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CHAPTER NOTES

Journey Inside the Atom Class 9 Notes | CBSE Science Chapter 8

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Ask me anything from this chapter. A concept, a numerical, or that one doubt you did not want to raise in class. Pick one to start:

Explain Journey Inside the Atom simply

What can the board exam ask?

Most important formulas

Quiz me on this chapter

Ask Ira anything...



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Journey Inside the Atom Class 9 Notes | CBSE Science Chapter 8

Journey Inside the Atom is Chapter 8 of CBSE Class 9 Science (Exploration, 2026-27). It follows how the idea of the atom grew, from Acharya Kanada's parmanus and the Greek atomos to Dalton's atomic theory, and then how Thomson, Rutherford, Bohr and Chadwick opened the atom up to find electrons, a nucleus, protons and neutrons.

The chapter then turns this picture into numbers you can work with: element symbols, atomic number (Z), mass number (A), the Bohr-Bury rules for filling shells, electronic configuration of the first eighteen elements, valency, isotopes, average atomic mass and isobars. School exams test both halves: reasoning about the gold foil experiment and the limits of each model, and short calculations of protons, neutrons, electrons, configurations and average atomic mass. The same ideas carry straight into Class 10 chemistry.

Quick Answer - Journey Inside the Atom

Atoms are divisible. Thomson found the atom as a sphere with electrons spread through it. Rutherford found a tiny, dense, positive nucleus with most of the mass. Electrons are in fixed shells K, L, M, N, where the maximum number of electrons is $2n^2$. The neutron was discovered in 1932. Atomic number Z = number of protons. The number of electrons lost, gained or shared to form an ion is called the valency.



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All Key Formulas at a Glance

Quantity or rule	Formula or statement	Example
Atomic number	$Z = \text{number of protons}$	Sodium: $Z = 11$
Electrons in a neutral atom	Number of electrons = Z	Sodium: 11 electrons
Mass number	$A = \text{number of protons} + \text{number of neutrons}$	Helium: $2 + 2 = 4$
Number of neutrons	Neutrons = $A - Z$	$^{12}_6\text{C}$: $12 - 6 = 6$
Standard notation	Mass number on top left, atomic number bottom left of the symbol	$^{12}_6\text{C}$
Maximum electrons in shell n	$2n^2$	
Outermost shell limit	At most 8 electrons	
Valency	Electrons lost, gained or shared to complete the outer shell	
Average atomic mass	Sum of (isotope mass $\times$ relative abundance)/100	
Size of atom and nucleus	Atom $\approx 10^{-10} \text{ m}$, Nucleus $\approx 10^{-15} \text{ m}$	
Charge of electron	$-1.602 \times 10^{-19} \text{ C}$	



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Key Concepts

1. Early Ideas of the Atom and the Discovery of the Electron

Everything around us, living or non-living, is matter, and matter is made of tiny particles called atoms. Atoms are far too small to see with the naked eye.

Ancient ideas (more than 2,000 years ago):

- **Acharya Kanada (India)** suggested that if matter (**dravya**) is divided again and again, you finally reach the smallest particles that cannot be divided further. He called them **parmanus**. His ideas are recorded in the Sanskrit text **Vaisesika Sutras**. A parmanu is infinitely small and cannot be perceived by the senses. Parmanus combine to form **dyads** (groups of two) and **triads** (groups of three), and the whole material universe, including living bodies, is made of such combinations.
- **Leucippus and Democritus (Greece)** proposed a similar idea and called the indivisible particles **atomos**, which means indivisible.

Both ideas came from thinking and imagination.

Dalton's atomic theory (1808): John Dalton proposed that all matter is made of tiny particles called atoms, the fundamental building blocks that cannot be created or destroyed. His theory was based on scientific description of how matter is made and our understanding of atomic structure.

Atoms are divisible: until the late 19th century, atoms were considered as the smallest units of matter. Then scientists found that atoms are made of even smaller particles called **radiation**. This phenomenon is **radioactivity**, which shows that atoms must contain smaller particles.

J. J. Thomson and cathode rays (1897)

Thomson studied how electric current passes through a glass tube with two electrodes sealed inside. He used a vacuum pump to remove most of the gas. He saw a glow (called **cathode rays**) coming from the **cathode** (negative electrode) towards the **anode** (positive electrode). These are called **cathode rays**.

- By studying cathode rays in electric and magnetic fields, Thomson concluded that they consist of **negatively charged particles** with a much smaller mass than that of a hydrogen atom.



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- These particles, later called **electrons**, came out of atoms, so atoms have smaller subatomic parts.
- The nature of cathode rays did not depend on the material of the cathode or on the gas in the tube. So electrons are a fundamental part of **every atom** of every element.
- The charge of an electron is -1.602×10^{-19} C. By convention and for convenience it is taken as **-1**.

2. Thomson's Model of the Atom

After finding electrons, Thomson faced a puzzle: atoms are electrically neutral, so some positive charge must balance the electrons. He proposed that the atom is a **sphere of positive charge with electrons distributed throughout it**.

- It was compared to a pudding with plums in it, so it is called the **plum pudding model**.
- A more familiar picture is a **watermelon**: the red pulp stands for the positively charged matter and the seeds stand for electrons spread throughout.
- The total positive charge equals the total negative charge.

This model was later replaced, but it was the first to show that the positive and negative charges of an atom stay balanced.

3. Rutherford's Gold Foil Experiment

In **1911**, **Geiger and Marsden**, working under Rutherford, aimed a narrow beam of **alpha (α) particles** at a thin sheet of gold foil. Alpha particles are tiny, positively charged particles emitted by the nucleus of a helium atom: 2 protons and 2 neutrons.

Expected result (from Thomson's model)

The beam of alpha particles should have passed straight through the foil. The positive spot can push strongly on an α -particle. The beam was deflected only slightly.

Actual observations: most α -particles passed straight through at small angles, some were deflected at large angles, and a few bounced back. Due to this, this is also called the **α -ray scattering experiment**. It showed large-angle deflections, and it could not explain Thomson's model.



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Observation	Conclusion drawn by Rutherford
Most α -particles passed straight through	Most of the atom is empty space
Some deflected through large angles	Positive charge is concentrated in a very small region, which repels the positive α -particles strongly when they come close
A few bounced back	This small region is dense and carries most of the mass; only a head-on approach to it turns a particle back

Rutherford's model (planetary model)

- The positive charge is concentrated in an extremely small central region called the **nucleus**.
- Most of the atom is **empty space**.
- The nucleus is **dense**, contains **all the mass**.
- Electrons **revolve around the nucleus**, the **planetary model**.

Size: the nucleus is about 10^5 (one lakh) times the size of an electron. The radius of the nucleus is $\approx 10^{-10}$ m and the diameter of the nucleus is $\approx 2 \times 10^{-10}$ m. If the nucleus is compared to a small ground (about 100 m across), the nucleus is like a small dot at the centre.

Limitation of Rutherford's model

A particle moving in a circle keeps changing its direction. A negatively charged electron circling the nucleus would move along a spiral path inward and the atom would collapse. Real atoms are stable, which is why Rutherford's model could not explain the **stability of the atom**.

Discovery of the proton

Rutherford showed that the positive charge is concentrated in a very small region, which he discovered and named. Protons have a positive charge equal and opposite to that of an electron (negative charge).

For an atom to be neutral, **number of protons** must be equal to **number of electrons**. For example, hydrogen has 1 proton and 1 electron; sodium has 11 protons and 11 electrons. Since the positive charge is equal to the negative charge, so the atom is neutral. All atoms are electrically neutral.



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4. Bohr's Model of the Atom

To explain why atoms are stable, **Niels Bohr** proposed a new model in **1913**. His postulates:

1. Electrons do not move randomly. They follow fixed circular paths called **stationary states**, **orbits** or **shells**.
2. In each shell an electron has a definite amount of energy, so shells are also called **energy levels**.
3. Shells are named **K, L, M, N, ...** or numbered **$n = 1, 2, 3, 4, \dots$** starting from the nucleus.
4. Electrons can revolve only in these allowed shells, never in between them.
5. While moving in a fixed shell, an electron **does not lose energy**.
6. The K shell ($n = 1$) is closest to the nucleus and has the **least energy**. Energy increases as we move away from the nucleus: an electron in L ($n = 2$) has more energy than one in K ($n = 1$).
7. An electron can move to another shell by **absorbing or releasing a fixed amount of energy** equal to the difference between the energy levels.
8. Each shell can hold only a certain number of electrons.

How Bohr explained stability: electrons do not spiral inward as in Rutherford's model. Bohr introduced **stationary states** where the energy of the electron stays constant even as it moves in a circular path, preventing an inward spiral and no collapse. Bohr's model explains the stability of atoms.

Beyond Bohr: Bohr's model also has limitations. Electrons do not follow well-defined paths around the nucleus, and we can only predict the behavior of electrons by studying this in higher classes.



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Model	Main idea	What it could not explain
Dalton (1808)	Atom as an indivisible particle	Radioactivity and electrons show atoms are divisible
Thomson (after 1897)	Positive sphere with electrons embedded	Large deflections and bounce-backs in the gold foil experiment
Rutherford (1911)	Tiny dense positive nucleus, electrons orbit, mostly empty space	Stability of the atom
Bohr (1913)	Electrons in fixed shells of definite energy, no energy loss	Has its own limitations; replaced by the quantum mechanical model
Quantum mechanical (modern)	Electron clouds, probability	

5. The Neutron, Element Symbols,

Discovery of the neutron (1932)

A puzzle: hydrogen has one proton and he heavy as a hydrogen atom, not twice. Some charge. In **1932, James Chadwick**, a student with mass nearly equal to that of a proton found in the nucleus of all atoms except h. The mass of an atom comes mainly from i their mass can be ignored.



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Subatomic particle	Symbol	Relative charge	Where found	Discovered by
Electron	e^{-}	-1	Shells around the nucleus	J. J. Thomson (1897)
Proton	p^{+}	+1	Nucleus	Ernest Rutherford
Neutron	n^0	0	Nucleus	James Chadwick (1932)

Neutrons and nuclear stability: protons all carry positive charge, so each repels every other. Neutrons, being neutral, reduce this repulsion by sitting between protons and increasing the distance between them. They also strengthen the **nuclear force** that binds all the particles of the nucleus. Heavier atoms need many more neutrons: iron has 26 protons and 30 neutrons, uranium has 92 protons and 146 neutrons.

Symbols of elements

By 1869 scientists knew about 69 elements naturally, and many more were created artificially.

- **Dalton (1803)** introduced the first pictorial symbols.
- **Berzelius (1813)** suggested that symbols be based on the first few letters of the element's name in alphabetic symbols.
- Today **IUPAC** (International Union of Pure and Applied Chemistry) uses the first few letters of the element's name in alphabetic symbols.

Rules for writing symbols:

- Many symbols are the first letter or first two letters of the element's name.
- The first letter is always a capital, the second letter is always a small letter. For example, CO, which is carbon monoxide, a compound of carbon and oxygen.
- Some use the first letter and a later letter. For example, Fe for iron.
- Some come from Latin, Greek or German. For example, Hg (hydrargyros), tungsten W (German wolle).



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Element	Symbol	Name the symbol comes from
Copper	Cu	Cuprum
Gold	Au	Aurum
Iron	Fe	Ferrum
Lead	Pb	Plumbum
Potassium	K	Kalium
Silver	Ag	Argentum
Sodium	Na	Natrium

Atomic number (Z)

The **number of protons in the nucleus** of an element is its atomic number (Z). The identity of an element and its chemical behaviour depend on the number of protons. Two atoms with different Z are different elements. The number of protons = Z. Hydrogen: Z = 1 (1 proton).

Mass number (A)

The **total number of protons and neutrons** in the nucleus of an atom is its mass number (A). Protons and neutrons together are called nucleons.

Mass number (A) = number of protons + number of neutrons

Element	Protons
Hydrogen	1
Helium	2
Lithium	3

Standard notation: write the symbol, with the atomic number as a subscript on the bottom left. Example: 6 protons, 6 electrons, 12 - 6 = 6 neutrons.



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6. Distribution of Electrons: Bohr-Bury Rules and Electronic Configuration

Bohr and Bury gave rules for how electrons fill shells:

1. The **maximum number of electrons in a shell** = $2n^2$, where n is the shell number. K ($n = 1$): $2 \times 1^2 = 2$. L ($n = 2$): $2 \times 2^2 = 8$. M ($n = 3$): $2 \times 3^2 = 18$.
2. The **outermost shell** can hold at most **8 electrons** (the first shell holds at most 2).
3. Shells fill **stepwise**, from the one closest to the nucleus outward, in the order K, L, M, N. The L shell fills only after K is complete, and so on.

The distribution of electrons among the shells is the **electronic configuration**. Build it by adding one electron to the correct shell each time the atomic number rises by 1.



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Element	Symbol	Z	K	L	M	Configuration
Hydrogen	H	1	1	-	-	1
Helium	He	2	2	-	-	2
Lithium	Li	3	2	1	-	2, 1
Beryllium	Be	4	2	2	-	2, 2
Boron	B	5	2	3	-	2, 3
Carbon	C	6	2	4	-	2, 4
Nitrogen	N	7	2	5	-	2, 5
Oxygen	O	8				
Fluorine	F	9				
Neon	Ne	10				
Sodium	Na	11				
Magnesium	Mg	12				
Aluminium	Al	13				
Silicon	Si	14				
Phosphorus	P	15				
Sulfur	S	16				
Chlorine	Cl	17				
Argon	Ar	18				



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Method in three steps: (1) find the number of electrons, then L up to 8, then M; (3) check the outer shell. Example: sulfur, Z = 16: K = 2, L = 8, leaving 6 electrons in M.

2,
Z.

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The number of neutrons does not affect the configuration. Only the number of electrons matters.

7. Valency: The Combining Capacity of an Atom

The number of hydrogen or chlorine atoms with which one atom of an element combines to form a compound is its **combining capacity**. Hydrogen and chlorine are used as the reference because each has a combining capacity of one. In H_2O , oxygen combines with 2 hydrogen atoms, so its combining capacity is 2. In NH_3 , nitrogen combines with 3 hydrogen atoms (combining capacity 3). In MgCl_2 , magnesium combines with 2 chlorine atoms (combining capacity 2).

- The outermost shell containing electrons is the **valence shell**; its electrons are **valence electrons**.
- Eight electrons in the outermost shell is an **octet**.
- Atoms with a complete octet (or 2 electrons in the case of helium) are **largely unreactive and more stable**. Atoms with incomplete valence shells are usually more reactive and less stable or **share** electrons to complete the octet.

Valency is the number of electrons gained or lost by an atom to achieve a complete octet.

- **Fewer than 4 valence electrons:** the atom loses electrons. Sodium 2, 8, 1 loses 1: valency 1.
- **More than 4 valence electrons:** the atom gains electrons. Oxygen 2, 6 gains 2: valency 2.
- **Exactly 4:** the atom cannot easily gain or lose electrons. Carbon 2, 4 has a valency 4.
- **Complete octet (or duplet for helium):** the atom is stable. Neon 2, 8 and argon 2, 8, 8 are examples.



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Element	Configuration	Valence electrons	What happens	Valency
Magnesium	2, 8, 2	2	Loses 2	2
Carbon	2, 4	4	Shares 4	4
Nitrogen	2, 5	5	Gains 3	3
Chlorine	2, 8, 7	7	Gains 1	1
Neon	2, 8	8	Stable octet	0

Some compounds appear to break this simple valency rule. You will study them in higher classes.

8. Isotopes, Average Atomic Mass

Isotopes

Dalton said all atoms of an element are identical. But atoms of the same element with the same number of protons, but different number of neutrons, and so different mass numbers. They are called isotopes. They have the same atomic number but different mass numbers.

Isotope	Notation	Protons
Protium	${}^1_1\text{H}$	1
Deuterium	${}^2_1\text{H}$	1
Tritium	${}^3_1\text{H}$	1

Carbon has three isotopes, ${}^{12}_6\text{C}$, ${}^{13}_6\text{C}$ and ${}^{14}_6\text{C}$. ${}^{12}_6\text{C}$ has 6 protons and 6 neutrons. ${}^{13}_6\text{C}$ has 6 protons and 7 neutrons. ${}^{14}_6\text{C}$ has 6 protons and 8 neutrons. ${}^{12}_6\text{C}$ is the most abundant in nature.

Properties: isotopes have similar chemical properties because they have the same number of electrons and the same electronic configuration. They differ in physical properties because they have different mass numbers.

Uses of isotopes:



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- $^{235}_{92}\text{U}$ (uranium): fuel in nuclear reactors to generate electricity in nuclear power plants.
- $^{60}_{27}\text{Co}$ (cobalt, radioactive): radiation treatment for cancer.
- $^{131}_{53}\text{I}$ (iodine): treatment of goitre and thyroid cancer.
- $^{14}_6\text{C}$ (carbon): finding the age of ancient fossils and artefacts in archaeology and geology.

Atomic mass unit and average atomic mass

Atoms are too tiny to weigh in grams or kilograms, so their masses are given in the **unified atomic mass unit (u)**. Earlier this was written as amu.

Chlorine occurs as two isotopes, of mass 35 u and 37 u, in the ratio 3 : 1 (about 75% and 25%).

- **Simple average** (ignores abundance): $(35 + 37)/2 = 36 \text{ u}$. This does not match nature, because the isotopes are not present in equal amounts.
- **Weighted average** (uses abundance): multiply each isotope's mass by its percentage abundance/100 and add.

$$= 35 \times 75/100 + 37 \times 25/100$$

$$= 105/4 + 37/4$$

$$= 142/4 = 35.5 \text{ u}$$

No single chlorine atom has a mass of 35.5 u. There are 3.5 lakh $^{35}_{17}\text{Cl}$ and 2.5 lakh $^{37}_{17}\text{Cl}$ atoms, and together they have an average mass of 35.5 u.

Isobars

Calcium ($Z = 20$), potassium ($Z = 19$) and argon ($Z = 18$) atoms of each with mass number 40 exist. These are atoms of **different elements** with the **same mass number**. Because their atomic numbers differ, isobars have different chemical properties.



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Feature	Isotopes	Isobars
Atomic number (Z)	Same	Different
Mass number (A)	Different	Same
Element	Same element	Different elements
Neutrons	Different	Different
Chemical properties	Similar	Different
Example	$^{35}_{17}\text{Cl}$ and $^{37}_{17}\text{Cl}$	$^{40}_{20}\text{Ca}$, $^{40}_{19}\text{K}$, $^{40}_{18}\text{Ar}$

Formula and Reaction Sh

This chapter has no chemical reactions. T numerical.

Quantity	
Atomic number	$Z = n$
Mass number	$A = p$
Neutrons	Numb
Maximum electrons in shell n	$2n^2$
Weighted average atomic mass	$\Sigma (ma$ abunc



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Important Definitions

- **Electron:** a negatively charged subatomic particle (relative charge -1) of very small mass, present in all atoms.
- **Proton:** a positively charged subatomic particle (relative charge +1) in the nucleus, much heavier than an electron.
- **Neutron:** a subatomic particle in the nucleus with no charge and mass nearly equal to that of a proton.
- **Nucleus:** the tiny, dense, positively charged centre of an atom that holds most of its mass.
- **Stationary state (orbit, shell, energy level):** a fixed circular path around the nucleus in which an electron has a definite energy and does not lose energy.
- **Atomic number (Z):** the number of protons in the nucleus of an atom.
- **Mass number (A):** the total number of protons and neutrons in the nucleus of an atom.
- **Electronic configuration:** the distribution of electrons in the shells of an atom.
- **Valence electrons:** the electrons in the outermost shell of an atom.
- **Valency:** the number of electrons gained or lost by an atom to achieve a stable electronic configuration.
- **Isotopes:** atoms of the same element with the same atomic number but different mass numbers.
- **Isobars:** atoms of different elements with the same mass number but different atomic numbers.



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Solved Examples (NCERT)

Example 1: Particles from Z and n

Question: An atom with an atomic number Z has n electrons, protons and neutrons.

Solution:

- Protons = Z = **26**.
- Electrons in a neutral atom = protons = Z .
- Nucleons = protons + neutrons = mass number.

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(This matches iron, which has 26 protons and 30 neutrons.)

Example 2: Neutrons from Z and A

Question: The nucleus of an atom contains 20 protons. If its mass number is 41, find the number of neutrons in it.

Solution: Neutrons = A - protons = 41 - 20 = **21**.

Question: An atom has 18 neutrons and an atomic number of 17. What is its mass number?

Solution: A = protons + neutrons = 17 + 18 = **35**.

Example 3: Magnesium atom

Question: For an atom of magnesium with a mass number of 24 and atomic number 12, determine the number of (i) protons, (ii) neutrons, (iii) electrons, and also illustrate the arrangement of electrons in a magnesium atom.

Solution:

1. Protons = Z = **12**.
2. Neutrons = A - Z = 24 - 12 = **12**.
3. Electrons = 12 (neutral atom).

Arrangement: K takes 2, L takes 8, leaving 2 valence electrons and tends to lose

Example 4: Neutrons and electron

Question: An atom has 79 protons and a mass number of 197. Find (i) the number of neutrons, and (ii) the number of electrons.

Solution: (i) Neutrons = 197 - 79 = **118**. (ii)

Question: An atom ^{70}X has 31 electrons. Find the number of protons and neutrons.

Solution: A neutral atom has electrons = protons = 31. Mass number = 70. Neutrons = 70 - 31 = 39.



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Example 5: Complete the table

Question: Complete Table 8.5. The given entries are: (a) atomic number 5, 6 neutrons; (b) mass number 14, 7 electrons, nitrogen; (c) mass number 24, 12 protons; (d) atomic number 15, 16 neutrons; (e) mass number 1, 0 neutrons.

Method: protons = electrons = Z ; $A = Z + \text{neutrons}$.

Z	A	Neutrons	Protons	Electrons	Element
5	$5 + 6 = 11$	6	5	5	Boron
7	14	$14 - 7 = 7$	7	7	Nitrogen
12	24	$24 - 12 = 12$	12	12	Magnesium
15	$15 + 16 = 31$	16			
1	1	0			



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Example 6: Electronic configuration

Question: Write the electronic configuration of an element with atomic number 18.

Solution:

- $Z = 12$ (magnesium): 2, 8, 2
- $Z = 16$ (sulfur): 2, 8, 6
- $Z = 18$ (argon): 2, 8, 8

Question: Identify the number of electrons in an element with atomic number 12.

Solution: Only Z matters. Carbon ($Z = 6$): 2, 8, 4, so **4**. The mass numbers 12, 19 and 23 are irrelevant for this configuration.

Example 7: Element X (case study)

Question: An element X has a mass number 24 and 12 protons. (i) How many electrons and protons does element X have? (ii) What is its atomic number? (iii) Identify the element.

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element X. (iv) Write its electronic configuration. (v) How many valence electrons does it have? (vi) What will be the mass number if two neutrons are added to its nucleus? (vii) What will be the relation of X with the new atom?

Solution:

1. Protons = $A - \text{neutrons} = 35 - 18 = 17$; electrons = **17**, protons = **17**.
2. Atomic number $Z = 17$.
3. $Z = 17$ is **chlorine (Cl)**.
4. Configuration **2, 8, 7**.
5. Valence electrons = **7** (so it gains 1; valency 1).
6. New mass number = $35 + 2 = 37$.
7. Same Z (17), different A (35 and 37): they are **isotopes** of chlorine.

Example 8: Isotopes and isobars

Question: The composition of the nuclei of two atoms is as follows: X has 17 protons and 19 neutrons; Y has 17 protons and 18 neutrons. Write the relation between (i) Y and Z, (ii) Z and X.

Solution: First find each mass number. X:

1. **Y and Z:** same atomic number (17), different mass numbers, so they are **isotopes** of the same element (chlorine).
2. **Z and X:** same mass number (37), different atomic numbers, so they are **isobars**.

Example 9: Average atomic mass

Question: If a bromine atom is available in the form of $^{81}_{35}\text{Br}$ (50.3%), calculate the average atomic mass of bromine.

Solution:

Average atomic mass = $79 \times 49.7/100 + 81 \times 50.3/100$

$$= (3926.3 + 4074.3)/100$$

$$= 8000.6/100 = \mathbf{80.006\text{ u} \approx 80\text{ u}}$$



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Check: the two abundances are close to equal, so the answer should be close to the midpoint of 79 and 81, slightly towards 81 because ^{81}Br is a little more common. 80.006 u fits.

Competency-Based Questions (with answers)

Q1. Assertion-Reason (textbook Pause and Ponder 7)

Assertion (A): Rutherford concluded that most of the mass of an atom is concentrated in a small region at the centre called the nucleus.

Reason (R): According to Thomson's model, electrons are embedded in a uniformly distributed positive charge sphere.

(i) Both A and R are true, and R is the correct explanation of A. (ii) Both A and R are true, but R is not the correct explanation of A. (iii) A is true, but R is false. (iv) A is false, but R is true.

Answer: (ii). A is true: from the gold foil experiment, the nucleus holds all the positive charge and most of the mass. R is false: Thomson's plum pudding model. R does not explain A, and Rutherford's conclusion came from the alpha particles being deflected or bounced back, a result that disproved Thomson's model.

Q2. Assertion-Reason: Isotopes

Assertion (A): The isotopes $^{35}_{17}\text{Cl}$ and $^{37}_{17}\text{Cl}$ have the same chemical properties.

Reason (R): Isotopes of an element have the same electronic configuration.

(i) Both A and R are true, and R is the correct explanation of A. (ii) Both A and R are true, but R is not the correct explanation of A. (iii) A is true, but R is false. (iv) A is false, but R is true.

Answer: (i). A is true: both atoms are chlorine, with 17 protons and 17 electrons, arranged 2, 8, 7. They have the same chemical properties, mainly on the valence electrons, and both differ only in their number of neutrons (18 and 20). R is true: isotopes have the same electronic configuration. The boiling point is a physical property, not a chemical one.

Q3. Case-based: Isotopes in daily life

A hospital uses one isotope to treat cancer. A scientist uses another isotope to date an ancient wooden object. A power plant generates electricity from a nuclear fuel.



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(a) Name the isotope used in each case. (b) Why do all isotopes of an element show similar chemical properties? (c) State one property in which isotopes differ.

Answer: (a) Cancer radiation treatment: cobalt-60 ($^{60}_{27}\text{Co}$). Goitre and thyroid cancer: iodine-131 ($^{131}_{53}\text{I}$). Dating: carbon-14 ($^{14}_6\text{C}$). Nuclear fuel: uranium-235 ($^{235}_{92}\text{U}$). (b) They have the same number of electrons and the same electronic configuration, and chemical properties depend mainly on valence electrons. (c) Physical properties such as melting point and boiling point.

Q4. Source-based: Chlorine's average mass

Chlorine occurs in nature as two isotopes of mass 35 u and 37 u, in the ratio 3 : 1.

(a) Find the simple average. (b) Find the weighted average. (c) Why is the weighted average the correct value? (d) In 1 million chlorine atoms, how many are ^{35}Cl ?

Answer: (a) $(35 + 37)/2 = 36$ u. (b) $35 \times 3/4 + 37 \times 1/4 = 105/4 + 37/4 = 142/4 = 35.5$ u. (c) The isotopes are not present in equal amounts: abundance, so it reflects the element as it is.

Q5. Reasoning: Why Rutherford failed

Both Rutherford's and Bohr's models have model fail to explain atomic stability, while

Answer: In Rutherford's model an electron losing energy, spiral inward and fall into the nucleus. Bohr postulated that energy stays constant even though it moves. Electrons change energy only by a fixed amount of energy.

Q6. Case-based: Heavy electrons

In an atom, there are 12 protons and 12 neutrons. If all the electrons are replaced with some hypothetical particles which are 500 times heavier. What effect will it have on (i) Atomic number (ii) Atomic mass (iii) Mass number

Answer: (i) No change: Z = number of protons. (ii) Atomic mass will no longer be ignored. (iii) No change: Atomic mass number



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change: the 12 particles carry the same total negative charge as 12 electrons, balancing the 12 protons, so the atom stays neutral.

Q7. Reasoning: Choosing the right statements (textbook exercise 1)

Choose the correct options and explain the reason for the correct and incorrect options in the context of Ernest Rutherford's gold foil experiment: (i) The experiment clearly showed the existence of neutrons in the nucleus. (ii) The results disproved the plum pudding model and led to the idea of a nucleus at the centre of the atom. (iii) The large deflection of a few alpha particles indicated that most of the mass of the atom and positive charge are packed into a tiny centre. (iv) The way alpha particles were deflected showed that electrons move around the nucleus.

Answer: Correct: **(ii) and (iii)**. (i) is incorrect: neutrons were discovered by Chadwick in 1932; the experiment gave no evidence of neutrons. (ii) is correct: an evenly spread positive charge cannot turn back α -particles, so the plum pudding model failed and a central nucleus was proposed. (iii) is correct: only a small, dense, positive centre can deflect α -particles through large angles. (iv) is incorrect: the deflection of α -particles is due to the nucleus, which is far too light to deflect α -particles. Orbiting electrons are too light to deflect α -particles. The scattering pattern gave no evidence of neutrons.

Q8. Reasoning: Correct or incorrect

Which of the following statements are correct? Give a reason for each statement. (i) Electrons slowly fall into the nucleus. (ii) Electrons can be found between energy levels. (iii) Electrons revolve around the nucleus. (iv) Electrons can be found between energy levels.

Answer: (i) Incorrect: in a stationary state electrons occupy only allowed shells, each with a fixed energy. (ii) Correct: this is the central postulate. (iv) Incorrect: electrons are found in shells, not between them.

Q9. Riddle (textbook Pause and Ponder)

Solve this riddle: I am an atom with a mass number of 23. I react vigorously with water. Who am I and why?

Answer: $Z = 11$ is **sodium (Na)**. Neutrons = 12. It has one valence electron, which it loses easily; this makes it highly reactive.



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Important Questions for Exams

1-Mark Questions

Q1. Who discovered the electron, and in which year?

Answer: J. J. Thomson, in 1897.

Q2. What is the maximum number of electrons that the M shell can hold?

Answer: $2n^2 = 2 \times 3^2 = 18$.

Q3. Write the symbols of potassium, silver and tungsten.

Answer: K, Ag, W.

Q4. Define isobars and give one example.

Answer: Atoms of different elements with the same mass number but different atomic numbers, for example $^{40}_{20}\text{Ca}$ and $^{40}_{18}\text{Ar}$.

Q5. What is the valency of an element with 3 valence electrons?

Answer: 3 (it loses its 3 valence electrons)

2-Mark Questions

Q6. How did the ancient idea of the atom change over time?

Answer: Kanada's parmanus and the Greek experiments. Dalton's theory (1808) proposed experiments of his time, so it was the first

Q7. Why did Thomson conclude that electrons are present in all atoms?

Answer: The nature of cathode rays was that they were made of whatever gas filled the tube. Since the same results were obtained for all gases, electrons must be a basic part of the atom.

Q8. Why does a helium atom have about four times the mass of a hydrogen atom?

Answer: Helium's nucleus also has 2 neutrons. Its mass is due to 4 nucleons, against 1 in hydrogen.

3-Mark Questions

Q9. State the Bohr-Bury rules for filling electron shells in phosphorus ($Z = 15$).



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Answer: (1) The maximum number of electrons in a shell is $2n^2$. (2) The outermost shell holds at most 8 electrons. (3) Shells fill stepwise from the innermost, K, L, M, and so on. Phosphorus: K = 2, L = 8, M = 15 - 10 = 5. Configuration 2, 8, 5.

Q10. What is valency? How is the valency of sodium, oxygen and carbon decided by their configurations?

Answer: Valency is the number of electrons gained, lost or shared by an atom to complete its octet. Sodium (2, 8, 1) has fewer than 4 valence electrons and loses 1: valency 1. Oxygen (2, 6) has more than 4 and gains 2: valency 2. Carbon (2, 4) cannot easily gain or lose 4, so it shares 4: valency 4.

Q11. Explain and arrange the following statements in the correct chronological order to show how atomic models have evolved over time. (i) Bohr's model proposed that electrons move in fixed orbits around the nucleus, each with a definite energy. (ii) Thomson's model depicted the atom as a 'plum pudding' with electrons embedded in a sphere of positive charge. (iii) Rutherford's model proposed that atoms have a dense central nucleus. (iv) Dalton's model described atoms as indivisible particles.

Answer: Order: (iv) → (ii) → (iii) → (i). Dalton's model (1808) described atoms as indivisible particles. Thomson's model (1897) showed atoms are divisible sphere. The gold foil experiment (1911) discovered a dense nucleus and most of the mass in a tiny nucleus. Bohr's model (1913) proposed fixed shells in which electrons move.

Q12. Electrons move around the nucleus in fixed orbits. Explain what keeps them attracted to the nucleus.

Answer: Electrons are negatively charged particles. Protons are positively charged. Opposite charges attract, so the electrons are bound to the atom as it moves around.

5-Mark Questions

Q13. Describe Rutherford's gold foil experiment. State the model proposed and one limitation of the model.

Answer:

- **Setup:** in 1911 Geiger and Marsden, uncovered the nature of α -particles (helium nuclei) at a distance.
- **Observations:** most particles passed through undeflected; some deflected at large angles; a few bounced back.



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- **Conclusions:** most of the atom is empty space; the positive charge and most of the mass are concentrated in an extremely small, dense nucleus.
- **Model:** a tiny positive nucleus at the centre, about 10^5 times smaller than the atom (atom $\approx 10^{-10}$ m, nucleus $\approx 10^{-15}$ m), with electrons revolving around it like planets around the Sun (planetary model).
- **Limitation:** a revolving electron is accelerating and should lose energy, spiral into the nucleus and make the atom collapse. Atoms are stable, so the model could not explain stability.

Q14. (a) Define isotopes. (b) Give the three isotopes of hydrogen with their numbers of protons, neutrons and electrons. (c) State two uses of isotopes. (d) Calculate the average atomic mass of chlorine, whose isotopes of mass 35 u and 37 u occur in the ratio 3 : 1.

Answer: (a) Atoms of the same element with the same atomic number but different mass numbers. (b) Protium ${}^1_1\text{H}$: 1 proton, 0 neutrons, 1 electron. Deuterium ${}^2_1\text{H}$: 1, 1, 1. Tritium ${}^3_1\text{H}$: 1, 2, 1. (c) Uranium-235 as fuel in nuclear reactors; carbon-14 for finding the age of fossils and artefacts (also cobalt-60 for cancer treatment and iodine-131 for goitre). (d) $35 \times 75/100 + 37 \times 25/100 = 26.25 + 9.25 = 35.5$ u.

Common Mistakes and Ex

1. **Mistake:** writing that Rutherford's experiment was discovered by Chadwick in 1932. The experiment was discovered by Rutherford in 1911.
2. **Mistake:** using mass number to write electron configuration. It is only on the number of electrons, which is equal to the atomic number.
3. **Mistake:** writing 2, 8, 9 or putting more than 8 electrons in the third shell. The third shell holds at most 8 electrons. For the first 18 elements, the third shell holds at most 8 electrons.
4. **Mistake:** writing symbols as AL, CO, CL. AL is an element, aluminium. CO is a compound, carbon monoxide. CL is an element, chlorine.
5. **Mistake:** writing the valency of chlorine as 7. Valency is the number of valence electrons, valency = 8 - valence electrons = 8 - 7 = 1.
6. **Mistake:** mixing up isotopes and isobars. Isotopes have the same atomic number but different mass numbers. Isobars have the same mass number but different atomic numbers.
7. **Mistake:** taking the simple average (36) instead of the weighted average with abundances: 35.5 u.
8. **Mistake:** saying one chlorine atom weighs 35.5 u. The average atomic mass of chlorine is 35.5 u, but each atom is either 35 u or 37 u.



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9. **Mistake:** saying Bohr's electrons are not moving. **Fix:** they move in circles around the nucleus; what stays constant in a stationary state is their energy.
10. **Mistake:** saying the K shell has the most energy. **Fix:** K ($n = 1$) is closest to the nucleus and has the least energy; energy increases outward.
11. **Tip:** in every numerical, write the relation first ($A = p + n$, electrons = Z) and then substitute. Show units (u) for atomic mass, and state the final answer clearly.

Quick Revision Points

- Acharya Kanada: parmanus, recorded in the Vaisesika Sutras; dyads and triads.
- Dalton (1808): first scientific atomic theory; atoms indivisible.
- Thomson (1897): cathode rays in a low-pressure gas tube are negatively charged particles, electrons, present in all atoms.
- Gold foil experiment (1911, Geiger and Marsden): some deflect sharply, a few bounce back.
- Rutherford: mostly empty space; tiny dense nucleus; electrons revolve (planetary model).
- Nucleus is about 10^5 times smaller than atom.
- Rutherford's limitation: could not explain stability of atom.
- Bohr (1913): fixed shells K, L, M, N ($n = 1, 2, 3, 4$).
- K has the least energy; energy increases with amount of energy.
- Z = number of protons = number of electrons.
- A = protons + neutrons (nucleons); neutrons = $A - Z$.
- Maximum electrons per shell = $2n^2$: K 2, L 8, M 18, N 32.
- Valency: electrons lost, gained or shared. Cl 1, O 2, N 3, C 4.
- Isotopes: same Z , different A ; similar chemical properties.
- Isobars: same A , different Z ; ^{40}Ca , ^{40}K .



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Weightage and Exam Pattern

Class 9 has no board exam. This chapter is assessed in your school's periodic tests, half-yearly and annual exams, and the paper pattern and marks are set by your school. It is also a foundation for Class 10 Science, where valency and electronic configuration are used again and again.

How the chapter is usually tested:

- **Numericals:** finding protons, neutrons and electrons from Z and A ; completing tables like Table 8.5; weighted average atomic mass.
- **Configurations:** writing electronic configurations of the first 18 elements and deducing valence electrons and valency.
- **Reasoning:** observations and conclusions of the gold foil experiment; why Thomson's and Rutherford's models were replaced; how Bohr explained stability.
- **Competency formats:** assertion-reason, multiple choice, true-or-false statements with reasons.



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