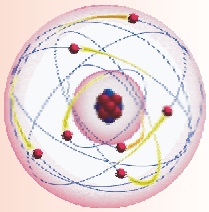
UNIT 1



Structure of the Atom

Unit Outcomes

After completing this unit, you will be able to:

- ***E*** *comprehend Dalton's atomic theory and modern atomic theory;*
- ***E*** *understand the discovery of the electron and the nucleus;*
- ***E*** *know the terms like atomic number, mass number, atomic mass, isotope, energy level, valence electrons, and electron configuration;*
- ***E*** *understand the Dalton, Thomson, Rutherford, Bohr and the quantum mechanical atomic models;*
- ***E*** *develop skills in:*
  - *determining the number of protons, electrons, and neutrons of atoms from atomic numbers and mass numbers,*
  - *calculating the atomic masses of elements that have isotopes,*
  - *writing the ground-state electron configurations of atoms using sub-energy levels and drawing diagrammatic representations of atoms; and*
- ***E*** *demonstrate scientific inquiry skills: observing, comparing and contrasting, communicating, asking questions, and applying concepts.*

MAIN CONTENTS

Start-up Activity

Scientists have been able to gather information about atoms without actually seeing them. Perform the following activity to get an idea about the structure of the atom.

Take an onion. It looks solid. Peel off a layer, and you will find another layer underneath. Layer after layer surfaces as the onion is peeled off. Keep on peeling to its core.

Analysis

1. How do you compare this with the atomic model?
2. What do the layers represent in the atomic model?
3. What does the core represent in the atom?

Two thousand years ago, ancient Greek philosophers developed a theory of matter that was not based on the experimental evidences. A notable Greek philosopher, namely, Democritus, believed that all matter was composed of very tiny, indivisible particles. He called them atomos. Hence, the word 'atom' came from the Greek word atomos, which means uncuttable or indivisible.

Aristotle was part of the generation that succeeded Democritus. He did not believe in atomos. Aristotle thought that all matter was continuous. That is, if one proceeded on breaking down a substance, it would be impossible to reach to the last indivisible particle. In other words, it would continue to divide infinitely. His opinion was accepted for nearly 200 years.

The early concept of atoms was simply a result of thinking and reasoning on the part of the philosophers, instead of experimental observations. In 1803, however, John Dalton proposed a completely different theory of matter. His theory was based on scientific experimental observations and logical laws. These scientific assumptions were

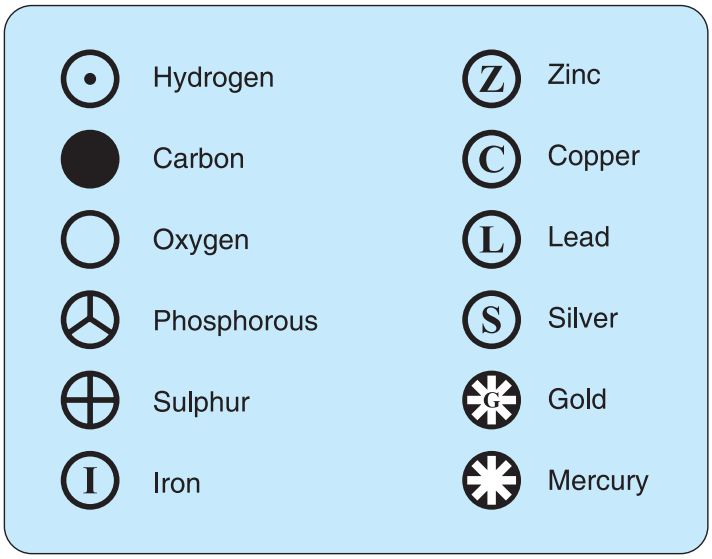


Figure 1.1 Dalton's atomic symbols.

very closely related to what is presently known about the atom. Due to this, John Dalton is often referred to as the father of modern atomic theory.

Dalton also worked on the relative masses of atoms and gave symbols to some elements as illustrated in Figure 1.1.

1.1 Atomic Theory

Competencies

By the end of this section, you will be able to:

- describe Dalton's atomic theory;
- describe modern atomic theory; and
- compare and contrast Dalton's atomic theory with modern atomic theory.

Activity 1.1

Form a group and discuss the following idea: Dalton's contribution was different from that of the ancient Greeks who postulated the existence of atoms. Point out the differences between the two ideas. Present your discussion to the class.

Even though the idea of the existence of atoms and atomic theory dates back to classical times, as discussed earlier, only Dalton's atomic theory ideas were the basis for the new era of science.

1.1.1 Dalton's Atomic Theory

Historical Note

John Dalton

British physicist and chemist John Dalton is best known for developing the atomic theory of elements and molecules, the foundation of modern physical science. While pondering the nature of the atmosphere during a meteorological study in the early 1800s, Dalton deduced the structure of carbon dioxide and proposed that an exact number of atoms constitute each molecule. He held that all atoms of a given element are identical and different from the atoms of every other element. The first to classify elements according to their atomic weights, Dalton set the stage for a revolution in scientific thought.

In 1804, John Dalton developed the first modern theory of atoms and proposed their existence. Dalton's atomic theory was based directly on the ideas of elements and compounds, and on the three laws of chemical combination. The three laws are:

i) The *law of conservation of mass* states that matter is neither created nor destroyed. This law is also called the law of indestructibility of matter. It means that the mass of the reactants is exactly equal to the mass of the products in any chemical reaction. A chemical reaction involves only the separation and union of atoms.

ii) The *law of definite proportions* states that a pure compound is always composed of the same elements combined in a definite ratio by mass. For example, water (H_2O) is composed of hydrogen and oxygen only. These elements are always in the proportion of 11.19% hydrogen to 88.81% oxygen by mass and in the proportion 2 : 1 by volume.

iii) The *law of multiple proportions* states that when two different compounds are formed from the same elements, the masses of one of the elements in the two compounds, compared to a given mass of the other element, is in a small whole-number ratio. For example, carbon and oxygen form two compounds: carbon monoxide and carbon dioxide. Carbon monoxide contains 1.3321 g of oxygen for each 1 g of carbon, whereas carbon dioxide contains 2.6642 g of oxygen for each 1g of carbon. Hence, carbon dioxide contains twice the mass of oxygen as does carbon monoxide.

Activity 1.2

Form a group and perform the following task. Present your findings to the class.

Using two chemical compounds as an example, describe the difference between the law of definite proportions and the law of multiple proportions.

Dalton proved that these laws are entirely reasonable if the elements are composed of tiny particles, which he called atoms**.** An atom is the smallest fundamental particle of an element.

The basic postulates of Dalton's Atomic Theory are summarized as follows:

1. All elements are made up of small particles called atoms.
2. Atoms are indivisible and indestructible.
3. All atoms of a given element are identical in mass and in all other properties.
4. Atoms are neither created nor destroyed in chemical reactions.
5. Compounds are formed when atoms of more than one element combine.
6. In a given compound, the relative numbers and types of atoms are constant.

Although some of these postulates, like postulate number 2 and number 3, have been shown to be incorrect by later work, Dalton's theory was a brilliant and logical explanation of many experimental discoveries and laws that were known at that time.

Activity 1.3

Form a group. Discuss the following laws and present to the class.
Use Dalton's atomic theory to explain:

a The law of conservation of mass
b The law of definite proportions
c The law of multiple proportions.

1.1.2 Modern Atomic Theory

Towards the end of the 19th century, various experimental discoveries revealed the existence of subatomic particles, isotopes, and so on. In light of these findings, Dalton's Atomic Theory was modified.

The Modern Atomic Theory can be summarized as follows:

1. Atoms are the smallest particles of all elements that can take part in a chemical reaction.
2. An atom is divisible. It can be subdivided into electrons, protons, and neutrons. An atom is also indestructible *i.e.*, atoms can neither be created nor destroyed during ordinary chemical reactions.
3. Atoms of the same element may not be identical in mass because of the existence of isotopes.
4. Atoms of the same elements have identical chemical properties.
5. Atoms of different elements have different chemical properties.
6. Atoms of two or more elements combine in simple whole-number ratios to form compounds.

Exercise 1.1

1. Name the postulates of Dalton's atomic theory.
2. Compare and contrast Dalton's atomic theory with modern atomic theory.

1.2 DISCOVERIES OF THE FUNDAMENTAL SUBATOMIC PARTICLES AND THE ATOMIC NUCLEUS

Competencies

By the end of this section, you will be able to:

- explain the discovery of the electron;
- explain the discovery of the nucleus;
- explain the discovery of the neutron.

Activity 1.4

Form a group and discuss the following ideas. Present your discussion to the rest of the class.

1. In your grade 7 chemistry lesson, you have learned that atoms consist of three fundamental subatomic particles. What are these? Where is their location in the atoms? What are their charges?
2. What were the consequences of the discoveries of subatomic particles on Dalton's atomic theory?

In the late 1880's, John Dalton thought atoms were indivisible. However, a series of investigations and astounding discoveries clearly demonstrated that atoms are made up of smaller particles, called subatomic particles. The three fundamental subatomic particles of an atom are electrons, protons, and neutrons.

1.2.1 Discovery of the Electron

Do you think that there is a similarity between cathode rays and electrons?

The electron was the first subatomic particle to be identified. In 1879, William Crooke studied electrical discharges in partially evacuated tubes called discharge tubes.

Two electrodes from a high-voltage source are sealed into a glass tube from which air has been evacuated. The negative electrode is the cathode and the positive one is the anode. When the high-voltage current is turned on, the glass tube emits a greenish light and a beam of light is seen at the anode. These rays flow from the cathode towards the anode in a straight line. They are called cathode rays. Later on, in 1897, J.J. Thomson found that the beam was deflected by both electrical and magnetic

fields. In an electric field, they bend toward the positive plate. From this he concluded that the cathode rays are made up of very small negatively charged particles, which he named electrons. He concluded that electrons are constituents of all matter, because he obtained the same results when he changed the gas and the electrode in the tube.

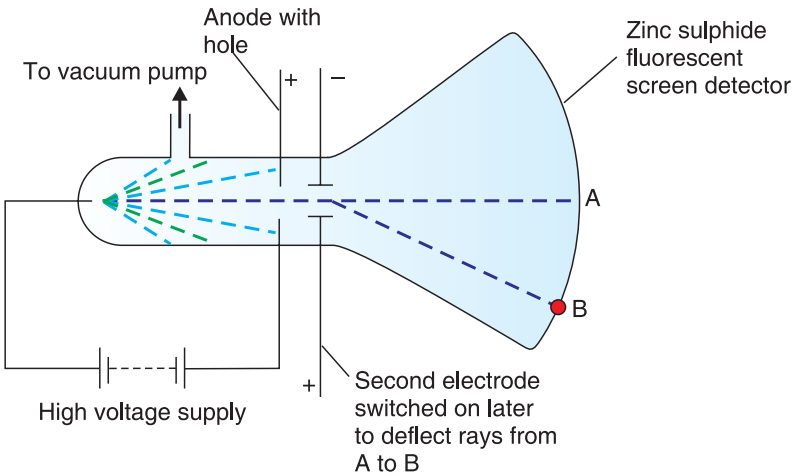


Figure 1.2 A simple cathode-ray tube.

Properties of Cathode Rays

Cathode rays possess the following properties:

1. An object placed between the cathode and the opposite end of the tube, casts a shadow on the glass. This shows that the cathode rays travel in a straight line.
2. A paddle wheel placed in the path of cathode rays rotates. This indicates that cathode rays are of particle nature-the particles strike the paddle and therefore move the wheel.

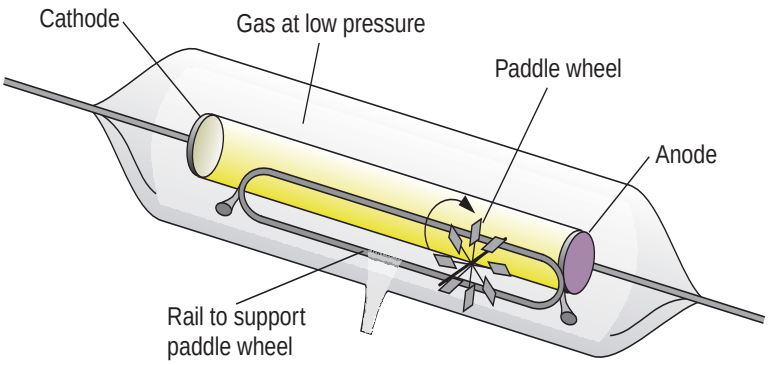


Figure 1.3 A paddle wheel placed in the path of cathode ray.

3. When an electric or magnetic field is applied in the path of cathode rays, they are deflected towards the positive plate. This shows that cathode rays are negatively charged.

4. The properties of cathode rays (like charge to mass ratio, e^-/m ratio) do not depend on the nature of the gas in the discharge tube or on the material of the cathode. This shows that electrons are present in the atoms of all elements.

Thus, cathode-ray experiments provided evidence that atoms are divisible into small particles, and that one of the atom's basic constituents is the negatively charged particle called electron.

PROJECT WORK

Produce a model of cathode ray tube from simple and locally available materials such as glass tube, copper wire and rubber stoppers. Using your model show how cathode rays are generated and explain their properties.

Charge and Mass of an Electron

J.J. Thomson studied the deflection of cathode rays under the simultaneous application of electric and magnetic fields, that are applied perpendicular to each other. His experiment led to the precise determination of the charge-to-mass ratio (e^-/m) of an electron, and this 'e^-/m' value was found to be 1.76×10^8 coulomb/g. Coulomb is a unit of electric charge.

Activity 1.5

Form a group and discuss how J.J Thomson's determination of the charge-to-mass ratio of the electron led to the conclusion that atoms were composed of sub atomic particles. Present your findings to the class.

In 1909, Robert Millikan determined the charge of an electron (e^-), using the oil drop experiment. He found the charge of an electron to be 1.60×10^{-19} coulombs. Combination of e^-/m and e^- values are used to determine the mass of an electron, which is found to be 9.11×10^{-31} kg.

From Thomson's experiment, $e^-/m = 1.76 \times 10^8$ C/g

From Millikan's experiment, $e^- = 1.60 \times 10^{-19}$ C

$$\therefore \text{ Mass of the electron} = m = \frac{e^-}{e^-/m} = \frac{1.60 \times 10^{-19}\ \text{C}}{1.76 \times 10^{8}\ \text{C/g}}$$

$$= 9.11 \times 10^{-28}\ \text{g} = 9.11 \times 10^{-31}\ \text{kg}$$

From the above discussion, it follows that:

An electron is a fundamental particle of an atom carrying a negative charge and having a very small mass. The mass of an electron is approximately $1/1837$ times the mass of a hydrogen atom.

Exercise 1.2

1. Explain how the electron was discovered.
2. Give explanation on the works of J.J. Thomson and R. Millikan.

1.2.2 Discovery of the Atomic Nucleus

Radioactivity

In 1895, the German physicist, Wilhelm Röntgen noticed that cathode rays caused glass and metals to emit very unusual rays. This highly energetic radiation penetrated matter, darkened covered photographic plates, and caused a variety of substances to fluoresce. Since these rays could not be deflected by a magnet, they could not contain charged particles as cathode rays do. Röntgen called them X-rays.

Not long after Röntgen's discovery, Antoine Becquerel, a professor of physics in Paris, began to study fluorescent properties of substances. He found that exposing thickly wrapped photographic plates to a certain uranium compound caused them to darken, even without the stimulation of cathode rays. Like X-rays, the rays from the uranium compound were highly energetic and could not be deflected by a magnet. Also, these rays were generated spontaneously. One of Becquerel's students, Marie Curie, suggested the name radioactivity to describe this spontaneous emission of particles and/or radiation. Consequently, any element that spontaneously emits radiation is said to be radioactive.

Further investigation revealed that three types of rays are produced by the "decay", or breakdown of radioactive substances such as uranium, polonium, and radium. These are alpha, beta, and gamma rays.

Alpha (α) rays consist of positively charged particles called α-particles. They are deflected by positively charged plates. Beta (β) rays or β-particles are electrons of nuclear origin that are deflected by negatively charged plates. The third type of radioactive radiation consists of high-energy rays called gamma (γ) rays. Like X-rays, γ-rays have no charge and are not affected by an external electric or magnetic field.

Activity 1.6

Form a group and discuss the following ideas. Present your discussion to the class.

Scientists discovered the three subatomic particles (electrons, protons and neutrons). Compare and contrast these fundamental sub atomic particles with alpha particles, beta particles and gamma rays in terms of the nature of the particles.

More details about the structure of an atom were provided in 1911 by Ernest Rutherford and his associates, Hans Geiger and Ernest Marsden. The scientists

bombarded a thin gold foil with fast moving alpha particles. Alpha particles are positively charged particles with about four times the mass of a hydrogen atom. Geiger and Marsden assumed that mass and charge were uniformly distributed throughout the atoms of the gold foil. Therefore, they expected the alpha particles to pass through the gold foil with only a slight deflection. However, when the scientists checked for the possibility of wide-angle deflections, they were shocked to find that roughly 1 in 8000 of the alpha particles had actually been redirected back toward the source.

After thinking about the observations for two years, Rutherford finally came up with an explanation. He reasoned that the rebounded alpha particles must have experienced some powerful force within the atom. He concluded that the source of this force must occupy a very small amount of space because very few of the total number of alpha particles had been affected by it. He also concluded that the force must be caused by a very densely packed bundle of matter with a positive electric charge. Rutherford called this small, dense, positive bundle of matter the nucleus.

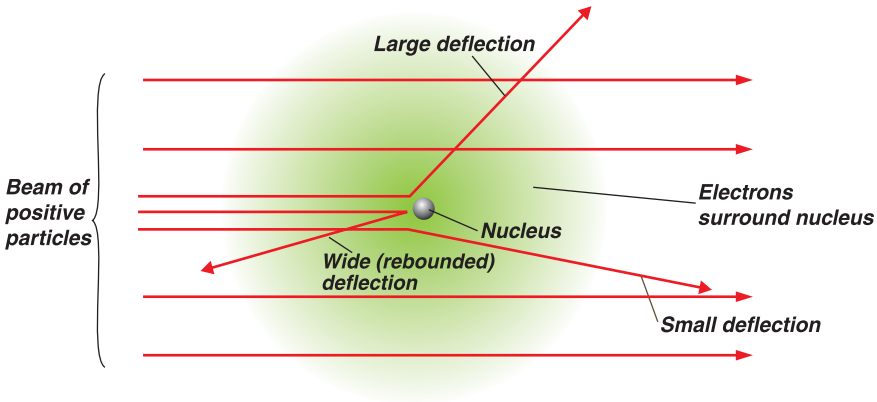


Figure 1.4 Rutherford's α – particles experiment.

Rutherford's Conclusion

1. Since most of the α-particles passed through the gold foil undeflected, most of the space in an atom is empty.
2. Some of the α-particles were deflected by small angles. This indicated the presence of a heavy positive centre in the atom, which Rutherford named the nucleus.
3. Only a few particles (1 in about a million) were either deflected by a very large angle or deflected back. This confirmed that the space occupied by the heavy positive centre must be very small.

Activity 1.7

Form a group and perform the following task. Present your findings to the class.

Predict what Rutherford might have observed if he had bombarded copper metal instead of gold metal with alpha particles.

From the angles through which alpha particles are deflected, Rutherford calculated that the nucleus of an atom has a radius of about 10^{-15} m. The radius of the whole atom is about 10^{-10} m. Therefore, the nucleus is about 1/10,000 (one ten-thousandth) of the size of the atom as a whole. Thus, if we magnified an atom to the size of a football stadium (about 100 m across), the nucleus would be represented by a pea placed at the centre of the pitch.

Rutherford suggested that the negatively charged electrons surrounding the positively charged nucleus are like planets around the sun revolving. He could not explain, however, what kept the electrons in motion around the nucleus.

1.2.3 Discovery of Neutrons

In 1932, an English scientist, James Chadwick, identified the existence of neutrons. He bombarded a thin foil of beryllium with α-particles of a radioactive substance. He then observed that highly penetrating rays, consisting of electrically neutral particles of a mass approximately equal to that of the proton, were produced. These neutral particles are called neutrons.

A neutron is a subatomic particle carrying no charge and having a mass of 1.675×10^{-24} g. This mass is almost equal to that of a proton or of a hydrogen atom.

Exercise 1.3

1. Which experimental evidence indicates that:
   i) electrons are negatively charged particles?
   ii) electrons are the constituents of all atoms?
2. In Rutherford's experiment, most of the alpha particles passed through the gold foil undeflected. What does this indicate?
3. In Rutherford's experiment, some of the α-particles were deflected by small angles. What does this indicate?
4. What was determined from Millikan's oil drop experiment?
5. Three particles are fired into a box that is positively charged on one side and negatively charged on the other side. The result is shown in Figure 1.5.

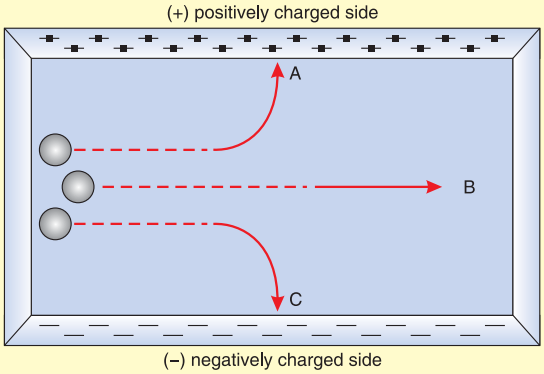


Figure 1.5 particles fired into a box.

a Which particle is a proton?

b Which particle is an electron?

c Which particle is a neutron?

Activity 1.8

Hydrogen is the most abundant element in the universe. It is the fuel for the sun and other stars. It is currently believed that there are roughly 2,000 times more hydrogen atoms than oxygen atoms. There are also 10,000 times more hydrogen atoms than carbon atoms.

Make a model of a hydrogen atom, using materials of your choice, to represent a hydrogen atom, including the proton and electron. Present the model to the class, and explain in what ways your model resembles a hydrogen atom.

1.3 COMPOSITION OF AN ATOM AND ISOTOPES

Competencies

By the end of this section, you will be able to:

- write the relative charges of an electron, a proton, and a neutron;
- tell the absolute and relative masses of an electron, a proton, and a neutron;
- tell the number of protons and electrons in an atom from the atomic number of the element;
- determine the number of neutrons from given values of atomic number and mass number;

- explain the terms atomic mass and isotope, and
- calculate the atomic masses of elements that have isotopes.

Activity 1.9

Based on your previous knowledge discuss the following concepts in groups, and share your ideas with the rest of the class.

1. Would you think that it is possible for someone to discover a new element that would fit between magnesium (atomic number 12) and aluminum (atomic number 13)?
2. You are given the following three findings:
   - a An atom of calcium has a mass of 40 a.m.u. Another atom of calcium has 44 a.m.u.
   - b An atom of calcium has a mass of 40 a.m.u. An atom of potassium has a mass of 40 a.m.u.
   - c An atom of calcium has a mass of 40 a.m.u. An atom of cobalt has a mass of 59 a.m.u.

Are these findings in agreement with Dalton's atomic theory? Discuss your conclusions in your group.

All atoms have two regions. The nucleus is a very small region located at the center of an atom. Except for the nucleus of the simplest type of hydrogen atom, all atomic nuclei are made of two kinds of particles, protons and neutrons. Surrounding the nucleus, there is a region occupied by negatively charged particles called electrons. This region is very large compared with the size of the nucleus. A proton has a positive charge, which is equal in magnitude to the negative charge of an electron. Atoms are electrically neutral because they contain equal numbers of protons and electrons.

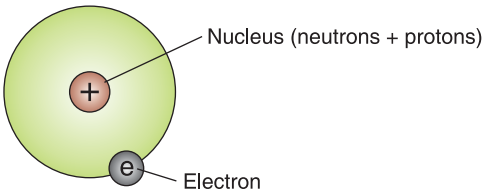


Figure 1.6 A simple representation of an atom.

A proton has a mass of 1.673×10^{-27} kg, and a neutron has a mass of 1.675×10^{-27} kg. Hence, protons and neutrons have approximately the same mass. Most of the mass of the atom is concentrated in the nucleus, assuming the mass of an electron to be negligible or almost zero.

Activity 1.10

Make a group and perform the following task: Sketch a modern atomic model of a chlorine atom and identify where each type of subatomic particles would be located. Submit the model to your teacher.

Activity 1.11

Make a group and discuss, "If atoms are primarily composed of empty space, why can't you pass your hand through a solid object"? Present your finding to the class.

Physicists have identified other subatomic particles. But particles other than electrons, protons, and neutrons have little effect on the chemical properties of matter. Table 1.1 gives a summary of the properties of electrons, protons, and neutrons. (The relative electric charge, relative mass, and actual mass are discussed in the next section.)

Table 1.1 Properties of Subatomic Particles.

Particle	Symbols	Relative electric charge	Relative mass (a.m.u)	Actual mass (kg)
Electron	e^-	–1	0.0005486 ≈ 0	9.109×10^{-31}
Proton	p^+	+1	1.007276 ≈ 1	1.673×10^{-27}
Neutron	n^o	0	1.008665 ≈ 1	1.675×10^{-27}

It is convenient to think of the region occupied by the electrons as an electron cloud, a cloud of negative charge. The radius of an atom is the distance from the center of the nucleus to the outer portion of this electron cloud. This is about 10^{-10} m = 100 pm = 0.1 nm. Because atomic radii are so small, they are expressed using a unit that is specifically convenient for the sizes of atoms. This unit can be picometer, nanometer, micrometer, etc. The abbreviation for the

picometer is pm (1 pm = 10^{-12} m = 10^{-10} cm)
nanometer is nm (1 nm = 10^{-9} m = 10^{-7} cm)
micrometer is μm (1μm = 10^{-6} m = 10^{-4} cm)

1.3.1 Atomic Number and Mass Number

Activity 1.12

Form a group and perform the following task. Present your findings to the class.

The table below gives the number of electrons, protons and neutrons in atoms or ions for a number of elements.

Atom or ion of elements	A	B	C	D	E	F	G
Number of electrons	5	7	9	12	7	6	9
Number of protons	5	7	10	10	7	5	9
Number of neutrons	5	7	10	10	8	6	10

Based on the table, answer the following questions:

i) Which of the species are neutral?

ii) Which of them are negatively charged?

iii) Give the symbolic representations for 'B' , 'D' and 'F'

All atoms are composed of the same basic particles. Yet all atoms are not the same. Atoms of different elements have different number of protons. Atoms of the same element have the same number of protons.

The Atomic Number (Z) of an element is equal to the number of protons in the nucleus of an atom. It is also equal to the number of electrons in the neutral atom.

$$Z = p^+ \quad \text{or} \quad Z = e^- \qquad \text{(in a natural atom)}$$

For different elements, the nuclei of their atoms differ in the number of protons they contain. Therefore, they differ in the amount of positive charge that their atoms possess. Thus, the number of protons in the nucleus of an atom determines the identity of an atom.

Each atom of an element is identified by its atomic number. In the periodic table, the atomic number of an element is indicated below its symbol. Notice that the elements are placed in order of increasing atomic number. Hydrogen, H, has atomic number 1, hence, atoms of the element hydrogen have one proton in the nucleus. Next is helium, He, which has two protons in its nucleus. Lithium, Li, has three protons. Beryllium, Be, has four protons, and so on.

Mass Number (A) is the sum of the number of protons and neutrons in the nucleus of an atom. Collectively, the protons and neutrons of an atom are called its nucleons.

$$\text{Mass Number } (A) = \text{Number of protons} + \text{Number of neutrons}$$

$$A = p^+ + n^\circ$$

The name of an element, its atomic number and mass number can be represented by a shorthand symbol. For example, to represent neutral atoms of magnesium with 12 protons and 12 neutrons, the symbol is $^{24}_{12}Mg$. The atomic number is written as a subscript to the left of the symbol of magnesium. The mass number is written as a superscript to the left of the symbol. This method can be represented in general by the following illustration, in which X stands for any chemical symbol.

$$^{\text{Mass Number}}_{\text{Atomic Number}}X \quad \text{or} \quad ^{A}_{Z}X$$

$^{24}_{12}Mg$ can be read as "Magnesium-24", and "Magnesium-25" represents $^{25}_{12}Mg$. In a similar manner, $^{23}_{11}Na$ can be read as "Sodium-23".

The number of neutrons in an atom is found by subtracting the atomic number from the mass number. For example, the number of neutrons for a silver atom ($^{108}_{47}Ag$) is:

Mass number – Atomic number = Number of neutrons

$$108 - 47 = 61$$

∴ The number of neutrons is 61.

Example 1

How many protons, electrons, and neutrons are found in an atom of bromine-80 if its atomic number is 35?

Solution:

Atomic number of bromine is 35.

Mass number of bromine is already given as 80. Now we can write bromine, Br in the form of:

$$^{A}_{Z}X \quad \text{as} \quad ^{A}_{Z}Br = ^{80}_{35}Br$$

Mass number – Atomic number = Number of neutrons

$$80 - 35 = 45 \text{ neutrons}$$

For a neutral atom: number of protons = number of electrons = atomic number.

∴ Number of protons = Number of electrons = 35

Activity 1.13

Discuss the following idea in groups and present your discussion to the class.

Why do all atoms of a chemical element have the same atomic number although they may have different mass numbers?

Exercise 1.4

Determine the atomic number, mass number, and number of neutrons in

a P b $^{39}_{19}K$ c $^{56}_{26}Fe$

1.3.2 Isotopes and Atomic Mass

The simplest atoms are those of hydrogen. Each hydrogen atom contains one proton only. However, hydrogen atoms contain different numbers of neutrons.

Three types of hydrogen atoms are known. The most common type of hydrogen is called protium. It accounts for 99.985% of the hydrogen atoms found on Earth. The nucleus of a protium atom consists of one proton only, and it has one electron moving around it. The second form of hydrogen atom is called deuterium, which accounts for 0.015% of the Earth's hydrogen atoms. Each deuterium atom has a nucleus containing

Activity 1.14

Form a group and discuss the following idea and present your discussion to the class.

Nitrogen has two naturally occurring isotopes, N-14 and N-15. The atomic mass of nitrogen is 14.007. Which isotope is more abundant in nature? Explain.

one proton and one neutron. The third form of hydrogen atom is known as tritium, and it is radioactive. It exists in very small amounts in nature, but it can be prepared artificially. Each tritium atom contains one proton and two neutrons. Protium, deuterium and tritium are the three isotopes of hydrogen.

Isotopes are the atoms of the same element that have the same atomic number but different masses number. The isotopes of a particular element have the same number of protons and electrons but different number of neutrons. Isotopes of an element may occur naturally, or they may be made in the laboratory (artificial isotopes). Since isotopes have the same number of protons and electrons, they have the same chemical properties.

Designating Isotopes

Isotopes are usually identified by specifying their mass number. There are two methods for specifying isotopes. In the first method, the mass number is written with a hyphen after the name of the element. Tritium, for example, is written as Hydrogen-3. We will refer to this method as hyphen notation. The second method for

designating an isotope is by indicating its isotopic nuclear composition. For example, Hydrogen-3 is written as $^{3}_{1}H$. The following table illustrates this designation of isotopes for hydrogen and helium.

Table 1.2 Isotopes of Hydrogen and Helium.

Isotope	Nuclear symbol	No. of protons	No. of electrons	No. of neutrons	Z	A
Hydrogen-1 (protium)	$^{1}_{1}H$	1	1	0	1	1
Hydrogen-2 (deuterium)	$^{2}_{1}H$	1	1	1	1	2
Hydrogen-3 (tritium)	$^{3}_{1}H$	1	1	2	1	3
Helium-3	$^{3}_{2}He$	2	2	1	2	3
Helium-4	$^{4}_{2}He$	2	2	2	2	4

Exercise 1.5

1. Write the nuclear symbol for carbon-13 and carbon-12.
2. The three atoms of elements X, Y and Z are represented as:

   $^{23}_{11}X$, $^{27}_{13}Y$ and $^{31}_{15}Z$. Determine the number of protons, electrons, and neutrons in the neutral atoms indicated.

Relative Atomic Masses

Actual masses of atoms, measured in grams, are very small. An atom of oxygen-16, for example, has a mass of 2.657×10^{-23}g. For most chemical calculations, it is more convenient to use relative atomic masses. In order to set up a relative scale of atomic mass, one atom has been arbitrarily chosen as the standard and assigned a relative mass value. The masses of all other atoms are expressed in terms of this defined standard.

The standard used by scientists to determine units of atomic mass is the carbon-12 atom. It has been arbitrarily assigned a mass of exactly 12 atomic mass units, or 12 a.m.u. One atomic mass unit, or 1 a.m.u, is exactly 1/12 the mass of a carbon-12 atom, or 1.660×10^{-24} g. The atomic mass of any atom is determined by comparing it with the mass of the carbon-12 atom. The hydrogen-1 atom has an atomic mass of about 1/12 of that of the carbon-12 atom, or about 1 a.m.u. The precise value of the atomic mass of a hydrogen-1 atom is 1.007 a.m.u. An oxygen-16 atom has about 16/12 (or 4/3 the mass of a carbon-12 atom). Careful measurements show the atomic mass of oxygen-16 to be 15.994 a.m.u. The mass of a magnesium-24 atom is found to be slightly less than twice that of a carbon-12 atom. Its atomic mass is 23.985 a.m.u.

The masses of subatomic particles can also be expressed on the atomic-mass scale. The mass of an electron is 0.0005486 a.m.u, that of the proton is 1.007276 a.m.u, and that of the neutron is 1.008665 a.m.u. Note that the proton and neutron masses are close to 1 a.m.u but not equal to 1 a.m.u. You have learned that the mass number is the total number of protons and neutrons in the nucleus of an atom. You can now see that the mass number and relative atomic mass of a given nucleus are quite close to each other. They are not identical because the proton and neutron masses deviate slightly from 1 a.m.u.

Activity 1.15

Form a group and perform the following task, based on the given information:

Magnesium has three naturally occurring isotopes, namely Mg-12, Mg-13, and Mg-14.

Make models of three isotopes of magnesium using locally available materials such as clay, colours and sticks, and depict the nucleus, the positions of protons, electrons and neutrons.

Present your model to the class.

Average Atomic Masses of Elements

Most elements occur naturally as mixtures of isotopes, as shown in the examples in Table 1.3. The percentage of each isotope in the naturally occurring elements is nearly always the same, no matter where the element is found. The percentage at which each isotope of an element occurs in nature is taken into account when calculating the average atomic mass of the element. Average atomic mass is the weighted average of the atomic masses of the naturally occurring isotopes of an element.

Table 1.3 Atomic masses and abundances of some natural occurring isotopes.

Isotopes	Mass number	% natural abundance	Atomic mass (a.m.u)	Average atomic mass of element (a.m.u)
Hydrogen-1	1	99.985	1.6078	1.00794
Hydrogen-2	2	0.015	2.0141	
Oxygen-16	16	99.762	15.994	15.9994
Oxygen-17	17	0.038	16.999	
Oxygen-18	18	0.200	17.999	

* Since tritium is radioactive and exists in a very small amount in nature, it is not included in the calculation of atomic mass of hydrogen.

The average atomic mass of an element is calculated by summing up the products of relative atomic mass and percentage abundance of each isotope and dividing by 100.

Example 2

Chlorine has two isotopic compositions: 75.77 % of $^{35}_{17}Cl$ and 24.23 % of $^{37}_{17}Cl$. Determine the average atomic mass of chlorine.

Solution:

75.77 % of natural chlorine exists as chlorine 35, and the rest 24.23 % as chlorine 37.

The average atomic mass of chlorine is calculated as follows:

$$= \frac{(75.77 \times 35) + (24.23 \times 37)}{100} = 35.45$$

Thus, the average atomic mass of chlorine is 35.45 a.m.u

Example 3

Calculate the average atomic mass of boron, using the following data:

Isotope	Relative masses	Percent abundance
$^{10}_{5}B$	10.0134	19.70
$^{11}_{5}B$	11.0094	80.30

Solution:

The average atomic mass of boron is obtained by taking the sum of the products of the relative atomic mass and the percentage abundance of the first and the second isotopes of boron and dividing by 100.

$$\text{Average atomic mass of B} = \frac{(10.0134 \times 19.70) + (11.0094 \times 80.30)}{100}$$

$$= 10.813 \text{ a.m.u}$$

Exercise 1.6

Give appropriate answers for the following questions.

1. What are isotopes? Explain their symbolic designation using the isotopes of hydrogen.

2. The three isotopes of uranium are: $^{234}_{92}U$, $^{235}_{92}U$ and $^{238}_{92}U$. How many protons, neutrons and electrons are present in each isotope?
3. Indicate the number of fundamental particles present in an atom of $^{206}_{82}Pb$.
4. The metal thallium occurs naturally as 30% thallium-203 and 70% thallium-205. Calculate the atomic mass of thallium.

Critical Thinking

5. Why is the gravitational force in the nucleus so small?
6. Could a nucleus of more than one proton but no neutron exist? Explain.

1.4 ATOMIC MODELS

Competencies

By the end of this section, you will be able to:

- name the five atomic models,
- describe Dalton's, Thomson's and Rutherford's atomic models,
- state Bohr's postulates,
- describe Bohr's model,
- describe the quantum mechanical model,
- describe main-energy level and sub-energy level,
- define the term electronic configuration,
- write the ground state-electronic configuration of the elements,
- draw diagrams to show the electronic configuration of the first 18 elements,
- write the electronic configuration of the elements using sub–energy levels,
- write electronic configuration of elements using noble gas as a core and
- describe valence electrons.

Activity 1.16

Discuss the following concept in groups and present your conclusion to the class. Have you heard the word "*model*" in your everyday life? Car makers produce different models as time passes. What is your understanding about the atomic model?

Atoms are too small to be observed directly. Hence, it is better to develop a tentative mental picture (*model*) of the atomic concept. The ideas about atoms have changed

many times. But, at present we can reduce these ideas into five models: Dalton's Model, Thomson's Model, Rutherford's Model, Bohr's Model, and the Quantum Mechanical Model.

i) Dalton's Atomic Model

This model of atoms was accepted for 100 years without serious challenge. He thought of atoms as solid indestructible spheres. He called it the billiard ball model.

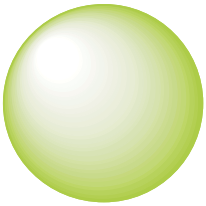


Figure 1.7 Dalton's model of an atom.

ii) Thomson Model

After learning that atoms contain electrons, Thomson proposed a new model of the atom, which is shown in Figure 1.8. It is sometimes called the plum-pudding model. In his model of an atom, Thomson proposed thought that electrons were embedded inside a positively charged sphere, just like plums in a pudding. Today we might call Thomson's model the *chocolate chip ice-cream model*.

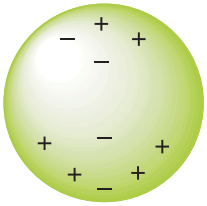


Figure 1.8 Thomson's model of an atom.

Activity 1.17

Perform the following tasks in groups and present your findings to the class.

1. Draw a sketch diagram of a carbon atom using Thomson's atomic model.
2. What was incorrect about Thomson's model of the atom?

iii) Rutherford Model

Why was Thomson's model of atom discarded and replaced by Rutherford's model?

By 1911, evidence was collected that allowed scientists to modify Thomson's model of an atom. Ernest Rutherford (1871–1937), a former student of Thomson, performed one of the classic experiments of scientific history. As indicated earlier in this unit, he bombarded different types of matter with high-energy, positively charged alpha particles. Based on the Thomson's model of an atom, Rutherford hypothesized that the alpha particles should go through very thin metal foils undeflected. However, after performing an experiment on gold leaf, he found that a small but significant number of the alpha particles were deflected through large angles by the gold atoms. These results showed Rutherford that the Thomson model of the atom was not valid. A uniform positive charge with embedded negatively charged electrons would not interact with the alpha particles in such a manner.

The Rutherford model of the atom has a small, dense, positively charged nucleus around which electrons whirl at high speeds and at relatively long distances from it. He compared the structure of an atom with the solar system, saying that the nucleus corresponds to the sun, while the electrons correspond to the planets. This picture of the atom is also called the planetary atom. In other words, Rutherford gave us the nuclear model of the atom.

Even though Rutherford showed that the atom has a nucleus, he did not know how the electrons were arranged outside the nucleus. This was the major drawback of Rutherford's model of an atom.

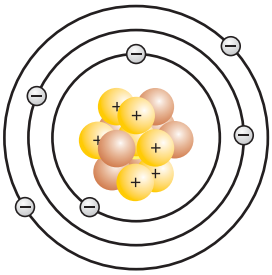
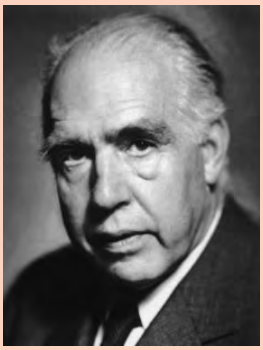


Figure 1.9 Rutherford model.

Historical Note



Niels Bohr

A Nobel Prize winner, Niels Bohr, was known not only for his own theoretical work, but also as a mentor to younger physicists who themselves made important contributions to physical theory. As the director of the Institute for Theoretical Physics at the University of Copenhagen, Bohr gathered together some of the finest minds in the physics community, including Werner Heisenberg and George Gawow. During the 1920's, the Institute was the source of many important works in quantum mechanics and in theoretical physics in general.

iv) Bohr Model

In 1913, the Danish physicist Niels Bohr (1885–1962) proposed an atomic model in which the electrons moved around the nucleus in circular paths called orbits. He assumed the electrons to be moving around the nucleus in a circular orbit as the planets move around the sun. Based on the Rutherford's atomic model, Bohr made the following modifications.

- The electrons in an atom can exist only in a restricted number of stable orbits with energy levels in which they neither absorb nor emit energy. These orbits designated by a number called the principal quantum number, n. The principal quantum number has the values of 1, 2, 3, . . . The energy levels are also called shells. They are represented by K, L, M, . . . etc. The K-shell is the first shell, the L-shell is the second shell, the M-shell is the third shell, etc.
- When an electron moves between orbits it absorbs or emits energy. When an electron jumps from lower to higher states it absorbs a fixed amount of energy. When an electron falls from a higher (excited) state to a lower (ground) state it emits a fixed amount of energy.
- The electrons move around the nucleus in energy levels.

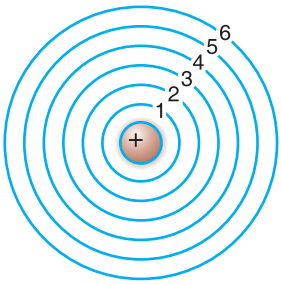


Figure 1.10 The first six energy levels in the Bohr model of an atom.

- According to Bohr, for each element, the number of energy levels or shells is fixed. Also, each different energy level or shell of an atom can only accommodate a certain number of electrons. The maximum number of electrons that the main energy level can have is given by the general formula $2n^2$, where n is equal to the main energy level.

Therefore, the maximum number of electrons in each shell in the Bohr model is:

first energy level	$(n = 1)$ is $2(1)^2$	= 2 electrons
second energy level	$(n = 2)$ is $2(2)^2$	= 8 electrons
third energy level	$(n = 3)$ is $2(3)^2$	= 18 electrons
fourth energy level	$(n = 4)$ is $2(4)^2$	= 32 electrons

It can also be summarized in the form of Table 1.4.

Table 1.4 Energy levels and electrons for the first four main energy levels.

Principal quantum number (*n*)	1	2	3	4
Shell	K	L	M	N
Maximum possible number of electrons	2	8	18	32

- When electrons fill the various energy levels in an atom, they first occupy the shell with the lowest energy level. When the lowest energy level is filled, according to the $2n^2$ rule, then the electrons enter the next higher energy level.

The outermost shell of an atom cannot accommodate more than eight electrons, even if it has the capacity to hold more electrons. This is because having more than eight electrons in the outermost shell makes the atom unstable.

For example, let us write the electron configuration of magnesium that has 12 electrons. The first 2 electrons occupy the K-shell (the first energy level). The L-shell (the second energy level) is occupied with 8 electrons. The remaining 2 electrons enter the M-shell (the third energy level). Hence, the electron configuration of magnesium atom becomes $\begin{matrix} K & L & M \\ 2 & 8 & 2 \end{matrix}$.

Bohr's diagrammatic representation, using shells, in some atoms are given below. Note that *p* = proton, *n* = neutron, and ● = electron.

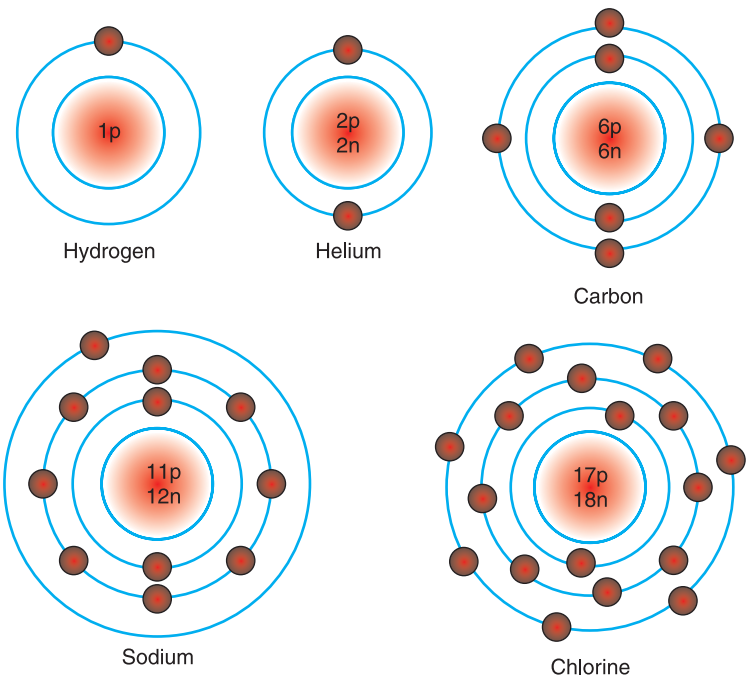


Figure 1.11 Diagrammatic representation of Bohr's model.

Activity 1.18

Form a group and try to compare Rutherford's atomic model with Bohr's atomic model? Present your discussion to the class.

Exercise 1.7

Draw Bohr's diagrammatic representations for the following atoms:

a Nitrogen (atomic number = 7)

b Sulphur (atomic number = 16)

c Potassium (atomic number = 19)

d Calcium (atomic number = 20)

v) The Quantum Mechanical Model

During the 1920's, the discovery of the wave-like properties of electrons led to the wave-mechanical model or quantum-mechanical model of atoms. In this model, the electrons are associated with definite energy levels, but their locations cannot be pinpointed. Instead, they are described in terms of the probability of being found in certain regions of space about the nucleus. These regions of space are called orbitals. An orbital is a particularly shaped volume of space where the probability of finding an electron is at a maximum.

According to this model, the electrons in an atom are in a series of energy levels or shells. Each energy level or shell is described by a number called the principal quantum number, *n*, which is related to the size of the energy level. The larger the energy value of *n*, the farther the electrons are found from the nucleus.

The quantum mechanical model introduces the concept of sublevel for each main energy level. For atoms of the known elements, there are four types of sublevels designated by the letters *s*, *p*, *d*, and *f*. The number of sublevels within each energy level or shell is equal to the numerical value of *n*.

Table 1.5 The principal quantum number specified with shell, sublevels and maximum number of electrons.

Principal quantum number (n)	Shell	Number and type of sublevels each shell ($2n^2$)	Maximum number of electrons in
1	K	1 (*s*)	2
2	L	2 (*s*, *p*)	8
3	M	3 (*s*, *p*, *d*)	18
4	N	4 (*s*, *p*, *d*, *f*)	32
5	O	5 (*s*, *p*, *d*, *f*, *g*)	50
6	P	6 (*s*, *p*, *d*, *f*, *g*, *h*)	72

Any orbital can accommodate a maximum of two electrons. Accordingly,

s-sublevel has only 1 orbital and can hold a maximum of 2 electrons,

p-sublevel has 3 orbitals and can hold a maximum of 6 electrons,

d-sublevel has 5 orbitals and can hold a maximum of 10 electrons, and

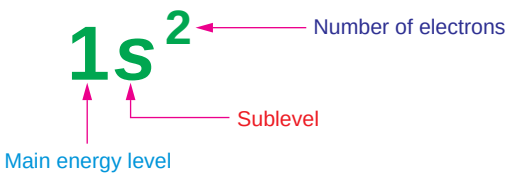
f-sublevel has 7 orbitals and can hold a maximum of 14 electrons.

Electron Configurations

The arrangement of electrons in an atom is known as the electron configuration of the atom. Because atoms of different elements have different numbers of electrons, a distinct electronic configuration exists for the atoms of each element. Like all systems in nature, electrons in atoms tend to assume arrangements that have the lowest possible energies. The lowest energy arrangement of the electrons in an atom is called the ground state electron configuration. A few simple rules, combined with the quantum number relationships discussed below, allow us to determine these ground-state electron configurations.

The quantum mechanical model is designated by the following notation: a coefficient which shows the main energy level, a letter that denotes the sublevel that an electron occupies, and a superscript that shows the number of electrons in that particular sublevel.

The designation is explained as follows



For example, the electron configuration of lithium ($^{5}_{3}Li$) is: $1s^2 2s^1$.

This indicates that there are 2 electrons in the first *s*-sublevel and 1 electron in the 2nd *s*-sublevel.

The configuration for sodium ($^{23}_{11}Na$) atom is: $1s^2 2s^2 2p^6 3s^1$. This indicate that there are 2 electrons in the first *s*-sublevel, 2 electrons in the second *s*-sublevel, 6 electrons in the second *p*-sublevel, and 1 electron in the third *s*-sublevel.

Activity 1.19

Form a group and perform the following task in groups and present your findings to the class.

The following ground state electron configurations are incorrect. Identify in each case the mistakes that have been made and write the correct electron configurations.

i) $1s^2\ 2s^2\ 2p^4\ 3s^2$

ii) $1s^2\ 2s^1\ 2p^6$

iii) $1s^2\ 2s^2\ 2p^6\ 3p^3\ 3s^2$

iv) $1s^2\ 2s^2\ 2p^6 3s^2\ 3p^3\ 4s^2\ 3d^4$

The following table lists the first 14 elements in the periodic table and shows the electron configuration of each.

Table 1.6 The ground state electron configuration of the first 14 elements.

Element	Electron Configuration
Hydrogen	$1s^1$
Helium	$1s^2$
Lithium	$1s^2\ 2s^1$
Beryllium	$1s^2\ 2s^2$
Boron	$1s^2\ 2s^2\ 2p^1$
Carbon	$1s^2\ 2s^2\ 2p^2$
Nitrogen	$1s^2\ 2s^2\ 2p^3$
Oxygen	$1s^2\ 2s^2\ 2p^4$
Fluorine	$1s^2\ 2s^2\ 2p^5$
Neon	$1s^2\ 2s^2\ 2p^6$
Sodium	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Magnesium	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Aluminium	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
Silicon	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

Exercise 1.8

Write the ground state electron configuration for the following elements.

a Phosphorus (atomic number = 15)
b Sulphur (atomic number = 16)
c Chlorine (atomic number = 17)
d Argon (atomic number = 18)

Rules Governing Electron Configuration

Most of the electronic configuration of an atom can be explained in terms of the building-up principle, or also known as the Aufbau principle. According to the Aufbau principle, an electron occupies the lowest energy orbital available before entering a higher energy orbital.

The atomic orbitals are filled in order of increasing energy. The orbital with the lowest energy is the 1*s* orbital. The 2*s* orbital is the next higher in energy, then the 2*p* orbitals, and so on. Beginning with the third main energy level, n = 3, the energies of the sublevels in the different main energy levels begin to overlap.

In Figure 1.12, for example, the 4*s* sublevel is lower in energy than the 3*d* sublevel. Therefore, the 4*s* orbital is filled before any electrons enter the 3*d* orbitals. Less energy is required for two electrons to pair up in the 4*s* orbital than for a single electron to occupy a 3*d* orbital. Which sublevel will be occupied after the 3*d* sublevel is fully occupied?

The electrons are arranged in sublevels according to Aufbau principle which is also known as the diagonal rule. The diagonal rule is a guide to the order of filling energy sublevels. It is particularly helpful for atoms with atomic numbers higher than 18. This is because for atoms with higher atomic numbers their sublevels are not regularly filled.

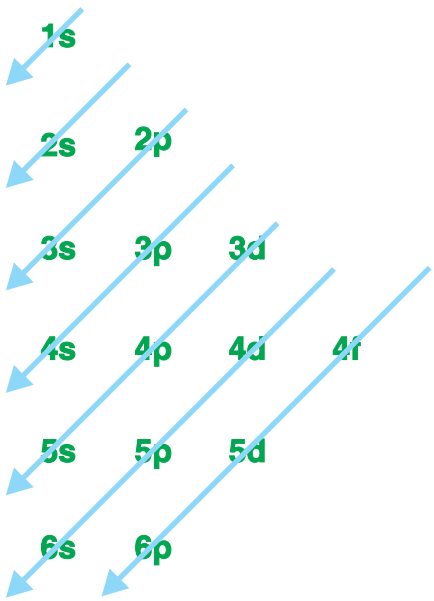


Figure 1.12 Diagonal rule for writing electron configuration.

If you list the orbitals following the direction of arrows, you can obtain the electron configuration of most atoms. Therefore, the order of filling the sublevels is given as:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6$...

The outermost occupied energy level of an atom is called the valence shell and the electrons that enter into these energy levels are referred to as valence electrons.

Example 1

What is the electron configuration of magnesium, $^{24}_{12}Mg$?

Solution:

$$^{24}_{12}Mg = \underbrace{1s^2}_{2}\ \underbrace{2s^2 2p^6}_{8}\ \underbrace{3s^2}_{2} \quad \begin{array}{l} \Rightarrow \text{in sublevel} \\ \Rightarrow \text{in main energy level} \end{array}$$

From this, we can write the configuration of elements by using a noble gas as a core and the electrons outside the core. Note that noble gases have complete outer electron shells, with 2 or 8 electrons. Therefore, the above electron configuration can be written as follows:

$$^{24}_{12}Mg = \underbrace{1s^2 2s^2 2p^6}_{[Ne]}\ 3s^2$$

$$= [Ne]\ 3s^2$$

Example 2

Write the electron configuration of copper (Z = 29).

Solution:

Following the diagonal rule, we can fill the sublevels, in order of increasing energy, until all its electrons are filled. Therefore, the configuration will be:

$$Cu = \underbrace{1s^2 2s^2 2p^6 3s^2 3p^6}_{[Ar]}\ 4s^1 3d^{10}$$

$$= [Ar]\ 4s^1 3d^{10}$$

Exercise 1.9

1. Write the electron configuration of the following elements using noble gases as a core.

   a O b Al c Cl d Ca

2. Write the electron configuration of the following elements.

a $^{84}_{36}Kr$ b $^{56}_{26}Fe$ c $^{65}_{30}Zn$

Check list

Key terms of the unit

- *Atom*
- *Atomic Model*
- *Aufbau principle*
- *Average atomic mass*
- *Bohr's Model*
- *Cathode ray*
- *Dalton's atomic theory*
- *Dalton's Model*
- *Electron*
- *Electron configuration*
- *Electron configuration*
- *Isotopes*
- *Main energy level*
- *Modern atomic theory*
- *Neutron*
- *Nucleus*
- *Proton*
- *Quantum Mechanical Model*
- *Radioactivity*
  - *α - particles*
  - *β - particles*
  - *γ - particles*
- *Rutherford's Model*
- *Sub-energy level*
- *Thomson's Model*

Unit Summary

The Greek philosopher Democritus proposed what we consider as the first atomic theory. He believed that all matter consists of very small, indivisible particles. He named the particles atoms (meaning uncuttable or indivisible).

John Dalton's work in atomic theory marked the start of modern chemistry. His concept of an atom was far more detailed and specific than Democritus'.

Dalton did not try to describe the structure or composition of atoms – he did not know what an atom really looked like. But he did realize that the different properties shown by elements such as hydrogen and oxygen can be explained by assuming that hydrogen atoms are not the same as oxygen atoms.

In our current view of atomic theory, atoms consist of smaller particles – electrons, protons, and neutrons. The structure of an atom is basically a cloud of electrons

surrounding a small, dense region, called the nucleus. An electron is a negatively charged particle with negligible mass.

The nucleus was discovered by Ernest Rutherford. Rutherford suggested that the charge and mass of an atom are concentrated in its nucleus, and that the nucleus has positive charge.

In current views, the nucleus of an atom contains protons and neutrons. Protons have positive charge, and neutrons have no charge. Each element has a unique number of protons in its atoms. This number is the atomic number of the element, and is designated by Z. The total number of protons and neutrons in the nucleus of an atom is the mass number of the element. Mass number is designated by *A*. Atomic number and mass number for an atom of an element, X, is designated by:

$$^{A}_{Z}\mathbf{X}$$

For a neutral atom,

- the number of protons = the number of electrons = atomic number.
- the number of protons + the number of neutrons = mass number.

Over centuries, scientists have developed these main atomic models:

- Dalton's Model: an atom is a solid, indestructible sphere.
- Thomson's Model: an atom is a sphere with a uniform positive charge. Negatively charged electrons are embedded in the positive charge.
- Rutherford's Model: Rutherford discovered that atoms are mostly empty space with a dense, positive nucleus.
- Bohr's Model: in an atom, electrons follow circular orbits around the nucleus. The Bohr Model was the first quantum mechanical model of the atom.
- The Quantum Mechanical Model: in an atom, the orbits of electrons around the nucleus are elliptical. This theory is a modification of the Bohr model.

The quantum mechanical description of an atom is the electron configuration of the atom – the arrangement of electrons in the atom. The maximum number of electrons in a main energy level is $2n^2$, where n is the principal quantum number.

The quantum mechanical model is governed by this principle:

The Aufbau principle is a guide to the order of filling energy levels and sublevels. According to this principle lower energy sub levels are completely filled before electrons enter higher energy sublevels.

REVIEW EXERCISE ON UNIT 1

Part I: True/False type questions

1. Identify whether each of the following statement is correct or incorrect. Give your reasons for considering a statement to be false.
   - a Atoms of the same element have the same number of protons.
   - b Atoms of different elements can have the same number of protons.
   - c Atoms of the same element have the same number of electrons.
   - d Atoms of the same element have the same number of neutrons.
   - e The *s*- sublevel contains 2 electrons.
   - f The *d*-sublevel contains 5 orbitals.
   - g The 4*s*-sublevel comes before the 3*d* sublevel.
   - h The fundamental particle not present in a hydrogen atom is the proton.
   - i Isotopic elements contain the same number of neutrons.
   - j An alpha particle is composed of He^{2+} ions.
   - k The mass of an electron is almost equal to the mass of a proton.

Part II: Write the missing words in your exercise book

2. The mass of an atom is mainly concentrated in _____.
3. The diameter of an atom is _____ times that of the nucleus.
4. The energy level in the third principal quantum number can accommodate _____electrons.

Part III: Give the appropriate answers for the following questions

5. Draw diagrammatic representation of the following atoms by using shell or energy level.

   a $^{9}_{4}Be$ b $^{19}_{9}F$ c $^{28}_{14}Si$ d $^{39}_{19}K$

6. Write the electron configuration of the elements: (use quantum mechanical model)

   a $^{32}_{16}S$ b $^{52}_{24}Cr$ c $^{56}_{26}Fe$ d $^{64}_{29}Cu$

7. For the elements in question number 6,
   - a give the valence electrons for each.
   - b give the number of neutrons for each.

8. In a certain atom, the last electron has a $3d^2$ electron configuration. On the basis of this:
   a write the electron configuration of the element using a noble gas as a core.
   b write the electron configuration using the main energy levels.
   c what is the atomic number of the atom?
   d what are the valence shell electrons of the atom?
   e how many protons does the atom have?
9. What is the relative atomic mass of an element whose isotopic composition is 90% of ^{20}X and 10% of ^{22}X?

Part IV: Short answer type questions

10. Comment on Dalton's atomic theory and the areas in which modifications were made.
11. John Dalton made a number of statements about atoms that are now known to be incorrect. Why do you think his atomic theory is still found in science text books?
12. If scientists had tried to repeat Thomson's experiment and found that they could not, would Thomson's conclusion still have been valid? Explain your answer.