2. Particulate Nature of Matter

2.1 Introduction

Matter is anything that has mass and volume. Solids such as sand, iron, and wood; liquids such as water, kerosene and ethanol; and gases, such as oxygen, nitrogen, carbon(IV) oxide and even air, are all matter.

If all these are matter, what then is not matter? Can light and sound be weighed? Can you measure their volumes? These things are not matter since they cannot be weighed and do not occupy space.

For many years scientists tried to find out what matter was made of. Their efforts led to the concept that matter is made up of particles. This concept is still held today. But what evidence do we have for this?

2.2 Evidences for the particulate nature of matter

There are several practical evidences that support the concept of particulate nature of matter. We shall examine some of these here.

Evidence from dissolving substances

Experiment 2.1 Dissolving Sodium Chloride in Water

Put water into a measuring cylinder up to the $20cm^3$ mark. Weigh out about 30g of sodium chloride, and with the aid of a spatula, add the salt to water a little at a time. Record the level of water in the measuring cylinder. Stir it well to dissolve completely (Figure 2.1).

You will initially notice that the level of water did not change after stirring. Continue the addition, stirring and recording the water levels each time until a rise in water level is observed. Weigh the remaining salt and thus determine how much was added before a 2 cm rise in water level was noticed.

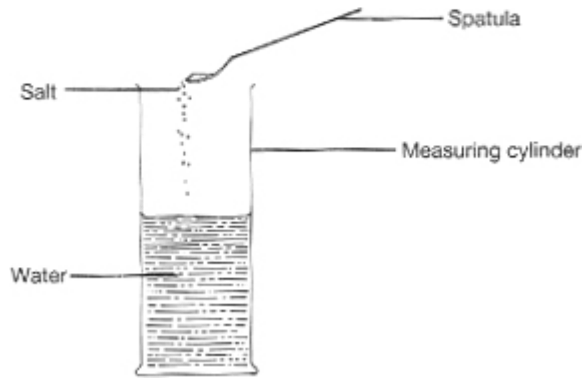


Figure 2.1 Dissolving sodium chloride in water inside a measuring cylinder.

Where has all the salt disappeared into?

Activity 2.1

Fill a beaker with glass beads (or any other such spherical object). Add sand to the beaker (Figure 1.2). Where does the sand enter? What does this suggest about the disappearance of salt into the water contained in a measuring cylinder without causing a rise in the water level (Experiment 2.1)?

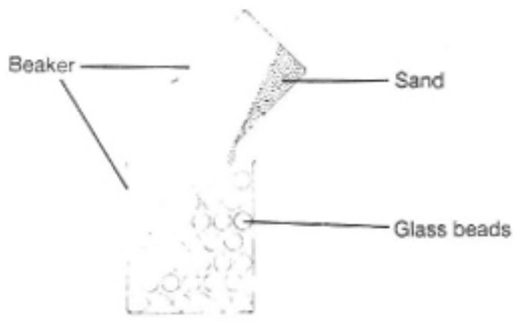


Figure 2.2 A beaker of glass beads and sand

Activity 2.1 illustrates what happens in Experiment 2.1. As the sand particles go into the space between the glass beads, so do particles of sodium chloride go into spaces between particles of water. That is, both water and sodium chloride are made up of tiny invisible particles.

Evidence from Diluting solutions

Experiment 2.2 Diluting a Solution of Potassium Tetraoxomanganate(VII)

You are provided with a solution of potassium tetraoxomanganate(VII) containing 2.0g of the crystals per dm^3. Measure out $10cm^3$ of the purple solution into a 100 cm^3 volumetric flask. Add distilled water to it till the level of the solution rises to the $100cm^3$ mark. Shake to mix well. Transfer $1cm^3$ of this solution into a test-tube labelled A.

Measure out $10cm^3$ of the diluted solution into another $100cm^3$

volumetric flask and dilute as before. Again transfer $1cm^3$ of the resultant solution into another test tube labelled B. Continue the progressive dilution until the colour of the solution is only just noticeable, putting $1cm^3$ of the solution from each step into test-tubes labelled C, D, E . . . respectively (Figure 2.3).

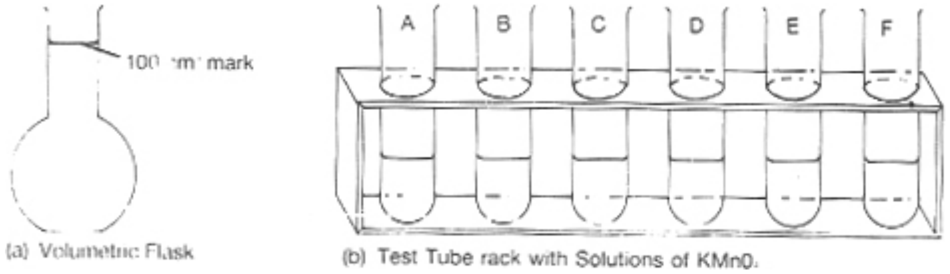


Figure 2.3

(a) Volumetric flask (b) Test-tube rack with solutions of KM_nO_1

View the solutions in test-tubes A, B, C . . . A gradual fading in colour is observed. The colour in the last test tube is hardly noticeable. Assuming that each step of dilution makes use of one-tenth of the original mass of potassium tetraoxomanganate(VII) in the preceeding solution, calculate the mass of this salt in the <u>$1cm^3$</u> of the last solution.

If dilution was done five times the mass in $100cm^3$ of the last diluent is (2/10)g.

∴The mass of the potassium tetraoxomaganate(VII) in $1cm^3$ of the solution is $(2/10^7)$g

= 2 Ã— 10^{-7}g.

How many times did you carry out your own dilution? Calculate the mass of the salt in $1cm^3$ of your last solution.

The result of the above calculation shows that some salt still remains in the last test tube even though the colour is no more noticeable. The colour of the solution is not noticeable because the salt particles are widely separated by the water particles. They are so far apart that the solution appears more like pure water.

Evidence from sublimation

Experiment 2.3 Sublimation of Iodine Crystals

Put about 5g of iodine crystals into an evaporating dish. Hold an inverted glass funnel about 5cm above the dish. Heat the dish on a wire gauze. Continue the heating until a reasonable quantity of violet solid is deposited on the funnel (Figure 2.4).

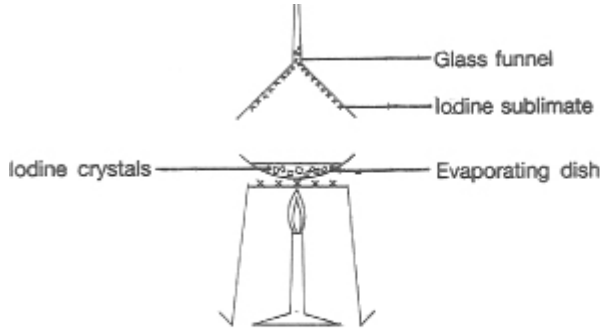


Figure 2.4 Sublimation of iodine

How have iodine crystals moved from the evaporating dish, through space, into the glass funnel?

In Experiment 2.3 iodine has sublimed. Very tiny particles of iodine in vapour form leave the evaporating dish which is hot, and settle on the colder glass funnel. The departing vapour can be seen on close observation while heating, as violet vapour in the space between the evaporating dish and the funnel.

An explanation of sublimation is that some of the particles of iodine crystals take up energy from the source of heat and move out as gas. On getting to the cold glass funnel they give up this energy and become very tiny solid particles. Many such particles recombine to form the visible solid iodine once more.

Brownian motion

Experiment 2.4 Demonstration of Brownian motion

Put a drop of Indian ink into 25cm^3 of water. Stir to mix well. Using a teat pipette, withdraw some of the resulting colloidal solution and put a drop of it onto a microscope slide. Place the cover-slip of the slide over it and keep on a slide mount. Pass a strong beam of light on to the slide from the side, as illustrated in Figure 2.5. Bring the colloidal solution on the slide to a focus using a powerful objective lens.

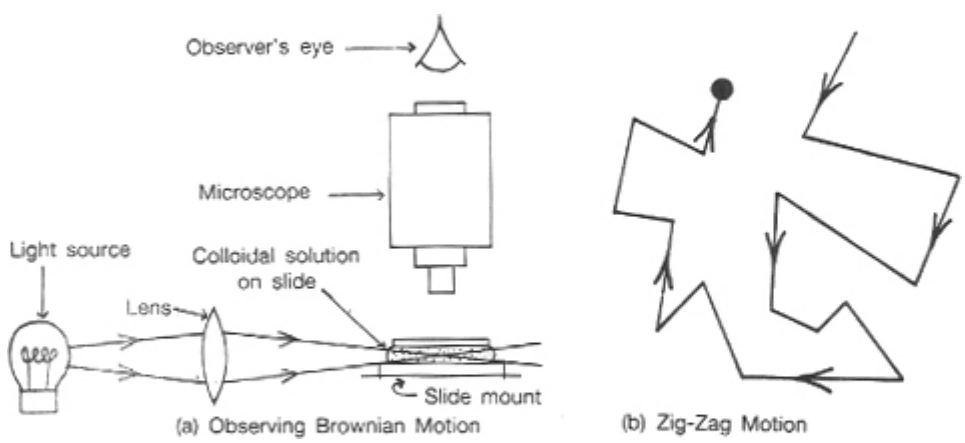


Figure 2.5

Particles of the Indian ink will be seen to be in constant random motion The motion is termed Brownian motion in honour of Robert Brown, who first noticed it among pollen grains suspended in water.

To explain Brownian motion we again use the particulate nature of matter. When mobile particles of the water strike a particle of the ink dispersed in it, the ink particle moves in the direction of the water particle. Because several water particles bombard the ink particles per second, the ink particles move in a zig-zag manner.

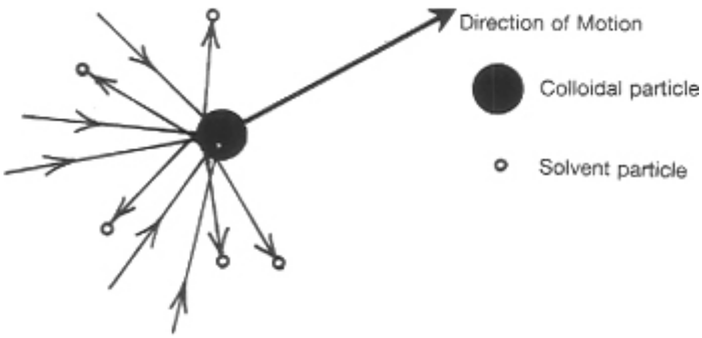


Figure 2.6 Explanation of Brownian Motion

Only particles of colloidal size can move in this fashion. Larger particles cannot be so moved by the force due to bombardments by small solvent particles.

2.3 Particle size

The particles that make up matter are too tiny to be seen even with a microscope. Their dimensions cannot therefore be measured directly. There must however be a way of estimating their sizes in order to compare them.

Experiment 2.5 Estimating the Size of a Particle

Fill a clean large glass trough with water. Sprinkle powdered animal charcoal thinly on top of the water, then wait for the water level to be steady. While waiting, fill a teat pipette with olive oil (turpentine oil, groundnut oil or any other vegetable oil may be used). Gently drop the oil into a small, clean, dry measuring cylinder, counting the drops. Determine how many drops make up $1cm^3$ of oil, hence calculate the volume of one drop of oil.

Place a drop of the oil into $100cm^3$ of petroleum ether and shake to dissolve. Wash the teat pipette, then rinse it with the oil solution in petroleum ether. With the teat pipette, put a drop of the oil solution at the centre of the surface of the water spinkled with powdered animal charcoal. The petroleum ether quickly evaporates leaving a circle of oil film on the surface of the water (Figure 2.7). The charcoal helps to mark the boundaries of the circle.

Determine the average diameter of the circle by measuring across different directions. Assuming that the oil spreads out so thinly that no particle of it is on top of the other, calculate the thickness of an oil particle.

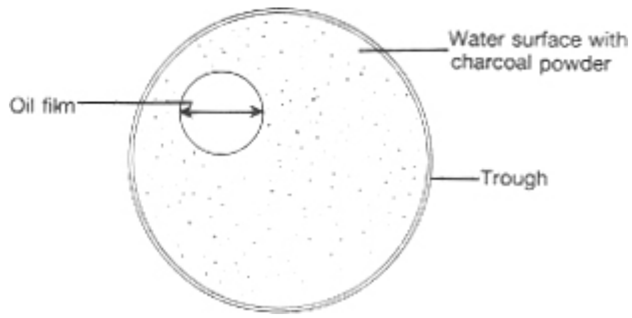


Figure 2.7 Oil film on the surface of water

Specimen Result And Calculations

50 drops of oil $= 1cm^3$

$\therefore$ Volume of 1 drop of oil $= \frac{1}{50}$ cm^3 $=$ 0.02 cm^3

Diameter of oil film $=$ 10 cm.

Since 0.02 cm^3 of oil is dissolved in 100 cm^3 of petroleum ether, and if one drop of the oil solution also has a volume of 0.02 cm^3 (since the same teat pipette was used) then the volume of oil in the one drop of solution

$$= \frac{0.02}{100} \times 0.02cm^3$$
$$= 0.000004 \ cm^3 = 4 \times 10^{-6} \ cm^3$$

Area covered by the oil film is given by Ï€.r^2, where r is the radius of the circle. If r = 5 cm

$\therefore$ Area $= \pi \times 5^2 \ cm^2 = \pi \times 25 \ cm^2$

$\therefore$ Thickness of the particle $= \frac{4 \times 10^{-6} \ cm^3}{\pi \times 25 \ cm^2}$, i.e. $\frac{\text{volume}}{\text{area}}$

$= 5.1 \times 10^{-8}$ cm.

This is a very small value, and may be smaller if the assumption that the particles spread out such that none is on top of the other, is not true. If the particles are spherical then thickness thus calculated is the diameter of the sphere.

Particles of all matter are however not of the same size. They are generally of the order of 10^{-7} cm in diameter. This unit, 10^{-7} cm, is known as one â€˜nanometerâ€™ (nm).

Exercise 2A

What is the figure 5.1 Ã— 10^{-8} cm in â€˜nanometersâ€™? You should have 5.1 Ã— 10^{-1} nm.

2.4 Daltonâ€™s atomic theory

The concept of particulate nature of matter was already on a firm footing when John Dalton put forward his famous atomic theory in 1808. The theory goes beyond merely stating that mater is made up of particles. The basic ideas of Daltonâ€™s atomic theory are expressed in the following statements:-

1. All matter are composed of tiny particles called atoms.
2. Atoms cannot be sub-divided.
3. Atoms of the same element are alike in all respects, especially in mass, but differ from atoms of other elements.
4. Atoms cannot be created or destroyed.
5. Atoms combine in small whole number ratios to form â€˜compound-atomsâ€™. Some of these theories are supported by the laws of chemical combination which are discussed later.

2.5 Particles of matter

The particles of matter can exist as atoms, molecules or ions.

In Integrated Science you learnt about atoms, elements, molecules and compounds.

An atom is the smallest particle of an element which can take part in a chemical reaction. The atom is the basic unit of an element

An element is defined as a substance which cannot be split into two or more different substances.

A molecule is the smallest electrically neutral particle of a substance which can exist on its own.

When atoms of the same element combine chemically, a molecule of the element is formed; but when atoms of different elements combine chemically, the molecule of a compound is formed.

The atomicity of an element is the number of atoms in a molecule of the element. In nature, elements exist either as single atoms or molecules. Elements that exist as single atoms are described as **monoatomic**, e.g. all metals and the noble gases. The atomicities of some common elements are indicated in Table 1.1.

TABLE 2.1 Atomicities of some common elements

Atomicity	Elements
1	He, Ar, Ne, all metals.
2	H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2.
3	O_3 (Ozone)
4	P_4 (Phosphorus)
8	S_8 (Sulphur)

2.6 Atomic structure

Daltonâ€™s atomic theory has undergone drastic modifications with the advent of modern instruments and technologies. Before discussing these modifications, we shall first consider the modern atomic structure. Details of experimental evidences for the modern atomic structure will be dealt with in a latter chapter.

Towards the end of the nineteenth century, J.J. Thompson produced experimental evidence for the existence of sub-particles known as **electrons** in the atom of every element. He found that the electron has a charge of â€“1 and a mass of $\frac{1}{1837}$ of the mass of a hydrogen atom.

At the beginning of the twentieth century Lord Rutherford demonstrated the nuclear structure of the atom. He showed that the atom consisted of a small positively charged **nucleus** surrounded by a negatively charged electron cloud (Figure 2.8).

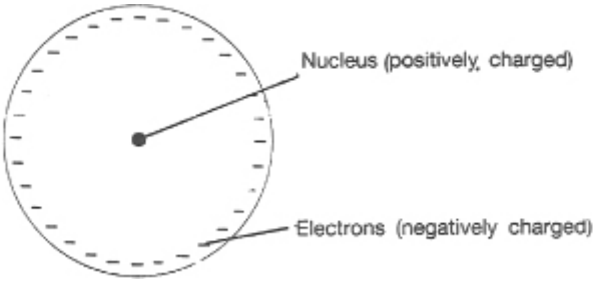


Figure 2.8 Rutherfordâ€™s structure of an atom

Further experiments showed that the nucleus contained two particles of approximately equal masses. These are the **proton** which has a charge of +1 and the **neutron** which has no charge. The proton is responsible for the positive charge of the nucleus. The proton and neutron have a mass of 1 **atomic mass unit** (a.m.u.) each. The mass of the atom is concentrated in the nucleus, as the electron has a negligible mass. The mass of the atom is therefore approximately equal to the sum of the masses of the protons and neutrons in the nucleus. The charges and masses of the three fundamental particles are summarised in Table 2.2.

TABLE 2.2 The Fundamental Particles of the Atom

Particle	Position in the Atom	Mass (in a.m.u.)	Charges (relative to electron)
Electron	Shells	$\frac{1}{1837}$ (negligible)	−1
Proton	Nucleus	1	+1
Neutron	Nucleus	1	zero

Electrons are located within **shells** around the nucleus. Starting with that nearest to the nucleus, the shells are named as **K, L, M, N.** These labels correspond to the **principal quantum numbers 1, 2, 3, 4** respectively as illustrated in Figure 2.9.

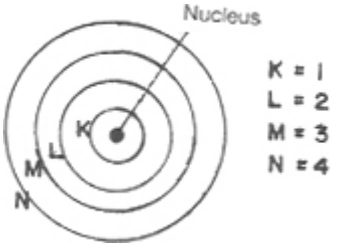


Figure 2.9 Modern atomic structure

The maximum number of electrons each shell can contain is specified by the formula $2n^2$, where n is the shell number or principal quantum number. Table 2.3 gives the maximum number of electrons each shell can contain.

TABLE 2.3 Maximum number of electrons in atomic shells

Shell number	Shell symbol	Maximum number of electrons
1.	K	2
2.	L	8
3.	M	18
4.	N	32

Representing the shells with circles, we arrive at the following structures for the first and simplest eleven elements (Figure 2.10).

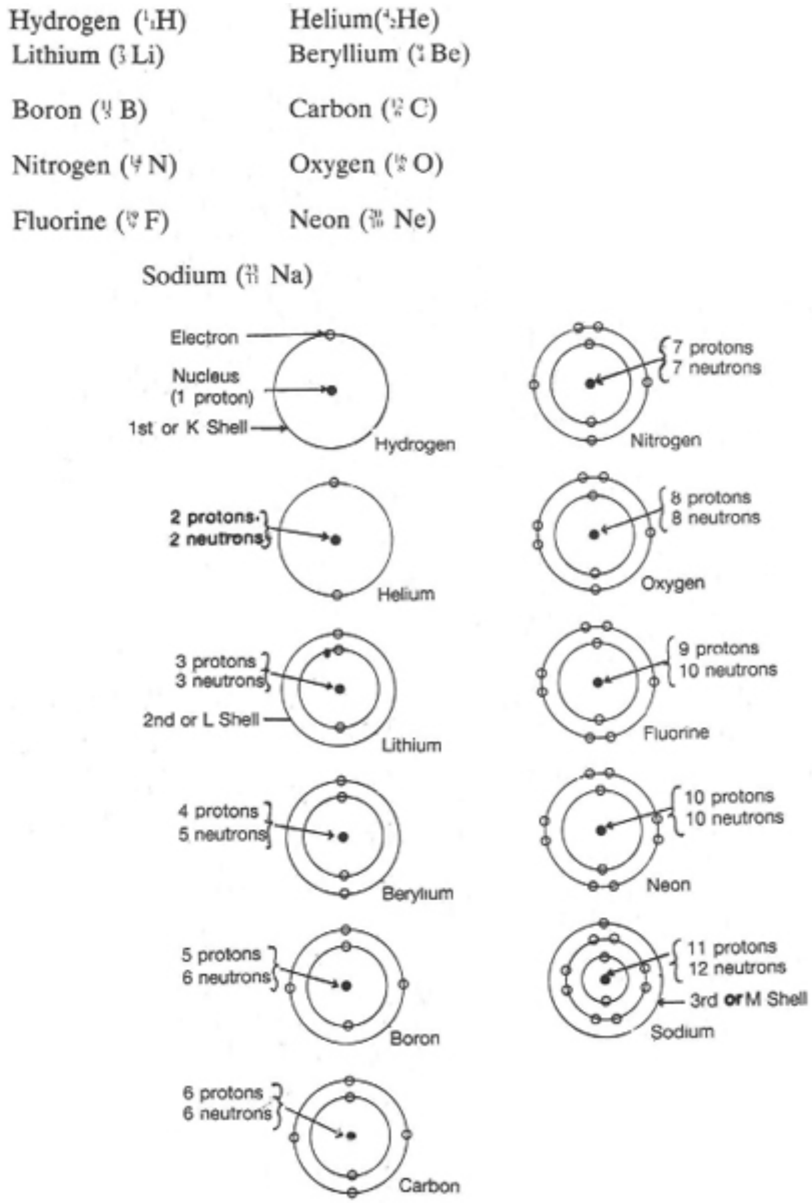


Figure 2.10 Atomic structures of the first eleven elements

A short-hand form of indicating the atomic structures is to represent the shells with the letters **K, L, M, N** . . . and to write the number of electrons in each shell in that order. Table 2.4 gives a summary of this for the first twenty elements.

TABLE 2.4 The Electronic Configurations of some Elements

Element	Atomic Number	Shells K	L	M	N
Hydrogen	1	1			
Helium	2	2			
Lithium	3	2	1		
Berylium	4	2	2		
Boron	5	2	3		
Carbon	6	2	4		
Nitrogen	7	2	5		
Oxygen	8	2	6		
Fluorine	9	2	7		
Neon	10	2	8		
Sodium	11	2	8	1	
Magnesium	12	2	8	2	
Aluminium	13	2	8	3	
Silicon	14	2	8	4	
Phosphorus	15	3	8	5	
Sulphur	16	2	8	6	
Chlorine	17	2	8	7	
Argon	18	2	8	8	
Potassium	19	2	8	8	1
Calcium	20	2	8	8	2

Activity 2.2

Build models for the hydrogen, carbon and sodium atoms. Use different colours of plasticine to indicate the protons, neutrons and electrons. Use wire to make the shells and to hold the shells and the nuclei together.

Exercise 2B

Draw the arrangements of electrons round the nuclei of the following atoms: magnesium, aluminium, sulphur, argon, potassium and calcium.

2.7 Atomic numbers

We have seen that each atom is characterised by a given number of protons. The numbers of protons and electrons in each atom are equal. All atoms are electrically neutral because the number of protons and electrons are equal in each atom. When the number of electrons is not equal to the number of protons, we have an **ion**. This is obtained by increasing or decreasing the number of electrons in a neutral atom.

The number of protons in an atom is equal to its atomic number, and it is a constant. The atomic number characterises the atom of every element. The elements hydrogen, carbon and oxygen which have one, six and eight protons in their nuclei have atomic numbers of 1, 6 and 8 respectively. The atomic numbers of the first twenty elements listed in

Table 2.4 correspond to the number of protons in each element.

Exercise 2c

What is the difference between a particle with nine protons and nine electrons, and another with nine protons but eight electrons? Which of these two particles is an atom?

2.8 Mass number

We mentioned earlier in this chapter that the. mass of the atom is due to the protons and neutrons in the nucleus.

The numerical sum of the masses of the protons and neutrons in the nucleus of an atom is equal to the mass number.

Since each of these particles is of mass equal to 1 a.m.u, the mass number (in a.m.u) is in effect equal to the total number of protons and neutrons in the nucleus.

For example, Carbon which has 6 protons and 6 neutrons in its nucleus has a mass number of 12; Aluminium which has 13 protons and 14 neutrons has a mass number of 27.

2.9 Isotopy

Daltonâ€™s atomic theory postulated that all atoms of an element were alike in all respects including the possession of equal masses. The use of modern instruments has revealed that there are atoms of the same element whose atomic masses are different. Such atoms are called **isotopes**. These atoms have the same number of protons but different numbers of neutrons. Hydrogen, the lightest element, has three isotopes. They are listed in Table 2.5.

TABLE 2.5 Isotopes of Hydrogen

Isotope	Number of Protons	Number of Neutrons	Mass Number
Protium	1	—	1
Deuterium	1	1	2
Tritium	1	2	3

It is customary to represent an isotope by writing the symbol of the element and indicating the atomic and mass numbers as shown below, for protium.

Mass number $\rightarrow$ 1
Atomic number $\rightarrow$ 1 $\quad {}^{1}_{1}H$ $\leftarrow$ Symbol of Atom.

The sub-script is the atomic number of the element, while the super-script is the mass number. For example, ${}^{1}_{1}H$, ${}^{2}_{1}H$, ${}^{3}_{1}H$. respectively represent protium, deuterium and tritium, the isotopes of hydrogen. When the element under reference is obvious, the sub-script may be omitted. Thus, we may simply write ^{1}H, ^{2}H, and ^{3}H for isotopes of hydrogen.

There are only a few elements which do not show isotopy. In Table 2.6 the isotopes of some common elements are indicated.

TABLE 2.6 Isotopes of some common elements

Element	Atomic Number	Number of Neutrons	Mass Number	Isotope
Oxygen	8	8	16	${}^{16}_{8}O$
	8	9	17	${}^{17}_{8}O$
	8	10	18	${}^{18}_{8}O$
Carbon	6	6	12	${}^{12}_{6}C$
	6	7	13	${}^{13}_{6}C$
	6	8	14	${}^{14}_{6}C$
Neon	10	10	20	${}^{20}_{10}Ne$
	10	11	21	${}^{21}_{10}Ne$
	10	12	22	${}^{22}_{10}Ne$
Chlorine	17	18	35	${}^{35}_{17}Cl$
	17	20	37	${}^{37}_{17}Cl$

Exercise 2D

Calculate the number of neutrons in the following isotopes:

$${}^{3}_{2}He,\ {}^{4}_{2}He,\ {}^{14}_{7}N,\ {}^{15}_{7}N,\ {}^{23}_{11}Na,\ {}^{24}_{11}Na.$$

Solution

To answer this question you subtract the number of protons (atomic number) from the mass number. As an example, for $^{3}_{2}He$ there is (3-2) = 1 neutron.

2.10 Relative atomic mass

The ^{12}C isotope has 6 protons and 6 neutrons in its nucleus with a mass number of 12 atomic mass units. It is used as a standard for the comparison of the masses of other atoms.

The relative atomic mass of an element is defined as the number of times the mass of one atom of the element is as heavy as one-twelfth of the mass of one atom of carbon 12 ($^{12}_{6}C$).

Relative atomic mass has no units. It is a ratio. For practical purposes however, a unit of measurement is needed. Chemists often want to know the masses of the elements taking part in a chemical reaction. They have added the gramme to the standard of reference, i.e. ^{12}C = 12g, and found that 12g of ^{12}C isotope contains **6.02 Ã— 10^{23} atoms.** This number of atoms, called the Avogadro number, is known as the **mole. The relative atomic mass of an element expressed in grammes contains one Avogadro number of atoms and is a mole of the element.**

For atoms that exhibit isotopy the relative atomic mass is the weighted average of the atomic masses of its different isotopes. For example, naturally occuring chlorine contains 75% by weight of $^{35}C1$ and 25% of $^{37}C1$. These percentages, called the **relative abundance** of isotopes, is constant for every sample of chlorine. The relative atomic mass of chlorine is found by calculating the relative average of the atomic masses thus:-

Relative atomic mass of chlorine =

$$\frac{(75 \text{ x } 35) + (25 \text{ x } 37)}{100}$$
$$= 35.5$$

Similarly lead, which contains 24% of lead-206, 23% of lead-207 and 53% of lead-208 has a relative atomic mass of

$$\frac{(24 \text{ x } 206) + (23 \text{ x } 207) + (53 \text{ x } 208)}{100}$$
$$= 207.29$$

The existence of isotopes is responsible for the relative atomic masses of elements being fractional. The mass numbers of the different isotopes are however whole numbers.

Worked Example

How many atoms are there in 3.55g of chlorine?

Solution

Note that 3.55 is one tenth of 35.5, i.e l/10th of the relative atomic mass of chlorine. Therefore, it should contain one-tenth the Avogadro number of atoms; i.e. 6.02 Ã— 10^{22} atoms.

Exercise 2 E

(a) Calculate the number of atoms in

(i) 1.2g of carbon (Relative atomic mass 12)

(ii) 2g of hydrogen (Relative atomic mass 1)

(iii) 8g of oxygen (Relative atomic mass 16).

(b) If the relative abundance of ^{20}Ne and ^{22}Ne are 91% and 9% respectively, calculate the relative atomic mass of neon. **(Answer 20.18)**

2.11 Relative molecular mass

The relative molecular mass of one molecule of an element or compound, like the relative atomic mass of one atom of an element is a ratio based on the ^{12}C isotope.

The relative molecular mass of a compound is defined as the number of times one molecule of the compound is as heavy as one-twelfth of the mass of an atom of ^{12}C isotope.

Like relative atomic mass, relativc molecular mass has no units since it is a ratio. It is equal to the sum of the relative atomic masses of all the atoms making up one molecule of the element or compound.

Given that one molecule of hydrogen chloride, HCl has one atom of hydrogen and one atom of chlorine combined together, then the relative molecular mass of hydrogen chloride is (1 + 35.5) = 36.5. Similarly, given that carbon(IV) oxide, CO_2, has one atom of carbon **(relative atomic** mass 12), joined to two atoms of oxygen (relative atomic mass 16), the relative molecular mass of carbon(IV) oxide is 12 + (16 Ã— 2) = 44.

The Avogadro number of molecules of a compound is one mole of the compound.

The mass of one Avogadro number of molecules, or one mole of any compound is numerically equal to the relative molecular mass of the compound expressed in grammes.

Thus, 44 g of carbon(IV) oxide contains 6.02 Ã— 10^{23} molecules; 4.4 g contain 6.02 Ã— 10^{22} molecules; and 0.44g contains 6.02 Ã— 10^{21} molecules.

Exercise 2 F

How many molecules are there in 0.73g of hydrogen chloride?

2.12 Modifications of Daltonâ€™s atomic

theory

Consider the statements of Daltonâ€™s atomic theory. Which of them is still true in the light of modern atomic structure? To answer this question we shall re examine these statements one by one.

1. *All matter are made up of tiny particles called atoms.* This statement is still true. The only addition is that atoms are themselves made up of smaller particles called electrons, protons and neutrons.
2. *Atoms cannot be sub-divided.* Certainly, this is no longer accepted. Some atoms with unstable nuclei disintegrate spontaneously, emitting radioactive rays and forming lighter atoms.
3. *Atoms of the same element are alike in all respects, especially in mass.* This has been disproved by the phenomenon of isotopy. However, atoms of the same element have the same atomic number and react alike. Chemical reactions depend on the electronic configurations of atoms, which are similar for atoms of the same element. But their masses depend on the number of neutrons in each atom. These are not always equal for different atoms of the same element, as illustrated in Table 2.6.
4. *Atoms cannot be created or destroyed.* This statement is no longer acceptable. A number of new elements have been made by man â€“ the trans-uranium elements. We shall describe how they are made in- a later chapter. Also, the disintegration of unstable nuclei yields atoms of lighter elements; that is, these lighter atoms are â€˜createdâ€™ from the unstable heavier ones.

   A large amount of energy is released during the disintegration of unstable nuclei. This energy comes from some of the mass of the atom which is converted to energy. The energy derived from atomic reactions usually comes from some lost mass. Small mass differences sometimes occur between reactants and products. This mass difference is changed to energy. For such reactions the law of conservation of mass does not hold, and cannot be used to show that atoms cannot be destroyed in the course of a chemical reaction. The loss of mass implies the possibility of destruction of matter.
5. *Atoms combine in small whole number ratios.* To a large extent this statement still holds. Polymeric carbon compounds however exist, in which the ratios of combining atoms is far from being simple. We shall discuss some of such compounds in a subsequent chapter.

2.13 Symbols of the elements

An important consequence of the atomic theory is that it became possible to represent an element with a symbol, a compound with a formula and a reaction with an equation. Since atoms of an element were said to be alike, a symbol could represent any atom of a particular element.

Dalton invented symbols which were later found unsuitable. They made the writing of equations cumbersome. Modern symbols are derived from the names of the elements. For some elements the capitalised first letters of their names are used to represent them. Examples include the following:

Element	**Symbol**
Boron	A
Carbon	C
Fluorine	F
Hydrogen	H
Iodine	I
Nitrogen	N
Oxygen	O
Phosphorus	P
Sulphur	S
Uranium	U

For some other elements the capitalised first letter of their names and a small form of another letter in their names are used to represent them. Examples are:

Element	**Symbol**
Aluminium	Al
Argon	Ar
Barium	Ba
Berylium	Be
Bromine	Br
Cadmium	Cd
Caesium	Cs
Calcium	Ca
Chlorine	Cl
Chromium	Cr
Cobalt	Co
Magnesium	Mg

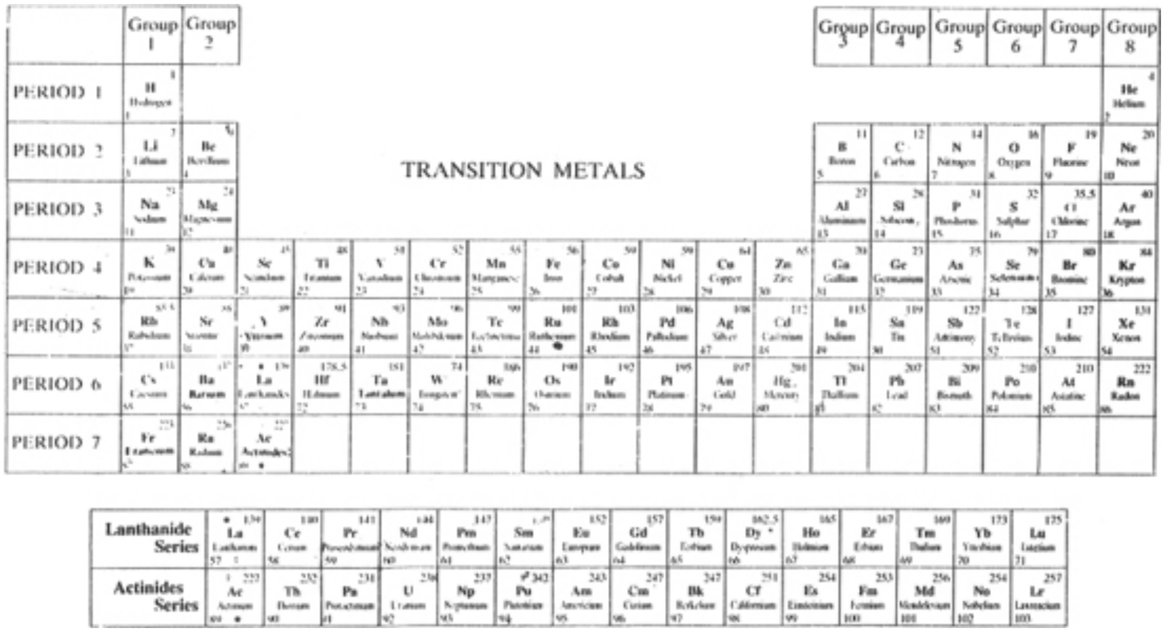
The symbols of a third group of elements are derived from their Latin names. Examples are: -

Element	**Latin Name**	**Symbol**
Copper	*Cuprum*	Cu
Gold	*Aurum*	Au

Iron	*Ferrum*	Fe
Lead	*Plumbum*	Pb
Mercury	*Hydragyrum*	Hg
Potassium	*Kalium*	K
Sodium	*Natrium*	Na
Silver	*Argentum*	Ag
Tin	*Stannum*	Sn

The symbols of most of the known elements are given in the **Periodic Table**.

PERIODIC TABLE



Chapter Summary

1. Matter is made up of atoms which are composed of protons, neutrons and electrons. These three fundamental particles have the characteristics tabulated below:

Particle	Mass (a.m.u)	Charge
Electron	Negligible	â€“1
Proton	1	+1
Neutron	1	0

2. Electrons are located in shells which surround the nucleus. The maximum number of electrons in each shell is equal to $2n^2$; where n is the shell number.
3. Protons and neutrons are located in the nucleus. The sum of the number of protons and neutrons in an atom is equal to its mass number.
4. The atomic number of an element is equal to the number of protons in its nucleus. The number is constant for each element. In every atom, the number of protons is equal to the number of electrons. When the number of electrons is greater or less than that of the protons, the atom is no more neutral and is called an ion.
5. Isotopes are atoms of the same element which have different mass numbers because they contain different numbers of neutrons.
6. The relative atomic mass of an element is the number of times an atom of the element is as heavy as $\frac{1}{12}$ of the mass of an atom of ^{12}isotope.

Assessment

1. What evidences do we have for the existence of atoms? Describe two experiments which illustrate these evidences.
2. State Daltonâ€™s atomic theory. Which statement of that theory is in conflict with the phenomenon of isotopy?
3. Give the charge and mass of each of the three fundamental particles in an atom. How many neutrons are present in the nuclei of the following isotopes:
   $^{31}_{15}P$, $^{37}_{17}Cl$, $^{238}_{92}U$?
4. Distinguish clearly between:
   (a) relative atomic mass and atomic mass.
   (b) atomic number and mass number.
   (c) mole and molecule.
5. Define the terms atomic number, mass number, isotopy and relative atomic number. Explain why chlorine has a fractional relative atomic mass of 35.5.
6. Write down the electronic structures of the following atoms:
   (a) an atom with atomic number 17 and mass number 35.
   (b) $^{27}_{13}X$
7. An atom has atomic number 16, and mass number 34. How many electrons, protons and neutrons has the atom? If another atom of the same element has a mass number of 35, and if the relative abundance of the heavier isotope is 80%, what is the relative atomic mass of the element?

8 (i) Distinguish between the atomic mass and the atomic number of an element.
   (ii) An atom has mass number 34, and 18 neutrons. Draw its electronic configuration.
   (iii) Explain why the atomic mass of chlorine is not a whole number.

9. Fill in the gaps in the following table.

Element	Atomic Number	Number of Neutrons	Mass Number	Electronic Configuration
$^{23}_{11}Na$	–	–	–	2, 8, 1
^{40}Ca	–	–	–	2, 8, 8, 2
$^{12}_{6}C$	–	–	–	–
X	–	20	–	2, 8, 7

10. Write down the electronic structures of the atoms X,Y,Z whose atomic numbers are shown below:

Atom	Atomic Number	Electronic Configuration
X	8	?
Y	12	?
Z	19	?