MODEL TEXTBOOK OF CHEMISTRY

Based on National Curriculum of Pakistan 2022-23

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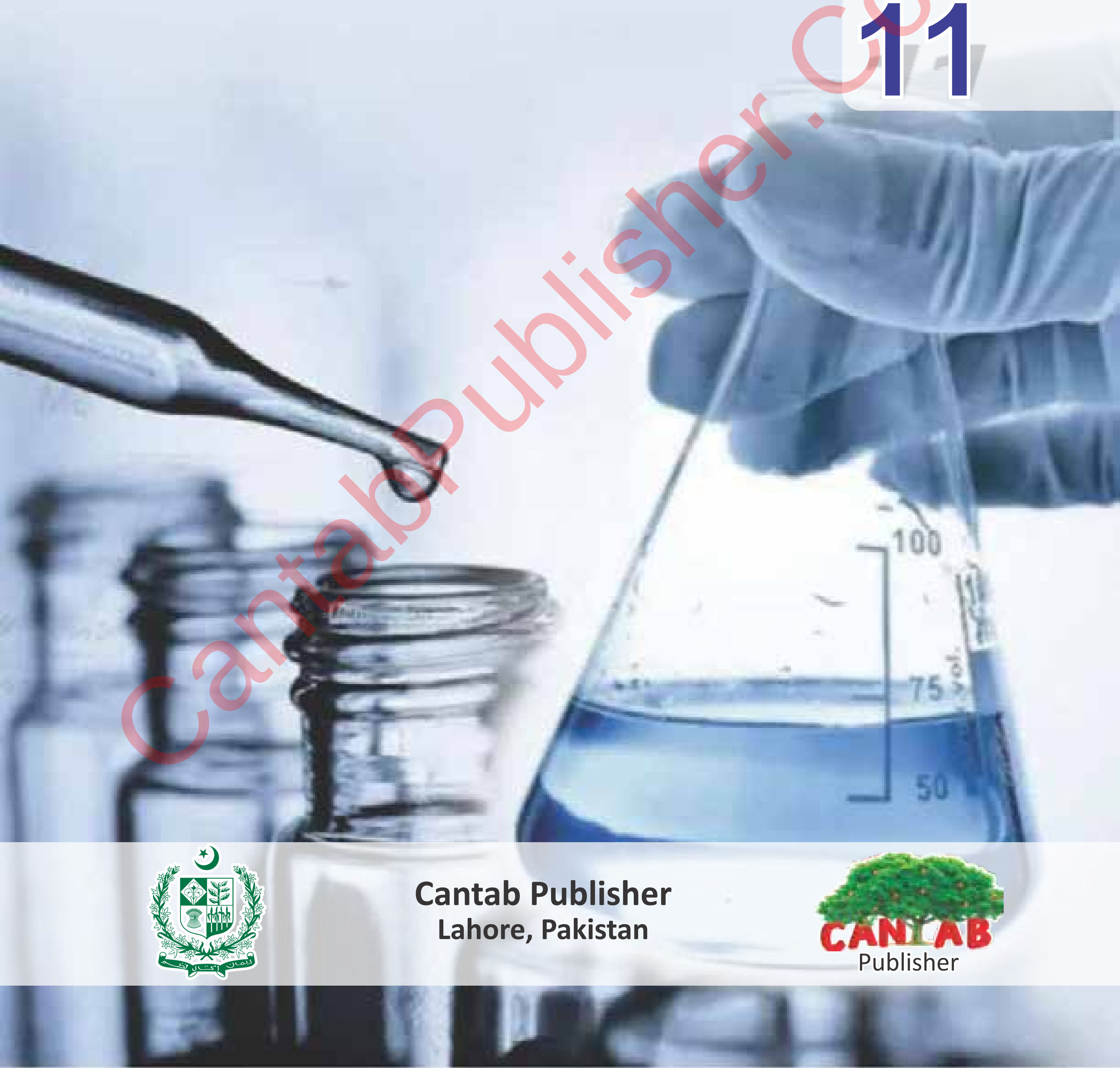

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A Textbook of Chemistry
for Grade 11

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Preface

This Grade 11 Chemistry textbook is designed to elevate your learning experience with advanced concepts and applications. It features detailed visual representations, real-world examples, and hands-on experiments to deepen your understanding. The book includes challenging exercises that encourage critical thinking, self-assessment sheets to gauge your comprehension, and collaborative group activities for interactive learning. Recorded lectures with dynamic animations and simulations bring complex topics to life, making them more accessible and engaging.Structured to support educators in crafting assessments aligned with Bloom's Taxonomy, this textbook ensures you are prepared to tackle higher-order thinking questions. A comprehensive glossary at the end provides quick reference to advanced terms, aiding in your mastery of Chemistry concepts.This educational tool is crafted to enrich your knowledge and foster a deeper appreciation of Chemistry, aligned with the 2023 curriculum standards.

2.2 Knowledge

Exploring Electronic Configurations: Quantum Numbers, Subshells, and Their Applications in Technology

What inside this QR Cord

Engaging Video Lectures
Explore engaging video lectures with dynamic PowerPoint presentations, simulations, and 3D animations for a deeper understanding of Chemistry concepts.

Tailored SLO-Based Question Bank
Prepare effectively with a question bank aligned to Student Learning Outcomes (SLOs), perfect for competitive exams like MDCAT, ECAT, and NUMS.

Latest Research Integration
Stay updated with the latest research integrated into your learning materials, enriching your knowledge base in Chemistry.

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Access answers to numerical and self-test questions instantly via QR codes, providing quick feedback and tracking your progress effortlessly.

Comprehensive Worksheets and Skill Sheets
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Specialized Exam Preparation
Prepare confidently for competitive exams such as MDCAT, ECAT, and NUMS with updated study materials tailored for intermediate students.

The QR codes in the Chemistry textbook provide easy access to video lectures for gaining knowledge and skill sheets for practical application. They make learning more interactive, letting students watch lectures and practice skills right when they need them, making studying chemistry more engaging and effective.

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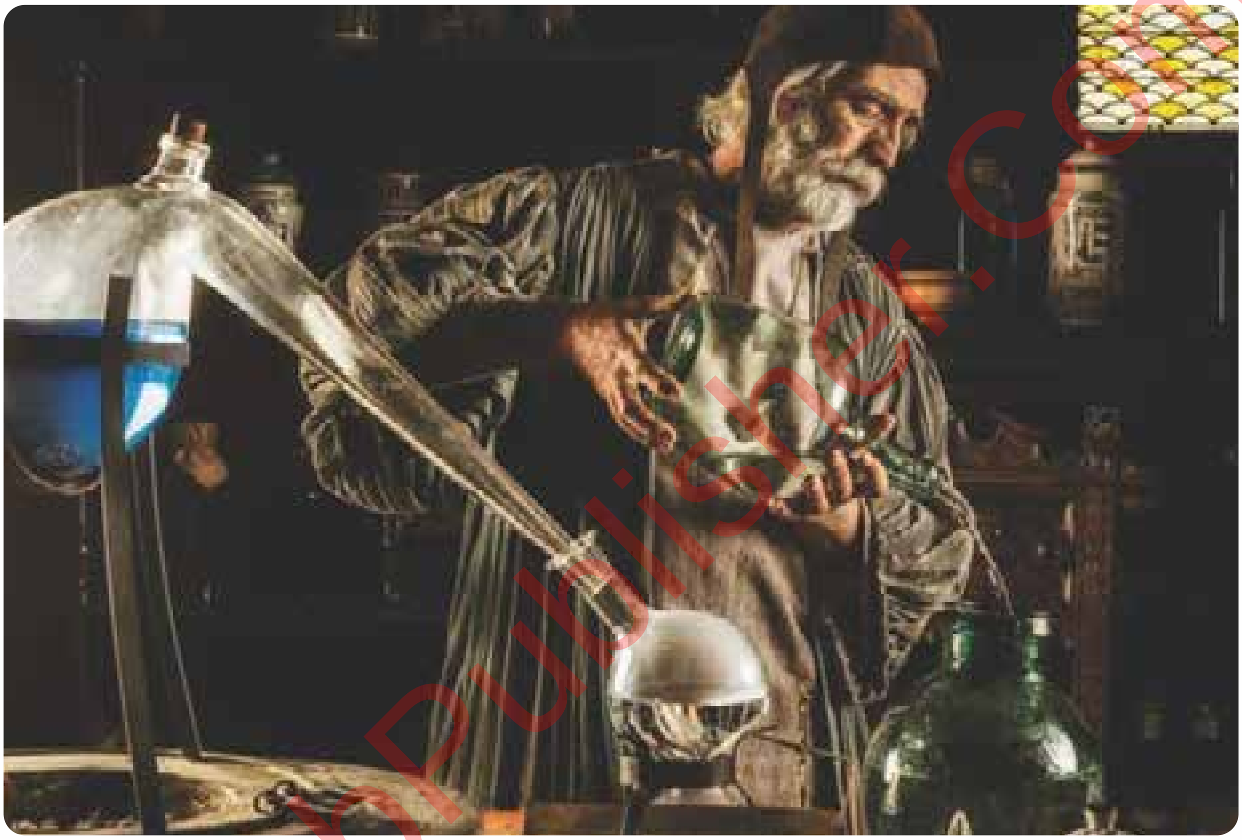
CHAPTER

1

History of Chemistry



Introduction

In this session, we will examine the captivating challenges of inductive reasoning using Al-Ghazali's renowned burning cotton thought experiment. By grasping this thought experiment, we will expose the constraints of forming conclusions solely based on repeated observations. Furthermore, we will explore the reliability of deductive reasoning, particularly in the context of chemistry, where established principles and laws enable us to make precise predictions and derive accurate conclusions. Throughout this exploration, we will develop a deeper understanding of the different types of reasoning employed in scientific inquiry and their pivotal role in shaping reliable scientific knowledge.

Student Learning Outcomes

- Describe how Al Ghazali's burning cotton thought experiment highlight the challenges of inductive reasoning examples of deductive reasoning with respect to Chemistry)

All the above mentioned SLOs are classified into knowledge and skills for the better understanding of students.

After studying this Unit, the students will be able to understand:

Knowledge

☑ **1.1** Understanding the definitions of inductive and deductive reasoning. Learning about historical and philosophical logics about inductive and deductive reasoning.

Skill

☑ **1.1** Developing the ability to identify and differentiate between inductive and deductive reasoning in scientific texts and experiments.

1.1 Knowledge

Challenges of Inductive Reasoning: Al-Ghazali's Burning Cotton Thought Experiment and Examples of Deductive Reasoning in Chemistry

Student Learning Outcomes

- Describe how Al Ghazali's burning cotton thought experiment highlight the challenges of inductive reasoning examples of deductive reasoning with respect to Chemistry)

What is Inductive Reasoning?

Inductive reasoning is a method of thinking that involves drawing general conclusions based on specific observations. It starts with specific data or examples and then arrives at a broad rule or conclusion. For instance, if you consistently see the sun rising in the east every morning, you might conclude that it always rises in the east. This type of reasoning is particularly useful for scientists when they are forming theories and hypotheses based on observed patterns in nature.

However, it's important to be aware of the limitations of inductive reasoning in scientific inquiry. One significant challenge is that it doesn't guarantee certainty. Just because something has happened repeatedly in the past doesn't mean it will always happen in the future. This assumption of uniformity can lead to incorrect conclusions if the observed patterns change.

For example, assuming all swans are white because every swan you have seen so far is white can be proven wrong by the discovery of a single black swan.

Another limitation of inductive reasoning is that it heavily depends on the observations made. If these observations are limited or biased, the conclusions drawn may be flawed.

Fig 1.1 The existence of black swans, scientifically known as Cyanaurates and native to Australia, offers a compelling illustration of the limitations of inductive reasoning. Prior to the discovery of black swans by Europeans, the prevailing belief in Europe was that all swans were white, based on the observed instances. The revelation of black swans underscored the potential for inductive reasoning based on limited observations to yield inaccurate conclusions

Inductive reasoning also tends to ignore potential exceptions to the observed patterns. This is why it's crucial for scientists to rigorously test their inductive conclusions. While inductive reasoning is useful for forming hypotheses, these hypotheses need to be verified through experiments and further observations.

Understanding these limitations is crucial in scientific inquiry because it reminds scientists to be cautious with their conclusions. Scientists often use inductive reasoning to identify patterns and develop initial theories.

However, they must also use deductive reasoning to test and confirm these theories. Deductive reasoning allows scientists to derive specific predictions from general principles and test them to ensure their validity. This process helps to construct more reliable and accurate scientific knowledge. In conclusion, while inductive reasoning is a valuable tool for scientists to identify patterns and make predictions, recognizing its limitations ensures that scientific theories are thoroughly tested and validated. This leads to a more robust and accurate understanding of the natural world.

Let's explore the limits of inductive reasoning by examining a well-known thought experiment by the Islamic philosopher Al-Ghazali and how his burning cotton experiments promote deductive reasoning.

Al-Ghazali, a distinguished Islamic philosopher and theologian of the 11th century, is well-regarded for his contributions to philosophy, theology, and logic. His burning cotton thought experiment is particularly noteworthy as it challenges the reliability of inductive reasoning. In this thought experiment, Al-Ghazali questions the certainty of causal relationships obtained from repeated observations. He illustrates this by using the example of cotton burning when exposed to fire.

Fig 1.2

Through inductive reasoning, one might conclude that fire causes cotton to burn. However, Al-Ghazali argues that this conclusion is not necessarily valid. He suggests that the connection between fire and burning is not inherently causal but rather a habit of the mind based on repeated observations. According to Al-Ghazali, an omnipotent being, like God in his philosophical context, could intervene and alter the outcome, making the burning of cotton a contingent event rather than a necessary one.

To illustrate this, Al-Ghazali draws a parallel to the story of Prophet Ibrahim (Abraham). When Prophet Ibrahim was thrown into the fire by his enemies, the fire did not burn him. Instead, by divine intervention, the fire became cool and safe, much like flowers. This story exemplifies the idea that what we perceive as natural laws can be altered by a higher power, challenging the assumption that the relationship between fire and burning is absolute.

Fig 1.3 The text below describes the early Islamic debate between reason and anti-reason and the role of Al-Ghazali and Abu Hanifa in it. Initially, there were no arguments about the essence of Islam during the time of the Prophet and the first four caliphs. As the Islamic state expanded and diverse people joined, new questions arose. This led to the classification of issues into theology and jurisprudence. Abu Hanifa, a Great Imam, developed an approach to solving new issues by using Quran and hadiths to find underlying reasoning for each verdict. His approach, known as "Hanafism" became one of the four schools of thought in Sunni Islam.

This thought experiment emphasizes the fundamental challenge of inductive reasoning, as it assumes that future events will follow past patterns without necessarily offering a basis for this assumption. Al-Ghazali's argument introduces a level of uncertainty, indicating that our inductive conclusions might not be as certain as we believe. It serves as a reminder that the relationships observed in the natural world might be subject to unseen forces or exceptions that have not been accounted for.

Understanding these limitations is crucial in scientific inquiry, as it prompts scientists to be cautious with their conclusions. While scientists employ inductive reasoning to identify patterns and develop initial theories, they must also use deductive reasoning to test and confirm these theories. Deductive reasoning enables scientists to derive specific predictions from general principles and test them to validate their accuracy. This process helps build more reliable and accurate scientific knowledge.

Example of Inductive Reasoning in Chemistry along with limitation.

Inductive reasoning is often employed to form hypotheses based on observed patterns. For example, the repeated observation of acids turning blue litmus paper red can lead to a broad inductive conclusion. For instance, hydrochloric acid (HCl), sulfuric acid (H_2SO_4), and nitric acid (HNO_3) all turn blue litmus paper red, which can lead to the inductive conclusion that all acids turn blue litmus paper red. Although this conclusion is often correct and valuable, it is drawn from repeated observations rather than an underlying principle, and thus is subject to certain limitations.

It's important to recognize that there may be exceptions, there are few acids that deviates from this pattern. In summary, while inductive reasoning is a useful tool for scientists to identify patterns and make predictions, understanding its limitations is crucial for proper testing and validation of scientific theories, resulting in a more robust and accurate understanding of the natural world.

Update Yourself

Sure, here are the names of some acids that might not change the color of litmus paper under certain conditions:

1. Boric Acid (H_3BO_3)
2. Carbonic Acid (H_2CO_3)
3. Very dilute Acetic Acid (CH_3COOH)
4. Very dilute Citric Acid ($C_6H_8O_7$)
5. Very dilute Phosphoric Acid (H_3PO_4)

Deductive reasoning allows scientists to derive specific predictions from general principles and test them to ensure their validity. This process helps in building more reliable and accurate scientific knowledge.

Deductive Reasoning and Its Applications in Chemistry

Deductive reasoning involves drawing specific conclusions from general principles or premises that are known to be true. Unlike inductive reasoning, which moves from specific observations to general conclusions, deductive reasoning starts with general statements and moves to specific conclusions. This method provides more certainty because the conclusions logically follow from the premises.

Examples of Deductive Reasoning are given in table

Table 1.1 Examples of Deductive Reasoning in Chemistry

Examples	General Principle	Deductive Application
Chemical Reactions and Stoichiometry	Law of Conservation of Mass	The Law of Conservation of Mass states that matter is neither created nor destroyed in a chemical reaction.
		Balanced Chemical Equation: $2H_2 + O_2 \rightarrow 2H_2O$
		Given: 4 grams of hydrogen and 32 grams of oxygen
		Deduction: We can deduce that 36 grams of water will be produced. The mass of the reactants equals the mass of the products, illustrating the conservation of mass.

Acid-Base Reactions	Brønsted-Lowry theory of acids and bases	According to the Brønsted-Lowry theory, acids donate protons (H_2) and bases accept protons.
		Given: Hydrochloric acid (HCl) and sodium hydroxide (NaOH)
		Deduction: We can deduce that $HCl+NaOH \rightarrow NaCl+H_2O$ will produce a salt (sodium chloride) and water. This prediction follows from the general principle of proton transfer between acids and bases.
Ideal Gas Law	Ideal Gas Law $PV=nRT$	The Ideal Gas Law relates the pressure (P), volume (V), temperature (T), and number of moles (n) of a gas with the gas constant (R).
		Given: Values of P (pressure), V (volume), and T(temperature)
		Deduction: Using the equation $PV=nRT$ we can deduce the number of moles n of the gas. For example, if $P=1$ atm, $V=22.4$ L and $T=273$ K, we can calculate n using the Ideal Gas Law.

This table format organizes the examples clearly, showing the general principles and their deductive applications in chemistry.

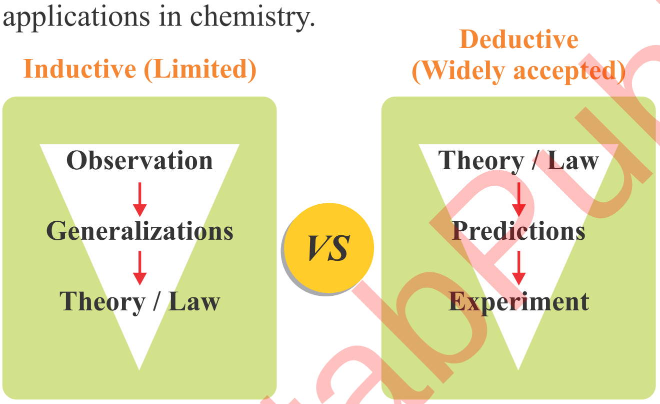


Fig 1.4

It's essential for students to grasp deductive reasoning. While inductive reasoning relies on observations and can lead to error-prone generalizations, deductive reasoning starts with a theory, makes predictions, and tests them through experiments, leading to more reliable conclusions. Emphasizing deductive reasoning helps students form conclusions based on sound evidence and logical progression. This fosters a critical approach to validating theories, which is indispensable for effective problem-solving and analysis throughout their academic journey.

Test your self

- What is inductive reasoning? Provide a definition and an example in your own words.
- Why is it important to recognize the limitations of inductive reasoning in scientific inquiry?
- Describe a scenario in chemistry where inductive reasoning might lead to an incorrect conclusion.
- Explain Al-Ghazali's burning cotton thought experiment. What point was he trying to make about causal relationships and inductive reasoning?
- What is the main difference between inductive and deductive reasoning?
- Provide an example of an inductive conclusion and a deductive conclusion in a chemical context.
- Why is deductive reasoning considered more reliable than inductive reasoning for making scientific predictions?
- Using the Brønsted-Lowry theory, predict the products of the reaction between hydrochloric acid (HCl) and sodium hydroxide (NaOH). Explain your deductive reasoning process.
- How can the insights from Al-Ghazali's thought experiment be applied to modern scientific practices, particularly in chemistry?
- Reflect on a time when you used inductive reasoning in a science experiment. What was the outcome, and how could deductive reasoning have helped?

CHAPTER

2

Atomic Structure

Introduction

This chapter provides a detailed exploration of atomic structure and electron behaviour, which are central to understanding the foundations of chemistry and material science. It delves into the intricacies of atomic shells and subshells, highlighting how these structures are composed of degenerate orbitals with identical energy levels. The chapter also introduces the fundamental particles of an atom protons, neutrons, and electrons – discussing their relative charges and masses and clarifies the concepts of atomic (proton) number and mass (nucleon) number. This understanding is crucial in determining the number of protons, neutrons, and electrons in atoms and ions, based on atomic number, mass number, and charge. Quantum numbers play a crucial role in this discussion as they relate to the electronic distribution of elements. The chapter describes the number of orbitals in s, p, d, and f subshells, and the maximum number of electrons these subshells can accommodate. It also applies fundamental principles such as the Aufbau principle, Pauli exclusion principle, and Hund's rule. The order of increasing energy of subshells (s, p, d, f) is crucial to understand, and the chapter discusses this in the context of electronic configurations, including the number of electrons in each shell, subshell, and orbital. It illustrates the significance of electronic configurations in the development of new materials, particularly in the field of electronics.

Student Learning Outcomes

- Describe that, each atomic shell and subshell are further divided into degenerate orbitals having the same energy.
- Describe protons, neutrons, and electrons in terms of their relative charges and relative masses
- Recognize that the terms atomic and proton number represent the same concept
- Recognize the terms mass and nucleon number represent the same concept
- Describe the behavior of beams of protons, neutrons, and electrons moving at the same velocity in an electric field
- Determine the numbers of protons, neutrons, and electrons present in both atoms and ions given atomic or proton number, mass/or nucleon number and charge
- Explain the change in atomic and ionic radius across a period and down a group
- Determine the electronic configuration of elements and their ions with proton numbers.
- (Some examples include a) simple configurations e.g. 2,8,b. subshells e.g. Is2, 2s2, 2p6, 2s1 c) students should be able to determine both of these from the periodic table and are not required to memorize these d. students should understand that the chemical properties of an atom are governed by valence electrons)
- Define terms related to electronic configuration (Some examples include shells, subshells, orbitals, principal quantum number (n), ground state),
- Relate Quantum Numbers to the Electronic Distribution of Elements.
- Describe the number of orbitals making up s, p d, and f sub-shells, and the number of electrons that can fill s, p d, and f subshells
- Apply the Aufbau principle, Pauli exclusion principle and Hund's rule to write the electronic configuration of elements
- Describe the order of increasing energy of the subshells (s,p,d, and f)
- Describe the electronic configurations to include the number of electrons in each shell, subshell and orbital.
- Explain the electronic configurations in terms of the energy of the electrons and inter-electron repulsion
- Determine the electronic configuration of atoms and ions given the proton or electron number and charge
- Illustrate the importance of electronic configurations in the development of new materials for electronic devices. (For example, semiconductors such as silicon have a specific electronic configuration that makes them ideal for use in electronic devices.)
- Describe the shapes of s, p, and d orbitals.
- Describe a free radical as a species with one or more unpaired electrons.
- Explain that ionization energies are due to the attraction between the nucleus and the outer electron.
- Explain how ionization energy helps account for the trends across a period and down a group of the Periodic Table.
- Account for the variation in successive ionization energies of an element.
- Explain the factors influencing the ionization energies of elements in terms of nuclear charge, atomic/ionic radius, shielding by inner shells and subshells, and spin pair repulsion.
- Deduce the electronic configurations of elements using successive ionization energy data.
- Deduce the position of an element in the Periodic Table using successive ionization energy data.
- Explain how a mass spectrometer can be used to determine the relative atomic mass of an element from its isotopic composition.
- Perform calculations involving non-integer relative atomic masses and abundance of isotopes from given data, including mass spectra.
- Explain the concept of emission spectra Use the concept of emission spectra to deduce the electronic configuration of elements.

All the above mentioned SLOs are classified into knowledge and skills for the better understanding of students.
After studying this Unit, the students will be able to understand:

Knowledge

☑ **2.1: Atomic Structure and Mass Spectrometry: Understanding Particles and Isotopic Analysis**

Students will learn about the fundamental particles of an atom protons, neutrons, and electrons focusing on their relative charges and masses. They will understand that "atomic number" and "proton number" are interchangeable terms that denote the number of protons in an atom, while "mass number" and "nucleon number" both refer to the total number of protons and neutrons. Additionally, the curriculum includes how a mass spectrometer operates and uses isotopic composition to calculate the relative atomic mass of elements.

☑ **2.2: Exploring Electronic Configurations: Quantum Numbers, Subshells, and Their Applications in Technology**

Students will delve into the fundamentals of electronic configuration, covering terms like shells, subshells, orbitals, and quantum numbers including the principal quantum number (n) and ground state. They will learn about the number of orbitals in s, p, d, and f subshells, and the maximum electrons each can hold. The curriculum will also introduce the sequence of subshell energy levels (s, p, d, f) and discuss electronic configurations in terms of electron energy and inter-electron repulsion. Students will explore how specific configurations, like those in semiconductors, play crucial roles in electronic device development, and how emission spectra can be used to determine electronic configurations.

☑ **2.3: Understanding Periodic Trends: Atomic Structure, Ionization Energies, and Electronic Configurations**

Students will investigate the variation in atomic and ionic radii across periods and down groups in the periodic table. They will explore the concept of electronic configurations, including simple configurations and subshell notations like 1s2, 2s2, 2p6. They will also learn about the nature of free radicals as species with unpaired electrons and delve into the principles governing ionization energies, including the factors that influence these energies nuclear charge, atomic radius, electron shielding, and spin-pair repulsion. This knowledge will extend to understanding how these properties influence chemical properties and trends in the periodic table.

Skill

☑ **Skills 2.1:**

Students will acquire skills in recognizing how protons, neutrons, and electrons behave when subjected to an electric field at the same velocity. They will learn to determine the specific counts of protons, neutrons, and electrons in atoms and ions based on given atomic numbers, mass numbers, and charges. Furthermore, students will perform calculations involving the relative atomic masses and isotopic abundances from mass spectra data, enhancing their analytical abilities in interpreting complex scientific data.

☑ **Skills 2.2**

Students will develop skills to apply foundational principles such as the Aufbau principle, Pauli exclusion principle, and Hund's rule to accurately write the electronic configurations of elements. They will be trained to determine the electronic configurations of both atoms and ions based on given atomic numbers and charges. Furthermore, students will analyze emission spectra to deduce electronic structures, enhancing their understanding of quantum mechanics in practical scenarios. They will also learn to describe the spatial shapes of s, p, and d orbitals, equipping them with the necessary tools to interpret and predict chemical and physical properties of elements.

☑ **Skills 2.3**

Students will develop the ability to determine the electronic configurations of elements and their ions from the periodic table without the need for memorization. They will analyze trends in ionization energy across periods and down groups, and interpret variations in successive ionization energies. Skills will also include using data on ionization energies to deduce both the electronic configurations and the positions of elements within the periodic table, enhancing their analytical capabilities in applying theoretical chemistry concepts to practical scenarios.

2.1 Knowledge

Atomic Structure and Mass Spectrometry: Understanding Particles and Isotopic Analysis

Student Learning Outcomes

- Describe protons, neutrons, and electrons in terms of their relative charges and relative masses
- Recognize that the terms atomic and proton number represent the same concept
- Recognize the terms mass and nucleon number represent the same concept
- Describe the behavior of beams of protons, neutrons, and electrons moving at the same velocity in an electric field
- Determine the numbers of protons, neutrons, and electrons present in both atoms and ions give atomic or proton number, mass/or nucleon number and charge
- Explain how a mass spectrometer can be used to determine the relative atomic mass of an element from its isotopic composition.
- Perform calculations involving non-integer relative atomic masses and abundance of isotopes from given data, including mass spectra.

Atoms are the basis of all of chemistry. You learned about them at grade 9th and they're here again. They're super important.

The structure of the atom

All elements are made of atoms. Atoms are made up of 3 types of particles protons, neutrons and electrons. Figure 2.1 shows how they are arranged in the atom.

Electrons have $^-1$ charge. They whizz around the nucleus in orbitals, which take up most of the volume of the atom. Most of the mass of the atom is concentrated in the nucleus, although the diameter of the nucleus is rather small compared to the whole atom. The nucleus is where you find the protons and neutrons. The mass and charge of these subatomic particles is really small, so relative mass and relative charge are used instead. Table 2.1 shows the relative masses and charges of protons, neutrons and electrons.

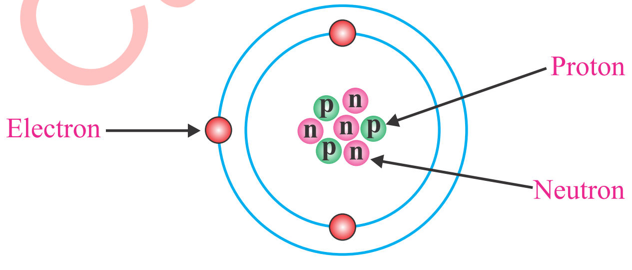


Fig: 2.1

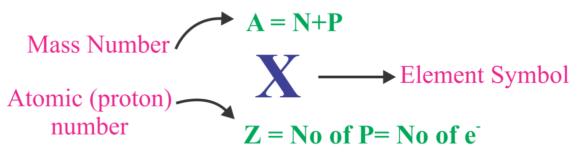
Table 2.1: Relative masses and charge of subatomic particles

Subatomic particle	Relative mass	Relative charge
Proton	1	+1
Neutron	1	0
Electron, e^-	$\frac{1}{2000}$	– 1

Mass number and atomic number

You can figure out the number of protons, neutrons and electrons in an atom from the nuclear symbol.

Mass number: This is the total number of protons and neutrons in the nucleus of an atom.



Atomic (proton) number: This is the number of protons in the nucleus of an atom it identifies the element. All atoms of the same element have the same number of protons. Sometimes the atomic number is left out of the nuclear symbol, e.g.7Li. You don't really need it because the element's symbol tells you, its value.

Atoms: For neutral atoms, which have no overall charge, the number of electrons is the same as the number of protons. The number of neutrons is just mass number minus atomic number Table shows some examples.

Table 2.2 Calculating the number of neutrons in atoms

Nuclear Symbol	Z	A	Protons	Electrons	Neutrons
$^{7}_{3}Li$	3	7	3	3	7 – 3 = 44
$^{79}_{35}Br$	35	79	35	35	79 – 35 = 44
$^{24}_{12}Mg$	12	24	12	12	24 – 12 = 12

Ions

Atoms form ions by gaining or losing electrons. Ions have different numbers of protons and electrons negative ions have more electrons than protons, and positive ions have fewer electrons than protons.

Example

Br^- is a negative ion: The negative charge means that there's 1 more electron than there are protons. Br has 35 protons (see table above), so Br^- must have 36 electrons. The overall charge +35 – 36 = –1.

Mg^{2+} is a positive ion: The 2+ charge means that there's 2 fewer electrons than there are protons. Mg has 12 protons (see table below), so Mg^{2+} must have 10 electrons. The overall charge = +12 – 10 = +2.

Isotopes

Isotopes of an element are atoms with the same number of protons but different numbers of neutrons.

Chlorine-35 and chlorine-37 are examples of isotopes. They have different mass numbers, so they have different numbers of neutrons. Their atomic numbers are the same both isotopes have 17 protons and 17 electrons.

Chlorine 35:	Chlorine 37:
$^{35}_{17}Cl$	$^{37}_{17}Cl$
35 – 17 = 18 neutrons	37 – 17 = 20 neutrons

Here's another example naturally occurring magnesium consists of 3 isotopes.

Table 2.3 Subatomic particles in Mg isotopes.

^{24}Mg (79%)	^{25}Mg (10%)	^{26}Mg (11%)
12 protons	12 protons	12 protons
12 neutrons	13 neutrons	14 neutrons
12 electrons	12 electrons	12 electrons

The number and arrangement of the electrons decides the chemical properties of an element. Isotopes have the same configuration of electrons, so they have the same chemical properties. Isotopes of an element do have slightly different physical properties though, e.g. **different densities and rates of diffusion.** This is because physical properties tend to depend more on the mass of the atom.

Atomic models

Models of the atom are useful for understanding loads of ideas in chemistry. But the accepted model of the atom has changed throughout history. The models of the atom that are currently accepted are based on the findings of numerous experiments and research, and have evolved over time to better fit the evidence at hand. One of the earlier models proposed by John Dalton in the early 19th century depicted atoms as solid spheres, with different spheres representing different elements (as shown in Figure 2.2). It's important to note that as new evidence is discovered, these models may continue to be revised and refined.

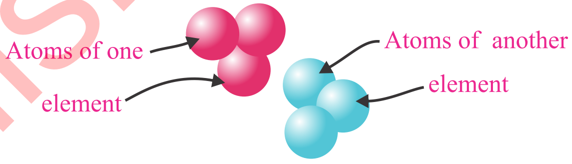


Fig: 2.2 Dalton's model of the atom

In 1897 J. J. Thomson concluded from his experiments that an atom must contain even smaller, negatively charged particles electrons. The 'solid sphere' idea of atomic structure had to be changed. The new model was known as the 'plum pudding model' see Figure 2.3.

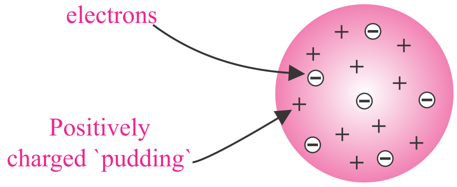


Fig: 2.3 Thomson's model of the atom

Rutherford's model

In 1909, Ernest Rutherford, along with his students Hans Geiger and Ernest Marsden, conducted a renowned experiment using gold foil. They aimed alpha particles, which have a positive charge, at an incredibly thin sheet of gold. It was expected, based on the plum pudding model, that the positive 'pudding' that made up most of the atom would slightly deflect most alpha particles. However, the results showed that most alpha particles passed straight through the gold atoms, and only a few were deflected backwards. This meant that the plum pudding model was flawed. As a result, Rutherford proposed a new idea - the nuclear model of the atom (as shown in the figure 2.4). This model suggests that there is a tiny, positively charged nucleus at the center, surrounded by a 'cloud' of negative electrons, and that most of the atom is actually empty space.

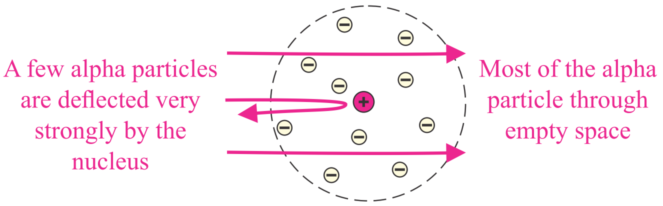


Fig: 2.4 Rutherford's model of the atom.

Bohr's model

Scientists realized that electrons in a 'cloud' around the nucleus of an atom, as Rutherford described would quickly spiral down into the nucleus, causing the atom to collapse. Niels Bohr proposed a new model of the atom with four basic principles:

- Electrons only exist in fixed orbits (shells) and not anywhere in between.
- Each shell has a fixed energy.
- When an electron moves between shells, electromagnetic radiation is emitted or absorbed.
- Because the energy of shells is fixed, the radiation will have a fixed Frequency

The frequencies of radiation emitted and absorbed by atoms were already known from experiments. The Bohr model fitted these observations.

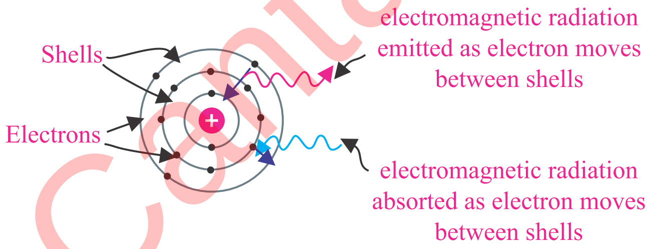


Fig: 2.5 Bohr's model of an atom.

Other atomic models

After further research, scientists discovered that electrons in a shell do not all have the same energy. As a result, they refined the Bohr model to include sub-shells, which we will delve into later on in this chapter. This updated model proved to be even more accurate than the original. Although the refined Bohr model is not without flaws, it remains a popular method for describing atoms due to its simplicity and ability to explain various experimental observations, such as bonding and ionization energy trends. However, the most precise model we have today is based on quantum mechanics. The quantum model accounts for observations that the Bohr model cannot explain. Nevertheless, it is more difficult to understand and visualize. In higher classes you will dis in discuss in higher classes.

Challange

- Describe how J. J. Thomson's model of the atom was different from Dalton's model.
- Explain how Rutherford's gold foil experiment provided evidence that Thomson's model was wrong.
- Describe Rutherford's model of the atom.
- Describe the main features of Bohr's model of the atom.

Relative Mass

The actual mass of an atom is very, very tiny. Don't worry about exactly how tiny for now, but it's far too small to weigh. So, you usually talk about the mass of an atom compared to the mass of one carbon atom instead this is its relative mass. You need to know about relative atomic mass, relative isotopic mass and relative molecular mass.

Relative atomic mass

The relative atomic mass, Ar, is the average mass of an atom of an element on a scale where an atom of carbon-12 is exactly 12. The relative atomic mass of each element is shown in the periodic table.

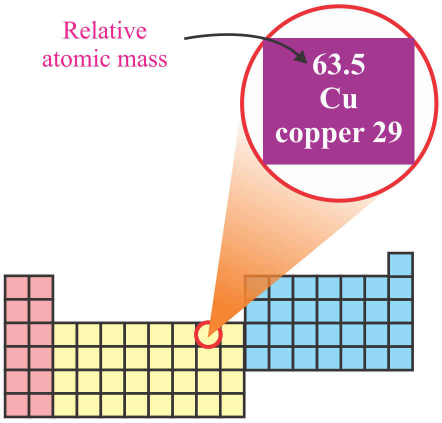


Fig: 2.6 Location of relative atomic masses on the periodic table.

Relative isotopic mass: Relative isotopic mass is the mass of an atom of an isotope of an element on a scale where an atom of carbon-12 is exactly 12.

Calculating relative atomic mass (Average atomic mass): Relative isotopic mass is usually a whole number. Relative atomic mass is an average, so it's not usually a whole number for more clearity see work example.

Worked Example

A natural sample of chlorine contains a mixture of ^{35}Cl and ^{37}CI, whose relative isotopic masses are 35 and 37. 75% of the sample is ^{35}CI and 25% is ^{37}CI. You need to take these percentages into account when you work out the relative atomic mass of chlorine.

we can find these % with the help of mass spectometer

= isotopic masses × percentages

$$\text{Relative atomic mass} = \frac{(35 \times 75) + (37 \times 25)}{100} \leftarrow \text{total percentage}$$

Relative atomic mass = 35.5 $(75\%\,^{35}\text{Cl} + 25\%\,^{37}\text{Cl})$

The Mass Spectrometer

A mass spectrometer is a device used for analyzing elements or compounds. Understanding how a time-of-flight (TOF) mass spectrometer works is crucial.

How a mass spectrometer works

A mass spectrometer can provide you with information about the relative atomic mass of an element, and the relative abundance of its isotopes or the relative molecular mass of a molecule, by analyzing a compound. When a sample is sprayed into a time-of-flight (TOF) mass spectrometer, four things occur.

1. **Ionization:** The sample must be ionized prior to entering the mass spectrometer. Two methods can be employed to achieve this.:

Electrospray ionization: Butt at here we will discuss only **Electron impact ionization** In this method, the sample is vaporized, and an electron gun fires high-energy electrons at it. This process knocks one electron off each particle, resulting in $^{+}1$ ions.

2. **Acceleration:** The positive ions are propelled forward by an electric field. This electric field provides equal amounts of kinetic energy to all ions, regardless of their weight. However, the lighter ions experience a greater acceleration as they are given the same amount of energy as the heavier ions, but their lighter weight allows them to accelerate more.

3. **Ion drift:** After passing through the region with an electric field, the ions move through a field-free region at the same speed. As a result, lighter ions move faster.

4. **Detection:** Lighter ions move faster through the drift region and arrive at the detector quicker than heavier ions. The detector measures the current generated when the ions strike it and records the time they take to pass through the spectrometer. This information is then utilized to determine the necessary mass/charge values for producing a mass spectrum.

Interpreting a mass spectrum

A mass spectrum is a type of chart produced by a mass spectrometer. It shows information about the sample that was passed through the mass spectrometer.

If the sample is an element, each line will represent a different isotope of the element. The y-axis gives the abundance of ions, often as a percentage. For an element, the height of each peak gives the relative isotopic abundance (the relative amount of each isotope presents in a sample). The x-axis units are given as a 'mass/charge' ratio.

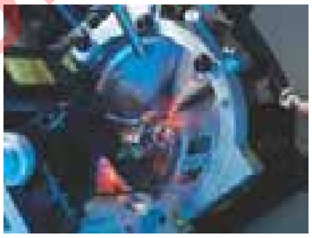


Fig: 2.7(a): Electros pray ionisation apparatus.

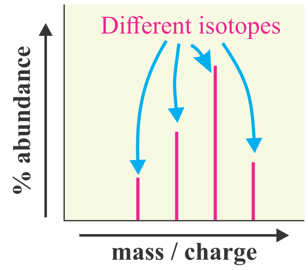


Fig: 2.7(b): The mass specturm of an element.

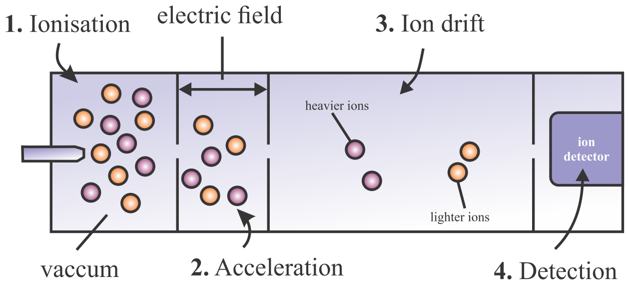
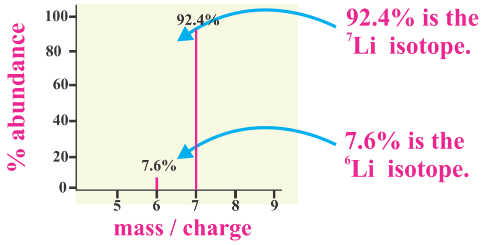


Fig: 2.7(c) Diagram showing how a TOF mass spectrometer work

Worked Example

The mass spectrum produced when a sample of lithium is passed through a mass spectrometer is shown below.



92.4% is the ^{7}Li isotope.

7.6% is the ^{6}Li isotope.

There are two peaks, so the sample contains two isotopes of lithium.One peak has a mass/charge ratio of 6, so the relative isotopic mass of one isotope is 6. (Since the lithium ions produced in a mass spectrometer have a charge of 1+ (Li^+) and $6 \div 1 = 6$.The other peak has a mass/charge ratio of 7, so the relative isotopic mass of the other isotope is 7. So, from the mass spectrum you can tell that there are two isotopes present, ^{6}Li and ^{7}Li. 92.4% of the sample is ^{7}Li. 7.6% of the sample is ^{6}Li.

How these peaks helps us to determine average atomic mass of an elements:

we can find these % with the help of mass spectometer

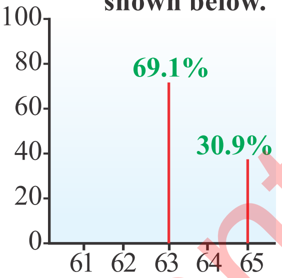
$$= \text{isotopic masses} \times \text{percentages}$$

$$\text{Relative atomic mass} = \frac{(6 \times 7.6) + (7 \times 92.4)}{100 \longleftarrow \text{total percentage}}$$

$$\text{Relative atomic mass} = 6.924 \qquad (7.6\%\,^6\text{Li} + 92.4\%\,^7\text{Li})$$

Challange

- **The mass spectrum for a sample of copper is shown below.**



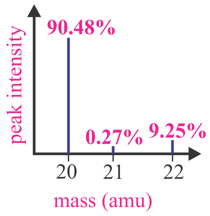
a) How many isotopes of copper were in the sample?

b) Give the relative isotopic mass of each isotope.

c) Give the relative isotopic abundance of each isotope.

Test your self

- Describe the relative charges and relative masses of protons, neutrons, and electrons.
- How does the atomic number of an element define its identity?
- If an atom has a nucleon number of 23 and an atomic number of 11, how many neutrons does it contain?
- If a proton and an electron enter an electric field at the same velocity, describe the direction in which each would deflect.
- Given an ion with an atomic number of 17, a mass number of 35, and a charge of -1, determines the number of protons, neutrons, and electrons in the ion.
- How would the number of neutrons change if the atomic number remains the same but the mass number increases by 2?
- Why might the relative atomic mass of an element be a non-integer value?
- How would you calculate the relative atomic mass of an element if you are provided with a mass spectrum showing multiple isotopic peaks?



Skill: 2.1

Objective:

Students will acquire skills in recognizing how protons, neutrons, and electrons behave when subjected to an electric field at the same velocity. They will learn to determine the specific counts of protons, neutrons, and electrons in atoms and ions based on given atomic numbers, mass numbers, and charges. Furthermore, students will perform calculations involving the relative atomic masses and isotopic abundances from mass spectra data, enhancing their analytical abilities in interpreting complex scientific data.

2.2 Knowledge

Exploring Electronic Configurations: Quantum Numbers, Subshells, and Their Applications in Technology

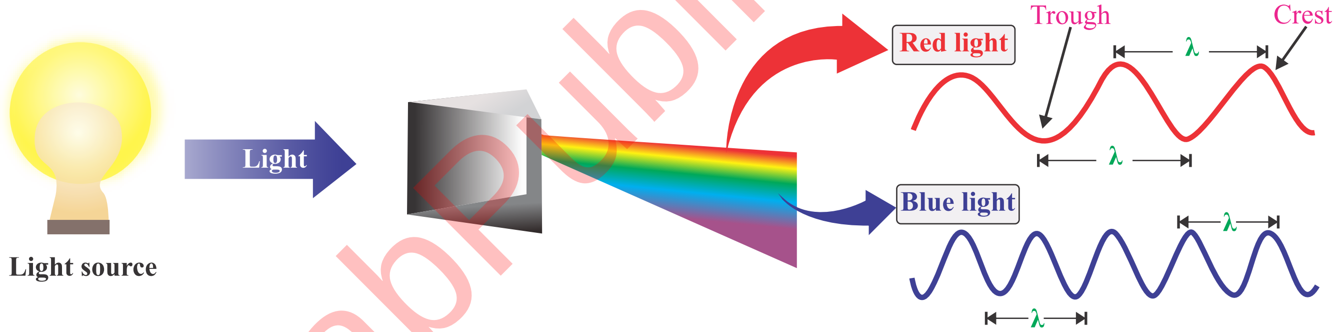
Student Learning Outcomes

- Describe that, each atomic shell and subshell are further divided into degenerate orbitals having the same energy
- Define terms related to electronic configuration (Some examples include shells, subshells, orbitals, principal quantum number (n), ground state),
- Relate Quantum Numbers to the Electronic Distribution of Elements.

- Describe the number of orbitals making up s, p d, and f sub-shells, and the number of electrons that can fill s, p d, and f subshells
- Apply the Aufbau principle, Pauli exclusion principle and Hund's rule to write the electronic configuration of elements
- Describe the order of increasing energy of the subshells (s, p, d, and f)
- Describe the electronic configurations to include the number of electrons in each shell, subshell and orbital.
- Explain the electronic configurations in terms of the energy of the electrons and inter-electron repulsion
- Determine the electronic configuration of atoms and ions given the proton or electron number and charge
- Illustrate the importance of electronic configurations in the development of new materials for electronic devices. (For example, semiconductors such as silicon have a specific electronic configuration that makes them ideal for use in electronic devices.)
- Describe the shapes of s, p, and d orbitals.
- Explain the concept of emission spectra Use the concept of emission spectra to deduce the electronic configuration of elements.

Electromagnetic Radiations (EMR)

Electromagnetic radiations (EMR) are waves of energy that travel through space at the speed of light. They encompass a wide range of wavelengths and frequencies, forming the electromagnetic spectrum. EMR includes radio waves, microwaves, infrared radiation, visible light, ultraviolet light, X-rays, and gamma rays. Each type of EMR has distinct characteristics and uses, from communication (radio waves) to medical imaging (X-rays).



Light passing through a prism is separated into a spectrum of colors we see in a rainbow.

The wavelength is the distance from a crest or trough on a wave to the next crest or trough.

Fig: 2.8: Wavelengths for light

Electromagnetic Spectrum

The electromagnetic spectrum is an arrangement of different types of electromagnetic radiation, ranging from longest to shortest wavelength. There is an inverse relationship between wavelength and frequency. The energy of electromagnetic radiation is directly proportional to the frequency and inversely proportional to the wavelength. Radio waves have long wavelengths and are used for various applications including AM and FM radio bands, cellular phones, and TV signals. Microwaves are also part of the electromagnetic spectrum.

Photons: Photons are small particles that make up light and heat. They have both wave and particle properties and travel at the speed of light. The energy levels of photons determine their wavelengths, with high-energy photons having shorter wavelengths and low-energy photons having longer wavelengths. Photons are crucial for technologies like lasers, which use a narrow range of wavelengths for various applications such as barcode scanning and medical imaging.

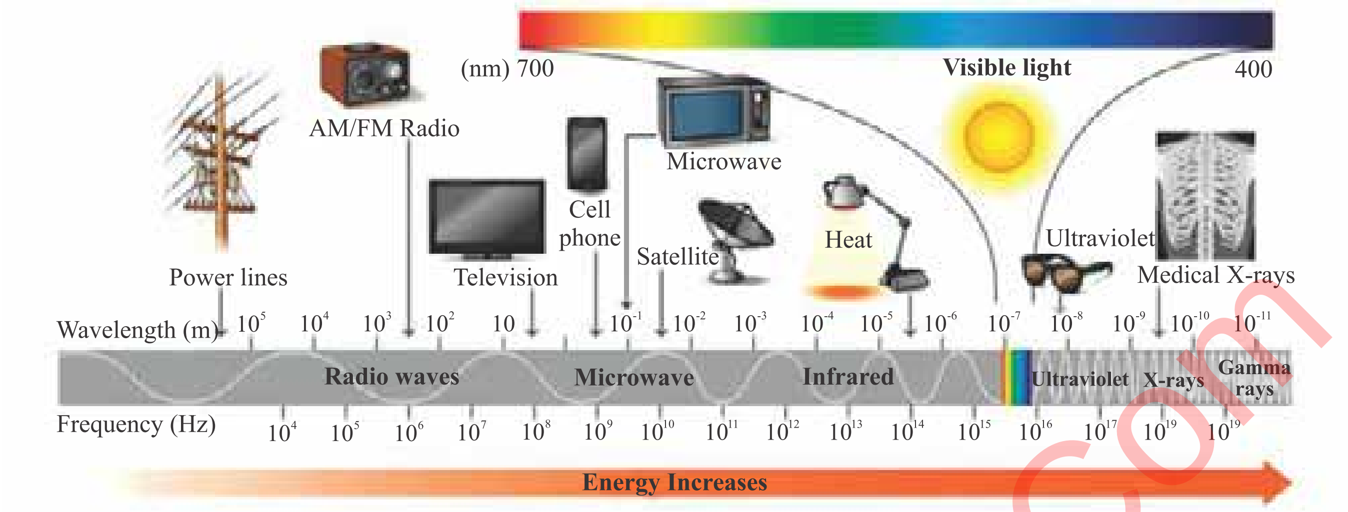


Fig: 2.9

Atomic Spectra

When light is emitted from heated elements and passed through a prism, it produces an atomic spectrum consisting of colored lines with dark areas in between. This separation of colors indicates that only particular wavelengths of light are produced when an element is heated, helping to identify each element uniquely.

Source of EMR in an Atom

In an atom, EMR is produced when electrons transition between different energy levels or shells. When an electron absorbs energy, it can move from a lower energy level (closer to the nucleus) to a higher energy level (further from the nucleus). This process is called excitation. Conversely, when an electron loses energy, it falls back to a lower energy level, emitting energy in the form of electromagnetic radiation. This emitted energy corresponds to specific wavelengths and frequencies, depending on the difference between the energy levels involved.

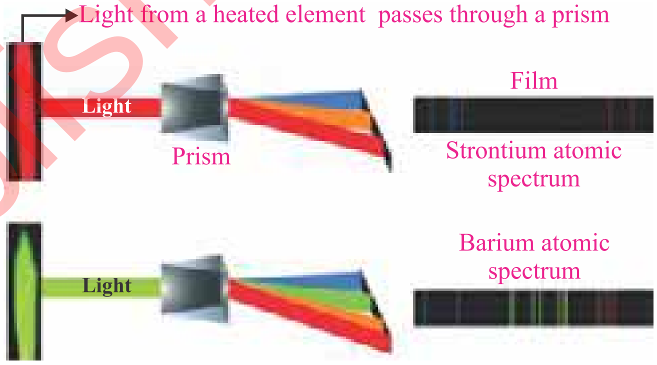


Fig: 2.10: In an atomic spectrum, light from a heated element separates into distinct lines.

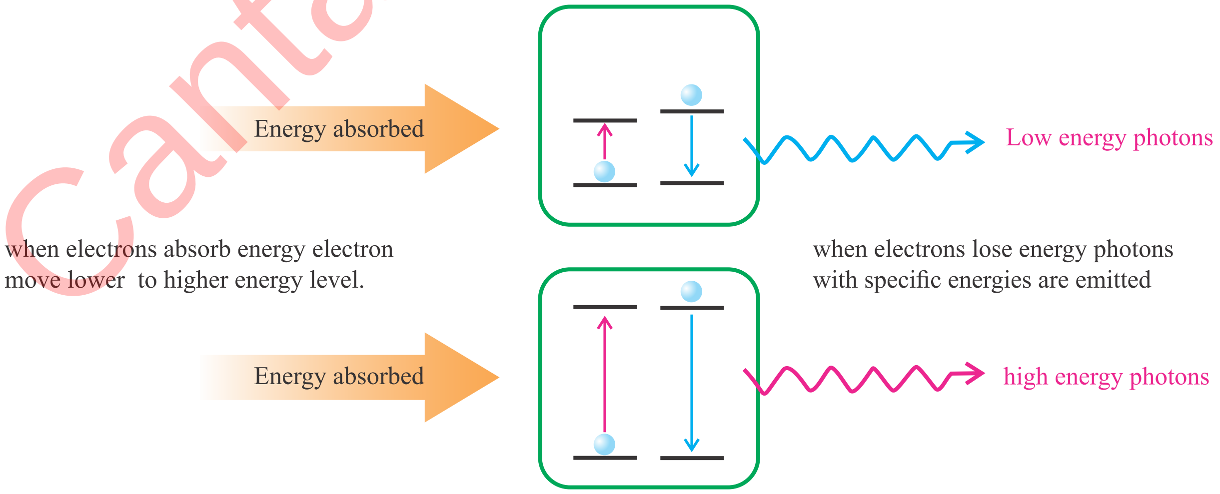


Fig: 2.11: Changes in energy levels

Subshells and Shells in an Atom

Atoms have distinct energy levels or shells, denoted by the principal quantum number (n). Each shell can contain one or more subshells, which are designated by the letters s, p, d, and f. These subshells have different shapes and capacities for holding electrons.

Shells (Energy Levels): These are the primary levels around the nucleus, represented by the principal quantum number n. The shells are labeled as K, L, M, N, etc., starting from the one closest to the nucleus.

Principal Quantum Number (***n***)
$1 < 2 < 3 < 4 < 5 < 6 < 7$

Lowest energy → Highest energy

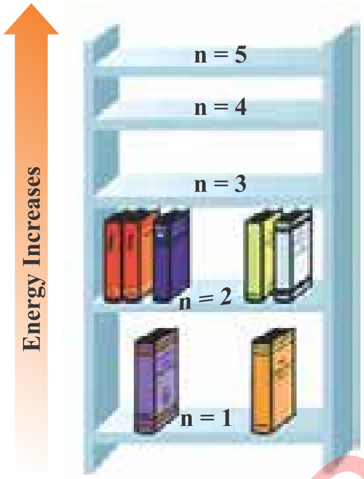


Fig: 2.12: Nucleus

Subshells: Within each shell, subshells (s, p, d, f) further define the energy levels of electrons. Each subshell has a different shape and a specific number of orbitals:

s subshell: 1 orbital, can hold 2 electrons
p subshell: 3 orbitals, can hold 6 electrons
d subshell: 5 orbitals, can hold 10 electrons
f subshell: 7 orbitals, can hold 14 electrons

Table 2.4

Energy level	Number of Sublevels	Types of Sublevels			
		s	*p*	*d*	*f*
n = 4	4	🟨	🟩🟩🟩	🟧🟧🟧🟧🟧	🟪🟪🟪🟪🟪🟪🟪
n = 3	3	🟨	🟩🟩🟩	🟧🟧🟧🟧🟧	
n = 2	2	🟨	🟩🟩🟩		
n = 1	1	🟨			

$s < p < d < f$

Lowest energy → Highest energy

Table 2.5

Energy Level (n)	1	2	3	4
n^2	$(1)^2$	$(2)^2$	$(3)^2$	$(4)^2$
Number of Orbitals	1	4	9	16

Relation of Subshells and Shells to EMR

When an electron transitions between subshells within the same shell or between different shells, it either absorbs or emits electromagnetic radiation. The energy change associated with these transitions determines the wavelength and frequency of the emitted or absorbed radiation. For instance, an electron moving from a 2p subshell to a 1s subshell releases a photon of light with a specific energy, producing visible or ultraviolet radiation. Conversely, an electron absorbing energy to jump from a 1s to a 2p subshell results in the absorption of specific wavelengths of light. These transitions and the resulting EMR are the basis for various spectroscopic techniques used to analyze the composition and properties of substances. Understanding the relationship between electron transitions, energy levels, and EMR helps explain the behavior of atoms and the nature of light and other forms of electromagnetic radiation.

Orbitals

There is no way to know the exact location of an electron in an atom. Instead, scientists describe the location of an electron in terms of probability. The orbital is the three-dimensional volume in which electrons have the highest probability of being found.

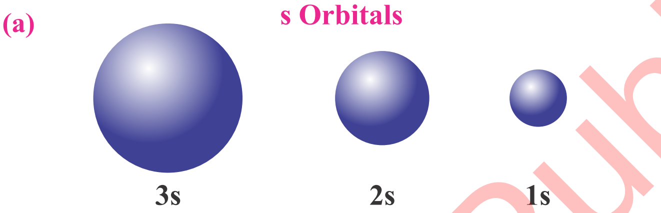
As an analogy

Imagine that you draw a circle with a 100-m radius around your chemistry classroom. There is a high probability of finding you within that circle when your chemistry class is in session. But once in a while, you may be outside that circle because you were sick

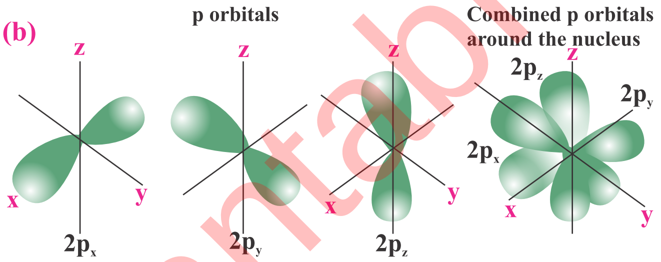
Shapes of Orbitals.

Each type of orbital has a unique three-dimensional shape. Electrons in an s orbital are most likely found in a region with a spherical shape. Imagine that you take a picture of the location of an electron in an s orbital every second for an hour. When all these pictures are overlaid, the result, called a probability density, would look like the electron cloud shown in figure 2.13(a). For convenience, we draw this electron cloud as a sphere called an s orbital. There is one s orbital for every energy level starting with n = 1. For example, in the first, second, and third energy levels, there are s orbitals designated as 1s, 2s, and 3s. As the principal quantum number increases, there is an increase in the size of the s orbitals, although the shape is the same see figure 2.13(a).

The orbitals occupied by p, d, and f electrons have three-dimensional shapes different from those of the s electrons. There are three p orbitals, starting with n = 2. Each p orbital has two lobes like a balloon tied in the middle. The three p orbitals are arranged in three perpendicular directions, along the x, y, and z axes around the nucleus (see Figure 2.13(b)). As with s orbitals, the shape of p orbitals is the same, but the volume increases at higher energy levels.



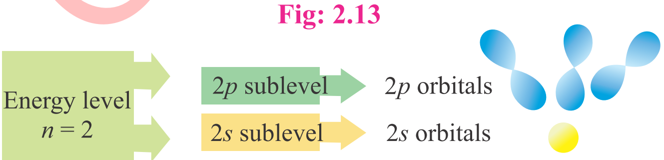
The sizes of the s orbitals increase because they contain electrons at higher energy levels.



The electron cloud of a p orbital shows the highest probability of finding a p electron.

The p orbitals in the same energy level are arranged on their axes around the nucleus.

Fig: 2.13



Energy level n = 3 consists of three sub-levels: s, p, and d. The d sub levels contain five d orbitals Energy level n = 4 consists of four sub levels s, p, d, and f. In the f sub level, there are seven f orbitals. The shapes of f orbitals are complex, and we have not included them in this text.

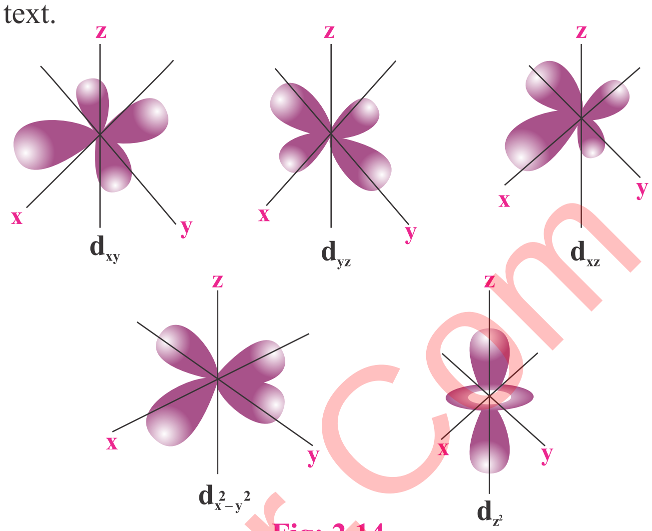
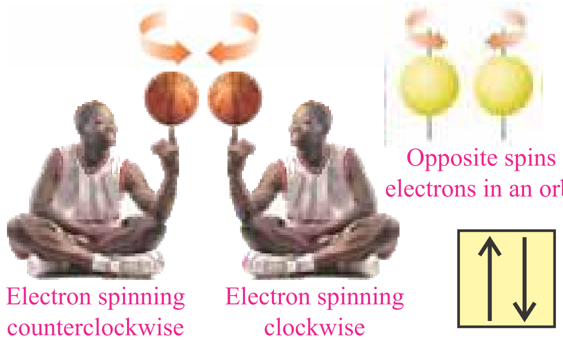


Fig: 2.14

Four of the five d orbitals consist of four lobes that are aligned along or between different axes. One d orbital consists of two lobes and a doughnut-shaped ring around its center.

Orbital Capacity and Electron Spin

The Pauli exclusion principle (detailed will be discuss in upcoming topics) states that each orbital can hold a maximum of two electrons. According to a model for electron behaviour, an electron is seen as spinning on its axis, which generates a magnetic field. When two electrons are in the same orbital, they will repel each other unless their magnetic fields cancel. This happens only when the two electrons spin in opposite directions. We can represent the spins of the electrons in the same orbital with one arrow pointing up and the other pointing down.



Opposite spins of electrons in an orbital

Electron spinning counterclockwise

Electron spinning clockwise

Fig: 2.15

Number of Electrons in Sublevels

There is a maximum number of electrons that can occupy each sublevel. An s sublevel holds one or two electrons. Because each p orbital can hold up to two electrons, the three p orbitals in a p sublevel can accommodate six electrons. A d sublevel with five d orbitals can hold a maximum of 10 electrons. With seven f orbitals, an f sub level can hold up to 14 electrons. As mentioned earlier, higher energy levels such as n = 5 , 6, and 7 would have 5, 6, and 7 sublevels, but those beyond sublevel f are not utilized by the atoms of the elements known today. The total number of electrons in all the sublevels adds up to give the electrons allowed in an energy level. The number of sublevels, the number of orbitals, and the maximum number of electrons for energy levels 1 to 4 are shown in table 1.2.

Table 2.6: The following table provides an overview of the possible quantum numbers.

Energy Level (n)	Number of Sublevels.	Type of Sublevel	Number of Orbitals	Maximum Number of Electrons	Total Electrons
1	1	1*s*	1	2	2
2	2	2*s*	1	2	
		2*p*	3	6	8
3	3	3*s*	1	2	
		3*p*	3	6	
		3*d*	5	10	18
4	4	4*s*	1	2	
		4*p*	3	6	
		4*d*	5	10	
		4*f*	7	14	32

Each orbital is specified by three interrelated quantum numbers:

- n, the principal quantum number
- ℓ, the angular momentum quantum number (sometimes called the azimuthal quantum number)
- mℓ, the magnetic quantum number

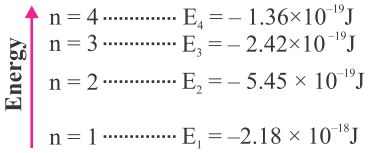
These quantum numbers all have integer values. A fourth quantum number, m_s, the spin quantum number specifies the orientation of the spin of the electron. We examine each of these quantum numbers individually .

The Principal Quantum Number (n)

The principal quantum number is an integer that determines the overall size and energy of an orbital. Its possible values are n = 1, 2, 3, 4 and so on. For the hydrogen atom, the energy of an electron in an orbital with quantum number n is given by the equation.

$$E_n = -2.18 \times 10^{-18} \times \left(\frac{1}{n^2}\right) J \quad (n=1,2,3,...)$$

Notice also that, as n increases, the spacing between the energy levels decreases as we will discuss in chapter 3 grade 9th.



The Angular Momentum Quantum Number (ℓ)

The angular momentum quantum number is an integer that determines the shape of the orbital. The possible values of ℓ are 0, 1, 2,…..(n – 1). In other words, for a given value of n, ℓ can be any integer (including 0) up to n - 1. For example, if n = 1, then the only possible value of ℓ is 0; if n = 2, the possible values of ℓ are 0 and 1. To avoid confusion between n and l, values of ℓ are often assigned letters as follows:

Table 2.7: The values of l beyond 3 are designated with letters in alphabetical order so that l = 4 is designated g, l = 5 is designated h, and so on these quantum no will be discuss in higher classes.

Value of /	Letter Designation
ℓ = 0	*s*
ℓ = 1	*p*
ℓ = 2	*d*
ℓ = 3	*f*

mℓ, the magnetic Quantum number (mℓ = 0, ±1, ±2, ±3, . . . , ±ℓ)

The magnetic quantum number, mℓ, is related to the orientation in space of the orbitals within a subshell. Orbitals in a given subshell differ in their orientation in space, not in their energy because they are de-generated orbitals (those orbitals which are same in energy).

The value of mℓ can range from +ℓ to -ℓ, with 0 included. For example, when ℓ = 2, mℓ can have five values: -2, -1, 0, +1, and +2. The number of values of mℓ for a given subshell (= 2ℓ + 1) specifies the number of orbitals in the subshell.

A Summary

Shells, Subshells, and Orbitals

Electrons in atoms are arranged in shells. Within each shell, there can be one or more electron subshells, each composed of one or more orbitals.

	Quantum Number
Shell	n
Subshell	*l*
Orbital	m_l

Shells and Subshells

Allowed values of the three quantum numbers are summarized in Table 2.8. By analyzing the sets of quantum numbers in this table, you will discover the following:

n = the number of subshells in a shell

2*l* + 1 the number of orbitals in a subshell = the number of values of m_ℓ.

n^2 the number of orbitals in a shell

The First Electron Shell, n = 1

When n = 1, the value of ℓ can only be 0, so me must also have a value of 0. This means that, in the shell closest to the nucleus, only one subshell exists, and that subshell consists of only a single orbital, the 1s orbital.

The Second Electron Shell, n = 2

When n = 2, l can have two values (0 and 1), so there are two subshells in the second shell. One of these is the 2s subshell (n = 2 and l = 0), and the other is the 2p subshell (n = 2 and 1). The values of m ℓ can be -1, 0, and +1 when ℓ = 1; three 2p orbitals exist. All three orbitals have the same shape. However, each has a different m ℓ value, that is, the three orbitals differ in their orientation in space.

The Third Electron Shell, n = 3

When n = 3, three subshells are possible for an electron; there are three values of ℓ = 0, 1, and 2. The first two subshells within the n = 3 shell are the 3s (ℓ = 0, one orbital) and 3p (ℓ = 1, three orbitals) subshells. The third subshell is labeled 3d (n = 3, ℓ = 2). Because m ℓ can have five values (-2, 1, 0, 1, and +2) for l = 2, there are five d orbitals in this d subshell.

The Fourth Electron Shell, n = 4

There are four subshells in the n = 4 shell. In addition to 4s, 4p, and 4d subshells, there is the 4f subshell for which ℓ = 3. Seven such orbitals exist because there are seven values of m ℓ when ℓ = 3 (–3, –2, –1,0, +1, +2, and +3).

Table 2.8: Summary of the Quantum Numbers, Their Interrelationships, and the Orbital Information Conveyed

Principal quantum number	Angular momentum quantum number	Magnetic quantum number	Number and type of orbitals in the subshell
Symbol = n values 1, 2, 3, ...	Symbol = ℓ Values = ...n – 1	Symbol = m ℓ Values = – ℓ ....0....+ ℓ	N= number of subshells No of orbitals in shell = n² No of Orbitals in sub shell = 2 ℓ +1
1	0	0	one 1s orbital (one orbital of one type in the n = 1 shell)
2	0 1	0 -1, 0, +1	one 2s orbital three 2p orbitals (four orbitals of two types in then n = 2 shell)
3	0 1 2	0 -1, 0, +1 -2, -1, 0, +1, +2	one 3s orbital three 3p orbitals five 3d orbitals nine orbitals of three types in the n = 3

4	0 1 2 3	0 - 1, 0, +1 - 2, - 1, 0, +1, +2 -3, - 2, - 1, 0, +1, +2, +3	one 4s orbital three 4p orbitals five 4d orbitals seven 4f orbitals 16 orbitals of four types in the n = 4

The Pauli exclusion principle

To make the quantum theory consistent with experiment, the Austrian physicist Wolfgang Pauli (1900–1958) stated in 1925 his **exclusion principle:** No more than two electrons can be assigned to the same orbital, and, if there are two electrons in the same orbital, they must have opposite spins. This leads to the general statement that no two electrons in an atom can have the same set of four quantum numbers (n, ℓ, mℓ, and m_s).

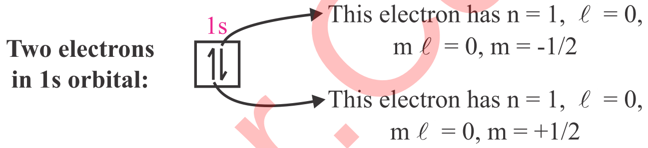
An electron assigned to the 1s orbital of the H atom may have the set of quantum numbers n = 1, ℓ = 0, mℓ = 0, and m_s = +1 ⁄ 2. If we represent an orbital by a box and the electron spin by an arrow (or   ), a representation of the hydrogen atom is then:

Electron in 1s orbital: $\boxed{\uparrow}$ Quantum number set

1s n = 1, $l = 0$, $m_l = 0$, m_s +1/2

The choice of m s (either + 1 ⁄ 2 or –1 ⁄ 2) and the direction of the electron spin arrow are arbitrary; that is, we could choose either value, and the arrow may point in either direction. Diagrams such as these are called **orbital box diagrams.**

A helium atom has two electrons. In the lowest energy (ground state) electronic configuration both electrons are assigned to the 1s orbital. The Pauli exclusion principle requires that each electron must have a different set of quantum numbers, so the orbital box diagram now is:



Do you Know

Orbitals are not boxes. Orbitals are not boxes in which electrons are placed. Thus, it is not conceptually correct to talk about electrons being in orbitals or occupying orbitals, although this is commonly done for the sake of simplicity.

By having opposite or "paired" spins, the two electrons in the 1s orbital of an He atom have different sets of the four quantum numbers. The relationship among the quantum numbers and the numbers of electrons is shown in Table

Table 2.9: Number of Electrons Accommodated in Electron Shells and Subshells with n = 1 to 6

Electron shell (n)	Subshells available	Orbitals available (2*l* + 1)	Number of electrons possible in subshell [2(2 + 1)]	Maximum electrons possible for nth shell ($2n^2$)
1	*s*	1	2	2
2	*s* *p*	1 3	2 6	8
3	*s* *p* *d*	1 3 5	2 6 10	18
4	*s* *p* *d* *f*	1 3 5 7	2 6 10 14	32
5	*s* *p* *d* *f* *g**	1 3 5 7 9	2 6 10 14 18	50

6	*s* *p* *d* *f* *g** *h**	1 3 5 7 9 11	2 6 10 14 18 22	72

*These orbitals are not occupied in the ground state of any known element detailed will be discuss in higher classes.

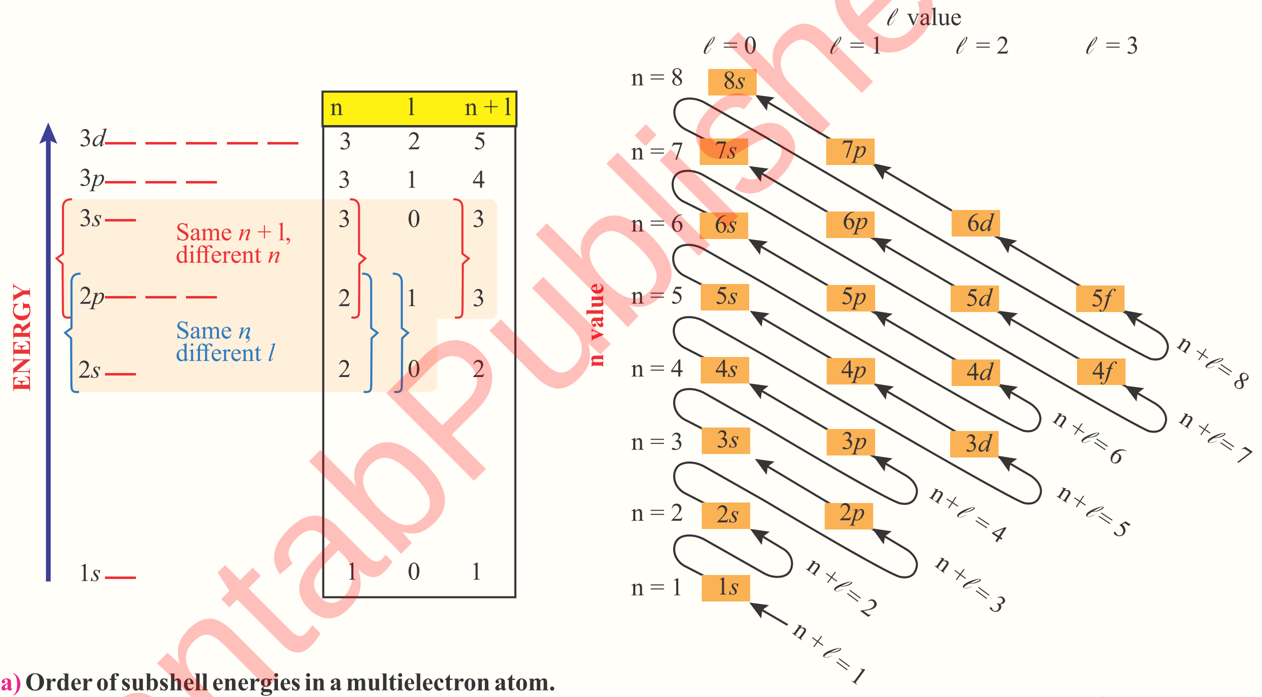
Atomic subshell energies and electron assignments

Our goal in this section is to understand and predict the distribution of electrons in atoms with many electrons. The procedure by which electrons are assigned to orbitals is known as the Aufbau principle (the German word Aufbau means "building up"). Electrons in an atom are assigned to shells (defined by the quantum number n) and subshells (defined by the quantum numbers n and ℓ) in order of increasingly higher energy. In this way, the total energy of the atom is as low as possible.

Order of Subshell Energies and Assignments

Based on theoretical and experimental studies of electron distributions in atoms, chemists have found there are two general rules that help predict these arrangements:

1. Electrons are assigned to subshells in order of increasing "n + ℓ" value.
2. For two subshells with the same value of "n + ℓ," electrons are assigned first to the subshell of lower **n**.



(a) Order of subshell energies in a multielectron atom. Energies of electron shells increase with increasing n, and, within a shell, subshell energies increase with increasing ℓ. (The energy axis is not to scale. The energy gaps between subshells of a given shell become smaller as n increases.)

(b) Subshell filling order. Subshells in atoms are filled in order of increasing n + l. When two subshells have the same n + l (value, the subshell of lower n is filled first. To use the diagram, begin at Is and follow the arrows of increasing n + l. (Thus, the order of filling is 1s 2s 2p 3s 3p 4s 3d and so on.)

Fig: 2.16: Summarizes the assignment of electrons according to increasing n + ℓ values. The discussion that follows explores the basis for this assignment of electron configurations

Hund's rule

When assigning electrons to p, d, or f orbitals, each successive electron is assigned to a different orbital of the subshell, and each electron has the same spin as the previous one, until the subshell is half full. Additional electrons must then be assigned to half-filled orbitals. This procedure follows Hund's rule, which states that the most stable arrangement of electrons in a subshell is that with the maximum number of unpaired electrons.

The order of stability while filling electrons in orbitals

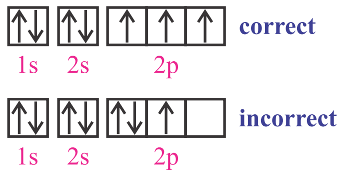
full filled > half filled > partial filled

Example: Nitrogen (N) and Oxygen (O) and elements of Groups 5a and 6a

An atom of nitrogen (Group 5A) has five valence electrons. Besides the two 2s electrons, it has three electrons, all with the same spin, in three different 2p orbitals.

Nitrogen: *spdf notation* $1s^22s^22p^3$ or $[He]2s^22p^3$

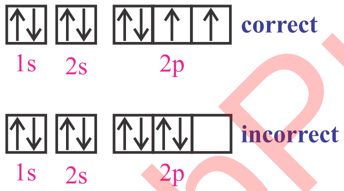
Box notation



An atom of oxygen (Group 6A) has six valence electrons. Two of these six electrons are assigned to the 2s orbital, and the other four electrons are assigned to 2p orbitals.

Nitrogen: **spdf notation** $1s^22s^22p^4$ or $[He]2s^22p^4$

Box notation

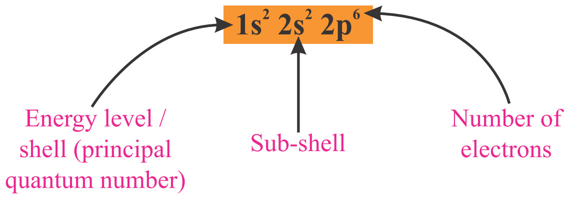


The fourth 2p electron must pair up with one already present. It makes no difference to which orbital this electron is assigned (the 2p orbitals all have the same energy), but it must have a spin opposite to the other electron already assigned to that orbital so that each electron has a different set of quantum numbers (the Pauli principle).

We begin our discussion of electron configuration

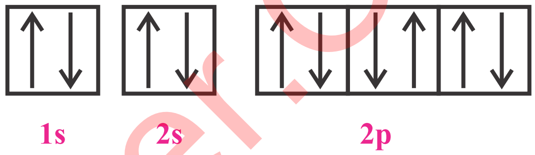
Showing electron configurations: The number of electrons that an atom or ion has, and how they are arranged, is called its electron configuration. Electron configurations can be shown in different ways. For example, an atom of neon has 10 electrons two electrons are in the 1s sub-shell, two are in the 2s sub-shell and six are in the 2p sub-shell. You can show this electron configuration in three ways.

1. Sub-shell notation



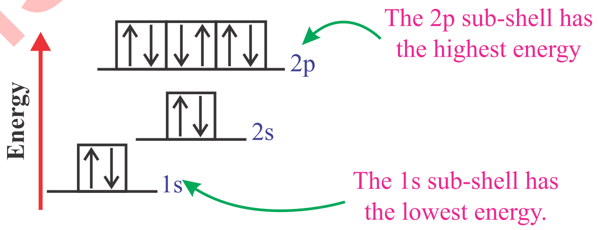
2. Arrows in boxes

Each of the boxes represents one orbital. Each of the arrows represents one electron. The up and down arrows represent the electrons spinning in opposite directions. Two electrons can only occupy the same orbital if they have opposite spin.



3. Energy level diagrams

These show the energy of the electrons in different orbitals, as well as the number of electrons and their arrangement.



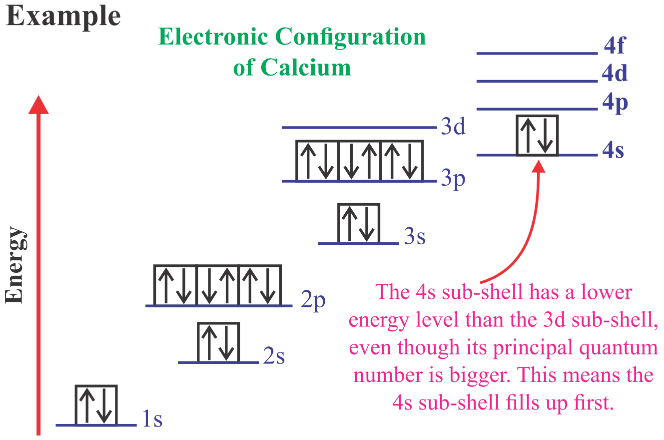
Working out electron configurations

You can figure out most electron configurations pretty easily, so long as you know a few simple rules:

Rule 1

Electrons fill up the lowest energy sub-shells first. As we discuss earlier in **Aufbau Principle**

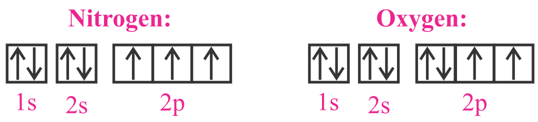
Example



Rule 2 (Hunds Rule)

Electrons fill orbitals in a sub-shell singly before they start sharing.

Example



Rule 3

For the configuration of ions from the s and p blocks of the periodic table, add or remove the electrons to or from the highest energy-occupied sub-shell

Mg atom: $1s^2\ 2s^2\ 2p\ 3s^2$ **Cl atom:** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Mg^{2+} ion: $1s^2\ 2s^2\ 2p^6$ **Cl^- ion:** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Shortened electron configurations:

Noble gas symbols in square brackets, such as [Ar], are sometimes used as shorthand in electron configurations. E.g. calcium ($1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6\ 4s^2$) can be written as $[Ar]4s^2$, where $[Ar] = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Electron configuration of transition metals:

Chromium (Cr) and copper (Cu) are badly behaved. They donate one of their 4s electrons to the 3d sub-shell. It's because they're happier with a more stable full or half-full d sub-shell as we discuss in hunds role. So, the electron configuration of a Cr atom is $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $3d^5$, $4s^1$ (not ending in $3d^4\ 4s^2$ as you'd expect). And the electron configuration of a Cu atom is $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $3d^{10}$, $4s^1$ (rather than finishing with $3d^9$, $4s^1$). **Here's another weird thing about transition metals: when they become ions, they lose their 4s electrons before their 3d electrons.**

Period 1: Hydrogen and Helium

We will draw orbital diagrams and write the corresponding electron configurations for the elements H and He in Period 1.

Table 2.10

Element	Atomic Number	Orbital Diagram	Electron Configuration
H	1	$1s^1$ ↑	$1s^1$
He	2	↑↓	$1s^2$

Period 2: Lithium to Neon

Table 2.11

Element	Atomic Number	Orbital Diagram	Electron Configuration	Abbreviated Electron Configuration
Li	**3**	1*s* 2*s* ↑↓ ↑	$1s^2$	$[He]2s^1$
Be	**4**	↑↓ ↑↓	$1s^22s^2$	$[He]2s^2$
B	**5**	2*p* ↑↓ ↑↓ ↑	$1s^22s^22p^1$	$[He]2s^22p^1$
C	**6**	↑↓ ↑↓ ↑ ↑	$1s^22s^22p^2$	$[He]2s^22p^2$
N	**7**	Unpaired electrons ↑↓ ↑↓ ↑ ↑ ↑	$1s^22s^22p^3$	$[He]2s^22p^3$
O	**8**	↑↓ ↑↓ ↑↓ ↑ ↑	$1s^22s^22p^4$	$[He]2s^22p^4$
F	**9**	↑↓ ↑↓ ↑↓ ↑↓ ↑	$1s^22s^22p^5$	$[He]2s^22p^5$
Ne	**10**	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	$1s^22s^22p^6$	$[He]2s^22p^6$

Period 3: Sodium to Argon

Table 2.12

Element	Atomic Number	Orbital Diagram (3s and 3p orbitals only)	Electron Configuration	Abbreviated Electron Configuration
Na	**11**	3*s* 2*p* [Ne] ↑ ☐☐☐	$1s^2 2s^2 2p^6 3s^1$	$[Ne] 3s^1$
Mg	**12**	[Ne] ↓↑ ☐☐☐	$1s^2 2s^2 2p^6 3s^2$	$[Ne] 3s^2$
Al	**13**	[Ne] ↓↑ ↑☐☐	$1s^2 2s^2 2p^6 3s^2 3p^1$	$Ne] 3s^2 3p^1$
Si	**14**	[Ne] ↓↑ ↑☐☐	$1s^2 2s^2 2p^6 3s^2 3p^2$	$Ne] 3s^2 3p^2$
P	**15**	[Ne] ↓↑ ↑ ↑ ↑	$1s^2 2s^2 2p^6 3s^2 3p^3$	$Ne] 3s^2 3p^3$
S	**16**	[Ne] ↓↑ ↓↑ ↑ ↑	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[Ne] 3s^2 3p^4$
CL	**17**	[Ne] ↓↑ ↓↑ ↓↑ ↑	$1s^2 2s^2 2p6 3s^2 3p^5$	$[Ne] 3s^2 3p^5$
Ar	**18**	[Ne] ↓↑ ↓↑ ↓↑ ↓↑	$1s^2 2s^2 2p^6 3s^2 3p^6$	$Ne] 3s^2 3p^6$

Electron Configurations and the Periodic Table

Up to now, we have written electron configurations using their energy level diagrams. However, the electron configurations of the elements are related to their position on the periodic table. Different sections or blocks within the periodic table correspond to the s, p, d, and f sub levels (see FIGURE). Therefore, we can "build" the electron configurations of atoms by reading the periodic table in order of increasing atomic number.

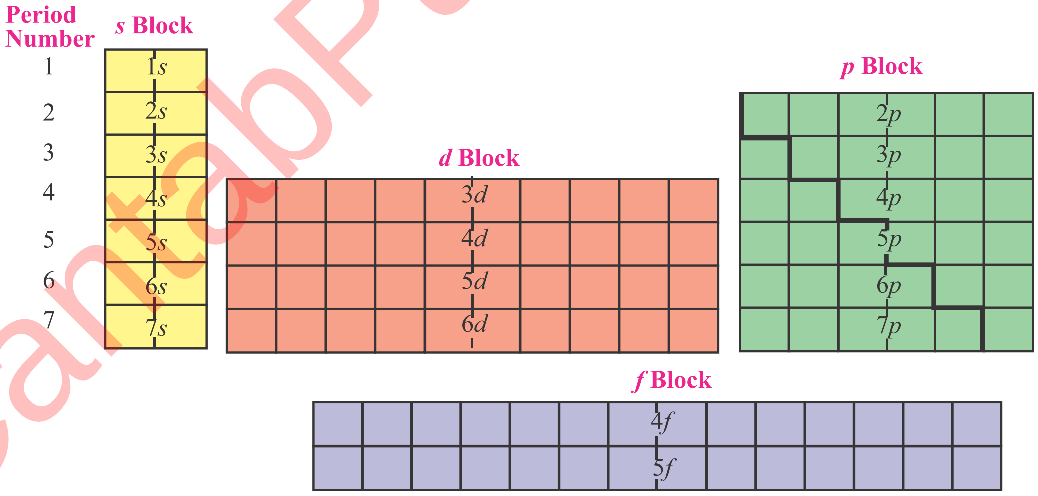


Fig: 2.17

Blocks on the Periodic Table

1. The s block includes hydrogen and helium as well as the elements in Group 1A (1) and Group 2A (2). This means that the final one or two electrons in the elements of the s block are located in an s orbital. The period number indicates the particular s orbital that is filling: 1s, 2s, and so on.

2. The p block consists of the elements in Group 3A (13) to Group 8A (18). There are six p-block elements in each period because three p orbitals can hold up to six electrons. The period number indicates the particular p sub-level

that is filling: 2p, 3p, and so on.

3. The d block, containing the transition elements, first appears after calcium (atomic number 20). There are 10 elements in each period of the d block because five d orbitals can hold up to 10 electrons. The particular d sub-level is one less (n - 1) than the period number. For example, in Period 4, the d block is the 3d sub-level. In Period 5, the d block is the 4d sub-level.

4. The f block, the inner transition elements, are the two rows at the bottom of the periodic table. There are 14 elements in each f block because seven f orbitals can hold up to 14 electrons. Elements that have atomic numbers higher than 57 (La) have electrons in the 4f block. The particular f sub-level is two less (n - 2) than the period number. For example, in Period 6, the f block is the 4f sub-level. In Period 7, the f block is the 5f sublevel.

Electron Configurations for Period 4 and Above

Period 4, we see that the 4s sublevel fills before the 3d sublevel. This occurs because the electrons in the 4s sublevel have slightly lower energy than the electrons in the 3d sublevel.

This order occurs again in Period 5, when the 5s sublevel fills before the 4d sublevel, in Period 6, when the 6s fills before the 5d, and in Period 7, when the 7s fills before the 6d. At the beginning of Period 4, the electrons in potassium (19) and calcium (20) go into the 4s sublevel. In scandium, the next electron added after the 4s sublevel is filled goes into the 3d block. The 3d block continues to fill until it is complete with 10 electrons at zinc (30). Once the 3d block is complete, the next six electrons, gallium to krypton, go into the 4p block

Table 2.13

Energy Lovel (n)	Atomic Number	Electron Configuration	Abbreviated Electron Configuration
4*s* Block			
K	19	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	[Ar] $4s^1$
Ca	20	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$	[Ar] $4s^2$
3*d* Block			
Se	21	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$	[Ar] $4s^2\ 3d^1$
Ti	22	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$	[Ar] $4s^2\ 3d^2$
V	23	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$	[Ar] $4s^2\ 3d^3$
*Cr	24	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$	[Ar] $4s^2\ 3d^5$ half-filled d sublevel is stable
Mn	25	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$	[Ar] $4s^2\ 3d^5$
Fe	26	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$	[Ar] $4s^2\ 3d^6$
Co	27	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$	[Ar] $4s^2\ 3d^7$
Ni	28	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$	[Ar] $4s^2\ 3d^8$
*Cu	29	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$	[Ar] $4s^2\ 3d^{10}$ filled d sublevel is stable
Zn	30	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$	[Ar] $4s^2\ 3d^{10}$
4*p* Block			
Ga	31	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$	[Ar] $4s^2\ 3d^{10}\ 4p^1$
Ge	32	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$	[Ar] $4s^2\ 3d^{10}\ 4p^2$
As	33	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$	[Ar] $4s^2\ 3d^{10}\ 4p^3$
Se	34	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$	[Ar] $4s^2\ 3d^{10}\ 4p^4$
Br	35	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$	[Ar] $4s^2\ 3d^{10}\ 4p^5$
Kr	36	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$	[Ar] $4s^2\ 3d^{10}\ 4p^6$
*Exceptions to the order of filling			

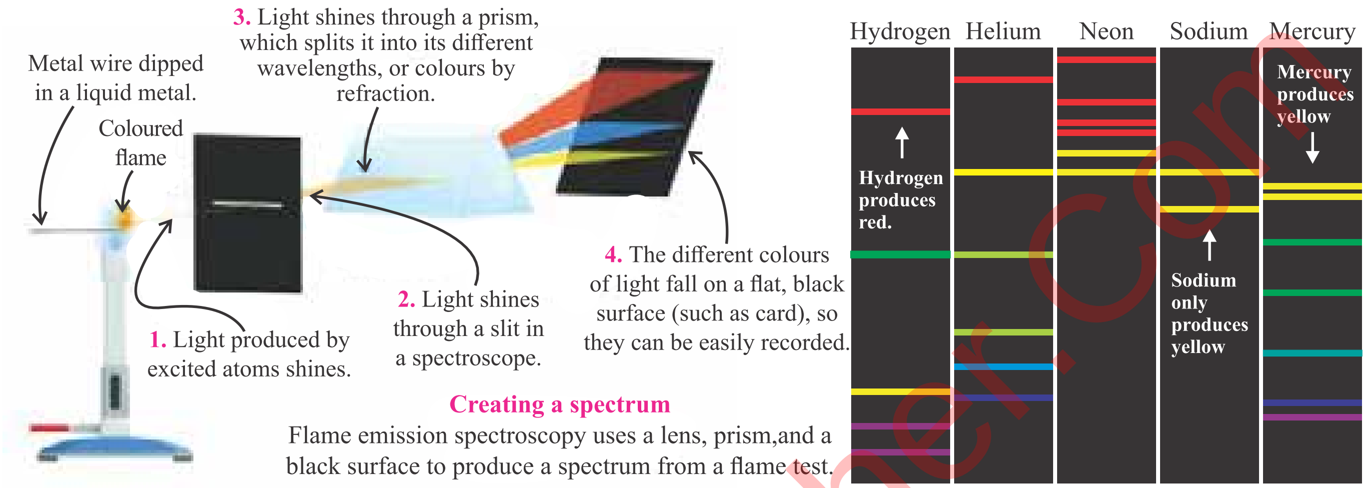
Concept of Emission Spectra

Emission spectra are produced when electrons in an atom absorb energy and move to higher energy levels, then release that energy as they return to their lower, original energy levels.

The light emitted during these electron transitions forms the emission spectrum of the element, which appears as discrete lines of different colors. Each line corresponds to a specific transition between energy levels and is unique to each element.

Deduction of Electronic Configuration Using Emission Spectra

The emission spectrum can be used to deduce the electronic configuration of an element by analyzing the wavelengths or frequencies of light emitted. These emissions correspond to specific electron transitions between energy levels, and from these transitions, the arrangement of electrons can be inferred.



Creating a spectrum

Flame emission spectroscopy uses a lens, prism,and a black surface to produce a spectrum from a flame test.

Fig: 2.18

Case Study: Sodium (Na) General Information:

Atomic Number of Sodium (Na): 11 . Electronic Configuration of Na: $1s^2 2s^2 2p^6 3s^1$

Sodium is well-known for its bright yellow doublet in its emission spectrum, which is especially prominent in street lighting.

Emission Spectrum Analysis

The most famous lines in the sodium emission spectrum are the double lines at about 589 nm. These lines are called the "Sodium D lines" and they are nearly identical in wavelength.These lines result from transitions in the outermost electron in the sodium atom.

Transitions Detail

Transition Process: The emission lines at 589 nm correspond to the transition of the electron from the 3p level back to the 3s level. When sodium atoms are excited, an electron in the 3s orbital can jump to the 3p orbital. As the electron returns to the 3s ground state, it emits light at this characteristic wavelength. as shown in Figure 2.18.

Energy Levels: The specific wavelengths help us determine the energy difference between the 3s and 3p levels in sodium.

Deduction of Electronic Configuration

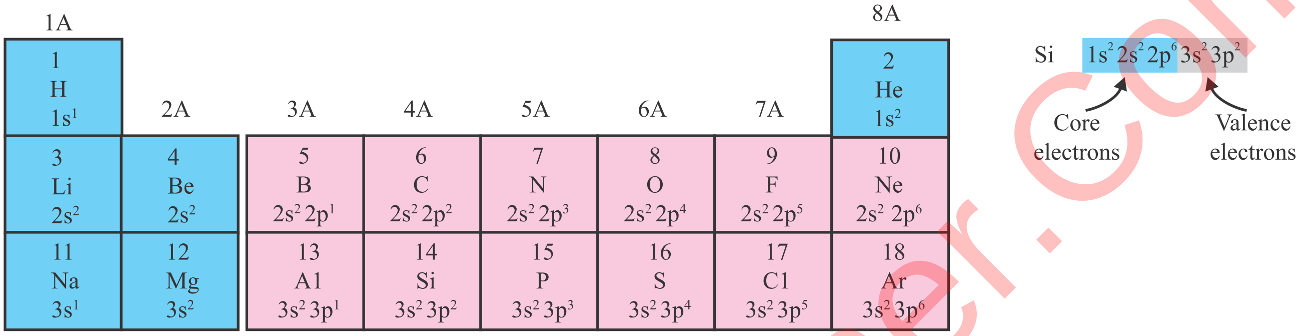
Given the transitions observed (from 3p to 3s), it is clear that the ground state configuration of sodium has a single electron in the 3s orbital. This single electron in the outermost shell ($3s^1$) defines many of the chemical properties of sodium, including its reactivity and its tendency to lose one electron and form Na^+ ions. The detailed analysis of these spectral lines not only confirms the ground state configuration ($1s^2 2s^2 2p^6 3s^1$) but also provides insights into the energy levels and the spacing between them in sodium atoms.

Conclusion

The emission spectrum of sodium, particularly the bright yellow doublet, is a direct result of the electron configuration and the behavior of electrons in the outer shell. By studying these spectral lines, scientists can deduce not only the electronic configuration but also understand better the quantum mechanical nature of electron energy levels. This case of sodium exemplifies how emission spectra serve as a powerful tool in atomic and quantum physics to explore and confirm the structure of atoms.

Electron Configurations, Valence Electrons, and the Periodic Table

Mendeleev arranged the periodic table so that elements with similar chemical properties are placed in the same column. By superimposing the electron configurations of the first 18 elements onto a partial periodic table, we can begin to understand the connection between an element's properties and its electron configuration. As we move right across a row, the orbitals fill in the correct order, and with each subsequent row, the highest principal quantum number increases by one. When moving down a column, the number of electrons in the outermost principal energy level remains the same. The outermost electrons are crucial in bonding and are known as an atom's valence electrons. For main-group elements, the valence electrons are those in the outermost principal energy level, while for transition elements, we count the outermost d electrons among the valence electrons. Elements in a column of the periodic table have similar chemical properties as they have the same number of valence electrons. We differentiate between valence electrons and other electrons in an atom, which are called core electrons. For instance, silicon has four valence electrons (in the n = 3 principal level) and ten core electrons.



1A	2A	3A	4A	5A	6A	7A	8A
1 H $1s^1$							2 He $1s^2$
3 Li $2s^2$	4 Be $2s^2$	5 B $2s^2 2p^1$	6 C $2s^2 2p^2$	7 N $2s^2 2p^3$	8 O $2s^2 2p^4$	9 F $2s^2 2p^5$	10 Ne $2s^2 2p^6$
11 Na $3s^1$	12 Mg $3s^2$	13 A1 $3s^2 3p^1$	14 Si $3s^2 3p^2$	15 P $3s^2 3p^3$	16 S $3s^2 3p^4$	17 C1 $3s^2 3p^5$	18 Ar $3s^2 3p^6$

Fig: 2.19: Outer Electron Configurations of the First 18 Elements in the Periodic Table

Role of Electronic Configuration in Determination of Type of Semiconductor

In silicon, the outermost energy level is the third energy level, composed of the 3s and 3p sublevels, which together can hold up to 8 electrons. However, in its ground state, silicon only has 4 valence electrons occupying these sublevels. This electronic configuration imparts unique properties to silicon, making it a pivotal material in semiconductor technology.

Silicon's electronic configuration leads to the formation of a band structure consisting of the valence band and the conduction band. In the valence band, electrons are tightly bound to their respective atoms, restricting their movement. The conduction band, on the other hand, is typically devoid of electrons in the ground state. The energy gap between these bands, known as the band gap, is relatively small in silicon. When sufficient energy is provided, electrons can be excited from the valence band to the conduction band, creating electron-hole pairs. This process, termed excitation, is fundamental to the functionality of semiconductors.

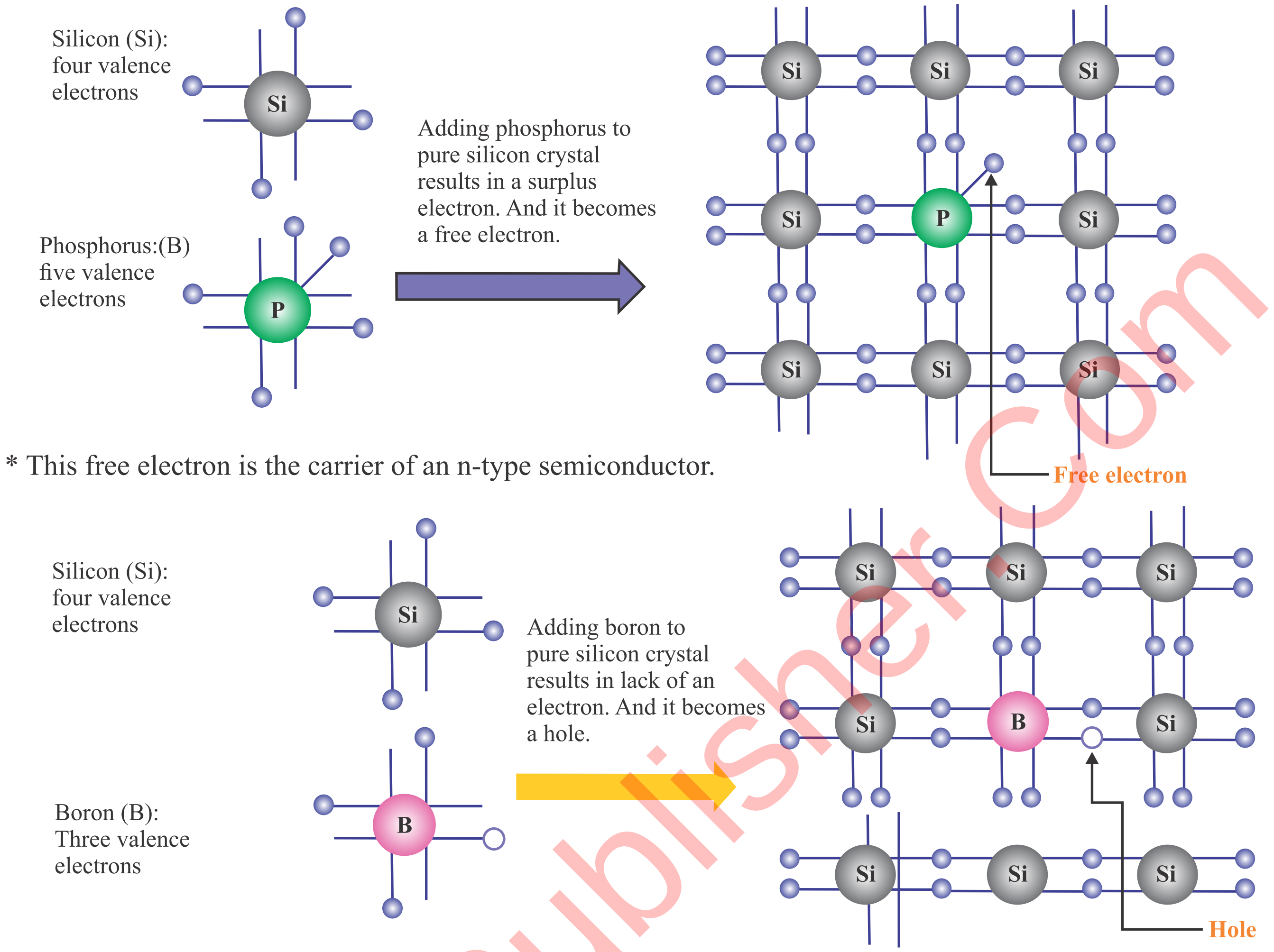
Semiconductors are classified as intrinsic or extrinsic. Intrinsic semiconductors are composed of pure elements like silicon or germanium, where the electrical properties are inherent to the material itself. Extrinsic semiconductors, however, are deliberately infused with impurities, known as dopants, to enhance their electrical conductivity. These dopants introduce either additional electrons (n-type) or create "holes" (p-type) that behave like positive charge carriers.

Types of Doping

N-Type Semiconductors: In n-type semiconductors, dopants such as phosphorus or arsenic are introduced. These elements have five valence electrons, one more than silicon. The extra electron becomes available for conduction, significantly enhancing the material's electrical conductivity. This additional electron elevates the semiconductor's energy level, enabling easier movement of charge carriers within the conduction band.

P-Type Semiconductors: In p-type semiconductors, dopants such as boron or gallium are used. These elements have three valence electrons, one less than silicon. The absence of one electron creates a "hole" in the valence band. These holes act as positive charge carriers, as they attract electrons from neighboring atoms, facilitating charge movement within the material.

The electronic configuration of silicon allows it to function as either an n-type or p-type semiconductor,



* This free electron is the carrier of an n-type semiconductor.

Fig: 2.20

depending on the doping agent used. This versatility is crucial in semiconductor device fabrication, enabling the creation of p-n junctions, which are the building blocks of many electronic components.

Applications of Semiconductors

By combining n-type and p-type materials, engineers can construct diodes, transistors, and other electronic components essential for controlling electrical currents. Diodes, for instance, allow current to flow in one direction only, making them integral to rectifying circuits. Transistors, which can amplify or switch electronic signals, are fundamental in all modern electronic devices, from computers to mobile phones.

Conclusion

Understanding the role of electronic configuration in determining the type of semiconductor is essential for grasping the underlying principles of modern electronics. Silicon's ability to be doped with various elements to modify its electrical properties forms the basis of semiconductor technology. This capability is exploited in creating devices that are indispensable in today's technological landscape, driving innovation and advancement in electronics and communication.

Skill: 2.2

Objective:

Students will develop skills to apply foundational principles such as the Aufbau principle, Pauli exclusion principle, and Hund's rule to accurately write the electronic configurations of elements. They will be trained to determine the electronic configurations of both atoms and ions based on given atomic numbers and charges. Furthermore, students will analyze emission spectra to deduce electronic structures, enhancing their understanding of quantum mechanics in practical scenarios. They will also learn to describe the spatial shapes of s, p, and d orbitals, equipping them with the necessary tools to interpret and predict chemical and physical properties of elements.

Test your self

- Describe what is meant by degenerate orbitals and provide an example within a subshell.
- Define the terms shell, subshell, orbital, and principal quantum number (n) in the context of electronic configuration.
- What is the ground state of an atom and how does it relate to its electronic configuration?
- Relate the quantum numbers to the position of an electron in an atom using an example of a specific element.
- Describe the number of orbitals in the s, p, d, and f subshells and the maximum number of electrons each can hold.
- Write the electronic configuration of an element with atomic number 14 using the Aufbau principle, Pauli Exclusion Principle, and Hund's rule.
- Describe the order of increasing energy for the s, p, d, and f subshells.
- How electronic configurations reflect the energy of electrons and the concept of inter-electron repulsion?
- Determine the electronic configuration of an ion with a proton number of 19 and a charge of +1.
- Use the concept of emission spectra to deduce the electronic configuration of an element like sodium.

2.3 Knowledge

Understanding Periodic Trends: Atomic Structure, Ionization Energies, and Electronic Configurations

Student Learning Outcomes

- Explain the change in atomic and ionic radius across a period and down a group
- Determine the electronic configuration of elements and their ions with proton numbers. (Some examples include a) simple configurations e.g. 2,8,b. subshells e.g. Is2, 2s2, $2p^6$, $2s^1$

  c) students should be able to determine both of these from the periodic table and are not required to memorize these d. students should understand that the chemical properties of an atom are governed by valence electrons)
- Describe a free radical as a species with one or more unpaired electrons.
- Explain that ionization energies are due to the attraction between the nucleus and the outer electron.
- Explain how ionization energy helps account for the trends across a period and down a group of the Periodic Table.
- Account for the variation in successive ionization energies of an element.
- Explain the factors influencing the ionization energies of elements in terms of nuclear charge, atomic/ionic radius, shielding by inner shells and subshells, and spin pair repulsion.
- Deduce the electronic configurations of elements using successive ionization energy data.
- Deduce the position of an element in the Periodic Table using successive ionization energy data.

The Sizes of Atoms and Ions

The sizes of atoms and ions are typically measured in terms of their radii. For elements that occur in nature as diatomic molecules such as Cl_2 and O_2, their atomic radius is simply half the distance between the centers of the nuclei in the molecule (as shown in the figure 2.21(a)). For metals, their atomic radius, also known as metallic radius, is half the distance between the centers of the nuclei in the solid metal like sodium (as shown in the figure 2.21(b)). The values of ionic radii are obtained from the distances between the centers of the nuclei in solid ionic compounds (as shown in the figure 2.21(c)).

The relative sizes of atoms follow periodic trends that are also shown in the figure. Several factors contribute to these trends.

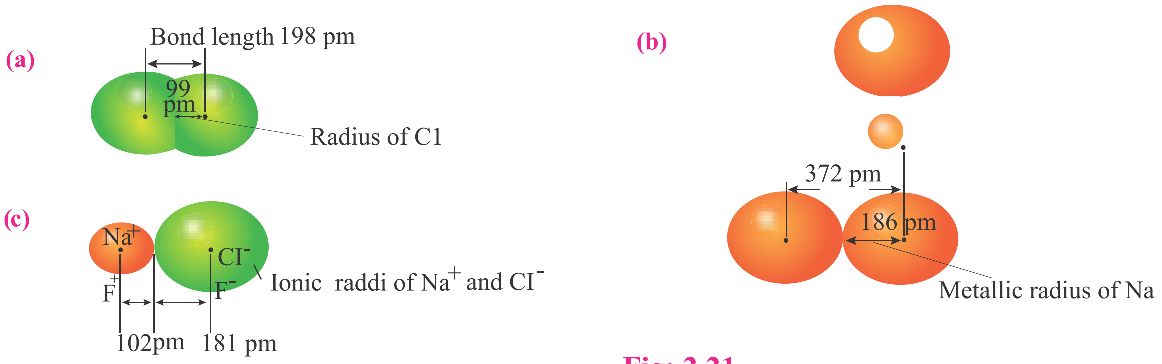


Fig: 2.21

Trends of Atomic Radius

Across a Period

As you progress through a period from left to right on the periodic table, the atomic radius diminishes. This is due to the escalating number of protons in the nucleus, resulting in an increased nuclear charge. The heightened positive charge exerts a stronger attraction on the electrons, drawing them nearer to the nucleus. Despite the fact that electrons are being added to the same principal energy level, the amplified nuclear charge is not entirely counteracted by electron-electron repulsions, ultimately causing a reduction in atomic size.

Down a Group

As you descend a group on the periodic table, the atomic radius expands. This phenomenon occurs because additional electrons occupy new principal energy levels, which are situated farther from the nucleus. Consequently, each successive energy level is larger than its predecessor, resulting in an overall increase in atomic size. Moreover, the inner electron shells create a shielding effect, diminishing the effective nuclear charge experienced by the outermost electrons. This enables them to exist at a greater distance from the nucleus.

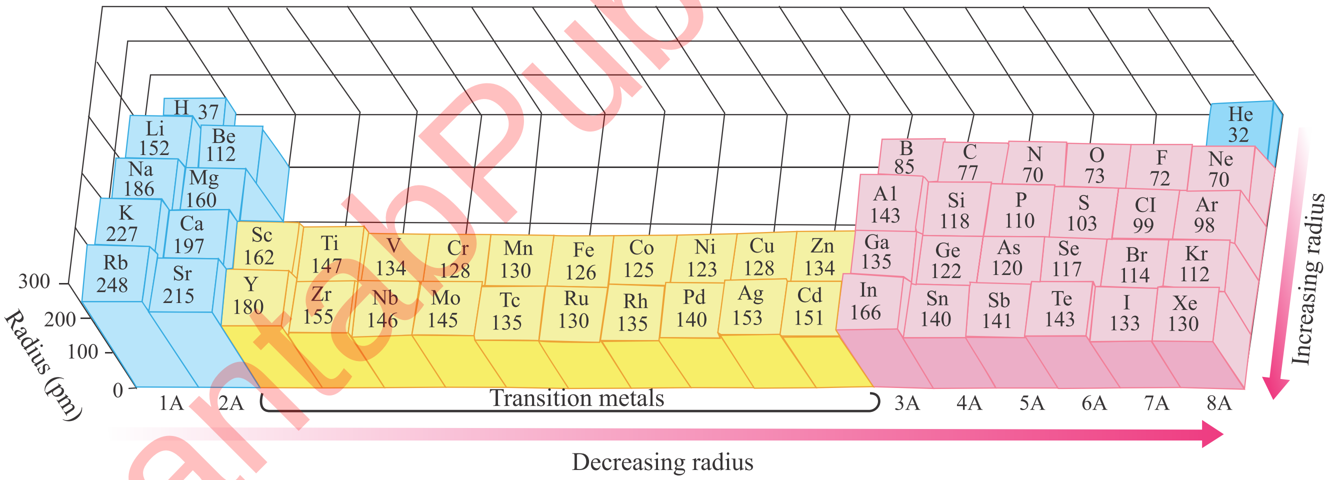


Fig: 2.22 Trends in Atomic Radius: In general, atomic radii increase as we move down a column and decrease as we move to the right across a period in the periodic table.

Trends of Ionic Radius

Across a Period

Cations (positively charged ions): Metals on the left side of the periodic table typically lose electrons to form cations. This loss of electrons reduces electron-electron repulsion and often leads to a decrease in the ionic radius due to the loss of an entire electron shell. The size of cation is always smaller than its neutral parent atom as shown in figure 2.23. Anions (negatively charged ions): Nonmetals on the right side of the period gain electrons to form anions. While gaining electrons increases electron-electron repulsion within the same shell, the overall increase in nuclear charge across the period leads to a smaller ionic radius for anions. The size of anion is always larger than its parents atom as shown in given figure 2.24.

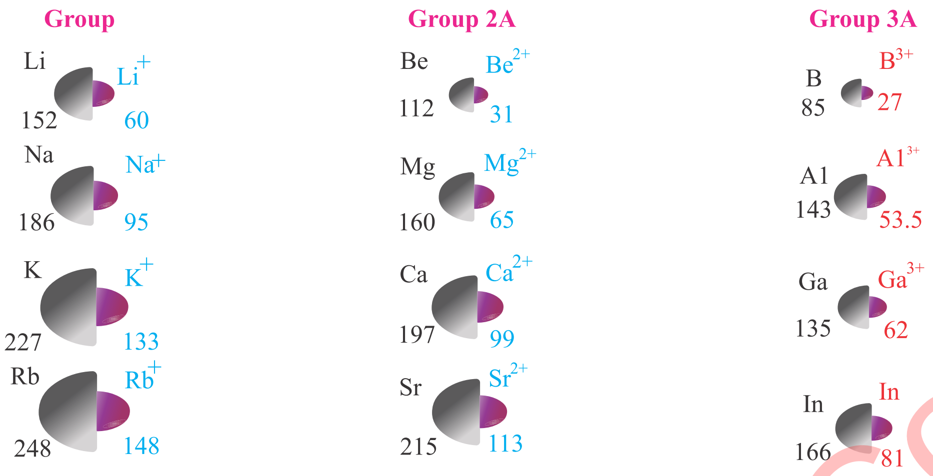


Fig 2.23: This view shows a comparative representation of the atomic radii of various atoms and their corresponding cations, measured in picometers (pm). The elements are grouped into three categories: Group 1A, Group 2A, and Group 3A. Each atom is depicted with a half-circle diagram where the larger gray portion represents the neutral atom, and the smaller purple portion represents the cation. The specific measurements of atomic radii for both the neutral atoms and their cations are listed next to each diagram. The visual representation clearly shows that cations are significantly smaller than their corresponding neutral atoms, demonstrating the reduction in size upon losing electrons and forming positively charged ions.

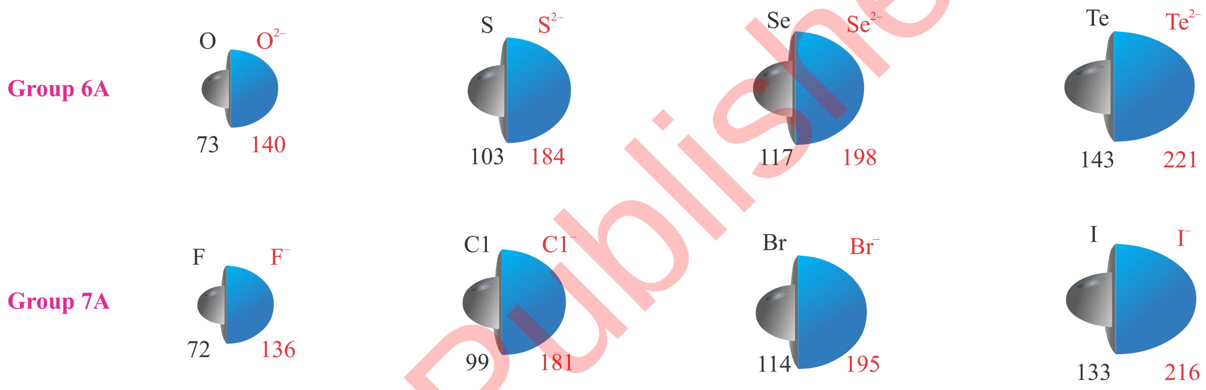


Fig 2.24: The image provides a comparative representation of the atomic radii of various atoms and their corresponding anions, measured in picometers (pm). The elements are grouped into two categories: Group 6A and Group 7A. Each atom is depicted with a half-circle diagram where the larger green portion represents the anion, and the smaller gray portion represents the neutral atom. The specific measurements of atomic radii for both the neutral atoms and their anions are listed next to each diagram. The visual representation clearly shows that anions are significantly larger than their corresponding neutral atoms, demonstrating the increase in size upon gaining electrons and forming negatively charged ions.

Down a Group

As you move down a group, both cations and anions increase in ionic radius similar to the atomic radius, the addition of electron shells as you move down a group increases the size of the ion. The shielding effect of inner electrons reduces the effective nuclear charge on the outermost electrons, allowing the ion to be larger.

Ionization Energy

The ionization energy (IE) of an atom or ion is the energy required to remove an electron from the atom or ion in the gaseous state. The ionization energy is consistently positive due to the fact that extracting an electron always requires energy. **(This process is akin to an endothermic reaction, which absorbs heat and consequently has a positive ΔH.)** The energy needed to extract the initial electron is known as the first ionization energy (IE1). For example, we represent the first ionization of sodium with the equation:

$$Na_{(g)} \longrightarrow Na^{+}_{(g)} + 1e^{-} \quad I.E_{1} = 496\ kJ/mol$$

The energy needed to eliminate the second electron is known as the second ionization energy (IE_2), and the energy needed to eliminate the third electron is referred to as the third ionization energy (IE_3), and so forth. The second ionization energy of sodium is denoted as:

$$\mathrm{Na^{+}_{(g)} \rightarrow Na_{2}+_{(g)} + 1\ e^{-}\ I.E_{2} = 4560\ kJ/mol}$$

It's important to note that the second ionization energy doesn't denote the energy needed to remove two electrons from sodium (that value is the sum of IE1 and IE2), but rather the energy required to remove one electron from Na+. We analyze trends in IE1 and IE2 individually.

Trends in First Ionization Energy

The trend for first ionization energy typically involves a decrease as we descend a group in the periodic table, as electrons in the outermost principal level are progressively more distant from the positively charged nucleus and consequently have weaker binding. Conversely, as we move horizontally across a period, in the periodic table, the first ionization energy generally rises due to the outermost electrons experiencing a higher effective nuclear charge (Zeff), resulting in stronger binding.

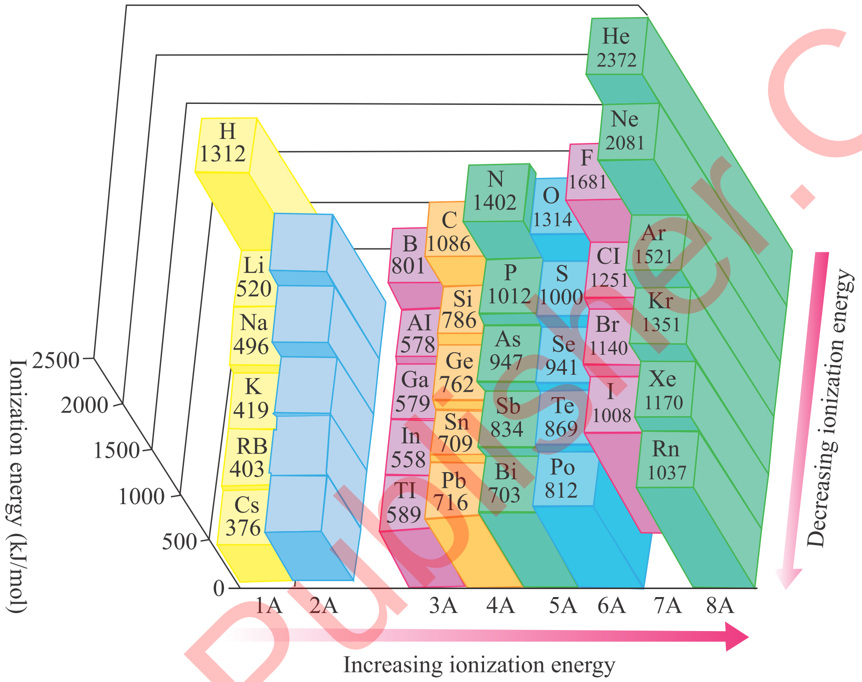
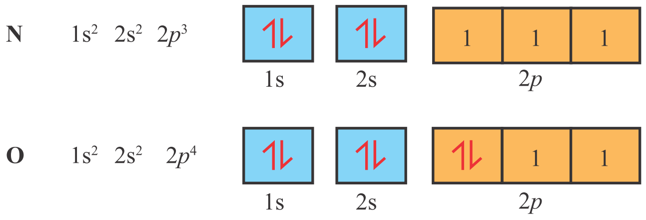


Fig: 2.25: Trends in Ionization Energy: First ionization energy increases as we move to the right across the period and decreases as we move down group in the periodic table.

Exceptions to trends in first ionization energy

Upon careful examination of figure 2.25 a, we can observe certain deviations from the trends in first ionization energies. For instance, boron exhibits smaller ionization energy than beryllium, despite being positioned to the right of beryllium in the same row. This anomaly is attributed to the disparity between the s block and the p block. Specifically, the 2p orbital penetrates into the nuclear region less than the 2s orbital. Consequently, the 1s electrons shield the electron in the 2p orbital from nuclear charge to a greater extent than they shield the electrons in the 2s orbital. As a result, the 2p orbitals possess higher energy, making the removal of the electron easier (resulting in lower first ionization energy). Similar exceptions arise for aluminum and gallium, both of which are situated directly below boron in group 3A.

There is another exception between nitrogen and oxygen. Despite oxygen being located to the right of nitrogen in the same row, it exhibits lower first ionization energy. This anomaly arises from the repulsion between electrons when they occupy the same orbital. Examine the electron configurations and orbital diagrams of nitrogen and oxygen shown here:



Nitrogen possesses three electrons distributed among three p orbitals, while oxygen has four electrons. In the case of nitrogen, the 2p orbitals are half-filled, leading to an exceptionally stable configuration. Conversely, one of oxygen's four 2p electrons must pair with another electron, resulting in an easier removal and a less stable configuration. Similar exceptions for analogous reasons can be observed for sulfur and selenium, which are positioned directly below oxygen in group 6A.

Trends in second and successive ionization energies

Observe the patterns in the first, second, and third ionization energies of sodium (group 1A) and magnesium (group 2A) as illustrated in the figure 2.26. Notably, a substantial increase is evident in the first and second ionization energies of sodium, while for magnesium, the ionization energy approximately doubles from the first to the second, followed by a significant leap between the second and third ionization energies. What accounts for these sharp increases? A comprehensive understanding of these trends can be attained by scrutinizing the electron configurations of sodium and magnesium

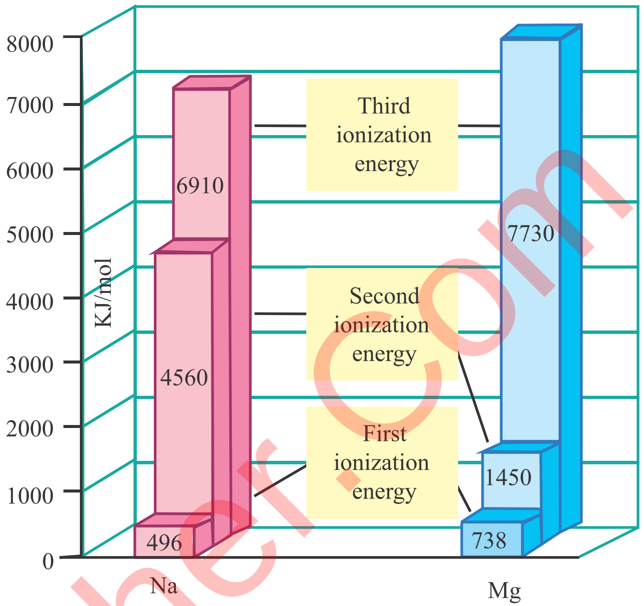


Fig: 2.26

Na [Ne] $3s^1$, Mg [Ne] $3s^2$

The initial ionization of sodium entails the removal of the valence electron in the 3s orbital. It is important to note that these valence electrons are more loosely held than the core electrons, resulting in the ion obtaining a noble gas configuration, which is known for its stability. Consequently, the first ionization energy is relatively low. On the other hand, the second ionization of sodium involves removing a core electron from an ion with a noble gas configuration, necessitating a significant amount of energy and resulting in a high value of IE2. Similarly, the first ionization of magnesium involves the removal of a valence electron in the 3s orbital, requiring slightly more energy than the corresponding ionization of sodium due to the discussed trends in Zeff (Zeff increases as we move right across a row). The second ionization of magnesium also entails removing an outer electron in the 3s orbital, but this time from an ion with a 1+ charge, instead of from a neutral atom. Removing the third electron from a magnesium ion with a noble gas configuration demands approximately double the energy needed to remove an electron from a neutral atom. Similarly, the third ionization of magnesium is comparable to the second ionization of sodium.

This process involves the removal of a core electron, requiring a substantial amount of energy and resulting in a high value for IE.

Table 2.14: Illustrates similar trends in the successive ionization energies of various elements. The ionization energy generally rises consistently with each subsequent removal of an outermost electron, but experiences a significant spike when the first core electron is removed.

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	496	4560					
Mg	738	1450	7730		Core electrons		
AI	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,100		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
CI	1251	2300	3820	5160	6540	9460	11,000
Ar	1521	2670	3930	5770	7240	8780	12,000

Free radical

A free radical is a chemical species that contains one or more unpaired electrons in its outer shell, making it highly reactive and unstable. These unpaired electrons are what give free radicals their characteristic properties, as they seek stability through electron pairing. This quest for electron pairing drives free radicals to engage in chemical reactions with other atoms or molecules, often causing chain reactions that can propagate rapidly through a substance.

Key Features of Free Radicals:

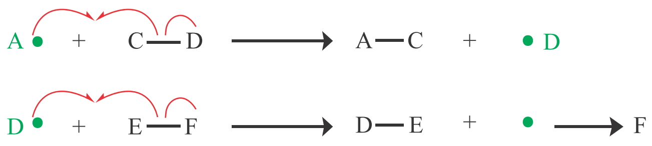
Free radicals are chemical species characterized primarily by the presence of one or more unpaired electrons in their outer shell. This unique feature makes them highly reactive and quite unstable compared to other chemical species that have paired electron configurations, which generally confer stability. The high reactivity of free radicals allows them to easily participate in chemical reactions by either donating or accepting electrons to achieve a more stable electron configuration. This property is crucial for numerous chemical processes, including combustion, polymerization, and various biological functions such as cell signaling.

Free radicals can be formed through several mechanisms. One common method is homolytic bond cleavage, where a bond between two atoms breaks evenly, and each atom retains one of the previously shared electrons, resulting in the formation of two free radicals.

Another source of free radicals is radiation exposure, which can provide enough energy to separate paired electrons within molecules. Additionally, certain redox reactions that involve the transfer of single electrons can also lead to the creation of free radicals as intermediate species.

Initiation:

Propagation:



Termination:

Role of free radicals

In both chemistry and biology, free radicals serve essential roles. They are indispensable for initiating polymerizations, oxidizing metals, and synthesizing various chemicals. Biologically, free radicals have dual roles, being both beneficial and harmful. They are vital for processes such as cell signaling and immune responses.

However, an excess of free radicals can cause oxidative stress, contributing to the aging process and the development of diseases like cancer and cardiovascular disorders.

Despite their inherent instability, free radicals can achieve stability under specific conditions. For instance, the delocalization of the unpaired electron over a larger molecular structure can help stabilize the radical. Furthermore, some molecules are specially designed to safely accommodate free radicals, which is particularly useful in industrial applications such as polymer production.

This ability to stabilize free radicals opens up numerous possibilities for their controlled use in scientific and commercial applications.

Skill: 2.3

Objective: Students will develop the ability to determine the electronic configurations of elements and their ions from the periodic table without the need for memorization. They will analyze trends in ionization energy across periods and down groups, and interpret variations in successive ionization energies. Skills will also include using data on ionization energies to deduce both the electronic configurations and the positions of elements within the periodic table, enhancing their analytical capabilities in applying theoretical chemistry concepts to practical scenarios.

Test your self

- Describe the trend in ionic radius as you move down a group in the periodic table and provide a reason for this trend.
- Determine the electronic configuration of an element with a proton number of 11 and its ion with a +1 charge.
- Why are the valence electrons crucial in determining the chemical properties of an atom?

- Define a free radical and provide an example of a common free radical in chemical reactions.
- Why free radicals are typically highly reactive species?
- Given the successive ionization energy data of an element, how you can determine the number of electrons in its outer shell.
- Compare the first ionization energies of two elements from the same period and write any differences in terms of the factors influencing ionization energy.
- How can ionization energy data help you deduce the position of an element in the periodic table?

Exercise

A detailed summary, chapter roadmap, multiple projects, and extensive exercises are accessible via QR code.

A Encircle the most suitable option against each statement.

1. Which quantum number primarily determines the energy of an electron in an atom?

a. Spin quantum number
b. Magnetic quantum number
c. Principal quantum number
d. Azimuthal quantum number

2. What is the maximum number of electrons that can occupy the p subshell?

a. 2 **b.** 6 **c.** 10 **d.** 14

3. The atomic number of an element also represents the number of:

a. Neutrons in the nucleus
b. Electrons in an ion
c. Protons in the nucleus
d. Electrons in the neutral atom

4. Which principle states that electrons fill the lowest energy orbitals first?

a. Pauli Exclusion Principle **b.** Hund's Rule
c. Aufbau Principle
d. Heisenberg Uncertainty Principle

5. If an atom has more protons than electrons, what is the charge of the ion?

a. Positive **b.** Negative **c.** Neutral **d.** Variable

6. What does high ionization energy suggest about an atom's outer electrons?

a. They are close to the nucleus
b. They are loosely bound
c. They are in a high energy state
d. They are shared with another atom

7. What information can be deduced from the electronic configuration of an element?

a. Atomic mass **b.** Chemical properties
c. Isotopic composition **d.** Molecular shape

8. Why does atomic radius decrease across a period from left to right?

a. Decrease in nuclear charge
b. Increase in nuclear charge and effective nuclear attraction
c. Increase in electron shielding
d. Decrease in proton number

9. What does the term 'degenerate orbitals' refer to in atomic structure?

a. Orbitals that are filled after electron excitation
b. Orbitals with electrons that have opposite spins
c. Orbitals having the same energy levels within the same subshell
d. Orbitals that belong to different shells

10. Given the atomic number 15 and charge +3, what is the electronic configuration of the resulting ion?

a. $1s^2\,2s^2\,2p^6\,3s^2\,3p^3$ **b.** $1s^2\,2s^2\,2p^6\,3s^2\,3p^2$
c. $1s^2\,2s^2\,2p^6\,3s^2\,3p^1$ **d.** $1s^2\,2s^2\,2p^6\,3s^1\,3p^1$

11. An element with successive ionization energies significantly increasing after the removal of two electrons likely belongs to which group in the periodic table?

a. Alkali metals **b.** Alkaline earth metals
c. Halogens **d.** Noble gases

12. How would the paths of beams of protons, neutrons, and electrons differ when placed in an electric field if all have the same velocity?

a. All would travel in straight lines
b. Protons and electrons bend towards each other, while neutrons travel straight
c. Electrons bend towards the field, protons away, and neutrons travel straight

d. All would bend towards the field

13. Calculate the number of neutrons in an ion with an atomic number of 17, mass number of 35, and a -1 charge.

a. 17 **b.** 18

c. 19 **d.** 20

14. Using mass spectrometry data, if the two isotopes of an element are found to be 79% at mass 35 and 21% at mass 37, what is the relative atomic mass of the element?

a. 35.42 **b.** 35.79

c. 36 **d.** None of these

B Restricted Response Question.

1. Write the concept of emission spectra and how it relates to electronic transitions in atoms.
2. Describe a potential new material that could be developed based on its electronic configuration and its application in technology.
3. Describe how the energy levels of electrons within an atom influence its chemical properties.
4. Why the 4s subshell is filled before the 3d subshell.
5. How Hund's rule affects the filling of orbitals within a subshell.
6. If a d subshell contains 10 electrons, how many orbitals are there and how are the electrons distributed among them?
7. How do the principal quantum number (n), azimuthal quantum number (l), magnetic quantum number (ml), and spin quantum number (ms) determine the electronic distribution in an atom?
8. Illustrate how the electronic configuration of silicon makes it suitable for use in semiconductors.
9. Given an element with an atomic number of 15 and a mass number of 31, determine the number of electrons in its +3 cation.
10. How would the calculated relative atomic mass change if a third isotope with a mass of 13 u and an abundance of 10% is discovered?
11. How a large jump between two successive ionization energies can provide information about an element's electronic configuration
12. Describe how nuclear charge, atomic radius, shielding effect, and spin-pair repulsion influence the ionization energy of an element.
13. Define the concept of ionization energy and why it generally increases across a period.

C Extended Question.

1. Explain how the proton number defines an element's identity and affects its chemical properties, using periodic table trends as examples.
2. Outline how to determine an atom's electronic configuration using the Aufbau rule, Pauli exclusion principle, and Hund's rule, with a transition example from the ground to an excited state.
3. Discuss the reasons for changes in atomic radius across a period and down a group, including the roles of effective nuclear charge and electron repulsion.
4. Describe how mass spectrometry is used to determine isotopic composition and calculate atomic masses, explaining the data interpretation process.
5. Explain the significance of ionization energy in predicting periodic table placements and its relationship with electron configuration and valence electrons.
6. Explain how a mass spectrometer determines the relative atomic mass of an element from its isotopic composition, covering ionization, acceleration, deflection, and detection processes, as well as the use of relative isotopic abundances for calculation.
7. Explain how emission spectra reveal an elemenAt's electronic configuration by correlating observed emission lines with transitions between energy levels, thereby deducing the arrangement of electrons within the element's atoms.