are formed from the first letter of the name and a letter, appearing later in the name. Examples are: (i) chlorine, Cl, (ii) zinc, Zn etc.

Other symbols have been taken from the names of elements in Latin, German or Greek. For example, the symbol of iron is Fe from its Latin name ferrum, sodium is Na from natrium, potassium is K from kalium. Therefore, each element has a name and a unique chemical symbol.

Table 3.1: Symbols for some elements

Element	Symbol	Element	Symbol	Element	Symbol
Aluminium	Al	Copper	Cu	Nitrogen	N
Argon	Ar	Fluorine	F	Oxygen	O
Barium	Ba	Gold	Au	Potassium	K
Boron	B	Hydrogen	H	Silicon	Si
Bromine	Br	Iodine	I	Silver	Ag
Calcium	Ca	Iron	Fe	Sodium	Na
Carbon	C	Lead	Pb	Sulphur	S
Chlorine	Cl	Magnesium	Mg	Uranium	U
Cobalt	Co	Neon	Ne	Zinc	Zn

(The above table is given for you to refer to whenever you study about elements. Do not bother to memorise all in one go. With the passage of time and repeated usage you will automatically be able to reproduce the symbols).

3.2.2 Atomic mass

The most remarkable concept that Dalton's atomic theory proposed was that of the atomic mass. According to him, each element had a characteristic atomic mass. The theory could explain the law of constant proportions so well that scientists were prompted to measure the atomic mass of an atom. Since determining the mass of an individual atom was a relatively difficult task, relative atomic masses were determined using the laws of chemical combinations and the compounds formed.

Let us take the example of a compound, carbon monoxide (CO) formed by carbon and oxygen. It was observed experimentally that 3 g of carbon combines with 4 g of oxygen to form CO. In other words, carbon combines with 4/3 times its mass of oxygen. Suppose we define the atomic mass unit (earlier abbreviated as 'amu', but according to the latest IUPAC recommendations, it is now written as 'u' – unified mass) as equal to the mass of one carbon atom, then we would assign carbon an atomic mass of 1.0 u and oxygen an atomic mass of 1.33 u. However, it is more convenient to have these numbers as whole numbers or as near to a whole numbers as possible. While searching for various atomic mass units, scientists initially took 1/16 of the mass of an atom of naturally occurring oxygen as the unit. This was considered relevant due to two reasons:

- oxygen reacted with a large number of elements and formed compounds.
- this atomic mass unit gave masses of most of the elements as whole numbers.

However, in 1961 for a universally accepted atomic mass unit, carbon-12 isotope was chosen as the standard reference for measuring atomic masses. One atomic mass unit is a mass unit equal to exactly one-twelfth ($1/12^{th}$) the mass of one atom of carbon-12. The relative atomic masses of all elements have been found with respect to an atom of carbon-12.

Imagine a fruit seller selling fruits without any standard weight with him. He takes a watermelon and says, “this has a mass equal to 12 units” (12 watermelon units or 12 fruit mass units). He makes twelve equal pieces of the watermelon and finds the mass of each fruit he is selling, relative to the mass of one piece of the watermelon. Now he sells his fruits by relative fruit mass unit (fmu), as in Fig. 3.4.

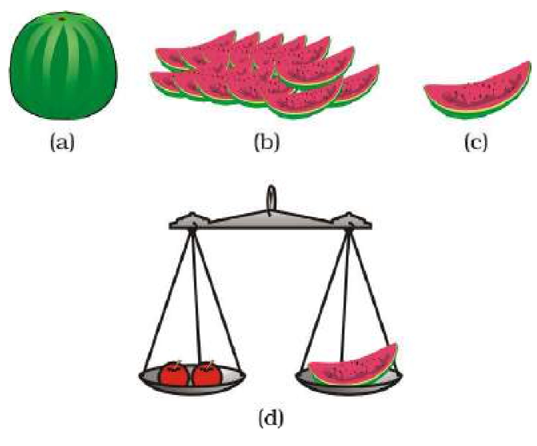


Fig. 3.4 : *(a) Watermelon, (b) 12 pieces, (c) 1/12 of watermelon, (d) how the fruit seller can weigh the fruits using pieces of watermelon*

Similarly, the relative atomic mass of the atom of an element is defined as the average mass of the atom, as compared to $1/12^{th}$ the mass of one carbon-12 atom.

Table 3.2: Atomic masses of a few elements

Element	Atomic Mass (u)
Hydrogen	1
Carbon	12
Nitrogen	14
Oxygen	16
Sodium	23
Magnesium	24
Sulphur	32
Chlorine	35.5
Calcium	40

3.2.3 How do atoms exist?

Atoms of most elements are not able to exist independently. Atoms form molecules and ions. These molecules or ions aggregate in large numbers to form the matter that we can see, feel or touch.

Questions

1. *Define the atomic mass unit.*
2. *Why is it not possible to see an atom with naked eyes?*

3.3 What is a Molecule?

A molecule is in general a group of two or more atoms that are chemically bonded together, that is, tightly held together by attractive forces. A molecule can be defined as the smallest particle of an element or a compound that is capable of an independent existence and shows all the properties of that substance. Atoms of the same element or of different elements can join together to form molecules.

3.3.1 Molecules of elements

The molecules of an element are constituted by the same type of atoms. Molecules of many elements, such as argon (Ar), helium (He) etc. are made up of only one atom of that element. But this is not the case with most of the non-metals. For example, a molecule of oxygen consists of two atoms of oxygen and hence it is known as a diatomic molecule, O_2. If 3 atoms of oxygen unite into a molecule, instead of the usual 2, we get ozone, O_3. The number of atoms constituting a molecule is known as its atomicity.

Metals and some other elements, such as carbon, do not have a simple structure but consist of a very large and indefinite number of atoms bonded together.

Let us look at the atomicity of some non-metals.

Table 3.3 : Atomicity of some elements

Type of Element	Name	Atomicity
Non-Metal	Argon	Monoatomic
	Helium	Monoatomic
	Oxygen	Diatomic
	Hydrogen	Diatomic
	Nitrogen	Diatomic
	Chlorine	Diatomic
	Phosphorus	Tetra-atomic
	Sulphur	Poly-atomic

3.3.2 Molecules of compounds

Atoms of different elements join together in definite proportions to form molecules of compounds. Few examples are given in Table 3.4.

Table 3.4 : Molecules of some compounds

Compound	Combining Elements	Ratio by Mass
Water	Hydrogen, Oxygen	1:8
Ammonia	Nitrogen, Hydrogen	14:3
Carbon dioxide	Carbon, Oxygen	3:8

Activity 3.2

- Refer to Table 3.4 for ratio by mass of atoms present in molecules and Table 3.2 for atomic masses of elements. Find the ratio by number of the atoms of elements in the molecules of compounds given in Table 3.4.
- The ratio by number of atoms for a water molecule can be found as follows:

Element	Ratio by mass	Atomic mass (u)	Mass ratio/ atomic mass	Simplest ratio
H	*1*	*1*	$\frac{1}{1} = 1$	*2*
O	*8*	*16*	$\frac{8}{16} = \frac{1}{2}$	*1*

- Thus, the ratio by number of atoms for water is H:O = 2:1.

3.3.3 What is an ion?

Compounds composed of metals and non-metals contain charged species. The charged species are known as *ions*. Ions may consist of a single charged atom or a group of atoms that have a net charge on them. An ion can be negatively or positively charged. A negatively charged ion is called an 'anion' and the positively charged ion, a 'cation'. Take, for example, sodium chloride (NaCl). Its constituent particles are positively charged sodium ions (Na^+) and negatively charged chloride ions (Cl^-). A group of atoms carrying a charge is known as a polyatomic ion (Table 3.6). We shall learn more about the formation of ions in Chapter 4.

Table 3.5: Some ionic compounds

Ionic Compound	Constituting Elements	Ratio by Mass
Calcium oxide	Calcium and oxygen	5:2
Magnesium sulphide	Magnesium and sulphur	3:4
Sodium chloride	Sodium and chlorine	23:35.5

3.4 Writing Chemical Formulae

The chemical formula of a compound is a symbolic representation of its composition. The chemical formulae of different compounds can be written easily. For this exercise, we need to learn the symbols and combining capacity of the elements.

The combining power (or capacity) of an element is known as its valency. Valency can be used to find out how the atoms of an element will combine with the atom(s) of another element to form a chemical compound. The valency of the atom of an element can be thought of as hands or arms of that atom.

Human beings have two arms and an octopus has eight. If one octopus has to catch hold of a few people in such a manner that all the eight arms of the octopus and both arms of all the humans are locked, how many humans do you think the octopus can hold? Represent the octopus with O and humans with H. Can you write a formula for this combination? Do you get OH_4 as the formula? The subscript 4 indicates the number of humans held by the octopus.

The valencies of some common ions are given in Table 3.6. We will learn more about valency in the next chapter.

Table 3.6: Names and symbols of some ions

Valency	Name of ion	Symbol	Non-metallic element	Symbol	Polyatomic ions	Symbol
1.	Sodium	Na^+	Hydrogen	H^+	Ammonium	NH_4^+
	Potassium	K^+	Hydride	H^-	Hydroxide	OH^-
	Silver	Ag^+	Chloride	Cl^-	Nitrate	NO_3^-
	Copper (I)*	Cu^+	Bromide	Br^-	Hydrogen carbonate	HCO_3^-
			Iodide	I^-		
2.	Magnesium	Mg^{2+}	Oxide	O^{2-}	Carbonate	CO_3^{2-}
	Calcium	Ca^{2+}	Sulphide	S^{2-}	Sulphite	SO_3^{2-}
	Zinc	Zn^{2+}			Sulphate	SO_4^{2-}
	Iron (II)*	Fe^{2+}				
	Copper (II)*	Cu^{2+}				
3.	Aluminium	Al^{3+}	Nitride	N^{3-}	Phosphate	PO_4^{3-}
	Iron (III)*	Fe^{3+}				

* *Some elements show more than one valency. A Roman numeral shows their valency in a bracket.*

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 number of ions present in the compound is indicated by enclosing the formula of ion in a bracket and writing the number of ions outside the bracket. For example, $Mg(OH)_2$. In case the number of polyatomic ion is one, the bracket is not required. For example, NaOH.

3.4.1 Formulae of simple compounds

The simplest compounds, which are made up of two different elements are called binary compounds. Valencies of some ions are given in Table 3.6. You can use these to write formulae for 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.

Examples

1. Formula of hydrogen chloride

Symbol H Cl

Valency 1 1

Formula of the compound would be HCl.

2. Formula of hydrogen sulphide

Symbol H S

Valency 1 2

Formula : H_2S

3. Formula of carbon tetrachloride

Symbol C Cl

Valency 4 1

Formula : CCl_4

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.

4. Formula of magnesium chloride

Symbol Mg Cl

Charge 2+ 1-

Formula : $MgCl_2$

Thus, in magnesium chloride, there are two chloride ions (Cl^-) for each magnesium ion (Mg^{2+}). The positive and negative charges must balance each other and the overall structure must be neutral. Note that in the formula, the charges on the ions are not indicated.

Some more examples

(a) Formula for aluminium oxide:

Symbol Al O

Charge 3+ 2-

Formula : Al_2O_3

(b) Formula for calcium oxide:

Symbol Ca O

Charge 2+ 2-

Here, the valencies of the two elements are the same. You may arrive at the formula Ca_2O_2. But we simplify the formula as CaO.

(c) Formula of sodium nitrate:

Symbol/ Formula	Na	NO_3
Charge	1+	1-

Formula : $NaNO_3$

(d) Formula of calcium hydroxide:

Symbol/ Formula	Ca	OH
Charge	2+	1-

Formula : $Ca(OH)_2$

Note that the formula of calcium hydroxide is $Ca(OH)_2$ and not $CaOH_2$. We use brackets when we have two or more of the same ions in the formula. Here, the bracket around OH with a subscript 2 indicates that there are two hydroxyl (OH) groups joined to one calcium atom. In other words, there are two atoms each of oxygen and hydrogen in calcium hydroxide.

(e) Formula of sodium carbonate:

Symbol/ Formula	Na	CO_3
Charge	1+	2-

Formula : Na_2CO_3

In the above example, brackets are not needed if there is only one ion present.

(f) Formula of ammonium sulphate:

Formula	NH_4	SO_4
Charge	1+	2-

Formula : $(NH_4)_2SO_4$

Questions

1. *Write down the formulae of*
   (i) *sodium oxide*
   (ii) *aluminium chloride*
   (iii) *sodium suphide*
   (iv) *magnesium hydroxide*
2. *Write down the names of compounds represented by the following formulae:*
   (i) $Al_2(SO_4)_3$
   (ii) $CaCl_2$
   (iii) K_2SO_4
   (iv) KNO_3
   (v) $CaCO_3$.
3. *What is meant by the term chemical formula?*
4. *How many atoms are present in a*
   (i) H_2S *molecule and*
   (ii) PO_4^{3-} *ion?*

3.5 Molecular Mass and Mole Concept

3.5.1 Molecular mass

In section 3.2.2 we discussed the concept of atomic mass. This concept can be extended to calculate molecular masses. The molecular mass of a substance is the sum of the atomic masses of all the atoms in a molecule of the substance. It is therefore the relative mass of a molecule expressed in *atomic mass units (u).*

Example 3.1 (a) Calculate the relative molecular mass of water (H_2O). (b) Calculate the molecular mass of HNO_3.

Solution:

(a) Atomic mass of hydrogen = 1u,
oxygen = 16 u
So the molecular mass of water, which contains two atoms of hydrogen and one atom of oxygen is $= 2 \times 1 + 1 \times 16$
$= 18$ u

(b) The molecular mass of HNO_3 = the atomic mass of H + the atomic mass of N + 3 × the atomic mass of O
= 1 + 14 + 48 = 63 u

3.5.2 Formula unit mass

The formula unit mass of a substance is a sum of the atomic masses of all atoms in a formula unit of a compound. Formula unit mass is calculated in the same manner as we calculate the molecular mass. The only difference is that we use the word formula unit for those substances whose constituent particles are ions. For example, sodium chloride as discussed above, has a formula unit NaCl. Its formula unit mass can be calculated as–

$$1 \times 23 + 1 \times 35.5 = 58.5 \text{ u}$$

Example 3.2 Calculate the formula unit mass of $CaCl_2$.

Solution:

Atomic mass of Ca
+ (2 × atomic mass of Cl)
= 40 + 2 × 35.5 = 40 + 71 = 111 u

Questions

1. *Calculate the molecular masses of H_2, O_2, Cl_2, CO_2, CH_4, C_2H_6, C_2H_4, NH_3, CH_3OH.*
2. *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.*

3.5.3 Mole concept

Take an example of the reaction of hydrogen and oxygen to form water:

$$2H_2 + O_2 \rightarrow 2H_2O.$$

The above reaction indicates that

(i) two molecules of hydrogen combine with one molecule of oxygen to form two molecules of water, or

(ii) 4 u of hydrogen molecules combine with 32 u of oxygen molecules to form 36 u of water molecules.

We can infer from the above equation that the quantity of a substance can be characterised by its mass or the number of molecules. But, a chemical reaction equation indicates directly the number of atoms or molecules taking part in the reaction. Therefore, it is more convenient to refer to the quantity of a substance in terms of the number of its molecules or atoms, rather than their masses. So, a new unit "mole" was introduced. One mole of any species (atoms,

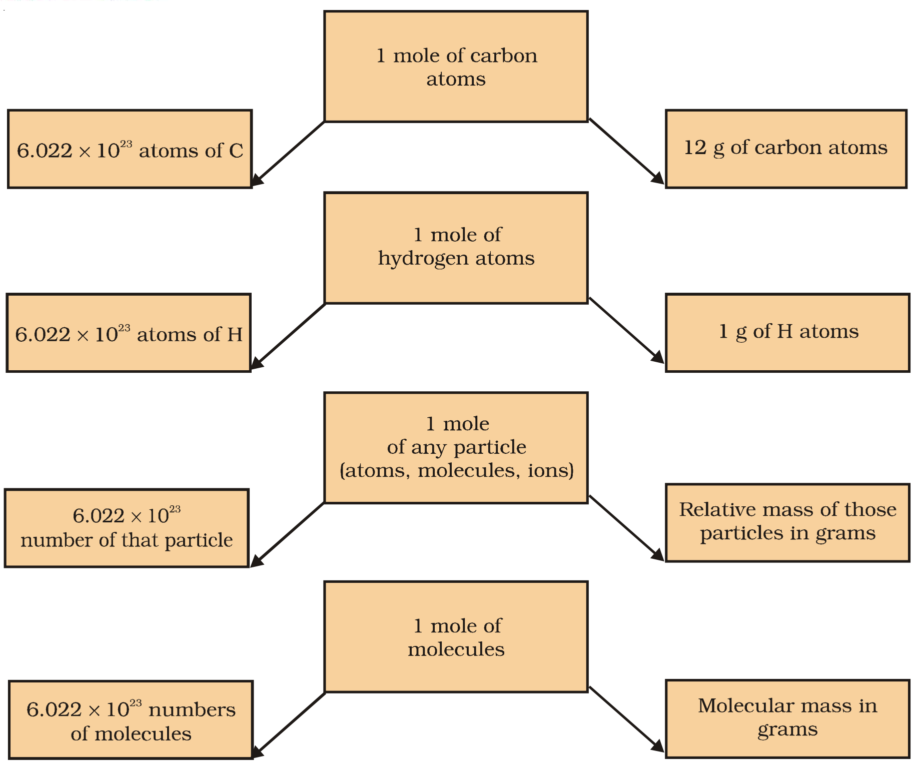


Fig. 3.5: *Relationship between mole, Avogadro number and mass*

molecules, ions or particles) is that quantity in number having a mass equal to its atomic or molecular mass in grams.

The number of particles (atoms, molecules or ions) present in 1 mole of any substance is fixed, with a value of 6.022×10^{23}. This is an experimentally obtained value. This number is called the Avogadro Constant or Avogadro Number *(represented by* N_0), named in honour of the Italian scientist, Amedeo Avogadro.

1 mole (of anything) = 6.022×10^{23} in number,
as, 1 dozen = 12 nos.
1 gross = 144 nos.

Besides being related to a number, a mole has one more advantage over a dozen or a gross. This advantage is that mass of 1 mole of a particular substance is also fixed.

The mass of 1 mole of a substance is equal to its relative atomic or molecular mass in grams. The atomic mass of an element gives us the mass of one atom of that element in atomic mass units (u). To get the mass of 1 mole of atom of that element, that is, molar mass, we have to take the same numerical value but change the units from 'u' to 'g'. Molar mass of atoms is also known as gram atomic mass. For example, atomic mass of hydrogen=1u. So, gram atomic mass of hydrogen = 1 g.

1 u hydrogen has only 1 atom of hydrogen
1 g hydrogen has 1 mole atoms, that is, 6.022×10^{23} atoms of hydrogen.
Similarly,
16 u oxygen has only 1 atom of oxygen,
16 g oxygen has 1 mole atoms, that is, 6.022×10^{23} atoms of oxygen.

To find the gram molecular mass or molar mass of a molecule, we keep the numerical value the same as the molecular mass, but simply change units as above from u to g. For example, as we have already calculated, molecular mass of water (H_2O) is 18 u. From here we understand that
18 u water has only 1 molecule of water,
18 g water has 1 mole molecules of water, that is, 6.022×10^{23} molecules of water.

Chemists need the number of atoms and molecules while carrying out reactions, and for this they need to relate the mass in grams to the number. It is done as follows:

1 mole = 6.022×10^{23} number
= Relative mass in grams.

Thus, a mole is the chemist's counting unit.

The word "mole" was introduced around 1896 by Wilhelm Ostwald who derived the term from the Latin word *moles* meaning a 'heap' or 'pile'. A substance may be considered as a heap of atoms or molecules. The unit mole was accepted in 1967 to provide a simple way of reporting a large number– the massive heap of atoms and molecules in a sample.

Example 3.3

1. Calculate the number of moles for the following:
   (i) 52 g of He (finding mole from mass)
   (ii) 12.044×10^{23} number of He atoms (finding mole from number of particles).

Solutions:

No. of moles	=	n
Given mass	=	m
Molar mass	=	M
Given number of particles	=	N
Avogadro number of particles	=	N_0

(i) Atomic mass of He = 4 u
Molar mass of He = 4g

Thus, the number of moles

$$= \frac{\text{given mass}}{\text{molar mass}}$$

$$\Rightarrow n = \frac{m}{M} = \frac{52}{4} = 13$$

(ii) we know,
1 mole = 6.022×10^{23}
The number of moles

$$= \frac{\text{given number of particles}}{\text{Avogadro number}}$$

$$\Rightarrow n = \frac{N}{N_0} = \frac{12.044 \times 10^{23}}{6.022 \times 10^{23}} = 2$$

Example 3.4 Calculate the mass of the following:

(i) 0.5 mole of N_2 gas (mass from mole of molecule)

(ii) 0.5 mole of N atoms (mass from mole of atom)

(iii) 3.011×10^{23} number of N atoms (mass from number)

(iv) 6.022×10^{23} number of N_2 molecules (mass from number)

Solutions:

(i) mass = molar mass × number of moles

$$\Rightarrow m = M \times n = 28 \times 0.5 = 14\,g$$

(ii) mass = molar mass × number of moles

$$\Rightarrow m = M \times n = 14 \times 0.5 = 7\,g$$

(iii) The number of moles, n

$$= \frac{\text{given number of particles}}{\text{Avogadro number}} = \frac{N}{N_0}$$

$$= \frac{3.011 \times 10^{23}}{6.022 \times 10^{23}}$$

$$\Rightarrow m = M \times n = 14 \times \frac{3.011 \times 10^{23}}{6.022 \times 10^{23}}$$

$$= 14 \times 0.5 = 7\,g$$

(iv) $n = \dfrac{N}{N_0}$

$$\Rightarrow m = M \times \frac{N}{N_0} = 28 \times \frac{6.022 \times 10^{23}}{6.022 \times 10^{23}}$$

$$= 28 \times 1 = 28\ g$$

Example 3.5 Calculate the number of particles in each of the following:

(i) 46 g of Na atoms (number from mass)

(ii) 8 g O_2 molecules (number of molecules from mass)

(iii) 0.1 mole of carbon atoms (number from given moles)

Solutions:

(i) The number of atoms

$$= \frac{\text{given mass}}{\text{molar mass}} \times \text{Avogadro number}$$

$$\Rightarrow N = \frac{m}{M} \times N_0$$

$$\Rightarrow N = \frac{46}{23} \times 6.022 \times 10^{23}$$

$$\Rightarrow N = 12.044 \times 10^{23}$$

(ii) The number of molecules

$$= \frac{\text{given mass}}{\text{molar mass}} \times \text{Avogadro number}$$

$$\Rightarrow N = \frac{m}{M} \times N_0$$

atomic mass of oxygen = 16 u

$\therefore$ molar mass of O_2 molecules = 16 × 2 = 32g

$$\Rightarrow N = \frac{8}{32} \times 6.022 \times 10^{23}$$

$$\Rightarrow N = 1.5055 \times 10^{23}$$

$$\approx 1.51 \times 10^{23}$$

(iii) The number of particles (atom) = number of moles of particles × Avogadro number

$N = n \times N_0 = 0.1 \times 6.022 \times 10^{23}$

$= 6.022 \times 10^{22}$

Questions

1. *If one mole of carbon atoms weighs 12 grams, what is the mass (in grams) of 1 atom of carbon?*
2. *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)?*

What you have learnt

- During a chemical reaction, the sum of the masses of the reactants and products remains unchanged. This is known as the Law of Conservation of Mass.
- In a pure chemical compound, elements are always present in a definite proportion by mass. This is known as the Law of Definite Proportions.
- An atom is the smallest particle of the element that cannot usually exist independently and retain all its chemical properties.
- A molecule is the smallest particle of an element or a compound capable of independent existence under ordinary conditions. It shows all the properties of the substance.
- A chemical formula of a compound shows its constituent elements and the number of atoms of each combining element.
- Clusters of atoms that act as an ion are called polyatomic ions. They carry a fixed charge on them.
- The chemical formula of a molecular compound is determined by the valency of each element.
- In ionic compounds, the charge on each ion is used to determine the chemical formula of the compound.
- Scientists use the relative atomic mass scale to compare the masses of different atoms of elements. Atoms of carbon-12 isotopes are assigned a relative atomic mass of 12 and the relative masses of all other atoms are obtained by comparison with the mass of a carbon-12 atom.
- The Avogadro constant 6.022×10^{23} is defined as the number of atoms in exactly 12 g of carbon-12.
- The mole is the amount of substance that contains the same number of particles (atoms/ ions/ molecules/ formula units etc.) as there are atoms in exactly 12 g of carbon-12.
- Mass of 1 mole of a substance is called its molar mass.

Exercises

1. 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.
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?

3. What are polyatomic ions? Give examples.
4. Write the chemical formulae of the following.
   (a) Magnesium chloride
   (b) Calcium oxide
   (c) Copper nitrate
   (d) Aluminium chloride
   (e) Calcium carbonate.
5. Give the names of the elements present in the following compounds.
   (a) Quick lime
   (b) Hydrogen bromide
   (c) Baking powder
   (d) Potassium sulphate.
6. Calculate the molar mass of the following substances.
   (a) Ethyne, C_2H_2
   (b) Sulphur molecule, S_8
   (c) Phosphorus molecule, P_4 (Atomic mass of phosphorus = 31)
   (d) Hydrochloric acid, HCl
   (e) Nitric acid, HNO_3
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 (Na_2SO_3)?
8. Convert into mole.
   (a) 12 g of oxygen gas
   (b) 20 g of water
   (c) 22 g of carbon dioxide.
9. What is the mass of:
   (a) 0.2 mole of oxygen atoms?
   (b) 0.5 mole of water molecules?
10. Calculate the number of molecules of sulphur (S_8) present in 16 g of solid sulphur.
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)

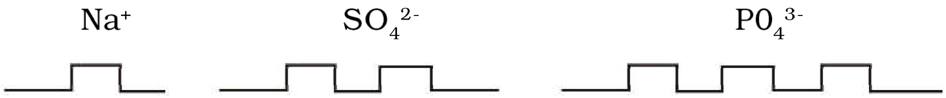
Group Activity

Play a game for writing formulae.

Example1 : Make placards with symbols and valencies of the elements separately. Each student should hold two placards, one with the symbol in the right hand and the other with the valency in the left hand. Keeping the symbols in place, students should criss-cross their valencies to form the formula of a compound.

Example 2 : A low cost model for writing formulae: Take empty blister packs of medicines. Cut them in groups, according to the valency of the element, as shown in the figure. Now, you can make formulae by fixing one type of ion into other.

For example:



Formula for sodium sulphate:

2 sodium ions can be fixed on one sulphate ion.

Hence, the formula will be: Na_2SO_4

Do it yourself :

Now, write the formula of sodium phosphate.

Chapter 4

Structure of the Atom

In Chapter 3, we have learnt that atoms and molecules are the fundamental building blocks of matter. The existence of different kinds of matter is due to different atoms constituting them. Now the questions arise: (i) What makes the atom of one element different from the atom of another element? and (ii) Are atoms really indivisible, as proposed by Dalton, or are there smaller constituents inside the atom? We shall find out the answers to these questions in this chapter. We will learn about sub-atomic particles and the various models that have been proposed to explain how these particles are arranged within the atom.

A major challenge before the scientists at the end of the 19th century was to reveal the structure of the atom as well as to explain its important properties. The elucidation of the structure of atoms is based on a series of experiments.

One of the first indications that atoms are not indivisible, comes from studying static electricity and the condition under which electricity is conducted by different substances.

4.1 Charged Particles in Matter

For understanding the nature of charged particles in matter, let us carry out the following activities:

Activity ______________ *4.1*

A. Comb dry hair. Does the comb then attract small pieces of paper?

B. Rub a glass rod with a silk cloth and bring the rod near an inflated balloon. Observe what happens.

From these activities, can we conclude that on rubbing two objects together, they become electrically charged? Where does this charge come from? This question can be answered by knowing that an atom is divisible and consists of charged particles.

Many scientists contributed in revealing the presence of charged particles in an atom.

It was known by 1900 that the atom was indivisible particle but contained at least one sub-atomic particle – the electron identified by J.J. Thomson. Even before the electron was identified, E. Goldstein in 1886 discovered the presence of new radiations in a gas discharge and called them canal rays. These rays were positively charged radiations which ultimately led to the discovery of another sub-atomic particle. This sub-atomic particle had a charge, equal in magnitude but opposite in sign to that of the electron. Its mass was approximately 2000 times as that of the electron. It was given the name of proton. In general, an electron is represented as 'e^-' and a proton as 'p^+'. The mass of a proton is taken as one unit and its charge as plus one. The mass of an electron is considered to be negligible and its charge is minus one.

It seemed highly that an atom was composed of protons and electrons, mutually balancing their charges. It also appeared that the protons were in the interior of the atom, for whereas electrons could easily be removed off but not protons. Now the big question was: what sort of structure did these particles of the atom form? We will find the answer to this question below.

Questions

1. *What are canal rays?*
2. *If an atom contains one electron and one proton, will it carry any charge or not?*

4.2 The Structure of an Atom

We have learnt Dalton's atomic theory in Chapter 3, which suggested that the atom was indivisible and indestructible. But the discovery of two fundamental particles (electrons and protons) inside the atom, led to the failure of this aspect of Dalton's atomic theory. It was then considered necessary to know how electrons and protons are arranged within an atom. For explaining this, many scientists proposed various atomic models. J.J. Thomson was the first one to propose a model for the structure of an atom.

4.2.1 Thomson's model of an atom

Thomson proposed the model of an atom to be similar to that of a Christmas pudding. The electrons, in a sphere of positive charge, were like currants (dry fruits) in a spherical Christmas pudding. We can also think of a watermelon, the positive charge in the atom is spread all over like the red edible part of the watermelon, while the electrons are studded in the positively charged sphere, like the seeds in the watermelon (Fig. 4.1).

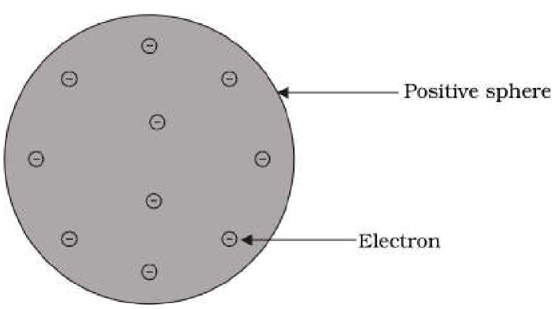
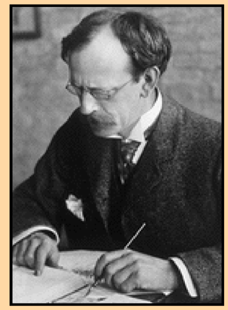


Fig.4.1: *Thomson's model of an atom*



J.J. Thomson (1856-1940), a British physicist, was born in Cheetham Hill, a suburb of Manchester, on 18 December 1856. He was awarded the Nobel prize in Physics in 1906 for his work on the discovery of electrons. He directed the Cavendish Laboratory at Cambridge for 35 years and seven of his research assistants subsequently won Nobel prizes.

Thomson proposed that:

(i) An atom consists of a positively charged sphere and the electrons are embedded in it.

(ii) The negative and positive charges are equal in magnitude. So, the atom as a whole is electrically neutral.

Although Thomson's model explained that atoms are electrically neutral, the results of experiments carried out by other scientists could not be explained by this model, as we will see below.

4.2.2 Rutherford's model of an atom

Ernest Rutherford was interested in knowing how the electrons are arranged within an atom. Rutherford designed an experiment for this. In this experiment, fast moving alpha (α)-particles were made to fall on a thin gold foil.

- He selected a gold foil because he wanted as thin a layer as possible. This gold foil was about 1000 atoms thick.
- α-particles are doubly-charged helium ions. Since they have a mass of 4 u, the fast-moving α-particles have a considerable amount of energy.
- It was expected that α-particles would be deflected by the sub-atomic particles in the gold atoms. Since the α-particles were much heavier than the protons, he did not expect to see large deflections.

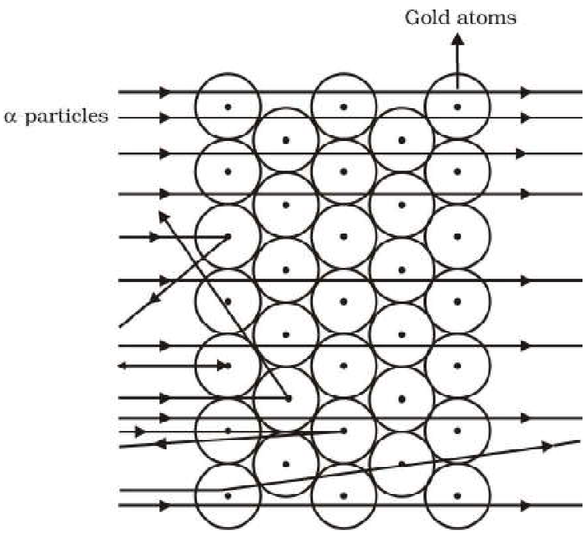


Fig. 4.2: *Scattering of α-particles by a gold foil*

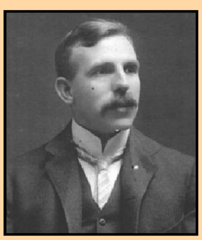
But, the α-particle scattering experiment gave totally unexpected results (Fig. 4.2). The following observations were made:

(i) Most of the fast moving α-particles passed straight through the gold foil.

(ii) Some of the α-particles were deflected by the foil by small angles.

(iii) Surprisingly one out of every 12000 particles appeared to rebound.

In the words of Rutherford, "*This result was almost as incredible as if you fire a 15-inch shell at a piece of tissue paper and it comes back and hits you*".



E. Rutherford (1871-1937) was born at Spring Grove on 30 August 1871. He was known as the 'Father' of nuclear physics. He is famous for his work on radioactivity and the discovery of the nucleus of an atom with the gold foil experiment. He got the Nobel prize in chemistry in 1908.

Let us think of an activity in an open field to understand the implications of this experiment. Let a child stand in front of a wall with his eyes closed. Let him throw stones at the wall from a distance. He will hear a sound when each stone strikes the wall. If he repeats this ten times, he will hear the sound ten times. But if a blind-folded child were to throw stones at a barbed-wire fence, most of the stones would not hit the fencing and no sound would be heard. This is because there are lots of gaps in the fence which allow the stone to pass through them.

Following a similar reasoning, Rutherford concluded from the α-particle scattering experiment that–

(i) Most of the space inside the atom is empty because most of the α-particles passed through the gold foil without getting deflected.

(ii) Very few particles were deflected from their path, indicating that the positive charge of the atom occupies very little space.

(iii) A very small fraction of α-particles were deflected by 180^0, indicating that all the positive charge and mass of the gold atom were concentrated in a very small volume within the atom.

From the data he also calculated that the radius of the nucleus is about 10^5 times less than the radius of the atom.

On the basis of his experiment, Rutherford put forward the nuclear model of an atom, which had the following features:

(i) There is a positively charged centre in an atom called the nucleus. Nearly all the mass of an atom resides in the nucleus.

(ii) The electrons revolve around the nucleus in circular paths.

(iii) The size of the nucleus is very small as compared to the size of the atom.

Drawbacks of Rutherford's model of the atom

The revolution of the electron in a circular orbit is not expected to be stable. Any particle in a circular orbit would undergo acceleration. During acceleration, charged particles would radiate energy. Thus, the revolving electron would lose energy and finally fall into the nucleus. If this were so, the atom should be highly unstable and hence matter would not exist in the form that we know. We know that atoms are quite stable.

4.2.3 Bohr's model of atom

In order to overcome the objections raised against Rutherford's model of the atom, Neils Bohr put forward the following postulates about the model of an atom:

(i) Only certain special orbits known as discrete orbits of electrons, are allowed inside the atom.

(ii) While revolving in discrete orbits the electrons do not radiate energy.

Neils Bohr (1885-1962) was born in Copenhagen on 7 October 1885. He was appointed professor of physics at Copenhagen University in 1916. He got the Nobel prize for his work on the structure of atom in 1922. Among Professor Bohr's numerous writings, three appearing as books are: (i) The Theory of Spectra and Atomic Constitution, (ii) Atomic Theory and, (iii) The Description of Nature.

These orbits or shells are called energy levels. Energy levels in an atom are shown in Fig. 4.3.

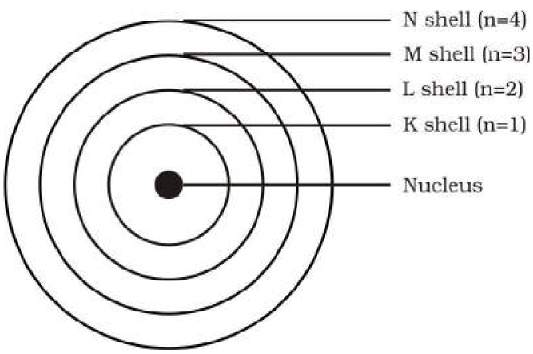


Fig. 4.3: *A few energy levels in an atom*

These orbits or shells are represented by the letters K,L,M,N,… or the numbers, n=1,2,3,4,….

Questions

1. *On the basis of Thomson's model of an atom, explain how the atom is neutral as a whole.*
2. *On the basis of Rutherford's model of an atom, which sub-atomic particle is present in the nucleus of an atom?*
3. *Draw a sketch of Bohr's model of an atom with three shells.*
4. *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?*

4.2.4 Neutrons

In 1932, J. Chadwick discovered another sub-atomic particle which had no charge and a mass nearly equal to that of a proton. It was eventually named as neutron. Neutrons are present in the nucleus of all atoms, except hydrogen. In general, a neutron is represented as 'n'. The mass of an atom is therefore given by the sum of the masses of protons and neutrons present in the nucleus.

Questions

1. *Name the three sub-atomic particles of an atom.*
2. *Helium atom has an atomic mass of 4 u and two protons in its nucleus. How many neutrons does it have?*

4.3 How are Electrons Distributed in Different Orbits (Shells)?

The distribution of electrons into different orbits of an atom was suggested by Bohr and Bury.

The following rules are followed for writing the number of electrons in different energy levels or shells:

(i) The maximum number of electrons present in a shell is given by the formula $2n^2$, where 'n' is the orbit number or energy level index, 1,2,3,.... Hence the maximum number of electrons in different shells are as follows:

first orbit or K-shell will be = $2 \times 1^2 = 2$, second orbit or L-shell will be = $2 \times 2^2 = 8$, third orbit or M-shell will be = $2 \times 3^2 = 18$, fourth orbit or N-shell will be = $2 \times 4^2 = 32$, and so on.

(ii) The maximum number of electrons that can be accommodated in the outermost orbit is 8.

(iii) Electrons are not accommodated in a given shell, unless the inner shells are filled. That is, the shells are filled in a step-wise manner.

Atomic structure of the first eighteen elements is shown schematically in Fig. 4.4.

- The composition of atoms of the first eighteen elements is given in Table 4.1.

Questions

1. *Write the distribution of electrons in carbon and sodium atoms.*
2. *If K and L shells of an atom are full, then what would be the total number of electrons in the atom?*

4.4 Valency

We have learnt how the electrons in an atom are arranged in different shells/orbits. The electrons present in the outermost shell of an atom are known as the valence electrons.

From the Bohr-Bury scheme, we also know that the outermost shell of an atom can accommodate a maximum of 8 electrons. It was observed that the atoms of elements, completely filled with 8 electrons in the outermost shell show little chemical activity. In other words, their combining capacity or valency is zero. Of these inert elements, the

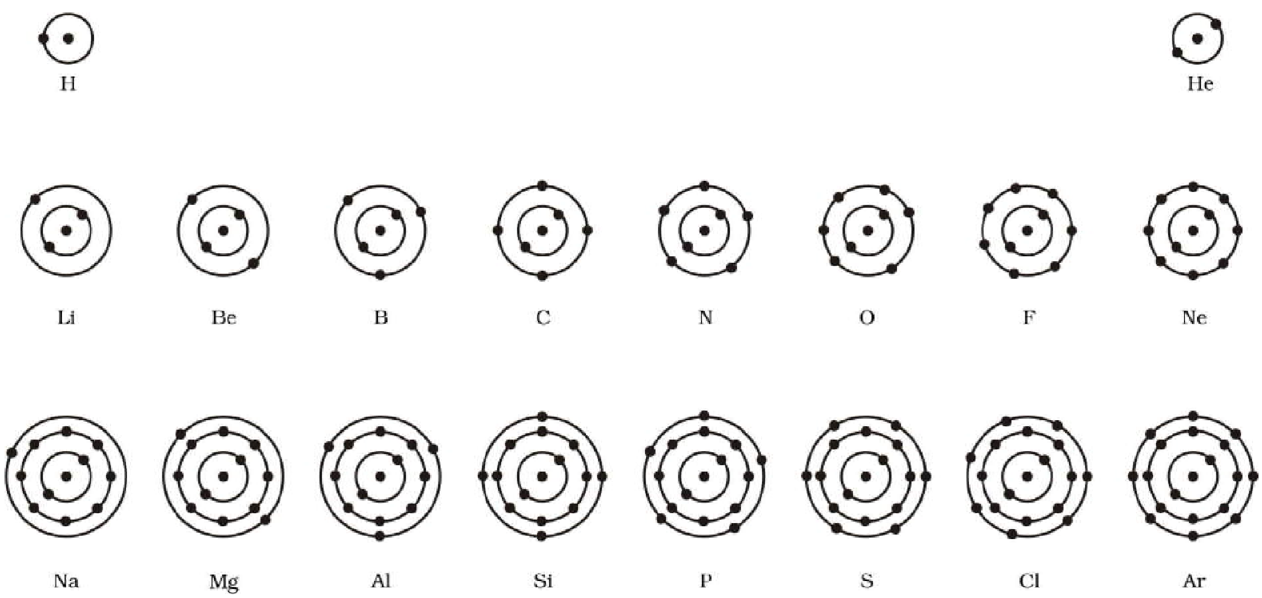


Fig.4.4: *Schematic atomic structure of the first eighteen elements*

Activity ______________ *4.2*

- Make a static atomic model displaying electronic configuration of the first eighteen elements.

Table 4.1: Composition of Atoms of the First Eighteen Elements with Electron Distribution in Various Shells

Name of Element	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Number of Electrons	Distribution of Electrons				Vale-ncy
						K	L	M	N	
Hydrogen	H	1	1	-	1	1	-	-	-	1
Helium	He	2	2	2	2	2	-	-	-	0
Lithium	Li	3	3	4	3	2	1	-	-	1
Beryllium	Be	4	4	5	4	2	2	-	-	2
Boron	B	5	5	6	5	2	3	-	-	3
Carbon	C	6	6	6	6	2	4	-	-	4
Nitrogen	N	7	7	7	7	2	5	-	-	3
Oxygen	O	8	8	8	8	2	6	-	-	2
Fluorine	F	9	9	10	9	2	7	-	-	1
Neon	Ne	10	10	10	10	2	8	-	-	0
Sodium	Na	11	11	12	11	2	8	1	-	1
Magnesium	Mg	12	12	12	12	2	8	2	-	2
Aluminium	Al	13	13	14	13	2	8	3	-	3
Silicon	Si	14	14	14	14	2	8	4	-	4
Phosphorus	P	15	15	16	15	2	8	5	-	3,5
Sulphur	S	16	16	16	16	2	8	6	-	2
Chlorine	Cl	17	17	18	17	2	8	7	-	1
Argon	Ar	18	18	22	18	2	8	8		0

helium atom has two electrons in its outermost shell and all other elements have atoms with eight electrons in the outermost shell.

The combining capacity of the atoms of elements, that is, their tendency to react and form molecules with atoms of the same or different elements, was thus explained as an attempt to attain a fully-filled outermost shell. An outermost-shell, which had eight electrons was said to possess an octet. Atoms would thus react, so as to achieve an octet in the outermost shell. This was done by sharing, gaining or losing electrons. The number of electrons gained, lost or shared so as to make the octet of electrons in the outermost shell, gives us directly the combining capacity of the element, that is, the valency discussed in the previous chapter. For example, hydrogen/ lithium/sodium atoms contain one electron each in their outermost shell, therefore each one of them can lose one electron. So, they are said to have valency of one. Can you tell, what is valency of magnesium and aluminium? It is two and three, respectively, because magnesium has two electrons in its outermost shell and aluminium has three electrons in its outermost shell.

If the number of electrons in the outermost shell of an atom is close to its full capacity, then valency is determined in a different way. For example, the fluorine atom has 7 electrons in the outermost shell, and its valency could be 7. But it is easier for fluorine to gain one electron instead of losing seven electrons. Hence, its valency is determined by subtracting seven electrons from the octet and this gives you a valency of one for fluorine. Valency can be calculated in a similar manner for oxygen. What is the valency of oxygen that you get from this calculation?

Therefore, an atom of each element has a definite combining capacity, called its valency. Valency of the first eighteen elements is given in the last column of Table 4.1.

Question

1. *How will you find the valency of chlorine, sulphur and magnesium?*

4.5 Atomic Number and Mass Number

4.5.1 Atomic number

We know that protons are present in the nucleus of an atom. It is the number of protons of an atom, which determines its atomic number. It is denoted by 'Z'. All atoms of an element have the same atomic number, Z. In fact, elements are defined by the number of protons they possess. For hydrogen, Z = 1, because in hydrogen atom, only one proton is present in the nucleus. Similarly, for carbon, Z = 6. Therefore, the atomic number is defined as the total number of protons present in the nucleus of an atom.

4.5.2 Mass number

After studying the properties of the sub-atomic particles of an atom, we can conclude that mass of an atom is practically due to protons and neutrons alone. These are present in the nucleus of an atom. Hence protons and neutrons are also called nucleons. Therefore, the mass of an atom resides in its nucleus. For example, mass of carbon is 12 u because it has 6 protons and 6 neutrons, 6 u + 6 u = 12 u. Similarly, the mass of aluminium is 27 u (13 protons+14 neutrons). The mass number is defined as the sum of the total number of protons and neutrons present in the nucleus of an atom. It is denoted by 'A'. In the notation for an atom, the atomic number, mass number and symbol of the element are to be written as:

Mass Number

$\boxed{\text{Symbol of element}}$ $\qquad$ $^{A}_{Z}\boxed{X}$

Atomic Number

For example, nitrogen is written as $^{14}_{7}N$.

Questions

1. *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?*
2. *With the help of Table 4.1, find out the mass number of oxygen and sulphur atom.*

4.6 Isotopes

In nature, a number of atoms of some elements have been identified, which have the same atomic number but different mass numbers. For example, take the case of hydrogen atom, it has three atomic species, namely protium ($^{1}_{1}H$), deuterium ($^{2}_{1}H$ or D) and tritium ($^{3}_{1}H$ or T). The atomic number of each one is 1, but the mass number is 1, 2 and 3, respectively. Other such examples are (i) carbon, $^{12}_{6}C$ and $^{14}_{6}C$, (ii) chlorine, $^{35}_{17}Cl$ and $^{37}_{17}Cl$, etc.

On the basis of these examples, isotopes are defined as the atoms of the same element, having the same atomic number but different mass numbers. Therefore, we can say that there are three isotopes of hydrogen atom, namely protium, deuterium and tritium.

Many elements consist of a mixture of isotopes. Each isotope of an element is a pure substance. The chemical properties of isotopes are similar but their physical properties are different.

Chlorine occurs in nature in two isotopic forms, with masses 35 u and 37 u in the ratio of 3:1. Obviously, the question arises: what should we take as the mass of chlorine atom? Let us find out.

The average atomic mass of chlorine atom, on the basis of above data, will be

$$\left[\left(35 \times \frac{75}{100} + 37 \times \frac{25}{100}\right)\right.$$

$$\left. = \left(\frac{105}{4} + \frac{37}{4}\right) = \frac{142}{4} = 35.5\,\mathrm{u}\right]$$

The mass of an atom of any natural element is taken as the average mass of all the naturally occuring atoms of that element. If an element has no isotopes, then the mass of its atom would be the same as the sum of protons and neutrons in it. But if an element occurs in isotopic forms, then we have to know the percentage of each isotopic form and then the average mass is calculated.

This does not mean that any one atom of chlorine has a fractional mass of 35.5 u. It means that if you take a certain amount of chlorine, it will contain both isotopes of chlorine and the average mass is 35.5 u.

Applications

Since the chemical properties of all the isotopes of an element are the same, normally we are not concerned about taking a mixture. But some isotopes have special properties which find them useful in various fields. Some of them are :

(i) An isotope of uranium is used as a fuel in nuclear reactors.

(ii) An isotope of cobalt is used in the treatment of cancer.

(iii) An isotope of iodine is used in the treatment of goitre.

4.6.1 Isobars

Let us consider two elements — calcium, atomic number 20, and argon, atomic number 18. The number of protons in these atoms is different, but the mass number of both these elements is 40. That is, the total number of nucleons is the same in the atoms of this pair of elements. Atoms of different elements with different atomic numbers, which have the same mass number, are known as isobars.

Questions

1. *For the symbol H,D and T tabulate three sub-atomic particles found in each of them.*
2. *Write the electronic configuration of any one pair of isotopes and isobars.*

What you have learnt

- Credit for the discovery of electron and proton goes to J.J. Thomson and E.Goldstein, respectively.
- J.J. Thomson proposed that electrons are embedded in a positive sphere.