CBSE Class 9 Science Notes Chapter 4: In Chapter 4 of CBSE Class 9 Science, we're going to talk about atoms. Atoms are like tiny building blocks that make up everything around us. We'll learn about how scientists discovered what atoms are made of and how they're put together. From the earliest ideas about atoms to the modern understanding, we'll cover it all.

By the end of this chapter, you'll have a better understanding of how everything in the world is made up of these tiny, invisible particles called atoms.

CBSE Class 9 Science Notes Chapter 4 Structure of the Atom Overview

These notes on Chapter 4, "Structure of the Atom," for CBSE Class 9 Science have been prepared by subject matter experts. They aim to provide students with a clear and detailed understanding of the fundamental concepts related to the structure of atoms.

With detailed explanations and illustrative examples, these notes are an invaluable resource to aid students in mastering the complexities of atomic theory.

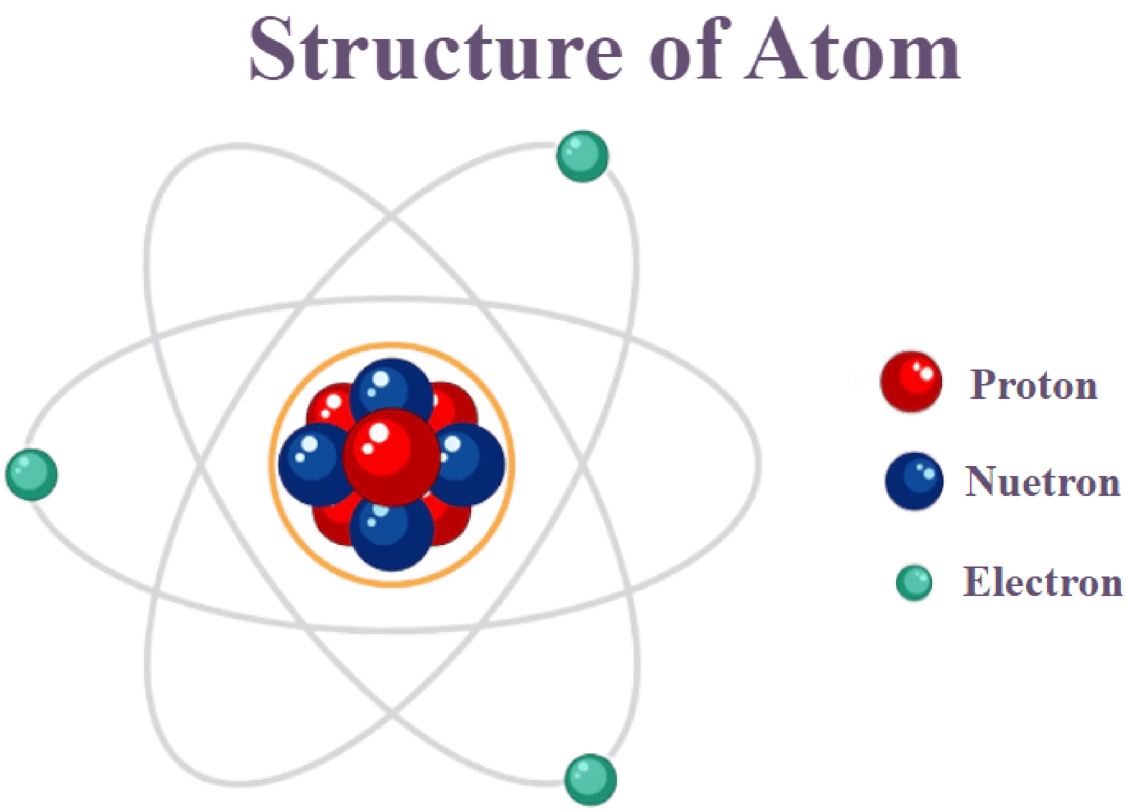
CBSE Class 9 Science Notes Chapter 4 PDF

The PDF link provided below for CBSE Class 9 Science Notes Chapter 4 provide students a convenient resource for studying this important topic. This comprehensive notes covers important concepts such as the states of matter, physical properties, and changes of state in a clear and concise manner.

CBSE Class 9 Science Notes Chapter 4 PDF

CBSE Class 9 Science Notes Chapter 4 Structure of the Atom

Atom

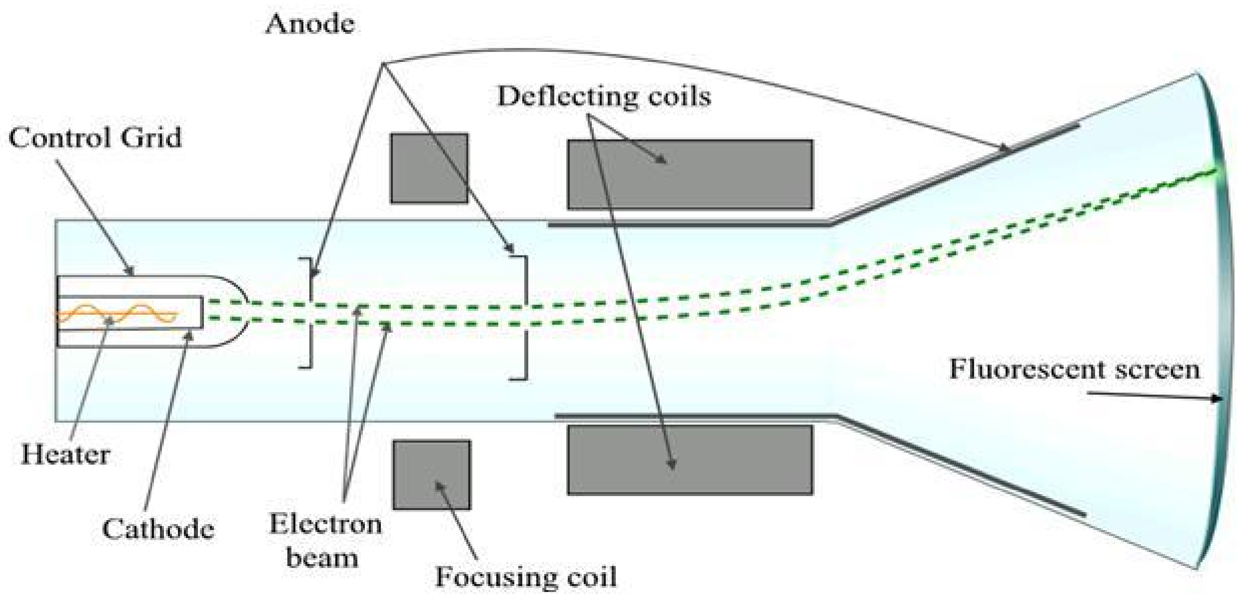


Atoms are the basic units of matter, forming the foundation of all substances in the universe. Comprising three subatomic particles—protons, neutrons, and electrons—atoms exhibit unique properties based on their composition. Protons, with a positive charge, and neutrons, which are neutral, reside in the atom's nucleus, while electrons, negatively charged, orbit around the nucleus in specific energy levels or shells. This intricate structure enables atoms to interact with one another, forming the diverse array of elements and compounds observed in nature.

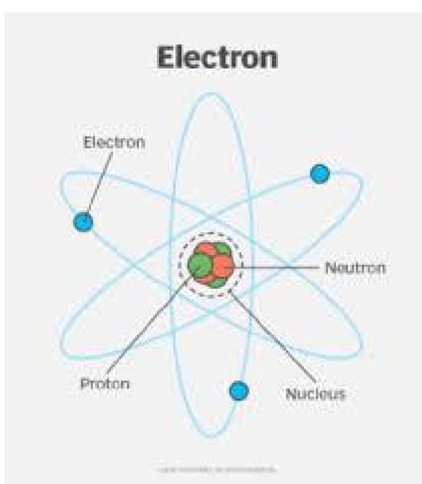
Cathode Ray Experiment

The Cathode Ray Experiment, conducted by J. J. Thomson, was a pivotal scientific investigation that led to the discovery of electrons. In this experiment, Thomson used a vacuum-sealed glass tube with electrodes at each end: a cathode (negative electrode) and an anode (positive electrode). When a high voltage was applied across the electrodes, a beam of particles traveled from the cathode to the anode.

Thomson observed that regardless of the type of gas in the tube or the material of the electrodes, the particles always moved in the same manner, indicating they were fundamental constituents of atoms. These particles were later identified as electrons, revealing that atoms were not indivisible, but composed of smaller subatomic particles. This experiment revolutionized our understanding of atomic structure and laid the foundation for modern physics.

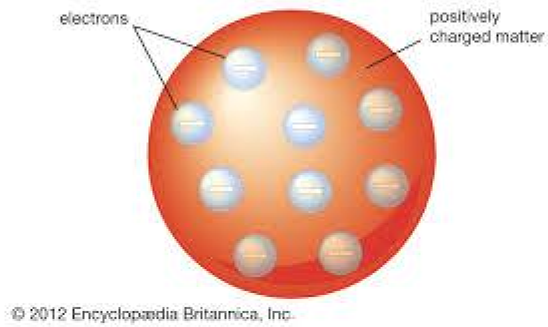


Electrons



Electrons are fundamental subatomic particles that carry a negative electric charge within an atom. They are characterized by their negligible mass, denoted as -1 charge, and represented by the symbol "e^-". Electrons are exceptionally tiny in size and are located outside the nucleus of an atom, where they move within specific energy levels or orbitals. These negatively charged particles play a crucial role in determining the chemical behavior and properties of atoms, as they participate in various chemical reactions and interactions with other atoms.

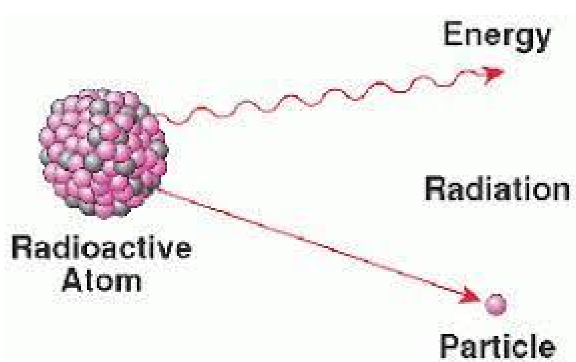
Thomson’s Model of an Atom



Thomson's model of an atom, proposed by physicist J.J. Thomson in 1904, is often referred to as the "plum pudding" model. According to this model, the atom is envisioned as a uniform, positively charged sphere with negatively charged electrons embedded within it, akin to plums in a pudding. In this representation, the positive charge is spread out evenly throughout the atom, while the electrons are scattered throughout the positively charged sphere.

This model implies that the atom has no overall charge and seeks to explain the stability of the atom despite the presence of negatively charged electrons. However, later discoveries and experiments led to the refinement of this model, ultimately paving the way for the development of more accurate atomic models, such as Rutherford's nuclear model and Bohr's planetary model.

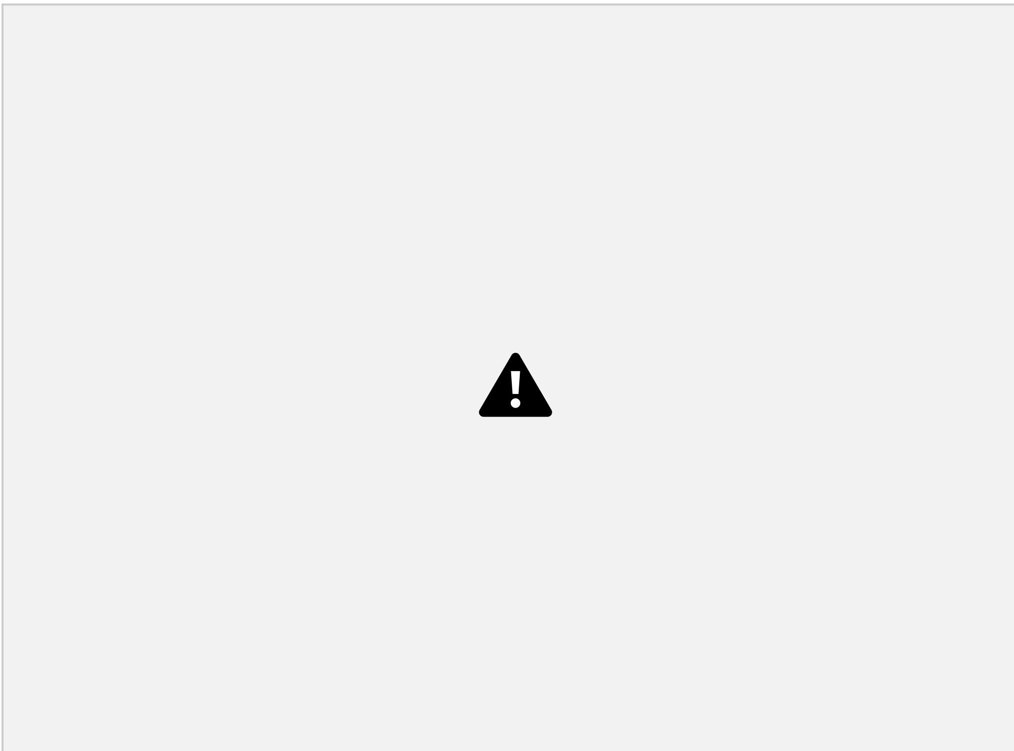
Radioactivity



Radioactivity is a natural process where the unstable nucleus of an atom emits energy in the form of particles or electromagnetic waves. This phenomenon occurs spontaneously in certain types of atoms, particularly those with an imbalance of protons and neutrons in their nuclei.

During radioactivity, particles such as alpha particles (consisting of two protons and two neutrons) or beta particles (electrons or positrons) are emitted from the nucleus. These emissions help the unstable atom achieve a more stable configuration. Since radioactivity is an intrinsic property of certain atomic nuclei, it occurs independently of external influences.

Rutherford Model



The Rutherford model, proposed by physicist Ernest Rutherford in 1911, revolutionized the understanding of atomic structure. In this model, Rutherford suggested that atoms have a dense central nucleus surrounded by orbiting electrons. The nucleus, which contains positively charged protons and neutral neutrons, occupies a very small volume compared to the overall size of the atom.

The electrons, which are negatively charged, orbit the nucleus in fixed paths or orbits, much like planets orbiting the sun. Rutherford's model also introduced the concept of the atomic number, representing the number of protons in the nucleus, which determines the identity of the element. Although the Rutherford model was groundbreaking, it was later refined with the development of quantum mechanics to explain the behavior of electrons in greater detail.

Rutherford’s Model of an Atom

Rutherford's conclusions from the α-particle scattering experiment shaped his model of the atom as follows:

(i) At the core of the atom lies a positively charged center known as the nucleus. The vast majority of an atom's mass is concentrated within this nucleus.

(ii) Electrons orbit the nucleus in distinct, well-defined paths.

(iii) The nucleus is remarkably tiny in comparison to the overall size of the atom.

Drawbacks of Rutherford’s Model

Despite its significant contributions to our understanding of atomic structure, Rutherford's model had several drawbacks:

1. Stability of Electrons: According to classical electromagnetic theory, electrons in motion should continuously emit radiation. This would cause them to lose energy and spiral into the nucleus, making atoms unstable.
2. Lack of Explanation for Spectral Lines: Rutherford's model couldn't account for the discrete lines observed in atomic spectra, such as those of hydrogen. These lines correspond to electrons transitioning between energy levels, which Rutherford's model couldn't explain.
3. Failure to Explain Chemical Properties: The model didn't provide an explanation for the chemical properties of elements. It couldn't clarify why atoms of different elements exhibit distinct chemical behaviors.
4. Absence of Electron Arrangement: Rutherford's model didn't specify the arrangement of electrons within the atom. It couldn't explain why certain atoms are more stable than others or why elements exhibit different valencies.

Properties of Electrons, Protons, and Neutrons

Electrons:

1. Charge: Negatively charged (-1).
2. Mass: Negligible compared to protons and neutrons.
3. Location: Found outside the nucleus in electron shells or orbitals.
4. Role: Involved in chemical bonding and determining the chemical properties of elements.

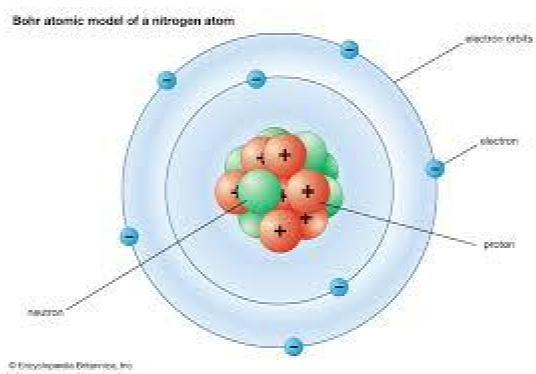
Protons:

1. Charge: Positively charged (+1).
2. Mass: Approximately equal to that of neutrons.
3. Location: Located in the nucleus of an atom.
4. Role: Determines the identity of the element and contributes to the atomic mass.

Neutrons:

1. Charge: Neutral (no charge).
2. Mass: Approximately equal to that of protons.
3. Location: Also located in the nucleus along with protons.
4. Role: Helps stabilize the nucleus and contributes to the atomic mass.

Neil Bohr Model



Bohr’s Model of an Atom: Niels Bohr proposed a model of the atom that addressed some of the limitations of Rutherford's model. According to Bohr's model:

1. Electrons revolve around the nucleus in stable orbits or energy levels.
2. Each orbit has a definite energy associated with it, and electrons do not emit radiant energy while in these stable orbits.
3. Electrons can jump from one orbit to another by either emitting or absorbing energy.
4. These orbits or energy levels are labeled as K, L, M, and N shells, with the K shell being closest to the nucleus and having the lowest energy.

Bohr's model provided a framework for understanding atomic spectra and the quantized nature of electron energies, laying the foundation for modern quantum mechanics.

Orbits

Orbits, also known as energy shells or energy levels, are regions around the nucleus of an atom where electrons are found. These orbits are arranged at different distances from the nucleus and are designated by letters such as K, L, M, and so on. Each orbit can hold a specific maximum number of electrons, determined by the formula 2n^2, where 'n' represents the orbit number.

Electrons fill these orbits in a step-wise manner, starting from the lowest energy level (K shell) and proceeding to higher energy levels. Orbits play a important role in understanding the electronic structure of atoms and their chemical behavior.

Electron Distribution in Different Orbits

According to Bohr and Bury, the maximum number of electrons in an orbit is given by the formula 2n^2, where 'n' represents the orbit number. The shells are filled stepwise, from lower to higher energy levels, and electrons are not filled in the next shell until previous shells are filled.

Valency

The valency of an atom is determined by the number of electrons in its outermost shell. Atoms with a completely filled outermost shell exhibit little chemical activity and have a valency of zero. For example, hydrogen has a valency of 1, while magnesium has a valency of 2.

Atomic Number: The atomic number of an atom is the number of protons present in its nucleus, denoted by 'Z'. It determines the identity of the element and is unique to each element.

Mass Number and Representation of an Atom

The mass number of an atom is the total number of protons and neutrons present in its nucleus. Isotopes are atoms of the same element with different mass numbers. Isobars are atoms of different elements with the same mass number.

Isotopes and Isobars

Isotopes and isobars are terms used in nuclear chemistry to describe different aspects of atomic structure.

Isotopes:

- Isotopes are atoms of the same element that have the same number of protons (atomic number) but different numbers of neutrons (and therefore different mass numbers).
- Since isotopes of the same element have the same number of protons, they exhibit similar chemical properties but may have different physical properties due to variations in mass.
- For example, carbon-12, carbon-13, and carbon-14 are isotopes of carbon, with the same number of protons (6) but different numbers of neutrons and different mass numbers (12, 13, and 14 respectively).

Isobars:

- Isobars are atoms of different elements that have the same mass number (total number of protons and neutrons) but different atomic numbers (number of protons).
- Isobars have different chemical properties since they belong to different elements with different numbers of protons and electrons.

- For example, calcium-40 and argon-40 are isobars, both having a mass number of 40 but different atomic numbers (20 for calcium and 18 for argon).

Calculation of Mass Number for Isotopic Elements

The mass number of an isotopic element can be calculated by considering the relative abundance of each isotope and its respective mass number. Here's how you can calculate it:

Determine the isotopes present and their respective percentages or relative abundances. For example, consider an element like carbon, which has two common isotopes: carbon-12 (with a natural abundance of about 98.9%) and carbon-13 (with a natural abundance of about 1.1%).

Convert the percentages to decimals by dividing by 100. For example, 98.9% becomes 0.989 and 1.1% becomes 0.011.

Multiply each isotope's relative abundance by its mass number. For carbon-12, the calculation would be 12 * 0.989 = 11.868, and for carbon-13, it would be 13 * 0.011 = 0.143.

Add the results of these calculations together to find the average mass number. For carbon, the calculation would be 11.868 + 0.143 = 12.011.

So, the calculated mass number for carbon, considering its isotopes carbon-12 and carbon-13, would be approximately 12.011 atomic mass units (u).

Benefits of CBSE Class 9 Science Notes Chapter 4 Structure of the Atom

- **Clarity and Understanding**: The notes are prepared by subject matter experts, ensuring clarity and accuracy in the explanation of concepts. This helps students gain a better understanding of the topic.
- **Easy Accessibility**: Students can access these notes anytime, anywhere, which allows for flexible learning. Whether at home or on the go, students can refer to these notes to clarify doubts or reinforce their understanding.
- **Exam Preparation**: With comprehensive coverage of the chapter's content, these notes serve as valuable study material for exam preparation. They highlight important points and concepts that are likely to be tested in exams.
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