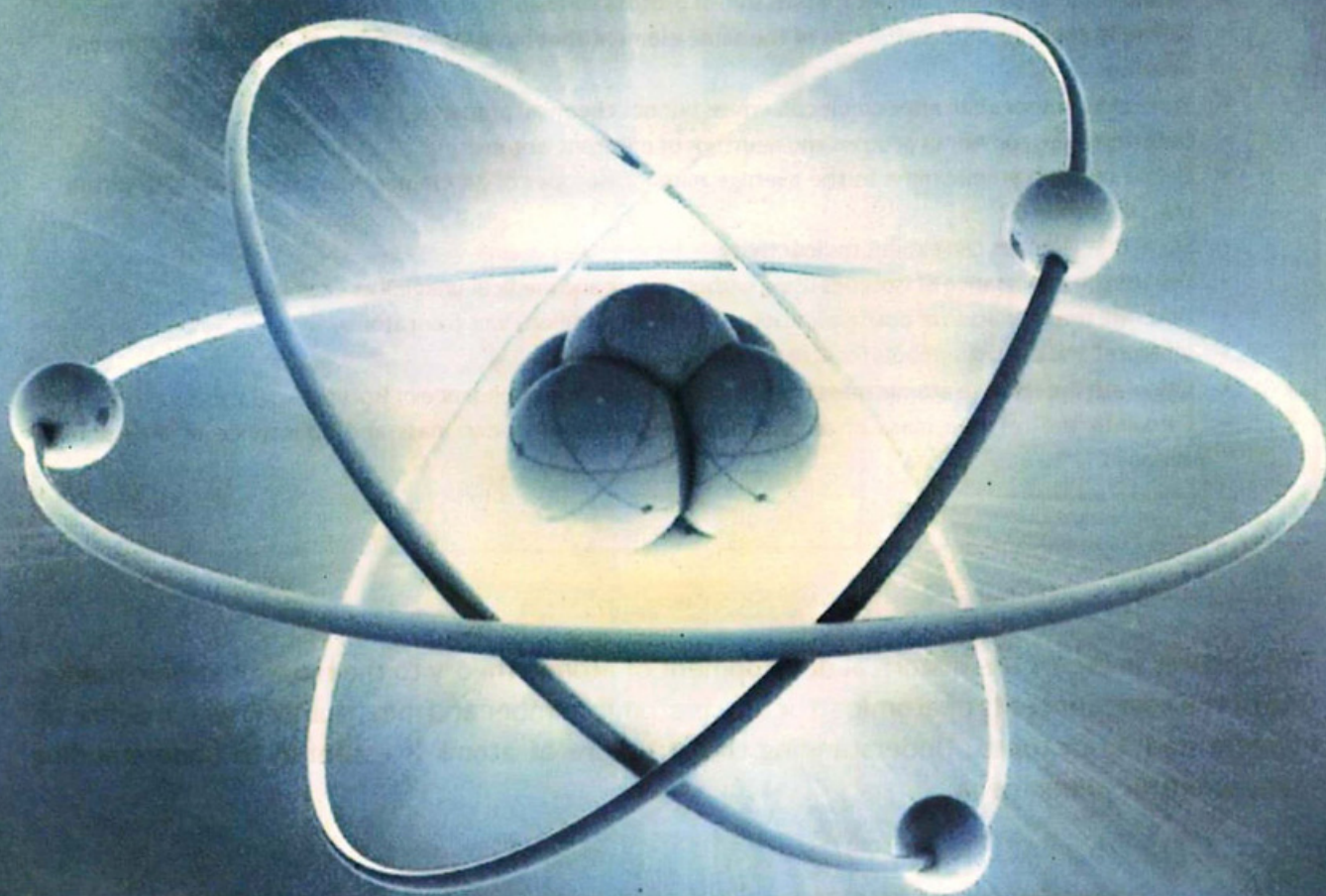




ATOMIC STRUCTURE

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Explain the structure of the atom as the central nucleus containing neutrons and protons surrounded by electrons in shells.
- State that, orbits(shells) are energy levels of electrons and a larger shell implies higher energy and greater average distance from nucleus.
- State the electrons are quantum particles with probabilistic paths whose exact paths and location cannot be mapped (with reference to uncertainty principle)
- Explain that nucleus is made up of protons and neutrons held together by strong nuclear force.
- Explain that an atomic model is an aid to understand the structure of an atom.
- State the relative charge and relative masses of a subatomic particles (an electron, proton, and neutron).
- Interpret the relationship between a subatomic particle, their mass, and charge.

- Illustrate the path that positively and negatively charged particles would take under the influence of a uniform electric field.
- Define proton number / atomic number as the number of protons in the nucleus of an atom.
- Explain that the proton number is unique to each element and use to arrange elements in periodic table.
- State that radioactivity can change the proton number and alter an atom's identity.
- Define nucleon number / atomic mass as sum of protons and neutrons in the nucleus of an atom.
- Define isotopes as different atoms of the same element that have same number of protons but different neutrons.
- State that isotopes can affect molecular mass but not chemical properties of an atom.
- Determine the number of protons and neutrons of different isotopes.
- Define relative atomic mass as the average mass of isotopes of an element compared to $1/12$ th of the mass of carbon-12.
- State that isotopes can exhibit radioactivity.
- Discuss the importance of isotopes using carbon dating and medical imaging as examples.
- Describe the formation of positive (cation) and negative (anion) ions from atoms.
- Interpret and use the symbols for atoms and ions.
- Calculate the relative atomic mass from relative masses and abundance of isotopes.
- Calculate the relative mass of an isotope given relative atomic mass and abundance of all stable isotopes.

INTRODUCTION

This chapter presents the historical development of atomic theory to the modern atomic model. One of the basic concepts of atomic structure is atomic number and mass number, which define an element and its isotopes. Understanding the structure of atoms is essential to understanding many scientific phenomena.

3.1 ATOMIC MODELS

The concept of the atomic model evolved over time as our understanding of atomic structure deepened through experimental observations and theoretical advances. Several important models of the atom had been proposed throughout history, each contributing to the understanding of atomic behaviour and properties. The most important atomic models are:

Dalton's model

In 1803, the British chemist John Dalton presented a scientific theory on the existence and nature of matter. This theory is called Dalton's atomic theory. Main postulates of his theory are as follows:

1. All elements are composed of tiny indivisible particles called atoms.
2. Atoms of a particular element are identical. They have same mass and same volume.
3. During chemical reaction atoms combine or separate or re-arrange. They combine in simple ratios.
4. Atoms can neither be created nor destroyed.

Dalton was able to explain quantitative results that scientists of his time had obtained in their experiments. He nicely explained the laws of chemical combinations. His brilliant work became

the main stimulus for the rapid progress of the chemistry during nineteenth century. However, series of experiments that were performed in 1850's and beginning of twentieth century clearly demonstrated that atom is divisible and consists of subatomic particles, electrons, protons and neutrons.

In 1911 Rutherford performed an experiment in order to know the arrangement of electrons and protons in atoms.

Rutherford's Experiment

Rutherford bombarded a very thin gold foil about 0.00004cm thickness with α -particles. (figure 3.1). He used α -particles obtained from the disintegration of polonium. α -particles are helium nuclei that are doubly positively charged (He^{++}). Most of these particles passed straight through the foil. Only few particles were slightly deflected. But one in 1 million was deflected through an angle greater than 90° from their straight paths. Rutherford performed a series of experiments using thin foils of other elements. He observed similar results from these experiments.

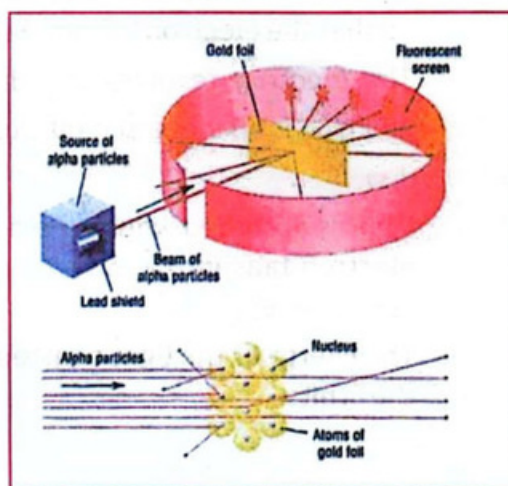


Figure 3.1 Rutherford's Experiment

Rutherford made the following conclusions

1. Since majority of the α -particles passed through the foil undeflected, most of the space occupied by an atom must be empty.
2. The deflection of a few α -particles through angles greater than 90° shows that these particles are deflected by electrostatic repulsion between the positively charged α -particles and the positively charged part of atom.
3. Massive α -particles are not deflected by electrons.

On the basis of conclusions drawn from these experiments, Rutherford proposed a new model for an atom. He proposed a planetary model (similar to the solar system) for an atom. An atom is a neutral particle. The mass of an atom is concentrated in a very small dense positively charged region. He named this region as nucleus. The electrons are revolving around the nucleus in circles. These circles are called orbits. The centrifugal force due to the revolution of electrons balances the electrostatic force of attraction between the nucleus and the electrons.

Defects in Rutherford's Atomic Model

Rutherford's model of an atom resembles our solar system. It has following defects:

1. Classical physics suggests that electron being charged particle will emit energy continuously while revolving around the nucleus. Thus the orbit of the revolving electron becomes smaller and smaller until it would fall into the nucleus. This would collapse the atomic structure.
2. If revolving electron emits energy continuously it should form a continuous spectrum.

Bohr's Atomic Theory

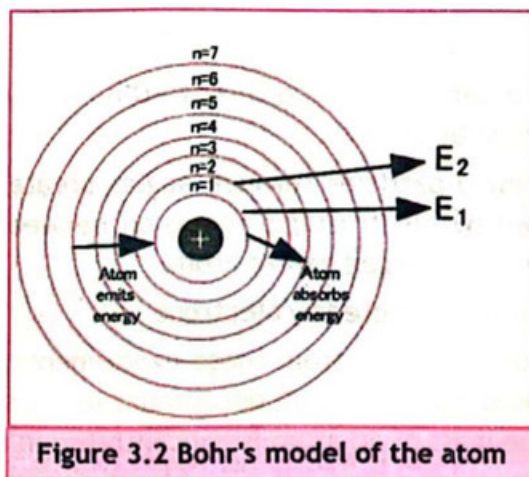
In 1913 Neil Bohr, proposed a model for an atom that was consistent with Rutherford's model. But

it also explains the observed line spectrum of the hydrogen atom. Main postulates of Bohr's atomic theory are as follows:

1. The electron in an atom revolves around the nucleus in one of the circular orbits. Each orbit has a fixed energy. So each orbit is also called energy level.
2. The energy of the electron in an orbit is proportional to its distance from the nucleus. The farther the electron is from the nucleus, the more energy it has.
3. The electron revolves only in those orbits for which the angular momentum of the electron is an integral multiple of $\frac{h}{2\pi}$ where h is Plank's constant (its value is 6.626×10^{-34} J.s).
4. Light is absorbed when an electron jumps to a higher energy orbit and emitted when an electron falls into a lower energy orbit. Electron present in a particular orbit does not radiate energy.
5. The energy of the light emitted is exactly equal to the difference between the energies of the orbits.

$$\Delta E = E_2 - E_1$$

Where ΔE is the energy difference between any two orbits with energies E_1 and E_2



Bohr model does not depict the three dimensional aspect of an atom.

Quantum Mechanical Model

This is the current model used by modern science to describe the structure of the atom. It incorporates the principles of quantum mechanics and treats electrons as wave-particle entities. Instead of exact orbits, it defines probability regions, called orbitals, where electrons are likely to be found.

The Heisenberg Uncertainty Principle

Heisenberg uncertainty principle is one of the fundamental concepts of quantum mechanics and is named after the German physicist Werner Heisenberg, who formulated it in 1927.

This principle states that it is impossible to simultaneously determine the exact location and future trajectory of an electron. As a result, plotting the electron orbit around the nucleus

becomes an irresistible challenge.

Imagine that you have a single hydrogen atom and you decide to observe the position of that single electron at a given moment. Shortly after you repeated this process, the electron moved to another position. This means that from the original location to the next one is completely unknown to you. Continuous repetition of this process allows the gradual construction of a three-dimensional map representing the likely locations where the electron is expected to exist. You cannot know for sure where an electron is and where it goes next. This makes it impossible to draw the orbit of the electron around the nucleus.

In hydrogen, the electron has the potential to exist anywhere in the spherical region surrounding the nucleus. 95% (or whatever you want) of the time, the electron will be in a relatively simple region of space close to the nucleus, called an orbital. An orbital is the region of space where the electron lives.

Louis de Broglie, a French physicist, in 1924 proposed dual nature of electrons. He suggested that sub-atomic particles like electrons, can exhibit both particle-like and wave-like behaviour. His idea opened the door for new possibilities in understanding behaviour of sub-atomic particles. This concept made a significant contribution to the development of quantum mechanics.

In 1927, Davisson and Germer, experimentally confirmed the de Broglie hypothesis that electron has wave like behaviour. This discovery laid the foundation for the Modern Quantum Mechanics.

Understanding Atomic Models

An atomic model is a tool for understanding the structure and behavior of atoms and their interactions in chemical reactions. Any atomic model helps us understand the structure of an atom. An atomic model is not a physical model, but represents a conceptual imagination. This helps to explain experimental observations of atomic behavior. The atomic model gives us a simplified representation of complex reality. As research and technology progress, scientists continue to improve our knowledge and atomic models.

A simple view of the structure of an atom

The nucleus of an atom is in the center. It contains protons and neutrons. Protons and neutrons are collectively called nucleons. The nucleus is surrounded by electrons in shells. Protons and neutrons are massive particles. The mass of electrons is so small. So, in practice, the mass of an atom is concentrated in the nucleus.

Nuclear Force

The nucleus contains protons and neutrons. Protons are positively charged and neutrons are neutral. The nucleus has no negative charge. The positively charged protons must cancel each other out and the nucleus must break apart. But atoms are stable and have existed for billions of years. Therefore, there must be some kind of attraction that connects them. No electrostatic or magnetic forces occur within the core. This is because these forces involve both attraction and repulsion. Therefore, the force that binds protons and neutrons together is a strong force. This force is called strong nuclear force. This is defined as the strong attractive force that binds protons and neutrons together. This force is stronger than electrostatic or magnetic forces. This force exists between neutrons and neutrons, protons and protons, and neutrons and protons.

3.2 SUBATOMIC PARTICLES

Subatomic particles are the fundamental particles that make up atoms. The three main subatomic particles are:

Proton

- Relative charge: +1
- Relative mass: Approximately 1 atomic mass unit (amu) or 1.6726×10^{-27} kg

Neutron

- Relative charge: 0 (neutral)
- Relative mass: Approximately 1 atomic mass unit (amu) or 1.6749×10^{-27} kg

Electron

- Relative charge: -1
- Relative mass: Approximately $1/1836$ amu or 9.11×10^{-31} kg

Protons and neutrons are found in the nucleus of an atom, whereas electrons orbit around the nucleus in energy levels or shells. They play crucial roles in determining the properties and behaviour of atoms and molecules. Neutrons and protons are held together in the nucleus by a strong nuclear force. This force exists between neutron-neutron, proton-proton, and neutron-proton.

Relationships between subatomic particles

Protons and neutrons have roughly the same mass, around 1 amu. This mass contributes significantly to the total mass of the atom. Electrons have much less mass, so their contribution to the total mass of an atom is usually negligible.

The interaction between the negatively charged electrons and positively charged protons in the nucleus is what holds the atoms together.

✓ The behavior of protons, neutrons and electrons in an electric field

What happens when a beam of these particles passes between two electrically charged plates?

figure 3.3

- Protons are positively charged and are deflected on a curved path toward the negative plate.
- Electrons are negatively charged and are deflected on a curved path toward the positive plate.
- Neutrons have no charge, go straight ahead.
- If the electrons and protons are traveling at the same speed, the electrons being lighter are deflected far more strongly than the heavier protons.

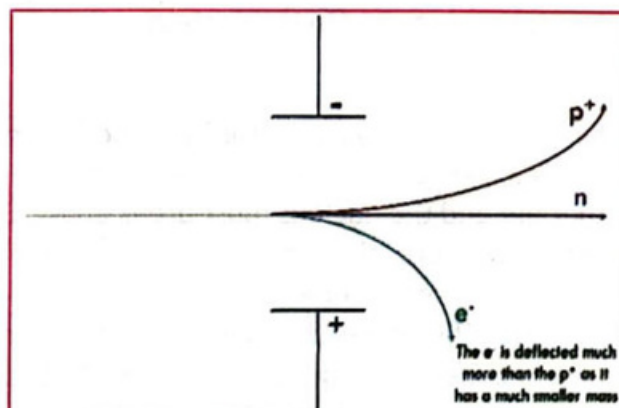


Figure 3.3: Path of positively and negatively charged particles through the uniform electric field.

Charge Neutrality

Atoms are electrically neutral because the number of protons (positively charged) in the nucleus is equal to the number of electrons (negatively charged) in the electron cloud. The charges balance each other so there is no net charge on the atoms.

Radioisotopes

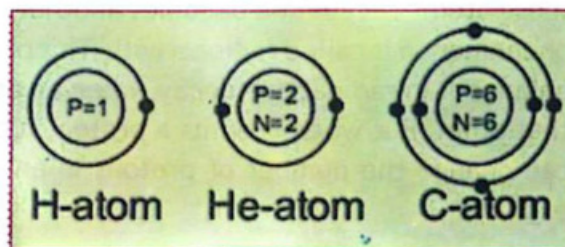
Different isotopes of the same element have the same number of protons in their atomic nuclei but differing numbers of neutrons. Some isotopes of an element are unstable and show radioactive decay. Radioactive isotopes of an element can be defined as atoms that contain an unstable combination of neutrons and protons, or excess energy in their nucleus. For example, hydrogen-3 (tritium), carbon-14, uranium-238 etc.

3.3 PROTON OR ATOMIC NUMBER

What determines the identity of an element?

Proton number refers to the number of protons in the nucleus of an atom. It is also known as the atomic number and is indicated by the symbol "Z".

Protons have a positive electrical charge. In neutral atoms, the number of protons is equal to the number of electrons. This balances the positive charge of the protons. This means that the proton number also indicates the number of electrons in the atom. For example, there is only one proton in the nucleus of a H atom; therefore its atomic number is 1. All the atoms of a given element have the same number of protons and therefore the same atomic number.



Do you think atomic number of He is 2? What is the proton number of C-atom?

Uniqueness of proton number

Each element has a unique proton number that distinguishes it from other elements. It determines the various properties of an element and its position in the periodic table. In the periodic table, elements are arranged based on their atomic or proton number. Therefore, the number of protons is related to the position of the element in the periodic table. Thus, the number of protons determines each particular element. This will tell you what element you are talking about.

For example, if an atom has a proton number of 6, it must be carbon. If an atom has 11 protons, it must be sodium. Similarly, each nitrogen atom has 7 protons, each oxygen atom has 8 protons, etc. You can identify each atom by the number of protons.

Nucleon number or Atomic mass

The total number of protons and neutrons in an atom is known as its mass number or nucleon number.

Some atoms of an element have different number of neutrons, such atoms are called isotopes.

No. of neutrons = mass number – atomic number

Example 3.1: Determining the number of protons and neutrons in an atom

Atomic number of an element is 17 and mass number is 35. How many protons and neutrons are in the nucleus of an atom of this element?

Problem Solving Strategy:

Number of protons are equal to atomic number and

Number of neutrons = mass number - atomic number

Solution:

Number of protons = atomic number = 17

Number of neutrons = mass number - atomic number
= $35 - 17 = 18$

Radioactivity

The proton number determines the identity of the element. In stable elements, the nuclear force is balanced. In some elements, the nuclear forces are not naturally balanced. The nucleus of these atoms decays and becomes another atom. This process is called radioactive decay and this phenomenon is called radioactivity. This process continues until the forces in the nuclear core are balanced. In radioactive decay, when an atom emits a neutron, it changes to another isotope of that atom. But when it emits a proton, it becomes another atom. This means that radioactivity can change the number of protons in an atom and thus change the identity of the atom. For example;

1. Carbon-14 is a radioactive isotope of carbon. It is naturally present in the atmosphere. When any living organism takes in carbon dioxide from the air, it incorporates both C-14 and C-12 atoms into its tissues. The radioactive C-14 undergoes radioactive decay, transforms into nitrogen-14.
2. Uranium-238 is a radioactive isotope of uranium. It decays over time and finally transforms into stable lead-206 atom.

3.4 RELATIVE ATOMIC MASS AND ATOMIC MASS UNIT

The first quantitative information about atomic masses came from the work of Dalton, Gay Lussac, Lavoisier, Avogadro and Berzelius. By observing the proportions in which elements combine to form various compounds, nineteenth century chemists calculated relative atomic masses. An atom is extremely small particle, therefore, we cannot determine the mass of a single atom. However, it is possible to determine the mass of one atom of an element relative to another experimentally. This can be done by assigning a value to the mass of one atom of a given element, so that it can be used as standard. By international agreement in 1961, light isotope of carbon C-12 has been chosen as a standard. This isotope of carbon (C-12) has been assigned a mass of exactly 12 atomic mass unit. This value has been determined accurately using mass spectrometer. The mass of atoms of all other elements are compared to the mass C-12. Thus **“the mass of an atom of an element relative to the mass of an atom of C-12 is called its relative atomic mass”**.

One atomic mass unit (amu) is defined as a mass exactly equal to one-twelfth the mass of one C-12 atom.

Mass of one C-12 atom = 12 amu

$$1 \text{ amu} = \frac{\text{mass of one C-12 atom}}{12}$$

A hydrogen atom is 8.40% as massive as the standard C-12 atom. Therefore, relative atomic mass of hydrogen.

$$= \frac{8.40}{100} \times 12 \text{ amu}$$

$$= 1.008 \text{ amu}$$

Similarly, relative atomic masses of O, Na, Al are 15.9994 amu, 22.9898 amu, 26.9815 amu respectively. Table 3.1 shows the relative atomic masses of some elements.

Table 3.1: relative atomic masses of some elements

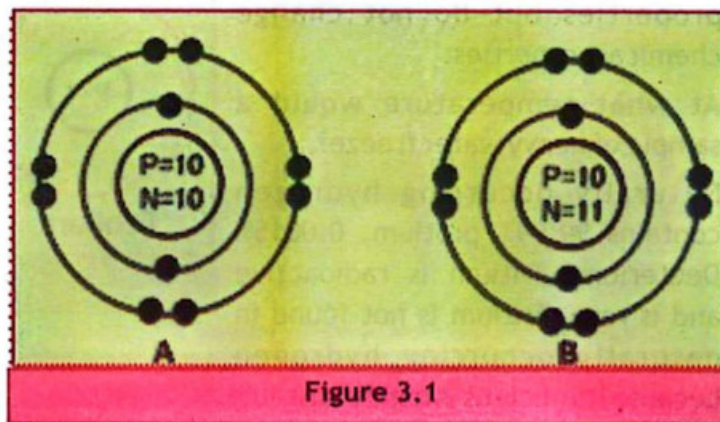
Element	Relative atomic mass	Element	Relative atomic mass
H	1.008 amu	Al	26.9815 amu
N	14.0067 amu	S	32.06 amu
O	15.9994 amu	Cl	35.453 amu
Na	22.9898 amu	Fe	55.847 amu

3.5 ISOTOPES

Figure 3.1 shows Bohr's Model for two atoms A and B. Can you identify three similarities and two differences in these atoms?

You will find,

- Both the atoms have same number of protons.
- Both the atoms have same number of electrons.
- Both have same atomic number.
- Both have different number of neutrons.
- Both differ in total number of protons and neutron. This means they have different mass numbers.



Since both the atoms have the same atomic number, they must be the atoms of same element and are called isotopes. The word isotope was first used by Soddy. It is a Greek word "isos" means same and "tope" means place.

Isotopes are atoms of an element whose nuclei have the same atomic number but different mass number. This is because atoms of an element can differ in the number of neutrons. Isotopes are chemically alike and differ in their physical properties.

How does the discovery of isotopes contradicted Dalton's atomic theory?

3.5.1 Isotopes of Hydrogen

Hydrogen has three isotopes. Hydrogen -1 (Protium) has no neutron. Almost all the hydrogen is Hydrogen -1. Its symbol is ${}^1\text{H}$. Hydrogen -2 (deuterium) has one neutron and hydrogen -3 (Tritium) has two neutrons. Their symbols are ${}^2\text{H}$ and ${}^3\text{H}$ respectively. Because hydrogen -1 also known as protium has only one proton, adding a neutron doubles its mass.

Protium / Hydrogen is a colourless, odourless, and tasteless gas. It is insoluble in water and is highly inflammable gas. Water that contains hydrogen-2 atoms in place of hydrogen-1 is called heavy water. Table 3.2 Shows some physical properties of ordinary water and heavy water.

Table 3.2 - Comparison of ordinary water and heavy water.

Property	Ordinary water	Heavy water
Melting Point	0.00°C	3.81°C
Boiling point	100°C	101.2°C
Density at 25°C	0.99701 g/cm ³	1.1044 g/cm ³

Isotopes affect molecular mass of a substance, can change physical properties but do not change chemical properties.

At what temperature would a sample of heavy water freeze?

Naturally occurring hydrogen contains 99.99% protium, 0.0015% Deuterium. Tritium is radioactive and is rare. Tritium is not found in naturally occurring hydrogen because its nucleus is highly unstable.

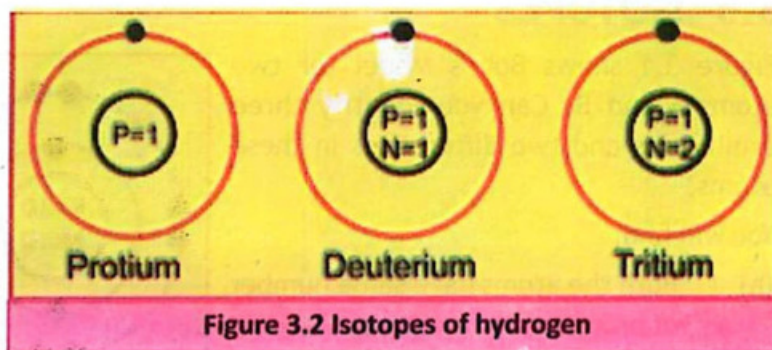


Figure 3.2 Isotopes of hydrogen

3.5.2 Isotopes of Carbon

Carbon has three isotopes. Carbon-12, carbon-13 and carbon -14. Almost all the carbon is carbon-12. Its symbol is ${}^{12}\text{C}$. It has six neutrons and six protons. Carbon-13 has symbol ${}^{13}\text{C}$. It has seven neutrons and six protons. Carbon-14 has eight neutrons and six protons. Its symbol is ${}^{14}\text{C}$. Different forms of carbon are black or greyish black solids except diamond. They are odourless and tasteless. They have high melting and boiling points and are insoluble in water.

Activity 3.1

Carbon has three isotopes ${}^{12}_6\text{C}$, ${}^{13}_6\text{C}$, ${}^{14}_6\text{C}$ Figure 3.3 shows incomplete structure of isotopes of carbon. Can you complete it?

Natural abundance of isotopes of carbon is as follows



Figure 3.3 Isotopes of Carbon

$${}^{12}_6\text{C} = 98.8\%, \quad {}^{13}_6\text{C} = 1.1\%, \quad {}^{14}_6\text{C} = 0.009\%$$

3.5.3 Isotopes of Chlorine

There are two natural isotopes of chlorine, chlorine-35 and chlorine-37. An atom of chlorine-35 has 17 protons and 18 neutrons. An atom of chlorine-37 has 17 protons and 20 neutrons. Chlorine-35 occurs in nature about 75% and chlorine-37 about 25%. Chlorine is a greyish yellow gas with sharp pungent irritating smell. It is fairly soluble in water

Activity 3.2

Chlorine has two isotopes. Figure 3.4 shows the structure of isotopes of chlorine. Can you write isotope symbol for each?

Isotope symbols:

Natural abundance 75.77%

24.23%

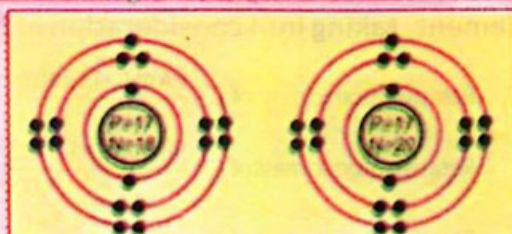
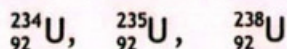


Figure 3.4 Isotopes of chlorine

3.5.4 Isotopes of Uranium

Activity 3.3

Uranium has three isotopes with mass number 234, 235 and 238 respectively.



The ${}^{235}_{92}\text{U}$ isotope is used in nuclear reactors and atomic bombs, whereas the ${}^{238}_{92}\text{U}$ isotope lacks the properties necessary for these applications. ${}^{234}_{92}\text{U}$ is rare. Natural abundance of Uranium isotopes is as follows

$${}^{234}_{92}\text{U} = 0.006\%, \quad {}^{235}_{92}\text{U} = 0.72\%, \quad {}^{238}_{92}\text{U} = 99.27\%$$

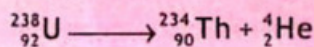
Fill in the blanks?

${}^{234}_{92}\text{U}$ has ___ protons, ___ electrons and ___ neutrons

${}^{235}_{92}\text{U}$ has ___ protons, ___ electrons and ___ neutrons

${}^{238}_{92}\text{U}$ has ___ protons, ___ electrons and ___ neutrons

When uranium-238 decays into thorium-234, it emits alpha particle. An alpha particle is doubly positively charged helium nucleus.



The fission of uranium-235 yields smaller nuclei, neutron and energy. The nuclear energy released by the fission of one kilogram of uranium-235 is equivalent to chemical energy produced by burning more than 17000 kg of coal.

Chemical properties of an element depend upon the number of protons and electrons. Neutrons do not take part in ordinary chemical reactions. Therefore, isotopes of an element have similar chemical properties.

3.5.5 Determination of Relative Atomic Mass

The relative atomic mass of an element can be calculated from the relative masses of its isotopes and their relative abundance.

Natural abundance of isotopes of carbon is as follows

$${}^1_6\text{C} = 98.8\%, \quad {}^{13}_6\text{C} = 1.1\%, \quad {}^{14}_6\text{C} = 0.009\%$$

Calculate relative atomic mass of carbon.

Solution:

The relative atomic mass is a weighed average of the all the naturally occurring isotopes of an element, taking into consideration of their natural abundance. Use general formula

$$\text{Relative atomic mass of C} = \frac{\text{RA of C-12} \times \text{at. mass of C-12} + \text{RA of C-13} \times \text{at. mass of C-13} + \text{RA of C-14} \times \text{at. mass of C-14}}{100}$$

$$\text{Relative atomic mass of C} = \frac{98.8 \times 12 + 1.1 \times 13 + 0.009 \times 14}{100}$$

$$\text{Relative atomic mass of C} = \frac{1185.6 + 14.3 + 0.126}{100}$$

$$\text{Relative atomic mass of C} = 12.00026 \text{ amu}$$

CONCEPT ASSESSMENT EXERCISE 3.1

An element has two isotopes A and B.

The relative atomic mass of element is 35.5 amu. Relative abundance of isotope A is 75.77 % and its isotopic mass is 35. Find the isotopic mass of B if its relative abundance is 24.23 %.

3.5.6 Uses of Isotopes

Stable and radioactive isotopes have many applications in science and medicines. Some of these are as follows:

- (i) Radioactive iodine -131 is used as a tracer in diagnosing thyroid problem.
- (ii) Na-24 is used to trace the flow of blood and detect possible constrictions or obstructions in the circulatory system.
- (iii) Iodine-123 is used to image the brain.
- (iv) Cobalt-60 is commonly used to irradiate cancer cells in the hope of killing or shrinking the tumors.
- (v) Carbon-14 is used to trace the path of carbon in photosynthesis. Radioactive

isotopes are used to determine the molecular structure e.g. sulphur-35 has been used in the structure determination of thiosulphate, $\text{S}_2\text{O}_3^{2-}$ ion.

- (vi) Radioactive isotopes are also used to study the mechanism of chemical reactions.
- (vii) Radioactive isotopes are used to date rocks, soils, archaeological objects, and mummies.

3.5.7 Carbon Dating

Carbon-14 is used to estimate the age of carbon-containing substances. Carbon atoms circulate between the oceans, and living organism at a rate very much faster than they decay. As a result the concentration of C-14 in all living things, keep on increasing. After death organisms no longer pick up C-14. By comparing the activity of a sample of skull or jaw bones, with the activity of living tissues, we can estimate how long it has been since the organism died. This process is called dating.

3.6 CATIONS AND ANIONS:

Cations:

Cations are positively charged ions that form when an atom loses one or more electrons. Cations are usually formed from metal atoms that tend to lose electrons to achieve a stable electron configuration similar to a noble gas. When an atom loses one or more electrons, it forms a cation. The resulting cation has the electronic configuration of a noble gas. Neutral atoms have equal number of protons and electrons. When an atom loses one or more electrons, the number of protons becomes greater than electrons, as a result atom acquires positive charge.

Example 3.1: Describing the formation of cations

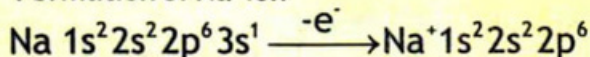
Describe the formation of Na^+ and Mg^{2+} cations.

Problem Solving Strategy:

1. Sodium belongs to Group IA on the periodic table. It has only one electron in the valence shell. Sodium atom loses its valence electron and is left with an octet. Represent this by drawing the complete electronic configuration or using an electron dot structure.
2. Magnesium belongs to Group IIA in the periodic table. It has two valence electrons. Magnesium atom loses these electrons to achieve noble gas configuration. Represent this by drawing the complete electronic configuration or using an electron dot structure. This number also corresponds to the Group number in the periodic table.

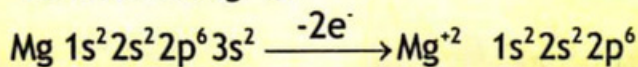
Solution:

(a) Formation of Na^+ ion

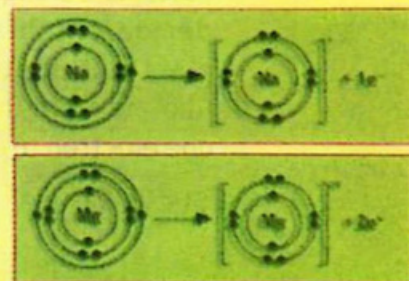


You can also represent this by following electron dot structure,

(b) Formation of Mg^{2+} ion



You can also represent this by electron dot structure,



CONCEPT ASSESSMENT EXERCISE 3.2

Describe the formation of cations for the following metal atoms:

- (a) Li (atomic no 3)
(b) Al (atomic no.13)

Anions

Anions are negatively charged ions that form when an atom gains one or more electrons. This process usually occurs when an atom has a relatively high electron affinity, meaning that it can easily attract and capture more electrons to achieve a stable electron configuration similar to a noble gas. When an atom gains one or more electrons, the number of electrons becomes greater than protons, so it acquires negative charge.

Example 3.1: Describing the formation of anions.

Describe the formation of anions for the following non-metal atoms:

- (a) Oxygen (atomic no.8) (b) Fluorine (atomic no. 9)

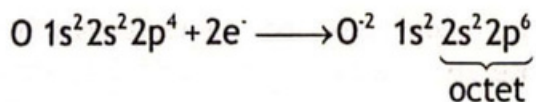
Problem Solving Strategy:

1. Write electronic configuration or dot structure.
2. Find the number of electrons needed to acquire eight electron configuration.
3. Represent addition of electrons.

Solution:

- (a) Formation of anion by oxygen atom.

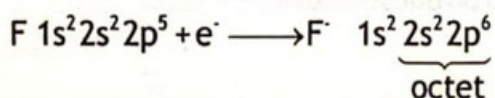
Oxygen belongs to Group VIA on the periodic table. So it has six electrons in its valence shell. It needs two electrons to achieve noble gas configuration.



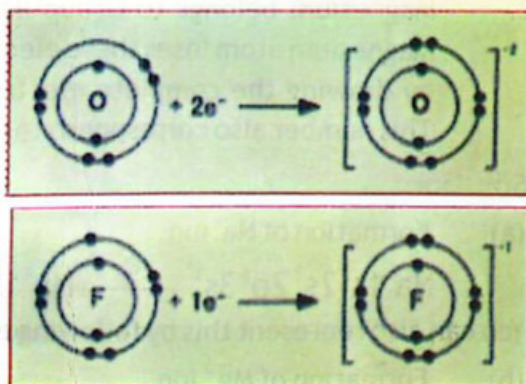
You can also represent this by electron dot structure,

- (b) Formation of anion by fluorine atom

Fluorine belongs to Group VIIA on the periodic table. So it has seven electrons in the valence shell. A fluorine atom therefore, requires only one electron to complete octet.



You can also represent this by electron dot structure,



CONCEPT ASSESSMENT EXERCISE 3.3

Describe the formation of anions by the following non-metals.

- (a) Sulphur (atomic No. 16)
- (b) Chlorine (atomic No. 17)

3.7 ELECTRONIC CONFIGURATION

To understand electronic configuration, you should know about shells and sub-shells.

Shells

According to Bohr's atomic theory, the electron in an atom revolves around the nucleus in one of the circular paths called shells or orbits. Each shell has a fixed energy. So each shell is also called energy level. Each shell is described by an n value. n can have values 1, 2, 3,

When,

$n = 1$, it is K shell

$n = 2$, it is L shell

$n = 3$, it is M shell etc.

As the value of n increases the distance of electron from the nucleus and energy of the shell increases.

Sub-Shells

A shell or energy level is sub divided into sub-shells or sub-energy levels. n value of a shell is placed before the symbol for a sub-shell. For instance,

$n = 1$, for K shell. It has only one sub-shell which is represented by 1s.

For L shell $n = 2$, L shell has two sub-shells, these are designated as 2s and 2p.

For M shell $n = 3$ So M shell has 3 sub-shells called 3s, 3p and 3d. While N shell has 4s, 4p, 4d and 4f sub-shells.

s sub-shell can accommodate maximum 2 electrons.

p sub-shell can accommodate maximum 6 electrons.

d sub-shell can accommodate maximum 10 electrons.

f sub-shell can accommodate maximum 14 electrons.

The increasing order of energy of the sub-shells belonging to different shells is given below.

$$1s < 2s < 2p < 3s < 3p < 4s < 3d \dots$$

The arrangement of electrons in sub-shells is called as the electronic configuration. We can fill the electrons present in various elements by using Auf Bau Principle. According to this principle, electrons fill the lowest energy sub-shell that is available first. This means electron will fill first 1s, then 2s, then 2p and so on.

Symbols for atoms and ions

The symbol for an atom represent the element . It consists of one or two-letters, the mass number as a left superscript, the atomic number as a left subscript, and the charge as a right superscript. For example;



This number is often omitted. This diagram shows symbol for magnesium “Mg” which stands for magnesium. The number to the upper left of the symbol is the mass number, which is 24. The number to the upper right of the symbol is the charge which is positive 2. The number to the lower left of the symbol is the atomic number which is 12.

KEY POINTS

- Rutherford proposed a planetary model for an atom. The nucleus of an atom is composed of protons. The electrons surround the nucleus and occupy most of the volume of the atom.
- According to Bohr's atomic model, the electron in an atom revolves around the nucleus in fixed circular orbits called shells. Isotopes are atoms of an element that differ in the number of neutrons.
- ${}^{235}_{92}\text{U}$ isotope is used in nuclear reactors and atomic bombs.
- Radioactive isotopes have many applications in science and medicines such as killing cancer cells, diagnosing thyroid problem, to image the brain, to detect obstruction in the circulatory system, to date rocks, soils, mummies etc.
- A shell or energy level is divided into sub-shells.
- There are four types sub-shell s, p, d, and f.
- The arrangement of electrons in sub-shells is called as the electronic configuration.
- According to the Auf Bau Principle, electrons fill the lowest energy levels first.
- **References for additional information**
- B.Earl and LDR Wilford, Introduction to Advanced Chemistry.
- Iain Brand and Richard Grime, Chemistry (11-14).

REVIEW QUESTIONS

1. Encircle the correct answer.

- (i) Chlorine has two isotopes, both of which have
 - (a) same mass number.
 - (b) same number of neutrons.
 - (c) different number of protons.
 - (d) same number of electrons.
- (ii) Number of neutrons in $^{27}_{13}\text{M}$ are
 - (a) 13
 - (b) 14
 - (c) 27
 - (d) 15
- (iii) Which isotope is commonly used to irradiate cancer cells?
 - (a) Iodine-123
 - (b) Carbon-14
 - (c) Cobalt-60
 - (d) Iodine-131
- (iv) M shell has sub-shells:
 - (a) 1s, 2s
 - (b) 2s, 2p
 - (c) 3s, 3p, 3d
 - (d) 1s, 2s, 3s
- (v) A sub-shell that can accommodate 6 electrons is
 - (a) s
 - (b) d
 - (c) p
 - (d) f
- (vi) $_{11}\text{Na}$ has electronic configuration:
 - (a) $1s^2, 2s^2, 3s^1$
 - (b) $1s^2, 2s^2, 2p^7$
 - (c) $1s^2, 2s^2, 2p^5, 3s^2$
 - (d) $1s^2, 2s^2, 2p^6, 3s^1$
- (vii) Which of the following statement is not correct about isotopes?
 - (a) they have same atomic number
 - (b) they have same number of protons
 - (c) they have same chemical properties
 - (d) they have same physical properties
- (viii) Which isotope is used in nuclear reactors?
 - (a) U-234
 - (b) U-238
 - (c) U-235
 - (d) All of these

2. Give short answer.

- (i) Distinguish between shell and sub-shell
- (ii) Why an atom is electrically neutral?
- (iii) How many sub-shells are there in N shell.
- (iv) Give notation for sub-shells of M shell.
- (v) List the sub-shells of M Shell in order of increasing energy
- (vi) Can you identify an atom without knowing number of neutrons in it?

3. The electronic configurations listed are incorrect. Explain what mistake have been made in each and write correct electronic configurations.
4. $x = 1s^2, 2s^2, 2p^4, 3p^2$, $y = 1s^2, 2s^1, 2p^1$, $z = 1s^2, 2s^2, 2p^5, 3s^1$
5. Which orbital in each of the following pairs is lower in energy?
6. (a) 2s, 2p (b) 3p, 2p (c) 3s, 4s
7. Draw Bohr's Model for the following atoms indicating the location for electron, protons and neutrons:
8. (a) Potassium (Atomic No 19, Mass No. 39)
(b) Silicon (Atomic No. 14 Mass No. 28)
(c) Argon (Atomic No. 18 Mass No. 39)
9. Write electronic configuration for the following elements:
10. (a) ${}_{14}^{28}\text{Si}$ (b) ${}_{12}^{24}\text{Mg}$ (c) ${}_{13}^{27}\text{Al}$ (d) ${}_{18}^{40}\text{Ar}$
11. State the importance and uses of isotopes in various fields of life.
12. The atomic number of an element is 23 and its mass number is 56.
a. How many protons and electrons does an atom of this element have?
b. How many neutrons does this atom have?
13. The atomic symbol of aluminium is written as ${}_{13}^{27}\text{Al}$. What information do you get from it?

