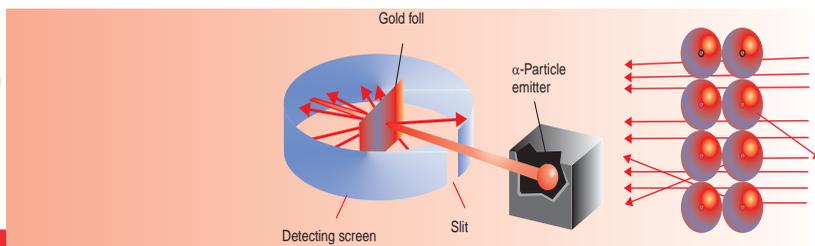


UNIT 2



Atomic Structure and the Periodic Table

Unit Outcomes

At the end of this unit, you should be able to:

- *understand the historical development of atomic structure;*
- *explain the experimental observations and inferences made by some famous scientists to characterize the atom;*
- *list and describe the subatomic particles;*
- *explain the terms “atomic mass” and “isotope”;*
- *understand electromagnetic radiation, atomic spectra and Bohr’s models of the atom;*
- *do calculations involving atomic structure;*
- *describe the quantum mechanical model of the atom, related postulates and principles;*
- *demonstrate an understanding of periodic law and how electronic configuration of atoms are related to the orbital diagrams and can explain periodic trends; and*
- *describe scientific inquiry skills along this unit: inferring, predicting, classifying, comparing, and contrasting, communicating, asking questions and making generalizations.*

MAIN CONTENTS

- 2.1 Historical Development of Atomic Nature of Substances
- 2.2 Dalton's Atomic Theory and the Modern Atomic Theory
 - Postulates of Dalton's Atomic Theory
 - How the Modern Theory Explains the Mass Laws
- 2.3 Early Experiments to Characterize the Atom
 - Discovery of the Electron
 - Radioactivity and Discovery of Nucleus
 - Discovery of Neutron
- 2.4 Makeup of the Nucleus
 - Constituents of the Nucleus
 - Atomic Mass and Isotopes
- 2.5 Electromagnetic Radiation (EMR) and Atomic Spectra
 - Electromagnetic Radiation
 - The Quantum Theory and Photon
 - Atomic Spectra
 - The Bohr Model of the Hydrogen Atom
 - Limitations of the Bohr Theory
- 2.6 The Quantum Mechanical Model of the Atom
 - The Heisenberg's Principle
 - Quantum Numbers
 - Shapes of Atomic Orbitals
- 2.7 Electronic Configurations and Orbital Diagrams
 - Ground State Electronic Configuration of the Elements

2.8 Electronic Configurations and the Periodic Table of the Elements

- The Modern Periodic Table
- Classification of the Elements
- Periodic Properties
- Advantages of Periodic Classification of the Elements

Start-up Activity

Form a group and try to explore the basic building blocks of the following:

- a proteins,
- b steel, and
- c paper.

Why is it important to understand the properties of different types of matter?

After the discussion in your group, share your findings with the rest of the class.

2.1 HISTORICAL DEVELOPMENT OF THE ATOMIC NATURE OF SUBSTANCES

At the end of this section, you should be able to:

- State briefly the historical development of atomic nature of substances.

Activity 2.1



1. Form a group and try to find the smallest constituent of following materials:
table salt, paper, sand, wood etc.
2. Discuss the following questions:
 - a What is matter made up of?
 - b Are all atoms similar in structure?
 - c When and how were these atoms discovered?

- d Can we see atoms with our naked eyes?
- e How small are these atoms?
- f Are atoms indivisible?

Share your ideas with rest of the class.:

This section will help us to get an understanding of the electronic structure of the atoms. A great deal of our understanding of electronic structure has come from studies of the properties of light or radiant energy. We begin our study by considering the historical development of the atomic nature of substances, and then we will learn how the atomic theories were developed.

A theory of the structure and behaviour of atoms has taken more than two millennia to evolve, from ancient Greek philosophers to the high-tech experiments of modern scientists. Prior to the scientific revolution and the development of the scientific method, starting in the 16th century, ideas about the atom were mainly speculative. It was not until the very end of the 19th century that technology became advanced enough to allow scientists a glimpse of the atom's constituent parts: the electron, proton and neutron.

Activity 2.2



In your Grade 9 chemistry lesson, you have learned about historical development of the atomic nature of substances.

1. Form a group and discuss on the belief of ancient Greek philosophers about the atomic nature of substances. Present your group discussion report to the class.
2. Refer to different chemistry reference books and, in group, discuss on how the beliefs about the structure of matter evolved. Start from early historical developments.

2.2 DALTON'S ATOMIC THEORY AND THE MODERN ATOMIC THEORY

At the end of this section, you should be able to:

- state postulates of Dalton's atomic theory;
- state postulates of Modern atomic theory; and

- state the laws of conservation of mass, definite proportions, multiple proportions and the basis of each of these laws.

Activity 2.3



Form a group and do the following activity:

Take 1g of iron fillings and 1.75g of sulphur powder. Heat them strongly to form a new compound iron (II) sulphide. Weigh the new compound formed.

NOTE: Perform the activity under the guidance of your teacher.

Discuss the following questions and share your findings with the rest of the class.

- What is the weight of new compound formed?
- Can we create or destroy mass?
- Is there any fixed ratio in which elements combine?
- What are the laws of chemical combination?
- Will the molecular formula of the synthesized iron (II) sulphide be same as that available in the laboratory?

In Grade 9, you learned about Dalton's and Modern Atomic Theories. Here again you will spend some more time on these theories.

2.2.1 Postulates of Dalton's Atomic Theory

Scientific laws usually develop based on previous scientific findings. The laws that are the basis for Dalton's atomic theory are the law of conservation of mass and the law of definite proportions.

What is the law that Dalton formulated based on these two laws and write its statement in your notebook? Illustrate this law using C and O to form CO.

Example 2.1

1. A nitrogen-oxygen compound is found to have an oxygen-to-nitrogen mass ratio of 1.142 g oxygen for every 1.000 g of nitrogen. Which of the following oxygen-to-nitrogen mass ratio(s) is/are possible for different nitrogen-oxygen compound(s)?

a 2.285	c 0.571
b 1.000	d 2.500
2. Determine, which of the following oxygen-to-nitrogen mass ratio(s) is/are also possible for nitrogen-oxygen compound. (Refer Example 2.1 No. 1 for required information)

a 0.612	c 1.713
b 1.250	d 2.856

Solution :

1. The given compound has 1.142 g of oxygen and 1.000 g of nitrogen. Response (a) has 2.285 g of oxygen for the same 1.000 g of nitrogen. The ratio of the masses of oxygen, 2.285:1.142, is almost exactly 2:1. Response (a) seems to be correct possibility, so is response (c). Here the ratio is $0.571:1.142 = 0.500 = 1:2$.

Responses (b) and (d) are not possibilities. They yield ratios of $1.000:1.142 = 0.875$ and $2.500:1.142 = 2.189$, respectively. Neither of these can be expressed as a ratio of small whole numbers.
2. By the same method,

a $0.612:1.142 = 0.536:1$ is not possible
b $1.250:1.142 = 1.095:1$ is not possible
c $1.713:1.142 = 1.5:1$ or $3:2$, is possible
d $2.856:1.142 = 2.5:1$ or $5:2$ is possible

Activity 2.4



In your Grade 9 chemistry lesson, you have learned about the laws of conservation of mass and the law of definite proportion. Form a group and discuss the following statements. Share your ideas with the rest of the class.

- The mass of a piece of wood before and after it is burnt to ashes is not the same, showing that mass can be created or destroyed. Is it a violation of law of conservation of mass?
What would be the mass of products if the burning of wood was carried out in a closed container?
- Sugar consists of C, H, and O atoms. When a certain amount of sugar is burned in a crucible, nothing remains in the crucible. Have the carbon, hydrogen and oxygen atoms of the sugar been destroyed? If so, how? If not, where have they disappeared?
- A sample of sodium chloride from Afar contains the same percent by mass of sodium as one from Tigray.
- Arsenic and oxygen can combine to form one compound, which is 65.2% by mass arsenic, and another compound, which is 75.8% by mass arsenic.

Exercise 2.1

- List the postulates of Dalton that continue to have significance (are retained in modern atomic theory).
- Match the atomic theory statements in **part A** with the matching items in **part B**.

Part A	Part B
I All matter is composed of extremely small, indivisible particles called atoms.	a Although graphite and diamond have different properties (due the nature of interatomic bonding) they are composed only of carbon. The carbon atoms are identical.
II All atoms of a given element are identical in mass and other properties, but atoms of one element differ from the atoms of every other element.	b $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, not CS_2 or NaCl .
III Atoms are not created, destroyed or converted into other kinds of atoms during chemical reactions. They are simply rearranged into new compounds.	c There are 6.02×10^{23} atoms in 55.85 g of iron.

3. How does the atomic theory account for the fact that when 1.00 g of water is decomposed into its elements, 0.111 g of hydrogen and 0.889 g of oxygen are obtained regardless of the source of the water?
4. Hydrogen peroxide is composed of two elements: hydrogen and oxygen. In an experiment, 1.250 g of hydrogen peroxide is fully decomposed into its elements.
 - a If 0.074 g of hydrogen are obtained in this experiment, how many grams of oxygen must be obtained?
 - b Which fundamental law does this experiment demonstrate?
 - c How is this law explained by the atomic theory?
5. A 15.20 g of nitrogen will react with 17.37 g, 34.74 g, or 43.43 g of oxygen to form three different compounds.
 - a Calculate the mass of oxygen per gram of nitrogen in each compound.
 - b How do the numbers in part (a) support the atomic theory?

2.2.2 Postulates of Modern Atomic Theory

Activity 2.5



Form a group and discuss the following questions:

Which of Dalton's postulates about atoms are inconsistent with later observations?

Do these inconsistencies mean that Dalton was wrong? Is Dalton's model still useful?

Share your ideas with the rest of the class.

Most of the experiments conducted during the development of the modern atomic theory will be presented and explained in **Sections 2.3 - 2.6**. In this section, generalizations derived from the experiments are presented as postulates of modern atomic theory.

Modern atomic theory is generally said to begin with **John Dalton**. Dalton's work was mainly about the chemistry of atoms – how they combine to form new compounds – rather than about the internal structure of atoms. However, Dalton never denied the possibility of atoms having a structure. Modern theories about the physical structure of atoms did not begin until **J.J. Thomson** discovered the electron in 1897.

In 1913, a Danish physicist, **Niels Bohr**, who had studied under both **Thomson** and **Rutherford**, further modified the nuclear model. He proposed that electrons move only

in restricted, successive shells and that the valence electrons determine the chemical properties of different elements. In the 1920s, **Bohr's theory** became the basis for quantum mechanics, which explained in greater detail the complex structure and behaviour of atoms.

Activity 2.6



Form a group and discuss the following questions:

1. Describe the limitations of Dalton's Atomic Theory.
2. Write the postulates of Modern Atomic Theory.

Share your discussion with the rest of the class.

2.2.3 How the Modern Theory Explains the Mass Laws

Activity 2.7



Form a group; use your Grade 9 knowledge and discuss the modern atomic theory to explain the:

- | | |
|--------------------------------|-----------------------------------|
| a law of conservation of mass | b law of definite proportions and |
| c law of multiple proportions. | |

Share your discussion with the rest of the class.

Chemical compounds are formed when atoms combine in whole-number ratios (*law of multiple proportions*). A given compound always has the same relative number and types of atoms (*law of constant composition*).

In a chemical reaction, the mass of the reactants (*starting materials*) equals the mass of the reaction products (*law of conservation of mass*), provided no reactant remains.

2.3 EARLY EXPERIMENTS TO CHARACTERIZE THE ATOM

At the end of this section, you should be able to:

- discuss the discovery of the electron;
- describe the properties of cathode rays;
- define the terms: radioactivity, radioactive decay and radio-isotope;
- describe the common types of radioactive emissions;
- discuss the alpha scattering experiment; and
- describe the major contribution of experiments by Thomson, Millikan and Rutherford concerning atomic structure.

2.3.1 Discovery of the Electron

Historical Note



Sir Joseph Thomson

J. J. Thomson (1856-1940) was a British physicist and Nobel laureate. Sir Joseph John Thomson was born near Manchester, England, and educated at Owens College (now part of Victoria University of Manchester) and Trinity College, University of Cambridge. At Cambridge he taught mathematics and physics, served as Cavendish Professor of Experimental Physics, and was (1918-40) master of Trinity College. He was also president of the Royal Society (1915-20) and professor of natural philosophy at the Royal Institute of Great Britain (1905-18).

Thomson was awarded the 1906 Nobel Prize in physics for his work on the conduction of electricity through gases. He is considered the discoverer of the electron through his experiments on the stream of particles (electrons) emitted by cathode rays. A theorist as well as an experimenter, in 1898, Thomson advanced the "plum-pudding" theory of atomic structure, holding that negative electrons were like plums embedded in a pudding of positive matter. Thomson was knighted in 1908.

Cathode Rays

In 1879, the English scientist **William Crookes** (1832-1919) experimented with gas-discharge tubes. When a very high electrical potential (~ 10,000 volts) is applied

across a gas taken in a discharge tube of a very low pressure (~ 0.001 torr) some radiations are emitted from cathode. These radiations are called cathode rays. Fig 2.1, shows emission of cathode rays in a discharge tube. At this stage the glass walls of the discharge tube opposite to the cathode starts glowing with a faint greenish light. It is now known that this greenish glow on the walls is due to the bombardment of the glass wall with the cathode rays.

Cathode rays normally travel in straight lines, but are deflected when a magnet is brought nearby (Figure 2.1b).

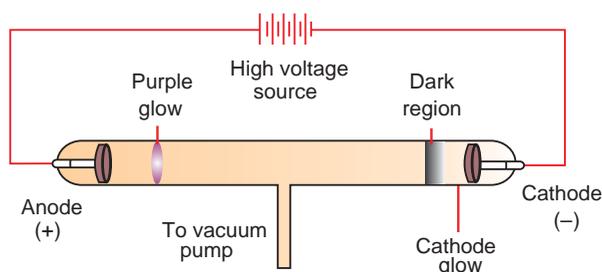
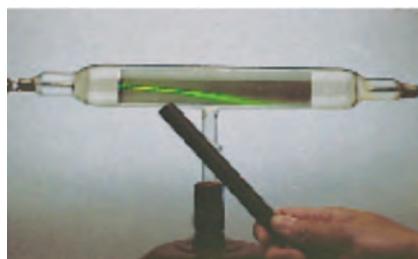


Figure 2.1 (a) Electric discharge in an evacuated tube.



(b) The cathode ray is bent in the presence of a magnet.

An English physicist Joseph John Thomson (1856-1940) in 1897 studied the behaviour of cathode rays in electric and magnetic fields, Thomson established clearly that the rays consist of negatively-charged particles. Moreover, his experiments showed that the particles were identical, regardless of the materials from which the electrodes were made or the type of gas in the tube. Thomson concluded that these negatively charged particles were constituents of every kind of atom. We now call these particles **electrons**, a term that had been coined by the Irish Physicist George Stoney in 1891 to describe the smallest unit of electric charge. **Cathode rays** are beams of electrons.

Perhaps Thomson's most significant experiment was the one illustrated and described in Figure 2.2. By measuring the amount of deflection of a cathode ray beam in electric and magnetic fields of known strengths, Thomson was able to calculate the ratio of the mass of an electron to its charge. If we denote the mass of the electron as m_e and its electrical charge as e , the mass-to-charge ration is m_e/e . This ratio has a value of -5.686×10^{-12} kg C⁻¹ (kilograms per coulomb).

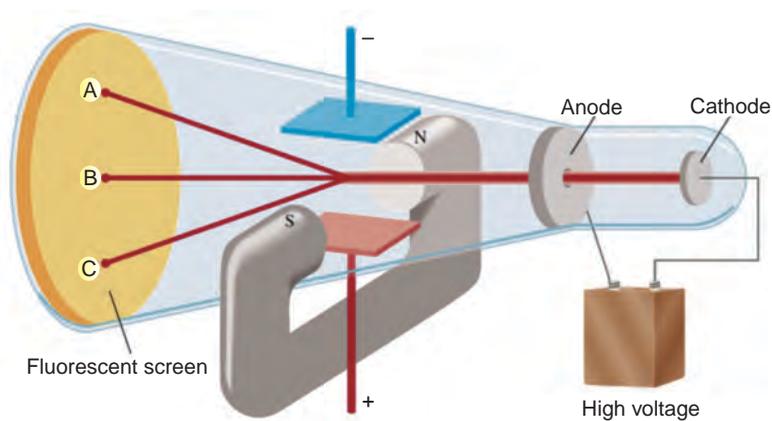


Figure 2.2 Thomson's apparatus for determining the charge-to-mass ratio.

In 1909, **Robert A. Millikan**, an American physicist, determined the charge on the electron by observing the behaviour of electrically-charged oil drops in an electric field. Based on careful experiments, Millikan established the charge on an electron as $e = -1.602 \times 10^{-19}$ C. From this value and the value for m_e/e , we can calculate the mass of an electron.

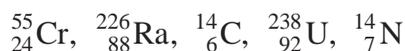
$$\begin{aligned} m_e &= \frac{m_e}{e} \times e = -5.686 \times 10^{-12} \text{ kg C}^{-1} \times -1.602 \times 10^{-19} \text{ C} \\ &= 9.109 \times 10^{-31} \text{ kg} \end{aligned}$$

In 1923, Millikan was awarded the Nobel Prize in physics for his oil-drop experiment.

2.3.2 Radioactivity and Discovery of Nucleus

Radioactivity

Radioactivity is the spontaneous emission of radiation from the unstable nuclei of certain isotopes. Isotopes that are radioactive are called **radioactive isotopes** or **radioisotopes**. For example,



are radioactive isotopes. Radioactive decay is defined as a nuclear breakdown in which particles or (electromagnetic) radiation is emitted. Shortly after the discovery of

radioactivity, three types of rays were identified in the emanations from radioactive substances. One type called **alpha** (α) particles which consist of particles that have a mass that is about four times that of a hydrogen atom. They also have a charge twice the magnitude of an electron but positive rather than negative. An alpha particle is now known to be a doubly-ionized helium atom, that is, He^{2+} .

A second type of radiation was shown to consist of negatively-charged particles, identical to cathode rays. These particles are called **beta** (β) particles, which are electrons, coming from inside the nucleus.

The third type of radiation, called **gamma** (γ) rays, is a form of electromagnetic radiation much like the **X-rays** but of even higher energy. Like **X-rays**, but unlike alpha and beta particles, gamma rays are a form of energy and not a form of matter.

Activity 2.8



Form a group and do the following. Share your ideas with the rest of the class.

In your notebook, make a table of the three kinds of radiation you have read about. Use the given table headings for the types of radiation, response to magnetic field, mass and charge as follows. Then fill in the cells.

Types of radiation	Greek letter	Mass Number	Charge	Deflected

Discovery of Nucleus

In 1910, the New Zealand chemist and physicist, Ernest Rutherford, who had studied with J.J. Thomson decided to use **α -particles** to probe the structure of atoms. Together with his associate Hans Geiger, Rutherford carried out a series of experiments using very thin foils of gold and other metals as targets for **α -particles** from a radioactive source.

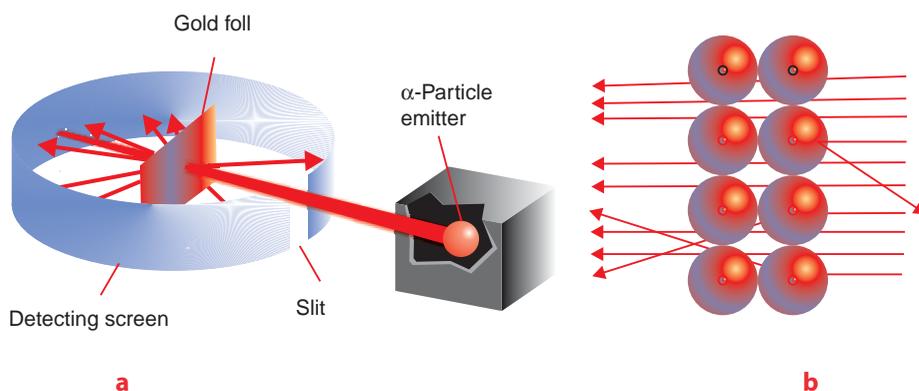


Figure 2.3 a Rutherford's experimental design for measuring the scattering of α -particles by a piece of gold foil.

b Magnified view of α -particles passing through and being deflected by the nucleus.

Activity 2.9



From a group and discuss Rutherford's experiment as shown in Figure 2.3 a and b answer the following questions:

1. Why majority of the α -particles penetrating the foil where undeflected?
2. Why small fraction of the α -particles showed slight deflection?
3. Why all α -particles did not bounced by an angle of 180° ?
4. Based on the finding of Rutherford's experiment how will you conclude that most of the space in the atom is empty?
5. Which observation led to conclusion that all the positive charge in the atom is concentrated in the nucleus?

Report your findings to the class.

2.3.3 Discovery of Neutron

Except for the lightest hydrogen isotope, protium (^1_1H), atoms have more mass than is indicated by the numbers of their protons (*Before the 1930's, protons were considered as the sole contributors to the mass of an atom*). For example, a helium nucleus, with two protons has a mass four times that of hydrogen. If all the mass came from the protons, a helium atom would have only twice the mass of a hydrogen atom. The reason for this "excess" mass puzzled scientists for several years.

One hypothesis was that the atomic nucleus also contained electrically neutral fundamental particles.

In the 1920s and early 1930s, alpha particles were used as projectiles to bombard a variety of materials. Bombardment of beryllium atoms produced a strange, highly penetrating form of radiation. In 1932, **James Chadwick** (1891-1972) showed that this radiation was best explained as a beam of neutral particles. These particles, called *neutrons*, were found to have about the same mass as protons but no electric charge. This discovery finally provided an explanation for the mysterious excess mass. A helium atom has two protons and two neutrons. Because protons and neutrons have roughly the same mass (and *electrons have almost no mass*) the helium atom should have about four times the mass of the hydrogen atom. The mass of a neutron, $m_n = 1.67493 \times 10^{-27}$ kg, is about 1840 times the mass of electron.

2.4 MAKE UP OF THE NUCLEUS

At the end of this section, you should be able to:

- describe make-up of the nucleus;
- define atomic mass;
- define isotope; and
- calculate the relative atomic mass (atomic mass) of naturally occurring isotopic elements.

2.4.1 Constituents of the Nucleus

Activity 2.10



Form a group and complete the following table by drawing in your notebook. Discuss your ideas with the rest of the class.

Symbol	Atomic Number	Number of neutrons	Mass Number	Number of electrons
${}^{12}_{6}\text{C}^{+}$	–	–	–	–
–	16	–	32	18
–	–	14	27	13
Pb	82	126	–	–

In 1914, Rutherford suggested that the smallest positive-ray particle was the fundamental unit of positive charge in all nuclei. He called this particle, which has a charge equal in magnitude but opposite in sign to that of an electron, a **proton**. A proton has a mass of $m_p = 1.67262 \times 10^{-27}$ kg, which is about 1840 times the mass of oppositely-charged electrons. Rutherford proposed that protons constitute the positively-charged matter in the nuclei of all atoms.

Table 2.1 The three subatomic particles

Particle	Symbol	Relative mass in atom	Approximate charge	Relative location
Proton	p^+	1	+1	Nucleus
Neutron	n^0	1	0	Nucleus
Electron	e^-	0.000545	-1	Outside nucleus

To form the picture of an atom in your mind, think of it as something like this: if an entire atom were represented by a football field, the nucleus would be only about as big as a lentil located at the centre of the field.

2.4.2 Atomic Mass and Isotopes

Not all atoms of an element are identical in mass. All carbon atoms have six protons in the nucleus ($Z = 6$) but only 98.89 % of naturally occurring carbon atoms have six neutrons in the nucleus ($A = 12$). A small percentage (1.11 %) have seven neutrons in the nucleus ($A = 13$), and even fewer (less than 0.01 %) have eight ($A = 14$).

Most elements found in nature are mixtures of isotopes. The average mass for the atoms in an element is called the atomic mass of the element and can be obtained as averages over the relative masses of the isotopes of each element, weighted by their observed fractional abundances. If an element consists of n isotopes, of relative masses $A_1, A_2 \dots A_n$ and fractional abundances of $f_1, f_2 \dots f_n$, then the average relative atomic mass (A) of the element is:

$$A = A_1 f_1 + A_2 f_2 + \dots + A_n f_n.$$

Activity 2.11



Form a group and discuss the following:

1. Are neutrons present in all atoms?

2. Can two atoms have the same number of electrons but different number of neutrons?
3. Can an atom have unequal number of electrons and protons?
4. Do the chemical properties of an atom depend on the number of electrons, protons or neutrons?
5. Why the atomic masses for most elements are not whole numbers?
6. Does an atom of zinc (Zn) have about the same mass as an atom of sulphur (S), about twice the mass, or about half the mass?
7. Nitrogen has two naturally occurring isotopes, N-14 and N-15. The atomic mass of nitrogen is 14.007 amu. Which isotope is more abundant in nature?

share your ideas with rest of the class.

Example 2.2

There are two isotopes of lithium found on earth. Isotope ${}^6\text{Li}$ ($6.01512 m_u$) accounts for 7.42% of the total, and isotope ${}^7\text{Li}$ ($7.01600 m_u$) accounts for the remaining 92.58%. What is the average atomic mass of lithium?

Solution:

$$\frac{7.42}{100} \times 6.01512 m_u + \frac{92.58}{100} \times 7.01600 m_u = 6.942 m_u$$

Exercise 2.2

1. What is the mass number of an isotope of tin that has 66 neutrons and 50 protons?
2. Calculate the number of protons and neutrons for ${}^{24}_{12}\text{Mg}$ and ${}^{88}_{38}\text{Sr}$.
3. Why do isotopes of an element have similar chemical properties?
4. Element **X** is toxic to humans in high concentration but essential to life at low concentrations. Identify element **X** whose atoms contain 24 protons and write the symbol for the isotope with 28 neutrons.

5. Copper (Cu: atomic mass $63.546 m_u$) contains the isotopes ^{63}Cu (mass = $62.9298 m_u$) and ^{65}Cu (mass = $64.9278 m_u$). What percent of a Cu atom is ^{65}Cu ?
6. The element chlorine contains two isotopes: ^{35}Cl , which has a mass of $34.97 m_u$, and ^{37}Cl , which has a mass of $36.97 m_u$. Calculate the percentage of each chlorine isotope. The average atomic mass of chlorine is $35.5 m_u$.
7. Carbon exists as the isotopes carbon-12, with a fractional abundance of 0.9890 and a mass of exactly $12 m_u$, and carbon-13, with a fractional abundance of 0.0110 and a mass of $13.00335 m_u$. Calculate the average atomic mass of carbon.

2.5 ELECTROMAGNETIC RADIATION (EMR) AND ATOMIC SPECTRA

At the end of this section, you should be able to:

- characterize electromagnetic radiation (EMR) in terms of wave length, frequency and speed;
- calculate the wave length and frequency of EMR;
- explain that light has both wave and particle natures;
- define photon as a unit of light energy;
- explain how photon theory explains the photoelectric effect and the relationship between photons absorbed and electrons released;
- explain that emission spectra of atoms consist of series of lines;
- state Bohr's assumption of energy of electrons in the hydrogen atom;
- explain that the line spectrum of hydrogen demonstrates the quantized nature of the energy of its electron;
- explain that atoms emit or absorb energy when they undergo transition from one energy state to another;
- explain the shortcomings of the Bohr theory; and
- calculate the radius of an electron orbit, velocity and the energy of the electron, using the Bohr model.

2.5.1 Electromagnetic Radiation

Activity 2.12



Form groups and discuss on the following questions. Share your ideas with the rest of the class.

The nature of the nucleus has been known for quite some time - since the days of Rutherford and his associates. What has been illusive however is the position and velocity of electrons.

1. Is it possible to know the exact location of an electron? Defend your suggestion.
2. Do electrons have a particle nature or a wave nature?
3. Explain why an electron does not enter the nucleus even though they are oppositely charged?
4. What is the velocity of an electron?

In 1873, James Clerk Maxwell proposed that light consists of electromagnetic waves. According to his theory, an electromagnetic wave has an electric field component and a magnetic field component. Further, his theory accurately describes how energy, in the form of radiation, propagates through space as electric and magnetic fields. Electromagnetic radiation is the emission and transmission of energy in the form of electromagnetic waves.

The wave properties of electromagnetic radiation are described by two interdependent variables, frequency and wavelength. **Wavelength** (λ , *Greek lambda*) is the distance between any point on a wave and the corresponding point on the next wave; that is, the distance the wave travels during one cycle.

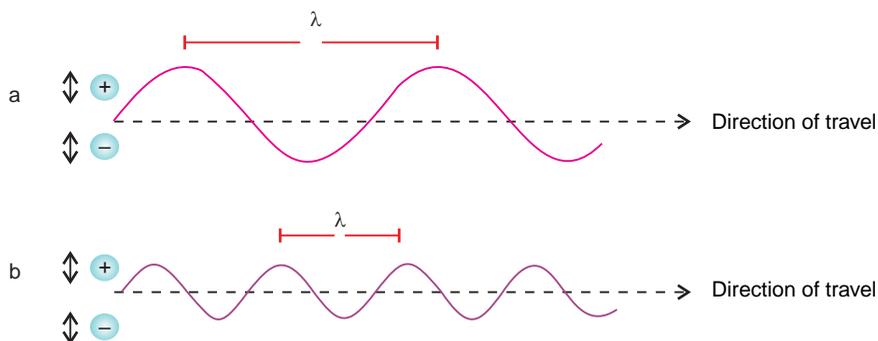


Figure 2.4 Electromagnetic waves.

In **Figure 2.4**, two waves with different wave lengths (λ) and thus different frequencies (ν) are shown. Wavelength is commonly expressed in meters, but since chemists often deal with very short wavelengths the nanometer, picometer and the angstrom are also used. **Frequency** (ν , Greek nu) is the number of cycles that pass a given point in space per second, expressed in units of s^{-1} or hertz (Hz).

The speed of the electromagnetic wave (light), c (distance travelled per unit time, in meters per second), is the product of its frequency (*cycles per second*) and its wavelength (*metres per cycle*),

$$c_0 = \nu \times \lambda \quad \dots(2.1)$$

In vacuum, light travels at a speed of $2.99792458 \times 10^8 \text{ m s}^{-1}$ ($3.00 \times 10^8 \text{ m s}^{-1}$ to three significant figures).

The speed of an electromagnetic wave depends on the nature of the medium through which the wave is travelling. The speed of an electromagnetic wave in medium (c) is the product of its wavelength and its frequency.

$$c = \lambda \times \nu \quad \dots(2.2)$$

Another characteristic of a wave is its amplitude, the height of the crest (or depth of the trough) of the wave. The amplitude of an electromagnetic wave is a measure of the strength of its electric and magnetic fields. Thus, amplitude is related to the intensity of the radiation, which we perceive as brightness in the case of visible light.

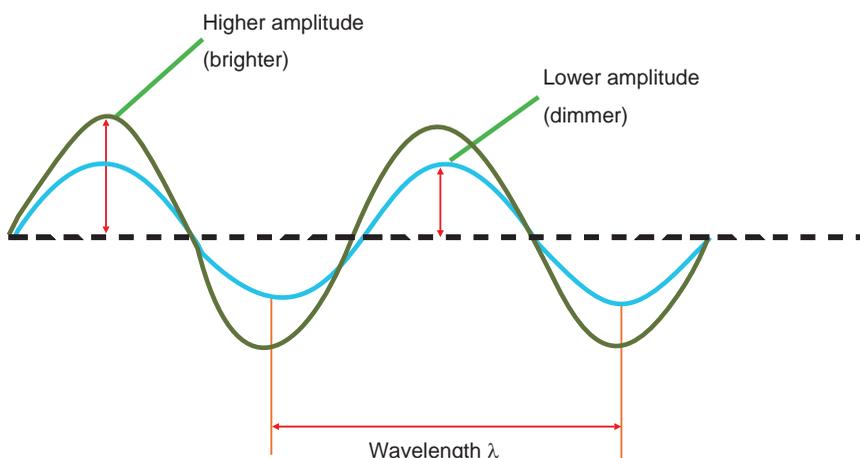


Figure 2.5 Amplitude (intensity) of waves.

Light of a particular shade of red, for instance, always has the same frequency and wavelength, but it can be dim (*low amplitude*) or bright (*high amplitude*).

Visible light occupies a small portion of the continuum of radiant energy which is known as the **electromagnetic spectrum** (Figure 2.6). The electromagnetic waves in the different spectral region travel at the same speed but differ in frequency and wavelength.

The long wavelength, low-frequency portion of the spectrum comprises the microwave and radiowave regions. The infrared (IR) region overlaps the microwave region on one end and the visible region on the other.

We perceive different wave lengths (*or frequencies*) of visible light as different colours, from red ($\lambda = 750$ nm) to violet ($\lambda = 400$ nm). Light of a single wavelength is called **monochromatic** (Greek “one colour”), whereas light of many wavelengths is polychromatic (Greek “many colours”). White light is polychromatic.

The region adjacent to visible light on the short-wavelength end consists of **ultraviolet** (UV) radiation. Still shorter wavelengths (*higher frequencies*) make up the X-ray and gamma ray (γ -ray) regions.

Thus, a TV signal, Infrared (IR) light, and a gamma ray emitted by a radioactive element differ principally in frequency and wavelength.

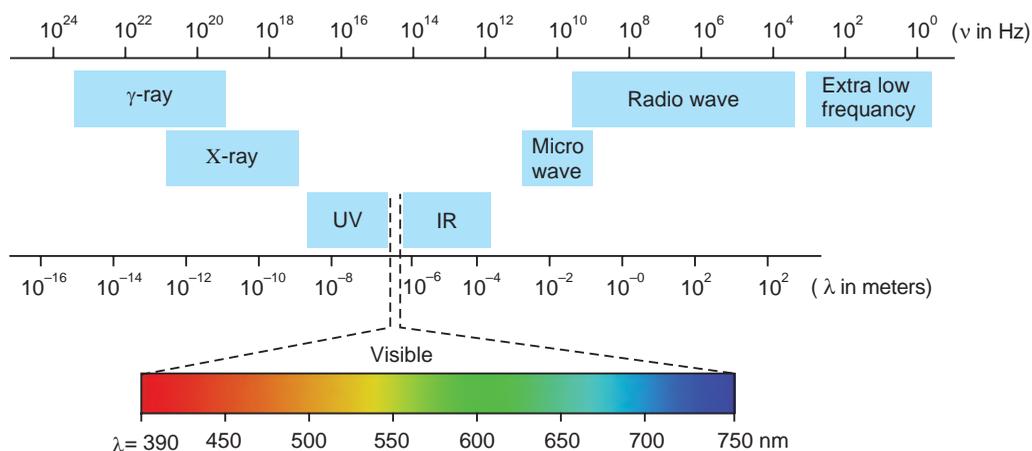


Figure 2.6 Regions of the electromagnetic spectrum.

Example 2.3

1. The yellow light given off by a sodium lamp has a wavelength of 589 nm. What is the frequency of this radiation?
2. A dental hygienist uses X-ray ($\lambda = 1.00 \text{ \AA}$) to take a series of dental radiographs while the patient listens to an FM radio station ($\lambda = 325 \text{ cm}$) and looks out the window at the blue sky ($\lambda = 473 \text{ nm}$). What is the frequency (in s^{-1}) of the electromagnetic radiation for each source?

Solution:

1. Rearranging Equation 2.1, and $\nu = c_o/\lambda$.

We insert the value for c and λ and then convert nm to m. This gives us

$$\nu = \frac{3.00 \times 10^8 \text{ m/s}}{589 \text{ nm}} \times \frac{10^9 \text{ nm}}{1 \text{ m}} = 5.09 \times 10^{14} \text{ s}^{-1}$$

2. Because we are provided with the wavelengths, we can find the frequencies from Equation 2.1. Since c_o has units of m s^{-1} , we first convert the entire wavelength to metres.

For X-ray,

$$\lambda = 1.00 \text{ \AA} \times \frac{10^{-10} \text{ m}}{1 \text{ \AA}} = 1.00 \times 10^{-10} \text{ m}$$

$$\lambda = c_o/\nu = \frac{3.00 \times 10^8 \text{ m/s}}{1.00 \times 10^{-10} \text{ m}} = 3.00 \times 10^{18} \text{ s}^{-1}$$

For the radio station, $\lambda = 325 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} = 3.25 \text{ m}$

$$\nu = \frac{3.00 \times 10^8 \text{ m/s}}{3.25 \text{ m}} = 9.23 \times 10^7 \text{ s}^{-1}$$

For the blue sky, $\lambda = 473 \text{ nm} \times \frac{10^{-9} \text{ m}}{1 \text{ nm}} = 4.73 \times 10^{-7} \text{ m}$

$$\nu = \frac{3.00 \times 10^8 \text{ m/s}}{4.73 \times 10^{-7} \text{ m}} = 6.34 \times 10^{14} \text{ s}^{-1}$$

Exercise 2.3

1. Some diamonds appear yellow because they contain nitrogeous compounds that absorb purple light of frequency $7.23 \times 10^{14} \text{ s}^{-1}$. Calculate the wavelength (in nm) of the absorbed light.
2. The FM station broadcasts traditional music at 102 MHz on your radio. Units for FM frequencies are given in megahertz (MHz). Find the wavelength of these radio waves in meters (m), nanometers (nm), and angstrom (\AA).

2.5.2 The Quantum Theory and Photon

Soon after Rutherford had proposed his nuclear model, a major problem arose with it. A nucleus and an electron attract each other, so if they are to remain apart, the energy of the electron's movement must balance the energy of attraction. However, the laws of physics had previously established that a charged particle moving in a curved path must give off energy. If this requirement is applied to an orbiting electron, why did not the electron continuously lose energy and spiral into the nucleus? Clearly, if electrons behaved the way as predicted in classical physics, all atoms would have collapsed a long time ago!

The breakthrough that soon followed forced a complete rethinking of the classical picture of matter and energy. In the macroscopic world, the two are distinct: matter occurs in pieces you can hold and weigh; you can change the amount of matter in a sample piece by piece. Energy is “*massless*”; its amount can be changed in a continuous manner. Matter moves in specific paths, whereas light travels in diffuse waves. As soon as 20th century scientists explored the subatomic world, however, these clear distinctions between particulate matter and wavelike energy began to fade. In the following sections, you will examine the theories and experiments that led to the view of a **quantized** or particulate nature of light.

The quantum theory is concerned with the rules that govern the gain or loss of energy from an object. In 1900, the German physicist **Max Planck** came to an entirely new view of matter and energy. He proposed that a hot glowing object could emit (or absorb) only certain amounts of energy.

$$E = nh\nu$$

where E is the energy of the radiation, ν is the frequency, n is a positive integer ($n = 1, 2, 3, \dots$) called **quantum number** and h is proportionality constant now called **Planck's constant**. With energy in joules (J) and frequency in second (s^{-1}), h has units of J s; ($h = 6.626 \times 10^{-34}$ J s)

Planck's contribution was to perceive that when we deal with the gain or loss of energy from objects in the atomic size or subatomic-size range, the rules “seem” to be different from those that apply when we are dealing with the energy gain or loss from objects of ordinary dimensions.

A Very Crude Analogy Might Best Illustrate What is Involved

Imagine a large truck loaded with about 2 tons of fine-grained sand. Let us assume that the amount of sand on the truck is measured by a supersensitive scale that can measure the weight of the 2-ton object to the nearest gram. With this scale as our measuring device, the gain or loss of a few grains of sand from the truck would be too small to be measured. A spade-ful of sand might be added or removed with no change in the scale reading.

Now imagine a tiny little truck operated by a driver of the size of tiny ant. For this little truck, a full load of sand would consist of, perhaps, a dozen grains of sand of the same size as those carried by the large truck. In this microscopic world, the load on the truck can be added to or decreased only by rolling on or off one or more grains of sand. On the scale that weighs this tiny truck, even one grain of sand represents a substantial fraction of the load and is easily measurable.

In the given analogy, the sand represents energy. An object of ordinary or macroscopic dimensions, like the large truck, contains energy in so many tiny pieces that the gain or loss of individual pieces is completely unnoticed. On the other hand, an object of atomic dimensions, such as our imaginary little truck, contains such a small amount of energy that the gain or loss of even the smallest possible piece makes a substantial difference.

The essence of **Planck's** quantum theory is that there is such a thing as a smallest-allowable gain or loss of energy. Even though the amount of energy gained or lost at one time may be very tiny, there is a limit to how small it may be. Planck termed the smallest allowed increment of energy gained or lost a **quantum of energy**. In the analogy, a single grain of sand represents a quantum of sand “energy”.

It should be kept in mind that rules regarding the gain or loss of energy are always the same, whether we are concerned with objects on the size scale of our ordinary

experience or with microscopic objects. However, it is only when dealing with matter at the atomic level of size that the impact of the quantum restriction is evident. Humans, being creatures of macroscopic dimensions, had no reason to suppose that the quantum restriction existed until they devised means of observing the behaviour of matter at the atomic level. The major tool for doing this at the time of **Planck's** work was the observation of the radiant energy absorbed or emitted by matter.

An object can gain or lose energy by absorbing or emitting radiant energy. Planck assumed that the amount of energy gained or lost at the atomic level by the absorption or emission of radiation had to be a whole-number multiple of a constant, times the frequency of the radiant energy.

$$\Delta E = hv, 2hv, 3hv, \dots(2.3)$$

where ΔE is the amount of energy gained or lost. The smallest increment of energy at a given frequency, hv is the quantum of energy.

Example 2.4

Calculate the amount of energy (*that is, the quantum of energy*) that an object can absorb from yellow light, whose wavelength is 589 nm.

Solution:

We obtain the magnitude of a quantum of energy absorbed from equation 2.3 ($\Delta E = hv$; $h = 6.626 \times 10^{-34}$ J s). The frequency, ν , is calculated from the given wavelength, $\nu = c_o/\lambda = 5.09 \times 10^{14}$ s⁻¹. Thus, we have:

$$\begin{aligned}\Delta E = hv &= 6.626 \times 10^{-34} \text{ J s} \times 5.09 \times 10^{14} \text{ s}^{-1} \\ &= 3.37 \times 10^{-19} \text{ J}\end{aligned}$$

At this stage, you may be wondering about the practical applications of **Planck's** quantum theory. **Planck's** theory has within it the seeds of a revolution for the way the physical world is perceived.

The Photoelectric Effect

Light shining on a clean metallic surface can cause the surface to emit electrons. This phenomenon is known as the photoelectric effect. For each metal, there is a minimum frequency of light (*threshold frequency*) below which no electrons are emitted, regardless of how intense the beam of light. In 1905, **Albert Einstein (1879–1955)**

used the quantum theory to explain the photoelectric effect. He assumed that the radiant energy striking the metal surface is a stream of tiny energy packets. Each energy packet, called a **photon**, is a quantum of energy, $h\nu$.

$$E_{ph} = h\nu \quad \dots(2.4)$$

where E_{ph} is the energy of a photon. Thus, radiant energy itself is considered to be quantized. Photons of high-frequency radiation have high energies, whereas photons of lower frequency radiation have lower energy.

When a photon is absorbed by the metal, its energy is transferred to an electron in the metal. A certain amount of energy is required for the electron to overcome the attractive forces that hold it within the metal. Otherwise, it cannot escape from the metal surface, even if the light beam is quite intense.

If a photon has sufficient energy, then an electron is emitted. If a photon has more than the minimum energy required to free an electron, the excess energy appears as the kinetic energy of the emitted electron. This situation is summarized by the equation

$$h\nu = E_k + E_b \quad \dots(2.5)$$

where E_k is the kinetic energy of the ejected electron, and E_b is the binding energy of the electron in the metal. Rewriting equation (2.5), using

$$E_k = \frac{1}{2} m_e v^2 \text{ and } E_b = h\nu_o$$

results in

$$h\nu = h\nu_o + \frac{1}{2} m_e v^2$$

where m_e is mass of an electron, and ν_o is the threshold frequency. Equation (2.5) shows that the more energetic the photon (*high frequency*), the higher the velocity of the ejected electron.

Now consider two beams of light having the same frequency (*which is greater than the threshold frequency*) but different intensities. The more intense beam of light consists of a larger number of photons. Consequently, it ejects more electrons from the metal's surface than the weaker beam of light. Thus, the more intense the light the greater the number of electrons emitted by the target metal. The higher the frequency of the light, the greater will be the kinetic energy of the emitted electrons.

Fortunately, we have equations to quantify these observations on the nature of the photoelectric effect. For example,

$$\Delta E = h\nu = h\nu_o + \frac{1}{2} m_e v^2 \quad \dots(2.6)$$

where ν_o is the threshold frequency, m_e is the mass of the electron and v is the velocity of the emitted electron. The energy (ΔE) of a quantum of radiation ($h\nu$) that

goes into ejecting a photoelectron out of the metal (*the work function $h\nu_o$*) and the binding energy gives us the kinetic energy ($\frac{1}{2}mv^2$). Equation 2.6 is one example of the first law of thermodynamics.

To understand the photoelectric effect, consider the analogy of a truck stuck in the mud. A stream of people can come by, and each of them individually can give a push without getting the truck unstuck. A small tractor also would have less than the threshold energy necessary to dislodge the truck. A farm tractor on the other hand, would be able to overcome the attractive forces of the mud and get the truck out. Larger tractors also could do the job. Some could even exert more pull than necessary thus giving the freed truck some kinetic energy.

Example 2.5

The maximum kinetic energy of the photoelectrons emitted from a metal is 1.03×10^{-19} J when light that has a 656 nm wavelength shines on the surface. Determine the threshold frequency, ν_o , for this metal. Given quantities: $h = 6.626 \times 10^{-34}$ J s, $\lambda = 656$ nm, kinetic energy of photoelectron $E_k = 1.03 \times 10^{-19}$ J.

Solution:

Solve for ν from $c_o = \nu \times \lambda$

$$\nu = c_o / \lambda = \frac{3.00 \times 10^8 \text{ m/s}}{656 \text{ nm} \times 10^{-9} \text{ m/nm}} = 4.57 \times 10^{14} \text{ s}^{-1}$$

Rearrange Equation 2.1 and solve for ν_o

$$\begin{aligned} \nu_o &= \frac{h\nu - E_k}{h} \\ &= \frac{(6.626 \times 10^{-34} \text{ J s} \times 4.57 \times 10^{14} \text{ s}^{-1}) - (1.03 \times 10^{-19} \text{ J})}{6.626 \times 10^{-34} \text{ J s}} = 3.02 \times 10^{14} \text{ s}^{-1} \end{aligned}$$

Therefore, a frequency of 3.02×10^{14} Hz is the minimum (threshold) required to evoke the photoelectric effect for this metal.

Note! The wavelength λ_o corresponding to ν_o is given by $\lambda_o = c/\nu_o$. For this example,

$$\lambda_o = \frac{3.00 \times 10^8 \text{ m/s}}{3.02 \times 10^{14} \text{ s}^{-1}} = 9.93 \times 10^{-7} \text{ m or } 993 \text{ nm.}$$

Photoelectrons will not be emitted from the surface of this metal unless the wavelength of the light is shorter than 993 nm. Remember that higher energies are associated with higher frequencies and shorter wavelengths.

Exercise 2.4

1. List the similarities between microwaves and ultraviolet radiation.
2. How does intensity of a radiation affect the kinetic energy of photons during photoelectric effect?
3. The threshold frequency for metallic potassium is $5.46 \times 10^{14} \text{ s}^{-1}$. Calculate the maximum kinetic energy and velocity that the emitted electron has when the wavelength of light shining on the potassium surface is 350 nm. (The mass of an electron is $9.11 \times 10^{-31} \text{ kg}$.)
4. A laser produces red light of wavelength 632.8 nm. Calculate the energy, in kJ, of one mole of photons of this red light.
5. Two members of the boron family owe their names to bright lines in their emission spectra. Indium has a bright indigo-blue line (451.1 nm), and thallium has a bright green line (535.0 nm). What are the energies of these two spectral lines?

2.5.3 Atomic Spectra

Atomic or line spectra are produced from the emission of photons of electromagnetic radiation (light).

When an element is vaporized, and thermally or electrically excited, it emits light with discrete frequencies. If dispersed by a prism, the light does not create a **continuous spectrum**, or rainbow, as sunlight does. Rather, it produces a line spectrum, a series of fine lines of individual colours separated by colourless (*black*) spaces.

The wavelengths at which the coloured lines occur are characteristic of the element (Figure 2.7). A line spectrum that consists of only relatively few wavelengths that is produced from light emitted by excited atoms is the unique feature of an element, and can be used for identification purposes.

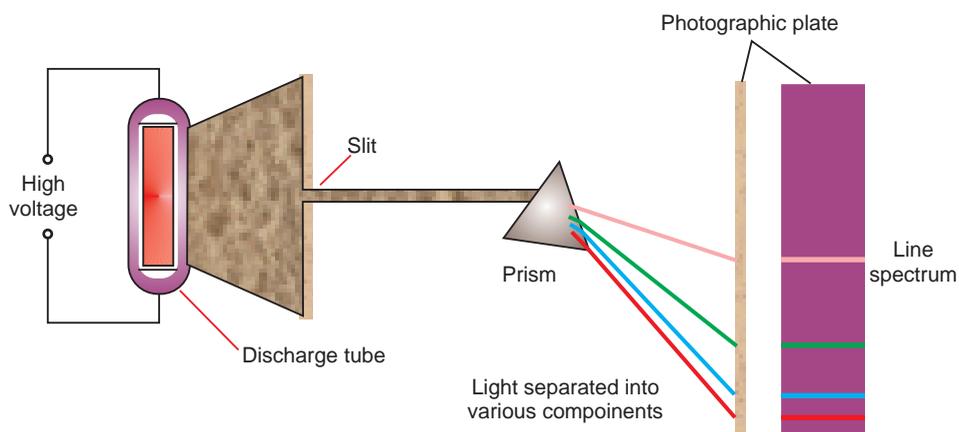


Figure 2.7 An experimental arrangement for studying the emission spectra of atoms and molecules.

2.5.4 The Bohr Model of the Hydrogen Atom

In 1913, **Niels Bohr**, a Danish physicist, who had worked with Rutherford, combined ideas from classical physics and the new quantum theory to explain the structure of the hydrogen atom. He suggested a model for the hydrogen atom that predicted the existence of line spectra. In doing so, he was able to explain the spectrum of radiation emitted by hydrogen atoms in gas-discharge tubes.

Based on the work of **Planck** and **Einstein**, **Bohr** made the revolutionary assumption that certain properties of the electron in a hydrogen atom – including energy, can have only certain specific values. That is to say, these properties are quantized. Bohr proposed the following three postulates for his model.

1. The hydrogen atom has only certain allowable energy levels, called **stationary states**. Each of these states is associated with a fixed circular orbit of the electron around the nucleus.
2. The atom does not radiate energy while in one of its stationary states. That is, even though it violates the ideas of classical physics, the atom does not change energy while the electron moves within an orbit.

3. The electron moves to another stationary state (*orbit*) only by absorbing or emitting a photon whose energy equals the difference in the energy between the two states.

$$E_{ph} = E_f - E_i = h\nu \quad \dots(2.7)$$

The subscripts *f* and *i* represent the final and the initial states, respectively. The **Bohr** radius, denoted by a_o ($a_o = 0.0529$ nm) can be calculated using the formula

$$a_o = \frac{r}{n^2} \quad \dots(2.8)$$

where n is a positive integer which is called **quantum number**.

r is the radius of the orbit and is given by:

$$r = \frac{n^2 \epsilon_o h^2}{\pi m_e e^2 z} \quad \dots(2.9)$$

where ϵ_o is the vacuum dielectric constant ($\epsilon_o = 8.854 \times 10^{-12}$ C V⁻¹ m⁻¹).

A spectral line results from the emission of a photon of specific energy (*and therefore, of specific frequency*), when the electron moves from a higher energy state to a lower one. An atomic spectrum appears as lines rather than as a continuum because the atom's energy has only certain discrete energy levels or states.

In **Bohr's** model, the quantum number n ($n = 1, 2, 3, \dots$) is associated with the radius of the electron's orbit, which is directly related to the atom's energy. The lower the quantum number, the smaller is the radius of the orbit and the lower is the energy level of the atom. When the electron is in the orbit closest to the nucleus ($n = 1$), the atom is in its lowest (first) energy level, which is called the ground state. By absorbing a photon whose energy equals the difference between the first and second energy levels, the electron can move to the next orbit. This second energy level (second stationary state) and all higher levels are called excited states. The hydrogen atom in the second energy level (first excited state) can return to the ground state by emitting a photon of a particular frequency:

$$\Delta E = E_e - E_g \quad \dots(2.10)$$

where E_g and E_e represent the ground and the excited energy states, respectively.

When a sample of atomic hydrogen absorbs energy, different hydrogen atoms absorb different amounts. Even though each atom has only one electron, so many atoms are present that all the allowable energy levels (orbits) are populated by electrons. When an electron drops from orbits with $n > 3$ (*second excited state*), the infrared series of spectral lines is produced *i.e.*, Paschen Series. The visible series arises from the photons emitted when an electron drops to the $n = 2$ orbit *i.e.*, Balmer Series (*first excited state*), and the ultraviolet series arises when these higher energy electrons drop to the $n = 1$ orbit (*ground state*).

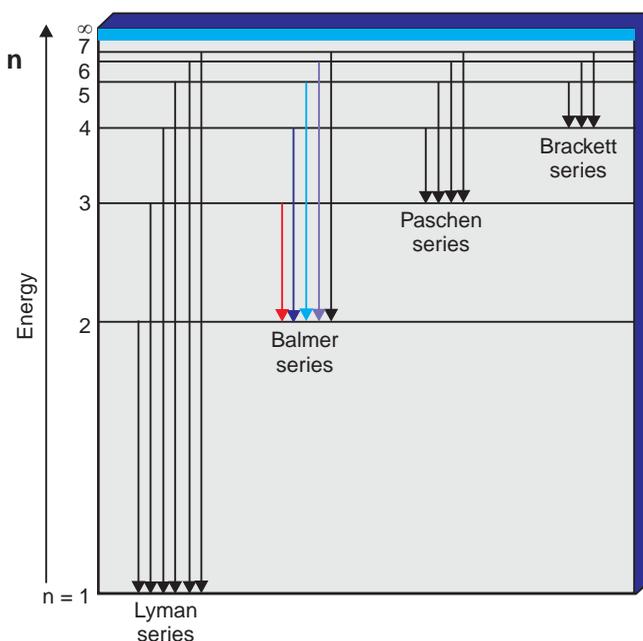


Figure 2.8 Representation of the observed spectral lines of the hydrogen atom.

Since a larger orbit radius means a higher atomic energy level, the farther the electron drops, the greater is the energy (higher ν , shorter λ) of the emitted photon. The spectral lines of hydrogen become closer and closer together in the short wavelength (*high energy*) region of each series because the difference in energy associated with the jump from initial state (n_i) to the final state (n_f) becomes smaller and smaller as the distance from the nucleus increases.

Having made this basic assumption, **Bohr** was then able to use classical physics to calculate properties of the hydrogen atom. In particular, he derived an equation for the electron energy (E_n). Each specified energy value (E_1, E_2, E_3, \dots) is called an **energy**

level of the atom. A very useful result from Bohr's work is an equation for calculating the energy levels of an atom,

$$E_n = -A \times \frac{Z^2}{n^2} \quad \dots(2.11)$$

where A is the **constant**, has a value of, $A = 2.18 \times 10^{-18}$ J. The number n is an integer called the **principal quantum number** ($n = 1, 2, 3, \dots$). Z is the charge of the nucleus. The negative sign in the equation appears because it is defined as zero energy when the electron is completely moved from the nucleus, i.e. $E_n = 0$ when $n = \infty$, so, $E_n < 0$ for any smaller n .

A can be expressed in terms of Rydberg constant as $A = hCR_H$, where R_H is Rydberg constant with a value of $1.0967 \times 10^7 M^{-1}$

For the H atom, $Z = 1$, so we have

$$E_n = -2.18 \times 10^{-18} \text{ J} \times \frac{1}{n^2} \quad \dots(2.12)$$

Therefore, the energy of the ground state $n = 1$ is -2.18×10^{-18} J. This equation is easily adapted to find the energy difference between any two levels:

$$\Delta E = hv = E_f - E_i = -2.18 \times 10^{-18} \text{ J} \times \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \quad \dots(2.13)$$

Exercise 2.5

1. The H atom and the Be^{3+} ion each have one electron. Does the Bohr model predict their spectra accurately? Would you expect their line spectra to be identical? Explain.
2. Calculate the energies of the states of the hydrogen atom with $n = 2$ and $n = 3$, and calculate the wavelength of the photon emitted by the atom when an electron makes a transition between these two states.
3. What is the wavelength of a photon emitted during a transition from the $n_i = 5$ state to the $n_f = 2$ state in the hydrogen atom?
4. How much energy, in kilojoules per mole, is released when an electron makes a transition from $n = 5$ to $n = 2$ in an hydrogen atom? Is this energy sufficient to break the **H-H bond** (436 kJ/mol is needed to break this bond)?

5. Five lines in the **H atom** spectrum have wavelengths (in Å):

- | | | | | | |
|---|--------|---|--------|---|--------|
| a | 1212.7 | b | 4340.5 | c | 4861.3 |
| d | 6562.8 | e | 10938d | | |

Three lines result from transition to $n_f = 2$ (*visible series*). The other two result from transitions in different series—one with $n_f = 1$, and the other with $n_f = 3$. Identify n_i for each line.

2.5.5 Limitations of the Bohr Theory

The **Bohr** model of an atom was successful in accounting for the spectral lines of **H-atom** which indicated that he was on the right track. Despite its great success in accounting for the spectral lines of the **H atom**, the Bohr model failed to predict the wavelengths of spectral lines of atoms more complicated than hydrogen, even that of helium, the next simplest element. It works beautifully for the hydrogen atom and for other one-electron species such as **He⁺** and **Li²⁺**. **Bohr** could not explain the further splitting of spectral lines in the hydrogen spectra on application of magnetic field and electric fields.

In the years following Bohr's development of a model for the hydrogen atom, the dual nature of radiant energy had become a familiar concept. Depending on the experimental circumstances, radiation might appear to have either a wave-like or particle-like (*photon*) character.

Is it possible that under proper circumstances, matter can behave as waves?

Louis de Broglie made a rather bold intuitive extension of this idea. Using **Einstein's** equation and **Planck's** equation for energy, he proposed that a particle with a mass, m , moving at a speed, v , has a wave nature consistent with a wavelength given by:

$$\lambda = \frac{h}{mv} \quad \dots(2.14)$$

where h is **Planck's constant**. The quantity mv for any object is called its momentum. The **matter waves** describe the wave characteristics of material particles.

Because **de Broglie's** hypothesis is perfectly general, any object of mass m and velocity v would give rise to a characteristic matter wave. However, it is easy to see from **Equation 2.14** that the wavelength associated with an object of ordinary size, such as a tennis ball is so tiny as to be completely out of the range of any possible observation. This is not so for electrons, because their mass is very small.

If de Broglie's concept is correct and all particles travel in waves, electrons should exhibit the typical wave-properties, diffraction and interference.

Electron particles with mass and charge create diffraction patterns, just as electromagnetic waves do. Although electrons do not have orbits of fixed radius, de Broglie thought that the energy levels of the atom are related to the wave nature of the electron.

The unsettling truth is that matter as well as energy show both behaviours: each possesses both “*properties*”. In some experiments, one property is observed, while in other experiments, the other property is observed. The distinction between a particle and a wave is only meaningful in the macroscopic world - it disappears at the atomic level. The distinction is in the minds and the limiting definitions of people, not in nature. This dual character of matter and energy is known as the **wave-particle duality**.

Example 2.6

1. Calculate the de Broglie wavelength of an electron that has a velocity of 1.00×10^6 m/s. (electron mass, $m_e = 9.11 \times 10^{-31}$ kg; $h = 6.626 \times 10^{-34}$ J s)
2. Calculate the frequency of the hydrogen line that corresponds to the transition of the electron from the $n = 4$ to the $n = 2$ state.

Solution:

$$1. \lambda = \frac{h}{m_e v} = \frac{6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}}{(9.11 \times 10^{-31} \text{ kg})(1.00 \times 10^6 \text{ m s}^{-1})}$$

$$= 7.27 \times 10^{-10} \text{ m}$$

$$2. \text{ We employ equation 2.13, } \Delta E = -R_H \times \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) = h\nu$$

$$\nu = -\frac{R_H}{h} \times \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

Substituting $n_i = 4$ and $n_f = 2$:

$$\nu = -\left(\frac{2.179 \times 10^{-18} \text{ J}}{6.626 \times 10^{-34} \text{ J s}} \right) \left(\frac{1}{4^2} - \frac{1}{2^2} \right)$$

$$v = - \left(\frac{2.179 \times 10^{-18} \text{ J}}{6.626 \times 10^{-34} \text{ J s}} \right) \times \left(-\frac{3}{16} \right) = 6.17 \times 10^{14} \text{ s}^{-1}$$

Exercise 2.6

1. What is the characteristic wavelength of an electron (in nm) that has a velocity of $5.97 \times 10^6 \text{ m s}^{-1}$ ($m_e = 9.11 \times 10^{-31} \text{ kg}$)?
2. Calculate the energy required for the ionization of an electron from the ground state of the hydrogen atom.
3. Calculate the wavelength (in nm) of a photon emitted when a hydrogen atom undergoes a transition from $n = 5$ to $n = 2$.

2.6 THE QUANTUM MECHANICAL MODEL OF THE ATOM

At the end of this section, you should be able to:

- state the Heisenberg uncertainty principle
- describe the significance of electron probability distribution;
- explain the quantum numbers n , l , m_l and m_s ;
- write all possible sets of quantum numbers of electrons in an atom; and
- describe the shape of orbitals that are designated by s , p , and d .

Activity 2.13



Make a group and discuss the following:

If particles have wavelike motion, why don't we observe that motion in the macroscopic world?

If electron possess particle nature it should be possible to locate electron. How can an electron be located?

Is there any wave associated with a moving elephant? Share your ideas with the rest of the class.

Bohr's model was very important because it introduced the idea of quantized energy states for electrons in atoms. This feature is incorporated in our current model of the atom, **the quantum mechanical model of the atom**.

The sophisticated mathematical description of atomic structure based on the wave properties of subatomic particles is called **wave mechanics or quantum mechanics**. Principally, the Austrian physicist **Erwin Schrödinger** developed a model of the hydrogen atom based on the wave nature of the electron in the late 1920s. Mathematical equations describing the nature of electron waves in atoms are fundamental to the modern picture of the atom. The wave equations that are acceptable solutions to the Schrödinger equation are called **wave functions**. To obtain one of these acceptable solutions, we must assign integral values called **quantum numbers** to three quantities in the wave equation. This requirement is similar to the requirement for an integral value of n in the **Bohr** equation for the hydrogen atom.

In contrast to the precise planetary orbits of the **Bohr atom**, the wave mechanics picture of the hydrogen atom is less certain. Instead of determining the exact location of the electron, we can only speak of the probability of the electron being found in certain regions of the atom. Or, if we adopt the view that the electron is just a cloud of negative electric charge, we can only speak of the charge densities in various parts of the atom.

2.6.1 The Heisenberg's Principle

Discovery of the wave properties of matter raised a new and very interesting question. If a subatomic particle can exhibit the properties of a wave, is it possible to say precisely where that particle is located? One can hardly speak of the precise location of a wave. A wave, as a whole, extends in space. Its location is therefore not defined precisely. To describe the limitation (problem) of trying to locate a subatomic particle that behaves like a wave, Heisenberg formulated what is known as the **Heisenberg uncertainty principle**, which states that it is impossible to know simultaneously both the momentum p ($p = mv$) and the position of a particle with certainty. Mathematically, the Heisenberg's uncertainty principle is given as

$$\Delta x \times \Delta p \geq \frac{h}{4\pi}$$

where Δx and Δp are the uncertainties in measuring the position and momentum, respectively.

The above equation says that if we measure the momentum of a particle precisely (*i.e.* if Δp is made very small), then the position will be correspondingly less precise

(i.e., Δx will become larger). Similarly, if the position of a particle is known more precisely, then its momentum will be less precise.

The German physicist **Werner Heisenberg** in 1927 concluded that there is a fundamental limitation to know both the location and the momentum of a particle simultaneously, which is due to dual behaviour of matter and radiation. Just as in the case of quantum effects, the limitation becomes important only when we deal with matter at the subatomic level, that is, with mass as small as that of an electron.

2.6.2 Quantum Numbers

An atomic orbital is specified first by three quantum numbers that are associated respectively, with the orbital's size (*energy*), shape, orientation and, later, independent of these three quantum numbers, the electron spin. The first three sets of quantum numbers have a hierarchical relationship: the size-related number limits the shape-related number, the shape-related number in turn limits the orientation-related number.

Three among the four quantum numbers characterize the orbitals in the atom. That is, they describe the orbital or the space the electron is supposed to occupy. The fourth quantum number is used to describe the spin of the electrons that occupy the orbitals.

The four quantum numbers are:

1. The **principal quantum number** (n) is a positive integer having values $n = 1, 2, 3, \dots$. It gives the following information:
 - (i) Relative size of the orbital or the relative distance of the electron from the nucleus. Size of orbital increases with the increase of principal quantum number n .
 - (ii) Energy of the orbital. Higher the n value, greater is the energy. For example: when the electron occupies an orbital with $n = 1$, the hydrogen atom is in its ground state and has lower energy than when the electron occupies an orbital with $n = 2$ (*first excited state*).
 - (iii) Maximum number of electrons present in any shell (given by the formula $2n^2$).
2. The **azimuthal quantum number** (l) is also known as angular momentum or subsidiary quantum number. It is an integer having values from 0 to $(n - 1)$. For an orbital with $n = 1$, l can have a value only of 0. For orbitals with $n = 2$, l can have a value of 0 or 1; for those with $n = 3$, l can be 0, 1 or 2; etc. So, the number of possible l values equals the value of n . For a given value of n , the maximum possible value of l is $(n - 1)$. The azimuthal quantum number gives the following information:

- (i) Number of subshell present within any shell.
- (ii) It describes the shape of the orbital and is sometimes also called the orbital-shape quantum number.
3. The **magnetic quantum number (m_l)** is also known as the orbital-orientation quantum number. It is an integer having values from $-l$ through 0 to $+l$. The possible values of an orbital's magnetic quantum number are set by its angular momentum quantum number (*that is, l determines m_l*). An orbital with $l = 0$ can have only $m_l = 0$. However an orbital with $l = 1$, can have m_l value of -1 , 0, or $+1$; thus there are three possible orbitals with $l = 1$ each with its own spatial orientation. *The number of possible m_l values or orbitals for a given l value is $(2l + 1)$. It prescribes the orientation of the orbital in the three-dimensional space about the nucleus.*
4. The **electron spin quantum number (m_s)** has only two possible values, $+\frac{1}{2}$ (represented by the arrow, \uparrow) and $-\frac{1}{2}$ (represented by the arrow \downarrow). The name electron spin quantum suggests that electrons have a spinning motion. However, there is no way to attach a precise physical reality to electron spin.

The quantum numbers specify the energy states of the atom.

- The atom's energy **levels** or **shells** are given by the n value.
- The atom's **sublevels** or **subshells** are given by the n and l values. Each level contains sublevels that designate the shape of the orbital.
- The atom's **orbitals** are specified by the n , l and m_l values. Thus, the three quantum numbers that describe an orbital express its size (energy), shape and spatial orientation. Each sublevel is designated by a letter:

$l = 0$, is an ***s*** sublevel

$l = 1$, is a ***p*** sublevel

$l = 2$, is a ***d*** sublevel

$l = 3$, is a ***f*** sublevel

The letters ***s***, ***p***, ***d***, and ***f*** are derived from the names of spectroscopic lines: ***s***, sharp; ***p***, principal; ***d***, diffuse; and ***f***, fundamental. Sublevels are named by joining the n value and the letter designation. For example, the sublevel (subshell) with $n = 2$, $l = 0$ is

called the $2s$ sublevel; the only orbital in this sublevel has $n = 2$, $l = 0$ and $m_l = 0$. A sublevel with $n = 3$, $l = 1$, is a $3p$ sublevel. It has three possible orbitals: one with $n = 3$, $l = 1$ and $m_l = -1$; another with $n = 3$, $l = 1$ and $m_l = 0$ and the third $n = 3$, $l = 1$, and $m_l = +1$.

For a given principal quantum number, n , the total number of orbitals is determined as:

Number of orbitals = n^2 in a shell.

Similarly, the number of orbitals in each subshell is determined as:

Number of orbitals in a subshell = $2l+1$.

Example 2.7

- What values of the angular momentum quantum number (l) and magnetic quantum number (m_l) are allowed for a principal quantum number (n) of 3? How many orbitals are allowed for $n = 3$?
- Give the name, magnetic quantum numbers, and numbers of orbitals for each sublevel with the following quantum numbers:
 - $n = 3$, $l = 2$
 - $n = 2$, $l = 0$
 - $n = 5$, $l = 1$
 - $n = 4$, $l = 3$
- What is wrong with each of the following quantum number designations and/or sublevel names?

	n	l	m_l	Name
a	1	2	0	$1p$
b	4	3	+1	$4d$
c	3	1	-2	$3p$

Solution:

- Determining l values:

For $n = 3$, $l = 0, 1, 2$

Determining m_l for each l value:

For $l = 0$, $m_l = 0$

For $l = 1$, $m_l = -1, 0, +1$

For $l = 2$, $m_l = -2, -1, 0, +1, +2$

Number of orbitals = n^2 , n is 3 and $n^2 = 3^2 = 9$ orbitals

3s : 1 orbital

3p : 3 orbitals

3d : 5 orbitals

Total = 9 orbitals

2.

	n	l	sublevel name	Possible m_l values	number of orbitals
a	3	2	3d	-2, -1, 0, +1, +2	5
b	2	0	2s	0	1
c	5	1	5p	-1, 0, +1	3
d	4	3	4f	-3, -2, -1, 0, +1, +2, +3	7

3. a A sublevel of $n = 1$ can have only $l = 0$, not $l = 1$. The only possible sub shell is 1s.

b A sublevel with $l = 3$ is an f sublevel, not a d sublevel. The sublevel name should be 4f.

c A sublevel with $l = 1$ can have only m_l of -1, 0, +1, not -2.

Exercise 2.7

1. Give the sublevel notation for each of the following sets of quantum numbers.

a $n = 3, l = 2$

c $n = 4, l = 1$

b $n = 2, l = 0$

d $n = 4, l = 3$

2. Indicate whether each of the following is a permissible set of quantum numbers. If the set is not permissible, state why it is not.

a $n = 3, l = 1, m_l = +2$

d $n = 0, l = 0, m_l = 0$

b $n = 4, l = 3, m_l = -3$

e $n = 3, l = 3, m_l = -3$

c $n = 3, l = 2, m_l = -2$

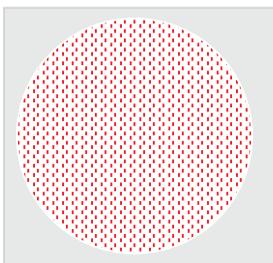
3. Consider the electronic configuration of an atom:

a What are the n, l and m_l quantum numbers corresponding to the 3s orbital?

- b List all the possible quantum number values for an orbital in the $5f$ subshell.
 - c In which specific subshell will an electron be present if the quantum numbers $n = 3$, $l = 1$, and $m_l = -1$?
4. Which of the quantum numbers relates to the electron only? Which relate (s) to the orbital?

2.6.3 Shapes of Atomic Orbitals

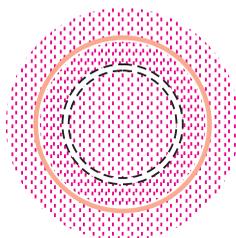
The wave mechanical picture of an electron in an s -orbital looks like a ball made of a raised soft hairy rug (*that is the orbital has spherical symmetry*).



1s

Figure 2.9 a Geometric shape of 1s-orbital.

The $2s$ -orbital also has spherical symmetry.



2s

Figure 2.9 b Geometric shape of 2s-orbital.

Higher level s -orbitals ($3s$, $4s$, ...) have similar overall shapes, that is, spherical symmetry. The second principal energy level ($n = 2$) consists of four different orbitals. One having $l = 0$, is the $2s$ orbital just described. The other three orbitals of the $n = 2$ level have $l = 1$. They are described by dumb bell-shaped regions (**Figure 2.10**).

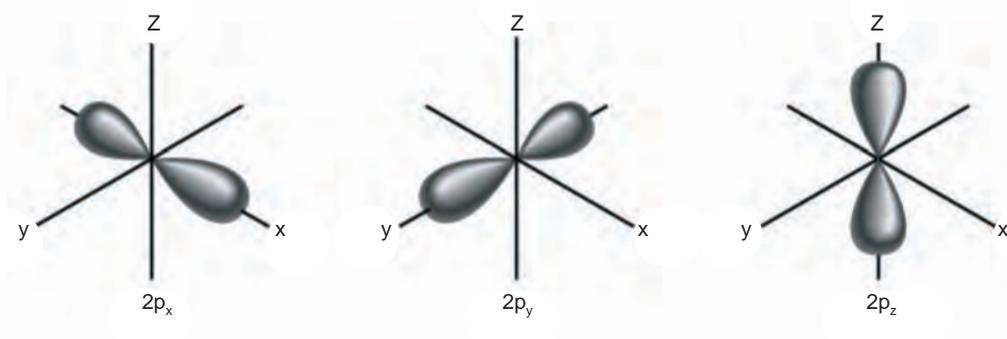


Figure 2.10 The three p -orbitals.

Two lobes lie along a line with the nucleus at the centre between the lobes. The different $2p$ orbitals are referred to as the $2p_x$, $2p_y$ and $2p_z$, orbitals, because they are perpendicular to one another and can be drawn along the x , y , and z coordinate axes (see Figure 2.10). Higher - level p -orbitals (that is $3p$, $4p$, ...) have similar overall shapes.

The third principal energy level ($n = 3$) is divided in to three sublevels: one $3s$ orbital, three $3p$ orbitals, and five $3d$ orbitals. The d orbitals are described in Figure 2.11. They are much more complex in shape than p orbitals.

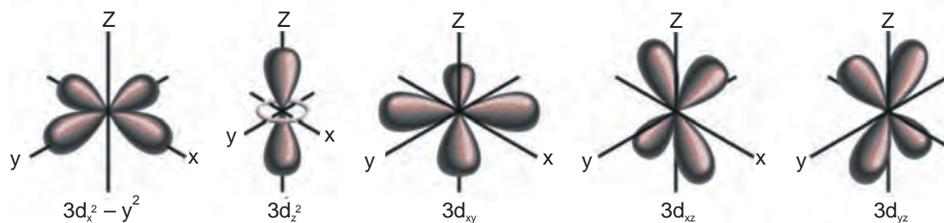


Figure 2.11 The five d orbitals.

The fourth principal energy level is divided in to four sublevels or subshells. There are one $4s$ subshell with one orbital, a $4p$ subshell with three orbitals, a $4d$ subshell with five orbitals, and a $4f$ subshell with seven orbitals. The f orbitals have complex shapes.

2.7 ELECTRONIC CONFIGURATIONS AND ORBITAL DIAGRAMS

At the end of this section, you should be able to:

- explain the Aufbau principle;
- explain the Pauli exclusion principle;
- explain Hund's rule; and
- write ground-state electronic configurations of multi-electron atoms.

Activity 2.14

Form a group and discuss the given questions. After discussion share your ideas with rest of the class.

1. Why are there never more than two electrons in an atomic orbital?
2. When there are two electrons in an orbital, why do these always have opposing spins?
3. Why are orbitals occupied singly first before the pairing of electrons occurs?
4. Why do the electrons in singly-occupied orbitals have parallel spins?

Several questions arise when you look carefully at the electron configuration of an atom. To answer these questions, it is necessary to know the basic principles that govern the distribution of electrons among atomic orbitals. The electron configuration for any atom follows the following three principles:

Aufbau Principle. In general, electrons occupy the lowest-energy orbital available before entering the higher energy orbital.

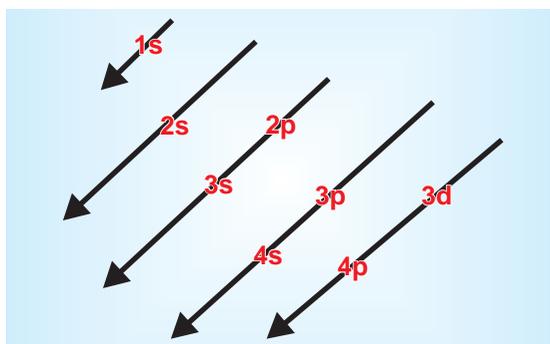


Figure: 2.12 Increasing order of filling orbitals.

Hund's Principle. Equal energy orbitals (degenerate orbitals) are each occupied by a single electron before the second electrons of opposite spin enters the orbital. In other words, each of the three $2p$ orbitals ($2p_x$, $2p_y$ and $2p_z$) will hold a single electron before any of them receives a second electron.

Pauli's Exclusion Principle. No two electrons can have the same four quantum numbers. *i.e.* they must differ in at least one of the four quantum numbers.

Note: *aufbau* is a German word, which means building-up, while *Pauli* and *Hund* are the names of scientists.

2.7.1 Ground State Electronic Configuration of the Elements

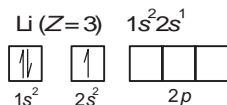
The electronic configuration of an atom describes the distribution of the electrons among atomic orbitals in the atom. Two general methods are used to denote electron configurations. The subshell (sublevel) notation uses numbers to designate the principal energy levels or principal quantum number and the letters *s*, *p*, *d* and *f* to identify the sublevels or subshells. A superscript number following a letter indicates the number of electrons in the designated subshell (*e*). The designation can be written as nl^e . The electron configurations for hydrogen (H; $Z = 1$), helium (He; $Z = 2$) and lithium (Li; $Z = 3$) are $1s^1$, $1s^2$, $1s^2 2s^1$ respectively.

Activity 2.15



Form a group and discuss the reason why the notation nl^e does not include electron spin quantum numbers.

The other way to present this information is through an orbital diagram, which consists of a box (or circle, or just a line) for each orbital available in a given energy level, grouped by sublevel, with an arrow indicating the electron's presence and its direction of spin. Traditionally \uparrow , $m_s = +\frac{1}{2}$ and \downarrow , $m_s = -\frac{1}{2}$, but these are arbitrary, so it is necessary only to be consistent. The orbital diagrams for the first three elements are:

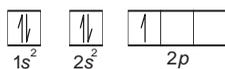
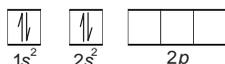
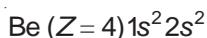


In Li, the $2s$ orbital is only half-filled, so the fourth electron of beryllium fills in with its spin paired: $n = 2$, $l = 0$, $m_l = 0$, $m_s = +\frac{1}{2}$.

For beryllium, the $2s$ sublevel is filled, and the next lower energy sublevel is the $2p$. The three orbitals in the $2p$ sublevel have equal energy (same n and l values), which means that the fifth electron of boron can go into any one of the $2p$ orbitals. Since a p sublevel has $l = 1$, the m_l (orientation) values can be -1 , 0 , or $+1$. For

convenience, the boxes are labelled from left to right, $-1, 0, +1$, and assume it enters the $m_l = -1$ orbital: $n = 2, l = 1, m_l = -1$, and $m_s = +\frac{1}{2}$.

For our purposes here, the designation of $m_l = -1$ for this boron electron (*and the box we chose to label -1 , in which we placed the arrow*) is arbitrary. The $2p$ orbital's have equal energy and differ only in their orientation.



Activity 2.16



Form a group and discuss the following question:

1. What does each box in an orbital diagram represent?
2. Which quantity is represented by the direction (either up or down) of the half arrows in an orbital diagram?
3. Is Hund's rule used in deriving the electron configuration of beryllium?

Share your ideas with the rest of the class.

Exercise 2.8

1. Write the ground-state electron configuration for potassium
2. Write the electron configuration of an element with atomic number 10.
3. Write the electron configuration of an element with ($Z = 20$ and $A = 40 m_u$).
4. Draw the orbital diagrams for the valence electrons of the atoms: C, N, O, F and Ne and determine the number of unpaired electrons for each atom.

Activity 2.17



Form a group and complete the electron configuration for the period 3 elements of periodic table.

Atomic Number	Element	Orbital diagram	Full electron configuration	Condensed electron configuration
11	Na	$3s$ $3p$ 	$[1s^2 2s^2 2p^6 3s^1]$	$[\text{Ne}] 3s^1$
12	Mg		$[1s^2 2s^2 2p^6 3s^2]$	$[\text{Ne}] 3s^2$
13	Al			
14	Si			
15	P			
16	S			
17	Cl			
18	Ar			

Group Assignment



Form a group and discuss on the following questions. Share your ideas with the rest of the class.

Select any ten elements. For each of your chosen elements, prepare an index card. In the upper left-hand corner of the card, place a box like the one on the periodic table, name, and the atomic mass of the element.

Do some research on the element and write a few sentences about its properties and uses. Attach either an object or a picture to the card that shows its properties. For example for carbon, you could attach a pencil lead. For mercury, you might have a picture of a thermometer. Make a large periodic table, using your cards. Show the result to your classmates.

Table 2.2 Maximum capacities of subshells and principal shells

n	1	2	3	4
l	0	0, 1	0, 1, 2	0, 1, 2, 3
Subshell designation	s	s, p	s, p, d	s, p, d, f
Orbital in subshell	1	1, 3	1, 3, 5	1, 3, 5, 7
Subshell capacity	2	2, 6	2, 6, 10	2, 6, 10, 14
Principal shell capacity	2	8	18	32

Activity 2.18



Form a group and complete the following table for the first row transition metals on your note book

Atomic Number	Element	Orbital Diagram	Electron configuration (ground state)	Condensed Electron configuration
21	Sc			
22	Ti			
23	V			
24	Cr			
25	Mn			
26	Fe			
27	Co			
28	Ni			
29	Cu			

Examine the electron configurations of chromium and copper. The expected configuration, those based on the **aufbau** principle, is not the ones observed through the emission spectra and magnetic properties of the elements.

	Expected	Observed
Cr ($Z = 24$)	$[\text{Ar}] 4s^2 3d^4$	$[\text{Ar}] 4s^1 3d^5$
Cu ($Z = 29$)	$[\text{Ar}] 4s^2 3d^9$	$[\text{Ar}] 4s^1 3d^{10}$

The reason for these exceptions to the **aufbau** principle are not completely understood, but it seems that the half-filled $3d$ subshell of chromium ($3d^5$) and the fully filled $3d$ subshell of copper ($3d^{10}$) lends a special stability to the electron configurations. Apparently, having a half-filled $4s$ subshell and a half-filled $3d$ subshell gives a lower energy state for a Cr atom than having a filled $4s$ subshell.

Because there is little difference between the $4s$ and $3d$ orbital energies, expected and observed electron configurations are quite close in energy.

At higher principal quantum numbers, the energy difference between certain subshell is even smaller than that between the $3d$ and $4s$ subshells. As a result, there are still more exceptions to the Aufabu principle among the heavier transition elements.

Exercise 2.9

1. Write the electron configuration for the Co^{3+} , Cl^- , Al^+ , Cr , As^- , and Cu .
2. Write the electron configuration and the orbital diagram of the first excited state of sodium. (Hint: The outermost electron is excited).
3. What is the electron capacity of the n^{th} energy level? What is the capacity of the fourth energy level?

2.8 ELECTRONIC CONFIGURATIONS AND THE PERIODIC TABLE OF THE ELEMENTS

At the end of this section, you should be able to:

- correlate the electron configuration of elements with the periodicity of elements;
- give a reasonable explanation for shape of the periodic table;
- classify elements as representative, transition and inner-transition elements;
- explain the general trends in atomic radius, ionization energy, electron affinity, electronegativity; and metallic character of elements within a period and group of the periodic table; and
- write the advantages of the periodic classification of elements.

2.8.1 The Modern Periodic Table

Activity 2.19



Form a group and assign codes to the first twenty elements in the periodic table by the letters from a to t .

The following are some of the properties of the coded elements:

1. b , g , j , o , p , s and t are the only metals.
2. p and o form amphoteric hydroxides.

3. *b, j, s* and *t* show characteristics colour of the flame test for their chlorides.
4. *a* and *r* can form covalent hydrides of empirical formula XH_2 , the hydroxides of *l* is a mild reducing agent, while that of *e* is not.
5. *i, c* and *q* do not form oxides, hydrides or chlorides; their boiling points are -260°C , -246°C and -186°C respectively.
6. The dissociation energy of molecular *a* is smaller than that of *r*.
7. *f* and *k* form oxides of empirical formula X_2O_5 . On hydrolysis, the oxide of *f* yields a much more acidic solution than that of *k*.
8. *n* forms an oxide which has high melting point and is an important raw material of glass.

Identify the coded elements and place them in the correct positions in the periodic table, and giving reasons for doing so.

9. Discuss on the drawbacks of Mendeleev's periodic table and the need for the modern one.

You have studied in Grade 9 that the modern periodic table of the elements is one of the great classifying schemes in science and has become an indispensable tool to chemists.

Periodic relationships can be summarized by the general statement called periodic law. In its modern form, the periodic law states that certain sets of physical and chemical properties recur at regular intervals (periodically) when the elements are arranged according to increasing atomic number.

Activity 2.20



Form a group and discuss the following questions:

1. Do the elements having last electron in *s* orbital form a family?
2. Is there some variations in physical and chemical properties of these elements?
3. Why are some elements placed out of the main body of periodic table?
4. Where are isotopes located?
5. Can the properties of a compound formed be predicted on the basis of the location of its combining elements in the periodic table?

2.8.2 Classification of the Elements

Activity 2.21



Form a group and discuss the following questions:

1. How can the 90 or so naturally occurring elements be classified?
2. Are they all separate and distinct?
3. Is there a need to learn about 90 different elements as well as some thousands of their compounds?
4. What could be the basis of classifying elements?

Elements are placed in the periodic table in accordance to valence electron entering the orbital of lowest energy. There are 18 groups and 7 periods in the modern periodic table.

Representative or main group elements: These consist of all *s*- and *p*-block elements. The chemical properties of the representative elements are determined by the number of valence electrons in their atoms.

Transition elements: These are *d*-block elements. There are four series of transitional elements, *3d*, *4d*, *5d* and *6d* depending on the energy levels of *d*-orbitals.

Inner transition elements: These are the *f*-block elements. There are two series of *f*-block elements, *4f* and *5f* series called **lanthanides** and **actinides**, respectively. The periodic table is unable to include the inner transition elements in its main frame. They have been allotted the same single place in the periodic table though their electronic configurations are not identical. Besides, the variation in their properties is not much.

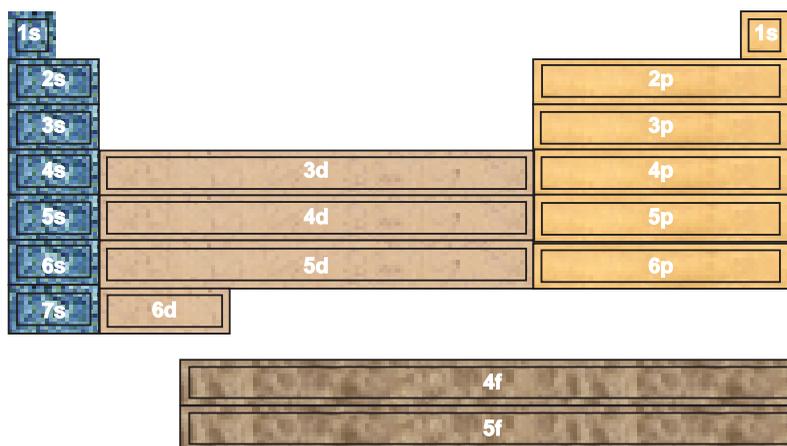


Figure 2.13 Classification of elements according to the type of subshells being filled.

2.8.3 Periodic Properties

Activity 2.22



Form a group and discuss the following questions:

1. How does size of elements vary across a period?
2. Why are elements of Group 1 and 2 called metals, while those of Group 17 non-metals?
3. Why elements of Group 18 are least reactive?
4. Define shielding effect and effective nuclear charge.

Some physical properties, such as thermal and electrical conductivity, density, and hardness are displayed only by bulk matter, that is, by large aggregations of atoms. In this section you will examine some periodic atomic properties like atomic radii, ionization energies, electron affinities, electronegativity, and metallic character.

Atomic Size (Atomic Radii)

Activity 2.23



Form a group and discuss on the following questions:

1. Why the sizes of atoms do not increase uniformly with increasing atomic number?
2. Why the difference in atomic radius between the elements $Z = 11$ (Na; 186 pm) and $Z = 12$ (Mg; 160 pm) is so large, where as between $Z = 24$ (Cr; 125 pm) and $Z = 25$ (Mn; 124 pm) the difference is negligible.
3. In which location in the periodic table would you expect to find the elements having the largest atoms? Explain.
4. Why isoelectronic ions do not have the same ionic radii?
5. Why does the quantum-mechanical description of multi-electron atoms make it difficult to define the term atomic radius?
6. How do the sizes of atoms change as we move:
 - a from left to right in a row in the periodic table?
 - b from top to bottom in a group in the periodic table? Explain.
7.
 - a Why does the He atom have a smaller radius than the H atom?
 - b Why is the He atom smaller than the Ne atom? Explain.

Share your ideas with the rest of the class.

Exact size of an isolated atom cannot be measured because its outermost electrons have a chance of being found at relatively large distances from the nucleus. What can be measured is the distance between the nuclei of two adjacent atoms, and can derive a property called the **atomic radius** from this distance.

One of the most common methods to determine atomic radius is to assume that atoms are spheres that touch each other when they are bonded together (Figure 2.14).

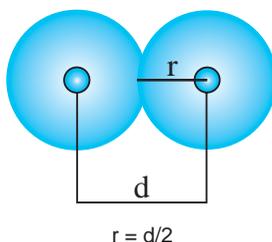


Figure 2.14 Atomic radius for a diatomic molecule.

Figure 2.14 depicts a way to define the radius of an atom using the distance d between the nuclei of two atoms on an element in a molecule. It is assumed that the atom is a hard sphere with a radius equal to half the bond distance, $1/2d$.

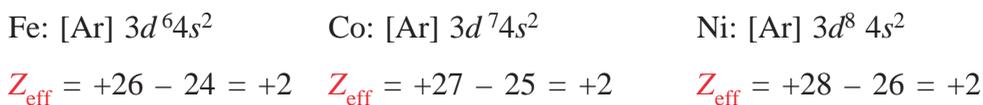
For example, the distance between iodine atoms in an iodine molecule I_2 , (I—I) is 2.66 \AA , so we can define the radius of an iodine atom to be half this distance: 1.33 \AA ; similarly, the distance between iron atoms (Fe) in iron metal is 2.48 \AA . Therefore, each iron atom in the metal has a radius of 1.24 \AA .

Within a vertical group of the periodic table, each succeeding member has one more principal shell occupied by electrons, thus, atomic radii increases from top to bottom within a group of the periodic table. To describe how atomic radii vary within a period of the periodic table, it is helpful to employ a new concept. The **effective nuclear charge** (Z_{eff}) acting on an electron is the actual nuclear charge less the screening effect of other electrons in the atom.

For example, consider a sodium atom. If the $3s$ valence electron were at all times completely outside the region in which the ten electrons of the neon core ($1s^2 2s^2 2p^6$) are found, the $3s$ electron would be perfectly screened or shielded from the positively charged nucleus; it would experience an attraction to a net positive charge of only $+11 - 10 = +1$. The corresponding situation for magnesium atom would be that of the two $3s$ electrons outside the neon core and a net positive charge of $+12 - 10 = +2$ acting on each of the $3s$ electrons. Similarly it is found that the net positive charge acting on the valence electrons would progressively increase across the third period.

The effective nuclear charge increases from left to right in a period of representative elements in the periodic table. Because the effective nuclear charge increases, valence electrons are pulled in toward the nucleus and held more tightly. Atomic radii of the A-group elements (*main group element*) tend to decrease from left to right in a period of the periodic table.

The restriction to A-group elements is an important one. In a series of B-group elements (*transition elements*), electrons enter an inner electron shell, not the valence shell. In this process, the effective nuclear charge remains essentially constant instead of increasing. For example, compare the effective nuclear charges, Z_{eff} , of iron, cobalt, and nickel.



Because the effective nuclear charges are almost the same, we conclude that the radii should also be almost the same. The actual values are 124, 124, and 125 pm, respectively, for Fe, Co and Ni.

Exercise 2.10

1. Arrange the following atoms in order of increasing atomic radius: F, P, S, As.
2. The atoms and ions Na, Mg^+ , Al^{2+} and Si^{3+} all have the same number of electrons. For which of these will the effective nuclear charge acting on the outer most electron be the smallest? For which will it be greatest? Explain

Ionization Energy (IE)

Activity 2.24



Form a group and discuss on the following:

1. The trend in successive ionization energies as electrons are removed one at a time from an aluminum atom. Why is there a big jump between IE_3 and IE_4 ?
2. Why does sulphur have a lower first ionization energy than phosphorus?
3. The second ionization energy of lithium is much greater than that of beryllium. Explain.
4. For strontium, which quantity will be greater, the difference between IE_1 and IE_2 or the difference between IE_2 and IE_3 ?

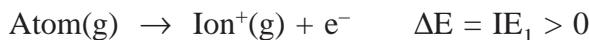
Share your ideas with the rest of the class.

The ionization energy (IE) is the amount of energy required to remove the outermost electron in an isolated gaseous atom or ion.

Multi-electron atoms can lose more than one electron, so the ionization energies required to remove each electron are numbered in sequence from the ground-state atom. Consider for example the boron atom, which has five electrons, two in an inner core ($1s^2$) and three valence electrons ($2s^22p^1$). The five ionization steps and their ionization energies, IE_1 through IE_5 , are:



The first ionization energy (IE_1) is the energy needed to remove an electron from the highest occupied sublevel of the gaseous atom.



The second ionization energy (IE_2) removes the second electron. Since the electron is being pulled away from a positively charged ion, IE_2 is always larger than IE_1 :



The first ionization energy is a key factor in an element's chemical reactivity because, atoms with a low IE_1 tend to form **cations** during reactions, whereas those with a high IE_1 , (*except the noble gases*) often form **anions**.

The elements exhibit a periodic change in first ionization energy. There is a roughly inverse relationship between IE_1 and atomic size.

The only significant deviations from this pattern occur in Group IIIA. Although IE_1 decreases as expected from boron (B) to aluminium (Al), no decrease occurs for the rest of the group. As with the atomic size trends, filling the d sublevels of the intervening transition elements in period 4, 5 and 6 causes a greater than expected Z_{eff} , which holds the outer electrons tightly in this larger IIIA members.

As we move across a period, Z_{eff} generally increases so atomic radii become smaller. As a result, the attraction between the nucleus and the outer electrons increases, so an electron becomes more difficult to remove. In general, ionization energy **increases**

across a period; it is easier to remove an electron from an alkali metal than from a noble gas.

There are two *exceptions* in the otherwise smooth increase in ionization energy trend, which occur at Groups IIIA and VIA in period 2 (at B and at O) and in period 3 (at Al and at S). The *first deviation* occurs because the np sublevel is higher in energy than the ns sublevel. The *second deviation* occurs because the np^4 electron occupies the same orbital as another np electron, the first such pairing, so electron repulsions raise the orbital energy. Removing this electron relieves the repulsions and leaves a stable, *half-filled* np sublevel; thus the fourth p electron is pulled off more easily.

For the representative elements, removing a core electron requires much more energy compared to removing a valence electron.

Group Assignment



Form a group and discuss on the following questions. Share your ideas with the rest of the class.

Plot the first ionization energy of the first 18 elements against atomic number. On the same graph, make a plot of the second ionization energy of the first 18 elements. Consult chemistry books for the required information.

Is second IE less than first IE for an element?

Compare the second IE values of sodium and magnesium.

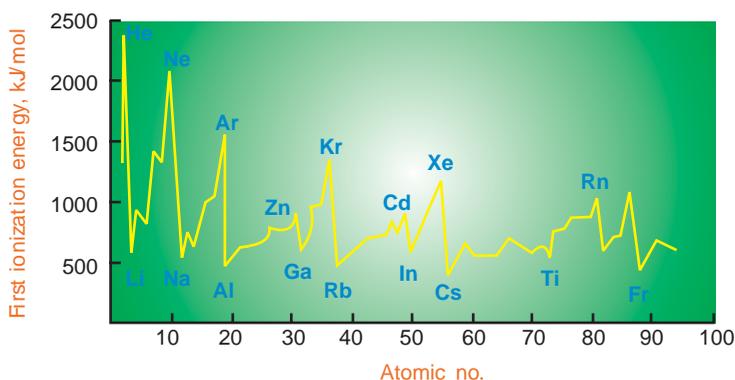


Figure 2.15 A graph of first ionization energy versus atomic number.

Example 2.8

Explain the irregularities in the trends across periods.

1. Boron, with $Z = 5$ and the electron configuration $1s^2 2s^2 2p^1$, has smaller first ionization energy (801 kJ mol^{-1}) than does beryllium ($Z = 4$), which has the electron configuration $1s^2 2s^2$ and $IE_1 = 900 \text{ kJ mol}^{-1}$.

Solution: The $2p$ electron of boron is at a higher energy than the $2s$ electron of beryllium and is therefore easier to remove. This kind of discontinuity occurs generally in proceeding from a Group IIA or Group IIB elements to a Group IIIA element.

2. IE_1 values for nitrogen ($[\text{He}]2s^2 2p_x^1 p_y^1 p_z^1$) is higher than oxygen ($[\text{He}]2s^2 2p_x^2 p_y^1 p_z^1$).

Solution: The observed IE_1 values are 1314 kJ/mol for oxygen and 1402 kJ/mol for nitrogen. In $2p$ orbitals, the repulsion between the paired electrons in the $2p_x$ orbital of oxygen makes the removal of one of those electrons easier to accomplish than the removal of an unpaired electron from the half-occupied $2p$ orbital of nitrogen.

Table 2.3 Ionization energies of some selected elements (in kJ mol^{-1})

	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
	Li	Be	B	C	N	O	F	Ne
IE_1	520	900	801	1086	1402	1314	1681	2081
IE_2	7298	1757						
	Na	Mg						
IE_1	496	738						
IE_2	4562	1451						
	K	Ca						
IE_1	419	590						
IE_2	3059	1145						
	Rb	Sr						
IE_1	403	550						
IE_2	2633	1064						

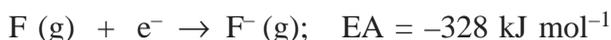
Exercise 2.11

1. Use the third period of the periodic table as an example to illustrate change in the first ionization energies of the elements as we move from left to right. Explain the trend.
2. Ionization energy measurements are usually carried out with atoms in the gaseous state. Why?
3. The first and second ionization energies of K are 419 kJ mol^{-1} and 3052 kJ mol^{-1} and those of Ca are 590 kJ mol^{-1} and 1145 kJ mol^{-1} , respectively, compare their value and comment on the differences.
4. Why does potassium have a lower first ionization than lithium?
5. Based on their positions in the periodic table, predict which atom of the following pairs will have larger first ionization energy:
 - a Ga, Ge
 - b Br, Sb
 - c K, Cr
 - d Mg, Sr
 - e O, Ne

Electron Affinity (EA)

Ionization energy refers to the process of forming a gaseous positive ion from a gaseous atom. The corresponding atomic property for the formation of a gaseous negative ion is electron affinity, the energy change that occurs when an electron is added to a gaseous atom or ion.

An electron approaching a neutral atom experiences an attraction for the positively charged nucleus. Repulsion of the incoming electron by electrons already present in the atom tends to offset this attraction. Still, in many cases the incoming electron is absorbed by the atom and energy is evolved as in the process:



When a fluorine atom gains an electron, energy is given off. The process is exothermic and the electron affinity is therefore a negative quantity.

Activity 2.25



Form a group and discuss on the following:

1. By using electron configurations, explain why the electron affinity of F is negative value whereas the electron affinity of Ne is a positive value.
2. Which group of elements has electron affinities with the largest negative values? Explain why.
3. Silicon has an electron affinity of -134 kJ/mol . The electron affinity of phosphorus is -72 kJ/mol . Give a plausible reason for this difference.

Share your ideas with the rest of the class.

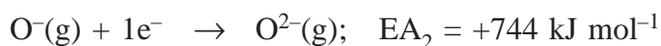
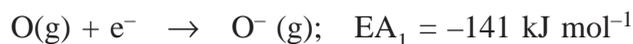
Table 2.4 Electron affinities of some selected elements (in kJ mol⁻¹)

IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
Li	Be	B	C	N	O	F	Ne
-60	0	-27	-154	-7	-141	-328	-
Na	Mg	Al	Si	P	S	Cl	
-53	0	-44	-20	-200	-200	-349	
K	Ca	Ga	Ge	As	Se	Br	
-48	0	-36	-116	-195	-195	324	
Rb	Sr	In	Sn	Sb	Te	I	
-46	0	-34	-121	-190	-190	-295	
Cs				Bi	Po	At	
-47				-183	-270	-270	

Table 2.4 lists several electron affinity values but we observe fewer clear-cut trends and more irregularities than in the Table 2.3, listing the ionization energies of some elements. The given information suggest a rough correlation between electron affinity and atomic size. Smaller an atom, more negative is its electron affinity. The smaller the atom the closer an added electron can approach the atomic nucleus and the more strongly it is attracted to the nucleus. This certainly seems to be the case for the Group IA elements, for Group VIA from **S** to **Po** and for Group VIIA from **Cl** to **At**. The first row elements present some problems. The electron affinity of **O** is not as negative as that of **S**, nor is that of **F** as negative as that of **Cl**. Here it may be that electron repulsions in the small compact atoms keep the added electron from being tightly bound as we might expect.

In most cases, the added electron goes into an energy sublevel that is already partly filled. For the Group IIA and VIIIA atoms, however, the added electron would be required to enter a significantly higher energy level, the *np* level for the Group IIA atoms and the *s* level for the next principal level for the Group VIIIA atoms. In these cases a stable anion does not form.

As the stepwise loss of multiple electrons in the formation of positive ions has been described, in the same manner, the stepwise addition of electrons in anion formation can be described, and we can write a separate electron affinity for each step. For oxygen atom, it can be written as:



Activity 2.26

Form a group and discuss the possible reasons why EA_2 is a positive quantity for oxygen atom.

Share your ideas with the rest of the class.

Electronegativity

Atoms of the elements in the upper right of the periodic table (small, non-metal atoms) attract bonding electrons most strongly. Therefore, they have the greatest electronegativities. Atoms of the elements to the left side of the table (large, metal atoms) have a weaker hold on electrons. They have the smallest electronegativities. On an electronegativity scale devised by Linus Pauling, the most non-metallic and hence most electronegative element, fluorine, is assigned a value of 4.0. Typical active metals have electronegativities of about 1.0 or less.

Within a period of the periodic table, elements become more electronegative from left to right. In the second period, the trend is regular.

Li	Be	B	C	N	O	F
1.0	1.5	2.0	2.5	3.0	3.5	4.0

That is, it increases by about 0.5 per element as we move from lithium at the far left to fluorine at the far right. In other periods the trend is in the same direction but less regular;

Na	Mg	Al	Si	P	S	Cl
0.9	1.2	1.5	1.8	2.1	2.5	3.0

Within a group, electronegativity decreases from top to bottom. Chlorine is less electronegative than fluorine and sulphur is less electronegative than oxygen. A comparison of electronegativities is not quite straight forward when considering two elements that are neither in the same period nor in the same group.

Activity 2.27

Form a group and discuss on the following:

1. What is the electronegativity of an atom? How is it different from electron affinity?
Discuss in groups.
2. In a given family of the periodic table, what is the general relationship between electronegativity and size?

Share your ideas with the rest of the class.

Metallic Character**Activity 2.28**

Form a group and summarize the trend in metallic character as a function of position in the periodic table. Is it the same as the trend observed for atomic size and ionization energy?

Share your ideas with the rest of the class.

Metallic character refers to the chemical properties associated with elements classified as metals. These properties arise from the elements ability to lose electrons. As one moves across a period from left to right in the periodic table, the metallic character decreases, as atoms are more likely to gain electrons to fill their valence shell rather than to lose them to remove the shell. Down a group, the metallic character increases, due to the lesser attraction from the nucleus to the valence electrons.

2.8.4 Advantages of Periodic Classification of the Elements**Activity 2.29**

Form a group and discuss the following:

- a In which region of the periodic table do you locate metals?
- b In which region are the elements with general electronic configuration of ns^2p^5 located? Give group number.
- c Write group number and period of an element with atomic number 34.

- d Is an element with the electronic configuration $1s^2 2s^2 2p^6 3s^1$ a metal or a non-metal? Will it form a cation or an anion readily? Give appropriate reason.
- e An element is located in Group 18 and 2nd Period of the modern periodic table. On the basis of this information predict the reactivity of the element.
- f An element X of 3rd Period has very low IE_1 and largest size in the period. Specify possible group numbers.

Some of the advantages of periodic classification of elements are:

1. The classification of elements is based on the atomic number, which is a fundamental property of an element.
2. The reason for placing isotopes at one place is justified as the classification is on the basis of atomic number.
3. It explains the periodicity of the properties of the elements and relates them to their electronic configurations.
4. The position of the elements that were misfits on the basis of mass number (anomalous pairs like argon and potassium) could be justified on the basis of atomic number.
5. The lanthanides and actinides are placed separately at the bottom of the periodic table.
6. The table is simple, systematic and easy way for remembering the properties of various elements as it is based on the electronic configuration.

Unit Summary

- Cathode rays (electrons) are produced when electricity passes through evacuated tubes. X-rays form when cathode rays strike matter.
- Radioactivity is the emission of radiation by unstable nuclei, and the most common types of radiations are alpha (α) particles, beta (β) particles and gamma (γ) rays. Alpha particles are helium nuclei; beta particles are electrons; and gamma rays are high frequency electromagnetic radiation similar to X-rays.
- Rutherford's atomic model is that of a very small positively charged nucleus and extra-nuclear electrons. The nucleus consists of protons and neutrons and contains practically all the mass of an atom. Atomic masses and relative abundances of the isotopes of an element can be established by mass spectrometry. The atomic mass of the element is the average of these mass numbers based on their percentage abundances.

- *Electromagnetic radiation is the transmission of electric and magnetic fields as a wave motion. The waves are characterized by their velocity in a medium: $c = v\lambda$. A light source that emits an essentially unbroken series of wavelength components has a continuous spectrum. Only a discrete set of wavelength components is present in the emission spectrum of an atom.*
- *Einstein's explanation of the photoelectric effect views light as packets of energy called photons. The energy of the photon (E_{ph}) is given by the expression $E = h\nu$ where h is Planck's constant.*
- *Bohr's theory requires the electron in a hydrogen atom to be in one of a discrete set of energy levels. The fall of an electron from a higher to a lower energy level releases a discrete amount of energy as a photon of light with a characteristic frequency.*
- *Bohr's theory accounts for the observed atomic spectrum of hydrogen atom.*
- *The electron in a hydrogen atom can be viewed as a matter-wave enveloping the nucleus. The matter-wave is represented by a wave equation, and solutions of the wave equation are wave functions. Each wavefunction is characterized by the value of four quantum numbers: the principal quantum number, n ; the angular momentum quantum number l ; the magnetic quantum number, m_l ; and the spin quantum number, m_s . Wave functions with acceptable values of the three are called atomic orbitals. An orbital describes a region in an atom that has a high probability of containing an electron or a high electron charge density. Orbitals with the same value of n are in the same principal energy level or principal shell. Those with the same value of n and of l are in the same sublevel or subshell. The shapes associated with orbitals depend on the value of l . Thus, the s orbital ($l = 0$) is spherical and the p orbital ($l = 1$) is dumbbell-shaped.*
- *The n , l and m_l quantum numbers define an orbital, but a fourth quantum number is also required to characterize an electron in an orbital - the spin quantum number, m_s . This quantum number may have either of two values: $+1/2$ or $-1/2$.*
- *The wave mechanical treatment of the hydrogen atom can be extended to multi-electron atoms, but with this essential difference: principal energy levels are (i) lower than those of the hydrogen atom and (ii) split, that is, having different energies for the different subshells.*

- *Electron configuration refers to the distribution of electrons among orbitals in an atom. Introduced here are the subshell notations (or "s,p,d, f") and the orbital diagram. Key ideas required to write a probable electron configuration are: (i) electrons tend to occupy the lowest energy orbitals available; (ii) no two electrons in an atom can have all four quantum numbers alike; and (iii) where ever possible, electrons occupy orbitals singly rather than in pairs.*
- *The Aufbau principle describes a hypothetical process of building up one atom from the atom of preceding atomic number. With this principle and the idea cited above, it is possible to predict probable electron configurations for many of the elements. In the Aufbau process, electrons are added to the s or p subshell of highest principal quantum number in the representative or main group elements. In transition elements, electrons go into the d subshell of the second last shell, and in the inner transition elements, into the f subshell of the third last shell.*
- *Elements with similar valence-shell electron configurations fall in the same group of the periodic table. For A-group elements, the group number corresponds to the number of electrons in the principal shell of highest quantum number. The period number is the same as the highest number of principal shell containing electrons (the outer shell). The division of the periodic table into s, p, d and f blocks greatly assists in the assignment of probable electron configurations.*
- *Certain atomic properties vary periodically, when atoms are considered in terms of increasing atomic number. The properties and trends considered in this unit are those of atomic radius, ionic radius, ionization energy and electron affinity. Values of these atomic properties strongly influence physical and chemical properties of the elements.*

Check List

Key terms of the unit

- | | |
|---------------------------------|------------------------------------|
| • <i>amplitude</i> | • <i>charge/mass ratio</i> |
| • <i>atomic mass</i> | • <i>Dalton's atomic theory</i> |
| • <i>atomic mass unit (amu)</i> | • <i>effective nuclear charge</i> |
| • <i>atomic number</i> | • <i>electronic configuration</i> |
| • <i>auf bau Principle</i> | • <i>electromagnetic radiation</i> |
| • <i>cathode</i> | • <i>exited state</i> |
| • <i>cathode rays</i> | • <i>ground state</i> |

- *frequency*
- *Hund's Rule*
- *inner-transition metal*
- *isotope*
- *isotope atom*
- *mass number*
- *metalloid*
- *non-metal*
- *Pauli's Exclusion Principle*
- *period*
- *periodic law*
- *periodic table*
- *photon*
- *proton*
- *quantum numbers*
- *representative element*
- *s-block elements*
- *Schrödinger equation*
- *transition metal*
- *uncertainty principle*
- *wave length*

REVIEW EXERCISE

Part I: Multiple Choice Type Questions

1. The number of neutrons in an atom of ${}^{226}_{88}\text{Ra}$ is:

a 88	c 138
b 82	d 314
2. Which of the following are usually found in the nucleus of an atom?
 - a Protons and neutrons only
 - b Protons, neutrons and electrons
 - c Neutrons only
 - d Electrons and neutrons only
3. An atom has an atomic number of 31 and a mass numbers of 70. How many electrons will it have in its valence shell?

a 5	c 3
b 4	d 2
4. Which of the following would produce a line spectrum rather than a continuous spectrum?

a Sunlight	c A normal filament light bulb
b Excited hydrogen atom	d A yellow (<i>sodium</i>) street light

- c In naturally occurring copper isotopes, ${}^{63}_{29}\text{Cu}$ contributes 69.09% and ${}^{65}_{29}\text{Cu}$, 30.91%. Calculate the relative atomic mass of copper. (Accurate mass determined; ${}^{63}_{29}\text{Cu} = 62.9298 m_u$, ${}^{65}_{29}\text{Cu} = 64.9278 m_u$)
19. Two particles **X** and **Y** have the following composition:
X: 17 protons, 18 neutrons, 17 electrons
Y: 17 protons, 18 neutrons, 18 electrons
- a What is the relationship between these particles?
b Will these two particles have similar chemical properties? Explain why?
20. Arrange the following in order of increasing ionization energy: Li, Na, Ne, N, O
21. Explain the following:
a The first ionization energy of beryllium is greater than that of boron.
b The first ionization energy of oxygen is less than that of nitrogen.
c The first ionization energy of lithium is greater than that of sodium.
22. The electron configuration of a particular metal cation M^{3+} is $[\text{Ar}] 3d^2$.
a Identify the corresponding metal.
b Write the electron configuration of the metal atom.
23. Arrange the following in order of increasing atomic radius Mg, Cs, Ca, Al, Ba.
24. Explain briefly, why potassium always occurs as a +1 ion in its compounds and calcium as a +2 ion.
25. Arrange the atoms (ions) in each of the following groups in order of increasing size based on their location in the periodic table.
a Mg^{2+} , O^{2-} , Na^+ , F^- , Al^{3+}
b Ne, N^{3-} , F^- , Na^+ , C^{4-}
c F, Be, C, B, Li
d K^+ , S^{2-} , As^{3-} , Cl^- , Ca^{2+}
26. Excited sodium atoms emit light with a wavelength of 589 nm. Calculate the:
a frequency of the light, and
b energy of one of these photons in joules

27. A hydrogen atom is excited to the $n = 8$ energy level. It emits a photon of light as it falls to the $n = 2$ energy level. Calculate the:
- wavelength of light emitted, and
 - frequency of the light emitted
28. The electron of a hydrogen atom is in the $n = 3$ level. What is its energy?
29. Calculate the wavelength of the light emitted when an electron falls from $n = 3$ to the $n = 1$ state in hydrogen atom.
30. The photon emitted by a cyclotron has a velocity of $1.50 \times 10^3 \text{ m s}^{-1}$. What is the wavelength of this photon? Given that the mass of photon = $1.676 \times 10^{-27} \text{ kg}$ and Planck's constant = $6.62 \times 10^{-34} \text{ J.s}$.
31. Write the number and the letter for the orbital that corresponds to the following pairs of n and l quantum numbers:
- | | |
|------------------|------------------|
| a $n = 3, l = 1$ | c $n = 3, l = 2$ |
| b $n = 4, l = 0$ | d $n = 5, l = 3$ |
32. Write the electron configurations for the following atoms and ions:
- | | |
|--------------------|--------------------|
| a Fe^{3+} | c Cr^{3+} |
| b V | d Al^+ |
33. Identify the transition element (s) from the following:
- | | |
|----------------------|----------------------|
| a ${}_{40}\text{Zr}$ | c ${}_{56}\text{Fe}$ |
| b ${}_{88}\text{Ra}$ | d ${}_{36}\text{Kr}$ |